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IGCSE Chemistry

Cambridge IGCSE Chemistry · 0620

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States of matter

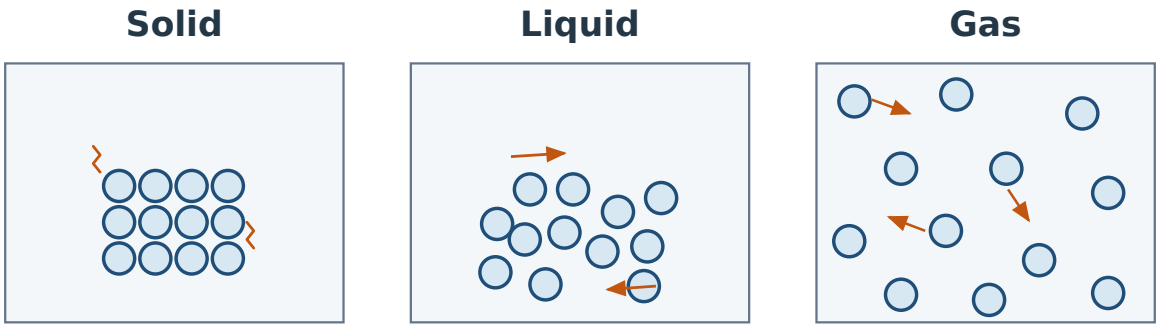
1.1 Solids, liquids and gases

Type 1: Explaining changes of state and heating and cooling curves using particle motion

Key Ideas A piece of ice has its own shape. When the ice melts, the water takes the shape of its container. When water becomes a gas, it can spread throughout the space available to it. Solid, liquid and gas are the three states of matter. We use the general term **particles** for the tiny units that make up a substance; the particles in water are water molecules. During a change of state, the water particles remain water particles. Their arrangement and movement change.

Compare the spacing first, then ask whether the particles can move away from fixed positions. In a typical solid, the particles are close together in a fairly regular arrangement. Attractions between them hold them near fixed positions, but they still vibrate. Liquid particles are also close together, but their arrangement is irregular and they can move and slide past neighbouring particles. Gas particles are far apart and move rapidly and randomly in all directions.

The same particles can have different arrangements and movements



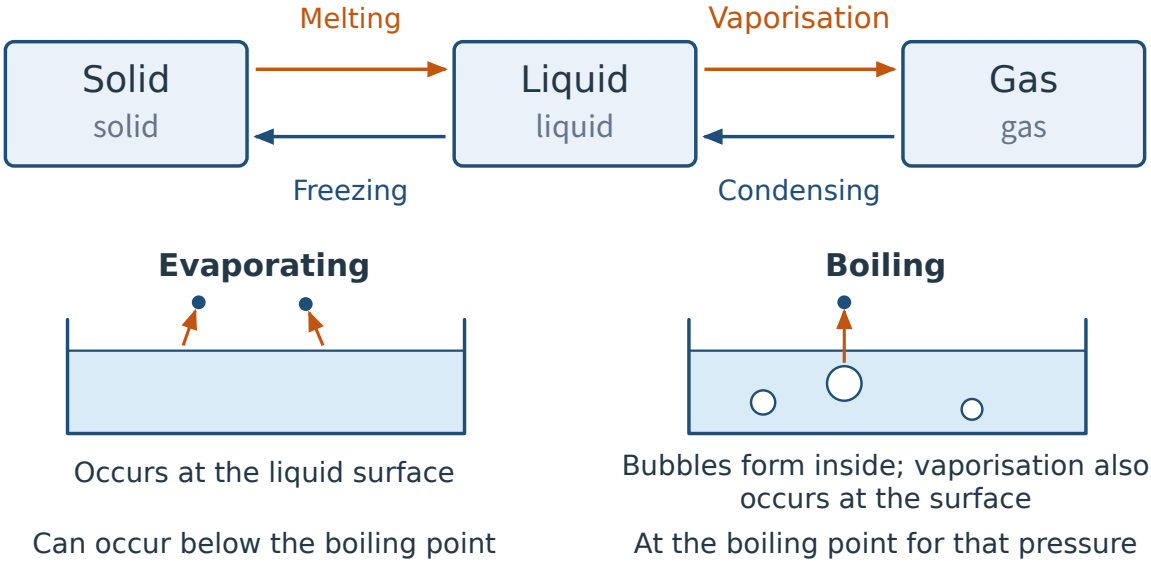
Short orange lines indicate movement. Particles in a solid are also moving.

The three boxes represent regions being observed. Each shows 12 particles of the same substance to compare their arrangement and movement; particle sizes and densities are not drawn to scale.

Whether particles can move from one position to another explains the visible properties of the material. A solid has a fixed shape and volume. A liquid has a fixed volume but takes the shape of its container. A gas has neither a fixed shape nor a fixed volume and fills the available space in its container. There is little empty space between the particles in solids and liquids, so these states are difficult to compress. Gases have much more space between their particles and are easier to compress. Compression mainly reduces the distance between particles; it does not make each particle smaller.

The name of a change of state depends on its starting and finishing states. A solid becoming a liquid is melting; a liquid becoming a solid is freezing; and a gas becoming a liquid is condensing. The general change from liquid to gas is vaporisation. Evaporation occurs at the surface of a liquid and can happen below its boiling point. During boiling, bubbles form within the liquid, while vaporisation also occurs at the surface. Water left out gradually disappears by evaporation: the whole beaker does not have to reach boiling point.

Identify the states before and after the change



Warm-coloured arrows show changes that absorb energy; blue arrows show changes that release energy. The large circles below represent bubbles, each containing many gas particles.

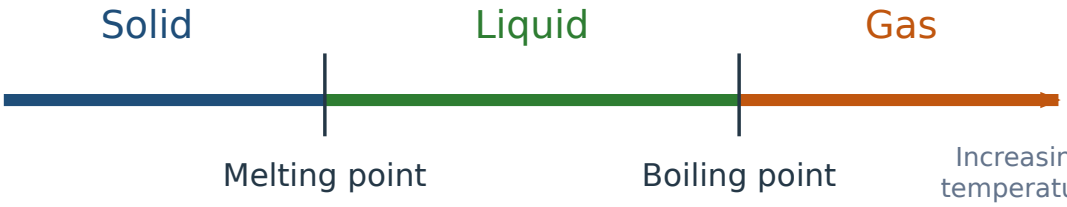
Melting and vaporisation require energy to be absorbed. Freezing and condensation release energy to the surroundings. During melting, particles gain energy, overcome the restrictions that hold them in fixed positions and begin to slide past one another. During freezing, they become restricted to fixed positions again, although they continue to vibrate.

For example, melted wax hardens as it cools. The change is liquid → solid, so it is freezing. Water vapour that cools and forms small droplets on the outside of a glass changes from gas → liquid, so it is condensing. Complete the change for condensation: _____→_____. Particles in a solid are still _____.

How can temperature tell us the state? The melting point is the temperature at which a solid melts; the boiling point is the temperature at which a liquid boils. A pure substance is a sample containing only one chemical substance. At the same external pressure, a pure substance has a fixed melting point and a fixed boiling point. A whole range of temperatures cannot be treated as one boiling point for a pure substance. Questions often use °C for temperature in degrees Celsius. Compare negative temperatures in the same order as numbers on a number line: for example, −7°C is higher than −77°C.

A pure substance's state depends on temperature relative to its melting and boiling points

Compare the same pure substance at the same external pressure



A liquid's temperature is above its melting point and below its boiling point.

During a change of state, solid and liquid coexist at the melting point; liquid and gas at the boiling point.

The comparison is for a pure substance at the same pressure. Distances on the diagram do not represent actual temperature differences. Negative temperatures are compared using their order on a number line.

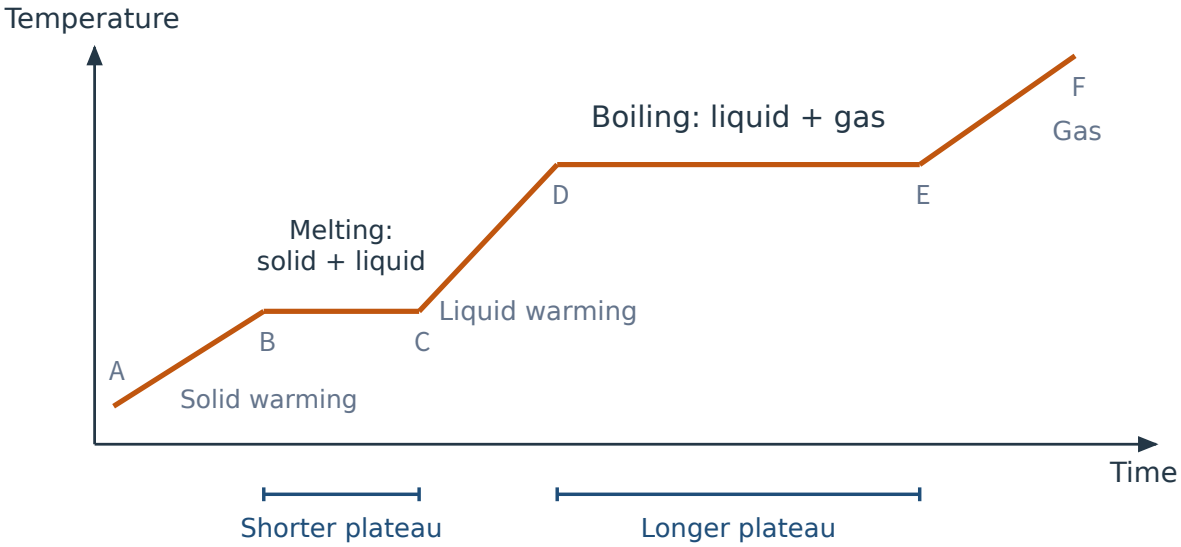
At the same pressure, a substance is normally solid below its melting point, liquid between its melting and boiling points, and gaseous above its boiling point. If the question states that a sample is liquid at room temperature, its boiling point must be above room temperature. Colour may help identify a substance, but it cannot replace this temperature condition.

For example, a pure substance is liquid at 20°C. The possible boiling points are −10°C, 50°C and 70–80°C. First reject −10°C because it is below 20°C. Then reject the wide temperature range. The remaining fixed temperature is 50°C. There is no need to guess the substance's name from its colour.

Judge temperature and energy absorption separately. The energy a particle has because it is moving is its kinetic energy. Temperature reflects the particles' average kinetic energy. When a substance warms up within one state, its particles generally move faster or vibrate more vigorously. Absorbed energy can also overcome the attractions that maintain the original state, changing the particles' arrangement and the state of the substance. Absorbing energy therefore does not mean that the temperature must rise at every moment.

Plotting a sample's temperature against time during heating produces a heating curve. First identify time on the horizontal axis and temperature on the vertical axis, then read the curve from left to right. In the diagram below, AB shows the solid warming up. At BC, the melting point has been reached and melting occurs, so solid and liquid coexist. CD shows the liquid warming up. At DE, the boiling point has been reached and boiling occurs, so liquid and gas coexist. EF shows the gas continuing to warm up.

Constant heating rate: temperature rises on slopes; energy is still absorbed on plateaux



Rising slope: mean kinetic energy increases. Plateau: temperature stays constant as the state changes.

At the same constant heating rate, a longer time means more energy absorbed.

A→F follows the same sample as heating continues. Energy is still absorbed during each plateau and is used to change state. A plateau does not mean that no energy is being absorbed.

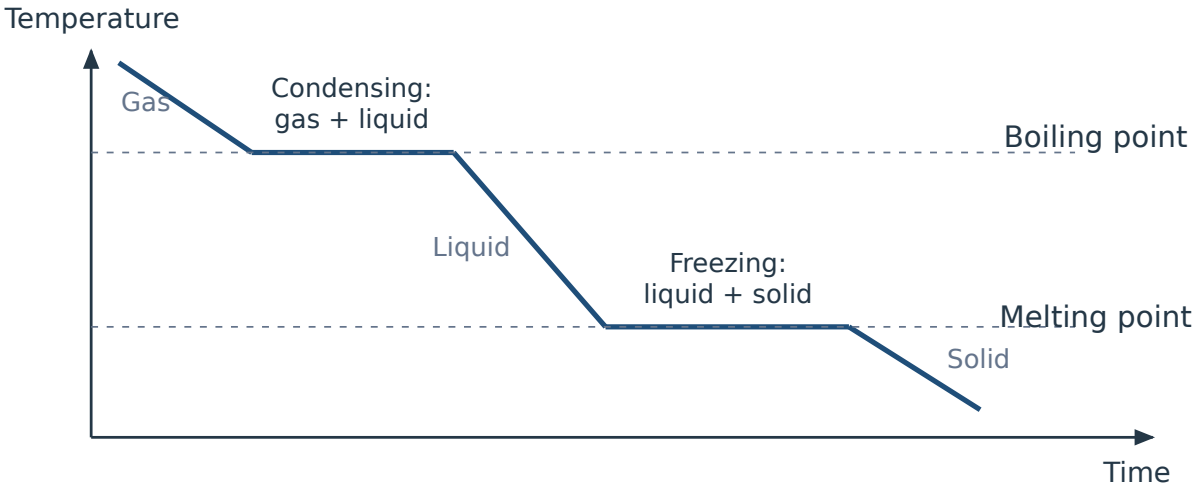
An upward-sloping section shows a rising temperature and increasing average kinetic energy. A horizontal plateau shows a temporarily constant temperature while the state changes. Energy is still entering the sample and is mainly used to overcome attractions between particles. A plateau does not mean that heating has stopped, nor that the particles' average kinetic energy keeps increasing. Changing from vibrating near fixed positions to sliding past one another is a change in the way particles move and in their freedom of movement. Distinguish this from an increase in temperature and average particle speed.

If a question says that heating occurs at a constant rate, the same amount of energy is supplied to the sample per unit time. You can therefore compare the total energy absorbed during two sections by comparing their durations on the time axis.

For example, a heater supplies 10 units of energy to a sample every second. In 20 seconds it supplies 200 units; in 50 seconds it supplies 500 units. More energy enters during the longer interval. Both intervals could be plateaux at constant temperature. Without a condition such as a constant rate of energy transfer, plateau length alone does not allow this comparison.

Read a cooling curve through the reverse sequence of changes. A gas first cools, then condenses at the temperature corresponding to its boiling point. Once condensation is complete, the liquid cools further and freezes at the temperature corresponding to its melting point. Once freezing is complete, the solid can continue to cool. For the same pure substance at the same pressure, the condensation temperature equals the boiling point and the freezing temperature equals the melting point.

Cooling: condensation comes before freezing, at a higher temperature



Same pure substance and pressure: condensing at the boiling point; freezing at the melting point.

The sample is shown cooling gradually from a hot gas. It first condenses at the higher temperature, then freezes at the lower temperature. No actual durations are specified for the plateaux.

For example, as hot water vapour cools, the gas-to-liquid plateau appears first. Only after the temperature falls further can a liquid-to-solid plateau occur. If a graph shows heating followed by cooling, always read the time axis from left to right. At the maximum-temperature turning point, change your interpretation from energy absorption during heating to energy release during cooling, then continue reading the cooling sequence to the right. The boiling and condensation plateaux on the same graph should lie at the same temperature; the freezing plateau should be lower. A curve may end at an intermediate stage: a question does not have to show the final section in which the solid cools.

Answers: gas; liquid; vibrating near fixed positions

Recognising the Question Identify the states before and after the change, then name the change of state. Read a curve from left to right in time, marking sloping sections and plateaux. Use the heating or cooling direction to identify the states, then check temperature, energy absorption or release, and duration separately. For a pure substance at a stated temperature, also check its fixed transition temperatures and the temperature range for each state.

Worked Example 1 (2025·May/June 0620/11 Q1) · Level 1

Which process happens when water vapour changes to rain?

A boiling

B condensing

C evaporating

D freezing

Solution

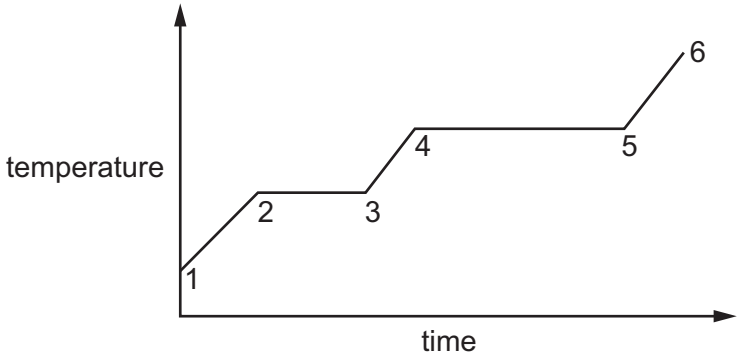
Water vapour is a gas, whereas rain consists of liquid droplets. The change is gas → liquid, which is condensing, so choose B.

Boiling and evaporation both change liquid into gas, which is the opposite direction. Freezing changes liquid into solid, so its final state does not match liquid raindrops either. Check both the starting and finishing states when naming a change.

Worked Example 2 (2025·May/June 0620/21 Q2) · Level 3

A sample of ethanol is heated at a constant rate.

The heating curve for this sample is shown.



Which statement about the heating curve is correct?

A Ethanol changes its physical state between 3 and 4.

B Ethanol condenses between 2 and 3.

C Energy is absorbed between 3 and 4 but **not** absorbed between 4 and 5.

D More energy is absorbed between 4 and 5 than between 2 and 3.

Solution

Step 1: read the axes and conditions. The horizontal axis shows time and the vertical axis shows temperature. Ethanol is heated at a constant rate, so equal time intervals supply equal amounts of energy. From 1→6, the graph shows a rise, a plateau, another rise, another plateau and a final rise.

Step 2: identify the state in each section. From 1→2, the solid warms up; 2→3 is the melting plateau; 3→4 shows the liquid warming up; 4→5 is the boiling plateau; and 5→6 shows the gas warming up. Changes of state occur at the two horizontal plateaux. The sloping sections show temperature changes within one state.

Step 3: check the four statements separately. A describes 3→4 as a change of state, but the liquid is warming up, so A is incorrect. B calls 2→3 condensation, but heating changes the solid into liquid here: it is melting, so B is incorrect. C says that no energy is absorbed during 4→5, but energy continues to be absorbed during the boiling plateau even though the temperature remains constant, so C is incorrect.

For D, compare the horizontal time intervals. The 4→5 plateau is clearly longer than the 2→3 plateau. Since energy is supplied at a constant rate, a longer time means more energy is absorbed. D is therefore correct: choose D. The higher temperature of the plateau is not the reason for this energy comparison. The relevant conditions are the same rate of energy supply and different durations.

Practice

1. (2025·May/June 0620/13 Q1) · Level 1

Liquid iron is cooled to form solid iron.

Which statement about the particles in iron is correct?

A The particles move further apart.

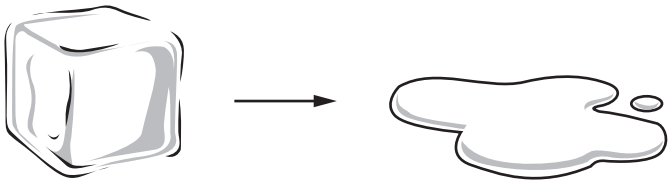
B The particles move faster.

C The position of the particles becomes fixed.

D The attractions between the particles become weaker.

2. (2025·May/June 0620/22 Q1) · Level 1

An ice cube melts.



ice cube pool of water

What happens to the molecules of water in the ice cube?

A They evaporate.

B They dissolve.

C They gain energy.

D They lose energy.

3. (2025·May/June 0620/12 Q1) · Level 2

Substance L melts at –7 °C and is a brown liquid at room temperature.

What is the boiling point of pure L?

A –77 °C

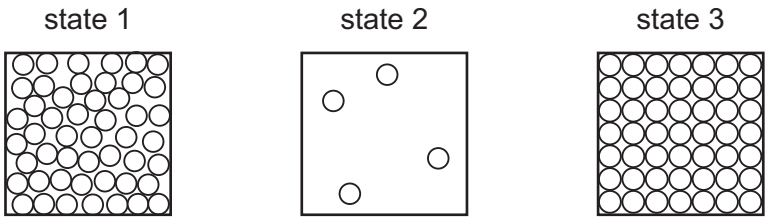
B –7 °C to +7 °C

C 59 °C

D 107 °C to 117 °C

4. (2025·May/June 0620/11 Q2) · Level 3

The diagrams show the arrangement of particles in three different states of matter.



Which row describes the change in energy of the particles and in particle motion for the given change in state?

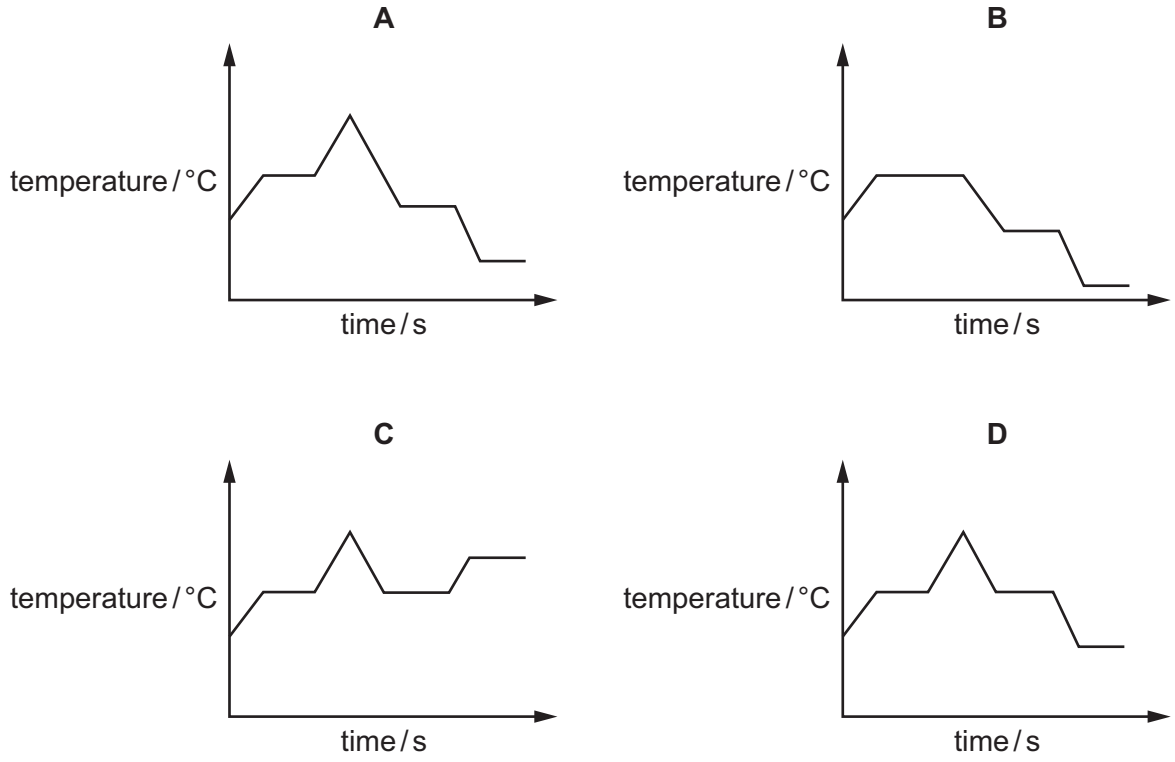
	change in state	energy of particles	particle motion
A	1 → 2	decreases	decreases
B	2 → 1	decreases	increases
C	3 → 1	increases	increases
D	1 → 3	increases	decreases

5. (2025·May/June 0620/22 Q2) · Level 3

A sample of a liquid, X, is heated to a temperature above its boiling point.

X is then cooled, so that it condenses and then freezes.

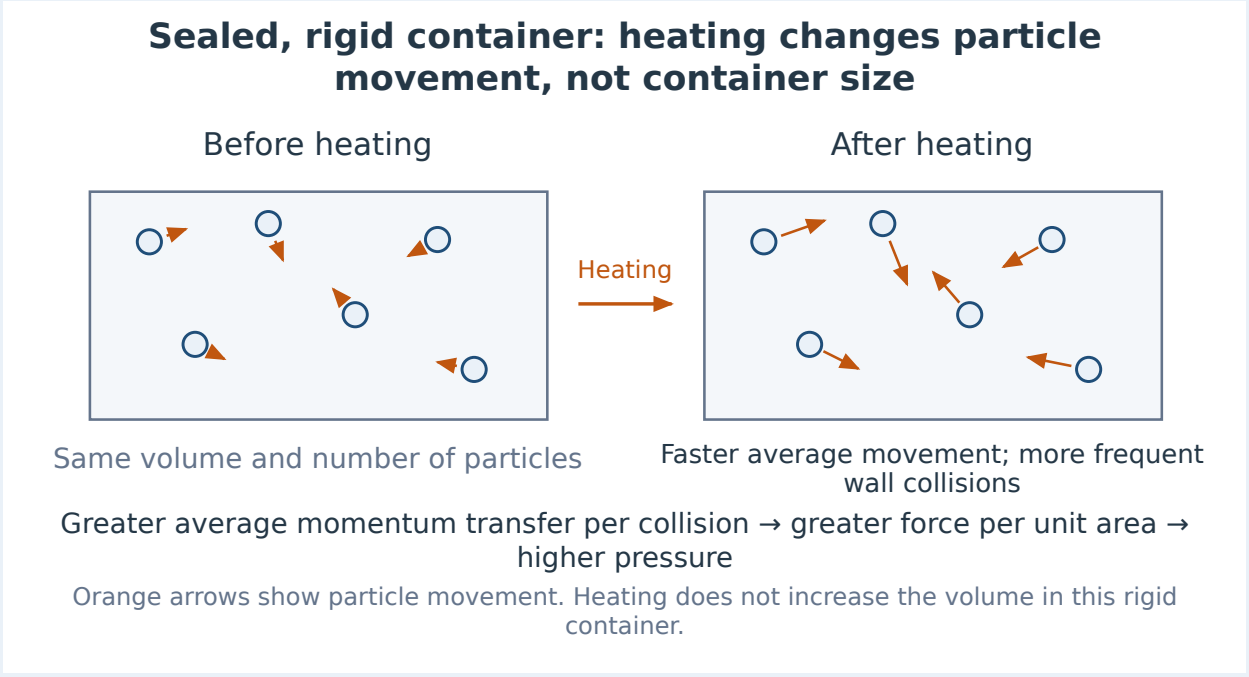
Which graph describes the change in temperature of X over time?



Type 2: Explaining gas particle behaviour when temperature, pressure or volume changes

Key Ideas Gas particles are always moving in all directions. When they collide with a container wall, they exert a force on it. The average force per unit area produces the gas **pressure**. Faster-moving gas particles generally strike the walls more often, and each collision produces a greater average transfer of momentum. Whether heating increases the pressure or the volume also depends on whether the container can expand.

First identify the conditions that remain constant. A fixed amount of gas means that no particles enter or leave during the comparison: the total number of particles remains the same. A sealed container does not allow gas to escape. A fixed volume means that the container cannot change size. When such a gas is heated, its average kinetic energy increases and its particles move faster. They strike the walls more frequently within the same space, with a greater average effect per collision, so the pressure increases. The volume remains unchanged.



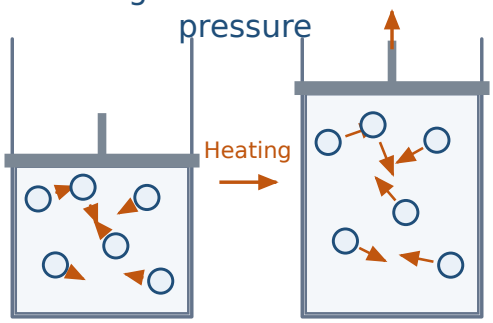
The comparison uses the same type and amount of gas. Both containers have the same volume. The heating arrow shows energy input; the particle arrows represent movement.

If the container can expand, check whether the external pressure is constant. Imagine gas in a cylinder fitted with a movable piston: a plate that can slide along the cylinder walls. If the external pressure on the piston remains the same, heating makes the gas particles move faster. The gas pushes the piston outwards and occupies more space. By the time the internal and external pressures balance again, the gas has a larger volume. Therefore, for a fixed amount of gas at constant pressure, heating increases volume and cooling decreases it. Do not also call this a fixed-volume situation: that is a different condition.

Another comparison keeps the **temperature constant**. Constant temperature means constant average kinetic energy. Increasing the external pressure moves the piston inwards and reduces the gas volume. The same number of particles now occupies less space, their average separation decreases, and they strike the walls more frequently until the gas pressure again balances the external pressure. Reducing the external pressure allows the gas to expand. Compression reduces the gaps between particles; it does not shrink individual particles or, in this constant-temperature comparison, automatically increase their average kinetic energy.

When volume can change, identify the condition held constant

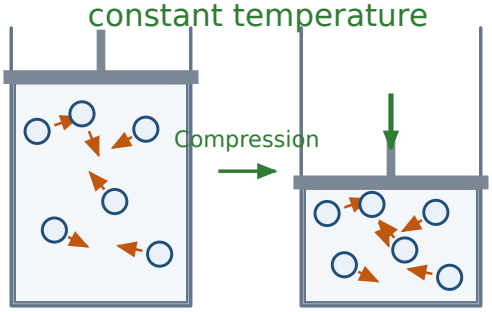
Heat the gas at constant external pressure



Faster particles push the piston outwards

Volume increases until pressures balance

Increase external pressure at constant temperature



Same number of particles in less space

Smaller gaps; more frequent wall collisions

Heating at constant pressure → larger volume. Higher pressure at constant temperature → smaller volume.
Compression does not shrink individual particles. At constant temperature, their mean kinetic energy is unchanged.

Each panel shows the same gas and particle number before and after. The grey bar is a movable piston.

External pressure is constant on the left; temperature is constant on the right. Particle size and total particle number remain unchanged. Piston positions show the volume changes.

For example, two equal amounts of the same gas are heated. Sample A is in a rigid, sealed bottle; sample B is in a cylinder with a movable piston at constant external pressure. A cannot expand, so the increased effect of particle collisions raises its pressure. B can push the piston outwards and eventually occupies a larger volume at the same pressure. Heating makes the particles move faster in both samples, but it does not produce the same conclusion about volume.

When two conditions change together, consider their effects separately before combining them. For a fixed amount of gas, reducing pressure and increasing temperature both favour an increase in volume. Increasing pressure and decreasing temperature both favour a decrease in volume. If one change favours expansion and the other favours contraction, the effects oppose one another. You cannot determine the final volume by reading only one condition, nor can you simply add or cancel percentage changes.

If two conditions change, consider each effect before combining them

Pressure change	Temperature change	Effect on volume
Decreases	Increases	Both increase the volume
Increases	Decreases	Both decrease the volume
Decreases	Decreases	Opposing effects; do not simply add them

Compare a fixed amount of gas. Without a stated relationship, percentage changes cannot be treated as volume changes.

For a fixed amount of gas, first identify the separate effects of pressure and temperature, then compare the combined conditions. For example, equal amounts of the same gas, A and B, begin in the same state. A undergoes a decrease in pressure and an increase in temperature. B undergoes an increase in pressure and a decrease in temperature. Both changes increase A's volume, while both decrease B's volume, so A has the larger final volume. If a question gives four combinations, first identify the combination in which both factors support the requested direction, then check whether any other row could exceed it. You do not need to introduce a gas equation when the required calculation relationship has not been provided.

In a table question, each column may represent an independent comparison. Treat heating at constant pressure and increasing pressure at constant temperature separately, then combine the two answers. For movement and pressure at constant volume, connect temperature → average particle motion → wall collisions → pressure. When no conversion relationship is given, compare increase, decrease or no change. Doubling a temperature in degrees Celsius does not mean that the gas volume doubles.

Recognising the Question Identify which of the amount of gas, temperature, pressure and volume remain constant. Then decide whether expansion or compression is possible. Explain the quantity that changes using particle movement, wall collisions and spacing.

Worked Example 1 (2025·May/June 0620/12 Q2) · Level 2

Which row describes how the volume of a gas changes when its temperature is increased at constant pressure and when its pressure is increased at constant temperature?

	temperature is increased at constant pressure	pressure is increased at constant temperature
A	volume decreases	volume decreases
B	volume decreases	volume increases
C	volume increases	volume decreases
D	volume increases	volume increases

Solution

In the first column, temperature is increased at constant pressure. The particles move faster on average, and the gas can expand until its pressure again balances the external pressure. The first entry is therefore volume increases.

In the second column, pressure is increased at constant temperature. A fixed amount of gas is compressed into a smaller volume and the spacing between particles decreases. The second entry is therefore volume decreases.

Combine these independent comparisons. Only row C gives volume increases; volume decreases, so choose C. The fixed-volume conclusion for a sealed, rigid container does not apply to the first column. The second column does not say that temperature increases with pressure.

Practice

1. (2025·May/June 0620/21 Q1) · Level 2

A gas is placed in a sealed container with a fixed volume.

The gas is heated.

Which statements are correct?

1 The gas molecules move around faster.

2 The gas particles collide with the walls of the container more frequently.

3 The pressure in the container decreases.

A 1 and 2 **B** 1 only **C** 2 and 3 **D** 2 only

2. (2025·May/June 0620/23 Q1) · Level 3

Which two changes, when applied **at the same time** to a fixed amount of gas, cause the greatest increase in volume of the gas?

	change in pressure	change in temperature
A	decrease by 10%	decrease by 10%
B	increase by 10%	decrease by 10%
C	decrease by 20%	increase by 10%
D	increase by 20%	increase by 10%

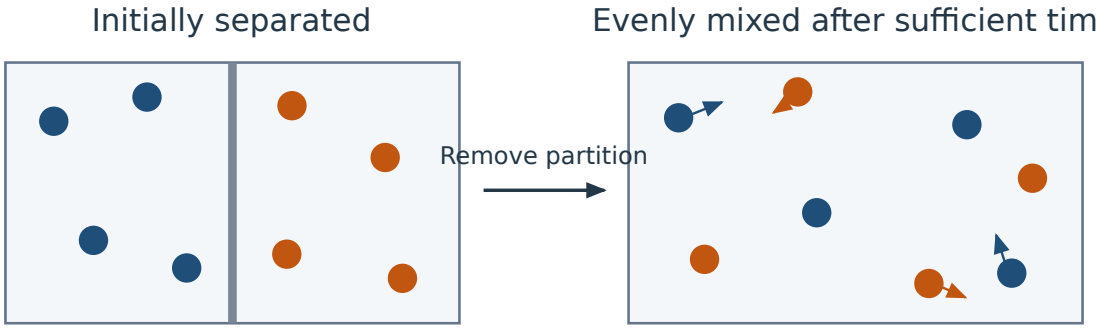
1.2 Diffusion

Type 1: Comparing diffusion using particle motion and relative molecular mass

Key Ideas If a substance with a smell is placed on one side of a room, the smell may eventually be detected further away. This does not happen because every particle chooses a particular direction. Gas particles constantly move randomly in all directions, gradually entering regions that initially contain none of those particles. The resulting mixing of particles of different substances is called **diffusion**.

To describe where there are more particles, we use **concentration**: the amount of a particular type of particle per unit volume. On the more concentrated side, more of those particles can randomly cross a boundary to the other side. Particles also move in the opposite direction, but fewer do so. Overall, there is therefore a **net movement** from a region of higher concentration of those particles to one of lower concentration. Net means the overall effect remaining when movements in opposite directions are compared. It does not mean that all particles travel in one direction.

Removing the partition allows the two gases to diffuse into each other



Each gas has a net movement towards its own lower concentration; the gases do not simply exchange sides.

Random motion continues after mixing. A greater gas density does not prevent upward diffusion.

Blue and orange identify the two gases. Each diagram contains the same 4 + 4 particles.

Colours identify the gases. Particle numbers are unchanged by mixing. Small arrows show some of the possible directions of random motion.

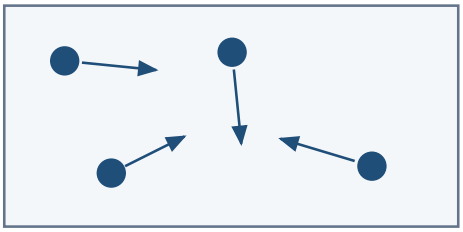
Gases can diffuse into one another, and particles in liquids can also mix gradually through random motion. Given enough time, substances in a sealed container usually become fairly evenly mixed. The particles are still moving, but roughly equal numbers move in opposite directions, so there is no continuing net movement that makes one side more concentrated.

In colour diagrams, compare the concentration of the same coloured gas in different places. Suppose the lower jar initially contains a coloured gas and the upper jar contains a colourless gas. When the barrier is removed, the coloured gas diffuses upwards and the colourless gas diffuses downwards. Eventually, both jars approach the same colour, which is paler than the original coloured gas. Being denser than air does not prevent a gas's particles from moving upwards: density is not a barrier separating the gases. A diagram that merely swaps the upper and lower colours does not show thorough mixing.

To compare rates, also compare the masses of the gas particles. A molecule is a group of atoms chemically bonded into one unit. **Relative molecular mass**, written Mr, compares the mass of one molecule against a common reference mass. It is a relative number with no unit. At the same temperature, a gas with a smaller Mr generally has a higher average particle speed and diffuses faster. Gas particles at the same temperature have comparable average kinetic energies, so lighter particles have a higher average speed. Comparing molecular masses is not the same as comparing how much gas is present: the latter is the total amount of the sample.

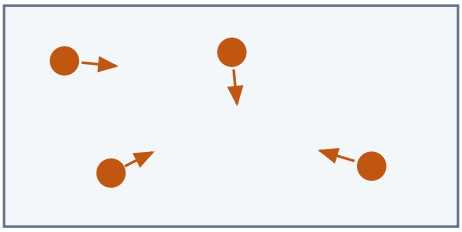
At the same temperature, lighter gas particles generally diffuse faster

Lower relative molecular mass



Higher average speed

Higher relative molecular mass



Lower average speed

Same temperature → comparable mean kinetic energy; lighter particles have a higher average speed.

Equal circle sizes identify particles. Arrow lengths indicate average speeds, not one speed for every particle.

The gases are compared at the same temperature. Arrows show only the relative speeds, not exact speed ratios. Circle size does not represent molecular mass.

If the question supplies the gases' Mr values, simply compare them. If it provides only formulae, first add the relative masses of every atom in each molecule. The letters in a formula identify elements; a small number at the lower right tells you how many atoms of that element are present in one particle. No small number means one atom. For example, in CO_2, C represents carbon and O represents oxygen, so the molecule contains one carbon atom and two oxygen atoms. Using carbon 12 and oxygen 16 from the Periodic Table:

$$M_{r(CO_2)} = 12 + 2 \times 16 = 44$$

Chapter 3 develops relative masses and chemical formulae more fully. For now, this addition is enough to compare diffusion rates.

The symbols and Periodic Table values needed here are listed together below for reference when you meet a new formula. H is hydrogen, C is carbon, N is nitrogen, O is oxygen and S is sulfur. The second letter of Cl is a lower-case l: the symbol means chlorine, not C joined to a second element.

H Hydrogen	C Carbon	N Nitrogen	O Oxygen	S Sulfur	Cl Chlorine
1	12	14	16	32	35.5

Convert formula subscripts into atom counts before adding masses

CO_2 → Read the subscripts → 1 C atom + 2 O atoms

$12 + 2 \times 16 = 44$

M_r refers to the whole molecule. An omitted subscript means 1.

This relative mass comparison has no unit such as g or kg.

The arrows convert the formula into atom counts. Carbon 12 and oxygen 16 are used to calculate Mr for the whole CO_2 molecule.

You can count atoms from a supplied formula even if the substance's name is unfamiliar. For example, CH_4 contains one C and four H atoms; C_2H_6 contains two C and six H atoms. These substances are methane and ethane, respectively, but a full understanding of organic naming is not needed to compare their diffusion. CO is carbon monoxide, SO_2 is sulfur dioxide, HCl is hydrogen chloride, H_2S is hydrogen sulfide and NO_2 is nitrogen dioxide. The exam questions in this group provide the molecular formulae needed.
For example, compare CO_2 and O_2 at the same temperature. O_2 contains two oxygen atoms, so its relative molecular mass is:

$$M_{r(O_2)} = 2 \times 16 = 32$$

CO_2 has an M_r of 44. Since 32 is less than 44, O_2 generally diffuses faster. The subscript 2 in CO_2 does not make its mass the same as that of O_2, and having more kinds of letter in a formula is not a reason for faster diffusion. For a gas consisting of single atoms, compare the relative masses of those atoms directly.

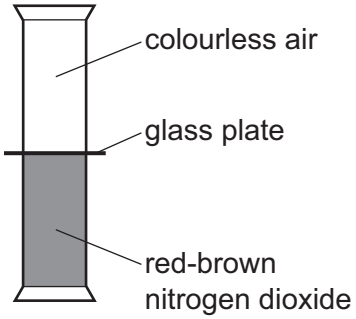
When answering diffusion questions, first decide whether the question asks where particles eventually appear or which gas arrives first or diffuses faster. The first requires constant random motion and concentration differences. The second also requires the stated temperature and particle-mass conditions. Colour intensity or density alone is insufficient.

Recognising the Question For the final distribution, use the mixing caused by random motion. For diffusion rate, compare the relative masses of whole gas particles at the same temperature. Do not substitute colour, density or total sample mass for molecular mass.

Worked Example 1 (2025·May/June 0620/12 Q3) · Level 2

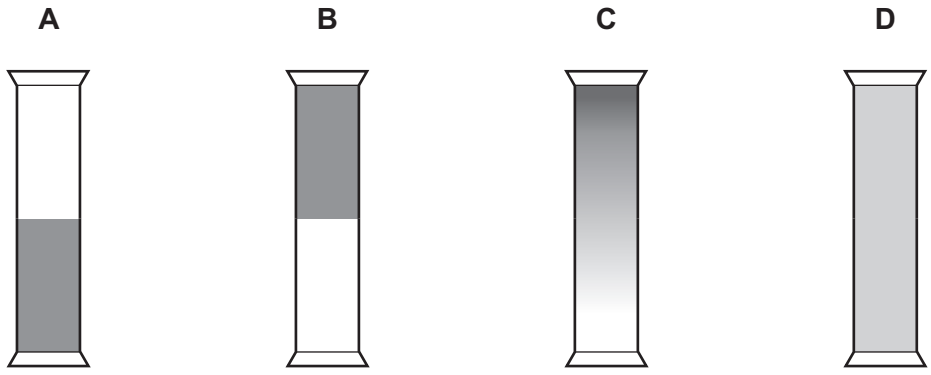
Nitrogen dioxide is a red-brown gas which is more dense than air.

The diagram shows the arrangement of two gas jars which contain nitrogen dioxide and air separated with a glass plate.



The glass plate is removed, and the gas jars are left for 24 hours.

Which diagram shows the gas jars after 24 hours?



Solution

Initially, the lower jar contains red-brown nitrogen dioxide and the upper jar contains colourless air. After the glass

plate is removed, particles of both gases continue moving randomly. Nitrogen dioxide can enter the upper jar, and air can enter the lower jar. After 24 hours, the gases will be fairly evenly mixed throughout both jars.

The nitrogen dioxide spreads through roughly twice its original space, so both jars should eventually have the same paler colour. Diagram D fits. Nitrogen dioxide being denser than air does not prevent its particles from diffusing upwards.

A largely retains the original layers, B merely swaps the gases' positions, and C still shows a clear colour gradient. None shows uniform mixing after sufficient diffusion. Choose D. The particles continue moving after the mixture becomes uniform.

Worked Example 2 (2025·October/November 0620/23 Q2) · Level 2

Which gas diffuses most quickly at room temperature and pressure?

A CH_4 **B** C_2H_6 **C** CO **D** SO_2

Solution

Room temperature and pressure are fixed, so the gases are compared under the same temperature and pressure conditions. Apply the relationship already introduced: a gas with a smaller relative molecular mass, Mr, generally diffuses faster.

Count the atoms in each formula, then calculate using H=1, C=12, O=16 and S=32 from the question paper's Periodic Table:

Gas	Atoms present	Relative molecular mass
CH_4	1 C, 4 H	$12 + 4 \times 1 = 16$
C_2H_6	2 C, 6 H	$2 \times 12 + 6 \times 1 = 30$
CO	1 C, 1 O	$12 + 16 = 28$
SO_2	1 S, 2 O	$32 + 2 \times 16 = 64$

Compare 16, 30, 28 and 64. The smallest is 16 for CH_4, so methane, CH_4, diffuses fastest. Choose A. Ethane, C_2H_6, also contains only C and H, but has different numbers of atoms. Nor does having fewer letters make CO lighter than CH_4. Calculate the mass of the whole molecule, rather than just counting the different elements.

Practice

1. (2025·October/November 0620/42 Q1(c)) · Level 2

The three states of matter are solid, liquid and gas.

(c) Gases diffuse.

(i) Describe, in terms of particles, why gases diffuse.

(ii) State what determines the relative rate of diffusion of gases at constant temperature.

2. (2025·February/March 0620/22 Q3) · Level 2

Samples of four gases are released in a room at the same time.

The gases are carbon dioxide, CO_2, hydrogen chloride, HCl, hydrogen sulfide, H_2S, and nitrogen dioxide, NO_2.

Which gas diffuses fastest?

A carbon dioxide

B hydrogen chloride

C hydrogen sulfide

D nitrogen dioxide

Atoms, elements and compounds

2.1 Elements, compounds and mixtures

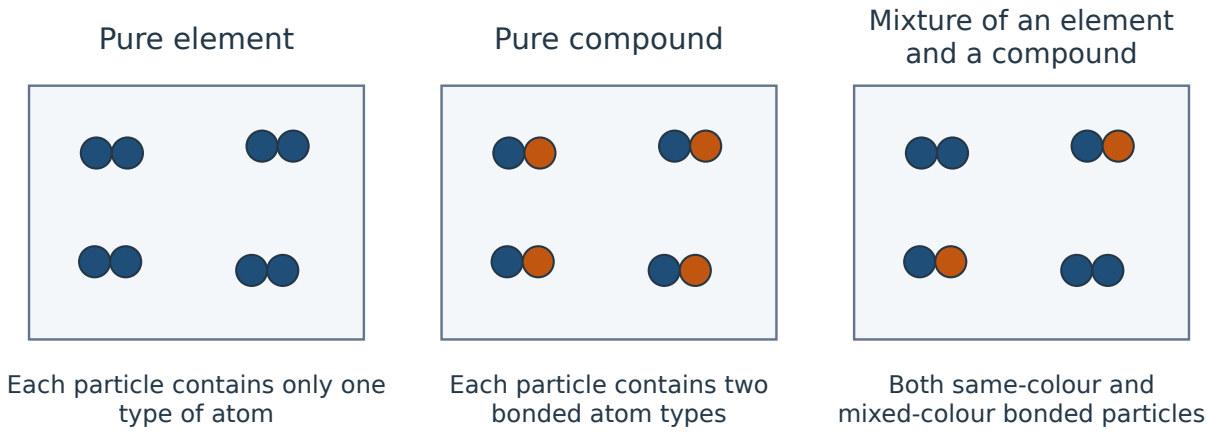
Type 1: Distinguishing elements, compounds and mixtures from particle diagrams

Key Ideas We have used particles as a general name for the tiny units that make up matter. Now distinguish two levels. An **atom** is a basic particle of an element. When atoms are joined by chemical bonds into a separate unit, they form a **molecule**. In this section's sphere diagrams, each colour represents one type of atom. A connected group of spheres represents chemically bonded atoms, not particles that happen to be close together.

An **element** contains only one type of atom. One oxygen molecule contains two oxygen atoms, but both atoms are of the same type, so oxygen is still an element. A **compound** contains two or more elements chemically combined in a fixed ratio. For example, a water molecule contains two hydrogen atoms and one oxygen atom. Hydrogen and oxygen are different elements, so water is a compound. The distinction depends on which elements are actually present and whether they are chemically combined, not simply on substances having different names.

A **mixture** consists of two or more substances mixed together. Mixing does not necessarily turn all of them into one new compound. A particle diagram often shows different types of the original particles present together. A pure compound can also contain atoms of different colours, so two colours in a box are not enough to identify a mixture.

Identify each particle's composition, then count the types of particles present



Colours identify atom types. Joined circles represent atoms chemically bonded into one particle.

Mixing does not reconnect substances into new particles. The original particle types can remain present.

Each connected group is one molecule. Every box contains four molecules. The diagrams compare composition and types of substance; they do not specify actual sizes or concentrations.

For example, a box containing only identical blue–orange two-atom molecules represents a pure compound. If both blue–blue and blue–orange molecules are present, there is one element and one compound, so the sample is a mixture. If there are only blue–blue and orange–orange molecules, it is a mixture of two elements. First

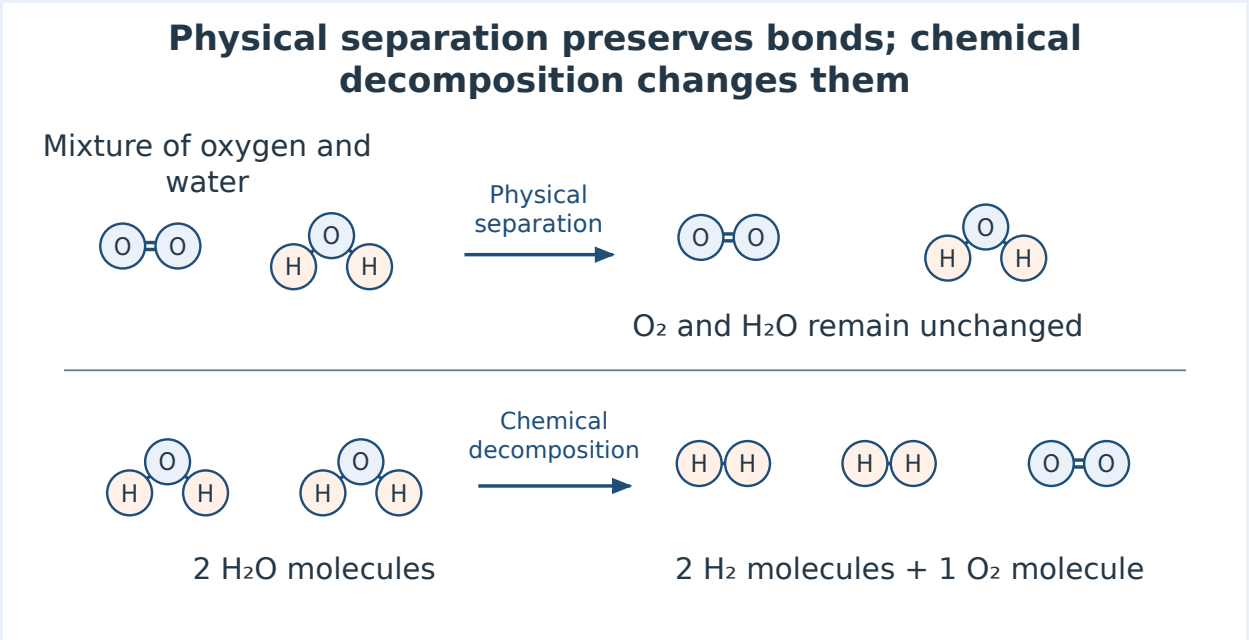
identify the composition of each particle, then count the types of substance in the whole box. Keep these two steps in that order.

A question may give substance names instead of sphere diagrams. **Aqueous** means dissolved in water, so include water as a component as well as the named solute. The solute is the substance being dissolved; water is the solvent that dissolves it here. Dissolving is a mixing process and does not mean that all the components have chemically combined into a different pure substance.

The following composition facts are needed for this group of exam questions. Hydrogen H, oxygen O, carbon C, nitrogen N, iron Fe and chromium Cr are elements. Oxygen gas O_2 and nitrogen gas N_2 consist of molecules containing only one type of atom, so they are still elements. Water H_2O contains hydrogen and oxygen, while glucose $C_6H_{12}O_6$ contains carbon, hydrogen and oxygen: both are compounds. Clean, dry air still contains different gases, including nitrogen, oxygen, a little argon Ar and carbon dioxide CO_2. Clean does not mean pure, and dry only means that water vapour has been removed. Stainless steel is an alloy of iron with chromium and other elements. An alloy is a mixture; appearing to be one continuous piece of metal does not prove that it is a pure element.

Apply these facts by listing the actual components. For example, aqueous glucose contains glucose and water, so it is a mixture of two compounds. Oxygen dissolved in water retains oxygen as an element and water as a compound. If the question asks for exactly one element and one compound, air must be rejected because it contains several elemental gases.

Another useful distinction concerns proportions and separation. In a compound, the elements are chemically combined in fixed proportions. The proportions of a mixture can vary, and its components can be separated using differences in physical properties. Distillation, for example, separates a mixture using differences in boiling point; it does not split water molecules into hydrogen and oxygen. Turning a compound into other substances requires a chemical change. This distinction will be used again in the chapters on reactions and separation.



In the top row, separating a mixture does not break up water molecules. The bottom row shows changes in bonding during chemical decomposition. The numbers of H and O atoms are the same before and after.

For example, match these statements to the columns for element, compound and mixture: all atoms belong to the same element; different elements are chemically combined; components can be separated by physical processes. The first describes an element, the second a compound and the third a mixture. If a multiple-choice row combines all three columns, every entry must be correct. Two correct columns do not make up for an incorrect third entry.

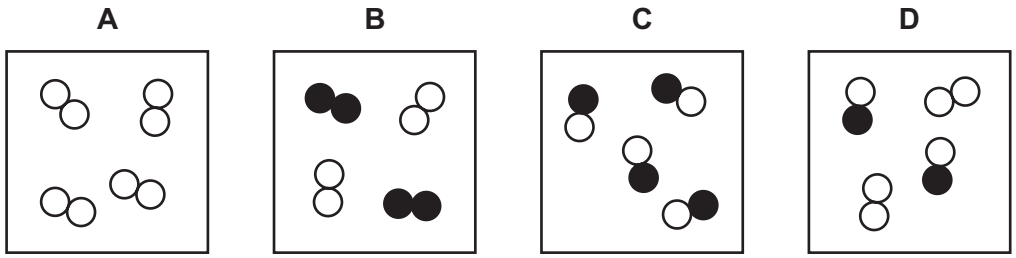
Compounds can be changed into other substances by chemical changes. Elements cannot be broken down into simpler substances by ordinary chemical processes. This refers to chemical processes, not other processes that

change atomic nuclei. Physical separation of mixtures uses differences such as boiling point or solubility, without requiring the components to become new substances.

Recognising the Question First identify how many types of atom each particle contains and how they are bonded. Then identify how many types of substance are in the whole sample. In worded questions, expand each substance name into its actual components.

Worked Example 1 (2025·May/June 0620/13 Q2) · Level 2

Which diagram shows a mixture of an element and a compound?



Solution

White and black circles represent two different types of atom. Each connected pair represents one molecule.

A contains only white–white molecules, so it contains one element and is a pure element. B contains both white–white and black–black molecules. Each molecular type contains only one type of atom, so B is a mixture of two elements. C contains only identical black–white molecules, each made from two elements, so it is a pure compound. D contains both white–white molecules and black–white molecules. The first is an element and the second a compound, exactly as required. Choose D. Do not classify C as a mixture simply because its pure compound is drawn using both black and white circles.

Practice

1. (2025·May/June 0620/12 Q4) · Level 2

Which substance is a mixture of **one** element and **one** compound?

A aqueous glucose

B clean, dry air

C oxygen dissolved in distilled water

D stainless steel

2. (2025·October/November 0620/22 Q4) · Level 3

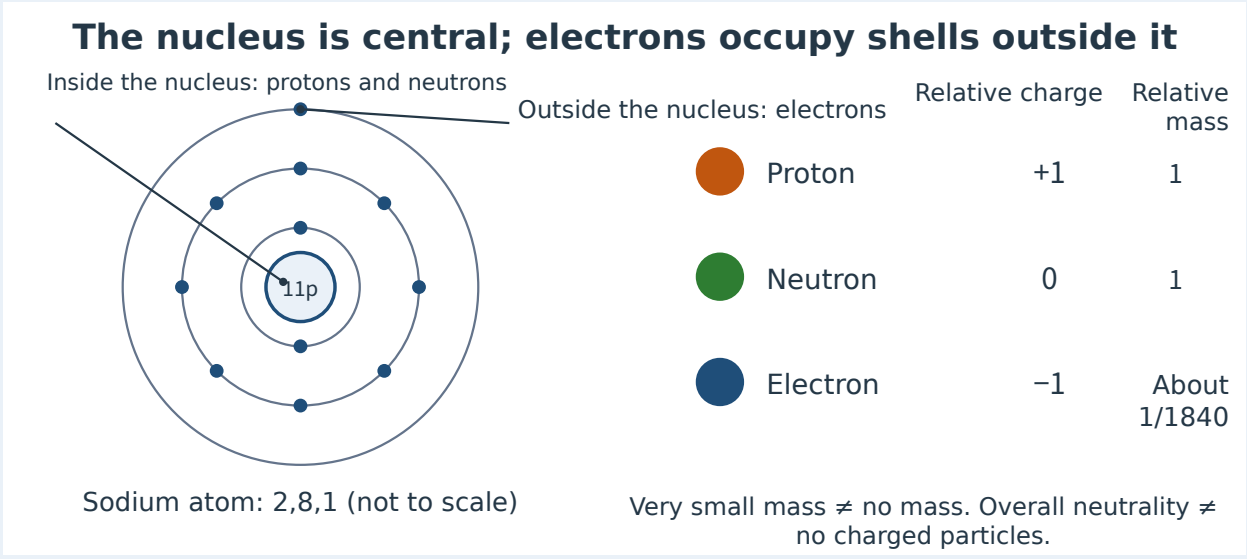
Which row describes elements, compounds and mixtures?

	elements	compounds	mixtures
A	all the atoms are the same type	contain different atoms chemically joined together	can be separated using physical processes
B	all the atoms are the same type	can be separated into new substances using chemical processes	contain different atoms chemically joined together
C	can be separated into new substances using chemical processes	all the atoms are the same type	contain different atoms chemically joined together
D	contain different atoms chemically joined together	all the atoms are the same type	can be separated using physical processes

2.2 Atomic structure and the Periodic Table

Type 1: Finding particle numbers from atomic number, charge and electron shells

Key Ideas We previously represented particles with dots. Now imagine magnifying one atom. The very small central region is the **nucleus**. It contains positively charged **protons** and uncharged **neutrons**. Negatively charged **electrons** occupy electron shells outside the nucleus. The nucleus and electrons are enlarged in the diagram to make their positions clear; it is not a scale drawing of an atom. The nucleus of the most common hydrogen atom contains one proton and no neutrons.



Protons in the nucleus are positive, neutrons are uncharged, and electrons outside the nucleus are negative. The electron's relative mass is recorded as 1/1840, as given in the first exam question.

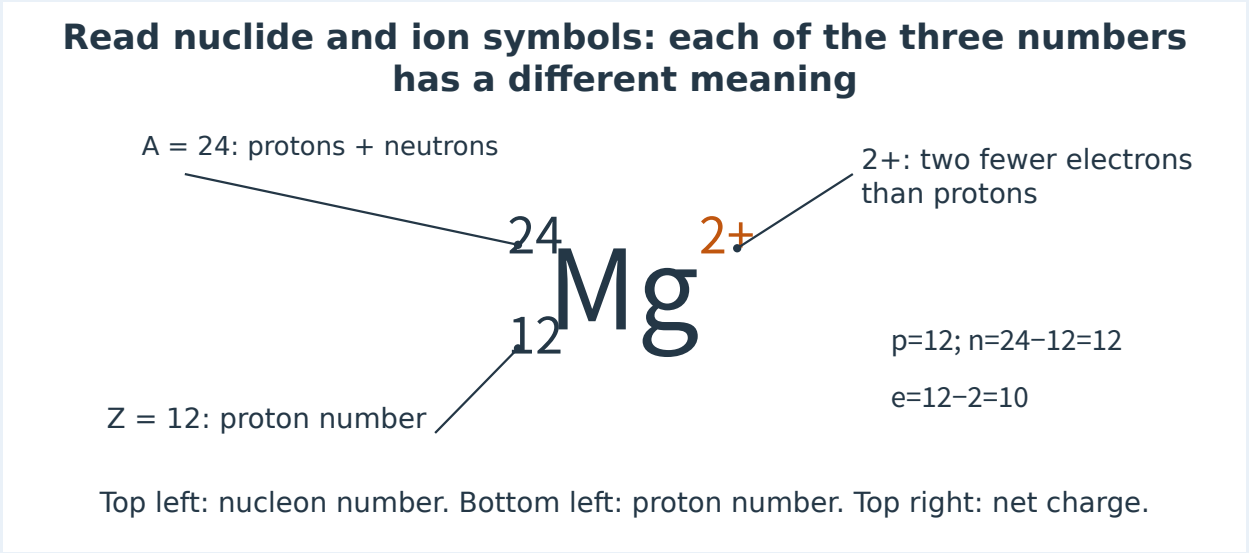
Charge and mass are different properties. If a proton's charge is assigned +1, an electron's charge is −1 and a neutron's relative charge is 0. A proton and a neutron each have a relative mass of approximately 1. An electron's mass is much smaller; the first exam question gives 1/1840. Electrons do have mass. Neutrons also have mass: they lack net charge, not mass. Read table headings carefully so that a charge of 0 does not end up in the relative-mass column.

Proton number determines the element. The number of protons in an atom's nucleus is its atomic number, also called its proton number, represented by Z. Elements are arranged by Z in the Periodic Table. To identify an element, use its proton number: for example, Z=11 is sodium, Na, and Z=19 is potassium, K. The exercises also use zinc, Zn, with Z=30, and rubidium, Rb, with Z=37. These numbers are also proton numbers. Adding or removing neutrons does not change the element; changing the proton number does.

Protons and neutrons in the nucleus are collectively called nucleons. The **mass number**, also called the nucleon number and represented by A, is the total number of protons and neutrons in a particular nucleus. If A and Z are known, the neutron number is:

$$\text{number of neutrons} = A - Z$$

Electrons are not included in the nucleon number. The average relative atomic mass in the Periodic Table cannot automatically be used as the mass number of every atom. An average mass and the total number of nucleons in one nucleus are different information.



Subtract proton number Z from mass number A to obtain the neutron number. The ionic charge changes the electron count in this question, not the nucleus.

The 24 at the upper left is mass number A; the 12 at the lower left is proton number Z. Mg identifies magnesium.

The 2+ at the upper right means that the ion has two units of positive charge.

Read the nucleus first. There are 12 protons, and the number of neutrons is:

$$\text{number of neutrons} = 24 - 12 = 12$$

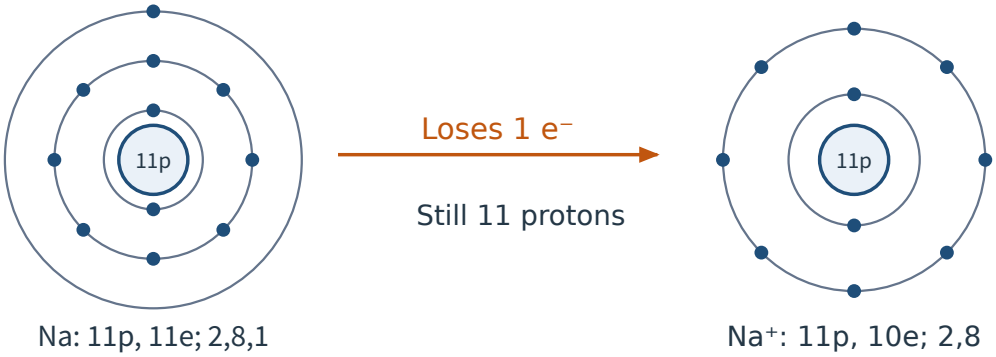
Then determine the electrons. A neutral magnesium atom has 12 electrons. It loses two when forming Mg^{2+}, so:

$$\text{number of electrons} = 12 - 2 = 10$$

Do not subtract the charge 2 from mass number 24. A charge of 2+ does not mean that the nucleus has lost two nucleons.

An atom is neutral because its positive and negative charges balance. In a neutral atom, the number of electrons equals the number of protons. If an atom or group of atoms gains or loses electrons, it forms a particle with a net charge, called an ion. This question type first deals with ions made from single atoms. Losing negative electrons produces a positive ion, called a cation. Gaining electrons produces a negative ion, called an anion. When electrons are gained or lost, proton and neutron numbers in the nucleus stay the same.

Losing electrons forms positive ions; gaining electrons forms negative ions



Na: 11p, 11e; 2,8,1 Na⁺: 11p, 10e; 2,8

Cl gains 1 electron: 17p, 18e → Cl⁻, with arrangement 2,8,8.

Electrons are negatively charged, so losing negative charge leaves an overall positive charge. The nucleus remains unchanged. In the chlorine example below, the electron change is in the opposite direction.

For each calculation, use Z to find the proton number, A–Z to find the neutron number, and finally the upper-right charge to find the electron number:

Ionic charge	+1	+2	−1	−2
Electrons	$Z-1$	$Z-2$	$Z+1$	$Z+2$

If the charge is written as only + or −, the omitted charge number is 1.
For example, F^- has A=19 and Z=9. Its nucleus contains nine protons, and its neutron number is:

$$\text{number of neutrons} = 19 - 9 = 10$$

An ion with a single negative charge has one more electron than the neutral atom, so:

$$\text{number of electrons} = 9 + 1 = 10$$

Nine positive charges and ten negative charges leave a net charge of −1, matching F^-.

If a question gives only an element's name and no mass number for a particular atom, its proton number can still be determined. If it also specifies an atom, its electron number can be determined. However, this information generally does not uniquely determine the neutron number. In a multiple-choice question, proton and electron numbers may already rule out every other row. There is no need to invent a particular nuclide that the question has not specified.

How are electrons arranged in shells? An electronic configuration lists the number of electrons in each shell, from the innermost outwards, separated by commas. For example, 2,8,1 describes three shells in one sodium atom, containing 2, 8 and 1 electrons respectively: 11 in total. It does not describe three atoms. For ground-state atoms of the first 20 elements covered here, use the arrangements below to read diagrams and write configurations. Do not treat a maximum of eight electrons in the third shell as a universal rule for every element.

Z and element	Configuration	Z and element	Configuration
1 H Hydrogen	1	11 Na Sodium	2,8,1
2 He Helium	2	12 Mg Magnesium	2,8,2
3 Li Lithium	2,1	13 Al Aluminium	2,8,3
4 Be Beryllium	2,2	14 Si Silicon	2,8,4
5 B Boron	2,3	15 P Phosphorus	2,8,5
6 C Carbon	2,4	16 S Sulfur	2,8,6
7 N Nitrogen	2,5	17 Cl Chlorine	2,8,7
8 O Oxygen	2,6	18 Ar Argon	2,8,8
9 F Fluorine	2,7	19 K Potassium	2,8,8,1
10 Ne Neon	2,8	20 Ca Calcium	2,8,8,2

For an ion, calculate the total electron number before writing the configuration. For example, Mg^{2+} has ten electrons, giving 2,8; Cl^- has 18 electrons, giving 2,8,8. When a positive ion forms, electrons are lost from the atom's outermost shell first. If that shell becomes empty, the ion has one fewer occupied shell than the atom. The change from Na, 2,8,1, to Na^+, 2,8, is an example.

What else can the Periodic Table tell us? For the main-group elements in Groups I to VII studied here, the group number gives the number of outer-shell electrons in the atom. Group I has one, and Group VII has seven. The Group VIII noble gases have full outer shells: He has two electrons in its first shell, while Ne and Ar have eight in their outer shells. A period is a horizontal row in the Periodic Table. The period number gives the number of occupied electron shells in the atom. Na, 2,8,1, has three shells and one outer electron, so it is in Period 3, Group I. This relationship identifies the position of the **neutral atom**. The two remaining shells in Na^+ do not move sodium to Period 2.

Group I metals commonly lose their one outer-shell electron to form +1 ions. Group VII non-metals commonly gain one electron to form −1 ions. This connects directly to the rubidium question in this type: knowing that Rb is in Group I is enough to deduce that it loses one electron and forms Rb^+, without memorising all its inner shells. Section 2.4 explains how ions combine to form compounds.

Atoms and ions of different elements can share an electron arrangement



Ne: 10p, 10e F^-: 9p, 10e Mg^{2+}: 12p, 10e

All three have 10 electrons, arranged 2,8. Their different proton numbers identify different elements.

Count electrons first, then arrange them in shells. Identical electron arrangements do not mean that particles with different nuclei are the same element.

When finding particles with the same configuration, distinguish the same electrons from the same element.
Ne has 10p and 10e; F^- has 9p and 10e; Mg^{2+} has 12p and 10e. All three have the configuration 2,8, but their nuclei have different proton numbers. If asked to give an atom and an ion, first calculate the target electron number. Then find candidates with exactly that many electrons and finally check whether a positive or negative ion is required. For example, Na^+ forms when a neutral sodium atom with 11 electrons loses one:

$$\text{electrons in } Na^+ = 11 - 1 = 10$$

A neutral atom with the same configuration must also have ten electrons, so its proton number is ten: it is Ne. If a negative ion is required, F^- can be used:

$$\text{electrons in } F^- = 9 + 1 = 10$$

Retain the upper-right negative charge when writing F^-. If the target changes to 18 electrons, use the same approach: count first, then find the required atom and ion separately.

Recognising the Question Distinguish atomic number, mass number and ionic charge. Count in the order p→n→e, then use the total electron number to write the configuration. To find an atom or ion with the same configuration, keep the target electron number fixed and separately check the proton number and the required positive or negative charge.

Worked Example 1 (2025·May/June 0620/41 Q2(a)) · Level 1

Atoms are made of electrons, neutrons and protons.

(a) Complete Table 2.1.

Table 2.1

particle	relative charge	relative mass
electron		$\frac{1}{1840}$
neutron	0	
proton		

Solution

The question already gives an electron's relative mass as 1/1840 and a neutron's relative charge as 0. Complete the same properties table by distinguishing charge from relative mass. An electron has relative charge −1; a neutron has relative mass approximately 1; and a proton has charge +1 and relative mass approximately 1.

The completed table is:

particle	relative charge	relative mass
electron	−1	1/1840
neutron	0	1
proton	+1	1

An electron is much lighter than a nucleon but still has mass. A neutron's zero charge does not mean zero mass. Retain the question's stated value of 1/1840 rather than replacing it with another approximation.

Worked Example 2 (2025·May/June 0620/43 Q2(b)(i)) · Level 3

Atoms are made of electrons, neutrons and protons.

(b) Atoms of the same element are known as isotopes.

$^{32}_{16}S$ and $^{34}_{16}S$ are isotopes of sulfur.

(i) Complete Table 2.1 to show the number of electrons, neutrons and protons in one atom or ion of these isotopes.

Table 2.1

isotope	electrons	neutrons	protons
$^{34}_{16}S$			
$^{32}_{16}S^{2-}$			

Solution

Step 1: read the proton number of each nucleus. Both sulfur symbols have 16 at the lower left, so both rows have 16 protons. The upper-left values, 34 and 32, are mass numbers; do not enter them in the proton column.
Step 2: calculate each neutron number. The upper-right charge 2− does not change the nucleus:

neutrons in $^{34}S = 34 - 16 = 18$

neutrons in $^{32}S^{2-} = 32 - 16 = 16$

Step 3: use the charge to find the electron number. ^{34}S has no charge shown, so it is a neutral atom with 16 electrons. $^{32}S^{2-}$ has gained two electrons compared with a neutral sulfur atom:

electrons in $^{32}S^{2-} = 16 + 2 = 18$

Fill the original table in its actual column order: electrons, neutrons, protons.

Particle	electrons	neutrons	protons
^{34}S	16	18	16
$^{32}S^{2-}$	18	16	16

Check the second row: 16 positive charges and 18 negative charges give −2 overall, matching S^{2-}. A negative charge does not mean fewer electrons.

Worked Example 3 (2025·May/June 0620/41 Q2(c)) · Level 3

Atoms are made of electrons, neutrons and protons.

(c) A calcium atom has the electronic configuration of 2,8,8,2.

Give the formula of one atom and one negative ion that has the same electronic configuration as Ca^{2+}.

- atom
- negative ion

Solution

First calculate the electron number of the target, Ca^{2+}. A calcium atom has configuration 2,8,8,2, giving 20 electrons in total. Ca^{2+} loses the two outermost electrons and has 18 remaining, with configuration 2,8,8.

Next find a neutral atom. For a neutral atom, proton number equals electron number, so look for Z=18. The Periodic Table identifies argon: write Ar.

> **Finally, find the negative ion specified.** Chlorine has Z=17. A chlorine atom gains one electron to form Cl^-, with 17+1=18 electrons and configuration 2,8,8. Cl^- is therefore suitable. Do not give a positive ion with 18 electrons: the question requires a negative ion.
>
> A suitable answer is atom: Ar; negative ion: Cl^-. S^{2-} also has 16+2=18 electrons, and P^{3-} has 15+3=18; both are accepted by the official mark scheme. Give one of each particle type as requested. An element name and another element symbol without its required charge do not constitute an atom-and-ion answer.

Practice

1. (2025·May/June 0620/43 Q2(a)) · Level 1

Atoms are made of electrons, neutrons and protons.

(a) State which of these particles are found in the nucleus of an atom.

2. (2025·May/June 0620/11 Q4) · Level 1

Which part of an atom has a relative mass of 1 and a relative charge of 0?

A electron

B neutron

C nucleus

D proton

3. (2025·May/June 0620/12 Q5) · Level 1

Which statement defines the atomic number of an element?

A the group number of the element

B the mass of one atom of the element

C the number of particles in the nucleus of one atom of the element

D the number of protons in the nucleus of one atom of the element

4. (2025·May/June 0620/11 Q3) · Level 1

An atom of element Q contains 19 electrons, 19 protons and 20 neutrons.

What is Q?

A calcium

B potassium

C strontium

D yttrium

5. (2025·May/June 0620/13 Q4) · Level 2

An atom of the element erbium is represented by $^{167}_{68}Er$.

Which row shows the number of protons, neutrons and electrons in this atom?

	protons	neutrons	electrons
A	68	99	68
B	68	99	99
C	99	68	68
D	99	68	99

6. (2025·May/June 0620/22 Q3) · Level 2

Which row gives the number of protons, electrons and neutrons in an atom of zinc?

	protons	electrons	neutrons
A	30	30	35
B	30	35	35
C	35	30	30
D	35	35	30

7. (2025·May/June 0620/13 Q5) · Level 2

Rubidium is in Group I of the Periodic Table and bromine is in Group VII of the Periodic Table.

Rubidium reacts with bromine to form an ionic compound.

Which row shows the electron change taking place for rubidium and the correct formula of the rubidium ion formed?

	electron change	formula of ion formed
A	electron gained	Rb^+
B	electron gained	Rb^-
C	electron lost	Rb^+
D	electron lost	Rb^-

8. (2025·May/June 0620/21 Q3) · Level 3

Which atom has an equal number of protons, neutrons and electrons?

A ^{40}Ar **B** ^{1}H **C** ^{23}Na **D** ^{14}N

9. (2025·May/June 0620/22 Q4) · Level 3

Which particles have the electronic configuration 2,8,8?

1 an argon atom, Ar

2 an aluminium ion, Al^{3+}

3 a sodium ion, Na^{+}

4 a chloride ion, Cl^{-}

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

10. (2025·May/June 0620/41 Q2(b)(i)) · Level 3

Atoms are made of electrons, neutrons and protons.

(b) Atoms of the same element are known as isotopes.

$^{39}_{19}K$ and $^{41}_{19}K$ are isotopes of potassium.

(i) Complete Table 2.2 to show the number of electrons, neutrons and protons in one atom or ion of these isotopes.

Table 2.2

isotope	electrons	neutrons	protons
$^{39}_{19}K$			
$^{41}_{19}K^{+}$			

11. (2025·May/June 0620/42 Q5(d)(ii)) · Level 3

Alkali metals are reactive elements.

(d) Rubidium has only two naturally occurring atoms, ^{85}Rb and ^{87}Rb.

(ii) Complete Table 5.1 to show the number of protons, neutrons and electrons in the atom and ion of rubidium shown.

Table 5.1

	^{85}Rb	$^{87}Rb^{+}$
protons		
neutrons		
electrons		

12. (2025·May/June 0620/43 Q2(c)) · Level 3

Atoms are made of electrons, neutrons and protons.

(c) Ions are atoms or groups of atoms that have gained or lost one or more electrons.

An oxygen atom has the electronic configuration of 2,6.

Give the formula of **one** atom and **one** positive ion that has the same electronic configuration as O^{2-}.

- atom
- positive ion

2.3 Isotopes

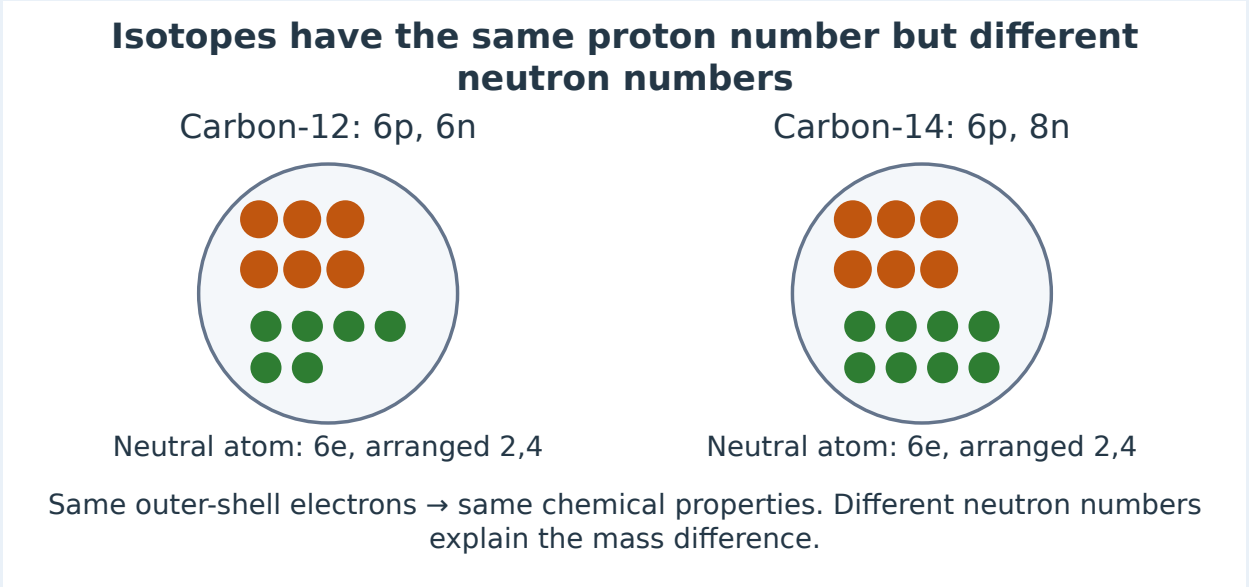
Type 1: Comparing isotopes and calculating relative atomic mass from abundance

Key Ideas Atoms of the same element have a fixed proton number, but their neutron numbers can differ.

Isotopes are **atoms of the same element with the same number of protons but different numbers of neutrons**.
Mass number depends on both types of particle in the nucleus:

$$\text{mass number} = \text{proton number} + \text{neutron number}$$

Isotopes therefore also have different mass numbers. However, different atomic masses alone are insufficient: first establish that the atoms belong to the same element.



Orange dots represent protons and green dots represent neutrons. This counting diagram does not show the actual internal arrangement of a nucleus.

Carbon-12 and carbon-14 each have six protons, so both are carbon. Carbon-12 has six neutrons; carbon-14 has eight. As neutral atoms, both also have six electrons and configuration 2,4. Chemical reactions mainly involve electrons, especially outer-shell electrons. Isotopes of an element therefore have the same electronic configuration and the same chemical properties. Their different masses do not make one react like oxygen while the other continues to react like carbon.

Read every position in the symbol. A at the upper left is mass number, Z at the lower left is proton number, and the upper-right symbol is ionic charge. When comparing two entries, first check whether Z is the same,

then whether A differs. If both Z and A are the same, the entries describe the same nuclide. If Z differs, the elements differ. If the upper-right charges differ, distinguish atoms from ions first: different electron numbers cannot replace the different neutron numbers required for isotopes. The previous question type's Mg^{2+} notation diagram introduced all these positions; now apply them to comparing two particles.

For example, chlorine-35 and chlorine-37 both have 17 protons, but have 18 and 20 neutrons respectively, so they are isotopes. Chlorine-35 and argon-40 have 17 and 18 protons respectively, so they are different elements. Forming Cl^- from chlorine-35 changes its electron number; it does not turn it into chlorine-37.

Why are relative atomic masses in the Periodic Table often not whole numbers? Atoms are extremely light, so comparing their masses directly in kilograms is inconvenient. Chemistry uses one-twelfth of the mass of one carbon-12 atom as a reference. Relative atomic mass, Ar, is the ratio of the **average mass** of an element's atoms, accounting for its isotopes, to this reference. It is a relative ratio, with no mass unit such as g. The reference is carbon-12, not an arbitrary carbon atom or the total mass of a carbon sample.

An isotope's abundance is its proportion of the total number of atoms in a sample. For example, 80% of the atoms having relative mass 35 means that about 80 out of every 100 atoms are that isotope. It does not mean that each such atom has a mass of 35×80. A more numerous isotope has a greater effect on the average, so weight the masses by their proportions.

Abundance is a number proportion: weight each mass by its population



8 lighter atoms 2 heavier atoms

Mean = (35 × 8 + 37 × 2) ÷ 10 = 35.4

Assume 80% mass-35 and 20% mass-37 in this sample; these are not stated natural abundances.

The more abundant isotope has the greater effect on the average. Averaging 35 and 37 once each incorrectly assumes that the two isotopes each make up half the sample.

For example, a sample contains eight atoms of relative mass 35 and two of relative mass 37. First calculate their total relative mass:

$$35 \times 8 + 37 \times 2 = 354$$

Then divide by the total of ten atoms to obtain the average:

$$A_r = \frac{354}{10} = 35.4$$

Adding 35 and 37 and dividing by two would give 36, but would silently assume equal numbers of the two atoms, contrary to the ratio 8:2. Use the stated relative mass of each isotope. Do not replace an isotope's mass with the average A_r of the whole element.

For percentage abundances, add each isotope's relative mass × percentage abundance, then divide by 100. The sample above has abundances of 80% and 20%, so:

$$A_r = \frac{35 \times 80 + 37 \times 20}{100} = 35.4$$

If decimal proportions 0.80 and 0.20 are used instead, do not divide by 100 again:

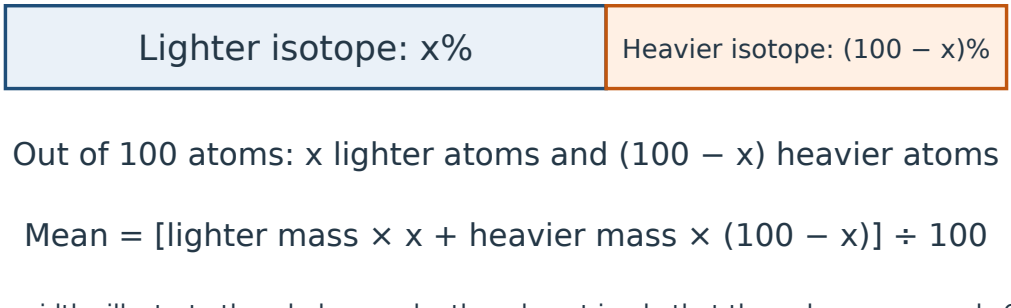
$$A_r = 35 \times 0.80 + 37 \times 0.20 = 35.4$$

Both forms give the same result. Check that the average lies between 35 and 37 and is closer to 35, the isotope with 80% abundance.

Calculate the unrounded value first, then round as requested. One decimal place means one digit after the decimal point, as in 39.2. For three significant figures, count three digits from the first non-zero digit, then check whether the next digit is at least 5. Both trailing zeros in 50.0 express the stated precision. The value 240.04 rounds to 240 at three significant figures; writing 2.40×10^2 makes those three figures explicit. Avoid heavily rounding individual products before adding them, as this can alter the final result.

How can abundance be found from the average? If there are only two isotopes, their percentage abundances must total 100%. If the lighter isotope has abundance x%, the other has abundance (100−x)%. Substitute these into the same weighted-average relationship to obtain an equation with one unknown.

Finding abundance: the two populations must add up to the whole sample



Writing the other abundance as 100−x leaves only one unknown in the weighted average. The illustrative bar widths are not the data to be calculated.

For example, two isotopes have relative masses 10 and 12, and their average is 10.6. Let the mass-10 isotope have abundance x%, so the mass-12 isotope has abundance (100−x)%. Substitute into the weighted average:

$$\frac{10x + 12(100 - x)}{100} = 10.6$$

Multiply both sides by 100, then expand the brackets:

$$10x + 1200 - 12x = 1060$$

Collect the x terms and rearrange:

$$1200 - 2x = 1060 \quad \Rightarrow \quad 2x = 140$$

Thus x=70: the lighter isotope has abundance 70% and the heavier one 30%. Substitute back to check:

$$10 \times 0.70 + 12 \times 0.30 = 10.6$$

The average agrees, and the two percentage abundances also total 100%.

If the average is exactly halfway between the two masses, the abundances are equal. Use this as a check; do not automatically write 50% when the average is not at the midpoint. The europium example demonstrates the midpoint case. The later rubidium exercise places the average closer to the lighter isotope.

An integer average does not prove that only one isotope is present



50% of each: 151 × 0.50 + 153 × 0.50 = 152

The average 152 is an integer, but the sample still contains two isotopes.

The values 151 and 153 in this section's europium question provide a direct counterexample. An element's known isotopic composition requires separate evidence.

A whole-number average alone does not prove that there is only one isotope. Equal abundances of 151 and 153 give an average of 152, while there are still two isotopes. Whether an average contains a decimal fraction and how many isotopes exist are separate questions.

There is additional composition information for aluminium specifically: CIAAW states that the standard atomic weight of normal terrestrial aluminium is determined by aluminium-27. This concerns its normal natural composition; it does not mean that other artificial or trace radioactive aluminium nuclides do not exist. In the aluminium question here, distinguish this known fact about a particular element from what can generally be deduced from the number 27 alone. Do not extend one mark scheme's conclusion into a rule for every element. Source: CIAAW · Aluminium.

Recognising the Question For isotopes, keep the proton number the same and change the neutron number. For an average, weight each mass using the stated abundance. To work backwards to abundance, also use the total of 100%. Retain unrounded values until the end. A whole-number average alone does not establish a single isotope.

Worked Example 1 (2025·May/June 0620/11 Q5) · Level 2

Which row identifies the number of electrons, neutrons and protons in a particle which is an isotope of $^{11}_{5}B$?

	electrons	neutrons	protons
A	5	5	5
B	5	6	5
C	6	5	6
D	6	6	6

Solution

Step 1: interpret the given boron nuclide. The lower-left 5 means five protons. The upper-left 11 means 11 protons plus neutrons, so it has 11−5=6 neutrons. Another isotope must still have five protons but a different neutron number.

Step 2: check each row. A has 5p and 5n: the same element with a different neutron number, boron-10. B has 5p and 6n, so it is still the given boron-11 rather than another isotope. C and D both have 6p, so their element has changed and they are not boron isotopes.

Choose A. The headings are electrons, neutrons, protons, so the last column is the proton number needed first to identify the element. A also has 5e, consistent with a neutral boron atom.

Worked Example 2 (2025·May/June 0620/41 Q2(b)(ii)) · Level 1

Atoms are made of electrons, neutrons and protons.

(b) Atoms of the same element are known as isotopes.

$^{39}_{19}$K and $^{41}_{19}$K are isotopes of potassium.

(ii) Table 2.3 shows the relative masses and the percentage abundances of the two isotopes in a sample of potassium.

Table 2.3

relative mass of isotope	percentage abundance of isotope
39	90
41	10

Calculate the relative atomic mass of this sample of potassium to **one** decimal place.

Solution

Interpret percentage abundance as a proportion of atoms. The isotope of relative mass 39 has abundance 90%; the isotope of relative mass 41 has abundance 10%. Imagine 100 atoms: 90 contribute 39 each and ten contribute 41 each.
First calculate the total relative mass of the 100 atoms:
$39 \times 90 + 41 \times 10 = 3510 + 410 = 3920$
Then divide by the total number of atoms:

$$A_r = \frac{3920}{100} = 39.2$$

The question requires one decimal place, so write **39.2**, without g. The answer lies between 39 and 41 and is closer to the more abundant 39, as expected. Directly averaging 39 and 41 to obtain 40 would wrongly assume equal abundances.

Worked Example 3 (2025·May/June 0620/21 Q4) · Level 3

The only naturally occurring isotopes in a sample of europium are ^{151}Eu and ^{153}Eu. This sample of europium has a relative atomic mass, A_r, of 152.

What is the percentage of ^{151}Eu present in this sample to three significant figures?

A 0.987% **B** 50.0% **C** 76.0% **D** 98.7%

Solution

Step 1: define one unknown abundance. Only two isotopes, ^{151}Eu and ^{153}Eu, are stated. Let ^{151}Eu have abundance x%, so ^{153}Eu has abundance (100−x)%.
Step 2: set the weighted average equal to the given value, 152.

$$\frac{151x + 153(100 - x)}{100} = 152$$

Multiply both sides by 100 and expand:
$151x + 15300 - 153x = 15200$
Collect the x terms and rearrange:
$15300 - 2x = 15200 \quad \Rightarrow \quad 2x = 100$
Hence x=50.

Step 3: give the answer to three significant figures. The percentage abundance is **50.0%: choose B**. Substitute into the original relationship to check:

$$151 \times 0.500 + 153 \times 0.500 = 152$$

Both relative masses are one unit from 152, so equal abundances also satisfy the midpoint check.

The whole-number average 152 comes from two isotopes contributing together. It does not imply that only one isotope of mass 152 exists.

Practice

1. (2025·May/June 0620/12 Q6) · Level 1

What is different for isotopes of the same element?

A nucleon number

B number of electron shells

C number of electrons in the outer electron shell

D proton number

2. (2025·May/June 0620/42 Q5(d)(i)) · Level 1

Alkali metals are reactive elements.

(d) Rubidium has only two naturally occurring atoms, ^{85}Rb and ^{87}Rb.

(i) State the term given to these naturally occurring atoms of rubidium.

3. (2025·May/June 0620/43 Q2(b)(iii)) · Level 1

(iii) The relative atomic masses of all elements are compared to one atom of an isotope.

Identify this isotope.

4. (2025·May/June 0620/13 Q10) · Level 1

What is the definition of relative atomic mass?

A It is the average mass of the isotopes of an element compared to $\frac{1}{12}$th of the mass of an atom of ^{12}C.

B It is the total mass of the isotopes of an element compared to $\frac{1}{12}$th of the mass of an atom of ^{12}C.

C It is the average mass of an element compared to $\frac{1}{12}$th of the mass of an atom of ^{12}C.

D It is the total mass of an element compared to $\frac{1}{12}$th of the mass of an atom of ^{12}C.

5. (2025·May/June 0620/43 Q2(b)(ii)) · Level 1

Atoms are made of electrons, neutrons and protons.

(b) Atoms of the same element are known as isotopes.

$^{32}_{16}S$ and $^{34}_{16}S$ are isotopes of sulfur.

(ii) Table 2.2 shows the relative masses and the percentage abundances of the two isotopes in a sample of sulfur.

Table 2.2

relative mass of isotope	percentage abundance of isotope
32	95
34	5

Calculate the relative atomic mass of this sample of sulfur to **one** decimal place.

6. (2025·May/June 0620/41 Q2(b)(iii)) · Level 2

(iii) An isotope of aluminium has a nucleon number of 27.

Aluminium has a relative atomic mass of 27.

State what conclusion can be made from this information.

7. (2025·May/June 0620/22 Q5) · Level 2

A sample of element Q has two isotopes.

Their relative masses and abundances are shown.

relative mass of isotope	abundance / %
238	66
244	34

What is the relative atomic mass of this sample of Q to three significant figures?

A 240 **B** 241 **C** 242 **D** 243

8. (2025·May/June 0620/42 Q5(d)(iii)) · Level 3

Alkali metals are reactive elements.

(d) Rubidium has only two naturally occurring atoms, ^{85}Rb and ^{87}Rb.

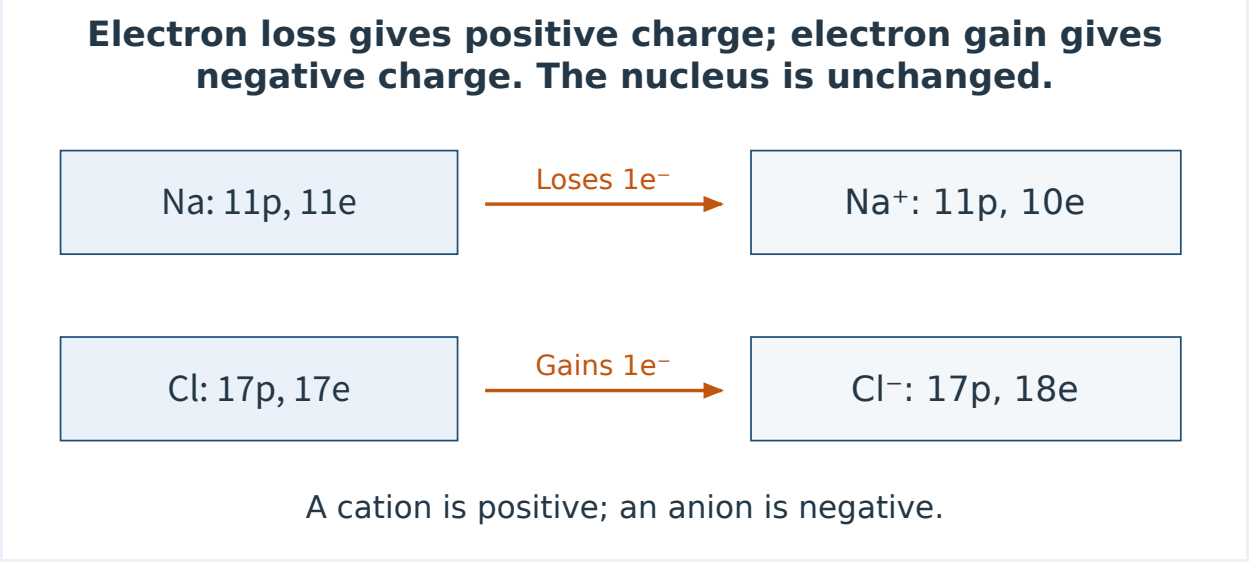
(iii) The relative atomic mass of rubidium to one decimal place is 85.5.

Determine the relative abundance of ^{85}Rb in rubidium. Express your answer as a percentage.

2.4 Ions and ionic bonding

Type 1: Using electron transfer to draw ions and dot-and-cross diagrams

Key Ideas A neutral atom has equal numbers of protons and electrons, so its positive and negative charges balance. An **ion** is an atom or group of atoms that has gained or lost electrons and therefore carries a charge. This question type first considers ions formed from single atoms. Losing electrons leaves more positive than negative charge and produces a **cation**. Gaining electrons gives more negative than positive charge and produces an **anion**. Determine charge by comparing protons and electrons, not by guessing from the name.

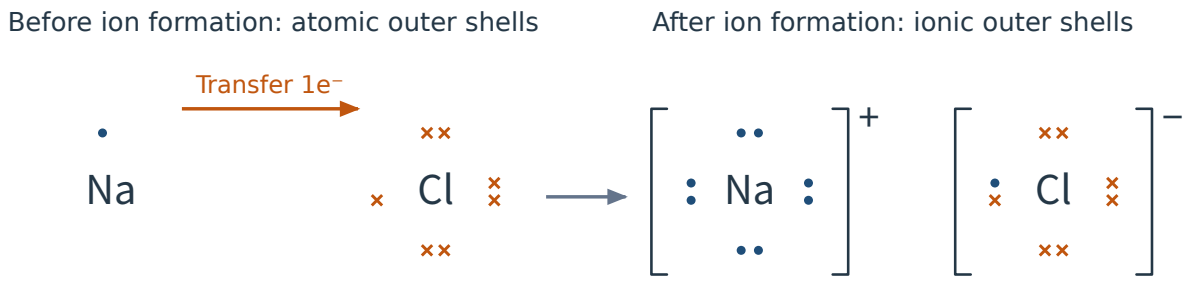


The arrows track changes in electron number. Each row retains its original proton number, so the identity of the element remains the same.

Forming an ion changes the electrons, not the nucleus. For example, Na has 11 protons. After losing one electron it still has 11 protons but only ten electrons, so it is Na^+. Cl has 17 protons; after gaining one electron it has 18 electrons, so it is Cl^-. An upper-right + means one electron fewer than in the atom, and 2+ means two fewer. A − means one extra electron, and 2− means two extra. These charges are neither mass numbers nor additional protons.

Begin with a simple pair to follow electron transfer. A sodium atom has configuration 2,8,1 and a chlorine atom 2,8,7. Sodium loses its one outer-shell electron and becomes 2,8. Chlorine gains that electron and becomes 2,8,8. Both have full outer shells. Group I metals usually lose one electron per atom, and Group VII non-metals usually gain one. For example, potassium and iodine form K^+ and I^-; the negative iodine ion is called iodide.

NaCl: one sodium atom transfers one electron to one chlorine atom



NaCl: total electrons lost = total electrons gained

The positive ion's outer shell was already a filled inner shell. Do not draw the transferred electrons back in.

Dots and crosses track electron origins. The ions do not share electron pairs.

Only the outer shells before and after ion formation are shown. Each ion has square brackets and a charge. The metal ion's full outer shell was an inner shell of the original atom.

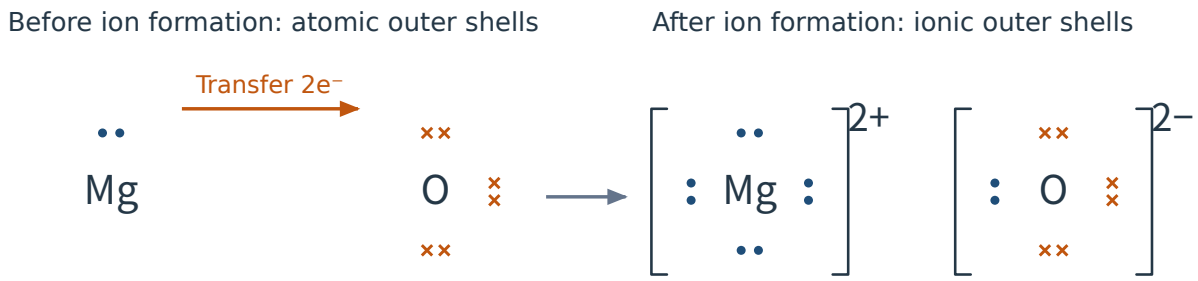
Dots and crosses in a dot-and-cross diagram track where electrons came from. They do not represent different kinds of electron. Here, the electron supplied by sodium is a dot and chlorine's original electrons are crosses. The Cl^- outer shell contains seven crosses plus one dot, giving eight electrons. Draw square brackets around each ion and write its charge at the upper right.

Remember that the outermost shell before and after ion formation may be a different shell. When Na loses the only electron in its outer shell, the already full shell inside becomes the new outer shell. Those eight electrons were present all along; they were not created in the reaction. When only outer shells are drawn, the number of visible dots before and after need not match. To check conservation of electrons, compare the complete electronic configurations.

Electron transfer forms charged ions; the ionic bond is the strong electrostatic attraction between oppositely charged ions. Electrostatic attraction means the attraction between positive and negative charges. Na^+ and Cl^- are held together by this attraction, not by sharing an electron pair. Many ions form a lattice extending in all directions, which the next question type examines. The dot-and-cross diagram here shows the ions' electrons and their simplest number ratio; it does not represent an isolated NaCl molecule.

If a metal loses two electrons, track each electron. Magnesium, Mg, has configuration 2,8,2. It loses two electrons to form Mg^{2+}, with configuration 2,8. Oxygen, O, has configuration 2,6. It gains two electrons to form O^{2-}, also with configuration 2,8. One Mg supplies two electrons to one O. The resulting charges have equal magnitudes and opposite signs, so magnesium oxide, MgO, has a simplest ion ratio of 1:1.

MgO: one magnesium atom transfers two electrons to one oxygen atom



MgO: total electrons lost = total electrons gained

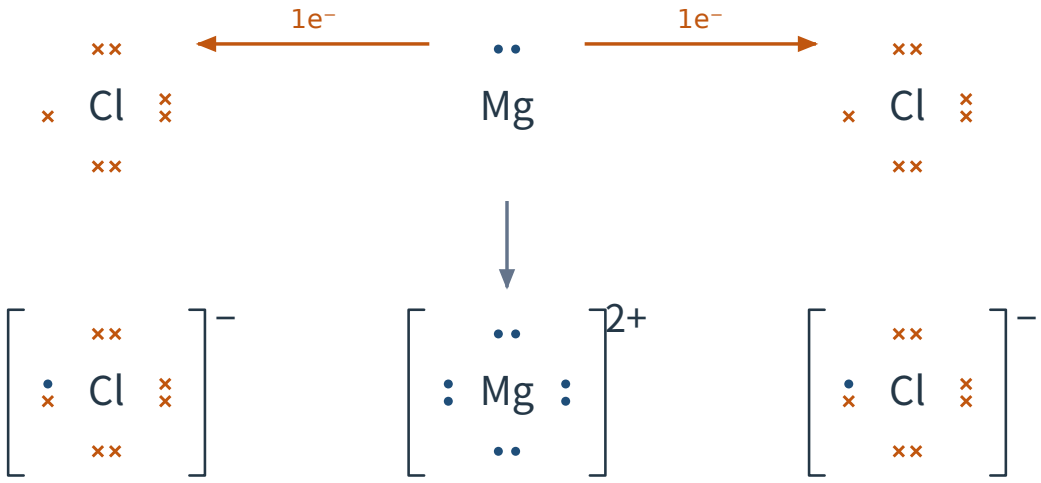
The positive ion's outer shell was already a filled inner shell. Do not draw the transferred electrons back in.

Dots and crosses track electron origins. The ions do not share electron pairs.

Only the outer shells before and after ion formation are shown. Each ion has square brackets and a charge. The metal ion's full outer shell was an inner shell of the original atom.

Replacing oxygen with chlorine changes the ratio. One Mg still loses two electrons, but each Cl needs only one. Two Cl atoms must therefore each accept one electron, forming one Mg^{2+} and two Cl^{-} ions. The compound is magnesium chloride, $MgCl_2$. The subscript 2 indicates two chloride ions; it does not change one chloride ion's charge to 2−.

$MgCl_2$: one magnesium loses $2e^-$; each of two chlorine atoms gains $1e^-$



One Mg^{2+} for two Cl^{-}: $(+2) + 2 \times (-1) = 0$

The subscript 2 counts two Cl ions; it does not give each Cl a 2− charge.

Each orange arrow transfers one electron. The charges on the three ions in the lower row add to zero. Only outer shells are shown.

For example, sulfur, S, has atomic number 16 and configuration 2,8,6. It gains two electrons to become S^{2-}. If it combines with Na, which loses only one electron per atom, two Na atoms must each supply one electron to form sodium sulfide, Na_2S. If it combines with Mg, which loses two electrons per atom, only one Mg is needed per S. First decide how many electrons each atom gains or loses, then make the total lost equal the total gained. One atom of each element is not always the correct combination.

The ratio must balance electrons lost and gained, and give zero total charge

NaCl	$1Na^+ + 1Cl^-$	$(+1)+(-1)=0$
MgO	$1Mg^{2+} + 1O^{2-}$	$(+2)+(-2)=0$
Na_2S	$2Na^+ + 1S^{2-}$	$2\times(+1)+(-2)=0$
$MgCl_2$	$1Mg^{2+} + 2Cl^-$	$(+2)+2\times(-1)=0$

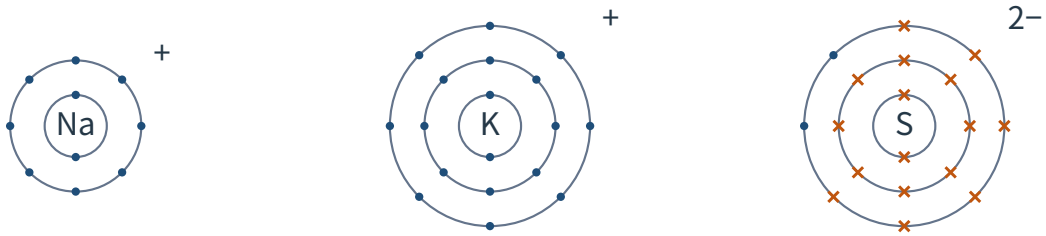
This is the simplest ion ratio, not a separate ionic molecule.

The same electrical-neutrality rule applies to different combinations of singly and doubly charged ions. Subscripts record ion numbers without changing each ion's charge.

Common Group II metal ions have charge 2+, while common Group VI non-metal ions have charge 2−. If other ionic charges are given directly, balance the total charge in the same way. For example, two Al^{3+} ions have total charge +6, exactly balancing three O^{2-} ions with total charge −6. This gives the 2:3 ratio in aluminium oxide, Al_2O_3. Chapter 3 develops the systematic writing of formulae from these ratios. Here, use them to decide how many ions to draw.

If part of an electron-shell diagram is provided, count the complete inner shells already printed. Na^+ ends with only two shells, 2,8. A K atom is 2,8,8,1 and loses one electron to form K^+ with three shells, 2,8,8. S^{2-} also has three shells, 2,8,8, but its charge comes from sulfur gaining two electrons. Having the same electron number as K^+ does not give it the same charge.

Before completing a template, determine the final number of shells and electrons in each



Na^+: 2,8; 10e in total

K^+: 2,8,8; 18e in total

S^{2-}: 2,8,8; 18e in total

Keep the printed inner shells. Add only the missing complete outer shell and charge.

All final electron shells are shown in the three diagrams. The two dots in sulfur's outer shell are transferred electrons. Do not add the missing eight electrons to an inner shell that is already full.

Suppose a K template already shows two electrons in its innermost shell and eight in the next. The final ion needs three shells, so add a complete third shell outside them with eight electrons. Do not put eight extra dots into the already full second shell. If a Na template shows only the first shell with two electrons, add a second shell with eight. Finally check each element symbol, the number of shells, the electrons in every shell, the dot-and-cross origins and the upper-right charge. Also check that the total charge on all ions is zero.

Recognising the Question Write the atoms' electronic configurations and how many electrons each must gain or lose, then balance the total transfer. Read the inner shells already printed on the template. Add only the missing outer shells, dots and crosses with the correct origins, and the charge at the upper right of each bracket.

Worked Example 1 (2025·February/March 0620/42 Q2(b)) · Level 3

This question is about ionic compounds.

(b) Potassium sulfide, K_2S, is an ionic compound.

Complete the dot-and-cross diagram in Fig. 2.1 of the ions in potassium sulfide.

Show the charges on the ions.

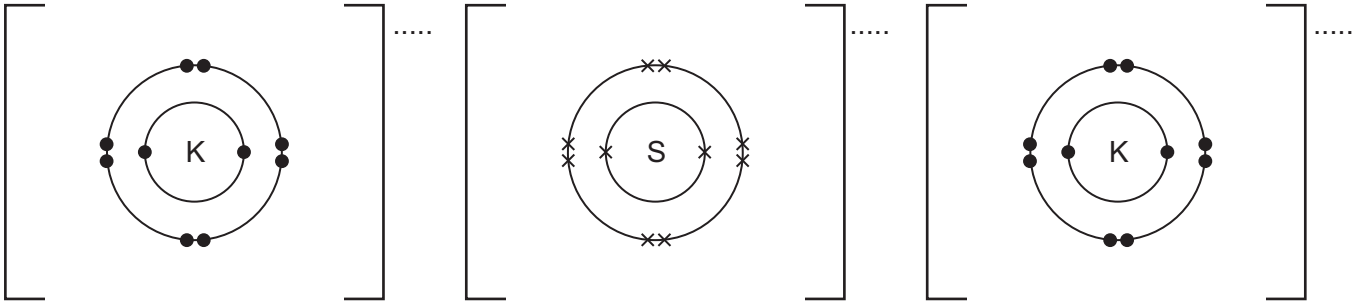


Fig. 2.1

Solution

First determine which ions form. Potassium, K, has atomic number 19, so its neutral atom has configuration 2,8,8,1. It loses one outer-shell electron to form K^+ with configuration 2,8,8. Sulfur, S, has atomic number 16 and neutral-atom configuration 2,8,6. It needs two electrons to form S^{2-}, also with configuration 2,8,8.

Second, balance the electron transfer. One K can supply one electron, while one S needs two. Two K atoms therefore each supply one electron to the same S atom. This forms two K^+ ions and one S^{2-} ion, with total charge $(+1)+(-2)+(+1)=0$, matching the 2:1 ratio in K_2S.

Third, inspect what the template already shows. Both K diagrams have two dots in the first shell and eight in the second; do not add electrons to these shells. Add a third shell around each K and place eight dots on it. The central S diagram also already has two inner shells. Add a third shell containing sulfur's six original crosses and the two dots transferred from potassium.

Fourth, write the charges at the marked upper-right positions outside the existing brackets: + for the left K, 2− for the central S and + for the right K. Do not turn the subscript 2 in K_2S into a 2+ charge on each potassium ion. Each of the two potassium ions has a single positive charge.

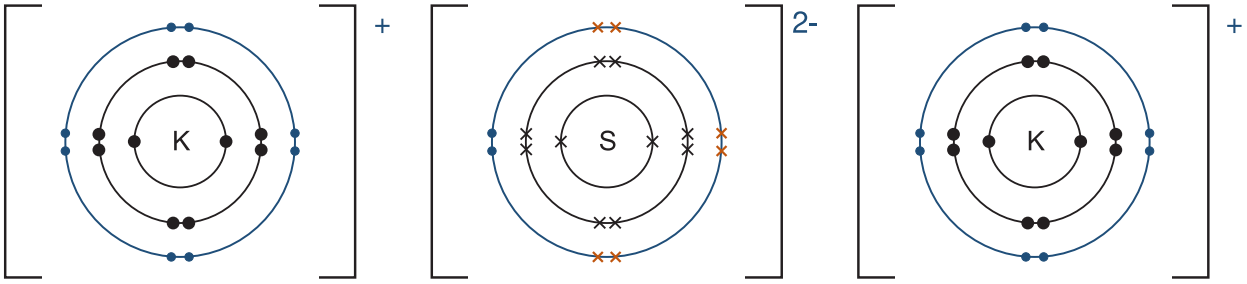


Fig. 2.1

Add the missing shells outside the original template: eight dots in the third shell of each K, and six crosses plus two dots in the third shell of S. The charges are +, 2− and + respectively.

Finally check electron conservation. Each of the three ions has 18 electrons, giving 54 in total. The original two K atoms and one S atom also had 19+16+19=54 electrons. Electrons have been transferred, not created or destroyed. The diagram shows ionic configurations and their number ratio, not a molecule in which three ions share electron pairs.

Practice

1. (2025·May/June 0620/12 Q7) · Level 1

When sodium chloride is formed from its elements, each chlorine atom .....1..... one .....2..... .

Which words correctly complete gaps 1 and 2?

	1	2
A	gains	electron
B	gains	proton
C	loses	electron
D	loses	proton

2. (2025·May/June 0620/42 Q5(c)) · Level 3

Alkali metals are reactive elements.

(c) Sodium sulfide, Na_2S, is an ionic compound.

Complete Fig. 5.1 to show the electronic configurations of the ions in sodium sulfide.

Show the charges on the ions.

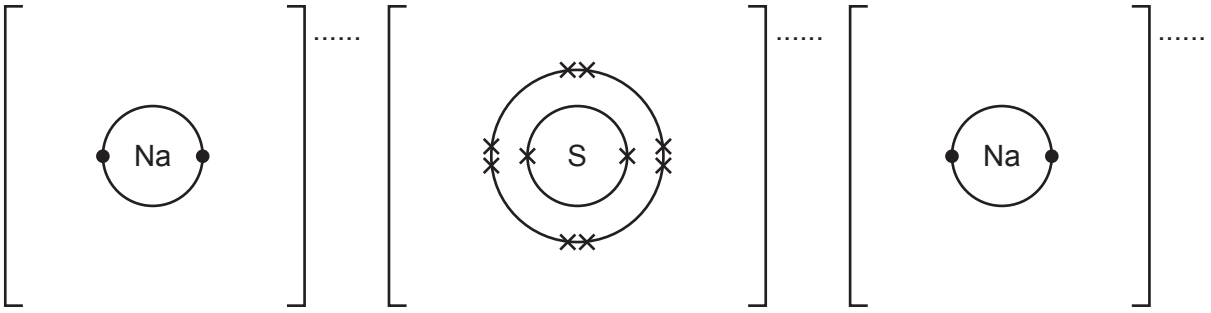
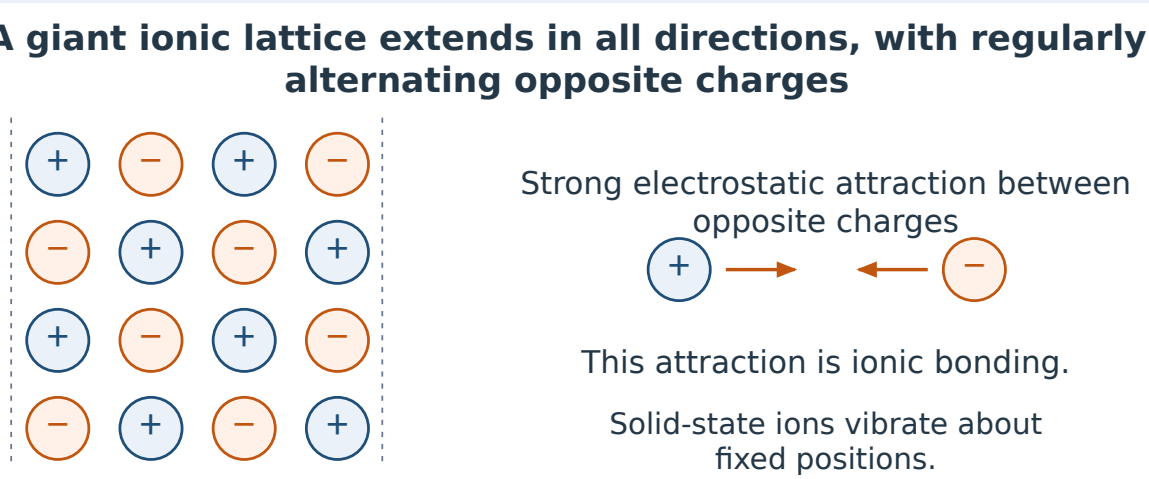


Fig. 5.1

Type 2: Explaining properties using ionic lattices and ion movement

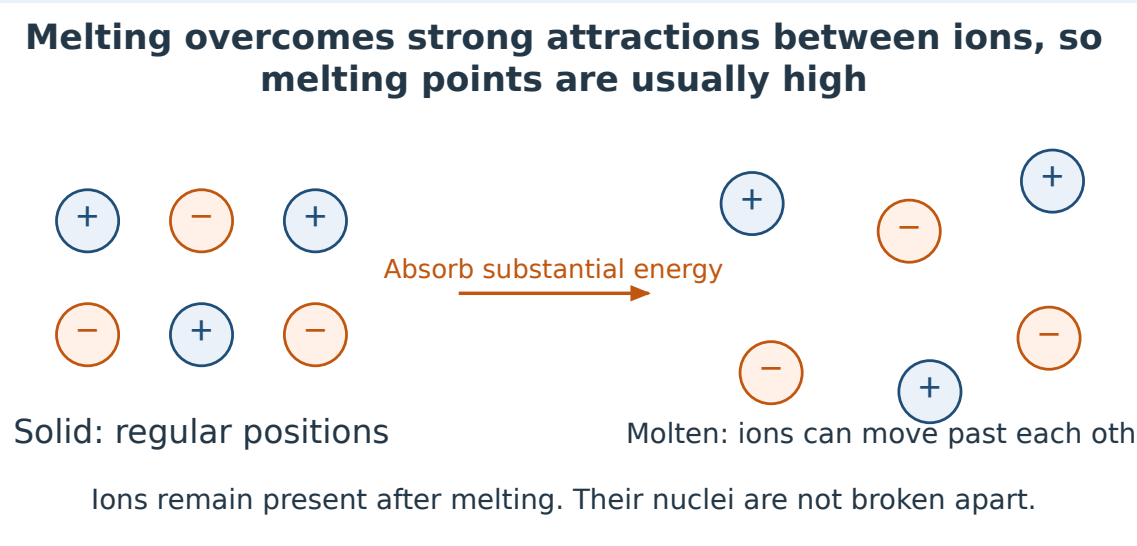
Key Ideas The previous question type established the ions' electron arrangements. Now widen the view from a few ions to the whole material. A **giant ionic lattice** is a regular, alternating arrangement of many positive and negative ions extending in all directions. Giant means a continuous structure rather than a collection of separate small molecules. Lattice refers to the regular arrangement.



The left-hand diagram shows the regular alternating arrangement. The right-hand arrows show only attraction between opposite charges; they do not mean that ions can move freely in the solid.

Strong electrostatic attractions between neighbouring oppositely charged ions are **ionic bonds**. Like charges repel. In a stable lattice, the arrangement allows positive and negative ions to attract one another and maintain regular positions. A flat grid represents only part of the three-dimensional lattice, not an entire crystal limited to the small section shown.

Melting the solid requires enough energy to overcome the strong attractions between ions in the lattice. These strong attractions require substantial energy, so ionic compounds usually have high melting and boiling points. Specify that the strong attraction is between oppositely charged ions. Do not describe strong covalent bonds or forces between ionic molecules.

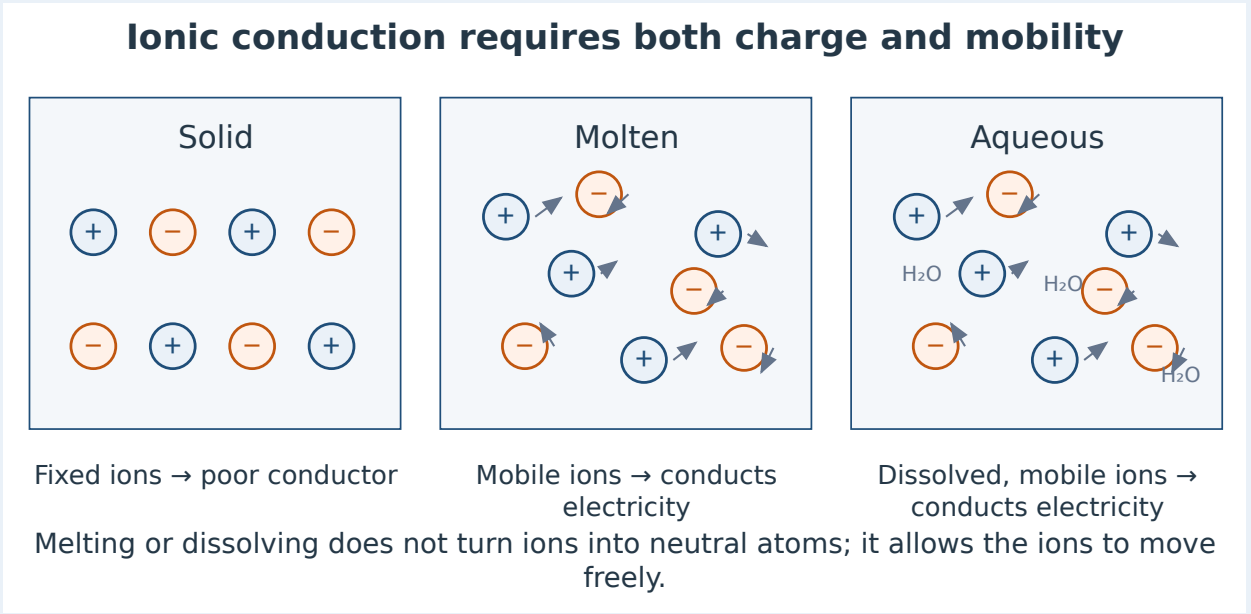


Orange arrows represent energy absorption to overcome strong electrostatic attractions in the lattice. The ions retain their charges after melting.

Melting allows ions to leave their fixed lattice positions and move relative to one another. They remain charged; they do not all turn back into neutral atoms. Boiling separates particles further and also requires considerable energy. Link the explanation as giant lattice → strong attractions between ions → large energy requirement.

For electrical conduction, ask another question: can the charged particles move? Electric current requires charge to move through the substance. Solid ionic compounds contain charged ions, but these only vibrate near fixed positions and cannot move freely through the lattice, so the solid is a poor conductor. **Molten** means that

the solid has melted into a liquid. Molten ions can move and carry charge, so the liquid conducts. The charge carriers are ions, not free electrons appearing from nowhere.

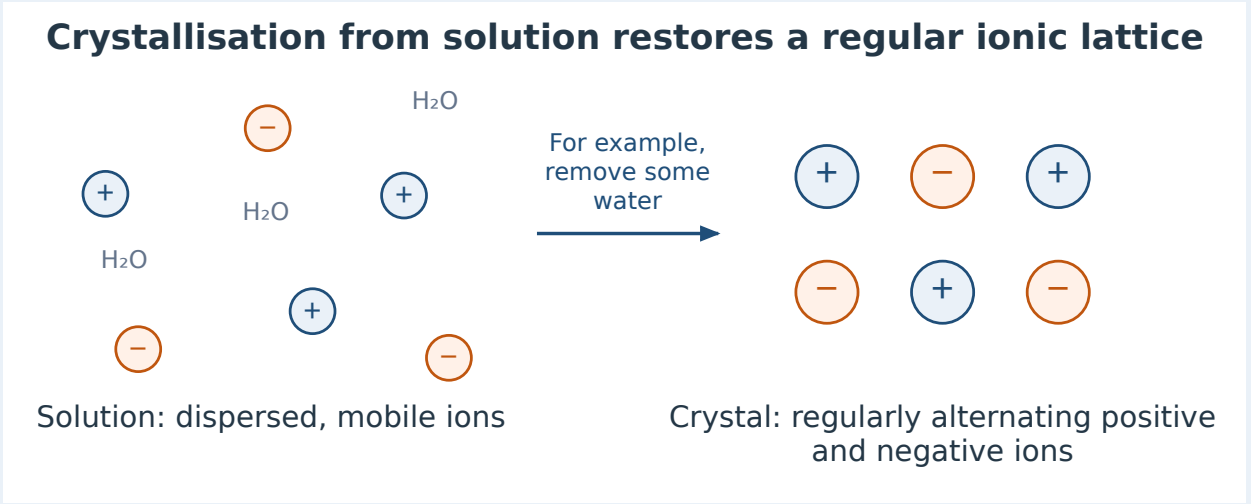


Arrows show that ions can move in the liquid or solution. When connected in a circuit, mobile ions carry charge. Ions in the solid cannot move freely through the lattice.

In an **aqueous solution**, the dissolved positive and negative ions can also move, allowing conduction. For example, potassium chloride KCl, sodium chloride NaCl and potassium iodide KI are soluble in water. Strong bonding in a solid is not enough to conclude that it is insoluble. Water molecules also attract ions during dissolving, and solubility must be judged for the particular substance. This does not mean that all ionic compounds are readily soluble or that adding water makes every substance a good conductor.

For example, divide sodium chloride into three samples: solid salt, molten salt and a clear salt solution. All contain Na^+ and Cl^-, but ions in the first cannot move freely through the lattice, whereas those in the other two can move. The solid is therefore a poor conductor, while the molten salt and solution conduct. The claim that solid salt must conduct because it contains positive and negative charges omits the essential condition that those charges must be able to move.

Crystallisation is the formation of a regular crystal arrangement. When an ionic compound crystallises from solution, for example after some water is removed, its ions return to a regular alternating positive-and-negative arrangement to form a giant lattice. Do not describe a collection of separate small NaCl or KCl molecules.



This example shows crystallisation after some water is removed. The crystal contains alternating ions, not paired small molecules.

Finally, distinguish two electron comparisons that are easily confused. When a metal atom loses electrons to form a positive ion, that ion has fewer electrons than its own atom. This does not prove that it has fewer electrons in total than a neighbouring negative ion.

For example, K has atomic number 19, so K^+ has 18 electrons. Cl has atomic number 17, so Cl^- also has 18 electrons. Both have full shells, but opposite charges. A reliable counterexample like this disproves a universal statement that metal ions in all ionic compounds have fewer electrons than their non-metal ions.

Recognising the Question Identify a lattice of regularly alternating positive and negative ions. Explain melting and boiling points using attraction strength and the energy needed to overcome it. For conductivity, check both that ions are charged and whether they can move in the solid, molten material or solution.

Worked Example 1 (2025·May/June 0620/21 Q5) · Level 2

Which statement about potassium chloride is correct?

A It has a high melting point because it is a giant structure with strong covalent bonds.

B It conducts electricity when solid because of its oppositely charged ions.

C It conducts electricity when molten because its ions can move.

D It is insoluble in water because of its strong bonding.

Solution

Potassium chloride, KCl, forms a giant ionic lattice of K^+ and Cl^- ions. In the solid, the ions are charged but restricted to fixed lattice positions, so they cannot move through the material to carry a current. Once molten, the ions can move and carry charge.

C correctly states that molten potassium chloride conducts because its ions can move. Choose C. A incorrectly calls the strong electrostatic attraction between ions strong covalent bonding. B ignores the inability of solid-state ions to move freely. D is also incorrect: KCl dissolves in water. Strong bonding in the solid alone does not establish that a compound is insoluble.

Always check two conditions for conduction: whether charged particles are present, and whether they can move in the stated physical state. Seeing the word ions is not enough to conclude that a solid conducts.

Practice

1. (2025·May/June 0620/11 Q6) · Level 3

Which statements about potassium iodide are correct?

1 It is formed from potassium anions and iodide cations.

2 It is a good electrical conductor when molten or in aqueous solution.

3 Potassium atoms share electrons with iodine atoms.

A 1 and 3 **B** 1 only **C** 2 and 3 **D** 2 only

2. (2025·May/June 0620/23 Q6) · Level 3

Which statement about ionic compounds is correct?

A Giant lattices are formed when ionic compounds crystallise from solution.

B In all ionic compounds, the metal ions have fewer electrons than the non-metal ions.

C In a giant lattice, ions of the same charge are closer together than ions of opposite charge.

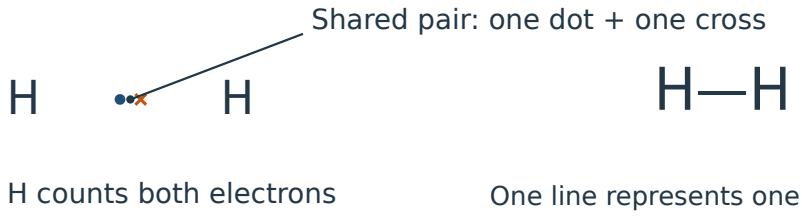
D When ionic compounds melt, strong intermolecular forces are broken.

2.5 Simple molecules and covalent bonding

Type 1: Drawing single and multiple bonds and lone pairs using shared electron pairs

Key Ideas Some atoms combine by **sharing electrons** rather than transferring electrons completely to another atom. A pair of electrons shared between two atoms forms a **covalent bond**. Both positively charged nuclei attract the shared pair, holding the atoms together. Include the shared electrons when checking the outer shell of either atom. The pair does not have to belong to only one atom at a time.

A covalent bond is a pair of electrons shared between two atoms



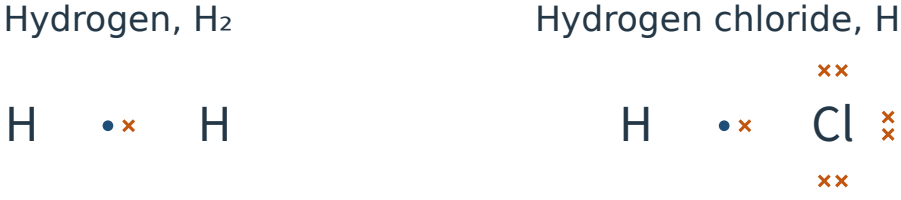
Dots and crosses show electron origins. They are not two different kinds of electron.

Both nuclei attract the shared electron pair, holding the atoms together. Include the same shared pair when checking the outer shell of each atom.

A dot-and-cross diagram represents one atom's original electrons as dots and the other's as crosses to track their origins. Electrons themselves are not different types. One dot plus one cross is one electron pair, not two bonds. A short line can represent this pair, as in H—H. If a question asks for outer-shell electrons only, do not add inner-shell electrons.

First establish the original outer-shell electron numbers. H has one, C four, N five, O six and Cl seven. These values follow from the earlier electronic-configuration table. Hydrogen's first shell is full with only two electrons. The other atoms considered here share electrons to achieve a noble-gas-like outer shell of eight electrons. A pair that does not take part in bonding and remains on one atom is called a **lone pair**. Lone pairs are still outer-shell electrons and must be included in the diagram.

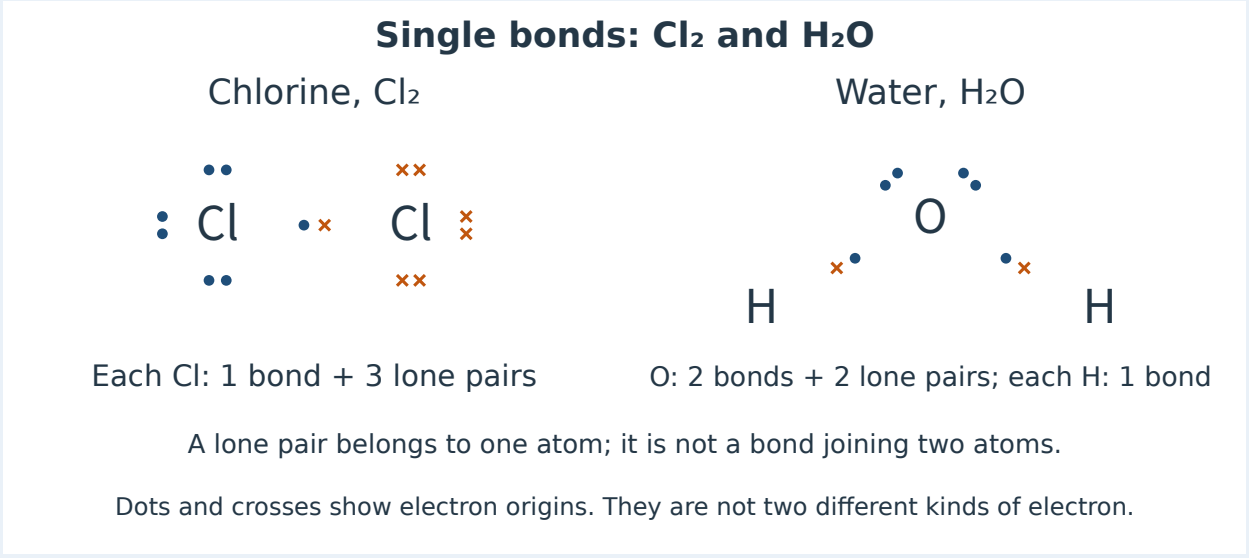
Single bonds: H_2 and HCl



Dots and crosses show electron origins. They are not two different kinds of electron.

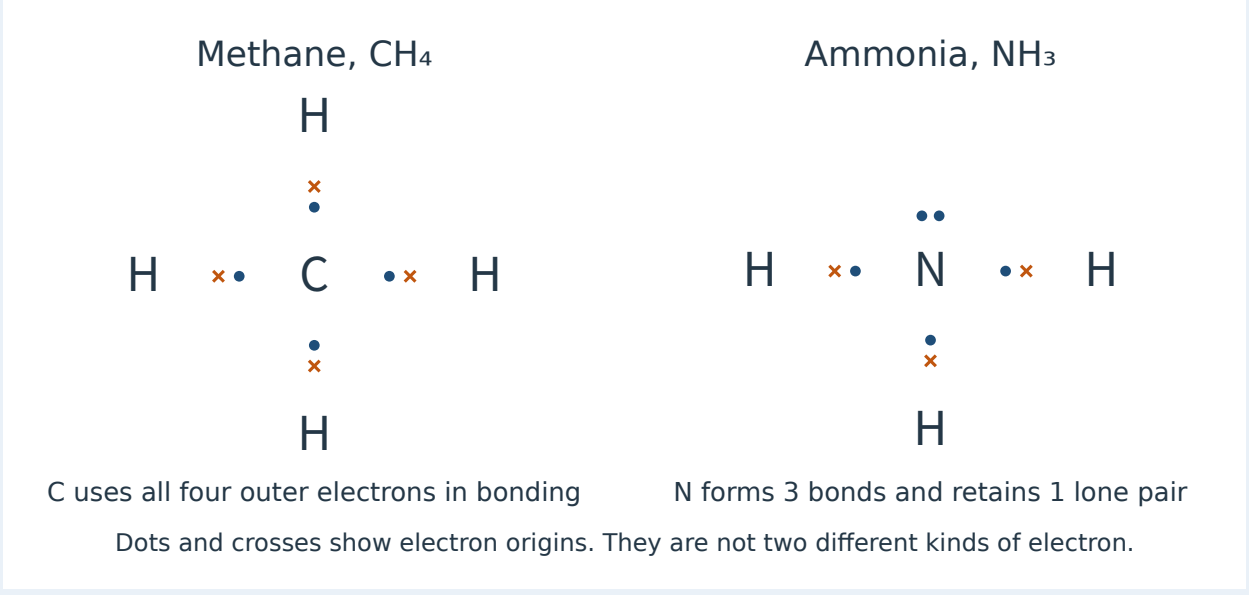
A hydrogen atom needs two outer-shell electrons. Chlorine starts with seven; after contributing to one shared pair, it retains three non-bonding pairs.

H_2 and HCl: begin with one shared region. In hydrogen, H_2, each H contributes one electron. Both atoms count the shared pair, filling their first shells. In hydrogen chloride, HCl, H contributes one electron and Cl contributes one of its seven. Chlorine's remaining six electrons form three lone pairs. H counts only two electrons. Cl counts six non-bonding electrons plus the two shared electrons, giving eight.



Check each atom after drawing. Cl starts with seven outer-shell electrons and O with six. Each bonded atom counts both electrons in its shared pair.

Cl_2 and H_2O: arrange the electrons outside the shared pairs too. In chlorine, Cl_2, each Cl starts with seven electrons and contributes one to a single bond. Each retains six electrons as three lone pairs. In water, H_2O, O starts with six electrons and contributes one to each of two shared pairs with H, forming two O—H single bonds. Oxygen's remaining four electrons form two lone pairs. Each H contributes only its own one electron. The whole water molecule contains 6+1+1=8 outer-shell electrons. O also counts eight around its own outer shell, including the H electrons in the shared regions.



Carbon uses each of its four electrons in a shared pair. Nitrogen uses three electrons in bonds, while its remaining two form a lone pair.

CH_4 and NH_3: check whether all the original electrons form bonds. In methane, CH_4, each of carbon's four original outer electrons pairs with one electron from an H. There are four single bonds and no lone pairs on C. In ammonia, NH_3, N starts with five outer electrons. Three form bonds with three H atoms; the remaining two form

one lone pair. N has three shared pairs plus one lone pair around it, totalling eight electrons. Each H still counts only two. Putting all five of nitrogen's electrons into bonding regions would create extra bonds or electrons.

For example, asking which atoms use all their original outer-shell electrons in bonding is different from asking which atoms finish with eight electrons. H uses its one original electron and C uses all four. O, N and Cl retain lone pairs. Compare where the original electrons go.

Use three checks when drawing. First count the atoms from the formula and arrange them using the supplied or previously taught connections. Next put one electron pair between the atoms of each bond and add lone pairs. Finally check that each H counts two electrons and each of the other atoms here counts eight. Also check that the total number drawn equals the sum of all atoms' original outer-shell electrons. A shared electron can be counted once for each atom when checking the two atoms separately, but only once in the total for the whole diagram.

Multiple bonds: two atoms can share more than one electron pair

Oxygen, O_2: O=O



Two shared pairs; two lone pairs per O

Nitrogen, N_2: N≡N



Three shared pairs; one lone pair per N

Each atom counts eight outer-shell electrons. Draw 12e in total for O_2, or 10e for N_2.

Dots and crosses show electron origins. They are not two different kinds of electron.

A double bond contains two shared pairs and a triple bond three. Include all non-bonding electrons when drawing multiple bonds.

Count shared electron pairs in double and triple bonds. In O_2, each O starts with six electrons. Sharing only one pair would not give both atoms eight. Each O instead contributes two electrons to make two shared pairs: the double bond O=O. Each O retains four electrons as two lone pairs. The complete diagram contains four bonding electrons plus eight non-bonding electrons, giving 12, exactly 6+6.

In N_2, each N starts with five electrons and contributes three to form three shared pairs: the triple bond N ≡ N. Each N retains two electrons as one lone pair. The diagram contains six bonding electrons plus four non-bonding electrons, giving ten, exactly 5+5. A triple bond requires six shared electrons, not just three.

In CO_2, carbon shares two pairs with each oxygen atom



O=C=O: no lone pairs on C; two lone pairs on each O

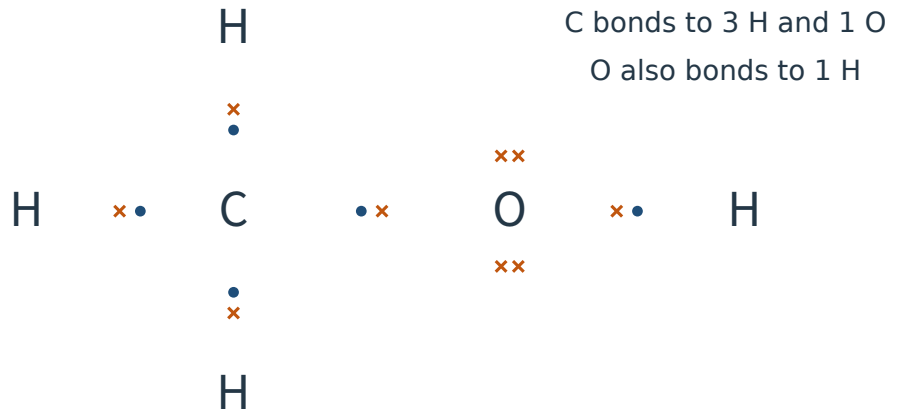
Total electrons: 4 + 6 + 6 = 16; four bonding pairs and four lone pairs.

Dots and crosses show electron origins. They are not two different kinds of electron.

The subscript 2 means two oxygen atoms; it is not a symbol for a double bond. Each C—O connection is a double bond because it contains two shared electron pairs.

Check the two double bonds in CO_2 separately. Carbon dioxide, CO_2, has the arrangement O=C=O. The central C starts with four electrons and contributes two to each double bond, leaving no lone pairs. Each O contributes two electrons to its double bond with C and retains four as two lone pairs. The diagram has 16 outer-shell electrons in total. C and each O count eight around themselves. The formula's subscript 2 counts oxygen atoms, whereas double bond means two shared pairs in each C—O connection. These two uses of 2 refer to different things.

Methanol, CH_3OH: establish the connections before placing outer-shell electrons



All five bonds are single bonds. O also retains two lone pairs.

Dots and crosses show electron origins. They are not two different kinds of electron.

In CH_3OH, CH_3 means one carbon bonded to three hydrogens, and OH means an oxygen bonded to another hydrogen. The connections are supplied here; prior knowledge of organic naming is not needed to read the diagram.

Methanol, CH_3OH, is shown with its connections supplied. The CH_3 part has one C bonded to three H atoms. That C is also bonded to an O, which is bonded to one further H. Carbon therefore forms four single bonds, while oxygen forms two single bonds and retains two lone pairs. Adding carbon's four electrons, oxygen's six and one from each of the four H atoms gives 14 outer-shell electrons for the whole diagram. The H bonded to O may use a different dot-or-cross symbol from the H atoms on the other side. The symbols simply distinguish electron origins within each neighbouring atom pair.

In C_2H_4, each carbon bonds to two hydrogens and shares two pairs with the other carbon



$H_2C{=}CH_2$: each C participates in four shared pairs and has no lone pairs.

Dots and crosses show electron origins. They are not two different kinds of electron.

A C=C double bond contains two electron pairs. Together with two C—H single bonds on each carbon, it allows each C to count eight outer-shell electrons.

For ethene, C_2H_4, first use the supplied structure $H_2C{=}CH_2$. The two C atoms share two pairs, and each C also shares one pair with each of two H atoms. Each carbon takes part in four shared pairs, counting eight electrons and retaining no lone pairs. The whole molecule has four C—H single bonds and one C=C double bond: six pairs or 12 bonding electrons. Counting pairs of neighbouring atoms is not enough to count electron pairs: the neighbouring C atoms in C=C share two pairs.

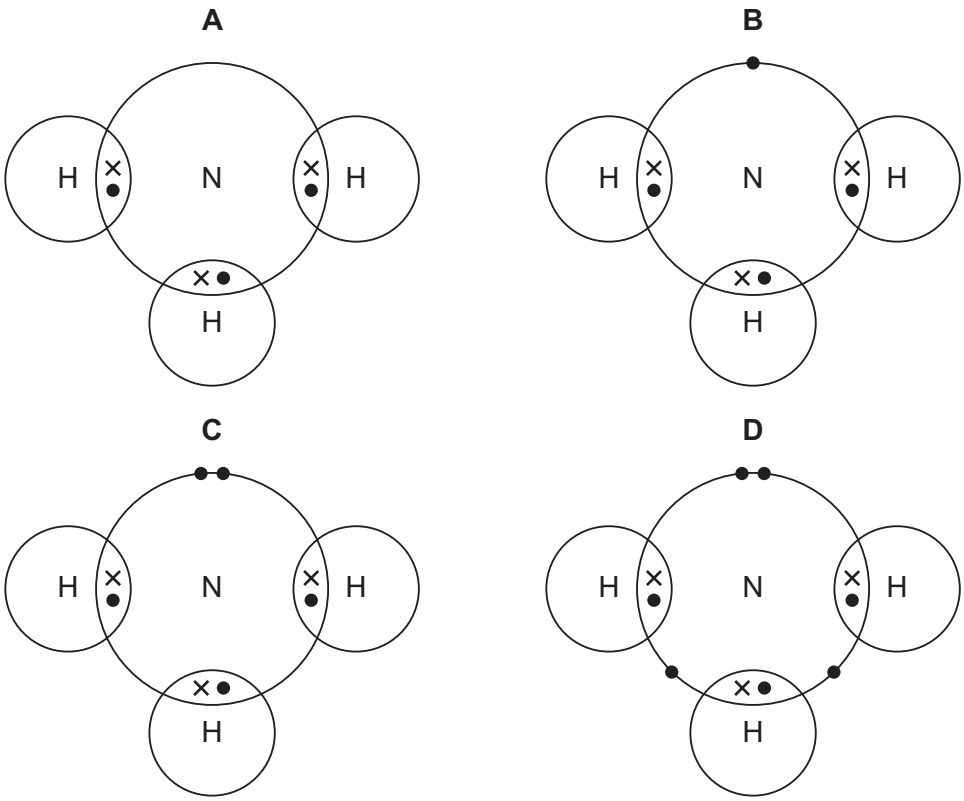
Apply the same rules to a new molecule with supplied connections. For example, if Cl—O—Cl and its two overlapping regions are already drawn, O forms one single bond with each Cl. O retains two lone pairs and each Cl three. Use the given connections first, then check the six electrons originally from O and seven from each Cl: 20 in total. The exercise asks you to add these electrons directly to the template.

Substances with shared electron pairs have covalent bonds, for example methane CH_4 and chlorine Cl_2. Argon, Ar, exists as single atoms with full outer shells and has no shared pairs between these atoms. Iron(III) oxide, Fe_2O_3, is an ionic compound of iron and oxygen, held together by electrostatic attraction between ions. Do not draw it as separate small molecules containing shared pairs. The Roman numeral III indicates a +3 charge on the iron ion. Formulae and more detailed ion combinations are covered in Section 3.1; this classification does not require deriving its entire structure from the name.

Recognising the Question Count the atoms and their original outer-shell electrons. Use the known or supplied connections to place shared pairs, then add lone pairs. Check that H counts two electrons and the other atoms here count eight. Also check the diagram's total electron count and dot-and-cross origins.

Worked Example 1 (2025·May/June 0620/13 Q6) · Level 3

Which dot-and-cross diagram represents the electronic configuration for the outer shell electrons in ammonia, NH_3?



Solution

First determine the electron total for NH_3. N starts with five outer-shell electrons and each of the three H atoms contributes one, giving eight in the whole diagram. Each of the three N—H connections contains one dot and one cross, forming three shared pairs. Of nitrogen's five original electrons, three form bonds and two remain as one lone pair.

Next check the atoms individually. Each H counts two electrons in its overlap and has a full first shell. N counts three shared pairs plus one lone pair, giving eight electrons. The total number of symbols representing electrons from N must be five; do not add electrons arbitrarily to fill the circle.

Finally compare the options. A omits nitrogen's lone pair. B leaves only one electron outside the bonds on N, so it is still one short. C has the complete lone pair above N and is correct. D includes that upper lone pair but also adds two lower non-bonding electrons, giving too many. Choose C. Checking the three H atoms is not enough: also check nitrogen's remaining electrons and the overall total.

Worked Example 2 (2025·May/June 0620/43 Q4(a)(ii)) · Level 3

(ii) Complete the dot-and-cross diagram in Fig. 4.1 to show the electronic configuration in an oxygen molecule, O_2. Show outer-shell electrons only.

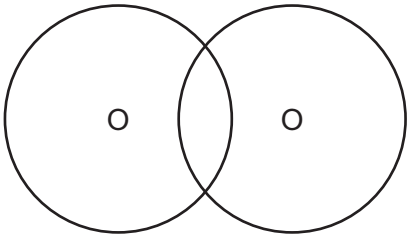


Fig. 4.1

Solution

Step 1: determine the shared electrons. Each O starts with six outer-shell electrons and contributes two, forming two shared pairs. Draw two dots and two crosses in the overlapping region of the circles.

Step 2: add each atom's remaining electrons. The left O has used two of its six electrons, leaving four dots. Arrange these as two pairs in the non-overlapping part of the left circle. The right O similarly retains four crosses as two pairs in the non-overlapping part of the right circle.

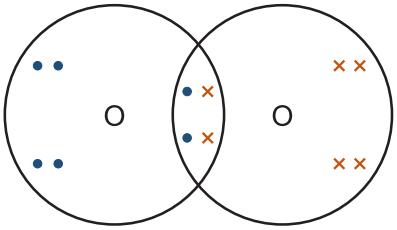


Fig. 4.1

The completed exam template has two shared pairs in the centre and two lone pairs on each oxygen.

Finally check the counts. There are six dots and six crosses, giving 12 electrons overall. Each O counts four non-bonding electrons+four shared electrons=8. Drawing only the two lines in O=O does not meet the request for a complete dot-and-cross diagram showing outer-shell electrons. Include the lone pairs too.

Worked Example 3 (2025·May/June 0620/41 Q4(a)) · Level 3

This question is about nitrogen and its compounds.

Nitrogen contains molecules with the formula N_2.

(a) Complete the dot-and-cross diagram in Fig. 4.1 to show the electronic configuration in a nitrogen molecule. Show outer-shell electrons only.

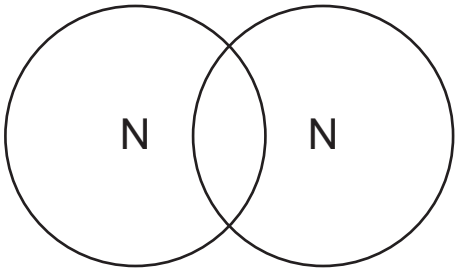


Fig. 4.1

Solution

Start with five original outer-shell electrons on each nitrogen. Each N contributes three to form three shared pairs. Draw three dots and three crosses in the overlap to represent the triple bond. The three means three pairs, not three electrons in total.

Then add the remaining electrons. Each N retains two electrons as one lone pair. Place two dots in the non-overlapping part of the left circle and two crosses in the non-overlapping part of the right circle:

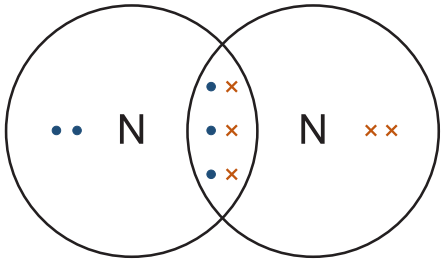


Fig. 4.1

Complete the two circles with three shared pairs and one lone pair at each end.

Check that the whole diagram contains five dots+five crosses=10 outer-shell electrons. Each N counts six shared electrons+two lone-pair electrons=8. All dots and crosses can be interchanged, provided the electron origins remain consistent.

Practice

1. (2025·May/June 0620/42 Q2(a)) · Level 1

This question is about covalent compounds.

(a) State what is meant by a covalent bond.

2. (2025·May/June 0620/11 Q7) · Level 2

Which substances contain one or more shared pairs of electrons?

1 argon

2 methane

3 iron(III) oxide

4 chlorine

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

3. (2025·May/June 0620/12 Q8) · Level 3

In the molecules CH_4, HCl and H_2O, which atoms use **all** of their outer shell electrons in bonding?

A C and Cl **B** C and H **C** Cl and H **D** H and O

4. (2025·May/June 0620/42 Q2(b)(ii)) · Level 3

This question is about covalent compounds.

(b) Chlorine(I) oxide, Cl_2O, is a simple molecule with covalent bonds.

(ii) Complete the dot-and-cross diagram in Fig. 2.1 of a molecule of chlorine(I) oxide.

Show outer electrons only.

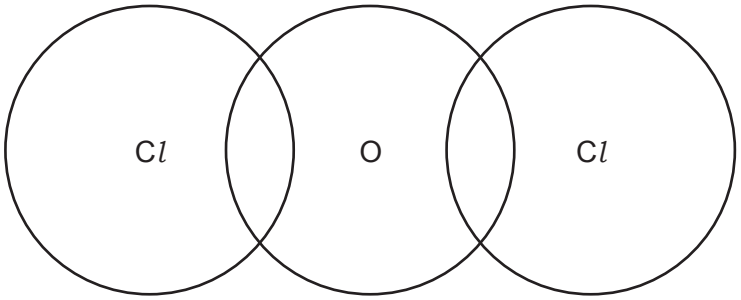
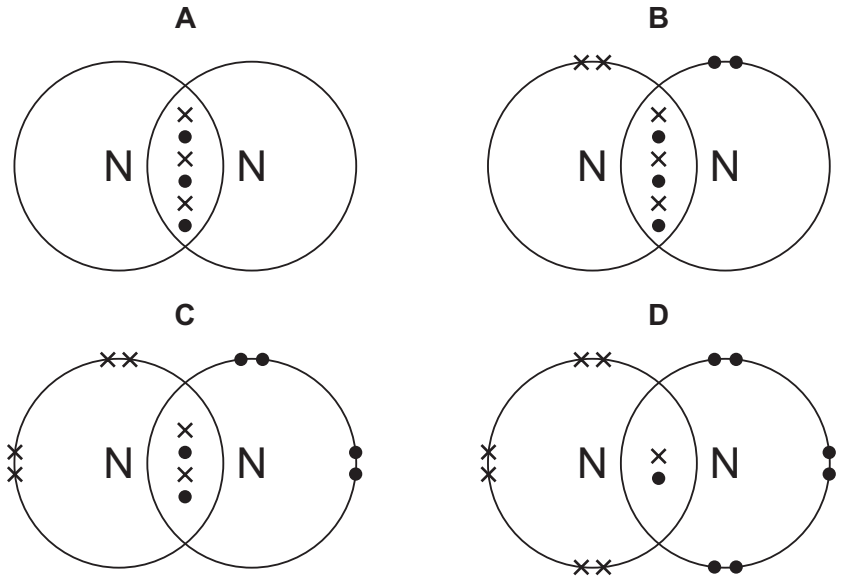


Fig. 2.1

5. (2025·May/June 0620/21 Q6) · Level 3

Which dot-and-cross diagram shows the outer shell electron arrangement in a molecule of nitrogen?

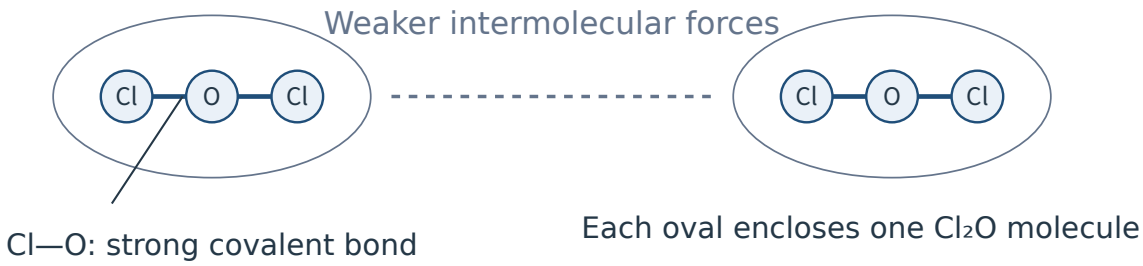


Type 2: Explaining properties by distinguishing covalent bonds from intermolecular forces

Key Ideas An electron diagram of a **covalent bond** shows two atoms sharing a pair of electrons. Now consider a substance made of many separate molecules. A **simple molecular substance** contains discrete molecules of limited size. Strong covalent bonds join atoms within each molecule, while weaker attractions between different molecules are called **intermolecular forces**.

Intermolecular means between molecules. To decide how much energy a change needs, first identify what is being separated: are whole molecules moving apart, or are the connections between atoms within a molecule changing? This course does not require further classification of intermolecular forces, but these two levels must be clearly distinguished.

Distinguish bonds within a molecule from forces between molecules



Cl—O: strong covalent bond

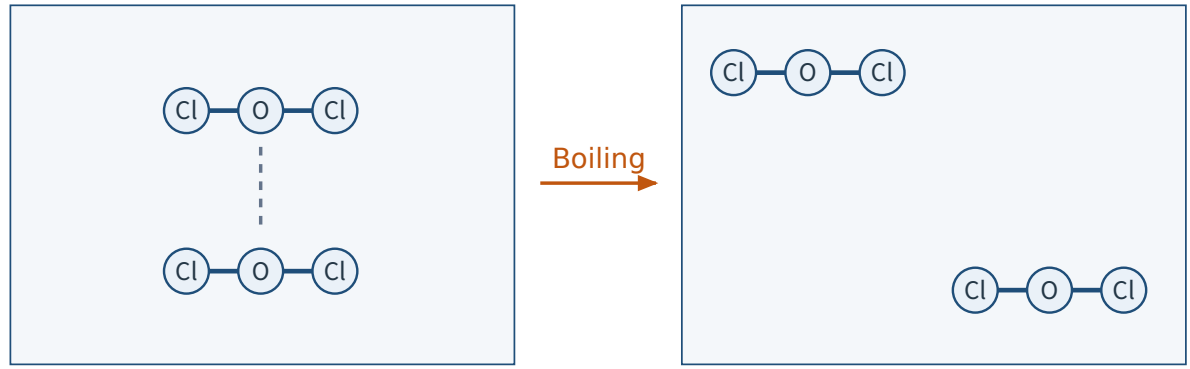
Each oval encloses one Cl_2O molecule

Ovals identify molecules. The connections do not specify actual bond angles.

Solid lines represent Cl—O covalent bonds within molecules; dashed lines represent weaker forces between different molecules. These two levels must not be interchanged.

Melting and boiling usually rearrange or separate whole molecules. In liquid chlorine(I) oxide, Cl_2O, each molecule remains a unit of bonded chlorine and oxygen atoms. Boiling it into a gas mainly overcomes the weaker attractions between different Cl_2O molecules, allowing intact molecules to become further apart. There is no need to break every Cl—O covalent bond.

Boiling separates molecules but preserves their internal bonds



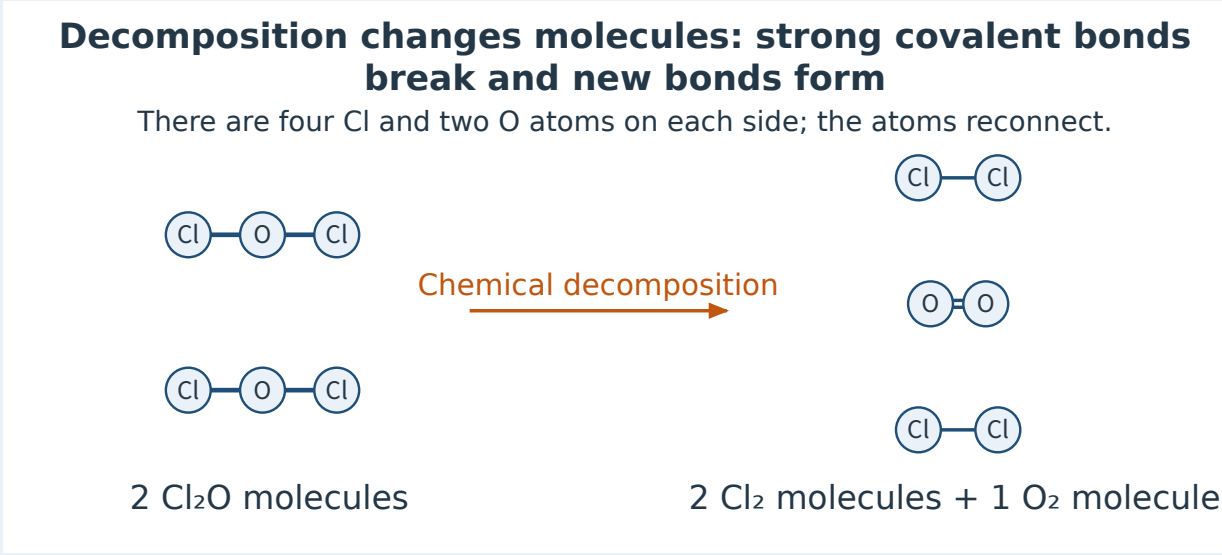
Liquid: molecules closer together

Gas: the same molecules farther apart

The arrow shows boiling from liquid to gas. The same two intact Cl_2O molecules remain before and after. Their separation changes, not the atomic connections within each molecule.

Because the intermolecular forces to be overcome are weak, relatively little energy is needed. Simple molecular substances therefore usually have low melting and boiling points. A low melting point does not mean weak covalent bonds, nor does it mean that molecules have no attractions between them.

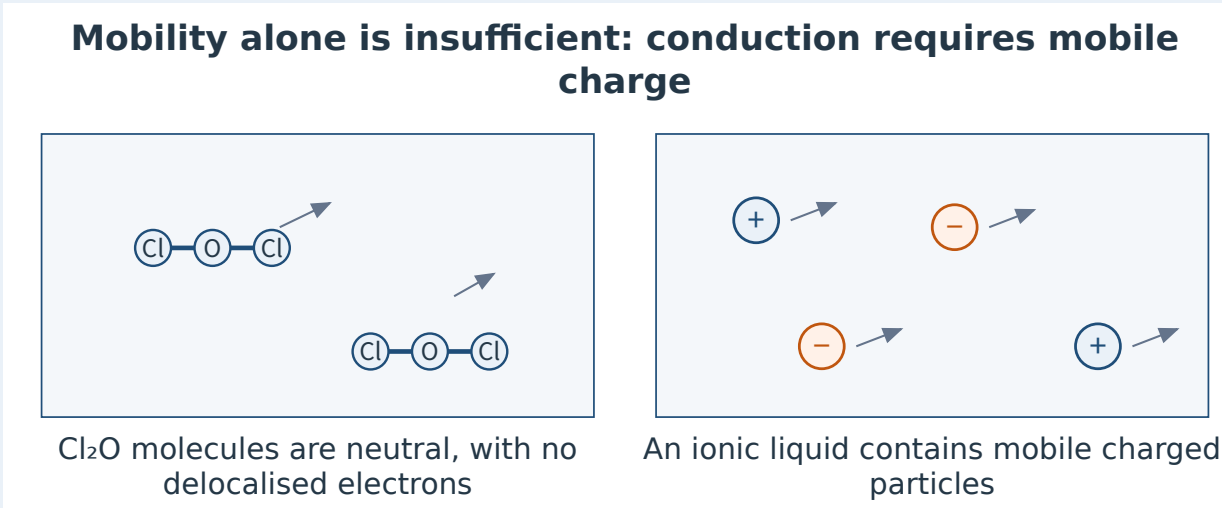
Thermal decomposition is a chemical change caused by heating in which the original substance becomes other substances. For example, decomposing Cl_2O into Cl_2 and O_2 changes connections within the original molecules. Strong Cl—O covalent bonds must break and bonds in the new substances form. Simply increasing the distances between intact Cl_2O molecules does not produce chlorine and oxygen.



The arrow shows changes in connections during chemical decomposition. Atom numbers are conserved, but the original Cl—O bonds must break. This is not an experimental procedure diagram and does not show actual bond angles.

For example, compare two containers. In A, the original molecules are further apart but their internal connections are unchanged. In B, atoms that were originally bonded have rearranged into new molecules. A may show boiling, mainly overcoming weaker intermolecular forces. B shows a chemical change involving the breaking of original covalent bonds and the formation of new bonds. A molecular substance boiling at a low temperature does not mean that the same temperature is sufficient to break strong covalent bonds and decompose it into elements.

Why are simple molecular substances usually poor electrical conductors? First, the molecules are neutral, so there are no ions available to move and carry charge. Second, there are no delocalised electrons free to move throughout the substance. Electrons are still present in atoms, covalent bonds and non-bonding regions of molecules; they simply do not provide mobile charge carriers throughout the material.



Particles can move on both sides, but the left-hand particles are neutral molecules and the right-hand particles are charged ions. The left-hand substance contains electrons, but none are free to move throughout it.

Molecules in a liquid can move, but they move as neutral units. Particle movement alone therefore does not explain electrical conduction. If two reasons are required, give no ions and no mobile electrons separately. Writing no electrons is incorrect. This comparison applies to substances that remain simple molecular in the stated condition. Some substances form ions when they interact with water; judge those aqueous solutions by the ions actually present.

When applying structural explanations to a data table, check both relevant conditions. A simple molecular substance usually has a low melting point and remains a poor conductor when molten. Failure to conduct alone

is insufficient: other structures can also be non-conducting but have high melting points. Choosing a low melting point while ignoring conduction when molten is also insufficient. Density describes the relationship between mass and volume, but is not a reliable sole criterion for identifying the bonding structure here.

Check both properties together

Substance	Melting point / °C	Conductivity when molten	Meets both simple-molecular conditions?
X	1450	Does not conduct	No: very high melting point
Y	−90	Does not conduct	Yes: low melting point and no conduction
Z	750	Conducts	No: conducts when molten

Do not use non-conduction alone, or substitute density for either required property.

Check both low melting point and no conduction when molten in every row. Only Y satisfies both.

Read the complete table. X does not conduct, but its melting point of 1450°C is high, so it fails the combined conditions. Y has a very low melting point of −90°C and does not conduct when molten, satisfying both. Z conducts when molten, so reject it. Only Y fits. If an exercise gives five rows and asks for two substances, check these two columns for each row, then give both letters and the two corresponding reasons. A negative melting point is below 0°C; do not remove the minus sign before comparing temperatures.

Recognising the Question Decide whether intermolecular separation or connections within molecules change. Explain melting and boiling points using weak intermolecular forces. For poor conductivity, check ions and mobile electrons separately. In data questions, require both a low melting point and no conduction when molten.

Worked Example 1 (2025·May/June 0620/42 Q2(b)(iii)) · Level 3

This question is about covalent compounds.

(b) Chlorine(I) oxide, Cl_2O, is a simple molecule with covalent bonds.

(iii) Explain, in terms of structure and bonding, why Cl_2O boils at a low temperature and does **not** thermally decompose into its constituent elements, Cl_2 and O_2.

Solution

First distinguish the two changes being compared. During boiling, liquid Cl_2O becomes gaseous and the original Cl_2O molecules move apart. During thermal decomposition, the original molecules change and form new substances consisting of Cl_2 and O_2.

Boiling overcomes attractions between different Cl_2O molecules. These intermolecular forces are weak and require relatively little energy to overcome, so Cl_2O has a low boiling point. The covalent bonds joining Cl to O within each molecule remain intact.

Decomposition into Cl_2 and O_2 requires the original Cl—O covalent bonds inside Cl_2O molecules to break and new bonds to form. The original covalent bonds are strong and require more energy to break than the intermolecular forces require to overcome. The energy sufficient for low-temperature boiling is therefore insufficient to decompose the molecules at the same time.

A complete answer can be written: During boiling, weak forces of attraction between Cl_2O molecules are overcome. Thermal decomposition would require the strong covalent bonds between Cl and O atoms to be broken. Therefore Cl_2O can boil at a low temperature while its molecules remain intact.

Weak describes the forces between molecules, not the Cl—O covalent bonds. Do not explain easy boiling by saying that the covalent bonds are weak.

Practice

1. (2025·May/June 0620/42 Q2(b)(iv)) · Level 2

This question is about covalent compounds.

(b) Chlorine(I) oxide, Cl_2O, is a simple molecule with covalent bonds.

(iv) Give **two** reasons why liquid Cl_2O is a poor conductor of electricity.

2. (2025·May/June 0620/31 Q6(d)) · Level 3

This question is about ionic and covalent compounds.

(d) Table 6.2 shows some properties of five compounds, **A**, **B**, **C**, **D** and **E**.

Table 6.2

compound	electrical conductivity when molten	density in g/cm^3	melting point in °C
A	does not conduct	3.53	1856
B	does not conduct	1.57	–157
C	conducts	1.93	110
D	conducts	3.03	1256
E	does not conduct	4.93	–83

State which **two** of the compounds, **A**, **B**, **C**, **D** and **E**, are simple molecules.
Give **two** reasons for your answer.

compounds and

reason 1

reason 2

3. (2025·May/June 0620/33 Q5(d)) · Level 3

This question is about ionic compounds and simple molecular compounds.

(d) Table 5.1 shows some properties of five compounds, **A**, **B**, **C**, **D** and **E**.

Table 5.1

compound	electrical conductivity when molten	density in g/cm^3	melting point in °C
A	conducts	1.85	115
B	does not conduct	2.65	1713
C	does not conduct	0.42	–182
D	conducts	2.99	1190
E	does not conduct	1.56	–101

State which **two** of the compounds, **A**, **B**, **C**, **D** and **E**, are simple molecules.

Give **two** reasons for your answer.

compounds and

reason 1

reason 2

2.6 Giant covalent structures

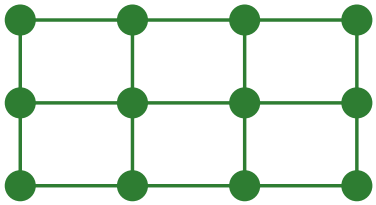
Type 1: Explaining graphite, diamond and silicon dioxide using giant covalent structures

Key Ideas Covalent bonds can join a few atoms into separate molecules, or continue joining atoms to form a network throughout a solid. The latter is a **giant covalent structure**. Giant describes the continuous extent of covalent bonding, not a separate molecule drawn unusually large.

Giant covalent structures extend continuously; they are not collections of small molecules



Separate H_2O molecules



Network: bonds join neighbouring atoms

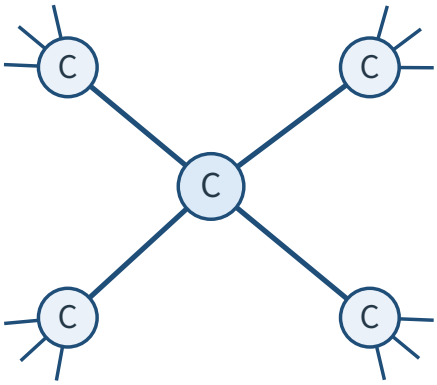
The right-hand diagram shows continuous bonding. Specific arrangements follow for diamond, graphite and SiO_2.

A giant structure has covalent bonds continuing throughout a network; it is not a small molecule drawn larger. The right-hand diagram illustrates connectivity without specifying one substance's geometry.

Water H_2O, ammonia NH_3 and carbon dioxide CO_2 consist of separate small molecules. Diamond, graphite and silicon dioxide are the continuous covalent networks studied here. Sodium chloride NaCl and magnesium chloride $MgCl_2$ also form giant lattices, but their bonding is electrostatic attraction between ions. The word giant alone does not place them among giant **covalent** structures.

Carbon atoms can be connected in different ways. Diamond and graphite both contain only carbon atoms, so they are different structural forms of the same element. Their different atomic arrangements and bonding produce different properties; another element does not have to be mixed in. Write their composition as C. The few carbon atoms drawn in a local structural diagram do not provide a subscript for a fixed small molecule.

Diamond: each carbon forms strong covalent bonds to four other carbons



Continuous three-dimensional network

To substantially change atom positions, many strong covalent bonds must break



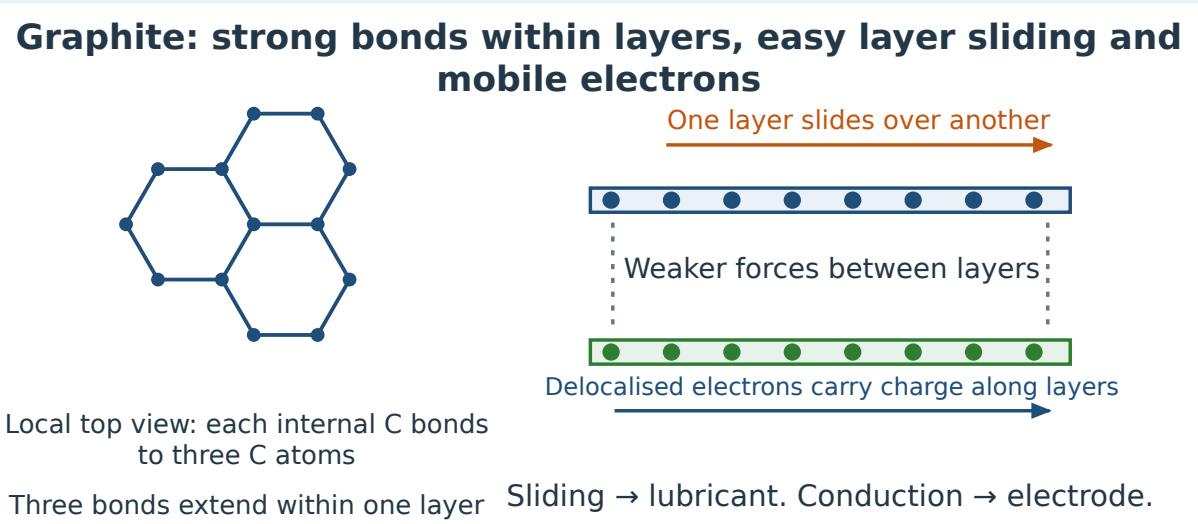
Hard → useful for cutting tools

Local two-dimensional projection. Boundary bonds continue beyond the diagram; this is not a C_5 molecule.

The central carbon in this local view has four bonds. The three-dimensional network cannot slide easily like graphite layers. All outer-shell electrons take part in bonding, leaving no freely mobile electrons.

In diamond, each C bonds to four other C atoms in a strong three-dimensional network. Each carbon uses its four outer-shell electrons in four covalent bonds, and the network continues in all directions. Moving carbon atoms substantially relative to one another requires many strong covalent bonds to break. Diamond is therefore very hard and suitable for cutting or drilling tools. Melting also requires a great deal of energy, giving a very high melting point. Diamond has no charged particles free to move throughout its structure and usually does not

conduct electricity. Hard means difficult to scratch or cut into; it does not mean impossible to fracture under any circumstances.



Upper orange arrows show layers sliding relative to one another; lower blue arrows show electron movement. These explain lubrication and electrical conduction respectively.

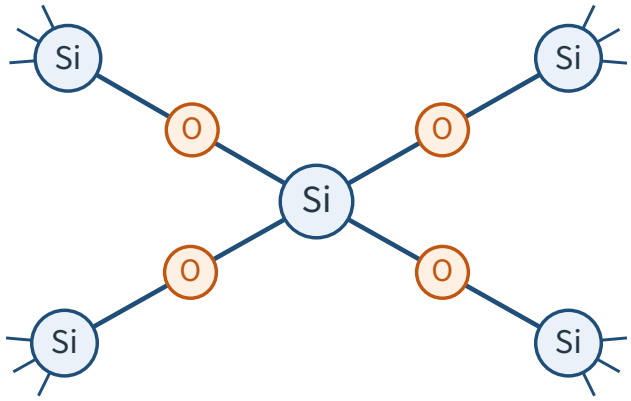
In graphite, each C bonds to three other C atoms within one layer. Strong covalent bonds form hexagonal networks, with many layers stacked together. Bonds within a layer are strong, while forces between layers are weaker. Layers can slide past one another, so graphite can act as a lubricant, helping contacting surfaces slide. Sliding mainly overcomes the weaker interlayer forces; it does not require all the strong C—C bonds in a layer to be broken.

Each carbon has four outer-shell electrons. After forming three local covalent bonds in graphite, an electron remains that is not confined to one bond or one atom. Such **delocalised electrons** can move along the layers and carry charge, so solid graphite conducts electricity. It can therefore be used as an electrode, carrying current into or out of an apparatus. Electrons carry the current; carbon atoms do not become freely moving ions.

For example, if asked why graphite helps machine parts slide, explain that weak forces between layers allow the layers to slide. Do not say that all graphite's covalent bonds are weak. If instead asked why it is suitable for an electrode, explain conduction by mobile delocalised electrons. Sliding layers do not explain an electric current.

Oxide means an oxygen-containing compound, but does not automatically mean an ionic structure. Silicon dioxide is also called silicon(IV) oxide, with formula SiO_2. It contains silicon, Si, and oxygen, O, so it is a compound, unlike the single-element structures diamond and graphite. It is held together by strong covalent bonds, not by an alternating lattice of silicon and oxide ions.

SiO_2 network: each Si bonds to four O; each O bonds to two Si



One O bridges two Si atoms
Four O neighbours per Si
Each O is shared by two Si

Overall O : Si = 4 ÷ 2 : 1

Hence the formula SiO_2

The network continues beyond the boundary. Do not use the five visible Si and four O atoms as the bulk formula.

Silicon dioxide is also called silicon(IV) oxide. Each short line is one Si—O single bond containing one shared electron pair. Silicon and oxygen are continuously connected in a three-dimensional network.

In silicon dioxide, each Si bonds to four O atoms and each O bonds to two Si atoms. Every Si—O connection is a single bond containing one shared pair. Oxygen links silicon atoms, so the continuous network is not a collection of separate SiO_2 molecules. Although four O atoms surround each Si, each O is shared between two Si atoms. The overall average is therefore 4÷2=2 O atoms per Si, giving SiO_2.

Like diamond, silicon dioxide has a three-dimensional network of strong covalent bonds. Disrupting this structure needs a great deal of energy, so it is hard and has a high melting point. It has no electrons free to move through the network and usually does not conduct. It lacks graphite's easily sliding layered structure. In property questions, connect local bonding → overall structure → property. The presence of silicon in the name does not establish electrical conduction.

To label an empty network with Si and O, count the bonds first. A node with four bonds is Si. A node between two Si atoms with only two bonds is O. Lines extending beyond the diagram show the network continuing outside the drawing and still count towards the bonds of edge atoms. Begin with a central node whose connections are clear, work outwards along each bond and check every node: each Si is adjacent to O atoms, and each O links two Si atoms.

For example, the centre in the diagram has four lines, so it is Si. Each intermediate node on a path to an outer node has two lines, so it is O. Each outer node has one bond to O and three more extending outwards, so these are also Si. The five Si and four O atoms visible are only a section cut from a larger network, with many atoms outside the drawing. They do not make the solid's formula Si_5O_4. The following exam question rotates this section of network; continue to label every node by its number of connections.

Recognising the Question Distinguish separate molecules, ionic lattices and continuous covalent networks, then connect local bonding to overall properties. Graphite's sliding layers and delocalised electrons explain different uses. In an unlabelled SiO_2 network, check each node: four connections for Si, two for O.

Worked Example 1 (2025·May/June 0620/13 Q7) · Level 3

Which statements about graphite are correct?

1 Each carbon atom is covalently bonded to four other carbon atoms.

2 Graphite is an electrical conductor because it contains ions that are free to move.

3 Graphite is used as a lubricant because it has layers that slide over each other.

A 1 and 2 **B** 1 only **C** 2 and 3 **D** 3 only

Solution

Connect each statement to the structure. Statement 1 says that each C is covalently bonded to four other C atoms. That is diamond's connectivity. In graphite, each C bonds to three C atoms within a layer, so 1 is incorrect.

Statement 2 attributes conduction to freely moving ions. Graphite's charge carriers are delocalised electrons. Carbon atoms remain within the covalent layers and are not freely moving ions, so 2 is incorrect.

Statement 3 links layers sliding past one another to use as a lubricant. This matches strong bonding within layers and weaker forces between them, so 3 is correct. Only 3 is correct: choose D. The statements test different relationships. Do not accept an incorrect explanation simply because its initial claim that graphite conducts is true.

Worked Example 2 (2025·May/June 0620/42 Q2(c)(iii)) · Level 3

This question is about covalent compounds.

(c) Carbon and silicon(IV) oxide both exist as giant covalent structures.

(iii) Silicon(IV) oxide contains silicon atoms, Si, and oxygen atoms, O.

Fig. 2.2 shows part of the giant covalent structure of silicon(IV) oxide.

Complete the diagram in Fig. 2.2 by adding the symbol for each of the 9 atoms shown.

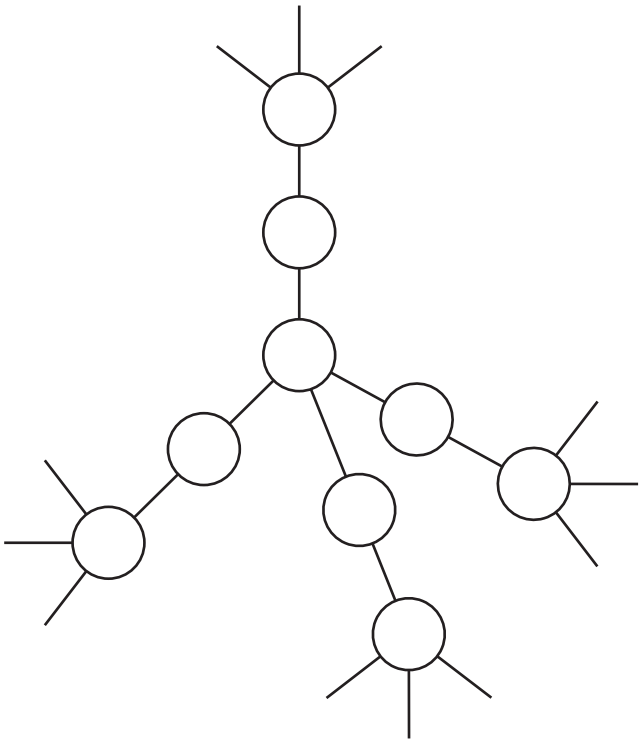


Fig. 2.2

Solution

Step 1: begin at the central node, where the connections are clear. Four lines extend from it, so it bonds to four neighbouring atoms and must be Si.

Step 2: follow the four paths outwards. Each of the four intermediate nodes next to the centre has two bonds, one to the centre and one to an outer node, so each is O. Each outer node has one bond towards O and three short lines continuing into the network beyond the diagram, giving four connections in total, so each is Si.

Step 3: complete and check every position in the original diagram. The upper vertical sequence, from top to bottom, is Si, O, Si at the centre. From the centre, the path down-left is O, Si; the path to the right is O, Si; and the path downwards and slightly right is also O, Si. The completed labels are shown below:

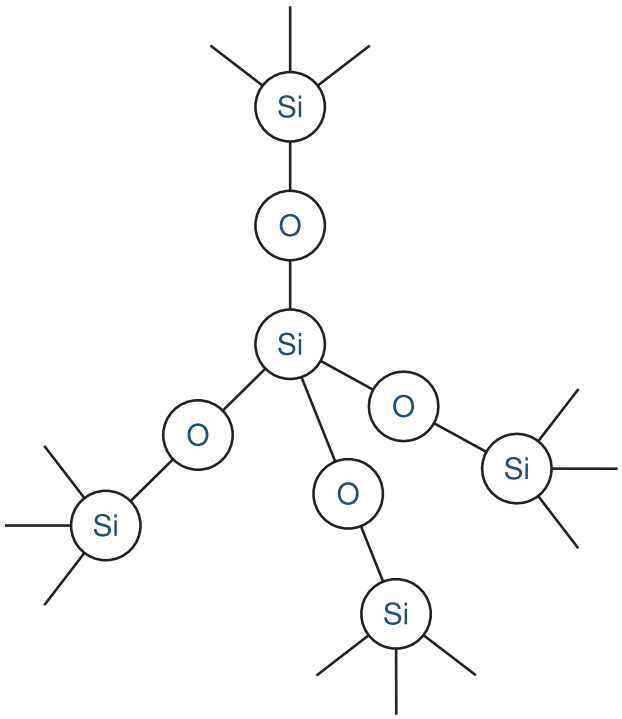


Fig. 2.2

All nine original nodes and the lines continuing beyond the edges are retained. Label the centre and four outer nodes Si, and the four bridging nodes O.

There are five Si labels and four O labels. Each O joins two Si atoms, and each Si has four bonds in total. The diagram is only part of a network: the visible 5:4 ratio does not change the compound's formula to Si_5O_4. The complete network remains SiO_2.

Practice

1. (2025·May/June 0620/11 Q8) · Level 1

Which substance has a giant covalent structure at room temperature and pressure?

A ammonia

B carbon dioxide

C diamond

D water

2. (2025·May/June 0620/12 Q9) · Level 1

Which substance has a giant covalent structure?

A sodium chloride

B ammonia

C magnesium chloride

D graphite

3. (2025·May/June 0620/41 Q1(b)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(b) contains carbon atoms only

4. (2025·May/June 0620/42 Q2(c)(i)) · Level 1

This question is about covalent compounds.

(c) Carbon and silicon(IV) oxide both exist as giant covalent structures.

(i) Name a giant covalent structure of carbon which conducts electricity.

5. (2025·May/June 0620/42 Q2(c)(ii)) · Level 1

This question is about covalent compounds.

(c) Carbon and silicon(IV) oxide both exist as giant covalent structures.

(ii) Identify the particles responsible for the conduction of electricity in this covalent structure of carbon.

6. (2025·May/June 0620/43 Q1(a)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(a) is a compound with a giant covalent structure

7. (2025·May/June 0620/43 Q1(d)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(d) is a good conductor of electricity when solid

8. (2025·May/June 0620/21 Q7) · Level 2

Which statement about silicon(IV) oxide is correct?

A It contains silicon ions and oxide ions arranged in a giant lattice.

B It is used in electrical circuits because it can conduct electricity.

C Its bonds each contain one shared pair of electrons.

D Its atoms are arranged in the same way as the atoms in graphite.

9. (2025·May/June 0620/22 Q6) · Level 2

Which statement about silicon(IV) oxide, SiO_2, is correct?

A It conducts electricity because it contains delocalised electrons.

B It has a giant covalent structure with each silicon atom bonded to four oxygen atoms.

C It is a simple covalent molecule.

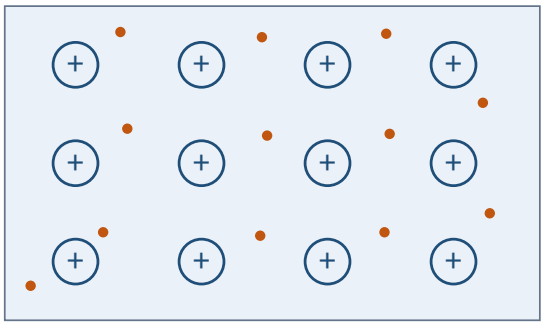
D Its structure is similar to that of graphite.

2.7 Metallic bonding

Type 1: Explaining metallic bonding using layers of metal ions and delocalised electrons

Key Ideas Metals provide another type of continuous structure. In metals such as copper Cu, zinc Zn and magnesium Mg, some outer-shell electrons are no longer confined to individual atoms and are distributed throughout the structure. These are **delocalised electrons**. Delocalised means not restricted to one atom or one particular chemical bond. The remaining positive metal ions form a regular arrangement called a **giant metallic lattice**.

Metallic bonding is electrostatic attraction between positive ions and delocalised electrons



Positive metal ions: regular arrangement
Delocalised electrons: spread throughout
Attraction between opposite charges holds the metal together

The electron sea describes delocalisation; it does not mean that a metal contains a liquid.

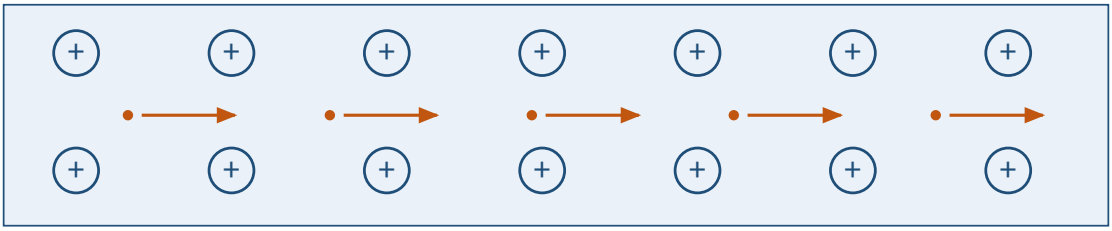
Circles marked + represent positive ions; orange dots represent delocalised electrons. Only part of the lattice is shown. Each electron is not separately bound to one atom.

Metallic bonding is the electrostatic attraction between positive ions in a metallic lattice and a sea of delocalised electrons. Electron sea describes electrons distributed throughout the structure; it is a metaphor, not an actual liquid inside the metal. The definition must identify both types of particle and their mutual attraction. Strong attraction between particles alone does not specify the particles involved.

Compare this with ionic bonding. In an ionic lattice, the negatively charged particles are anions; in a metallic lattice, they are delocalised electrons. Both involve electrostatic attraction, but that does not make them the same type of bond. Identify both objects of the attraction.

Solid metals already contain mobile charge carriers. Delocalised electrons can move through the entire metallic structure. When the metal is connected in a circuit, these electrons carry charge, so metals are usually good electrical conductors. Positive ions are charged too, but during conduction in a solid they continue vibrating near their fixed positions; they do not move freely through the material like ions in a liquid.

In a conducting solid metal, delocalised electrons move through the structure



In a circuit, electrons carry charge. Positive ions do not travel freely through the lattice.

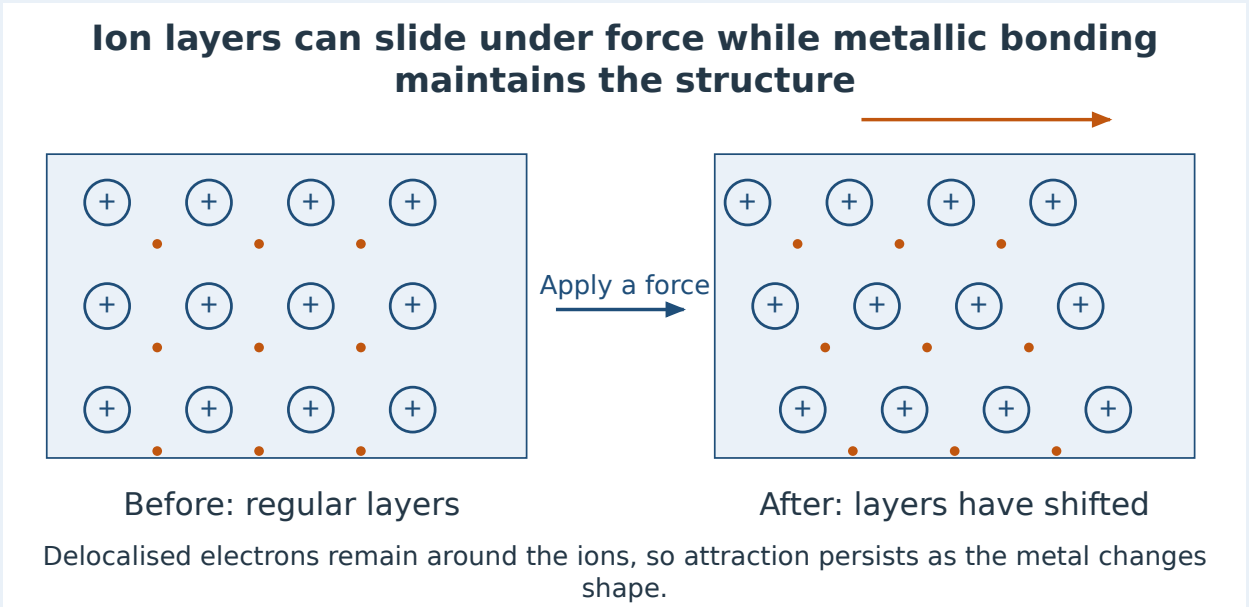
Orange arrows show electron drift; the opposite direction of conventional current is not also drawn.

Electrons can move through the metallic structure while positive ions remain near their original positions. The arrows specifically show electron movement.

For example, solid copper and solid sodium chloride both contain charged particles. Why do they conduct differently? Copper has delocalised electrons that can move through the solid, so it conducts. Solid sodium

chloride has ions fixed in a lattice and unable to carry charge freely through the material, so it conducts poorly. The deciding factor is the presence of mobile charge in the stated condition, not whether a metal appears in the substance's name.

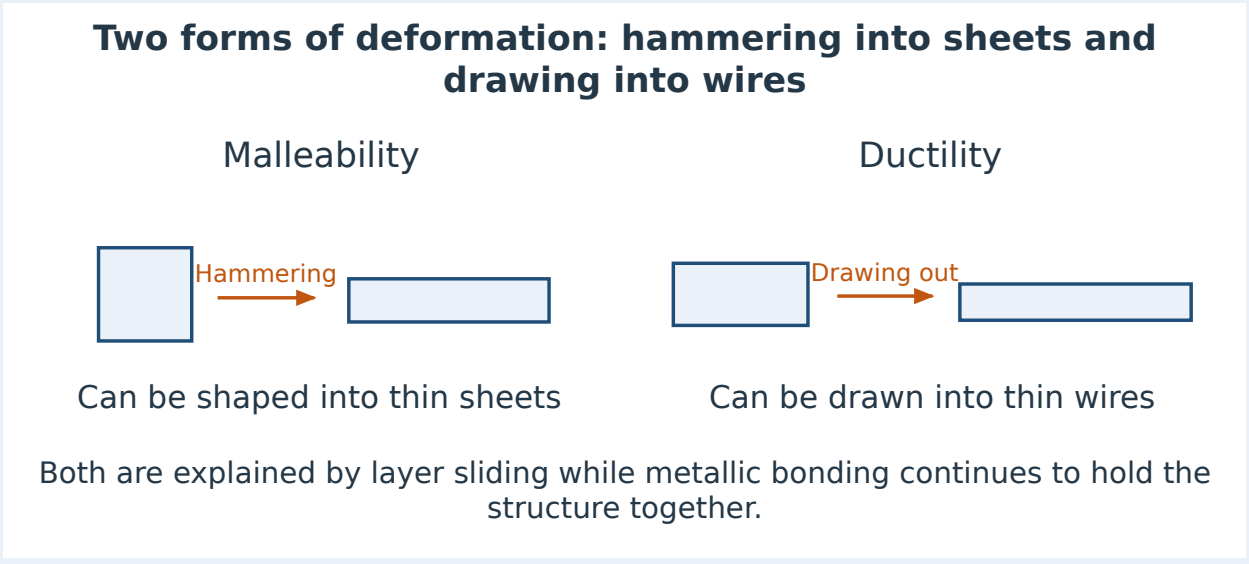
Metals change shape because layers can slide under force. Metal ions form regular layers. Under a suitable mechanical force, one layer can slide relative to another. Delocalised electrons remain distributed around the positive ions, so attraction between the ions and electrons continues. The metal can therefore deform without immediately breaking apart.



These arrows show layers sliding under a mechanical force, unlike the electron arrows in the conduction diagram. Attraction between positive ions and electrons remains after sliding.

This movement of ion layers under force differs from electrons moving through the solid in the previous conduction diagram. Do not explain conduction by flowing ion layers. For hammering or stretching, connect the applied force to layer sliding and the continued metallic attraction afterwards.

Malleability is the ability to be hammered into thin sheets. **Ductility** is the ability to be drawn into wires. They describe different visible results but can both be explained by the same structural mechanism. Ductile does not mean brittle. Simply saying that metals are strong also fails to explain how their structure can adjust its positions.



Orange arrows show hammering and stretching, making the two property terms concrete. The drawings illustrate deformation and do not provide geometric measurements.

For example, if asked why copper can be made into wire, identify ductility and explain that layers of ions can slide while delocalised electrons continue holding the structure together. If asked why the copper wire conducts afterwards, explain that delocalised electrons can move and carry charge. Being drawn into wire and conducting electricity are different properties of the same object and need the corresponding particle movements in their explanations.

A new metal name does not require memorising the copper example as a rule that applies only to copper. For an ordinary solid metal, identify positive metal ions, delocalised electrons and their interactions. Replacing copper with zinc or magnesium does not turn the charge carriers into anions or freely flowing metal ions.

Recognising the Question Define metallic bonding using positive ions, delocalised electrons and electrostatic attraction. Explain conduction through electron movement. Explain hammering into sheets or drawing into wires through layer sliding and the attraction that remains afterwards.

Worked Example 1 (2025·May/June 0620/41 Q3(a)(i)) · Level 2

This question is about copper and compounds of copper.

(a) (i) Describe the bonding in a metallic element such as copper.

Solution

First identify the two types of charged particle. Copper's giant metallic lattice contains positive metal ions and delocalised electrons that can move throughout the metal. These electrons are often described as a sea of delocalised electrons.

Then describe how they are held together. Strong electrostatic attraction between positive ions and negatively charged delocalised electrons maintains the entire metallic structure. This attraction is metallic bonding.

A complete answer is: Metallic bonding is the strong electrostatic attraction between positive metal ions in a giant metallic lattice and a sea of delocalised electrons.

This identifies positive ions, delocalised or mobile electrons, and the attraction between them. Saying only that copper has strong bonds does not identify the bonded particles. Saying only that electrons move mainly explains conduction, not the complete meaning of bonding.

Practice

1. (2025·May/June 0620/41 Q3(a)(ii)) · Level 1

This question is about copper and compounds of copper.

(ii) Explain how solid copper conducts electricity.

2. (2025·May/June 0620/43 Q3(a)(i)) · Level 1

This question is about zinc and compounds of zinc.

(a) Zinc is held together by electrostatic forces of attraction between particles.

(i) Name the type of bonding in zinc.

3. (2025·May/June 0620/43 Q3(a)(ii)) · Level 1

This question is about zinc and compounds of zinc.

(a) Zinc is held together by electrostatic forces of attraction between particles.

(ii) Name the **two** types of particles held together by the bonding in **(a)(i)**.

4. (2025·May/June 0620/43 Q3(a)(iii)) · Level 1

This question is about zinc and compounds of zinc.

(a) Zinc is held together by electrostatic forces of attraction between particles.

(iii) Name the type of particle whose movement allows zinc to conduct electricity.

5. (2025·October/November 0620/21 Q7) · Level 2

Which statement explains why solid magnesium is a good conductor of electricity?

A Magnesium has delocalised electrons that can move through the lattice.

B Magnesium has strong attractions between ions.

C Magnesium consists of a giant lattice of anions in a ‘sea’ of delocalised electrons.

D Magnesium consists of metal ions that move through the lattice.

Stoichiometry

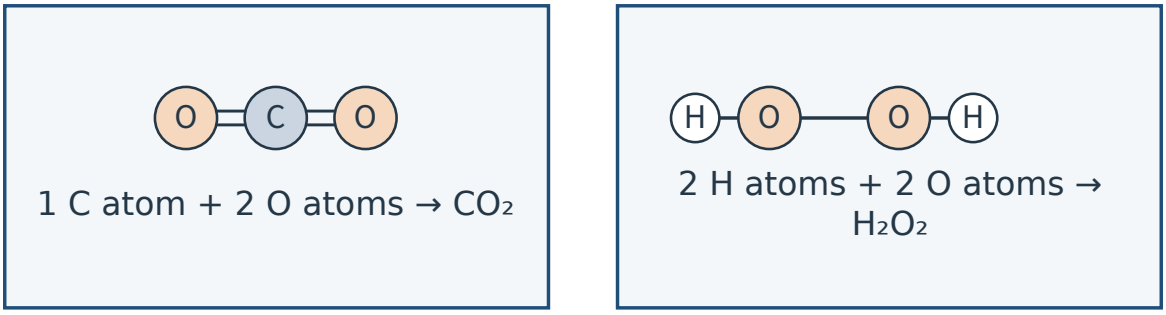
3.1 Chemical formulae and equations

Type 1: Writing formulae from atom ratios or ionic charges

Key Ideas A chemical formula first tells us what a substance is and how many of the relevant particles it contains. An element symbol begins with a capital letter; any second letter is lower case. For example, Co is cobalt and Cu is copper, whereas CO combines the element symbols C and O and represents carbon monoxide. Mixing up capitals and lower-case letters changes the substance.

A **molecular formula** gives the types and actual numbers of atoms in one molecule. A small number at the lower right is a **subscript** and counts only the element immediately before it. No subscript means one. CO_2 contains one C and two O atoms; NH_3 contains one N and three H atoms. If a model is supplied, first identify one complete molecule, then count each atom type. Do not use the total across several molecules as the formula of one molecule.

A molecular formula counts atoms in one whole molecule; coefficients count molecules separately



Molecular formulae retain actual atom counts. H_2O_2 cannot be reduced to HO to represent that molecule.

Identify one molecule before counting its atoms. Circles and bond lengths are not drawn to actual scale.

For example, if a diagram shows three separate CO_2 molecules, each still contains one C and two O atoms and its molecular formula remains CO_2. To indicate three such molecules, write $3CO_2$. Molecular formulae give actual numbers, so hydrogen peroxide, H_2O_2, cannot be cancelled down to HO as if it were a fraction. Section 3.3 introduces empirical formulae, which give only the simplest ratio.

Some common elements exist as diatomic molecules. Their elemental formulae are H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2 for hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine respectively. Noble gases such as helium He, neon Ne, argon Ar and xenon Xe are monatomic, so do not add 2. Metallic elements usually use the element symbol directly, such as Mg, Al, Fe and Cu. Ionic compounds need the charge-balancing method below to determine their ratios.

The formula of an ionic compound gives the simplest whole-number ratio of positive and negative ions. As the giant lattice showed earlier, it does not consist of isolated small salt molecules. For the material to be electrically neutral overall, total positive and negative charges must cancel. Write each ion's charge first, then find the smallest numbers needed to balance them. Do not copy group numbers or charges straight into a formula.

Ionic compounds: find the smallest whole-number ratio giving zero total charge



Al : O = 2 : 3. Total charge: +6 − 6 = 0. Formula: Al_2O_3.

The diagram gives the smallest number ratio; it does not show an isolated Al_2O_3 molecule within the crystal.

For example, Ca^{2+} and O^{2-} combine in a 1:1 ratio and their charges exactly cancel. Calcium oxide is therefore CaO, not Ca_2O_2.

Al^{3+} and O^{2-} have different charge magnitudes. The lowest common multiple of 3 and 2 is 6: two Al^{3+} ions have total charge +6, and three O^{2-} ions have total charge −6. Their simplest ratio is therefore 2:3, giving Al_2O_3.

If a model shows four Ca^{2+} and four O^{2-} ions, count the ratio 4:4 and then simplify it to 1:1. The total number in a finite lattice fragment is not the composition of a separate molecule.

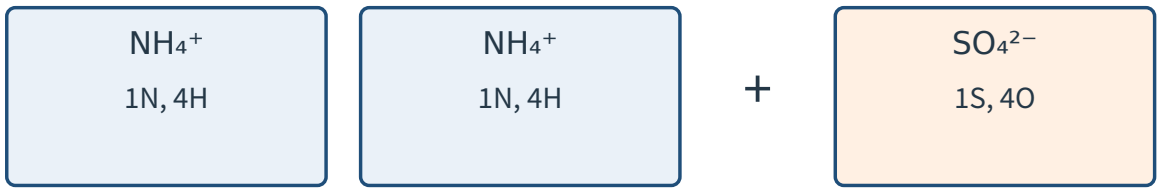
Roman numerals specify a metal ion's charge, not its atom count. Copper(II) means Cu^{2+}, cobalt(II) means Co^{2+}, iron(II) means Fe^{2+} and iron(III) means Fe^{3+}. For example, Co^{2+} needs two Cl^- ions, so cobalt(II) chloride is $CoCl_2$. Cu^{2+} combines with one O^{2-}, so copper(II) oxide is CuO. The (II) does not mean that there must always be two metal atoms.

Some ions contain several atoms and are called **polyatomic ions**. They carry a charge as a whole. Do not change their internal atom numbers to balance charges. Common negative ions include chloride Cl^-, bromide Br^-, iodide I^-, oxide O^{2-} and sulfide S^{2-}. Learn the names and formulae of the grouped ions below as well.

Ion name	Term in questions	Ion formula
Hydroxide ion	hydroxide	OH^-
Nitrate ion	nitrate	NO_3^-
Carbonate ion	carbonate	CO_3^{2-}
Sulfate ion	sulfate	SO_4^{2-}
Ammonium ion	ammonium	NH_4^+

Ammonium, NH_4^+, is a positive ion even though it contains no metal. Two ammonium ions combine with one sulfate ion, giving total charge +2−2=0 and ammonium sulfate, $(NH_4)_2SO_4$. When two or more identical groups are needed, put brackets around the whole group and write its number outside. One group usually needs no brackets. For example, copper(II) sulfate, $CuSO_4$, contains one Cu^{2+} and one SO_4^{2-}. The internal O_4 in sulfate is not its ionic charge.

Brackets enclose a whole group: the outside subscript multiplies every atom inside



$(NH_4)_2SO_4$: N=2, H=8, S=1, O=4

The +2 charge of two NH_4^+ ions balances the −2 charge of one SO_4^{2-} ion.

NH_4^+ is ammonium and SO_4^{2-} is sulfate. The 2 outside the brackets multiplies both N and H in NH_4.

For example, calcium hydroxide needs two OH^- ions to balance one Ca^{2+}, giving $Ca(OH)_2$. It contains one Ca, two O and two H atoms. In $CaOH_2$, the subscript 2 would apply only to H and would change the composition.

For sodium carbonate, two Na^+ ions combine with one CO_3^{2-} to give Na_2CO_3. For magnesium chloride, one Mg^{2+} combines with two Cl^- to give $MgCl_2$. Check the charge balance each time before fixing the ion numbers.

Later calculations in this chapter also use these positive ions: H^+, Na^+, K^+ and Ag^+ have charge 1+; Mg^{2+}, Ca^{2+}, Ba^{2+} and Zn^{2+} have charge 2+; Al^{3+} has charge 3+. Ag is silver, Ba is barium and Zn is zinc. Check both ionic charge and the capitalisation of element symbols.

Common monatomic main-group ions: use the group to find charge, then balance the ratio



Transition metals can have different charges. Use the Roman numeral in names such as copper(II).

This table relates common ions to their charges. Do not mechanically copy an arbitrary element's group number into a formula as a subscript.

If only a Periodic Table position is given, return to Chapter 2's link between groups and electronic configurations. Common Group I metals lose one electron and Group II metals lose two; Group VII non-metals commonly gain one and Group VI non-metals gain two. Then balance total charge. For example, a salt formed by a Group II metal and a Group VII non-metal needs one 2+ ion for every two 1− ions. Do not assign transition-metal charges simply from their columns: use the Roman numeral in the name or the value supplied.

Prefixes such as mono and di can directly specify atom numbers in a molecule. Carbon monoxide is CO and carbon dioxide is CO_2; sulfur dioxide is SO_2 and sulfur trioxide is SO_3. Water is H_2O, ammonia NH_3, nitric acid HNO_3, sulfuric acid H_2SO_4 and hydrochloric acid HCl(aq). These names and formulae are used directly in later reaction descriptions. Anhydrous means without water, as in anhydrous copper(II) sulfate, $CuSO_4$. Do not add water of crystallisation when the question specifies only the anhydrous salt.

Recognising the Question For a molecular model, count the atoms in one complete molecule. For an ionic compound, write each ion's charge and find the simplest whole-number ratio with total charge zero. A number outside brackets multiplies the whole group. Roman numerals indicate charge.

Worked Example 1 (2025·February/March 0620/22 Q9) · Level 2

Which formula for the named compound is correct?

A calcium oxide, CaO

B cobalt(II) chloride, Co_2Cl

C sulfur dioxide, S_2O_2

D anhydrous copper(II) sulfate, $Cu(SO_4)_2$

Solution

Start with A. Calcium oxide contains Ca^{2+} and O^{2-} ions. The +2 from one Ca^{2+} cancels the −2 from one O^{2-}, giving the smallest ratio 1:1 and formula CaO. **Choose A**.

Check the errors in the other options individually. In B, cobalt(II) chloride needs Co^{2+} with two Cl^- ions, so its formula is $CoCl_2$, not Co_2Cl. In C, sulfur dioxide contains one S and two O atoms per molecule, giving SO_2. In D, anhydrous copper(II) sulfate combines Cu^{2+} and SO_4^{2-} in a 1:1 ratio, giving $CuSO_4$. $Cu(SO_4)_2$ would give total charge +2−4=−2 and cannot represent a neutral salt.

Practice

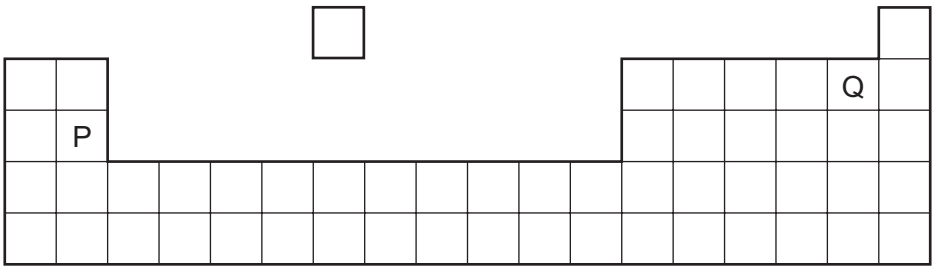
1. (2025·May/June 0620/11 Q9) · Level 2

Which row shows the correct formula for the named substance?

	substance	formula
A	cobalt(II) chloride	Cu_2Cl
B	sodium carbonate	Na_2CO_3
C	xenon	Xe_2
D	ammonium sulfate	NH_4SO_4

2. (2025·May/June 0620/23 Q5) · Level 3

The diagram shows the positions of elements P and Q in the Periodic Table.



What is the formula of the compound that is formed between P and Q?

A PQ **B** P_2Q **C** PQ_2 **D** P_2Q_7

Type 2: Balancing equations by conservation and writing state symbols and ionic equations

Key Ideas

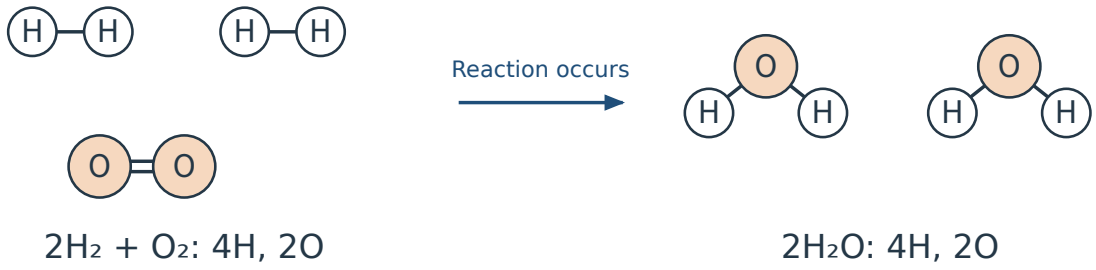
A **chemical equation** shows reactants changing into products. Reactants appear on the left, products on the right, and the arrow gives the direction of change. For example, magnesium reacts with oxygen to form magnesium oxide. Begin with the word equation:

magnesium + oxygen → magnesium oxide

For a symbol equation, replace these names with Mg, O_2 and MgO. If a reaction description or products are supplied, use that information; do not add other products from a similar question.

Ordinary chemical reactions change connections between atoms but **do not create or destroy atoms of an element**. The total number of atoms of each element must therefore be the same on both sides of a complete equation. This is **conservation of atoms**. Rearrangement can change the substances present, but cannot make an atom disappear.

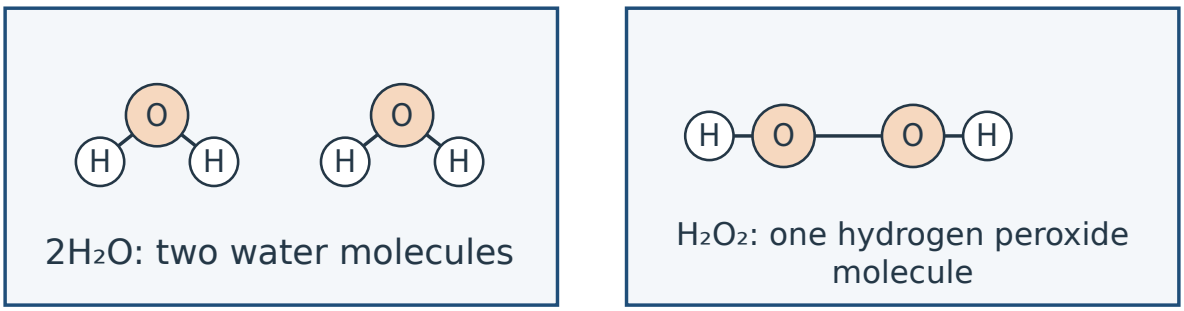
Chemical reactions reconnect atoms while conserving each element's total atom count



The arrow shows atoms rearranging into new molecules. No H or O atoms are created or lost.

A number before a formula is a **coefficient** and gives the relative number of those particles. For example, $2H_2O$ means two H_2O units, containing four H and two O atoms in total. H_2O_2 is a different substance, hydrogen peroxide, with two H and two O atoms in each molecule. **Balancing** chooses coefficients in front of already correct formulae to make each atom count equal. Do not alter subscripts and thereby change the reactants or products.

A coefficient changes quantity; a subscript changes one molecule's composition



Balance using coefficients only. Do not change H_2O to H_2O_2 to balance oxygen atoms.

Both sides contain only H and O, but the numbers of particles and the composition of each particle differ.

For example, hydrogen reacts with oxygen to form water. First write the correct formula for each substance:

$H_2 + O_2 \rightarrow H_2O$

The left has two O atoms, while each H_2O on the right has one, so put 2 before H_2O. The right now has four H atoms, requiring two H_2 units on the left:

$2H_2 + O_2 \rightarrow 2H_2O$

Check that H totals four on each side and O totals two on each side. Even if O was balanced first, check H again at the end.

With several elements, it is often helpful to balance those appearing in only a few formulae first, then deal with O or H appearing throughout. If a coefficient such as 1/2 or 3/2 appears, multiply **every coefficient in the whole equation by two** to obtain the smallest whole-number coefficients. If a question fixes $4NH_3$ or $4FeS_2$, keep that supplied 4 and balance the remaining terms. Some questions use the reversible arrow $\rightleftharpoons$, meaning that the reaction can proceed in both directions. Balancing still checks conservation of atoms on the two sides; retain the supplied arrow. The effects of conditions on equilibrium are discussed later.

You may also need **state symbols** after the formulae and coefficients. The brackets immediately after a formula describe the substance's state; they are not brackets around a polyatomic group.

State symbols follow formulae: distinguish an aqueous solution from a pure liquid

(s)	(l)	(g)	(aq)
Solid	Liquid	Gas	Aqueous solution
$Mg(s)$	$H_2O(l)$	$H_2(g)$	$NaCl(aq)$

$HCl(aq)$ is hydrogen chloride dissolved in water, forming hydrochloric acid; it is not pure liquid $HCl(l)$.

s, l and g give physical states; aq means that a substance is present in aqueous solution.

(s) means solid, (l) liquid, (g) gas and (aq) dissolved in water. $H_2O(l)$ is liquid water itself, whereas $NaCl(aq)$ is sodium chloride in water, not pure liquid sodium chloride, $NaCl(l)$. Dilute hydrochloric acid is therefore $HCl(aq)$.

Preserve information such as gas, solid precipitate or aqueous solution in the equation. A **precipitate** is an insoluble solid formed when solutions are mixed and is labelled (s).

For example, sodium hydroxide solution reacts with dilute hydrochloric acid to form sodium chloride solution and water. Begin with the word equation:

sodium hydroxide + hydrochloric acid

$\rightarrow$ sodium chloride + water

The formulae are NaOH, HCl, NaCl and H_2O respectively. Count Na, O, H and Cl: all are conserved with coefficients of one. Add the states for the solutions and liquid water:

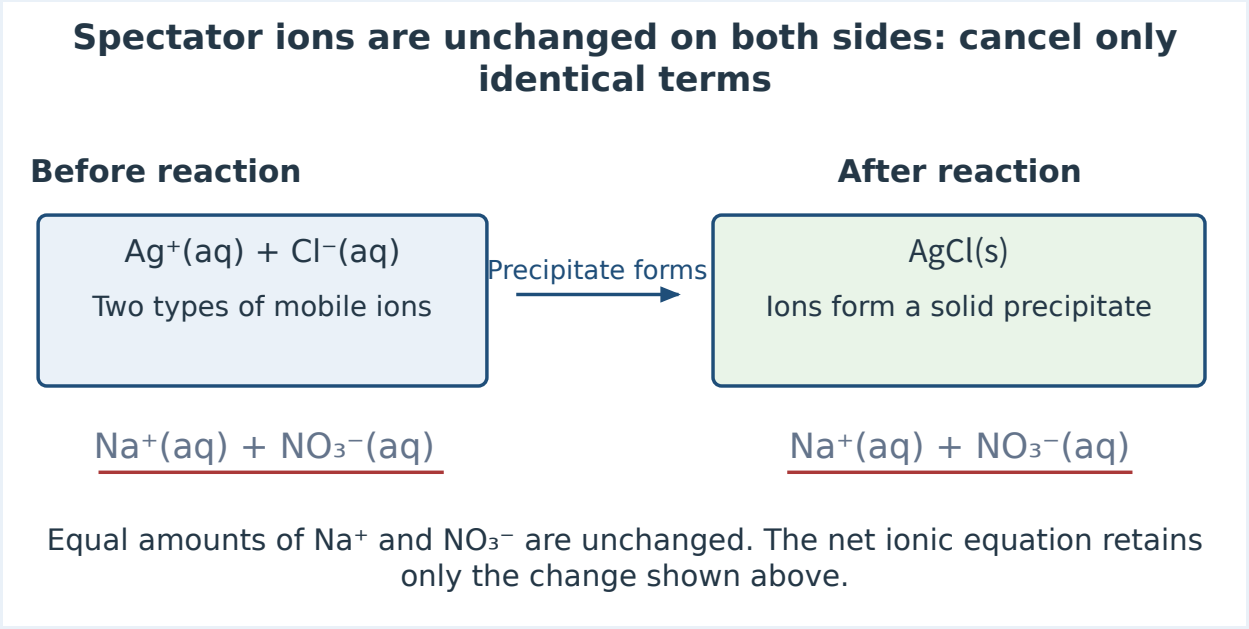
$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

The two H atoms in water have identifiable sources: one from NaOH and one from HCl.

An **ionic equation** focuses on the particles that actually change in solution. Chapter 2 established that soluble ionic salts in water contain mobile positive and negative ions. In a complete ionic equation, write these aqueous salts as their ions.

For example, write $NaCl(aq)$ as $Na^+(aq)$ and $Cl^-(aq)$, and $CaCl_2(aq)$ as $Ca^{2+}(aq)$ and $2Cl^-(aq)$. Any coefficient before the formula must multiply all the ions. The dilute strong acids HCl and HNO_3, and soluble strong bases NaOH and $Ba(OH)_2$ used here, are also written as ions, supplying H^+ or OH^- respectively.

Do not split every substance marked (aq) into ions: some dissolved substances remain neutral molecules. Solid precipitates, pure liquid water and neutral molecules such as Cl_2 and Br_2 remain intact in these ionic equations. Do not split $Fe(OH)_3(s)$ into freely moving aqueous ions.



The reaction is supplied: silver nitrate and sodium chloride form an AgCl precipitate. You do not have to infer solubility from missing information.

Mixing silver nitrate solution and sodium chloride solution forms a silver chloride precipitate. $AgNO_3$ is silver nitrate, AgCl silver chloride and $NaNO_3$ sodium nitrate. The given reaction is:

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Write the dissolved soluble ionic salts as ions, but retain the solid precipitate AgCl intact, giving the complete ionic equation:

$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq)$

$\rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

Na^+ and NO_3^- have the same identities, quantities and states on both sides. These are **spectator ions**. Cancel them from both sides to leave the net ionic equation:

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Cancel **identical particles present in equal amounts on both sides**, not all positive ions or all metal ions. Ag^+ enters the precipitate and cannot be cancelled; Na^+ remains unchanged and can. Finally check both atom counts and total charge. For example, $Ag^+ + Cl^-$ gives charge $+1-1=0$ on the left, matching neutral AgCl on the right. In a later reaction where Fe^{3+} forms a precipitate with OH^-, three OH^- are required to make the total charge zero.

For example, the following halogen reaction is given:

$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$

Follow the stated direction, write the dissolved potassium salts as ions and cancel identical K^+ ions from both sides. Cl_2 is a neutral diatomic molecule, whereas Cl^- is a chloride ion: do not interchange them. Check atoms and total charge at the end. Since the reaction direction is supplied, you need not determine which halogen is more reactive in advance.

Recognising the Question Use the names or reaction information to write correct formulae. Change only coefficients to conserve each element, then add states. For ionic equations, split soluble electrolytes according to the particles actually present, cancel equal amounts of spectator ions from both sides, and check both atoms and charge.

Worked Example 1 (2025·May/June 0620/22 Q8) · Level 3

Aluminium reacts with iron(III) oxide to form aluminium oxide and iron.

Which chemical equation for the reaction between aluminium and iron(III) oxide is correct?

A $Al + FeO \rightarrow AlO + Fe$

B $2Al + Fe_2O \rightarrow Al_2O + 2Fe$

C $Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$

D $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

Solution

Aluminium reacts with iron(III) oxide to form aluminium oxide and iron. Fe^{3+} and O^{2-} balance in a 2:3 ratio, giving iron(III) oxide, Fe_2O_3. Al^{3+} and O^{2-} also balance in a 2:3 ratio, giving aluminium oxide, Al_2O_3. Write the skeleton equation:

$Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$

Fe_2O_3 and Al_2O_3 each contain three O atoms, so oxygen is already balanced. Al_2O_3 on the right contains two Al atoms, so put 2 before Al on the left. Fe_2O_3 on the left contains two Fe atoms, so put 2 before Fe on the right:

$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

Choose D.

Check each element: both sides have two Al, two Fe and three O atoms. A and B use incorrect formulae for the compounds. C has correct compound formulae but has not balanced Al and Fe.

Worked Example 2 (2025·May/June 0620/41 Q4(c)(iv)) · Level 3

(iv) In step 2, oxygen and water react with NO to produce nitric acid as the only product.

Write a symbol equation for this chemical reaction.

Solution

First convert the names to formulae. Oxygen is O_2 and water is H_2O. NO is nitrogen monoxide and nitric acid is HNO_3. The question explicitly gives HNO_3 as the only product, so the skeleton is:

$NO + O_2 + H_2O \rightarrow HNO_3$

Second balance nitrogen and hydrogen. Start with $2HNO_3$ on the right, giving two N and two H atoms. This requires 2NO and $1H_2O$ on the left. Those two terms already supply three O atoms, while $2HNO_3$ needs six. Three more O atoms are therefore needed, corresponding to a coefficient of 3/2 for O_2.

Third multiply all coefficients by two to remove the fraction:

$4NO + 3O_2 + 2H_2O \rightarrow 4HNO_3$

The final totals on both sides are N=4, H=4 and O=12. Only a symbol equation is requested, so state symbols are not required. Nitric acid being soluble in water is not a reason to add an unspecified product.

Worked Example 3 (2025·May/June 0620/23 Q7) · Level 3

The equation for the reaction between iron(III) chloride and sodium hydroxide is shown.

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$$

What is the ionic equation for this reaction?

A $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe^{3+}(s) + 3OH^-(s)$

B $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

C $3Na^+(aq) + 3Cl^-(aq) \rightarrow 3Na^+(aq) + 3Cl^-(aq)$

D $3Na^+(aq) + 3Cl^-(aq) \rightarrow 3NaCl(aq)$

Solution

$FeCl_3(aq)$ and $NaOH(aq)$ supply Fe^{3+}, Cl^-, Na^+ and OH^- in water. The product $Fe(OH)_3(s)$ is a solid precipitate and must remain intact. First write the complete ionic equation:

$Fe^{3+}(aq) + 3Cl^-(aq) + 3Na^+(aq) + 3OH^-(aq)$

$\rightarrow Fe(OH)_3(s) + 3Na^+(aq) + 3Cl^-(aq)$

The $3Na^+$ and $3Cl^-$ have identical identities, amounts and states on both sides. They are spectator ions and cancel in pairs. The actual change is:

$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

Choose B. The left-hand charge is +3−3=0, matching the neutral precipitate. Fe=1, O=3 and H=3 are also conserved.

A splits the solid precipitate into separate ions and loses the precipitation change. C retains only unchanged spectator ions. D incorrectly treats the Na^+ and Cl^- that remain in solution as forming a new solid.

Practice

1. (2025·May/June 0620/41 Q4(c)(i)) · Level 2

(i) Balance the symbol equation for this reaction.

$$4NH_3 + ...O_2 \rightleftharpoons .....NO + ...H_2O$$

2. (2025·May/June 0620/42 Q4(a)(i)) · Level 2

(a) The main ore of chromium is chromite. Chromite contains $FeCr_2O_4$.

$FeCr_2O_4$ reacts with carbon.

(i) Complete the equation for this reaction.

$$FeCr_2O_4 + 4C \rightarrow Fe + 2Cr + ...$$

3. (2025·May/June 0620/43 Q4(a)(i)) · Level 2

(i) Iron pyrite is an ore containing sulfur. The ore contains a compound with the formula FeS_2.

FeS_2 reacts with oxygen in the air to produce sulfur dioxide.

Balance the symbol equation for this reaction.

$$4FeS_2 + ...O_2 \rightarrow .....SO_2 + ...Fe_2O_3$$

4. (2025·May/June 0620/13 Q8) · Level 3

Which equation represents the reaction of magnesium with dilute hydrochloric acid?

A $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

B $Mg(s) + 2HCl(l) \rightarrow MgCl_2(s) + H_2(g)$

C $2Mg(s) + 2HCl(aq) \rightarrow 2MgCl(aq) + H_2(g)$

D $2Mg(s) + 2HCl(l) \rightarrow 2MgCl(s) + H_2(g)$

5. (2025·May/June 0620/41 Q6(d)(i)) · Level 3

(d) When bromine is added to aqueous potassium iodide a displacement reaction occurs.

The equation for the reaction is shown.

$$Br_2 + 2KI \rightarrow 2KBr + I_2$$

(i) Write an ionic equation for the reaction.

6. (2025·May/June 0620/43 Q6(e)(ii)) · Level 3

(ii) When chlorine is passed through aqueous potassium bromide, a displacement reaction occurs.

The equation for the reaction is shown.

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$$

Write an ionic equation for the reaction.

3.2 Relative masses

Type 1: Calculating relative molecular or formula mass and using reaction mass ratios

Key Ideas An atom's actual mass is extremely small, so expressing it directly in grams is inconvenient. **Relative atomic mass, A_r**, compares the average mass of an element's atoms, accounting for its isotopes, with one-twelfth of the mass of a carbon-12 atom. This common reference puts different elements on the same comparison scale. Chapter 2 explained that an element may have several isotopes. Use the average that accounts for their composition, not an arbitrarily chosen mass number.

Relative atomic masses use the same reference mass

Common reference
1/12 of the mass of one ^{12}C atom

Compare →

For example, $A_r(O) = 16$
The average mass of an O atom is about 16 times this reference

A_r is a mass ratio, with no g unit. It is not the mass number of every isotope.

The reference is one-twelfth of a mass; a carbon atom is not physically cut up. The average mass must account for isotopic composition.

An A_r of 16 for O means that oxygen's average atomic mass is 16 times this reference. It does not mean that one oxygen atom weighs 16g. A_r is a **ratio with no unit of g**. Similarly, chlorine's A_r can be 35.5 without an atom having half a neutron. Relative average mass and the mass number of one atom are different quantities.

Relative molecular mass, M_r, is the sum of the A_r values of all atoms in a molecular formula. For example, CO_2 contains one C and two O atoms:

$$M_{r(CO_2)} = 12 + 2 \times 16 = 44$$

Ionic compounds have no separate molecules, so use the term **relative formula mass, M_r**. The calculation still adds every atom's relative-mass contribution in the formula, for example:

$$M_{r(NaCl)} = 23 + 35.5 = 58.5$$

Count atoms, multiply, then add. Do not simply add a few Periodic Table values while overlooking subscripts or brackets. $Mg(NO_3)_2$ contains one Mg and two NO_3 groups, so it has two N and six O atoms.

Expand brackets to count atoms, then add each element's mass contribution

$Mg(NO_3)_2$

Mg: 1 atom	N: 2 atoms	O: 6 atoms
1×24	2×14	6×16
24	28	96

$M_r = 24 + 28 + 96 = 148$

The 2 in magnesium nitrate applies to the entire NO_3 group: two N and six O atoms are present. M_r is relative formula mass.

For example, $Ca(OH)_2$ contains only one Ca, while the 2 outside the brackets multiplies both O and H:

$$M_{r(Ca(OH)_2)} = 40 + 2 \times 16 + 2 \times 1 = 74$$

Alternatively, calculate the relative mass of one OH group first, then add two groups to Ca:

$$M_{r(Ca(OH)_2)} = 40 + 2 \times (16 + 1) = 40 + 2 \times 17 = 74$$

Both methods account for every atom. Multiplying Ca by two as well, or multiplying only H by two, calculates a different composition.

In a table question, each row may give the number of an element's atoms, that element's A_r and the product of the two. A supplied oxygen row still contributes to the total; it is not an example to leave out afterwards. Complete every row, then add all the contributions in the right-hand column to obtain M_r. M_r still has no unit of g.

Simple reaction mass ratios can be used before introducing moles. A balanced equation gives particle-number ratios. Knowing each particle's relative mass lets you convert those number ratios to mass ratios, then scale them together to the mass given.

Equation coefficients give particle ratios; multiply by relative masses for mass ratios

$$2Mg + O_2 \rightarrow 2MgO$$

Mg	O_2	MgO
2 portions × 24	1 portion × 32	2 portions × 40
48g	32g	80g

48 : 32 : 80 scales proportionally: 6 g Mg reacts with 4 g O_2 to give 10 g MgO.

The number ratio 2:1:2 differs from the mass ratio 48:32:80. When scaling masses, multiply all three by the same factor.

The balanced equation for magnesium reacting with oxygen is:

$2Mg + O_2 \rightarrow 2MgO$

The particle-number ratio of Mg to O_2 is 2:1. Each Mg has relative mass 24 and each O_2 relative mass 32. Multiply the numbers by the respective relative masses to obtain the mass ratio:

$$m(\mathrm{Mg}) : m(\mathrm{O_2}) = (2 \times 24) : (1 \times 32) = 48 : 32$$

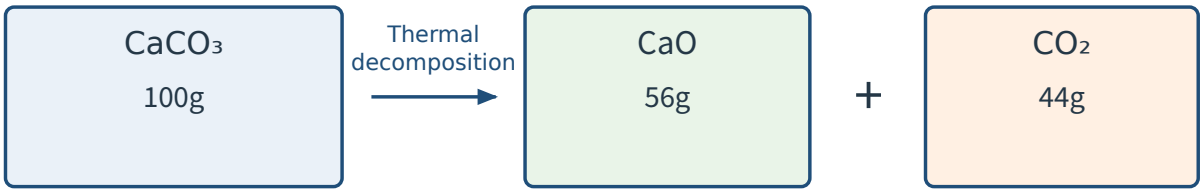
Thus 48g of Mg reacts exactly with 32g of O_2. For only 6g of Mg, reduce both masses by the same factor:

$$m(\mathrm{O_2}) = 32 \times \frac{6}{48} = 4\,\mathrm{g}$$

The method first establishes a mass ratio, then scales it proportionally. It does not treat the coefficient ratio 2:1 as a ratio of grams.

Conservation of mass must include **all** reactants and products. If one reactant decomposes into two products and only one product's mass is requested, do not assign the entire reactant mass to that product.

Not all the reactant mass remains in the one product being considered



100 g = 56 g + 44 g. Add both product masses to recover the reactant mass.

The reaction $CaCO_3 \rightarrow CaO + CO_2$ is supplied. Use A_r(Ca)=40, C=12 and O=16 to establish the mass ratio.

For example, the thermal decomposition of calcium carbonate is given as:

$CaCO_3 \rightarrow CaO + CO_2$

The relative formula or molecular masses are 100, 56 and 44 respectively, with all coefficients equal to one, so the mass ratio is 100:56:44. For complete decomposition of 25g of $CaCO_3$, multiply each mass by the same factor:

$$m(\mathrm{CaO}) = 56 \times \frac{25}{100} = 14\,\mathrm{g}$$

$$m(\mathrm{CO_2}) = 44 \times \frac{25}{100} = 11\,\mathrm{g}$$

The products total 14+11=25g, matching the original total mass. Section 3.3 expresses this proportional method more systematically using moles.

Recognising the Question Use subscripts and brackets to count each type of atom. Multiply by A_r and add to obtain M_r. For reaction masses, first use coefficient × relative mass to establish the mass ratio, then multiply the corresponding masses by the same scaling factor.

Worked Example 1 (2025·May/June 0620/31 Q3(c)) · Level 2

(c) A compound of sulfur has the formula $Al_2(SO_4)_3$.

Complete Table 3.2 to calculate the relative formula mass of $Al_2(SO_4)_3$.

Table 3.2

atom	number of atoms	relative atomic mass	
oxygen	12	16	12 × 16 = 192
aluminium		27	
sulfur		32	

Solution

In $Al_2(SO_4)_3$, the subscript 2 applies only to Al. The 3 outside the brackets means three SO_4 groups. There are therefore two Al, 3×1=3 S and 3×4=12 O atoms.

The table already gives oxygen's contribution as 12×16=192. Complete the other two rows:

Atom	Number	A_r	Contribution
oxygen	12	16	12×16=192
aluminium	2	27	2×27=54
sulfur	3	32	3×32=96

Add every mass contribution in the table:

$$M_r = 192 + 54 + 96 = 342$$

Relative formula mass has no unit of g.

Giving only the two atom counts is incomplete: include the products in the right-hand column and add every row's contribution to calculate the relative formula mass.

Worked Example 2 (2025·May/June 0620/12 Q11) · Level 3

The equation for the reaction of magnesium with oxygen is shown.

$$2Mg + O_2 \rightarrow 2MgO$$

Which mass of oxygen reacts exactly with 6 g of magnesium?

A 2 g **B** 4 g **C** 8 g **D** 16 g

Solution

The balanced equation is:

$2Mg + O_2 \rightarrow 2MgO$

Two Mg atoms correspond to one O_2 molecule, but the coefficient ratio 2:1 is not the mass ratio. Mg has A_r=24 and O_2 has M_r=32, so the corresponding mass ratio is:

$$m(\mathrm{Mg}) : m(\mathrm{O_2}) = (2 \times 24) : 32 = 48 : 32$$

48g of Mg uses exactly 32g of O_2. For the stated 6g of Mg, multiply the oxygen mass by the same factor:

$$m(\mathrm{O_2}) = 32 \times \frac{6}{48} = 4\,\mathrm{g}$$

Choose B. As a mass check, the MgO produced is 6+4=10g. Dividing 6 directly by coefficient 2 to obtain 3g would incorrectly use a particle-number ratio as a mass ratio.

Practice

1. (2025·May/June 0620/32 Q3(c)) · Level 2

(c) A compound of phosphorus has the formula $Ca_3(PO_4)_2$.

Complete Table 3.2 to calculate the relative formula mass of $Ca_3(PO_4)_2$.

Table 3.2

atom	number of atoms	relative atomic mass	
oxygen	8	16	8 × 16 = 128
calcium		40	
phosphorus		31	

2. (2025·May/June 0620/33 Q4(c)) · Level 2

(c) A compound of nitrogen has the formula $Al(NO_3)_3$.

Complete Table 4.2 to calculate the relative formula mass of $Al(NO_3)_3$.

Table 4.2

atom	number of atoms	relative atomic mass	
oxygen	9	16	9 × 16 = 144
aluminium		27	
nitrogen		14	

3. (2025·May/June 0620/11 Q10) · Level 3

The equation shows the thermal decomposition of magnesium carbonate.

$$MgCO_3 \rightarrow MgO + CO_2$$

[M_r: $MgCO_3$, 84]

Which mass of magnesium oxide is formed when 21.0 g of magnesium carbonate is completely decomposed?

A 1.9 g **B** 4.0 g **C** 10.0 g **D** 40.0 g

3.3 Amount of substance and quantitative calculations

Type 1: Converting between mass, amount of substance and particle number

Key Ideas In everyday life, a dozen means 12 items. Chemistry also needs a fixed counting unit for vast numbers of tiny particles. The **mole**, unit mol, is the unit of **amount of substance**. One mole contains 6.02×10^{23} specified particles. This number is the **Avogadro constant** used in this course's calculations. Specify what is being counted: Ar atoms, H_2 molecules or a particular ion, rather than an undefined collection of particles.

One mole counts the same number of entities, but identify which entities are being counted

1 mol Ar	1 mol H_2
6.02×10^{23} Ar atoms total atom count = 6.02×10^{23}	6.02×10^{23} H_2 molecules total H atom count = 1.204×10^{24}

The mole measures amount of substance. Counting molecules and counting all their atoms are distinct tasks.

Each H_2 molecule contains two H atoms. Equal amounts in mol mean equal numbers of the specified particles, but not necessarily equal total atom numbers.

1 mol of Ar contains 6.02×10^{23} Ar atoms. 1 mol of H_2 contains the same number of H_2 molecules, but each molecule contains two H atoms, so the total H atom number is $2\times6.02\times10^{23}$. The mole is neither a gram nor a volume unit. Equal amounts of different substances in mol contain equal numbers of specified particles, but may have different masses.

Molar mass, M, is the mass of one mole of a substance, commonly measured in g/mol. Its numerical value corresponds to the substance's A_r or M_r, but the units differ:

Substance	Relative mass, no unit	Molar mass
Ar	$A_r = 40$	$M = 40\,\mathrm{g/mol}$
CO_2	$M_r = 44$	$M = 44\,\mathrm{g/mol}$

Relative mass is a ratio; molar mass tells you the number of grams per mole.

Molar mass M is the mass of one mole of a substance

Ar	H_2O	CO_2
A_r=40	M_r=18	M_r=44
M=40g/mol	M=18g/mol	M=44g/mol

n = m ÷ M; rearrange to m = n × M or M = m ÷ n. Use the units to choose the operation.

A_r and M_r are relative numbers without units. M has the corresponding numerical value but the unit g/mol.

Divide the sample's total mass, m, by its mass per mole, M, to obtain its amount of substance:

$$n = \frac{m}{M}$$

Here n is in mol, m in g and M in g/mol. If the amount is known and mass is required, multiply the mass per mole by the number of moles:

$$m = nM$$

If mass and amount are known, find the mass per mole using:

$$M = \frac{m}{n}$$

These are three forms of the same proportional relationship. Identify the unknown first, then choose the required direction.

For example, a 9.0g sample of water has molar mass 18g/mol. Divide total mass by molar mass to find its amount:

$$n(\mathrm{H_2O}) = \frac{9.0}{18} = 0.50\,\mathrm{mol}$$

For 0.25mol of CO_2 with molar mass 44g/mol, use multiplication to find the mass:

$$m(\mathrm{CO_2}) = 0.25 \times 44 = 11\,\mathrm{g}$$

If 0.20mol of a molecular substance has mass 8.8g, its molar mass is:

$$M = \frac{8.8}{0.20} = 44\,\mathrm{g/mol}$$

Its corresponding relative molecular mass, M_r, is 44. If instead 0.100mol of a monatomic element has mass 4.00g, the same calculation gives:

$$M = \frac{4.00}{0.100} = 40.0\,\mathrm{g/mol}$$

Here the corresponding value is A_r=40.0. A relative-mass value alone does not uniquely identify a molecular formula or an element.

Each mole contains 6.02×10^{23} specified particles. For a known amount n, particle number is:

$$N = n \times 6.02 \times 10^{23}$$

If particle number is known, divide by the Avogadro constant to find n. For example, 3.01×10^{23} water molecules correspond to:

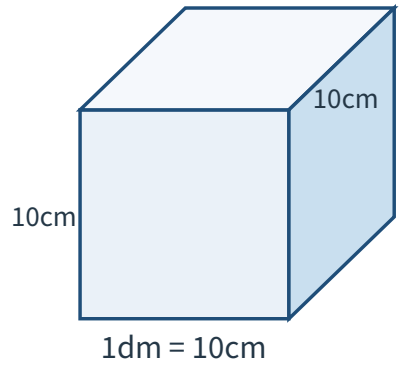
$$n(\mathrm{H_2O}) = \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}} = 0.50\,\mathrm{mol}$$

0.50mol of Na^+ likewise contains 3.01×10^{23} Na^+ ions; the specified particles are now ions. If mass is given, first divide by molar mass to obtain mol, then multiply by the Avogadro constant. Do not multiply grams directly by the constant.

Gas questions sometimes give volume rather than mass. Before using a volume relationship, check the cubic units. 1dm equals 10cm, and all three dimensions of a cube are multiplied by ten:

$$1\,\mathrm{dm^3} = (10\,\mathrm{cm})^3 = 1000\,\mathrm{cm^3}$$

1 dm = 10 cm, so this volume is 10 × 10 × 10 cm³



$$1\text{dm}^3 = 1000\text{cm}^3$$

cm³ → dm³: divide by 1000

dm³ → cm³: multiply by 1000

Converting cubic units requires the length conversion in all three dimensions. 1dm³ is neither 10cm³ nor 100cm³.

At **room temperature and pressure, r.t.p.**, this course uses a gas **molar volume of 24dm³/mol**, equivalent to 24000cm³/mol.

For volume V in dm³, amount and volume are related by:

$$n = \frac{V}{24} \quad \text{or} \quad V = 24n$$

For V in cm³, use the corresponding value 24000:

$$n = \frac{V}{24000} \quad \text{or} \quad V = 24000n$$

In both forms, n is in mol. This molar volume applies only to gases under the stated conditions, not to salt solution, liquid water, solids or arbitrary temperatures and pressures.

At the stated room temperature and pressure, use 24 dm³ per mole of gas

$H_2(g)$	$CO_2(g)$
$1\text{mol}; 24\text{dm}^3$	$1\text{mol}; 24\text{dm}^3$
Mass: 2 g	Mass: 44 g

24 dm³/mol = 24 000 cm³/mol; use only for the stated r.t.p. gas conditions.

Equal gas volumes do not mean equal masses. Do not apply the value 24 to solids, liquids or arbitrary temperatures and pressures.

For example, 12.0cm³ of O_2 is at r.t.p. First convert its volume to the amount of oxygen molecules:

$$n(O_2) = \frac{12.0}{24000} = 0.000500\,\text{mol}$$

Then multiply by the Avogadro constant to obtain the number of oxygen molecules:

$$N(O_2) = 0.000500 \times 6.02 \times 10^{23} = 3.01 \times 10^{20}$$

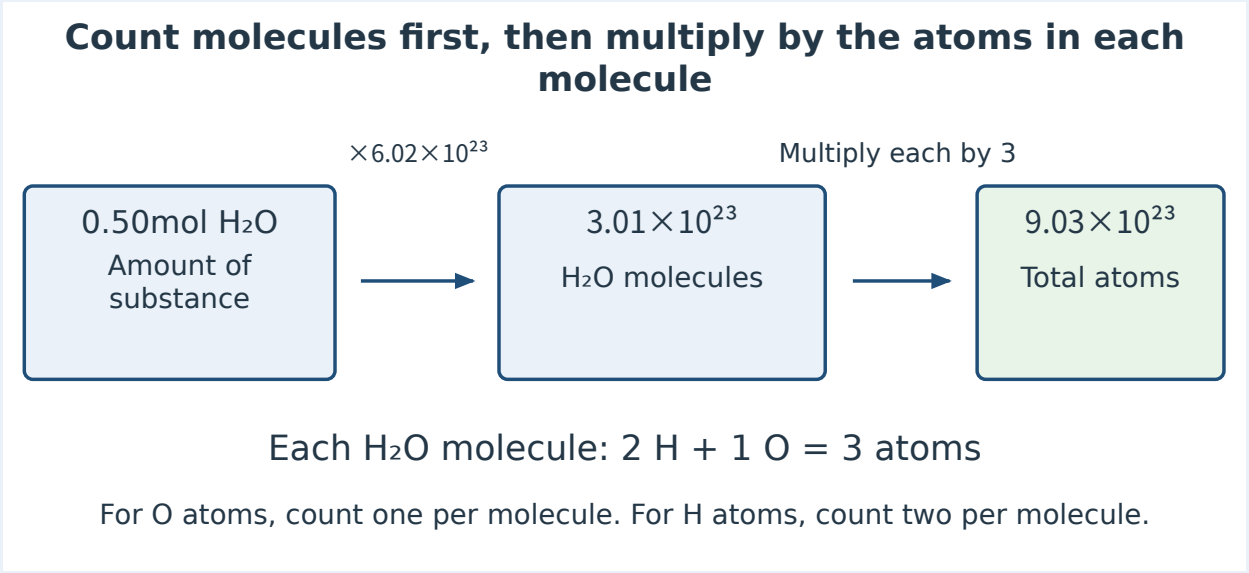
If O atoms are requested, each O_2 molecule contains two O atoms, so multiply by two again to obtain 6.02×10^{20} O atoms.

Conversely, for 0.0500mol of gas, its volume at r.t.p. can first be calculated in dm³:

$$V = 0.0500 \times 24 = 1.20\,\text{dm}^3 = 1200\,\text{cm}^3$$

Although 12.0cm³ and 12.0dm³ have the same numerical value, their volumes differ by a factor of 1000. Always check units before calculating.

Only at the final step decide whether to convert molecules into atoms. Read whether the question asks for atoms or molecules. One NH_3 molecule contains 1+3=4 atoms; one H_2O molecule contains 2+1=3. If only one element's atoms are required, multiply only by that element's atom count in each molecule.



The three sample quantities refer to different objects. The Avogadro constant first gives the number of the specified molecules.

Large numbers are often written in **standard form**:

$$a \times 10^k \quad (1 \le a < 10, \quad k \text{ is an integer})$$

For example, the leading number in 0.301×10²³ is below one, so rewrite it as:

$$0.301 \times 10^{23} = 3.01 \times 10^{22}$$

Multiplying the leading number by ten requires reducing the exponent by one to keep the same value. In multiplication, handle the ordinary numbers and powers of ten, then put the result into standard form. Round according to the supplied data only at the end; do not truncate the amount in mol prematurely.

Recognising the Question Identify whether the input is mass in g, amount in mol, gas volume or particle number. Convert through mol as needed. Finally check whether molecules, one element's atoms or all atoms are requested, and verify cm³/dm³, standard form and units.

Worked Example 1 (2025·October/November 0620/41 Q1(e)) · Level 3

(e) Calculate the number of atoms in 2.00 g of argon.

Give your answer in standard form.

Solution

Argon, Ar, consists of single atoms. Its Periodic Table value A_r(Ar)=40 corresponds to molar mass 40g/mol. First convert 2.00g to amount of substance:

$$n(\text{Ar}) = \frac{2.00}{40} = 0.0500\,\text{mol}$$

Every mole of Ar contains 6.02×10²³ Ar atoms, so:

$$N(\text{Ar}) = 0.0500 \times 6.02 \times 10^{23} = 0.301 \times 10^{23}$$

Standard form requires a leading number of at least one and less than ten. Multiply 0.301 by ten and reduce the exponent from 23 to 22:

$$N(\text{Ar}) = 3.01 \times 10^{22}$$

The answer is **3.01×10²² atoms**. Ar is monatomic, so unlike H_2 or NH_3, no further multiplication by an atom count per molecule is needed.

Practice

1. (2024·October/November 0620/43 Q4(f)(ii)) · Level 3

(ii) In another experiment, 48.0 cm^3 of hydrogen gas, H_2, is produced. The experiment is carried out at room temperature and pressure, r.t.p.

Calculate the number of molecules in 48.0 cm^3 of H_2 gas measured at r.t.p.

The value of the Avogadro constant is 6.02×10^{23}.

2. (2025·February/March 0620/22 Q12) · Level 3

The value of the Avogadro constant is 6.02×10^{23}.

What is the total number of atoms in 2.00 mol of ammonia gas?

A 1.20×10^{24} **B** 2.41×10^{24} **C** 4.82×10^{24} **D** 2.89×10^{25}

Type 2: Using mole ratios and limiting reactants to calculate mass or gas volume

Key Ideas The previous question type connected mass, mol, gas volume and particle number. Reaction calculations add one step: **converting the amount of one substance in mol into the amount of another**. Balanced coefficients give particle-number ratios. Multiplying all particle numbers by the same Avogadro constant also gives **mole ratios**.

Balanced coefficients also give mole ratios; mass ratios require molar masses

$$N_2 + 3H_2 \rightarrow 2NH_3$$

N_2	H_2	NH_3
1mol	3mol	2mol
28g	6g	34g

Coefficients 1 : 3 : 2 give the mole ratio 1 : 3 : 2, but the mass ratio is 28 : 6 : 34.

The reaction is supplied. Multiplying all specified particle numbers by the same Avogadro constant leaves their ratio unchanged.

For example, the equation for ammonia synthesis is given as:

$N_2 + 3H_2 \rightarrow 2NH_3$

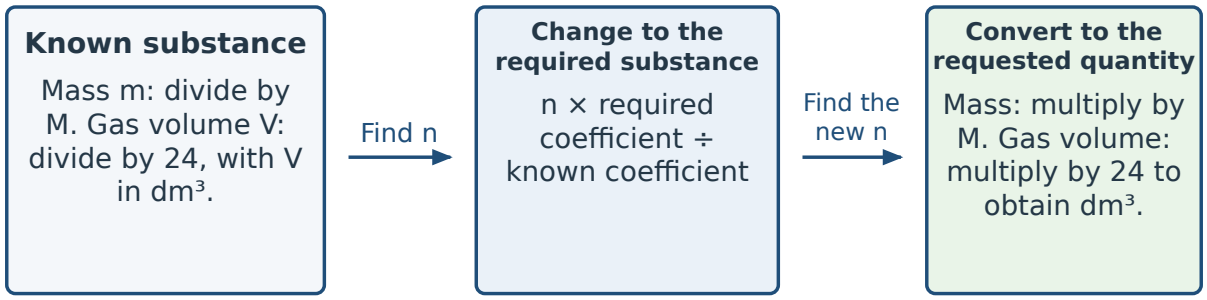
This means that 1mol of N_2 reacts exactly with 3mol of H_2 to form 2mol of NH_3. Complete reaction of 2mol of N_2 needs 6mol of H_2 and produces 4mol of NH_3. Scaling the equation by two scales the amounts of all three substances together.

To convert from the known substance to the required substance, use their equation coefficients to determine the mole ratio:

$$n_{\text{required}} = n_{\text{known}} \times \frac{\text{coefficient of required substance}}{\text{coefficient of known substance}}$$

The two coefficients must correspond to the two substances involved. If both are two, the ratio 2:2 is 1:1. Do not multiply a known amount by two simply because the product has coefficient two.

Convert the given quantity to moles before using the equation to change substance



The gas-volume relationship requires r.t.p.
For volumes in cm³, replace 24 with 24 000.

The coefficient ratio applies to amounts in mol. It cannot directly convert grams of one substance into dm³ of another.

For example, 4.8g of Mg reacts with sufficient oxygen according to:

$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

First find the amount of Mg. Its molar mass is 24g/mol:

$$n(\text{Mg}) = \frac{4.8}{24} = 0.20\,\text{mol}$$

Mg and MgO both have coefficient two, so the amount of MgO formed is:

$$n(\text{MgO}) = 0.20 \times \frac{2}{2} = 0.20\,\text{mol}$$

MgO has molar mass 40g/mol, so the product mass is:

$$m(\text{MgO}) = 0.20 \times 40 = 8.0\,\text{g}$$

If oxygen required is asked instead, use the O_2:Mg coefficient ratio of 1:2:

$$n(\text{O}_2) = 0.20 \times \frac{1}{2} = 0.10\,\text{mol}$$

At r.t.p., the corresponding oxygen volume is:

$$V(\text{O}_2) = 0.10 \times 24 = 2.4\,\text{dm}^3$$

Label the substance at each step to avoid confusing the product mass with the oxygen volume.

A reactant in **excess** is present in more than the amount required for the reaction, allowing the other reactant to be used up. If an acid is in excess, use the other known reactant to determine the yield. Excess does not mean infinitely large mass or amount; it means that the acid will not run out first in this calculation.

If quantities of both reactants are supplied, find the **limiting reactant**: the reactant used up first according to the reaction ratio, thereby limiting the total product. The reactant with the smaller initial mass or smaller initial amount in mol is not necessarily limiting; its coefficient also matters.

Calculate the following ratio for each reactant and compare:

$$\frac{\text{amount of reactant in mol}}{\text{coefficient of that reactant}}$$

The smaller ratio supports less reaction and identifies the reactant that runs out first.

When both reactant amounts are given, compare how much reaction each can support

$2H_2 + O_2 \rightarrow 2H_2O$; available: 5 mol H_2 and 2 mol O_2



O_2 gives the smaller value, 2, so it runs out first. 4 mol H_2 reacts to form 4 mol H_2O.

H_2 remaining = 5 − 4 = 1 mol. Adding more of the already excess H_2 produces no further product.

Each small box represents 1mol of the corresponding gas. Compare n divided by each reactant's coefficient, not only the original amounts in mol.

The diagram has 5mol of H_2 and 2mol of O_2, reacting according to:

$$2H_2 + O_2 \rightarrow 2H_2O$$

Divide each reactant's amount by its own coefficient:

$$H_2: \frac{5}{2} = 2.5 \quad ; O_2: \frac{2}{1} = 2$$

O_2 gives the smaller ratio, so it is limiting. Using all 2mol of O_2 consumes 4mol of H_2 and produces 4mol of H_2O. Hydrogen remaining is:

$$n_{\text{remaining}}(H_2) = 5 - 4 = 1\,\text{mol}$$

Saying only that two is smaller than five has not considered coefficients. In another reaction with different coefficients, comparing the original amounts directly could give the wrong conclusion.

After finding the product amount, calculate how much of the other reactant is required and check that it does not exceed the amount supplied. Find the excess reactant remaining using the same unit for both terms:

$$\text{amount remaining} = \text{initial amount} - \text{amount actually consumed}$$

Once the limiting reactant is used up, adding more of the already excess reactant does not increase the product. The available stoichiometric reaction has finished; no limiting reactant remains waiting to react.

Gas volume ratios can sometimes be read directly from coefficients, but two conditions must both hold: **all the substances compared are gases, and their temperatures and pressures are the same**. Then V is proportional to n, so the common molar-volume factor cancels and leaves the coefficient ratio.

Only gases at the same temperature and pressure have volume ratios matching equation coefficients

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Mole ratio: 1 : 3 : 2

r.t.p. volume ratio: 24 : 72 : 48 (dm^3)

Do not apply this volume ratio to solids, liquids or gases at different temperatures or pressures.

When all gases use the same 24dm³/mol, that common factor cancels and leaves the coefficient ratio.

Do not scale numbers directly by coefficients if one quantity is a solid mass and the other a gas volume, or if gases are at different temperatures and pressures. The reliable route is to convert each given quantity correctly to mol, then use the equation to reach the target substance. In particular, liquid $H_2O(l)$ does not have molar volume 24dm³/mol. If another gas is requested, calculate that gas alone instead of mixing in every product's subscripts or coefficients.

Recognising the Question Convert the known quantities to mol. If both reactant amounts are supplied, compare n divided by coefficient to find the limiting reactant. Use the appropriate coefficient ratio to find the target amount, then convert to mass or gas volume at r.t.p. Finally check units and any amount remaining.

Worked Example 1 (2025·May/June 0620/41 Q4(d)) · Level 3

(d) Ammonia is converted into ammonium sulfate, $(NH_4)_2SO_4$.

$$2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(s)$$

360 dm^3 of ammonia gas, measured at r.t.p., reacts with excess sulfuric acid.

Calculate the mass of ammonium sulfate produced, using the following steps.

- Calculate the number of moles of $NH_3(g)$ in 360 dm^3.

  One mole of any gas occupies 24 dm^3 at r.t.p.
- Calculate the number of moles of $(NH_4)_2SO_4$ produced.
- Calculate the mass of $(NH_4)_2SO_4$ produced.

  The M_r of $(NH_4)_2SO_4$ is 132.

Solution

First find the amount of NH_3. At r.t.p., one mole of gas occupies 24dm³. The given NH_3 volume is 360dm³:

$$n(NH_3) = \frac{360}{24} = 15\,\text{mol}$$

Next use the reaction ratio. The supplied reaction is:

$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$

Sulfuric acid is in excess, so all 15mol of NH_3 can react. Every 2mol of NH_3 produces 1mol of ammonium sulfate:

$$n((NH_4)_2SO_4) = 15 \times \frac{1}{2} = 7.5\,\text{mol}$$

Finally find the product mass. The given M_r of ammonium sulfate is 132, corresponding to molar mass 132g/mol:

$$m((NH_4)_2SO_4) = 7.5 \times 132 = 990\,\text{g}$$

The three requested answers are **15mol, 7.5mol and 990g**. The gas value 24dm³/mol cannot be used to calculate the solid product's volume.

Worked Example 2 (2025·May/June 0620/42 Q4(c)(iii)) · Level 3

(iii) When $(NH_4)_2Cr_2O_7$ is heated the following reaction occurs.

$$(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + 4H_2O + N_2$$

Calculate the volume of nitrogen gas produced at r.t.p., in cm^3, when 1.26 g of $(NH_4)_2Cr_2O_7$ is heated using the following steps.

The M_r of $(NH_4)_2Cr_2O_7$ is 252.

- Calculate the number of moles of $(NH_4)_2Cr_2O_7$ used.
- Determine the number of moles of N_2 formed.
- Calculate the volume of N_2 formed at r.t.p. in cm^3.

Solution

First, $(NH_4)_2Cr_2O_7$ has the given M_r of 252 and hence molar mass 252g/mol. The amount used is:

$$n((NH_4)_2Cr_2O_7) = \frac{1.26}{252} = 0.00500\,\text{mol}$$

Second, read the coefficients in the supplied reaction:

$(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + 4H_2O + N_2$

The reactant and N_2 both have coefficient one, so **$n(N_2)$=0.00500mol**. The coefficient 4 belongs to H_2O and must not multiply the amount of N_2.

Third, nitrogen volume is requested in cm^3 at r.t.p., so use 24000cm^3/mol:

$$V(N_2) = 0.00500 \times 24000 = 120\,cm^3$$

Using 24dm^3/mol first would give 0.120dm^3. Multiplying by 1000 still gives **120cm^3**. The final unit must match the question.

Practice

1. (2025·May/June 0620/22 Q7) · Level 2

When 65 g of zinc reacts with 32 g of sulfur, 97 g of zinc sulfide is produced.

When 65 g of zinc reacts with 40 g of sulfur, the mass of zinc sulfide produced is still 97 g.

Which statement explains this observation?

A Some of the zinc sulfide evaporates.

B The reaction rate is slow.

C The reaction stops before it is complete.

D Zinc is the limiting reactant.

2. (2025·May/June 0620/43 Q4(d)) · Level 3

(d) Dilute sulfuric acid reacts with sodium hydrogencarbonate, $NaHCO_3$, to produce carbon dioxide gas, CO_2.

$$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$$

4.20 g of $NaHCO_3$ is added to excess dilute sulfuric acid.

Calculate the volume of $CO_2(g)$, measured at r.t.p., produced using the following steps.

- Calculate the number of moles of $NaHCO_3$ in 4.20 g.

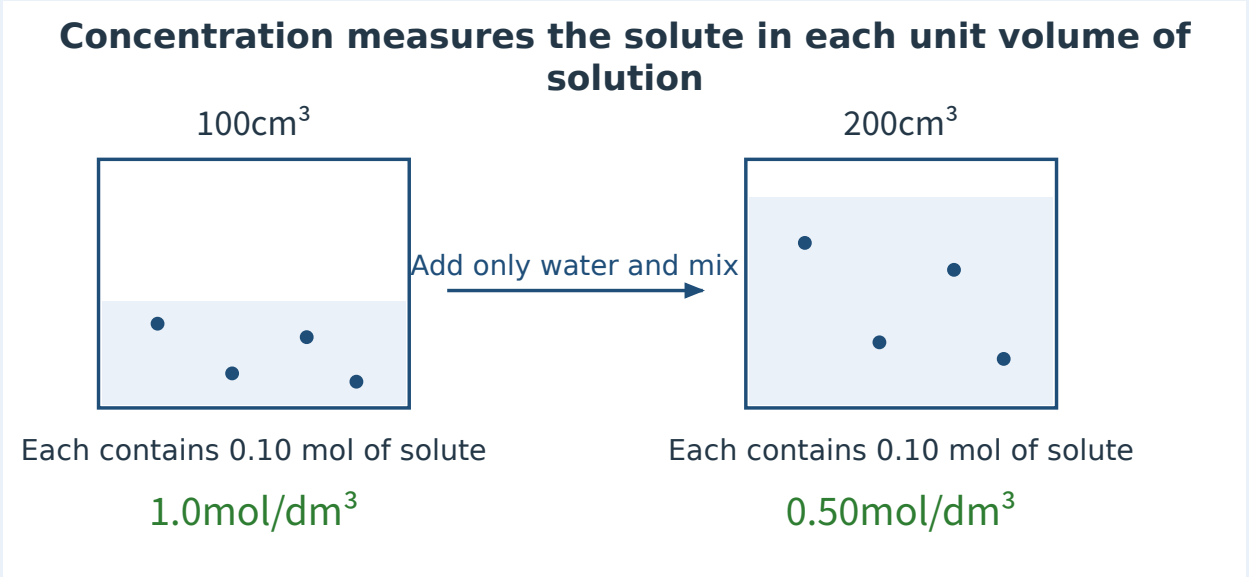
  The M_r of $NaHCO_3$ is 84.
- Deduce the number of moles of $CO_2(g)$ that are produced.
- Calculate the volume of $CO_2(g)$ produced in cm^3.

  One mole of any gas occupies 24 000 cm^3 at r.t.p.

Type 3: Combining concentration, volume and reaction ratios in solution and titration calculations

Key Ideas A **solute** is the substance being dissolved, and a **solvent** is the medium that dissolves it. Together they form a **solution**. For example, in salt solution, NaCl is the solute and water the solvent. The volume used for concentration is the **final volume of the whole solution**, not just the volume of water added.

Concentration describes the amount of solute per unit volume of solution. Diluting the same amount of solute with water leaves its amount in mol unchanged but spreads it through a larger volume, so each unit volume contains less. Concentration and the total amount of solute in a sample are different quantities. A higher concentration alone does not establish that a whole bottle contains more solute.



Blue dots represent the same amount of solute. Dilution leaves the solute amount in mol unchanged while increasing the final solution volume.

Two concentration units are used in this course. **g/dm³** gives the grams of solute per 1dm³ of solution:

$$\text{mass concentration} = \frac{m}{V}$$

Here m is in g and V in dm³. **mol/dm³** gives the amount of solute per 1dm³ of solution, so molar concentration is:

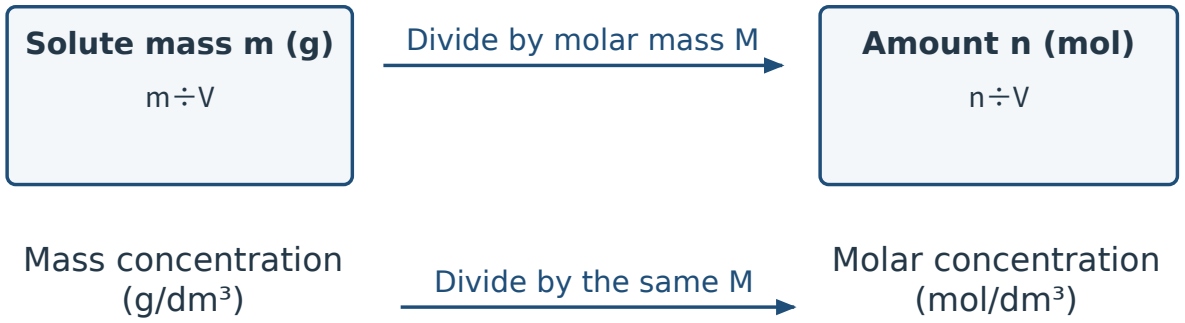
$$c = \frac{n}{V}$$

To find amount from concentration and volume, or volume from amount and concentration, rearrange respectively:

$$n = cV \quad ; \quad V = \frac{n}{c}$$

In these equations n is in mol, c in mol/dm³ and V must be in dm³. Divide a volume in cm³ by 1000 first.

For the same solution, g/dm³ counts mass and mol/dm³ counts amount of substance



V is the volume of the whole solution in dm³, not just the volume of water added.

For the same solution, mass concentration equals molar concentration × molar mass. Units in the division can also help check the correct direction.

For example, take 250cm³ of NaOH solution at 0.200mol/dm³. Convert the volume to 0.250dm³, then calculate the amount of solute:

$$n(\mathrm{NaOH}) = cV = 0.200 \times 0.250 = 0.0500\,\mathrm{mol}$$

NaOH has molar mass 40g/mol, from the relative-mass sum for Na, O and H: 23+16+1. The solute mass is:

$$m(\mathrm{NaOH}) = 0.0500 \times 40 = 2.00\,\mathrm{g}$$

This 2.00g of solute occupies 0.250dm³ of solution, so its mass concentration is:

$$\text{mass concentration} = \frac{2.00}{0.250} = 8.00\,\mathrm{g/dm^3}$$

Alternatively, convert directly from molar concentration:

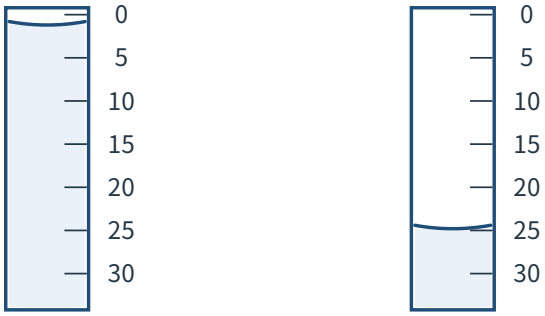
$$\text{mass concentration} = cM = 0.200 \times 40 = 8.00\,\mathrm{g/dm^3}$$

Both concentrations describe the same solution. Multiplying by molar mass simply converts moles per unit volume into grams per unit volume.

Titration uses a solution of known concentration to measure the volume needed to react exactly with another solution, allowing an unknown quantity to be calculated. In **neutralisation**, H^+ supplied by the acid reacts with OH^- supplied by the alkali to form H_2O. However, a formula unit or molecule of acid or alkali may provide more than one H^+ or OH^-. Balance the full equation first, then use its coefficients for the mole ratio.

In practice, two burette readings are often taken. The scale increases downwards because it measures volume delivered. **Final reading minus initial reading** gives the volume actually added. Read the bottom of a concave meniscus at eye level to avoid the error caused by looking from above or below.

Burette readings increase downwards: volume used = final reading − initial reading



Read the bottom of a concave meniscus at eye level. Use the same scale convention for both readings.

The two readings are labelled; fine graduations are omitted. As the liquid level falls, the reading increases. Their difference is the volume delivered.

Convert the diagram's 23.60cm³ to 0.02360dm³ before multiplying by a concentration in mol/dm³. If a required volume or titre is given directly, it already is the volume used; do not invent another subtraction. If several reliable, closely agreeing titres are supplied and an average is requested, calculate the mean from the appropriate trials before continuing. Use the question's instructions to decide whether rough or clearly inconsistent results should be included.

For example, 20.0cm³ of 0.150mol/dm³ H_2SO_4 reacts exactly with 0.200mol/dm³ NaOH solution. To find the required NaOH volume, begin with the balanced equation:

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

Convert the acid volume to 0.0200dm³, then multiply concentration by volume to find its amount:

$$n(H_2SO_4) = 0.150 \times 0.0200 = 0.00300\,\text{mol}$$

The acid:NaOH coefficient ratio is 1:2, so the amount needed is:

$$n(NaOH) = 2 \times 0.00300 = 0.00600\,\text{mol}$$

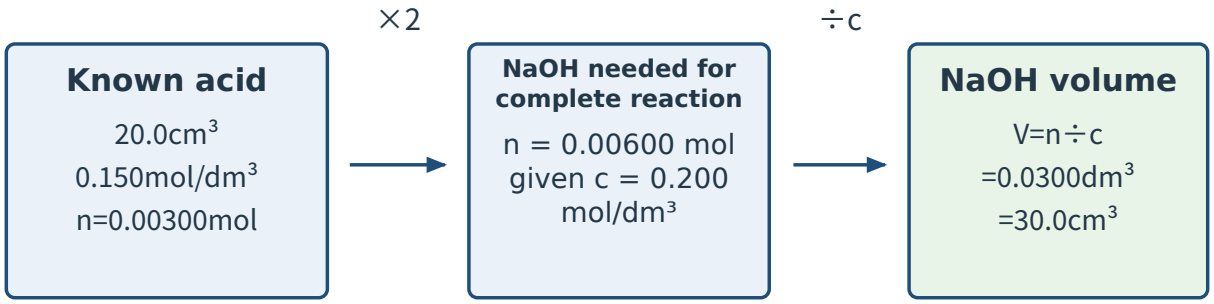
Finally divide the amount of NaOH by its own concentration:

$$V(NaOH) = \frac{n}{c} = \frac{0.00600}{0.200} = 0.0300\,\text{dm}^3$$

Converting back to the commonly requested unit gives 30.0cm³ of NaOH solution.

At the titration endpoint: calculate moles, apply the ratio, then find the required quantity

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$



The coefficients give a 1 : 2 mole ratio, not necessarily a 1 : 2 volume ratio. Concentrations also matter.

This route calculates a titration volume. If concentration is requested instead, use c=n÷V at the final step; the earlier mole ratio is unchanged.

For neutralisation of $Ba(OH)_2$ with HNO_3, one $Ba(OH)_2$ provides two OH^-, while each HNO_3 provides one H^+, so nitric acid needs coefficient two.

Continue using known concentration and volume to calculate mol, then coefficients to convert to the other substance's mol. If concentration is required, divide by that substance's own solution volume at the end. To find a solution's concentration before mixing, do not add the two liquid volumes and use their sum as its original volume.

If a solid mass and an acid solution's concentration and volume are both supplied, do not assume that either is automatically in excess. Calculate each amount first:

$$n_{\text{solid}} = \frac{m}{M} \quad ; \quad n_{\text{solute}} = cV$$

Then divide each amount by its own equation coefficient and compare to find the limiting reactant. Use that reactant to calculate a gas or other product. The two inputs use different routes to mol, but the same mole-ratio rules apply within the equation.

Recognising the Question Identify whether the unknown is solute mass, amount in mol, concentration or solution volume. Convert cm^3 to dm^3. For titration, find the actual volume used, calculate mol with cV, then use the balanced equation to find the other substance. If both reactants are quantified, still check which is limiting.

Worked Example 1 (2025·May/June 0620/23 Q9) · Level 3

Aqueous barium hydroxide, $Ba(OH)_2$, reacts with dilute nitric acid, HNO_3. The products are aqueous barium nitrate, $Ba(NO_3)_2$, and water.

In a titration, $15.2\,cm^3$ of $0.75\,mol/dm^3$ aqueous barium hydroxide is required to neutralise $20.0\,cm^3$ of dilute nitric acid.

What is the concentration of the nitric acid?

A $0.29\,mol/dm^3$

B $0.57\,mol/dm^3$

C $1.14\,mol/dm^3$

D $2.03\,mol/dm^3$

Solution

First balance the given neutralisation reaction:

$$Ba(OH)_2 + 2HNO_3 \rightarrow Ba(NO_3)_2 + 2H_2O$$

One $Ba(OH)_2$ supplies two OH^- and needs two HNO_3 for neutralisation, giving a 1:2 mole ratio. Convert $15.2cm^3$ to $0.0152dm^3$, then find the amount of barium hydroxide:

$$n(Ba(OH)_2) = 0.75 \times 0.0152 = 0.0114\,mol$$

The nitric acid required for exact neutralisation is:

$$n(HNO_3) = 2 \times 0.0114 = 0.0228\,mol$$

This amount of nitric acid is in $20.0cm^3$, or $0.0200dm^3$, of solution. Its concentration is:

$$c(HNO_3) = \frac{0.0228}{0.0200} = 1.14\,mol/dm^3$$

Choose C. The denominator is the nitric acid solution's own volume, not the total volume after mixing. Omitting coefficient two would incorrectly give $0.57mol/dm^3$.

Practice

1. (2025·May/June 0620/22 Q9) · Level 3

The concentration of a solution of aqueous sodium hydroxide is $0.50\,mol/dm^3$.

Which mass of sodium hydroxide is used to make $500\,cm^3$ of this solution?

A 10 g **B** 20 g **C** 40 g **D** 160 g

2. (2025·May/June 0620/21 Q9) · Level 3

Calcium carbonate reacts with dilute hydrochloric acid.

The equation for the reaction is shown.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

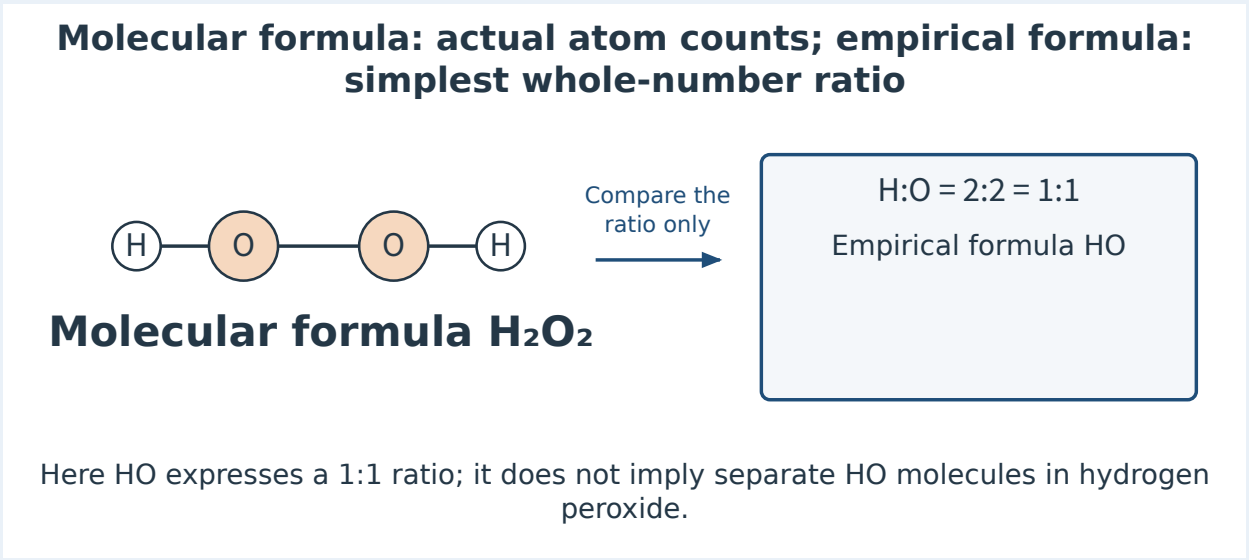
1.00 g of calcium carbonate is added to $50.0\,cm^3$ of $0.0500\,mol/dm^3$ hydrochloric acid.

Which volume of carbon dioxide is made in this reaction?

A $30\,cm^3$ **B** $60\,cm^3$ **C** $120\,cm^3$ **D** $240\,cm^3$

Type 4: Finding empirical formulae from composition and then determining molecular formulae

Key Ideas An **empirical formula** gives the simplest whole-number ratio of different atoms or ions in a compound. A **molecular formula** gives the actual number of each type of atom in one molecule. The two formulae may be identical or differ by a whole-number multiplier. An empirical formula alone does not establish how large a molecule is.

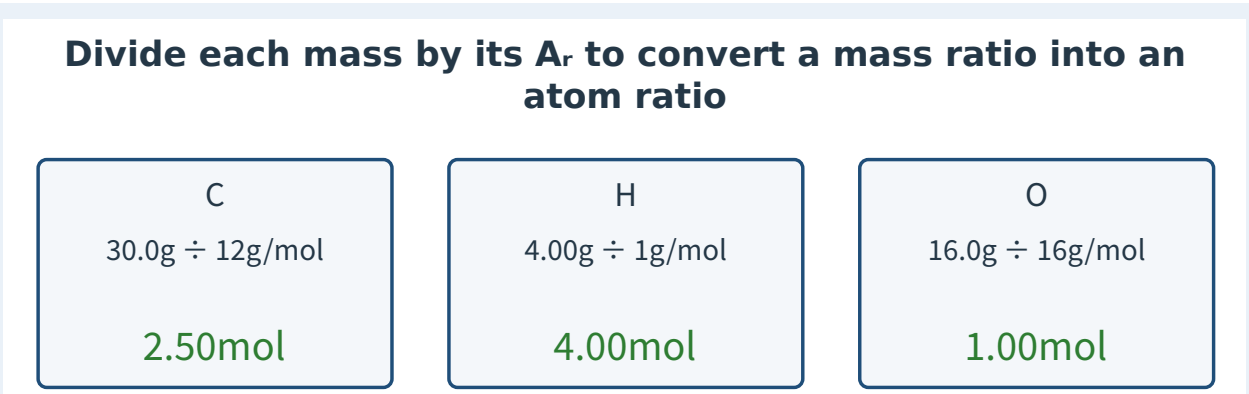


Count all atoms in one actual H_2O_2 molecule, then simplify its H:O ratio. The two formulae serve different purposes.

For example, ethane, C_2H_6, has atom ratio 2:6, simplifying to 1:3 and empirical formula CH_3. This does not mean an ethane molecule contains only one C and three H atoms. Water's ratio 2:1 is already simplest, so both its molecular and empirical formulae are H_2O. Ionic NaCl expresses the simplest 1:1 ratio of Na^+ to Cl^- in its lattice; there is no separate NaCl molecule whose size must be found.

When a molecular formula is given, count every element, then divide all subscripts by their highest common factor. In $C_6H_{12}O_6$, divide 6:12:6 by six to obtain 1:2:1 and CH_2O. In $C_3H_6O_2$, C and H are divisible by three but O is not. There is no factor common to all three, so the empirical formula remains $C_3H_6O_2$. Do not simplify different pairs independently.

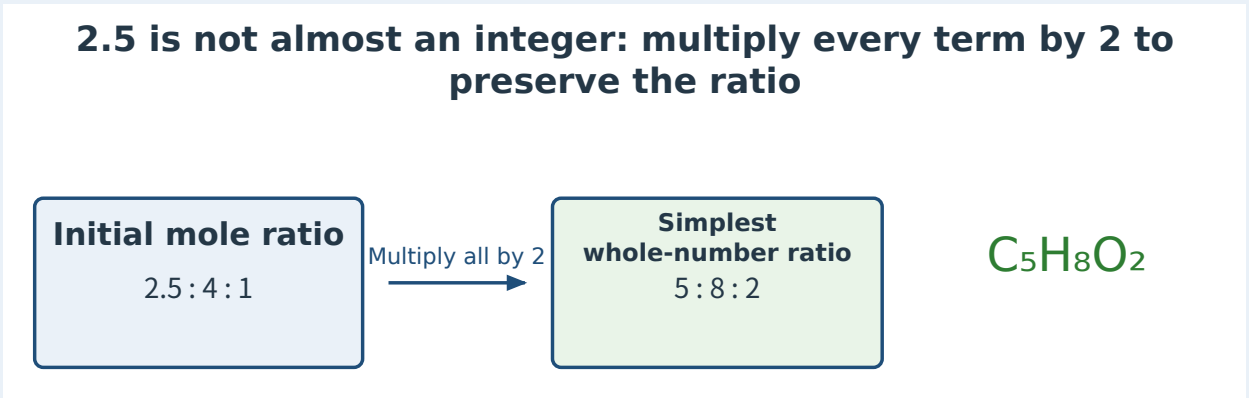
When composition masses are given, convert the mass ratio into an atom-number ratio. Equal masses of C and H contain different atom numbers because an average C atom is about 12 times as heavy as an H atom. To find atom-number proportions, divide each element's mass by the molar mass of its atoms: numerically, by its A_r.



The sample masses of C, H and O are given. Divide each by the molar mass of its atoms first.

For percentage masses, **assume a 100g sample** and turn each percentage into grams. This does not claim that the original experiment used 100g; scaling a sample does not change its elemental proportions. For example, 40.0%C, 6.7%H and 53.3%O can be treated as 40.0g, 6.7g and 53.3g. If one element's mass must be found by subtracting the others from the total, complete that step before dividing by A_r.

Next, divide **all** the calculated element amounts by the smallest one, giving a ratio with one as its smallest term. Values very close to integers such as one, two or three may differ slightly because of data precision. However, 2.5 is a clear half-integer and must not be rounded arbitrarily to three. To retain the composition and make every term an integer, multiply all terms by two. For fractions such as 1/3 or 2/3, consider multiplying the whole ratio by three.



The calculated ratio 2.5:4:1 from the previous diagram is used. Multiplying every term by two preserves all proportions.

For example, a sample contains 30.0g C, 4.00g H and 16.0g O. Divide each mass by the molar mass of its atoms:

$$n(\mathrm{C}) = \frac{30.0}{12} = 2.50\,\mathrm{mol}$$

$$n(\mathrm{H}) = \frac{4.00}{1} = 4.00\,\mathrm{mol}$$

$$n(\mathrm{O}) = \frac{16.0}{16} = 1.00\,\mathrm{mol}$$

The amount ratio is 2.50:4.00:1.00. Dividing by the smallest value, 1.00, leaves 2.5:4:1. Multiply the whole ratio by two to make every term an integer:

$$\mathrm{C{:}H{:}O} = 2.5 : 4 : 1 = 5 : 8 : 2$$

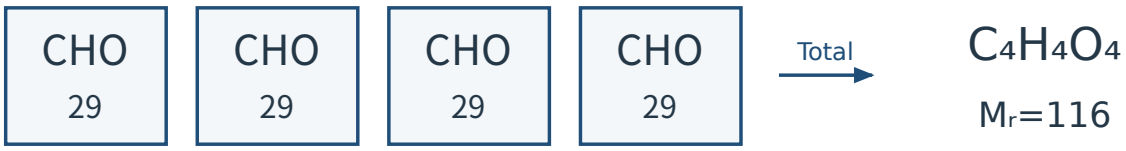
There is no common factor, so the empirical formula is $C_5H_8O_2$. Changing only 2.5 to three would change C:O to 3:1 and no longer match the data.

Finding a molecular formula also requires the actual relative molecular mass. Calculate the empirical formula mass first, then find the whole-number multiplier:

$$k = \frac{\text{actual molecular } M_r}{\text{empirical formula mass}}$$

Multiply every subscript in the empirical formula by k. The multiplier tells you how many empirical-ratio units occur in one actual molecule; do not simply append it to the end of the formula.

Empirical to molecular formula: actual M_r ÷ empirical formula mass = multiplier



$C_4H_4O_4$

M_r=116

116 ÷ 29 = 4; multiply the subscripts of C, H and O by 4.

Each box is a ratio unit for counting atoms, not a separate CHO molecule.

The actual molecular M_r is supplied here. An empirical formula alone cannot determine the multiplier.

For example, the empirical formula is CH_2O. Add the relative masses of the atoms in this simplest ratio:

$$\text{empirical formula mass} = 12 + 2 \times 1 + 16 = 30$$

If the actual molecular M_r is 90, the multiplier is:

$$k = \frac{90}{30} = 3$$

Multiply all three subscripts in CH_2O by three to obtain $C_3H_6O_3$. Check the relative molecular mass:

$$M_{r(\mathrm{C_3H_6O_3})} = 3 \times 12 + 6 \times 1 + 3 \times 16 = 90$$

The atom ratio 3:6:3 still simplifies to 1:2:1. If only CH_2O is given, without M_r or other sufficient information, a unique molecular formula cannot be determined.

Some questions supply a condensed structural formula, such as $CH_3CH(OH)CH_3$. Atoms of the same element appear in different positions, so expand each group and add the totals. This step uses the supplied symbols only; prior knowledge of organic names or reactions is unnecessary.

An element may appear in several parts of a condensed formula: count each part, then add

CH_3	$CH(OH)$	CH_3
C=1, H=3	C=1, H=2, O=1	C=1, H=3

Totals: C = 3, H = 3 + 2 + 3 = 8, O = 1 → C_3H_8O

The central CH(OH) contributes two H atoms: one in CH and one in OH.

The full condensed formula is supplied. Only element counting is needed here, not prior knowledge of organic naming or reactions.

The left CH_3 contributes one C and three H. The central CH(OH) contributes one C and one H, plus one O and one H from OH. The right CH_3 contributes another C and three H, giving C_3H_8O overall. H in OH or COOH is part of the molecule too; do not count only places labelled H_2 or H_3. To compare empirical formulae across several molecules, independently count all atoms and simplify the whole ratio for each. Compare complete empirical formulae, not just the count of one element.

Recognising the Question For mass composition, divide by A_r to obtain amount ratios, then divide all terms by the smallest. Convert fractional ratios to whole numbers together before writing the empirical formula. A molecular formula also needs actual M_r to determine the multiplier for the entire formula. For condensed structures, count each element across all groups first.

Worked Example 1 (2025·May/June 0620/41 Q5(e)) · Level 3

(e) An organic compound has the following composition by mass:

C, 64.87%; H, 13.51%; O, 21.62%.

Calculate the empirical formula of the compound.

Solution

The percentages describe mass, so 64.87, 13.51 and 21.62 cannot be used directly as atom subscripts. Assume 100g of sample: C, H and O then have masses 64.87g, 13.51g and 21.62g. This makes ratio conversion convenient without changing the composition.

Element	C	H	O
Mass/g	64.87	13.51	21.62
Mass÷A_r	64.87÷12	13.51÷1	21.62÷16
Amount in mol	5.4058…	13.51	1.35125
Divide all by 1.35125	4.000…	9.998…	1

Within the precision of the data, the atom ratio is 4:10:1. No whole-number factor can reduce all three terms, so the empirical formula is $\mathbf{C_4H_{10}O}$. The one O has no written subscript. Do not simplify H:O independently and disrupt the overall C:H:O ratio.

Worked Example 2 (2025·October/November 0620/43 Q6(b)) · Level 2

(b) Organic compound **R** has the empirical formula CHO and a relative molecular mass of 116.

Determine the molecular formula of compound **R**.

Solution

The empirical formula CHO means C:H:O=1:1:1. Calculate the relative mass of this simplest formula:

$$\text{empirical formula mass} = 12 + 1 + 16 = 29$$

The actual molecular M_r is 116, so the multiplier is:

$$k = \frac{116}{29} = 4$$

Multiply all three implicit subscripts of one in CHO by four to obtain the **molecular formula $C_4H_4O_4$**. Check its relative molecular mass:

$$M_{r(C_4H_4O_4)} = 4 \times 12 + 4 \times 1 + 4 \times 16 = 116$$

The ratio 4:4:4 still simplifies to 1:1:1. Apply the multiplier to every element, not just one.

Practice

1. (2024·May/June 0620/41 Q7(a)(ii)) · Level 2

(ii) Compound **B** has the empirical formula CH_2 and a relative molecular mass of 70.

Determine the molecular formula of compound **B**.

2. (2025·May/June 0620/21 Q8) · Level 3

Three organic molecules are described.

- The molecular formula of propanoic acid is CH_3CH_2COOH.
- The molecular formula of lactic acid is $CH_3CH(OH)COOH$.
- A glucose molecule consists of 6 carbon atoms, 6 oxygen atoms and 12 hydrogen atoms only.

Which molecules have the same empirical formula?

A propanoic acid, lactic acid and glucose

B propanoic acid and lactic acid only

C propanoic acid and glucose only

D lactic acid and glucose only

3. (2025·May/June 0620/43 Q5(e)) · Level 3

(e) An organic compound has the following composition by mass:

C, 58.82%; H, 9.80%; O, 31.38%.

Calculate the empirical formula of the compound.

4. (2025·October/November 0620/22 Q11) · Level 3

A compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen only.

The relative molecular mass, M_r, of the compound is 60.

Which row shows the empirical formula and the molecular formula of the compound?

	empirical formula	molecular formula
A	CH_2O	CH_2O
B	CH_2O	$C_2H_4O_2$
C	$C_3H_7O_3$	$C_3H_7O_3$
D	C_3H_8O	C_3H_8O

Type 5: Calculating percentage composition, purity and yield with the correct quantities

Key Ideas

A **percentage** describes how many parts out of every 100 parts of the whole are present. Divide the part by the corresponding whole, then multiply by 100%. For example, 2g as a percentage of 10g is:

$$\frac{2}{10} \times 100\% = 20\%$$

Matching mass units cancel in the division, so the percentage itself has no unit of g. If one mass is in kg and the other in g, convert them to the same unit first.

Three common chemical percentages look similar but use different parts and wholes. **Percentage composition by mass** compares an element's mass contribution with the mass of the whole compound. **Percentage purity** compares the target substance with the whole sample, including impurities. **Percentage yield** compares the product actually obtained with the maximum theoretical amount from the equation. Identify the denominator before calculating.

Identify the part and the whole: these three percentage denominators differ

Percentage by mass of an element

$$\frac{\text{Mass contributed by that element}}{\text{Mass of the whole compound}}$$

Then multiply by 100%

Percentage purity

$$\frac{\text{Mass of the required substance}}{\text{Mass of the whole impure sample}}$$

Then multiply by 100%

Percentage yield

$$\frac{\text{Actual mass of product collected}}{\text{Theoretical mass of product}}$$

Then multiply by 100%

State which whole contains the required part, then substitute masses in the same units.

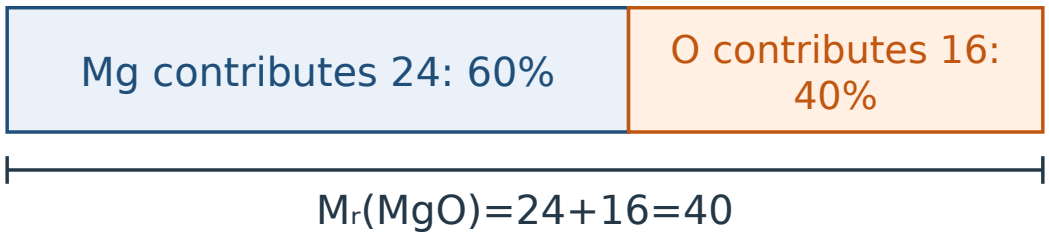
All three percentages involve division, but compare different quantities. The numerator and denominator of each fraction use the same mass unit.

1. Percentage composition from a formula. The formula gives atom numbers and A_r gives each atom type's relative mass. An element's total contribution is its atom count multiplied by A_r. The entire formula contributes M_r:

$$\text{percentage by mass of an element} = \frac{\text{total mass contribution of that element}}{M_r} \times 100\%$$

These are relative masses, so do not add g. Do not confuse atom-number ratios with mass ratios.

Mass contributions in MgO: 24 from Mg + 16 from O = 40



These are elemental mass contributions within pure MgO, not separate portions of magnesium and oxygen.

The bar is divided exactly in the mass-contribution ratio 24:16. It does not imply that MgO is a mixture of two elements.

For example, Mg and O have A_r values of 24 and 16 in MgO:

$$M_{r(\text{MgO})} = 24 + 16 = 40$$

Magnesium contributes 24 and oxygen 16 to each formula mass. Divide each by the total, 40:

$$\text{percentage by mass of magnesium} = \frac{24}{40} \times 100\% = 60\%$$

$$\text{percentage by mass of oxygen} = \frac{16}{40} \times 100\% = 40\%$$

The percentages total 100%. Although the atom ratio Mg:O is 1:1, each does not contribute half the mass because the atoms have different masses.

If the whole compound's mass is given, find the element's mass fraction first, then multiply by sample mass. For example, the mass of magnesium in 200g of pure MgO is:

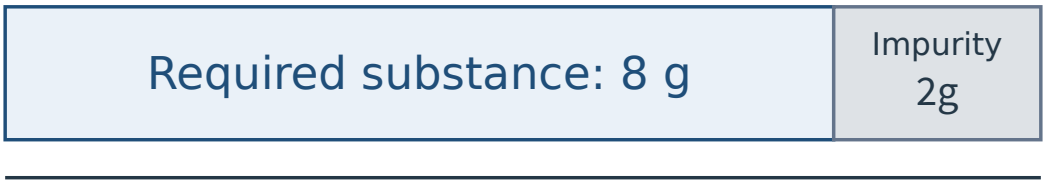
$$m(\text{magnesium}) = 200 \times \frac{24}{40} = 120\,\text{g}$$

This is the mass of magnesium chemically combined in MgO, not a separate 120g of magnesium powder mixed into the solid. If several atoms of an element occur in a formula, include their number first. For example, iron contributes 2×56 in Fe_2O_3, not just one 56.

2. Sample purity. A pure compound can itself contain several elements. When it is mixed with other substances, purity asks what fraction of the sample is the whole target compound. For an impure salt, use the target salt's mass, not the mass of one element within that salt:

$$\text{purity} = \frac{\text{mass of target substance}}{\text{total mass of original impure sample}} \times 100\%$$

Purity: count the entire required compound separately from the impurities



Whole sample: 8 + 2 = 10 g; purity = 8 ÷ 10 × 100% = 80%

The 8 g is the mass of the whole required substance, not just one element within it.

The bar follows the sample's actual ratio of 8g target substance to 2g impurities. The target substance can itself be a pure compound.

For example, a 10.0g sample contains 8.00g of the target salt and 2.00g of impurities. Its purity is:

$$\text{purity} = \frac{8.00}{10.0} \times 100\% = 80.0\%$$

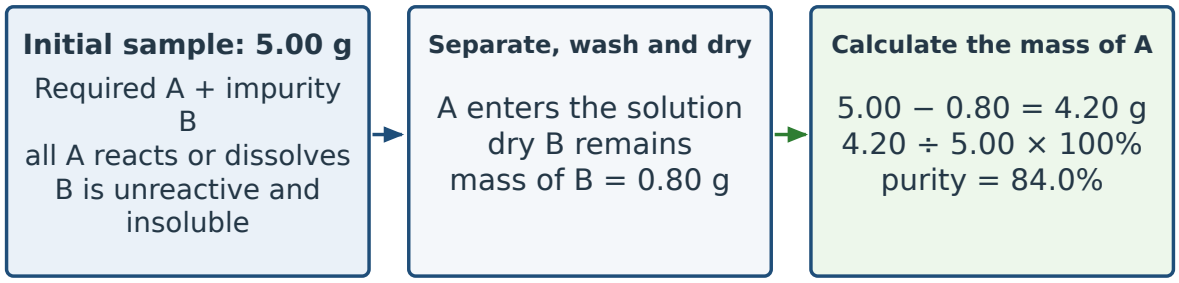
If impurity mass is given, first subtract it from the total. The 20% impurity fraction is not the target salt's percentage purity.

Conversely, if purity and total mass are given, convert the percentage to a fraction to find target mass:

$$\text{mass of target substance} = \text{total sample mass} \times \frac{\text{purity}}{100\%}$$

Some questions do not weigh the target substance directly. Instead, they react or dissolve it and weigh the impurities left behind. Read the conditions carefully: **all the target reacts or dissolves, the impurities neither react nor dissolve, and the remaining solid is washed and dried**. Only then does the dry residue represent the original impurities, allowing original total mass minus dry residue to give original target mass. Residue is the material left after filtration. Dry means that the measured mass should not include water.

Using a dry residue: first identify exactly what remains



Sample mass − dry impurity mass gives the required substance's mass only under these selective conditions.

A wet residue includes water, so its mass cannot be used as the impurity mass.

Arrows show sample treatment and the calculation of target mass from dry impurity mass. A and B are substance labels.

In the diagram, a 5.00g original sample leaves 0.80g of dry impurities. The target mass is:

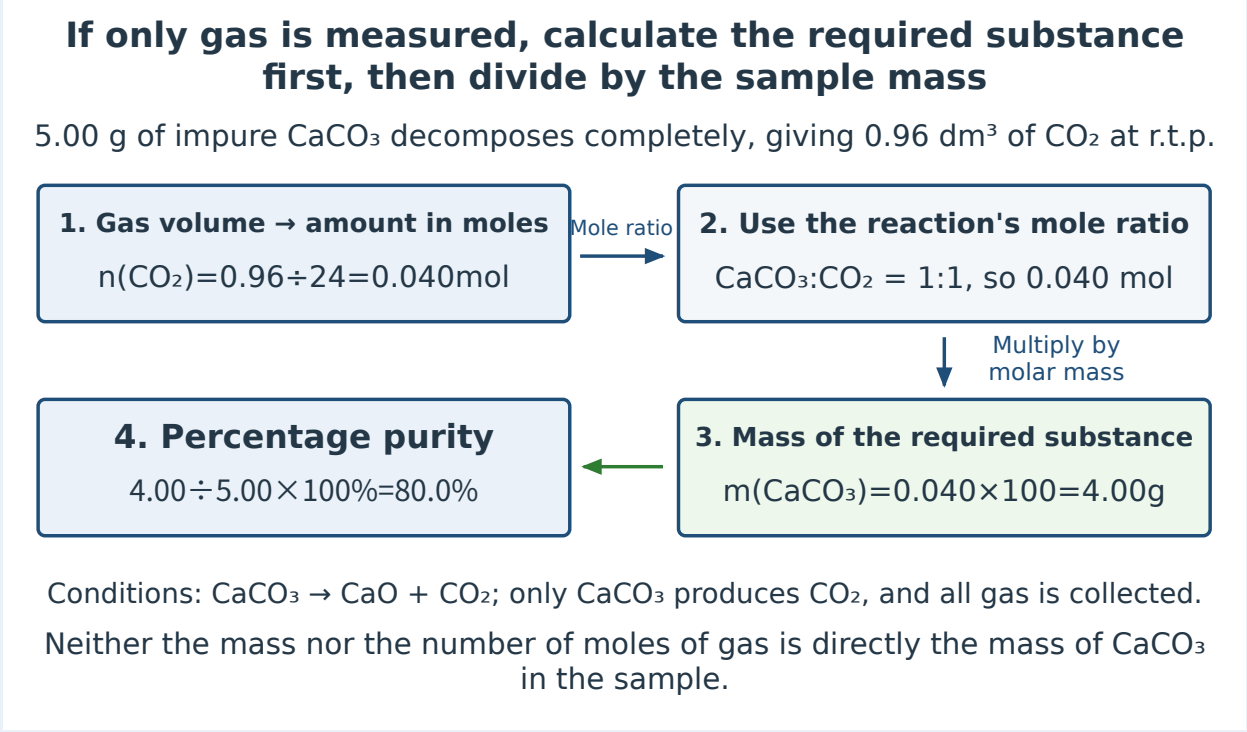
$$m_{\text{target}} = 5.00 - 0.80 = 4.20\,\text{g}$$

Compare this with the original total sample mass:

$$\text{purity} = \frac{4.20}{5.00} \times 100\% = 84.0\%$$

If impurities also dissolve, or some target fails to react, subtraction of the dry residue no longer isolates the target mass. If the residue is wet, its water makes it too heavy, making the calculated target mass and purity too low. The experimental conditions determine whether this subtraction is valid.

Other questions measure only a gas produced by reaction. Use the route already learned in Section 3.3: gas volume → gas amount in mol → target amount in mol using coefficients → target mass in g using molar mass. **Only then** divide by the original impure sample mass. Gas mass, gas amount in mol and target-substance mass are different quantities.



Arrows convert volume to mol, use coefficients to find target mol, multiply by molar mass and finally calculate purity. The supplied conditions specify that the impurities produce no gas.

For example, only $CaCO_3$ in a 5.00g sample produces CO_2, and reaction and collection are complete. The supplied reaction is:

$$CaCO_3 \rightarrow CaO + CO_2$$

The CO_2 collected has volume 0.96dm³ at r.t.p. Convert the gas volume to amount:

$$n(CO_2) = \frac{0.96}{24} = 0.040\,\text{mol}$$

The $CaCO_3$:CO_2 coefficient ratio is 1:1, so 0.040mol of $CaCO_3$ reacted. Its molar mass is 100g/mol, giving target mass:

$$m(CaCO_3) = 0.040 \times 100 = 4.00\,\text{g}$$

Only now divide by the original sample mass:

$$\text{purity} = \frac{4.00}{5.00} \times 100\% = 80.0\%$$

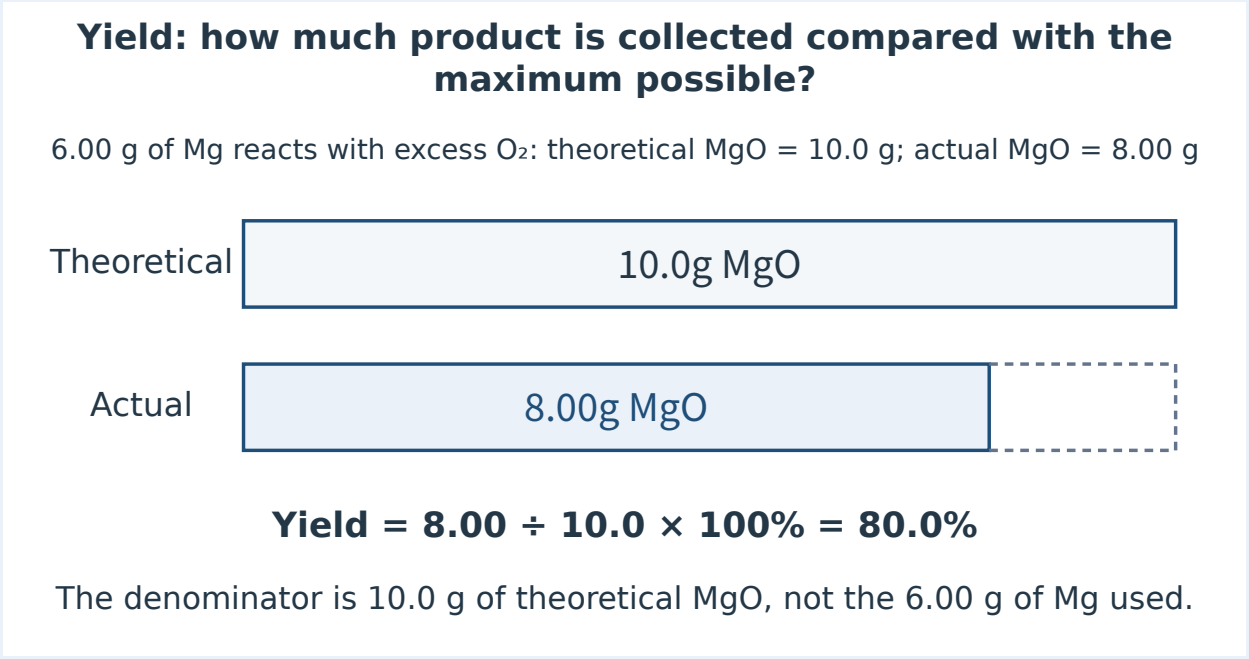
If the coefficient ratio changes to 2:1, the conversion to target mol must change too. Do not always copy the gas amount straight across.

3. Actual and theoretical yield. The **theoretical yield** is the product amount possible from the starting amounts, limiting reactant and balanced equation, assuming the specified reaction is complete and all product is collected. The **actual yield** is the amount collected and measured experimentally. It may be smaller because reaction is incomplete, other reactions occur or product is lost during transfer or filtration. Do not assume that atoms have disappeared.

Percentage yield compares actual yield with theoretical yield:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

The numerator and denominator must refer to the **same product in the same unit**. Calculate theoretical yield from the equation first. Dividing actual product mass directly by reactant mass is not valid, because the reaction may take in or release the mass of other substances.



Both mass axes use the same scale. The dashed section is theoretical product not collected; it does not claim that every shortfall is mechanical loss of product already formed.

For example, 6.00g of Mg reacts with sufficient oxygen and 8.00g of MgO is actually collected. The balanced equation is:

$$2Mg + O_2 \rightarrow 2MgO$$

First calculate the amount of Mg:

$$n(\text{Mg}) = \frac{6.00}{24} = 0.250\,\text{mol}$$

The Mg:MgO coefficient ratio is 1:1, so the theoretical amount of MgO is 0.250mol. Use its molar mass of 40g/mol to find the theoretical product mass:

$$m_{\text{theoretical}}(\text{MgO}) = 0.250 \times 40 = 10.0\,\text{g}$$

Compare the actual 8.00g with the theoretical 10.0g:

$$\text{yield} = \frac{8.00}{10.0} \times 100\% = 80.0\%$$

Using 8.00÷6.00 compares two different substances, product and reactant, and does not give percentage yield.

If theoretical yield is already supplied, calculate percentage yield directly. If both reactant quantities are given, first compare each amount in mol divided by its coefficient to identify the limiting reactant, then calculate theoretical product. Excess means that the other reactant is sufficient.

If percentage yield and theoretical yield are known, actual yield is:

$$\text{actual yield} = \text{theoretical yield} \times \frac{\text{yield}}{100\%}$$

If actual yield and percentage yield are known, calculate theoretical yield by:

$$\text{theoretical yield} = \frac{\text{actual yield}}{\frac{\text{yield}}{100\%}}$$

Retain enough digits during the calculation and round only at the end as requested.

Finally check the quantities and ranges. The denominator for composition is the compound mass, for purity the original sample mass, and for yield the theoretical product amount. Every selected mass must match the stated conditions. Purity and percentage composition must lie between 0% and 100%. If an experimental yield exceeds

100%, check for water or impurities in the product, weighing errors or an incorrect denominator. Do not interpret it as creating extra atoms.

Recognising the Question Name the denominator first: the whole compound, the original sample or the theoretical product. For elemental composition use subscript × A_r. For purity find the target mass first. For yield calculate the theoretical amount from the equation before dividing.

Worked Example 1 (2024·February/March 0620/42 Q1(b)(i)) · Level 2

(i) Calculate the percentage by mass of iron in iron(III) oxide, Fe_2O_3.

Solution

Identify the part and the whole. The numerator is the mass contribution of two Fe atoms; the denominator is the relative formula mass of the whole Fe_2O_3 formula. The Periodic Table gives A_r(Fe)=56 and A_r(O)=16, so iron contributes:

$$2 \times 56 = 112$$

The relative formula mass of iron(III) oxide is:

$$M_{r(Fe_2O_3)} = 2 \times 56 + 3 \times 16 = 160$$

The percentage by mass of iron is therefore:

$$\frac{112}{160} \times 100\% = 70.0\%$$

The answer is **70.0%**. Using 2÷5 to obtain 40% confuses the atom ratio 2:3 with a mass ratio. Using only 56 in the numerator omits the second Fe atom in each Fe_2O_3 formula.

Worked Example 2 (2008·May/June 0620/02 Q5(f)) · Level 1

(f) A sample of 20 g of impure magnesium sulphate contains 19.5 g of magnesium sulphate.
Calculate the percentage purity of the magnesium sulphate.

Solution

The 20g is the whole impure sample, while 19.5g is the target substance, magnesium sulfate. Both masses are in g, so no unit conversion or preliminary reaction equation is needed.

Use target magnesium sulfate mass as the numerator and original total sample mass as the denominator:

$$\text{purity} = \frac{19.5}{20} \times 100\% = 97.5\%$$

Check that the impurity mass is 20−19.5=0.5g, or 2.5%. The target salt's 97.5% and impurities' 2.5% total 100%. Do not mistake the impurity percentage for magnesium sulfate purity.

Worked Example 3 (2024·October/November 0620/21 Q10) · Level 3

Calcium carbonate is heated. Calcium oxide and carbon dioxide gas are formed.

The equation for the reaction is shown.

$$CaCO_3 \rightarrow CaO + CO_2$$

225 kg of calcium carbonate is heated until there is no further change in mass.

The yield of calcium oxide is 85 kg.

What is the percentage yield?

A 37.8% **B** 47.2% **C** 67.5% **D** 85.0%

Solution

CaO yield is required, so the denominator must be the **theoretical mass of CaO**, not the initial 225kg of $CaCO_3$.

First calculate relative formula masses:

$M_{r(CaCO_3)} = 40 + 12 + 3 \times 16 = 100$

$M_{r(CaO)} = 40 + 16 = 56$

The supplied reaction is:

$CaCO_3 \rightarrow CaO + CO_2$

$CaCO_3$ and CaO have coefficient ratio 1:1. Every 100 mass units of $CaCO_3$ can theoretically produce 56 mass units of CaO:

$$m_{\text{theoretical}}(CaO) = 225 \times \frac{56}{100} = 126\,\text{kg}$$

This is a mass-ratio calculation, so both masses can remain in kg. The actual CaO obtained is 85kg, giving:

$$\text{yield} = \frac{85}{126} \times 100\% = 67.460...\%$$

Round to **67.5%: choose C**. A's 37.8% comes from 85÷225, using the wrong substance in the denominator. D treats the number 85kg as 85% without comparing it with the theoretical yield.

Practice

1. (2025·May/June 0620/13 Q9) · Level 2

The relative formula mass, M_r, of calcium carbonate, $CaCO_3$, is 100.

What is the mass of carbon present in 100 g of calcium carbonate?

A 12 g **B** 36 g **C** 40 g **D** 60 g

2. (2024·October/November 0620/43 Q3(a)) · Level 2

(a) Tin(IV) oxide has the formula SnO_2.

The relative formula mass, M_r, of SnO_2 is 151.

Calculate the percentage by mass of tin in SnO_2.

3. (2025·October/November 0620/62 Q1(b)(iii)) · Level 2

Eggshells are made from a mixture of calcium carbonate and other insoluble substances.

Calcium carbonate reacts with dilute hydrochloric acid. The equation for the reaction is shown.

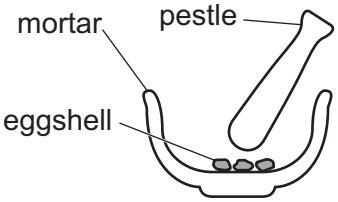
$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

The other substances in eggshells do **not** react with dilute hydrochloric acid.

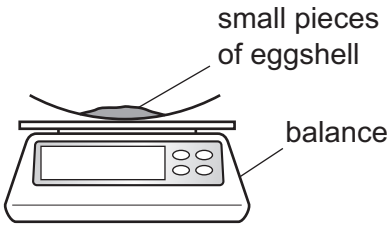
A student finds the percentage by mass of calcium carbonate in an eggshell.

Fig. 1.1 shows the first three steps of the method the student uses.

step 1
Grind the eggshell into small pieces.



step 2
Find the mass of the small pieces of eggshell.



step 3
Add the small pieces of eggshell to excess dilute hydrochloric acid and warm the mixture while stirring it with a glass rod.

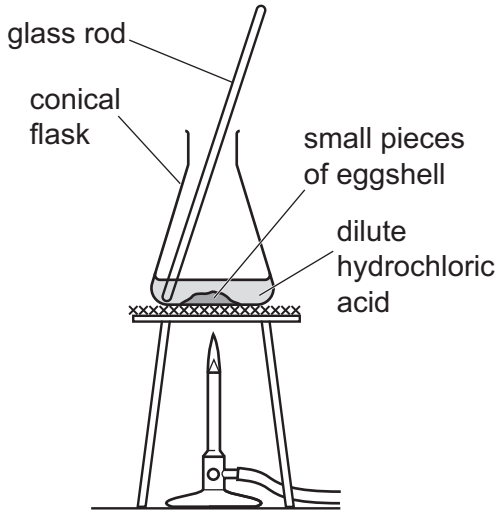


Fig. 1.1

(b) After **step 3**, the student filters the mixture to obtain the unreacted solid left from the eggshell.

(iii) After washing the residue, the student dries the residue in an oven.
The student then measures the mass of the dry residue.

The masses the student records are shown in Table 1.1.

Table 1.1

mass of small pieces of eggshell / g	2.00
mass of dry residue after filtration / g	0.11

Use the data in Table 1.1 to calculate the percentage by mass of calcium carbonate in the eggshell.

percentage of calcium carbonate =

4. (2010·May/June 0620/32 Q8(c)) · Level 3

(c) A 5.00 g sample of impure lead(II) nitrate was heated. The volume of oxygen formed was 0.16 dm^3 measured at r.t.p. The impurities did not decompose.
Calculate the percentage of lead(II) nitrate in the sample.

$$2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$$

Number of moles of O_2 formed =

Number of moles of $Pb(NO_3)_2$ in the sample =

Mass of one mole of $Pb(NO_3)_2$ = 331 g

Mass of lead(II) nitrate in the sample = g

Percentage of lead(II) nitrate in sample =

5. (2024·May/June 0620/21 Q10) · Level 3

When heated, copper(II) oxide, CuO, reacts with ammonia, NH_3.

$$3CuO + 2NH_3 \rightarrow 3Cu + N_2 + 3H_2O$$

8.5 g of ammonia reacts with an excess of copper(II) oxide to produce 26.4 g of copper.

What is the percentage yield of copper in this reaction?

A 27.5% **B** 32.2% **C** 55.0% **D** 82.5%

6. (2024·May/June 0620/23 Q7) · Level 3

The equation for the reaction of iron(III) oxide with carbon monoxide is shown.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

What is the percentage yield of iron when 16.8 g of carbon monoxide reacts completely with iron(III) oxide to form 8.96 g of iron?

A 26.7% **B** 40.0% **C** 53.3% **D** 80.0%

Electrochemistry

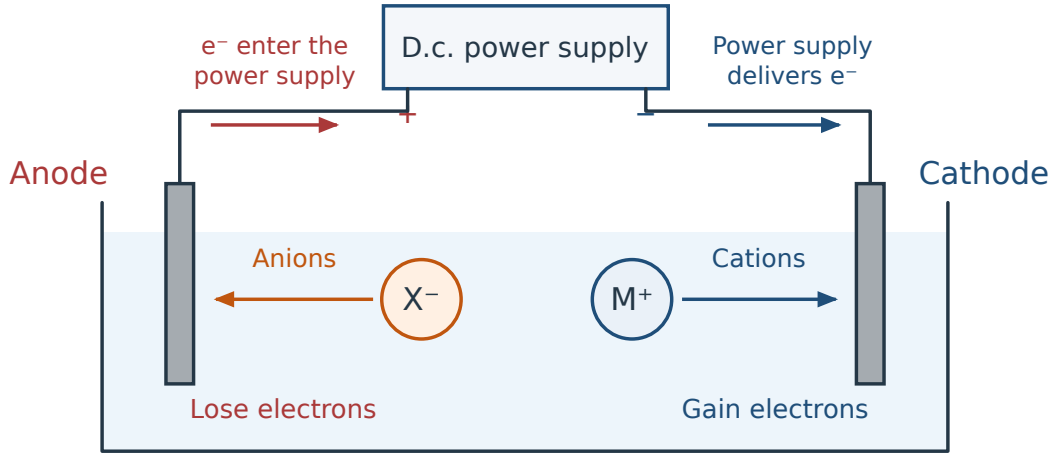
4.1 Electrolysis

Type 1: Tracking charge movement and predicting products of molten electrolysis

Key Ideas **Electrolysis** is the decomposition of an ionic compound, when molten or in aqueous solution, by the passage of an electric current. Read all three conditions together: the substance decomposed is ionic, it is molten or in aqueous solution, and the decomposition is caused by electric current. Molten means melted; aqueous means dissolved in water. These are different conditions.

The apparatus holding the melt or solution is an **electrolytic cell**. A conductor dipping into it, allowing charge transfer between the wire and the electrolyte, is an **electrode**. The melt or aqueous solution being electrolysed is the **electrolyte**. Electrolysis uses an external direct-current supply. The electrode connected to its positive terminal is the **anode** and that connected to its negative terminal is the **cathode**. Do not memorise left and right positions: follow the wires to the supply's + and −. In a single-cell symbol, the longer line is the positive terminal and the shorter line the negative terminal.

One circuit: electrons carry charge in wires; moving ions carry charge in the electrolyte



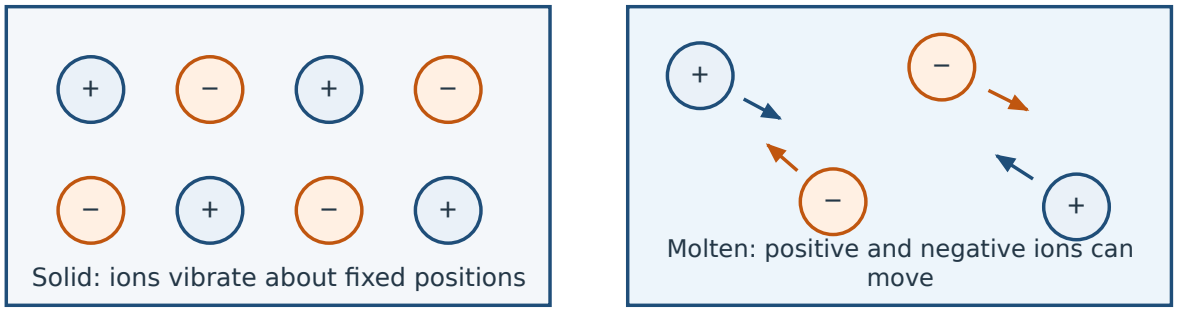
Electrons move through the metal wires; they are not the charge carriers moving through the solution.

X^- and M^+ represent oppositely charged ions. Arrows distinguish directed electron movement in wires from ion migration in solution.

Both the wires and electrolyte carry charge, but different particles move. **Electrons move in metal wires**: the supply removes electrons from the anode and supplies electrons to the cathode. **Ions move in the electrolyte**: positive ions, or cations, are attracted to the negative cathode; negative ions, or anions, are attracted to the positive anode. A cation goes to the cathode, not the anode. On reaching an electrode, ions gain or lose electrons and form new products. Do not explain conduction by electrons passing directly through the solution.

Chapter 2 established that ionic solids already contain positive and negative ions. In the solid, these are restricted to fixed lattice positions and cannot migrate freely throughout the substance, so sustained electrolysis cannot occur in this way. Melting or dissolving allows ions to move; it does not turn all neutral atoms into electrons.

The ions already exist; melting allows them to move throughout the liquid



Melting is a physical change; decomposition when current flows is the chemical change called electrolysis.

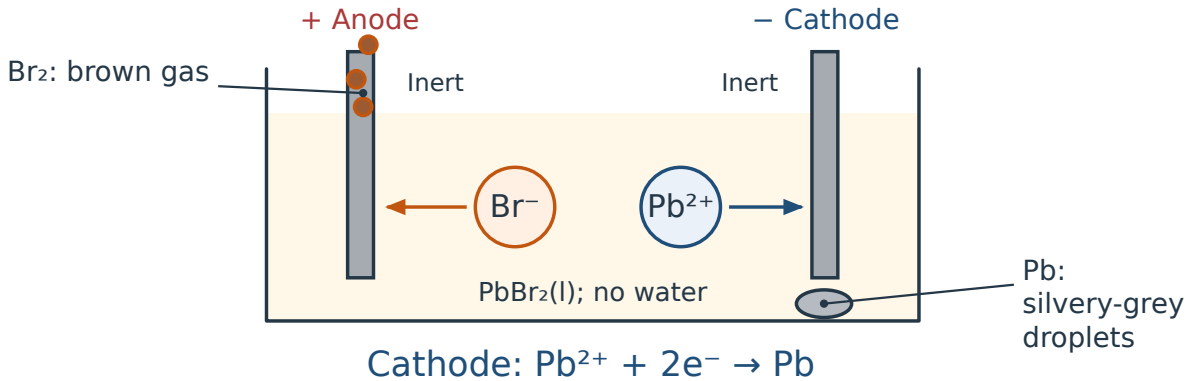
The same ionic substance is compared before and after melting. Arrows show mobile ions, not newly created electrons.

An **inert electrode** does not significantly take part in the reaction under the stated conditions. Platinum, Pt, and carbon or graphite, C, conduct and can act as inert electrodes in the experiments specified here. They provide surfaces for electron transfer. A carbon electrode does not make carbon the product in every question. Electrode inertness is a condition: when a copper anode is used later, the copper itself reacts.

For example, consider electrolysis of molten lead(II) bromide, $PbBr_2$, which contains Pb^{2+} and Br^-. Heating first melts it so that ions can move. An electric current is still required for the decomposition.

Pb^{2+} migrates to the cathode and gains two electrons to become lead, Pb. Br^- migrates to the anode, loses electrons and forms diatomic Br_2 molecules. Anhydrous $PbBr_2$ contains no hydrogen-bearing particles, so hydrogen cannot be produced.

Molten lead(II) bromide: identify Pb^{2+} and Br^-, then follow the electrode charges



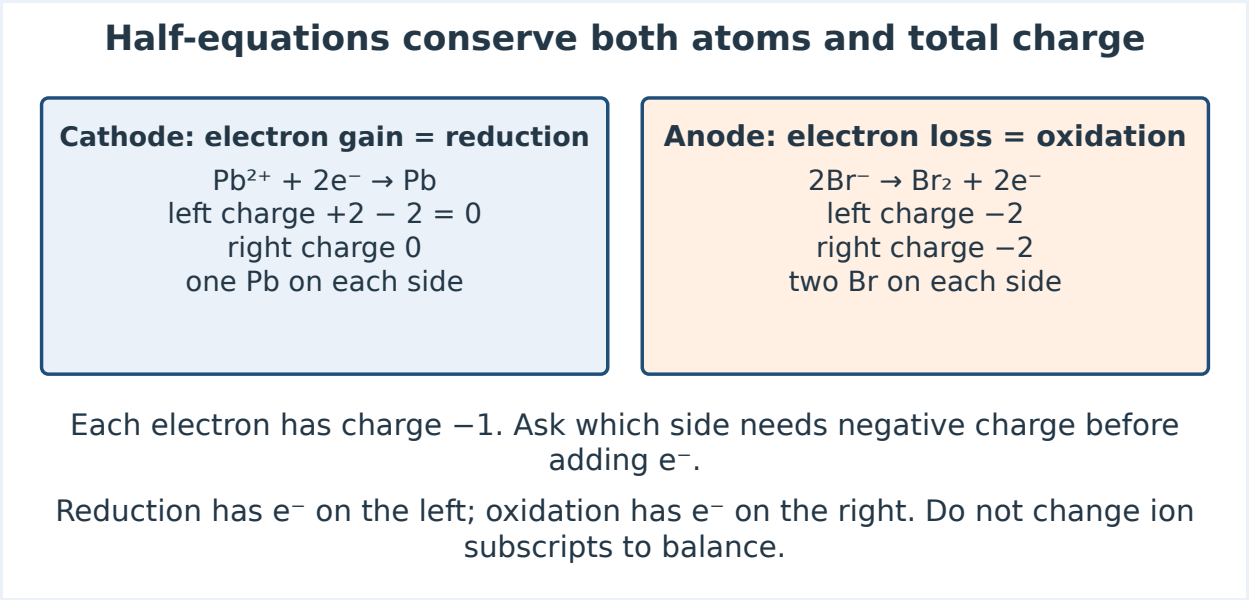
Cathode: $Pb^{2+} + 2e^- \rightarrow Pb$

Anode: $2Br^- \rightarrow Br_2 + 2e^-$

Ion arrows point towards their respective electrodes. Lead droplets form at the electrolysis temperature and become solid on cooling. Particle sizes are not drawn to scale.

The **products** are lead at the cathode and bromine at the anode. The corresponding **observations** are silvery-grey metal droplets at the cathode and brown bromine gas at the anode. Lead can form liquid droplets at the electrolysis temperature and solidify on cooling. Gas colour or the appearance of a metal is visible. Ions gaining electrons is an explanation, not a naked-eye observation.

An **ionic half-equation** records electron gain or loss at one electrode. The symbol e^- represents an electron with charge −1. **Gain of electrons is reduction; loss of electrons is oxidation**. Reduction therefore occurs at the cathode, with electrons on the left. Oxidation occurs at the anode, with electrons on the right. Use these electron definitions first; oxidation and reduction do not require oxygen gas to occur in every reaction.



The two electrodes belong to the same electrolysis of molten $PbBr_2$. The e^- terms directly record electron gain or loss.

First check the cathode half-equation:

$$Pb^{2+} + 2e^- \rightarrow Pb$$

Each side has one Pb. On the left, +2 balances the two electrons' −2 to give total charge zero, matching neutral Pb on the right.

The anode produces Br_2, so two Br^- are needed on the left. Their charge is −2. Adding $2e^-$ on the right also gives charge −2:

$$2Br^- \rightarrow Br_2 + 2e^-$$

Check both atom numbers and total charge in every half-equation. Balanced atoms alone are insufficient.

For another molten ionic compound containing two elements, use the same approach: list its ions, assign their destinations and identify the resulting elements. Metal ions gain electrons at the cathode to form the metal. Negative ions lose electrons at the anode to form the corresponding non-metal. In these questions, if hydrogen is produced, it is at the cathode, not the anode.

A molten binary compound gives products from its own ions

Example: $MgCl_2(l)$

$Mg^{2+} \rightarrow$ cathode $\rightarrow Mg$
$Cl^- \rightarrow$ anode $\rightarrow Cl_2$
$Mg^{2+} + 2e^- \rightarrow Mg$
$2Cl^- \rightarrow Cl_2 + 2e^-$

Example: $Na_2O(l)$

$Na^+ \rightarrow$ cathode $\rightarrow Na$
$O^{2-} \rightarrow$ anode $\rightarrow O_2$
$Na^+ + e^- \rightarrow Na$
$2O^{2-} \rightarrow O_2 + 4e^-$

There is no water or hydrogen-containing ion: do not automatically predict H_2 at the cathode.

These predictions assume the stated inert electrodes; metal is not formed at the anode here.

Arrows first identify each ion's destination, then its product element. Anhydrous conditions rule out using hydrogen as an automatic answer.

For example, molten $MgCl_2$ contains Mg^{2+} and Cl^-. Mg^{2+} gains electrons at the cathode to produce magnesium:
$Mg^{2+} + 2e^- \rightarrow Mg$
Cl^- loses electrons at the anode to produce chlorine:
$2Cl^- \rightarrow Cl_2 + 2e^-$
For a molten oxide, the negative ions are O^{2-}, distinct from hydroxide, OH^-. To form one O_2, begin with two O^{2-} to supply two O atoms. Their total charge is −4, so add four electrons on the right:
$2O^{2-} \rightarrow O_2 + 4e^-$
For other molten oxides, continue checking both conservation of atoms and conservation of charge.

Recognising the Question Identify molten or aqueous conditions, then follow wires to the positive and negative electrodes. For a molten binary compound, list its own ions. Electrons are gained at the cathode and lost at the anode. Check atoms and charge in every half-equation.

Worked Example 1 (2025·May/June 0620/42 Q3(a)) · Level 1

This question is about electrolysis.

(a) Complete the definition of electrolysis by filling in the missing words.

The decomposition of _____ compounds when

_____ or _____ by the passage of an electric current.

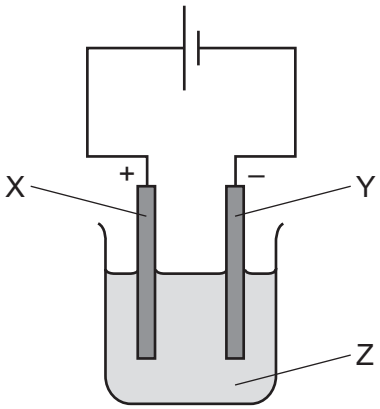
Solution

The question completes the definition of electrolysis. First identify what is decomposed: an **ionic** compound. Then identify when its ions can move: when **molten** or **in aqueous solution**.

The three gaps are **ionic; molten; in aqueous solution**. The complete meaning is the decomposition of an ionic compound, molten or in aqueous solution, by the passage of an electric current. The phrase by the passage of an electric current is already supplied, so do not substitute heating as the decomposition method. Heating may first melt the substance, but electrolytic decomposition still requires electric current.

Worked Example 2 (2025·May/June 0620/12 Q12) · Level 2

The diagram shows a simple electrolytic cell for an electrolysis experiment.



Three parts of the electrolytic cell are labelled X, Y and Z.

Which row shows the correct labels to replace X, Y and Z in the diagram?

	X	Y	Z
A	anode	cathode	aqueous salt
B	anode	cathode	solid salt
C	cathode	anode	aqueous salt
D	cathode	anode	solid salt

Solution

Use the supply's polarity rather than memorising left and right. X points to the electrode marked +, so X is the anode. Y points to the electrode marked –, so Y is the cathode.

Z points to the electrolyte in the beaker. Its ions must be mobile for this electrolysis. An aqueous salt is a salt solution and meets that condition. A solid salt has ions restricted to lattice positions and cannot conduct continuously in this way.

The row must therefore give **X=anode, Y=cathode, Z=aqueous salt: choose A**. B incorrectly specifies solid salt. C and D interchange the electrode names. All three labels must be correct in the same row.

Practice

1. (2025·May/June 0620/13 Q11) · Level 2

Which row describes the electrolysis of molten potassium bromide?

	product at anode	product at cathode
A	bromine	hydrogen
B	bromine	potassium
C	hydrogen	bromine
D	potassium	bromine

2. (2025·May/June 0620/22 Q10) · Level 2

Molten zinc oxide is electrolysed using inert electrodes.

Which row identifies the product at each electrode?

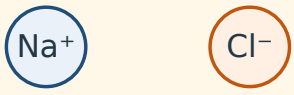
	anode	cathode
A	zinc	oxygen
B	hydrogen	oxygen
C	oxygen	zinc
D	oxygen	hydrogen

Type 2: Predicting aqueous electrolysis using concentration and electrode material

Key Ideas The molten salts in the previous question type contained no water. An **aqueous solution** contains water as well as ions from its solute, including small amounts of H^+ and OH^-. Acids and alkalis alter the concentrations of these ions. H^+ is the hydrogen ion and OH^- the hydroxide ion. Water can also react at an electrode, so NaCl(aq) cannot be treated as containing only Na^+ and Cl^- and assigned the same products as NaCl(l).

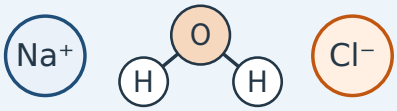
Water adds possible reacting particles, so molten-salt answers cannot simply be reused

NaCl(l): molten, no water



Candidates: Na^+ and Cl^-

NaCl(aq): dissolved in water



Also consider water and small amounts of H^+ and OH^-

Particle types only; not concentration ratios

Both are sodium chloride, but (l) and (aq) give different possible electrode reactants.

Water contains small amounts of H^+ and OH^-; acids and alkalis alter their concentrations. The numbers drawn only identify the particle types present.

Separate three conditions first: **what the solution contains, whether it is concentrated or dilute, and whether the electrodes are inert**. Then judge the cathode and anode independently. The main products below follow the concentration and electrode conditions supplied in this chapter's questions. Actual experiments also depend on the precise concentrations, current and electrodes. A salt name alone does not determine a unique answer independently of conditions.

At the cathode, compare metal deposition with hydrogen formation. The reactivity boundary needed here is as follows: potassium K, sodium Na, calcium Ca, magnesium Mg, aluminium Al, zinc Zn and iron Fe are more reactive than hydrogen; copper Cu and silver Ag are less reactive. In the usual aqueous prediction questions here, the first group of metal ions gives H_2 rather than direct deposition of Na or K as in molten salts. Cu^{2+} and Ag^+ can gain electrons and deposit Cu or Ag. Chapter 9 develops the complete reactivity comparison; use this stated boundary for now.

At an inert anode, **concentrated aqueous chlorides, bromides and iodides** produce chlorine Cl_2, bromine Br_2 and iodine I_2 respectively. Cl^-, Br^- and I^- are **halide ions**, distinct from neutral halogen molecules Cl_2, Br_2 and I_2. In this course's prediction questions that explicitly change the condition to a **dilute halide solution**, the main anode product becomes O_2. Do not overlook the deliberately stated dilute condition. Sulfate SO_4^{2-} and nitrate NO_3^- usually remain in solution here while water-related particles form oxygen at the anode; sulfate does not simply produce sulfur.

Read the state and electrodes first, then identify each electrode's product

These rules apply to the aqueous solutions and inert electrodes in this chapter

Cathode: metal or H_2?

Cu^{2+}, Ag^+ → Cu, Ag
with reactive metal ions such as Na^+ and K^+, water supplies H_2
H^+ in acids can also form H_2

Anode: halogen or O_2?

Concentrated Cl^- / Br^- / I^- → Cl_2 / Br_2 / I_2
stated dilute halide solutions → O_2
sulfate and nitrate solutions → O_2

With a copper anode, first consider whether copper loses electrons and dissolves; it is not inert.

Concentrated and dilute halides are the condition categories given in these course questions. Actual reactions also depend on concentration, electrodes and operating conditions.

Compare concentrated and dilute aqueous NaCl. Both contain water, and Na is more reactive than H, so both produce H_2 at the cathode. The H^+ form commonly used in questions is:

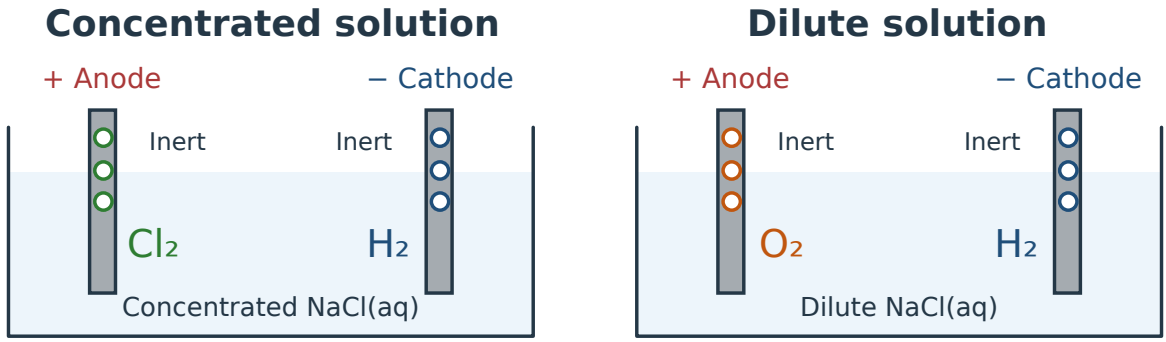
$$2H^+ + 2e^- \rightarrow H_2$$

An inert anode in concentrated NaCl solution produces chlorine:

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

When dilute NaCl and inert electrodes are explicitly specified, use the O_2 branch for the anode. Concentration changes the anode product here; it does not change the cathode product to sodium.

Changing only NaCl concentration keeps the cathode product but changes the main anode product



Both cathodes: $2H^+ + 2e^- \rightarrow H_2$

Anode: concentrated solution loses electrons from $2Cl^-$; dilute solution gives O_2 under the stated conditions.

The salt and inert electrodes remain the same while concentration is compared. Bubbles locate the products; they do not show actual relative production rates.

During electrolysis of concentrated brine, colourless gas bubbles appear at the cathode and chlorine bubbles at the anode. Sufficient chlorine has a yellow-green colour. The remaining solution contains Na^+ and OH^-, forming alkaline NaOH solution.

Writing the cathode reaction with water molecules shows that hydrogen comes from water while OH^- remains in solution:

$$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$$

If the options use H^+, the following form also records the electron number required for hydrogen formation:

$$2H^+ + 2e^- \rightarrow H_2$$

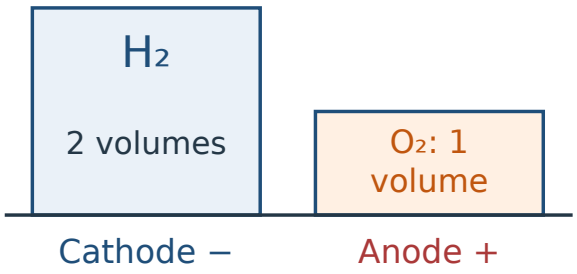
Neither form deposits Na^+ as metallic sodium.

Now consider **dilute sulfuric acid, $H_2SO_4(aq)$**. Its solution supplies conducting ions. With inert electrodes, hydrogen forms at the cathode and oxygen at the anode, with colourless bubbles at both. The overall change is:

$$2H_2O \rightarrow 2H_2 + O_2$$

At equal temperature and pressure, gas volume is proportional to amount, so the hydrogen volume is twice the oxygen volume. This follows from this reaction's coefficients; it does not mean that the negative electrode always produces more bubbles.

Dilute sulfuric acid: water decomposes, giving cathode H_2 and anode O_2 in a 2:1 volume ratio



Colourless gas bubbles at both electrodes

Cathode: $2H^+ + 2e^- \rightarrow H_2$

Anode: $4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

Overall: $2H_2O \rightarrow 2H_2 + O_2$. Compare gas volumes at the same temperature and pressure.

Column heights represent gas volumes at the same temperature and pressure, not electrode lengths. Dilute H_2SO_4 enables conduction; sulfate does not become sulfur.

For example, 12cm³ of oxygen collected during electrolysis of dilute sulfuric acid corresponds to 24cm³ of hydrogen at the same temperature and pressure. Locate oxygen at the anode and hydrogen at the cathode first, then use H_2:O_2=2:1. Sulfate does not become sulfur, and with no metal ions present, no arbitrary metal can form at the cathode. The gas tests needed here are: hydrogen gives a squeaky pop with a lighted splint; oxygen relights a glowing splint; chlorine bleaches damp litmus paper.

Electrode material is a separate condition. The blue colour of $CuSO_4(aq)$ comes from Cu^{2+}. With inert platinum or graphite electrodes, the cathode reaction is:

$Cu^{2+} + 2e^- \rightarrow Cu$

A red-brown copper layer forms and increases the cathode's mass. Water-related particles produce O_2 at the anode, seen as colourless bubbles. The inert anode does not replace Cu^{2+} in solution. As Cu^{2+} is removed, the blue fades and can approach colourless if enough is removed.

If the electrodes are changed to **copper**, the cathode still gains deposited copper and increases in mass:

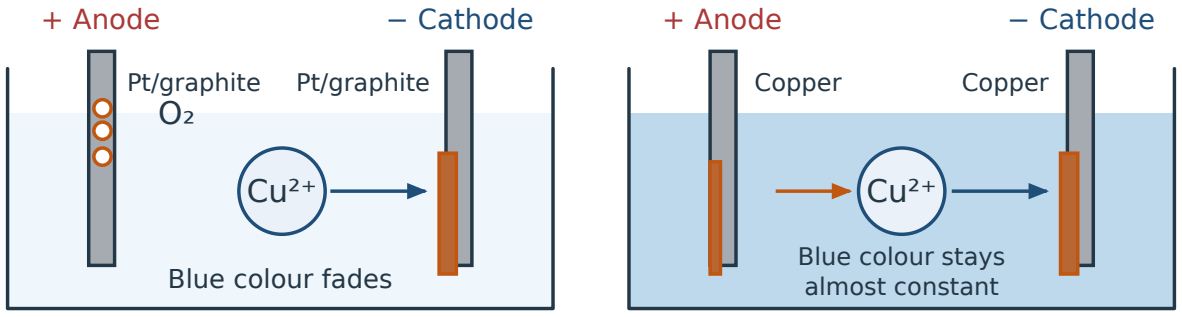
$Cu^{2+} + 2e^- \rightarrow Cu$

At the anode, however, copper itself loses electrons to form Cu^{2+}:

$Cu \rightarrow Cu^{2+} + 2e^-$

The copper anode loses atoms, becoming thinner and lighter while supplying Cu^{2+} to the solution. Both electrodes exchange two electrons per copper atom or ion. Cu^{2+} consumed at the cathode is therefore replaced in equal amounts at the anode, so the blue colour remains approximately unchanged under these conditions.

Same $CuSO_4$ solution: the anode material determines whether Cu^{2+} is replaced



Cathode removes Cu^{2+}; anode does not replace it

Cathode gains mass; bubbles at anode

Cathode removes Cu^{2+}; copper anode replaces the same amount

Cathode gains mass; anode loses mass

Right: $Cu \rightarrow Cu^{2+} + 2e^-$ and $Cu^{2+} + 2e^- \rightarrow Cu$ occur together

Both experiments use aqueous $CuSO_4$. Arrows show Cu^{2+} entering from the anode or moving to the cathode for deposition. Blue intensity indicates the concentration trend.

Draw the complete conclusion from the diagrams. The cathode gains mass in both $CuSO_4$ experiments because copper deposits in both. The solution becomes paler only in the inert-anode arrangement here, where Cu^{2+} is not replaced. Copper appearing at the cathode does not by itself prove that the solution must fade. Nor should a copper anode be treated as an inert electrode releasing oxygen: copper itself is now reacting.

The aqueous oxygen half-equation can be written in the common OH^- form. The products are O_2 and H_2O. Four OH^- contain four O and four H atoms, so O_2 and $2H_2O$ are needed on the right to conserve atoms.
The left-hand charge is now −4, while the neutral products have charge zero. Add four electrons on the right:
$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
Electrons on the right show oxidation at the anode.

OH^- to oxygen: balance atoms first, then charge

Identify the products: O_2 and H_2O

$$4OH^- \rightarrow O_2 + 2H_2O$$

Left: O = 4, H = 4

Left charge: −4

Right: O = 2 + 2, H = 4

Right charge is initially 0; add $4e^-$

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

Both sides now have charge −4. Electrons are on the right, so this is oxidation.

The downward arrow shows completion of the equation by charge conservation. OH^- has not been replaced by the O^{2-} found in a molten oxide.

The corresponding oxygen-forming anode reaction can also be written using water molecules:
$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

Atoms and total charge are conserved in this form too. Do not replace a reaction specifically written with OH^- by the O^{2-} half-equation for a molten oxide: the oxygen-containing particles differ between these media.

Finally, distinguish the command words. State the product asks for a substance name. Describe what is seen asks for an observation. Explain also requires the particle changes responsible. Bubbles alone do not prove oxygen is present, and naming oxygen does not describe the bubbles.

Observation, substance identity and particle explanation answer three different questions

Visible observation	Product or change	Particle explanation
Colourless bubbles	H_2 or O_2: a test is needed	Particles react at the electrode
Reddish-brown coating	Copper, Cu, is deposited	Cu^{2+} gains electrons to form Cu
Blue solution fades	Cu^{2+} concentration decreases	Cu^{2+} is removed faster than it is replaced

Tests: H_2 gives a squeaky pop with a lighted splint; O_2 relights a glowing splint; Cl_2 bleaches damp litmus.

Bubbles are an observation; a gas name is an identification. Changes in blue colour must be linked to aqueous Cu^{2+}.

For example, at the inert anode during $CuSO_4$ electrolysis, what is seen can be answered with bubbles appear. If the gas is requested, write O_2. If asked why the blue fades, explain that Cu^{2+} gains electrons and deposits at the cathode, while the inert anode does not replenish Cu^{2+}, so its solution concentration decreases. These answers correspond respectively to observation, naming and explanation.

Recognising the Question Identify aqueous conditions, concentrated or dilute solution, and electrode material, then judge each electrode separately. For $CuSO_4$, track whether Cu^{2+} is replenished. Describe bubbles, deposits or colour when observations are requested. Check both atoms and charge in half-equations.

Worked Example 1 (2025·May/June 0620/42 Q3(b)(i)) · Level 2

This question is about electrolysis.

(b) Electrolysis of aqueous copper(II) sulfate is carried out using the apparatus shown in Fig. 3.1. The electrodes are made of platinum.

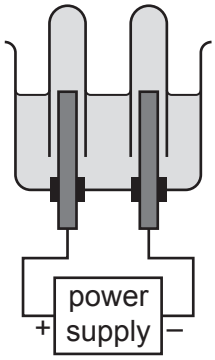


Fig. 3.1

(i) State whether the mass of the cathode increases, decreases or remains the same when platinum electrodes are used.

Solution

The question gives aqueous $CuSO_4$ and platinum electrodes. The cathode connects to the negative supply terminal and undergoes reduction:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Copper forms a red-brown coating on the platinum cathode.

The platinum has not changed into a different element, but the weighed cathode includes the attached copper, so its mass **increases**. Do not answer no change merely because platinum is inert. Inert means that platinum does not significantly react; it does not prevent new material depositing on its surface.

Worked Example 2 (2025·May/June 0620/42 Q3(b)(iv)) · Level 3

This question is about electrolysis.

(b) Electrolysis of aqueous copper(II) sulfate is carried out using the apparatus shown in Fig. 3.1. The electrodes are made of platinum.

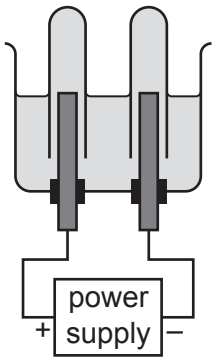


Fig. 3.1

(iv) Write the ionic half-equation for the reaction at the anode.

Solution

Check the electrode conditions first: aqueous $CuSO_4$ with a platinum anode. Platinum is inert, so do not write Cu → Cu^{2+}. Sulfate does not produce sulfur here either. Water-related particles produce oxygen at the anode; in the OH^- form introduced for this question type, the products are O_2 and H_2O.

Balance atoms first. $4OH^-$ contains four O and four H atoms. $O_2 + 2H_2O$ on the right gives 2+2=4 O and four H, conserving all atoms. The left now has charge −4 and the neutral products on the right have charge zero, so add $4e^-$ on the right.

The complete half-equation is:

$$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$$

Check O=4, H=4 and charge=−4 on both sides. Electrons on the right correctly show oxidation, or electron loss, at the anode. $OH^- \rightarrow O_2$ alone omits water, coefficients and electrons.

Worked Example 3 (2025·May/June 0620/42 Q3(c)(ii)) · Level 3

This question is about electrolysis.

(b) Electrolysis of aqueous copper(II) sulfate is carried out using the apparatus shown in Fig. 3.1. The electrodes are made of platinum.

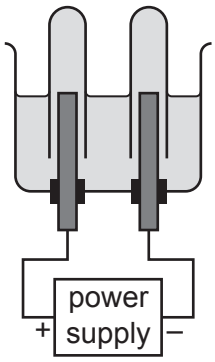


Fig. 3.1

(c) The electrolysis is repeated using copper electrodes.

(ii) Describe the change in appearance, if any, of the electrolyte when copper electrodes are used.

Solution

Read the new condition in part (c): platinum electrodes have been replaced with copper. Do not carry over the colour change for an inert anode.

The cathode still consumes Cu^{2+} from solution and deposits copper:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Meanwhile, the copper anode supplies fresh Cu^{2+}:

$$Cu \rightarrow Cu^{2+} + 2e^-$$

Both processes involve two electrons per copper atom or ion, so the anode replaces exactly the copper ions consumed at the cathode.

The blue colour of $CuSO_4$ solution comes from Cu^{2+}. Its concentration stays approximately constant under the stated conditions, so the electrolyte shows **no change** in appearance and remains blue. Copper deposition at the cathode does not automatically make the solution paler: the replenishment source is decisive.

Practice

1. (2025·May/June 0620/42 Q3(b)(iii)) · Level 1

This question is about electrolysis.

(b) Electrolysis of aqueous copper(II) sulfate is carried out using the apparatus shown in Fig. 3.1. The electrodes are made of platinum.

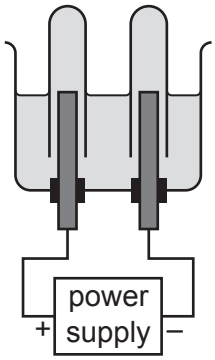


Fig. 3.1

(iii) Describe what is seen at the anode when platinum electrodes are used.

2. (2025·May/June 0620/11 Q11) · Level 2

Concentrated aqueous sodium chloride is electrolysed using inert electrodes.

What is the main product formed at the positive electrode (anode)?

A chlorine

B hydrogen

C oxygen

D sodium

3. (2025·May/June 0620/21 Q11) · Level 2

Concentrated aqueous sodium chloride is electrolysed using inert electrodes.

Which equation shows the correct ionic half-equation for the reaction at the anode?

A $Cl^- + e^- \rightarrow Cl$

B $2Cl^- \rightarrow Cl_2 + 2e^-$

C $H^+ \rightarrow H + e^-$

D $2H^+ + 2e^- \rightarrow H_2$

4. (2025·May/June 0620/22 Q11) · Level 2

Dilute aqueous sodium chloride is electrolysed using platinum electrodes.

What is the half-equation for the reaction at the cathode?

A $2H^+ + 2e^- \rightarrow H_2$

B $Na^+ + e^- \rightarrow Na$

C $2Cl^- \rightarrow Cl_2 + 2e^-$

D $4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

5. (2025·May/June 0620/42 Q3(b)(ii)) · Level 2

This question is about electrolysis.

(b) Electrolysis of aqueous copper(II) sulfate is carried out using the apparatus shown in Fig. 3.1. The electrodes are made of platinum.

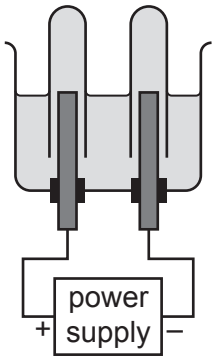


Fig. 3.1

(ii) Describe the change in appearance, if any, of the electrolyte when platinum electrodes are used.

6. (2025·May/June 0620/42 Q3(c)(i)) · Level 2

This question is about electrolysis.

(b) Electrolysis of aqueous copper(II) sulfate is carried out using the apparatus shown in Fig. 3.1. The electrodes are made of platinum.

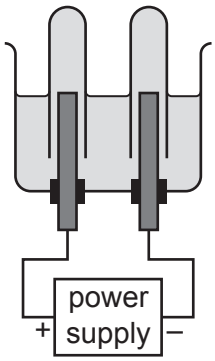


Fig. 3.1

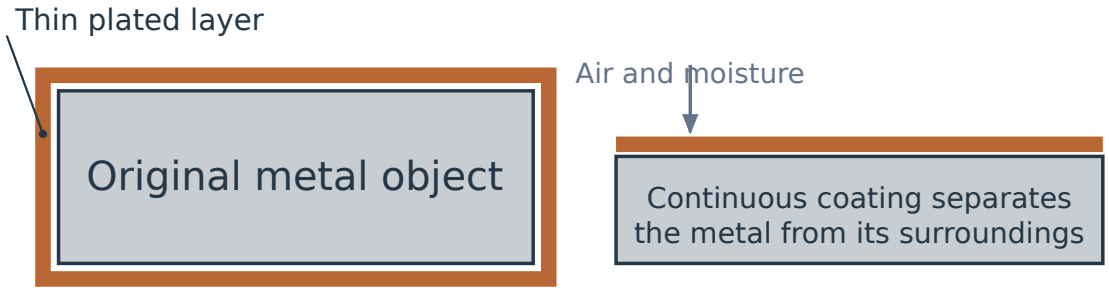
(c) The electrolysis is repeated using copper electrodes.

(i) State whether the mass of the cathode increases, decreases or remains the same when copper electrodes are used.

Type 3: Choosing electroplating apparatus and predicting electrode changes from the required coating

Key Ideas **Electroplating** uses electrolysis to deposit a thin metal layer on an object's surface. The object retains its original body but gains a surface of another metal. This can improve appearance by changing colour or lustre. A complete coating can also improve corrosion resistance. **Corrosion** is the gradual damage caused when a metal reacts with its surroundings. A continuous coating reduces contact between the underlying metal and air, moisture or other environmental substances.

Electroplating forms a layer of another metal on an object's surface



The appearance changes; an intact coating can improve corrosion resistance.

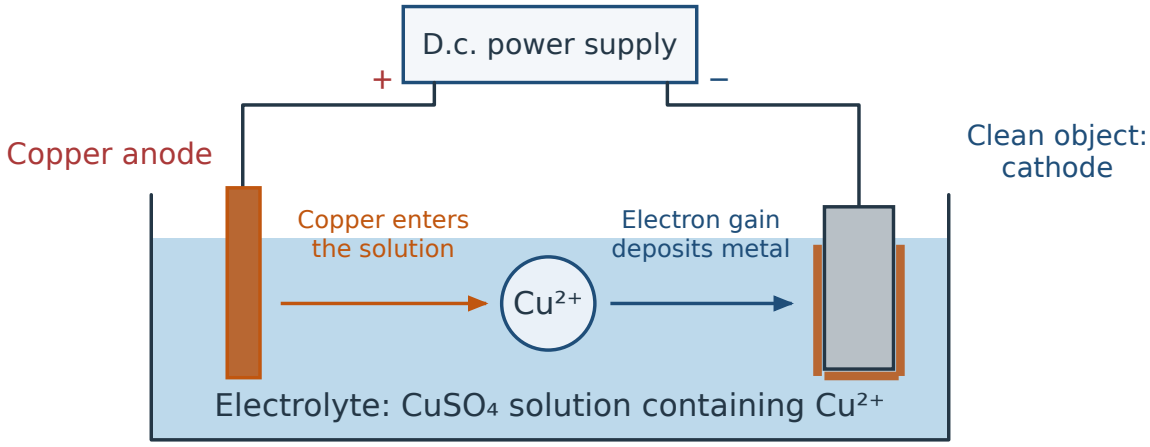
The coating is still a conductor: electroplating does not add an insulating plastic layer.

Coating thickness is enlarged in this cross-section. Arrows show environmental contact, and the horizontal coating acts as a barrier; they do not show current direction.

A metal coating still conducts, so making a metal object electrically insulating is not a general purpose of electroplating. Nor does electroplating make the whole object soft. Ask what a thin surface layer actually changes rather than assuming it alters every property of the original object.

How is the apparatus chosen? To deposit metal on the object, metal ions in solution must gain electrons at its surface. Connect the object to the negative terminal as the cathode. Use the required coating metal as the positive anode, with a solution containing that metal's ions. Anode metal loses electrons and replenishes ions; ions gain electrons at the cathode object to form the coating. The copper-electrode changes from the previous question type now become a useful process.

To deposit metal on the object, connect it to the negative terminal to receive electrons



For silver plating, use an Ag anode and Ag^+ solution; the object still connects to the negative terminal.

The left arrow shows the copper anode supplying Cu^{2+} to solution; the right arrow shows Cu^{2+} reaching the cathode and gaining electrons to form a coating.

For example, to copper-plate a steel strip, first clean its surface to remove grease and oxide so the coating can adhere. Connect the steel to the negative terminal of a direct-current supply and copper to the positive terminal, with both immersed in $CuSO_4$ solution.

When current flows, Cu^{2+} gains electrons at the steel surface and forms a red-brown copper coating:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

The copper anode continuously replaces copper ions:

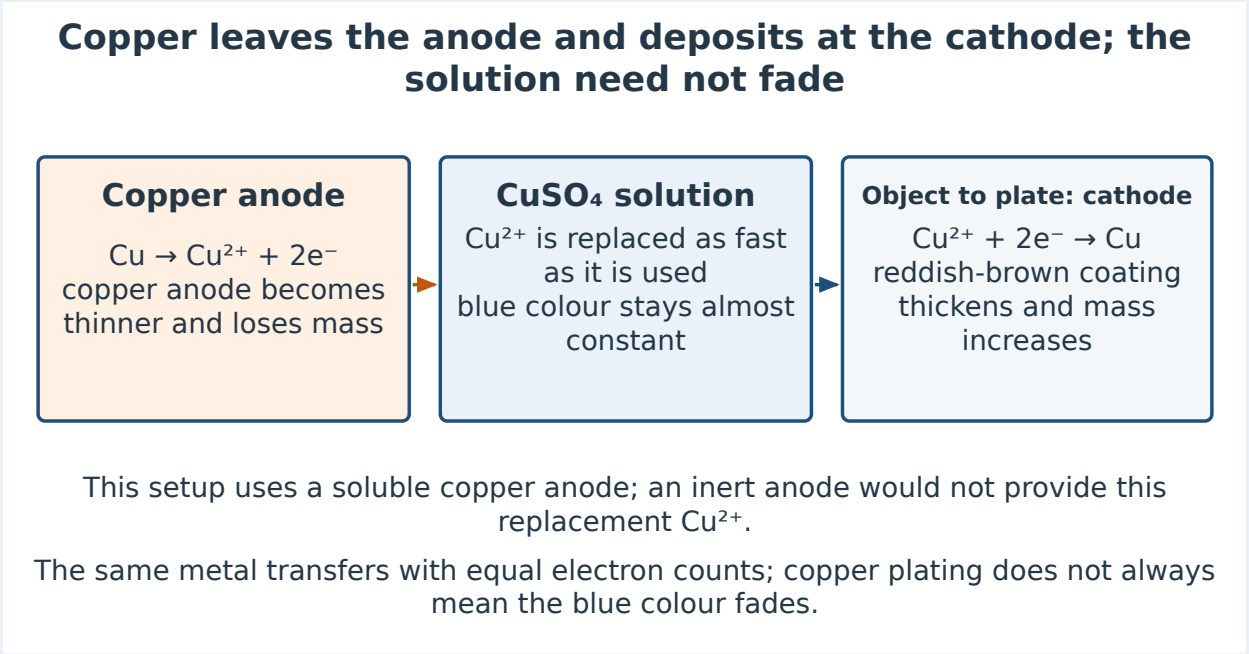
$$Cu \rightarrow Cu^{2+} + 2e^-$$

Current and time affect the amount deposited. Simply soaking two metals in water is not electroplating.

For silver plating, the object remains the cathode, but the anode becomes silver, Ag, and the solution must contain Ag^+. For example, aqueous silver nitrate, $AgNO_3$, supplies Ag^+. The cathode reaction is:

$$Ag^+ + e^- \rightarrow Ag$$

Whether the original object is nickel, Ni, or steel does not determine the coating metal. The coating comes from the ions reduced and deposited on its surface.



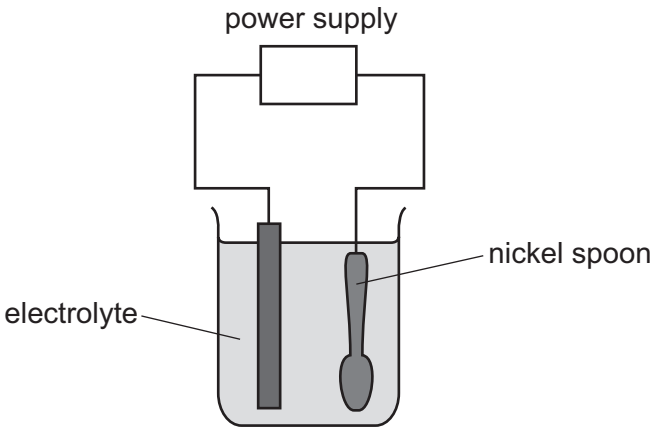
Arrows track copper from the anode, through ions in solution, to deposition on the object being plated.

Connect changes in all three places. Copper coating the steel gives a red-brown surface and increases its mass. Copper entering solution from the anode makes the anode thinner and lighter. Cu^{2+} consumed at the cathode is replaced in equal amounts at the anode, so the blue solution remains approximately unchanged. **Cathode mass gain, anode mass loss and unchanged solution colour can occur together**. They are not mutually exclusive.

Recognising the Question Identify the required coating metal. Connect the object to the negative terminal and the coating metal to the positive terminal, using a solution containing that metal's ions. Then track the object's surface, anode mass and whether solution ions are replenished.

Worked Example 1 (2025·February/March 0620/12 Q13) · Level 3

The diagram shows an experiment to electroplate a nickel spoon with silver.



Which row correctly describes the positive electrode, the negative electrode and the electrolyte?

	positive electrode	negative electrode	electrolyte
A	nickel spoon	pure nickel	silver nitrate solution
B	nickel spoon	pure silver	nickel nitrate solution
C	pure nickel	nickel spoon	silver nitrate solution
D	pure silver	nickel spoon	silver nitrate solution

Solution

The nickel spoon is to be coated with silver. The new silver layer must form on the spoon, so connect the spoon to the negative terminal as the cathode. Ag^+ can then gain electrons at its surface.
Use pure silver at the positive electrode so that it continually forms Ag^+:

$$Ag \rightarrow Ag^+ + e^-$$

The electrolyte must also supply silver ions. Silver nitrate solution, $AgNO_3(aq)$, is suitable. At the nickel spoon cathode:

$$Ag^+ + e^- \rightarrow Ag$$

Silver therefore deposits on the spoon.

The three positions require **pure silver at the positive terminal, the nickel spoon at the negative terminal, and silver nitrate solution as the electrolyte: choose D**. A and B already fail by connecting the spoon to the positive terminal. C connects the spoon correctly but uses a pure nickel anode, which cannot replenish Ag^+ as required by this silver-anode arrangement.

Practice

1. (2025·October/November 0620/13 Q12) · Level 2

Which statements explain why metal objects are electroplated?

1 Electroplating improves the appearance of the object.

2 Electroplating insulates the metal object.

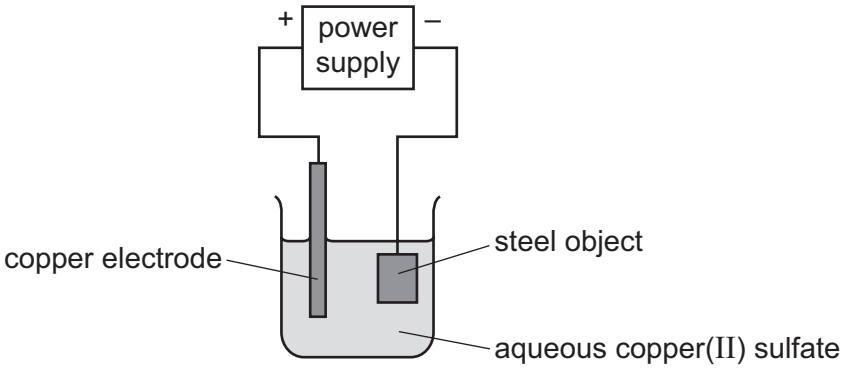
3 Electroplating makes the metal object more resistant to corrosion.

4 Electroplating makes the metal object softer.

A 1 and 2 **B** 1 and 3 **C** 2 and 4 **D** 3 and 4

2. (2025·May/June 0620/21 Q10) · Level 3

The diagram shows the electroplating of a steel object.



A student makes three statements.

1 The steel object turns a reddish-brown colour.

2 The aqueous copper(II) sulfate changes to a paler blue colour.

3 The copper electrode becomes thinner.

Which statements are correct?

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

4.2 Hydrogen–oxygen fuel cells

Type 1: Explaining hydrogen–oxygen fuel cells and comparing fuel options

Key Ideas A **hydrogen–oxygen fuel cell** converts chemical energy from the reaction of hydrogen with oxygen into electrical energy to power an external device such as a motor. The fuel is hydrogen, which must be supplied continuously. Oxygen must also be supplied, and water leaves the cell. The device does not endlessly reuse the same small initial quantity of hydrogen.

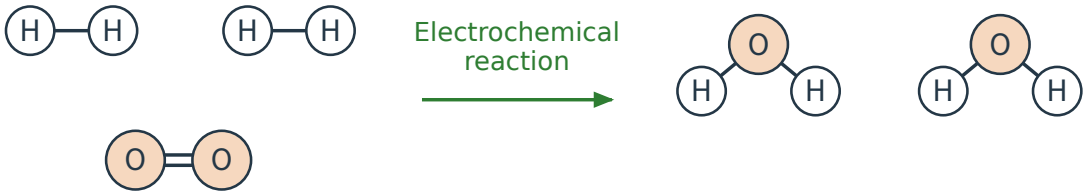
The overall reactants are H_2 and O_2, and **the only chemical product is H_2O**. Begin with the skeleton:

$H_2 + O_2 \rightarrow H_2O$

O_2 supplies two O atoms, so the right needs $2H_2O$. This gives four H atoms on the right, requiring $2H_2$ on the left. The balanced equation is:

$2H_2 + O_2 \rightarrow 2H_2O$

Overall hydrogen–oxygen fuel cell reaction: $2H_2 + O_2 \rightarrow 2H_2O$



Reactants: 4 H atoms, 2 O atoms

Only chemical product: water

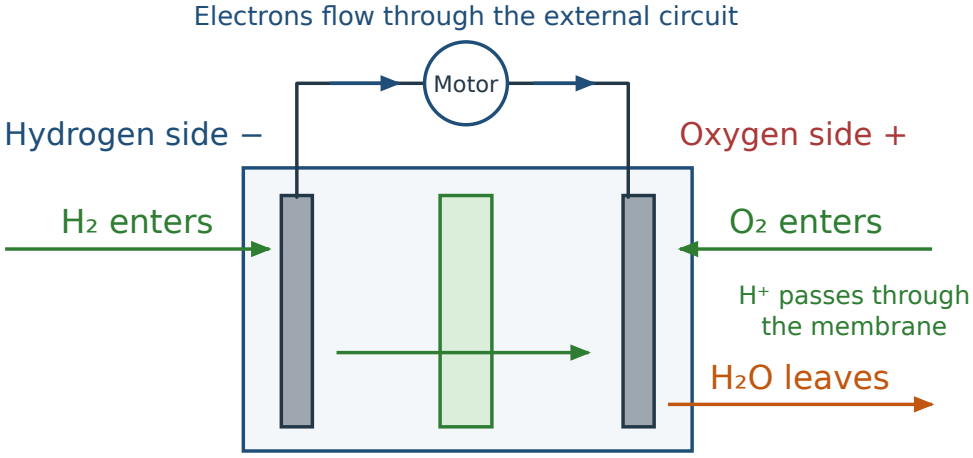
Electrical energy is an energy output, not another substance; H_2O_2 is hydrogen peroxide, not water.

The arrow shows the overall chemical change in the fuel cell. Sphere models represent atom numbers and connections, not a combustion flame or actual scale.

H_2O_2 is hydrogen peroxide, a different substance from H_2O. Do not change the fuel-cell product to H_2O_2 simply to avoid a coefficient. Electrical energy is an energy output, not another chemical product. Water being the only chemical product is consistent with simultaneous electrical output and some heat release.

How does the cell power an external circuit? A **proton exchange membrane** cell provides an example. Here a proton is H^+, what remains after a hydrogen atom loses its electron. The membrane allows H^+ through, while electrons flow through external wires and the electrical device. Hydrogen loses electrons on one side. Oxygen on the other side receives electrons from the external circuit and combines with H^+ to form water. Linking chemical reaction to directed charge transfer allows some reaction energy to reach the external device.

Chemical energy becomes electrical energy: fuel flows in and electrons do work through the external circuit



Proton-exchange membrane fuel cell: H^+ passes through the membrane; electrons take the external circuit.

Continuous H_2 and O_2 supply is needed for continuous electricity generation; some chemical energy also becomes heat.

Arrows distinguish gas and water flow, H^+ migration through the membrane, and electron flow in wires. The diagram represents a proton exchange membrane cell.

On the hydrogen side of this proton exchange membrane cell, electrons are lost:

$$2H_2 \rightarrow 4H^+ + 4e^-$$

On the oxygen side, electrons arrive from the external circuit and react with H^+ to form water:
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$
Add the two equations and cancel equal amounts of $4H^+$ and $4e^-$ from both sides to obtain the overall reaction:
$2H_2 + O_2 \rightarrow 2H_2O$
H^+ crosses the membrane while electrons flow through the external circuit. Do not draw both following the same pipe as water.

Oxidation and reduction define the anode and cathode. An electrolytic cell uses an external supply to drive its reaction, with a positive anode and negative cathode. The fuel cell shown supplies energy to the outside: its hydrogen anode loses electrons and is negative, while its oxygen cathode gains electrons and is positive. Do not generalise a positive anode to every cell. Both devices have oxidation at the anode and reduction at the cathode; the way energy is supplied or released differs.

Energy direction: electrolysis uses electricity; a fuel cell generates it

Electrolysis of water

$2H_2O \rightarrow 2H_2 + O_2$

Electrical energy in

Fuel cell

$2H_2 + O_2 \rightarrow 2H_2O$

Electrical energy out

The overall reactions run in opposite directions; apparatus, electrode materials and conditions need not be the same.

Diagonal arrows show energy input or output. Arrows in the chemical equations show the direction of substance change.

For example, a device using electrical energy to convert water into hydrogen and oxygen is carrying out electrolysis. A device taking in hydrogen and oxygen, producing water and supplying electricity is a fuel cell. Examine substance and energy arrows first; two electrodes alone do not identify the device.

When comparing with a petrol engine, keep the comparison's scope clear. Petrol, also called gasoline, is a mixture of carbon-containing substances. Burning it produces CO_2 and water; incomplete combustion may also produce harmful carbon monoxide, CO. A hydrogen–oxygen fuel cell produces only water **during vehicle use**, and its own reaction does not release CO_2. This is a clear advantage.

Comparing vehicles: separate emissions during use from hydrogen production and storage

Hydrogen–oxygen fuel cell in use

$2H_2 + O_2 \rightarrow 2H_2O$
water is the only chemical product
the cell reaction produces no CO_2
fuel must be supplied continuously

Petrol engine in use

Carbon-containing fuel burns to form CO_2 and H_2O
incomplete combustion can also form CO
liquid fuel is easier to store and transport

Hydrogen requires high-pressure storage, heavier tanks and refuelling facilities

CO_2 from hydrogen production depends on feedstock and energy source; cell emissions are not total process emissions.

The top row compares vehicle use; the lower row separately considers hydrogen production and storage. Petrol's mixture is not reduced to one fixed molecular formula here.

Hydrogen storage and transport present another practical condition. Carrying enough hydrogen on a vehicle usually requires high pressure and strong, relatively heavy tanks. It is harder to store than liquid petrol at ordinary temperature, and refuelling infrastructure also affects use. These are difficulties relevant to comparing options. Being flammable does not establish an advantage over petrol. A fuel needs an appropriate storage and supply system.

Also distinguish **using hydrogen** from **producing hydrogen**. Hydrogen is not a fossil fuel extracted like coal or petroleum. It can be produced from different feedstocks and energy sources. A **hydrocarbon** contains only carbon and hydrogen. Some hydrocarbons can undergo chemical reactions to produce hydrogen, but this differs from **fractional distillation** of petroleum. Fractional distillation physically separates existing components by differences in boiling point; it is not the chemical conversion of hydrogen atoms in hydrocarbons into H_2. Later organic chemistry chapters develop both processes. Use this distinction to read the questions here.

Producing hydrogen from fossil feedstocks or using electricity associated with emissions can generate CO_2 upstream. Electrolysis of water using renewable energy provides another route. A fuel cell releasing no CO_2 during operation therefore does not establish zero emissions throughout production and use.

If one advantage and one disadvantage are requested, answer each with a complete causal statement. An advantage is that, during vehicle use, water is the cell's only chemical product and the cell reaction produces no CO_2.

A disadvantage is that carrying enough hydrogen usually requires high-pressure storage, with strong, relatively heavy tanks, making storage and transport more difficult. These describe different aspects and can both be true.

Recognising the Question Check the direction hydrogen + oxygen → water, then distinguish the chemical product from electrical energy. When comparing vehicles, consider emissions during use, hydrogen production and hydrogen storage separately.

Worked Example 1 (2025·May/June 0620/12 Q13) · Level 2

Which equation represents the overall reaction that produces electricity in a hydrogen–oxygen fuel cell?

A $H_2 + O_2 \rightarrow H_2O_2$

B $2H_2O \rightarrow 2H_2 + O_2$

C $2H_2 + O_2 \rightarrow 2H_2O$

D $2H_2O_2 \rightarrow 2H_2O + O_2$

Solution

The fuel cell consumes H_2 and O_2 while supplying electricity, so its product must be water. Start with the skeleton:

$H_2 + O_2 \rightarrow H_2O$

One O_2 supplies two O atoms, requiring $2H_2O$ on the right. That gives four H atoms on the right, so $2H_2$ is also needed on the left.

The balanced equation is:

$2H_2 + O_2 \rightarrow 2H_2O$

Choose C. Both sides have four H and two O atoms.

A forms hydrogen peroxide, H_2O_2, so it has the wrong product. B decomposes water into hydrogen and oxygen, reversing the required direction. D decomposes hydrogen peroxide, which is also not a fuel-cell reaction starting from H_2 and O_2.

Worked Example 2 (2024·October/November 0620/41 Q5(a)(ii)) · Level 2

(a) Hydrogen is used in fuel cells to produce electricity in vehicles.

(ii) Give **one** advantage and **one** disadvantage of using fuel cells instead of gasoline in vehicle engines.

advantage

disadvantage

Solution

State the advantage first: **water is the cell reaction's only chemical product, and no CO_2 is produced during operation**. A suitable answer is Water is the only chemical product. Petrol contains carbon and produces CO_2 when burned. The overall H_2 and O_2 reaction involves only H and O, giving a clear basis for this product difference.

Then state the disadvantage: **hydrogen needs high-pressure storage, which makes storage more difficult and usually requires strong, relatively heavy tanks**. A suitable answer is Hydrogen needs to be stored at high pressure. Give one advantage and one disadvantage in their respective places; storage difficulty does not belong in the advantage box. The emissions statement applies to the cell reaction during vehicle use. Feedstocks and energy used to produce hydrogen may still cause CO_2 emissions, so do not extend it to zero emissions throughout the entire process.

Practice

1. (2025·May/June 0620/13 Q12) · Level 1

What is produced in a hydrogen–oxygen fuel cell?

A hydrogen and oxygen

B hydrogen only

C carbon dioxide

D water

2. (2025·May/June 0620/22 Q12) · Level 1

Hydrogen–oxygen fuel cells can be used to power vehicles.

Which statement about hydrogen–oxygen fuel cells is correct?

A The equation for the overall reaction is $H_2 + O_2 \rightarrow H_2O_2$.

B The only chemical products are water and carbon dioxide.

C Chemical energy in the fuel is converted into electrical energy.

D The hydrogen used in hydrogen–oxygen fuel cells is a fossil fuel.

3. (2025·May/June 0620/21 Q12) · Level 2

Which equation represents the overall reaction in a hydrogen–oxygen fuel cell?

A $2H_2 + O_2 \rightarrow 2H_2O$

B $2H_2O \rightarrow 2H_2 + O_2$

C $H_2 + O_2 \rightarrow H_2O_2$

D $H_2O_2 \rightarrow H_2 + O_2$

4. (2024·October/November 0620/21 Q13) · Level 2

Which statement describes an advantage of using a hydrogen–oxygen fuel cell in a car compared to a gasoline engine?

A The hydrogen is difficult to store.

B The hydrogen is highly flammable.

C The hydrogen used is made from hydrocarbons.

D The only chemical product is water.

5. (2024·May/June 0620/23 Q11) · Level 2

Hydrogen–oxygen fuel cells and gasoline are each used to power cars.

Which statement describes an advantage of using hydrogen–oxygen fuel cells in cars in comparison with gasoline engines?

A Hydrogen is a non-renewable resource.

B Hydrogen is produced during the fractional distillation of petroleum.

C Hydrogen–oxygen fuel cells do **not** release carbon dioxide.

D Hydrogen–oxygen fuel cells need electricity to work.

Chemical energetics

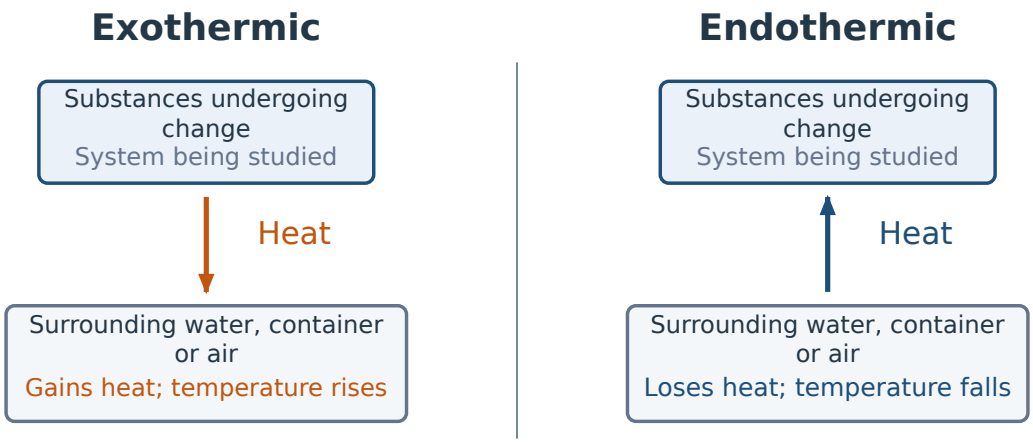
5.1 Exothermic and endothermic changes and bond energies

Type 1: Using heat flow and energy diagrams to identify energy changes and activation energy

Key Ideas If a reaction warms the surrounding water, heat has been transferred to that water. First distinguish the substances undergoing change from their surroundings. The former are the substances being studied; the latter may be water, the container or the air. The direction of heat transfer determines how the surroundings change.

An exothermic reaction transfers heat to the surroundings, which gain heat and rise in temperature. An endothermic reaction absorbs heat from the surroundings, which lose heat and fall in temperature. Absorbed and released are described from the viewpoint of the reacting substances.

Distinguish the changing substances from their surroundings



Identify the direction of heat transfer before deciding how the surroundings' temperature changes.

Arrows show heat transfer between the changing substances and their surroundings, not the direction of substance movement.

Thus, the surroundings become warmer means that the reaction releases heat to them. The temperature of the surroundings decreases means that the reaction absorbs heat from them. Do not link the reaction absorbing energy to warmer surroundings: the reacting substances are the side gaining energy.

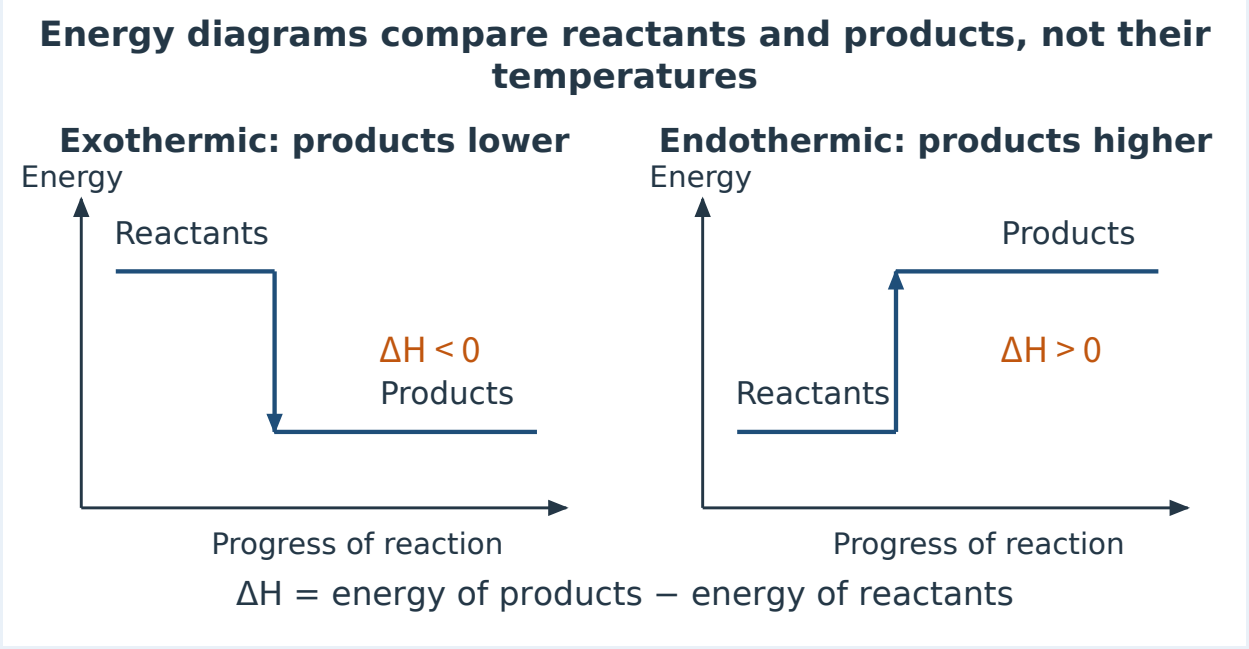
How is this shown on an energy diagram? Reactants are the starting substances and products are those formed. The vertical energy axis compares their energies. The horizontal progress of reaction axis follows the change from reactants to products without specifying a number of seconds. It is not a temperature–time graph.

In an exothermic reaction, some energy transfers to the surroundings, leaving the products at a lower energy level than the reactants. In an endothermic reaction, the substances gain energy from the surroundings, leaving the products at a higher level. Find the left and right ends and compare these levels; do not judge the whole reaction from one rising or falling section in the middle.

The net heat transfer of a reaction is described by its enthalpy change, ΔH. Δ means change and H denotes enthalpy. Here it is compared using the energy difference before and after reaction, calculated as final minus initial:

ΔH = energy of products − energy of reactants

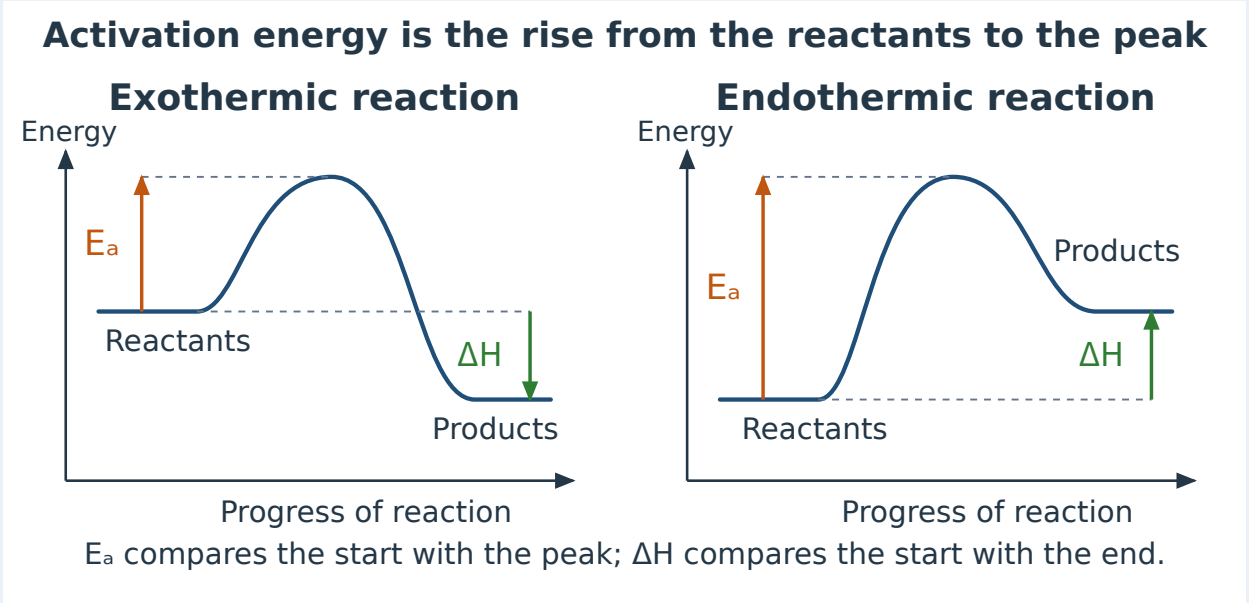
For an exothermic reaction, the final level is lower and ΔH is negative. For an endothermic reaction, the final level is higher and ΔH is positive.



The vertical axis is energy and the horizontal axis is progress of reaction. The height difference between the horizontal lines gives the net energy change. Only the starting and finishing levels are shown here.

For example, products above reactants mean net energy absorption. The reaction takes heat from the surroundings, lowering their temperature. Conversely, if the surroundings warm up, first identify an exothermic reaction, then choose a diagram with reactants above products. These are forward and reverse uses of the same relationship.

Some diagrams show a peaked pathway between the two levels. Reactant particles must collide with at least a certain energy before reaction can occur. This minimum requirement is the activation energy, E_a; the subscript a comes from activation. On the diagram, measure E_a upwards from the reactant level to the highest point of the pathway. Its endpoint differs from the endpoint used for ΔH.

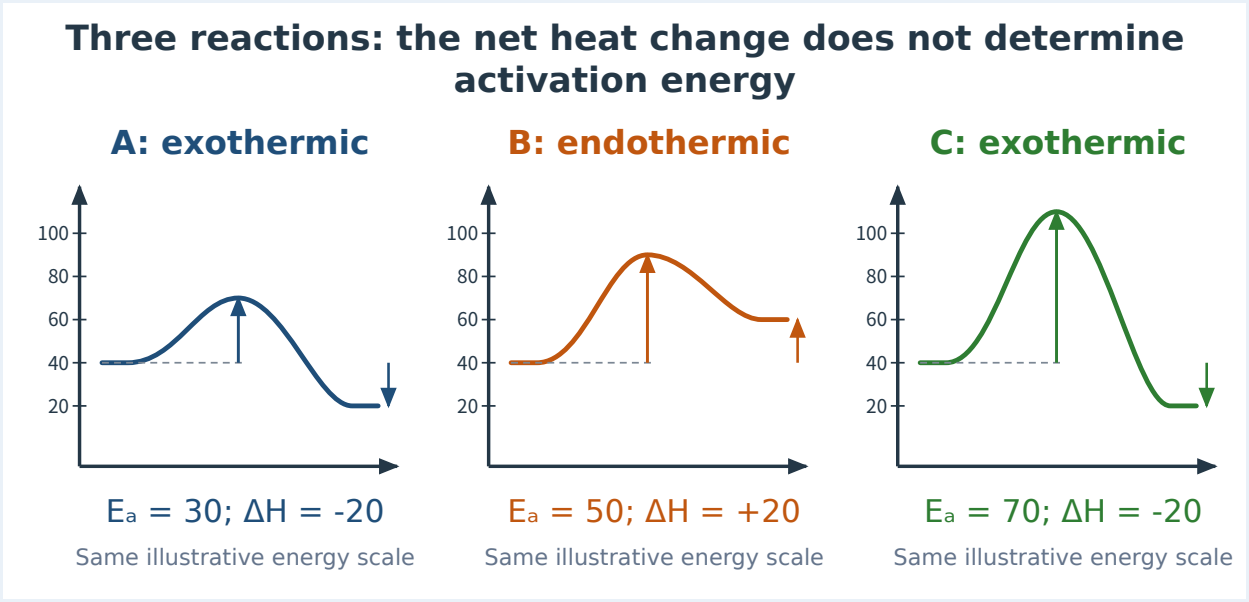


Dashed lines align the energy levels. Both reaction types must overcome an energy barrier. Even an exothermic reaction may need ignition or heating before it proceeds noticeably.

To draw an exothermic diagram, place the product line below the reactant line. For an endothermic diagram, place it above. Draw the peak above both lines and connect the starting point, peak and finishing point with

a curve. Mark E_a upwards from the reactant line to the peak. Mark ΔH from the reactant line to the product line: downwards for exothermic and upwards for endothermic. Even if the curve rises and then falls, compare its starting and finishing levels to judge the overall heat change.

Keep activation energy separate from the net heat effect. Reaching the activated state requires energy input, so overcoming the barrier is an endothermic energy-input step. After the peak, however, more energy may be released, making the overall reaction exothermic. A requirement to absorb activation energy does not make the whole reaction necessarily endothermic.



The three reactions share the same illustrative energy scale and all begin at 40. Exothermic A has a smaller activation energy than endothermic B, while exothermic C has a larger one. Exothermic or endothermic labels alone cannot rank E_a for different reactions.

For example, on the supplied illustrative scale, reaction A has reactants at 40, products at 20 and a peak at 70. Draw both horizontal levels on the same vertical axis. The lower products show net energy release. Place the peak at 70 and connect the two ends.

A's activation energy runs from reactants to the peak:

$$E_a = 70 - 40 = 30$$

A's enthalpy change compares products with reactants:

$$\Delta H = 20 - 40 = -20$$

B also has reactants at 40, but its peak is 90 and products 60. Calculate the two differences in the same way:

$$E_a = 90 - 40 = 50$$

$$\Delta H = 60 - 40 = +20$$

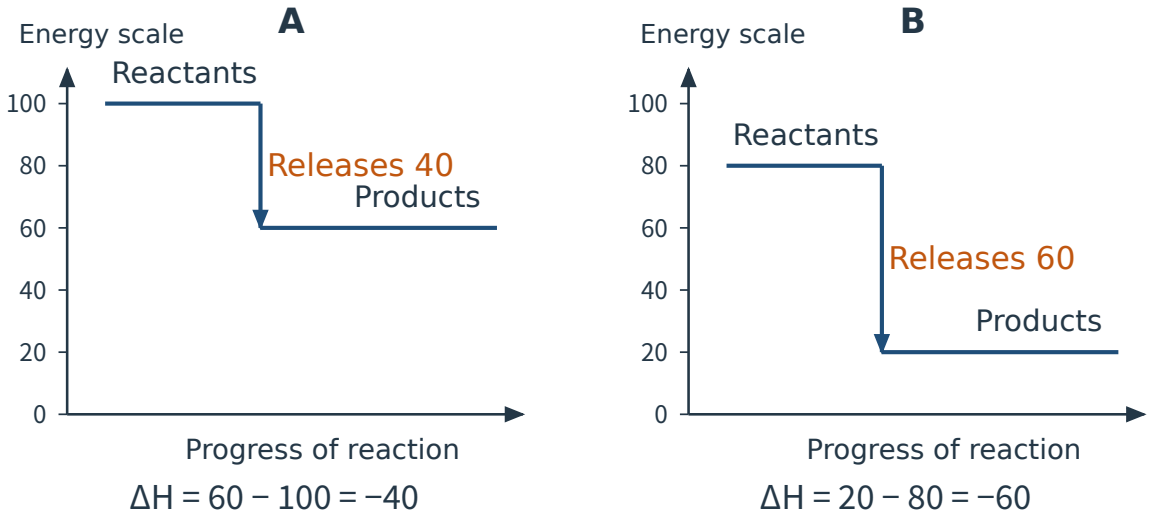
Although A is exothermic, its activation energy of 30 is below endothermic B's 50. C is also exothermic, but its activation energy of 70 is above B's. The claim that exothermic reactions always have greater activation energies than endothermic reactions is therefore false.

When naming a diagram's arrow, always check both endpoints. An arrow to the peak is activation energy; one to the product level is enthalpy change.

To compare temperature rises in the surroundings, also compare the net energy-level differences. For the same surrounding material, mass and heating conditions, more released heat causes a larger temperature rise. Vertical distances in different diagrams are directly comparable only if their energy axes have the same scale.

With the same energy scale, compare the reactant-to-product drop

Same energy scale and same conditions for heating the surroundings



$\Delta H = 60 - 100 = -40$ $\Delta H = 20 - 80 = -60$

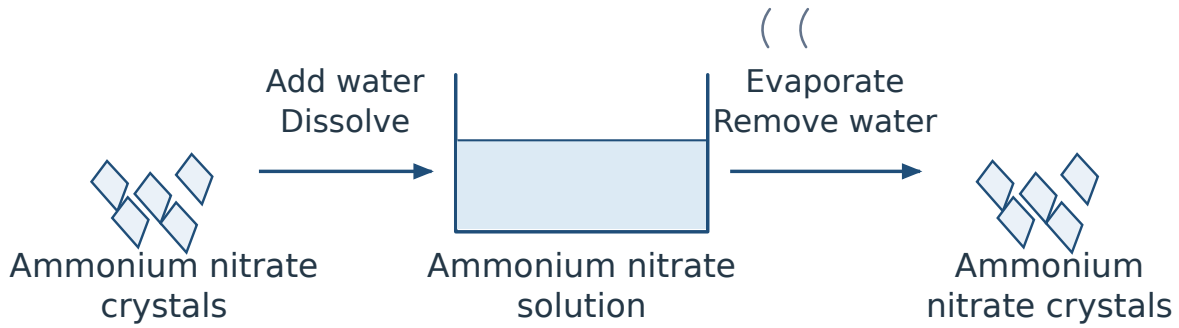
B has the larger net drop; the starting height alone does not determine the heat released.

Simple illustrative values show the net energy drops. Both diagrams use the same energy scale and the same heating conditions for the surroundings.

In the diagrams, A falls from 100 to 60 and B from 80 to 20. Calculate the energy released by each:
A: 100 − 60 = 40 illustrative energy units
B: 80 − 20 = 60 illustrative energy units
B starts at a lower reactant level than A but has a greater net drop. Under the same heating conditions shown, B therefore warms the surroundings more. Measure the reactant-to-product difference, not simply either level's distance from the horizontal axis.

Temperature changes also occur during processes such as dissolving, so endothermic alone does not establish a chemical reaction. Adding ammonium nitrate, NH_4NO_3, to water cools the mixture. Removing the water allows ammonium nitrate to be recovered. The first observation identifies an endothermic change; the second shows that the solid dissolved rather than all becoming a different substance in this experiment.

Evaporating the water recovers the original soluble solid



Recovery of the original substance supports the conclusion that dissolving occurred here.

Blocks represent visible crystal shapes. Individual molecules and ions in water are not shown. Temperature change is separate evidence for whether the process is endothermic.

If options combine the process and its heat effect, match the evidence separately: recovery of this original solid supports dissolves; cooling of the mixture supports endothermic. Both must occur in the same option.

An exothermic reaction makes the surroundings' temperature _____. Its products have a _____ energy level than its reactants, and ΔH has a _____ value. The difference from reactants to the peak represents _____.

Answers: rise; lower; negative; activation energy E_a

Recognising the Question Identify whether the question describes the surroundings' temperature or the reacting substances' energy. Find the starting and finishing levels on an energy diagram first. To compare temperature rises, also check the vertical scales and compare net energy drops.

Worked Example 1 (2025·May/June 0620/11 Q13) · Level 2

Three statements about energy changes in chemical reactions are listed.

1 In an endothermic reaction, the temperature of the surroundings increases.

2 In an exothermic reaction, thermal energy is taken in from the surroundings.

3 In the reaction pathway diagram for an exothermic reaction, the energy level of the products is lower than the energy level of the reactants.

Which statements are correct?

A 1 and 2 **B** 1 only **C** 2 and 3 **D** 3 only

Solution

The statements concern the surroundings' temperature, direction of heat flow and reactant/product energy levels. Identify what each is comparing.

Statement 1: an endothermic reaction absorbs heat from the surroundings. The surroundings lose heat and should cool, not warm, so 1 is incorrect.

Statement 2: an exothermic reaction transfers heat to the surroundings, not from them into the reaction, so 2 is incorrect.

Statement 3: an exothermic reaction transfers some energy to the surroundings, leaving the products at a lower energy level than the reactants. Statement 3 is correct.

Only 3 is correct, so choose D, 3 only.

Worked Example 2 (2025·May/June 0620/13 Q13) · Level 2

When ammonium nitrate is added to water, the temperature of the mixture decreases.

The ammonium nitrate can be recovered by evaporating the water.

Which statement explains these observations?

A The ammonium nitrate dissolves in the water, and the process is endothermic.

B The ammonium nitrate reacts with the water, and the process is endothermic.

C The ammonium nitrate dissolves in the water, and the process is exothermic.

D The ammonium nitrate reacts with the water, and the process is exothermic.

Solution

The question gives two different observations; explain each separately.

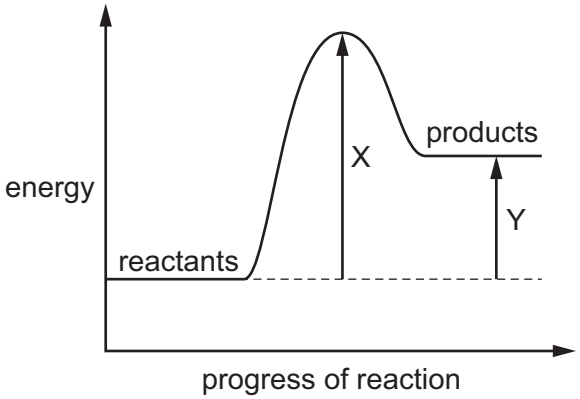
After ammonium nitrate is added, the mixture's temperature falls. The process therefore absorbs heat from the surroundings and is endothermic. Reject C and D, which say exothermic.

The second observation is that ammonium nitrate can be recovered after evaporating the water. In this experiment, the solid dissolves and is recovered when water is removed, so choose dissolves rather than reacts with the water. Of A and B, only A fits.

Together, the observations describe dissolving with an endothermic change. Choose A. A temperature change demonstrates energy transfer, but cooling alone does not identify the process as a chemical reaction.

Worked Example 3 (2025·May/June 0620/23 Q14) · Level 3

A reaction pathway diagram is shown.



Which row shows the correct labels for X and Y and the type of reaction taking place?

	X	Y	type of reaction
A	enthalpy change	activation energy	exothermic
B	activation energy	enthalpy change	endothermic
C	activation energy	enthalpy change	exothermic
D	enthalpy change	activation energy	endothermic

Solution

Check the endpoints of X first. It runs from the reactant level to the curve's peak, so it represents activation energy.

Y runs from the reactant level to the product level, so it represents enthalpy change.

Finally compare the two horizontal levels. Products are above reactants, so ΔH is positive and the overall reaction is endothermic. These assignments for X, Y and reaction type match B.

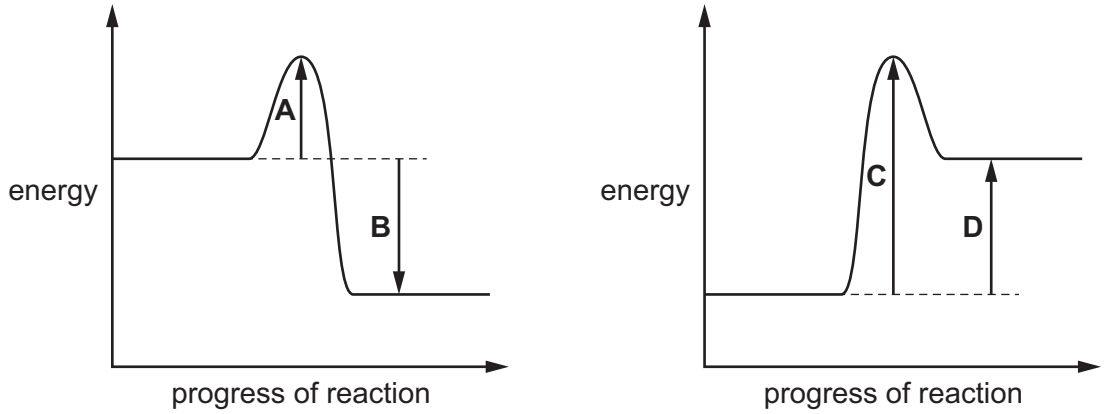
The descending section after the peak does not by itself make the reaction exothermic. Compare reactants with products to judge the net heat effect.

Practice

1. (2024·October/November 0620/21 Q14) · Level 2

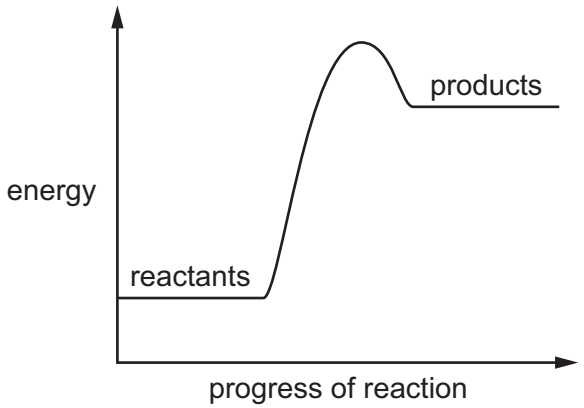
Two reaction pathway diagrams are shown.

Which arrow represents the activation energy for a reaction which releases thermal energy?



2. (2025·May/June 0620/22 Q13) · Level 3

A reaction pathway diagram is shown.



Which row identifies the type of reaction and how the temperature of the surroundings changes during the reaction?

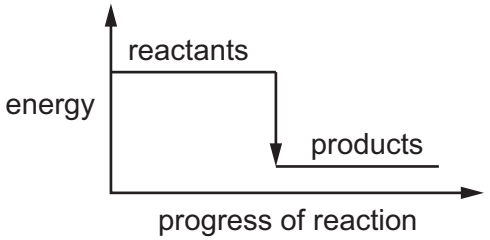
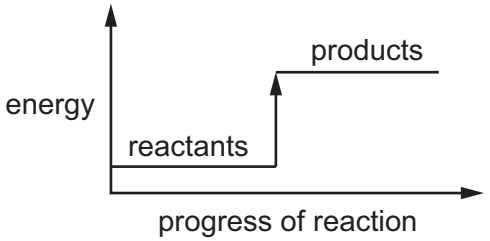
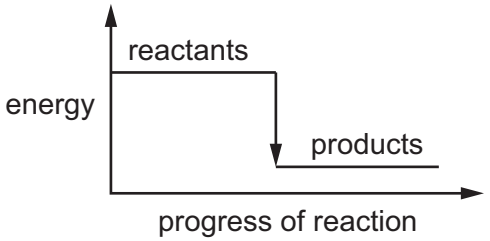
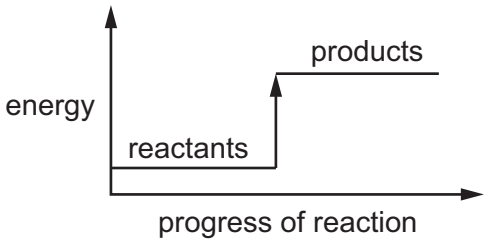
	type of reaction	temperature of the surroundings
A	endothermic	decreases
B	endothermic	increases
C	exothermic	decreases
D	exothermic	increases

3. (2025·May/June 0620/21 Q13) · Level 3

Water is added to anhydrous copper(II) sulfate in a test-tube.

The mixture becomes hot.

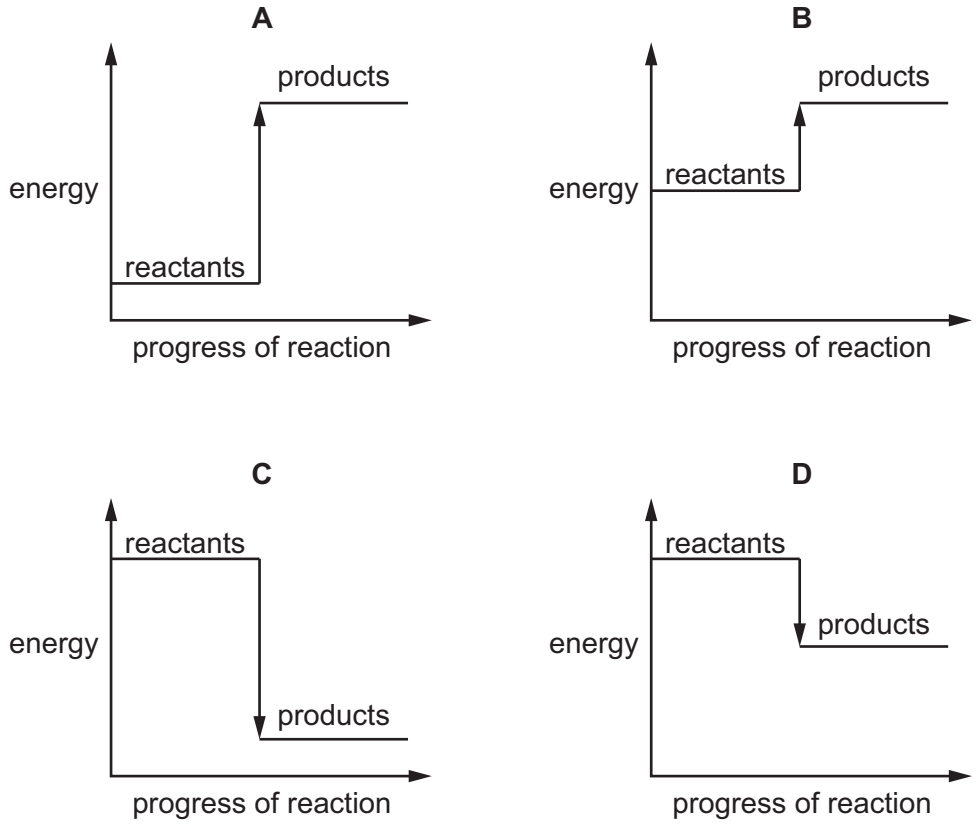
Which type of reaction and reaction pathway diagram apply to this reaction?

	type of reaction	reaction pathway diagram
A	endothermic	
B	endothermic	
C	exothermic	
D	exothermic	

4. (2025·May/June 0620/12 Q14) · Level 3

Which reaction pathway diagram shows the largest increase in the temperature of the surroundings?

(The scale on the *y*-axis is the same in each diagram.)



5. (2024·May/June 0620/22 Q12) · Level 3

Three statements about activation energy, E_a, are listed.

1 Colliding particles must have at least E_a before they can react.

2 E_a for exothermic reactions is always greater than for endothermic reactions.

3 E_a is always endothermic.

Which statements are correct?

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

Type 2: Calculating enthalpy change from bond breaking and bond formation

Key Ideas Separating the hydrogen and chlorine atoms in an HCl molecule requires the attraction between bonded atoms to be overcome. Breaking a chemical bond therefore absorbs energy. Conversely, forming a chemical bond lowers the system's energy and releases energy to the surroundings. Bond breaking is endothermic; bond formation is exothermic.

Breaking bonds absorbs energy; making bonds releases energy

Break the H—Cl bond



Form the H—Cl bond



For the same bond, breaking it requires the same amount of energy as forming it releases.

Horizontal arrows show bonded molecules changing to separate atoms, or separate atoms forming bonded molecules. The direction of energy transfer is labelled beside each arrow.

Breaking here means breaking the specified chemical bond, not increasing the distance between two molecules. To use bond energies for a molecular reaction's enthalpy change, list the bonds broken in the reactants and those formed in the products.

Bond energy is usually given in kJ/mol: the kilojoules required to break one mole of a specified bond in gaseous molecules. If average bond energies are supplied, calculate using those values. The value is per mole of that bond, which is not necessarily per mole of the molecule.

For example, each H_2 molecule has one H—H bond. Each CH_4 molecule has carbon connected to four hydrogens and therefore contains four C—H bonds. Thus 1mol of H_2 contains 1mol of H—H bonds, whereas 1mol of CH_4 contains 4mol of C—H bonds. Count the connections in each molecule first, then multiply by equation coefficients or the stated amount of substance.

Count bonds per molecule, then scale up to 1 mol

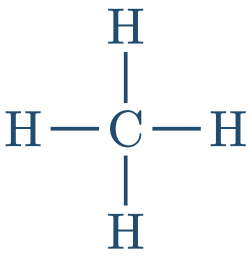
Hydrogen, H_2



Per molecule: 1 H—H bond

1 mol of H_2 contains 1 mol of H—H bonds

Methane, CH_4



Per molecule: 4 C—H bonds

1 mol of CH_4 contains 4 mol of C—H bonds

Multiply the number of molecules in the equation by the number of each bond per molecule.

Flat structural formulae show atom connections for counting bonds. The drawn angles are not used to determine three-dimensional molecular shape.

A reaction usually includes both bond breaking and bond formation. Breaking requires energy input and formation releases energy. Their difference gives the reaction's net energy change:

ΔH = total energy absorbed in breaking bonds − total energy released in forming bonds

If more energy is released, the result is negative and the reaction exothermic. If more is absorbed, the result is positive and the reaction endothermic.

For example, hydrogen reacts with chlorine to form hydrogen chloride:

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

The (g) symbols indicate gases. The supplied H—H, Cl—Cl and H—Cl bond energies are 436, 242 and 431kJ/mol respectively.

First find the energy absorbed in breaking bonds. For 1mol of H_2 and 1mol of Cl_2 as written, break 1mol each of H—H and Cl—Cl bonds:

Energy absorbed in bond breaking = 436 + 242 = 678 kJ

Next find the energy released in bond formation. The products contain 2mol of HCl, requiring 2mol of H—Cl bonds to form:

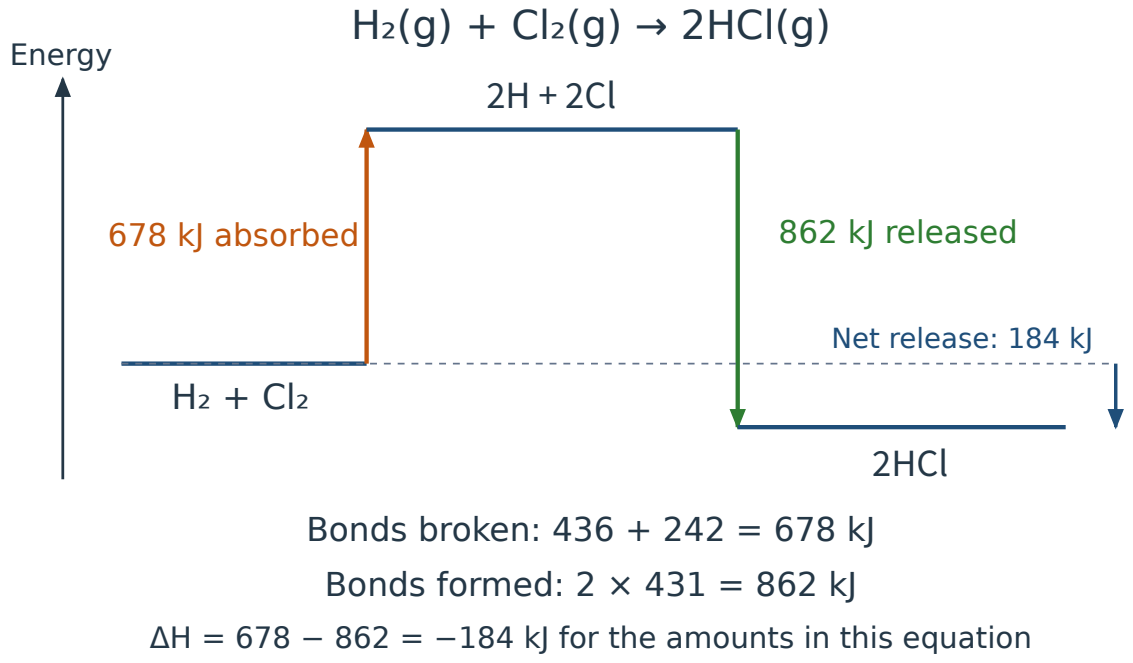
Energy released in bond formation = 2 × 431 = 862 kJ

Finally subtract energy released from energy absorbed:

$\Delta H = 678 - 862 = -184$ kJ

For these reactant amounts, the net energy released is therefore 184kJ.

Use bond breaking and bond making to calculate the net energy change



Bonds broken: 436 + 242 = 678 kJ

Bonds formed: 2 × 431 = 862 kJ

ΔH = 678 − 862 = −184 kJ for the amounts in this equation

Energy absorbed in bond breaking and released in bond formation is accounted for separately. The intermediate free atoms are an accounting device; the actual reaction need not first convert everything into separate atoms.

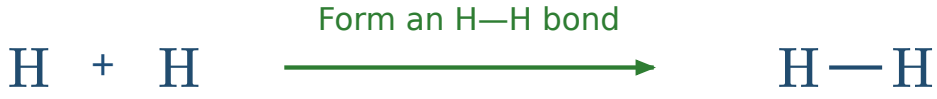
This calculation separates the energy contributions. It does not claim that every molecule must first become free atoms in the actual reaction. Nor is the total bond-breaking energy in the diagram the activation energy E_a from the previous question type. Activation energy measures the difference from reactants to the peak of the actual pathway. Bond-energy calculations concern the energies for breaking and forming the listed bonds.

If energy required and energy released are asked separately, give the magnitudes of those two amounts first. Only when calculating ΔH subtract released from absorbed and retain the sign. Match the requested units: total energy for specified amounts is in kJ; for a requested reaction enthalpy, follow the given equation and its kJ/mol requirement.

One starting condition needs separate attention. If molecules form directly from separate atoms, there are no corresponding starting bonds to break, so the bond-breaking term is zero. For example, 2mol of separate H atoms form 1mol of H_2, with H—H bond energy 436kJ/mol. No initial bonds need breaking. Forming 1mol of H—H bonds releases 436kJ:

$$\Delta H = 0 - 436 = -436 \text{ kJ}$$

Bond formation from separate atoms: bond-breaking energy = 0



2 mol of H atoms form 1 mol of H_2, making 1 mol of H—H bonds.

Bond breaking absorbs 0 kJ; bond making releases 436 kJ.

$\Delta H = 0 - 436 = -436$ kJ

The supplied H—H bond energy is 436kJ/mol. The calculation is for forming 1mol of H_2 from separate atoms.

Apply the same check when separate atoms form a molecule with several bonds: confirm the atomic starting point, then count the types and amounts of bonds formed. If the reactants are molecules such as H_2, include their original H—H bonds; the bond-breaking term is no longer zero.

Breaking bonds _____ energy; forming bonds _____ energy. To calculate net enthalpy change, take the total energy _____ and subtract the total energy _____.

Answers: absorbs; releases; absorbed in bond breaking; released in bond formation

Recognising the Question Check whether the reaction starts from atoms or molecules. Count bonds using structures and equation coefficients. Calculate energy absorbed in breaking bonds and released in forming bonds separately, then subtract released from absorbed and check the sign.

Worked Example 1 (2025·May/June 0620/41 Q6(d)(ii)) · Level 3

(ii) Iodine and bromine react at high temperatures to form iodine monobromide, IBr.

The equation is shown.

$$I_2(g) + Br_2(g) \rightarrow 2IBr(g)$$

The structures of the molecules involved in the reaction are I–I, Br–Br and I–Br.

Table 6.1

bond	bond energy in kJ / mol
I–I	150
Br–Br	193
I–Br	175

Calculate the enthalpy change, ΔH, for the reaction using the bond energies in Table 6.1.

Use the following steps.

- Calculate the **total** amount of energy required to break the bonds in 1 mol of $I_2(g)$ and 1 mol of $Br_2(g)$.
- Calculate the total amount of energy released when the bonds in 2 mol of IBr(g) are formed.
- Calculate the enthalpy change, ΔH, for the reaction.
  Your answer should include a sign.

Solution

Read the reaction first:

$I_2(g) + Br_2(g) \rightarrow 2IBr(g)$

1mol of I_2 contains 1mol of I—I bonds, and 1mol of Br_2 contains 1mol of Br—Br bonds. The product coefficient two means that 2mol of IBr forms, producing 2mol of I—Br bonds. The (g) symbols indicate gases.

Step 1: calculate the energy absorbed in bond breaking. The table gives I—I as 150kJ/mol and Br—Br as 193kJ/mol:

Energy absorbed in bond breaking = 1 × 150 + 1 × 193 = 343 kJ

Step 2: calculate the energy released in bond formation. Forming 1mol of I—Br bonds releases 175kJ. Here 2mol of these bonds form:

Energy released in bond formation = 2 × 175 = 350 kJ

This part asks for the magnitude of released energy, so give a positive number.

Step 3: calculate the net enthalpy change. Subtract energy released in formation from energy absorbed in breaking bonds, giving the requested kJ/mol:

$$\Delta H = 343 - 350 = -7 \text{ kJ/mol}$$

The 350kJ released exceeds the 343kJ absorbed, giving a net release of 7kJ. The reaction is exothermic and the final negative sign must remain.

Do not add the first two results. Do not divide the final answer by two merely because the product coefficient is two: the calculation is for the equation as supplied.

Practice

1. (2025·May/June 0620/43 Q6(e)(iii)) · Level 3

(iii) Iodine and chlorine react at high temperatures to form iodine monochloride, ICl.

The equation for the reaction is shown.

$$I_2(g) + Cl_2(g) \rightarrow 2ICl(g)$$

The structures of the molecules involved in the reaction are I–I, Cl–Cl and I–Cl.

Table 6.1

bond	bond energy in kJ / mol
I–I	150
Cl–Cl	242
I–Cl	218

Calculate the enthalpy change, ΔH, for the reaction using the bond energies in Table 6.1.

Use the following steps.

- Calculate the **total** amount of energy required to break the bonds in 1 mol of $I_2(g)$ and 1 mol of $Cl_2(g)$.
- Calculate the total amount of energy released when the bonds in 2 mol of $ICl(g)$ are formed.
- Calculate the enthalpy change, ΔH, for the reaction.
  Your answer should include a sign.

2. (2025·May/June 0620/22 Q14) · Level 3

The average bond energy for the C–H bond is 413 kJ / mol.

What is the enthalpy change when 1.0 mol of methane molecules is formed from carbon and hydrogen atoms?

A –1652 kJ **B** –413 kJ **C** +413 kJ **D** +1652 kJ

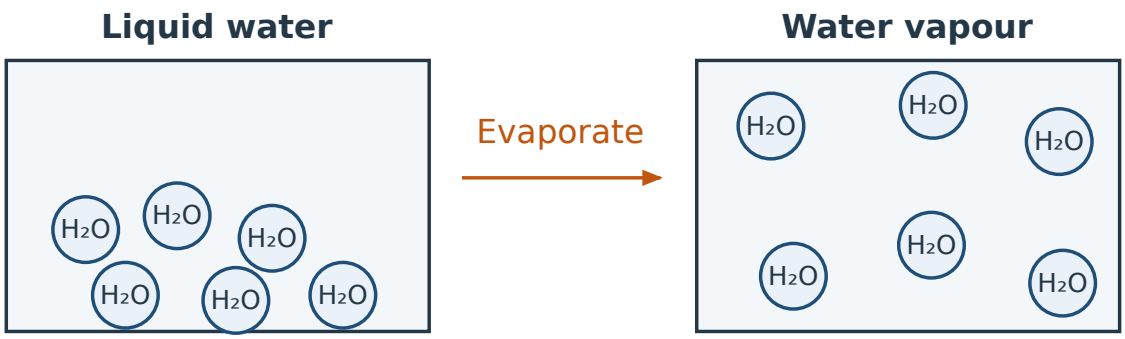
Chemical reactions

6.1 Physical and chemical changes

Type 1: Distinguishing physical and chemical changes by substance identity

Key Ideas Heating water to form water vapour, or cutting a solid into smaller pieces, changes its state, shape or distribution without necessarily producing a different substance. Such a change is called a physical change. First ask what the substance is before and after the change, then consider whether its appearance has changed.

Physical change: evaporated water molecules are still H_2O



All six symbols represent water molecules before and after; only their state and distribution change.

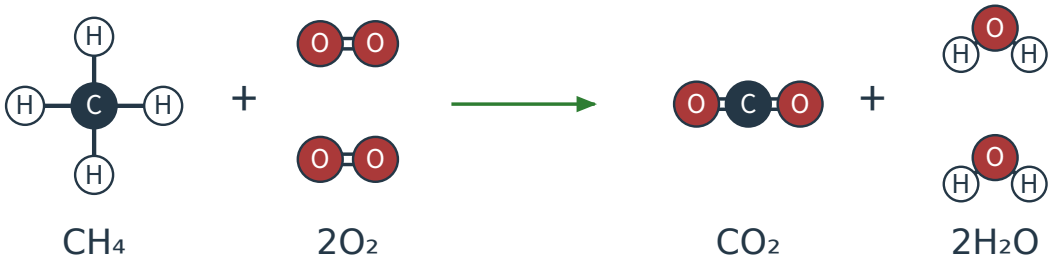
The diagram uses the same six H_2O labels to compare a liquid and a gas. Gas particles are further apart on average, but their identity has not changed. Particle sizes and separations are not drawn to scale.

For example, water changes from $H_2O(l)$ to $H_2O(g)$. The state symbols (l) and (g) mean liquid and gas respectively. The formula is H_2O before and after the change, so this is a physical change. Cooling water vapour to form liquid water is also a physical change. Distillation involves evaporating a liquid and then condensing the vapour, so it also uses changes of state. For example, when ethanol, C_2H_5OH, is distilled, the substance collected is still ethanol.

For dissolving, check whether the particular substances form new reaction products. In this section, dissolving sodium chloride or copper(II) sulfate in water, and mixing ink with water, are treated as physical changes: the original substances become dispersed in the water, and suitable separation methods can recover the solute. The disappearance of a visible solid alone does not prove that a chemical reaction has occurred.

A chemical change produces new substances. For example, complete combustion of methane, CH_4, in oxygen, O_2, produces carbon dioxide, CO_2, and water, H_2O. Their composition and properties differ from those of the original methane and oxygen. The original atoms are still present, but they have been rearranged into different combinations.

Chemical change: atoms rearrange to form new substances



C: 1 → 1; H: 4 → 4; O: 4 → 4 atoms.

The particle model for methane combustion represents $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. The types and numbers of atoms are conserved, but methane and oxygen have become carbon dioxide and water. The arrangement does not show the actual bond angles in methane.

For example, count the 1 C, 4 H and 4 O atoms on the left of the diagram, then count each type on the right. The totals are the same. However, CH_4 and O_2 on the left have become CO_2 and H_2O on the right. New substances have formed, so methane combustion is a chemical change. Rusting of iron, Fe, also produces new substances and is therefore a chemical change.

Neutralisation also produces new substances. For example, black copper(II) oxide, CuO, reacts with dilute sulfuric acid, H_2SO_4, to form copper(II) sulfate, $CuSO_4$, and water:

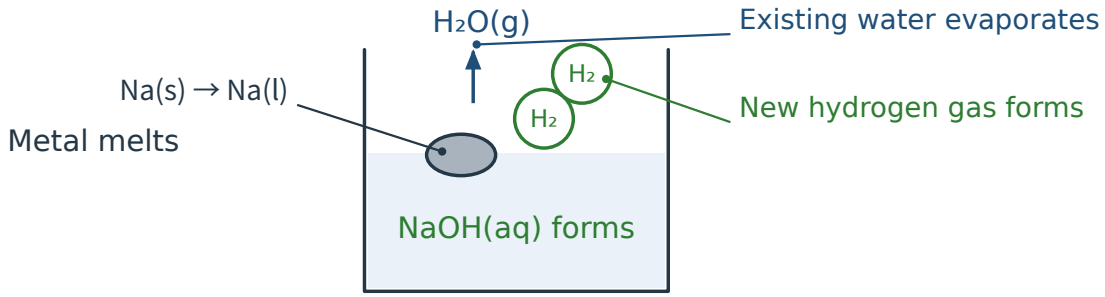
$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

The CuO has been converted into a salt and water, rather than simply transferred unchanged into the water. This is therefore a chemical change.

Bubbles, a colour change or a temperature rise may indicate a reaction, but no single observation replaces the test of whether new substances have formed. Boiling water also produces bubbles, and adding water to a coloured solution makes its colour paler. Physical changes can usually be reversed by physical methods, but reversibility is not the sole criterion, since some chemical reactions are themselves reversible.

One experiment may involve both types of change. The Group I metal sodium, Na, reacts with water to form sodium hydroxide, NaOH, and hydrogen, H_2. The heat produced may also melt the metal and evaporate some of the water already present. The pH is a numerical measure of how acidic or alkaline a solution is. NaOH makes the initially near-neutral solution alkaline, so its pH rises.

Physical and chemical changes can be observed during the same reaction



$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Water vapour is still H_2O; hydrogen is a different substance, H_2.

When sodium reacts with water, melting of the metal and evaporation of water already present are physical changes; formation of hydrogen and sodium hydroxide are chemical changes. The blue arrow only shows water vapour leaving the surface, not water decomposing into hydrogen.

For example, the first observation, “the metal melts”, describes Na(s) changing to Na(l), a physical change. The second, “H_2 is produced”, identifies a new substance and therefore a chemical change. The third, “water vapour forms”, describes $H_2O(l)$ changing to $H_2O(g)$, another physical change. In this experiment, the pH rises because the reaction forms alkaline NaOH. However, not every pH change automatically indicates a chemical change: dilution with water may also change pH. When several observations are listed, identify what changes in each one before combining the correct choices.

Answers: A physical change produces no new substances; a chemical change produces new substances. In a single experiment, assess melting, evaporation and the formation of new substances separately.

Recognising the Question For each observation, ask whether the original substance has only changed state or distribution, or whether a new substance has formed. Then combine the selected observations. Bubbles or a pH change alone are not universal tests.

Worked Example 1 (2025·May/June 0620/12 Q15) · Level 1

Which process involves a chemical change?

A dissolving copper(II) sulfate

B distilling ethanol

C freezing water

D neutralising copper(II) oxide

Solution

Check the substances before and after each of the four changes.

In A, copper(II) sulfate dissolves in water. The original solute is dispersed in the water; this question treats it as a physical change.

In B, distilling ethanol involves evaporation followed by condensation. The substance remains ethanol.

In C, water freezes. H_2O changes from liquid to solid but is still water.

In D, copper(II) oxide is neutralised by an acid, forming a salt and water. For example:

$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

New substances form in this reaction, so D is the chemical change.

Practice

1. (2025·May/June 0620/32 Q7(a)) · Level 2

(a) Select **two** processes that show a chemical change.

Tick (✓) **two** boxes.

boiling water	☐
burning methane	☐
dissolving salt	☐
mixing ink and water	☐
rusting of iron	☐

2. (2025·May/June 0620/11 Q14) · Level 2

When a small piece of a Group I metal is placed into a large beaker of cold water, a reaction occurs.

Four statements about this reaction are listed.

1 The metal melts.

2 Hydrogen is produced.

3 Steam is produced.

4 The pH of the solution increases.

Which statements about this reaction describe a physical change?

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

6.2 Rate of reaction

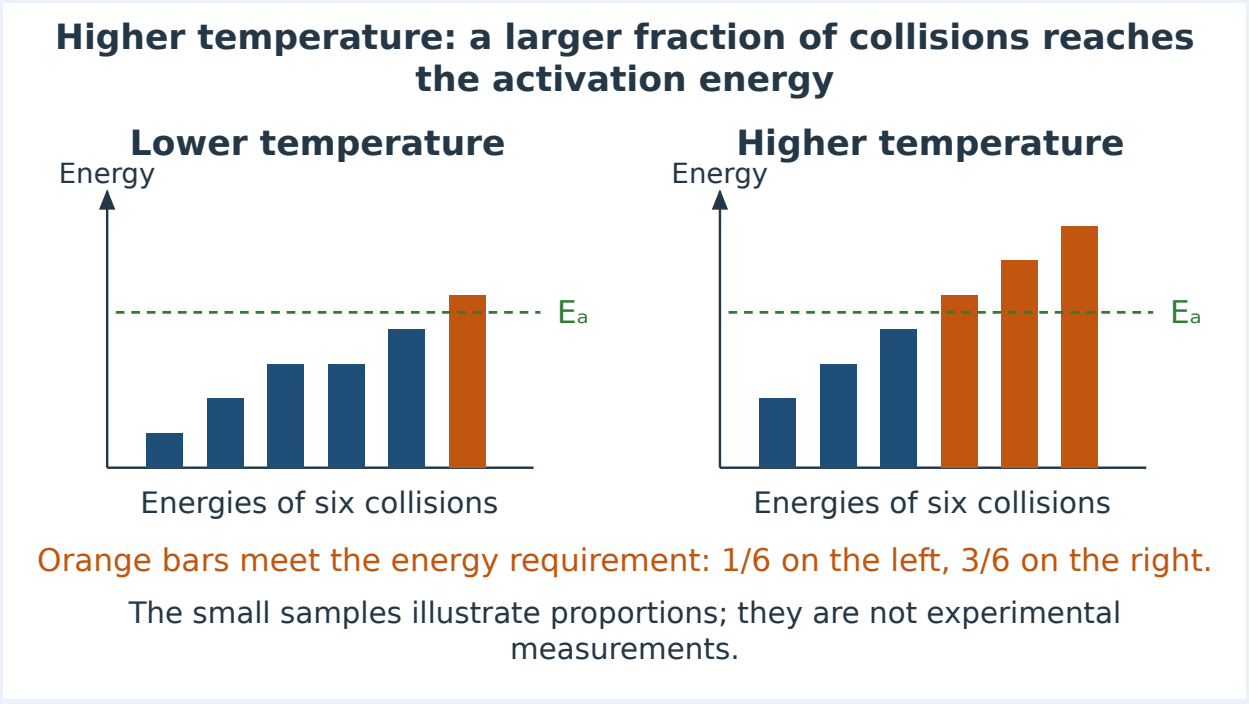
Type 1: Explaining reaction rate using collision frequency and successful collisions

Key Ideas The rate of reaction describes how quickly a reaction proceeds. The more product formed, or reactant consumed, in a given time, the greater the average rate. For example, when collecting a gas released by a reaction, a greater increase in the total gas volume over the same time interval indicates a faster reaction during that interval.

A particle explanation must distinguish two questions: how often particles collide, and how many of those collisions cause a reaction. Collision frequency is the number of collisions per unit time. Successful collisions are collisions that actually cause a chemical reaction. A collision must have at least the activation energy, E_a, to meet the energy requirement for reaction. After many low-energy collisions, the particles simply separate without forming products.

First consider temperature. Kinetic energy is the energy particles have because they are moving. At a higher temperature, particles have greater average kinetic energy and move faster on average. This increases the number of collisions per unit time. Particle energies are not all the same, however: increasing temperature also increases the proportion of particles or collisions that reach or exceed E_a. The diagram uses two hypothetical sets of

collisions to explain “proportion”. Each bar represents the energy of one collision; bars crossing the same E_a line meet the energy requirement.

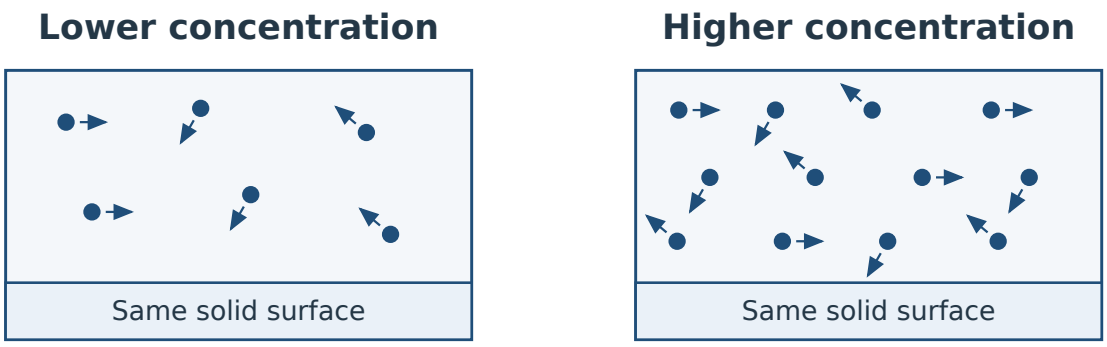


Both diagrams use the same schematic energy scale and the same E_a line. Each bar represents the energy of one collision. The numbers of bars illustrate proportions; they are not used to compare collisions per second.

For example, the temperature of a reaction rises from 30°C to 50°C, with all other conditions unchanged. First, the higher temperature increases average kinetic energy, so particles move faster. This increases collision frequency. Also state that a greater proportion reaches the activation energy. There are therefore more successful collisions per unit time, and the reaction is faster. If the question instead compares 30°C with 50°C, average kinetic energy and collision frequency are lower at 30°C. Explain both effects explicitly; “the particles are more active” is insufficient.

Next consider solution concentration. A higher concentration means more solute particles in the same volume. If these particles take part in the reaction, they collide with the other reactant more frequently when temperature and other conditions are unchanged. For an acid reacting with a solid, compare the number of acid particles per unit volume and their frequency of collision with the solid surface.

Higher solution concentration: more reacting particles in the same volume



Same temperature, box volume and solid surface area; blue dots represent reacting particles.

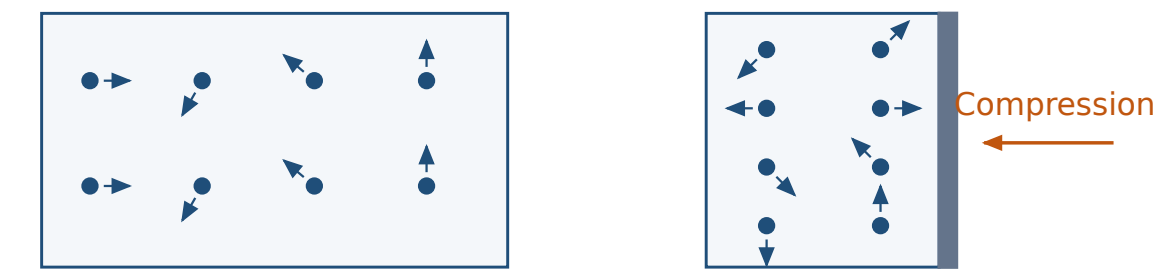
Higher concentration → more particles per unit volume → more frequent collisions with the solid

Equal volumes of two solutions contact identical solid surfaces. Short arrows show particle motion. At a fixed temperature the arrows have the same length; they do not imply that increasing concentration makes particles move faster.

For example, a more concentrated acid contacts the same solid powder. First identify “more reacting particles per unit volume”, then explain “more frequent collisions with the solid surface”, and conclude “a faster reaction”. The temperature has not changed, so average kinetic energy does not increase simply because concentration increases. Doubling concentration also does not automatically prove that the rate doubles exactly; here we are identifying whether the reaction is faster or slower and explaining why.

Gas pressure can also be explained using the number of particles per unit volume. Compressing the same sample of reacting gas at constant temperature reduces its volume. The total number of gas particles has not increased, but they occupy less space. There are more particles per unit volume, collisions are more frequent, and the rate increases. Conversely, decreasing the pressure of reacting gases, with other conditions unchanged, usually spreads the particles out and slows the reaction.

Compressing a gas at constant temperature puts its particles in a smaller volume



Same gas sample and temperature; the containers have the same height and depth.

Higher pressure → more particles per unit volume → more frequent collisions

Both sides contain the same number of gas particles, but the volume on the right is smaller. At the same temperature, this difference in volume does not itself increase average kinetic energy.

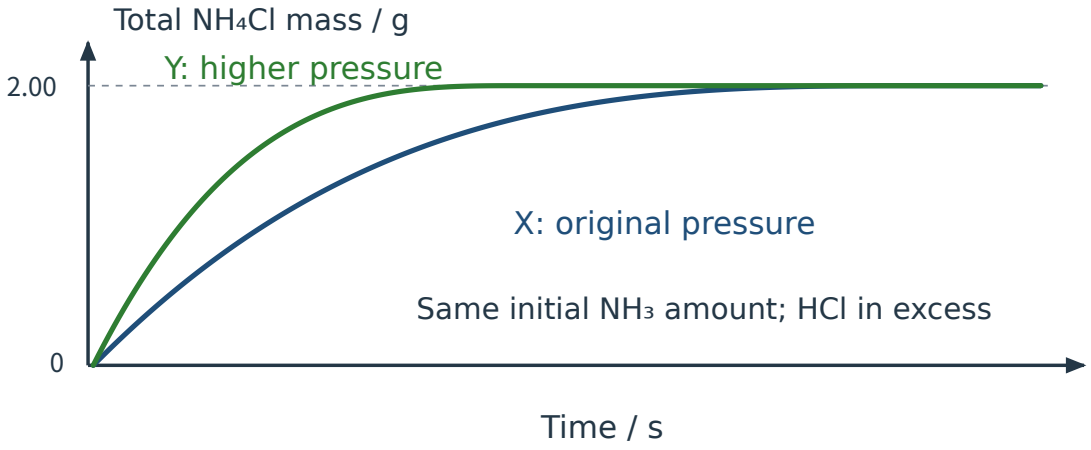
For example, the volume of the same gas sample is halved at constant temperature. Keep the total particle number unchanged, then compare the volumes: the number of particles per unit volume increases, so collision frequency rises. The change concerns how closely particles are packed; average kinetic energy still depends on temperature. This pressure explanation applies to gases. It does not imply that compressing a solid necessarily speeds up every reaction.

Increasing pressure does not mean increasing the original total amount of reactant. Ammonia, NH_3, and hydrogen chloride, HCl, can form solid ammonium chloride, NH_4Cl:

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

If each experiment starts with the same amount of NH_3 and HCl remains in excess, that amount of NH_3 determines the total NH_4Cl formed when the reaction is complete.

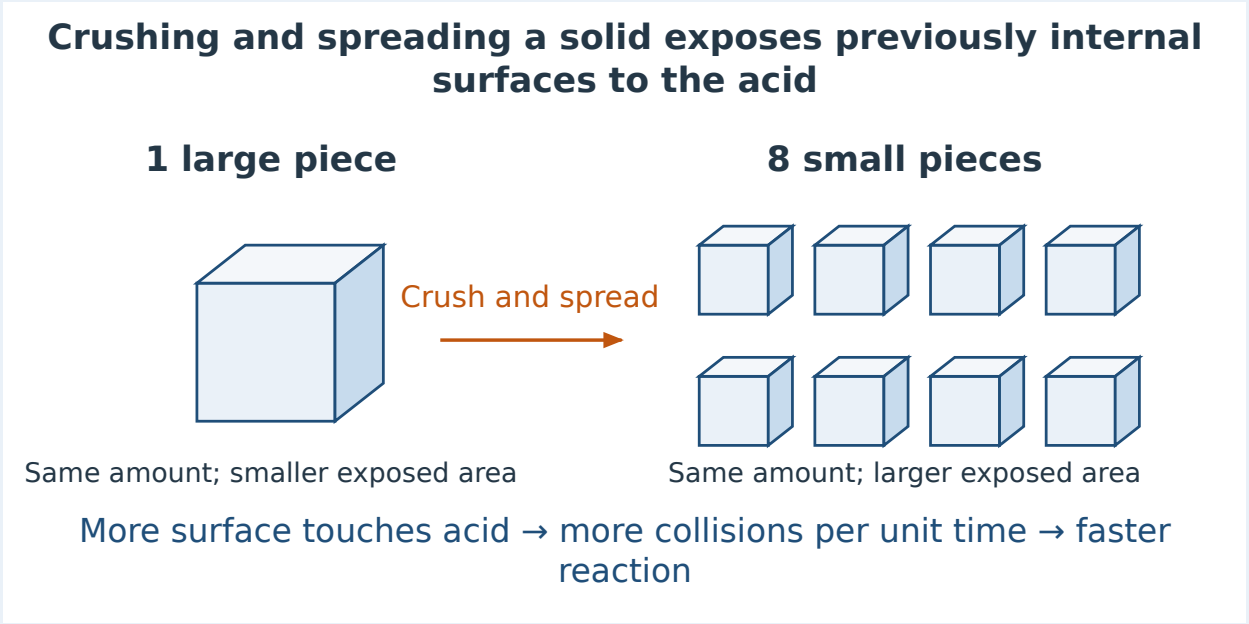
Higher gas pressure speeds the reaction; the same ammonia gives the same final NH_4Cl mass



The shared endpoint of 2.00 g illustrates the relationship between the curves. Temperature, the initial amount of ammonia and other conditions are unchanged; only pressure increases. Hydrogen chloride remains in excess, so higher pressure makes the plateau arrive sooner without increasing the final mass of ammonium chloride after complete reaction.

For example, only the pressure is increased. More particles per unit volume cause more frequent collisions, so the initial part of Y is steeper than X and reaches its plateau sooner. Then check the final amount: the initial NH_3 amount is unchanged and HCl is not limiting, so both plateaus have the same height. The chosen graph must show both “faster” and “the same endpoint”. A higher curve alone is insufficient. Here the vertical axis is the cumulative mass of solid product, not the volume of the compressed reacting gases.

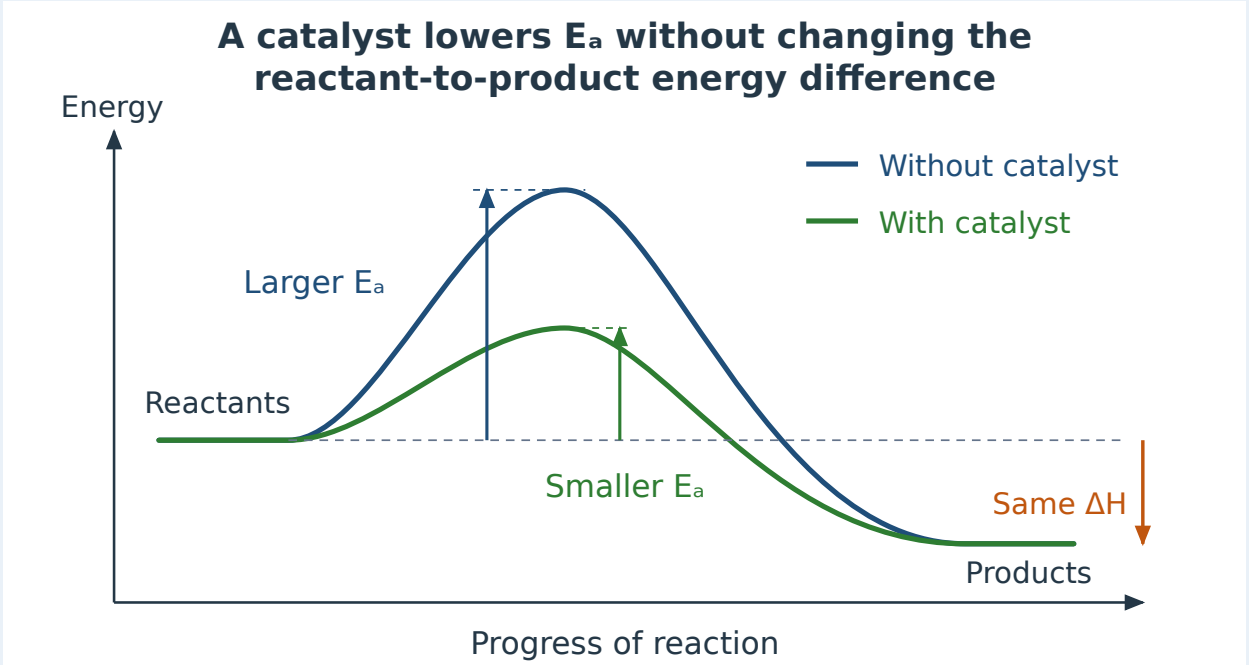
When a solid reacts with a solution, solution particles first encounter the solid surface. Breaking up and separating the same total amount of solid exposes surfaces that were previously inside the larger piece. The total exposed surface area increases. More reacting particles can therefore collide with the solid surface per unit time, making the reaction faster.



Each of the eight small cubes has half the edge length of the large cube. Their combined volume equals that of the large cube. The diagram compares exposed surface areas only; it does not establish an exact multiple for the increase in a real reaction rate.

For example, replace a powder with large pieces of the same total amount of solid, keeping acid concentration and temperature unchanged. Read the diagram in reverse: large pieces have a smaller total exposed surface area, so acid particles collide with the surface less often per unit time and the reaction is slower. There is less surface in contact with the acid. Temperature and the total amount of solid are unchanged; particles inside the solid have not become fewer or slower.

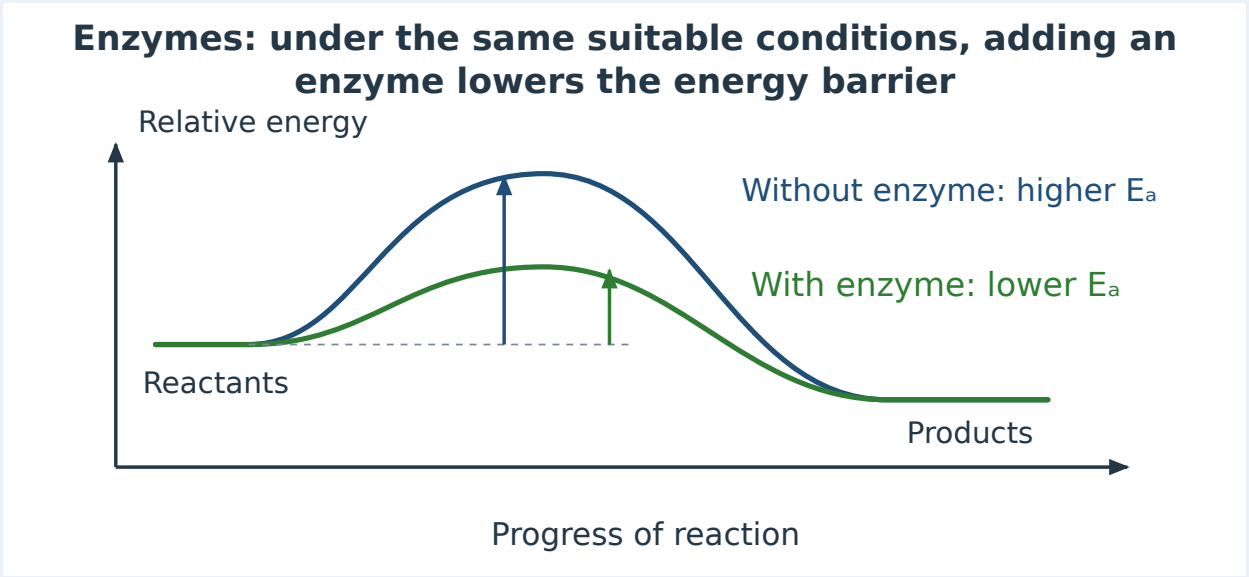
A catalyst increases the rate of a reaction and remains unchanged at the end. It provides an alternative reaction pathway with a lower activation energy. At the same temperature, more collisions meet this lower energy requirement, so the reaction is faster. Removing the catalyst removes the lower-activation-energy pathway it provides.



The same reaction is compared at the same temperature. The catalysed pathway has a lower peak, while the reactant and product energy levels are respectively identical for both pathways. This example is exothermic; the horizontal axis is not reaction time.

For example, measure from the reactant energy level to the peak of each curve. The difference is smaller with a catalyst, showing a lower E_a. Then compare the height difference between reactants and products: it is identical for both pathways, so ΔH is unchanged. The catalyst has not heated the particles. At constant temperature, its effect cannot be explained as an increase in average kinetic energy.

Enzymes are biological catalysts and also lower the activation energy of the reactions they catalyse. Adding a suitable enzyme under suitable conditions can speed up its reaction; the enzyme is not used up as a reactant. Excessive temperature may cause an enzyme to lose the structure it needs to work and become inactive. Therefore, the usual temperature explanation cannot be extended indefinitely to claim that an enzyme-catalysed reaction always gets faster at every higher temperature.



The same reaction is compared with identical temperature, pH and initial reactant conditions, all suitable for the enzyme. The enzyme provides a lower-activation-energy pathway. It changes neither the energy difference between reactants and products nor the average kinetic energy at the same temperature.

For example, two tubes contain equal volumes of H_2O_2 at the same concentration, temperature and pH, with conditions suitable for the enzyme. Catalase, an enzyme that catalyses its decomposition, is added to one tube but not the other. The tube with catalase reacts by a lower-E_a pathway. At the same temperature, a greater proportion of collisions meets this energy requirement, so more O_2 forms per unit time. The cause is neither higher average kinetic energy nor conversion of the enzyme itself into O_2. Starting under the same conditions without this catalyst gives a slower reaction, but the reaction need not stop completely. If all the H_2O_2 in both tubes eventually decomposes, the total amount of O_2 is the same. When assessing enzyme removal, keep other conditions constant and exclude effects caused by simultaneous changes in temperature or pH.

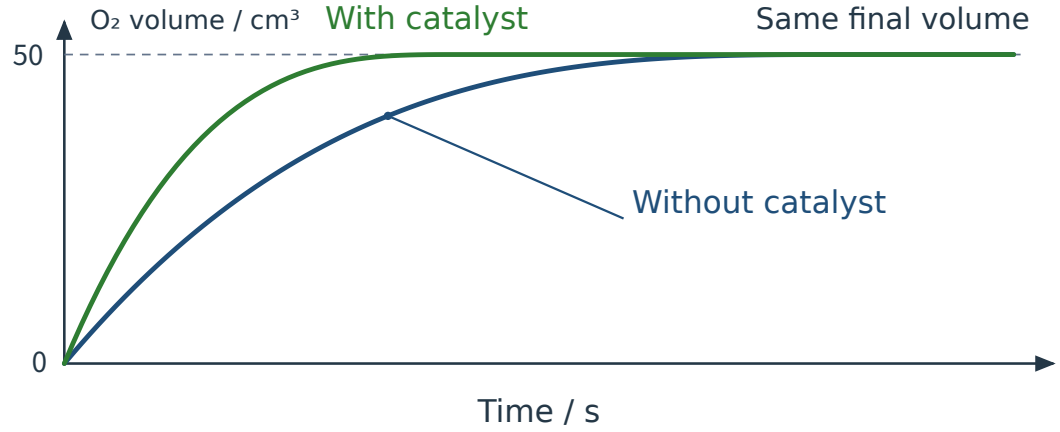
Finally, distinguish a faster reaction from a greater final amount of product. Hydrogen peroxide has the formula H_2O_2 and decomposes to form water, H_2O, and oxygen, O_2:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

With a catalyst, oxygen forms more quickly. If the same initial amount of H_2O_2 decomposes completely in each experiment, the final oxygen volume is the same when measured under identical gas conditions.

A catalyst reaches the same final amount sooner

Same initial H_2O_2 amount, temperature and gas measurement conditions



Complete decomposition of the same amount of H_2O_2 gives the same final amount of O_2.

The catalyst is unchanged after the reaction; its mass is unchanged if completely recovered.

These curves illustrate the relationship, with a hypothetical shared endpoint of 50 cm³. A steeper initial curve means oxygen forms faster. Equal final amounts do not mean equal amounts have been collected at every earlier time.

For example, both curves eventually reach the hypothetical 50 cm³ of oxygen. The catalysed curve reaches its plateau sooner, so the reaction is faster. The plateau heights are equal, so the final yield has not increased. The catalyst remains unchanged at the end; if completely recovered, its mass is also unchanged. Check catalyst statements in this order: rate increases, E_a decreases, ΔH is unchanged, the final yield from complete reaction of the same reactants is unchanged, and the catalyst remains unchanged afterwards.

If an equation contains the reversible arrow $\rightleftharpoons$, the forward reaction proceeds from the substances on the left to those on the right. For example, C_2H_4 is ethene and C_2H_5OH is ethanol. $C_2H_4(g)$ and $H_2O(g)$ can react towards the right to form $C_2H_5OH(g)$; (g) means gaseous. To explain why heating speeds up the forward reaction, still use average kinetic energy, collision frequency and the proportion reaching the activation energy. There is no need first to calculate the final yield.

Increasing temperature causes average kinetic energy to _____, collision frequency to _____, and the proportion reaching the activation energy to _____. A catalyst causes activation energy to _____, but the enthalpy change of the same reaction _____.

Answers: increase; increase; increase; decrease; remains unchanged

Recognising the Question Identify the single condition changed. Explain whether it changes particles per unit volume, exposed surface area, average kinetic energy or activation energy. Link this to collision frequency or the proportion of successful collisions, then explain the rate.

Worked Example 1 (2025·May/June 0620/11 Q15) · Level 2

The equation for the decomposition of hydrogen peroxide is shown.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

The reaction is exothermic.

When a small amount of a catalyst is added, the oxygen is produced more quickly.

Which statement about the catalyst is correct?

A The catalyst makes the reaction more exothermic.

B The mass of catalyst is the same before and after the reaction.

C The catalyst increases the final volume of oxygen produced.

D All of the catalyst is used up in the reaction.

Solution

First read the reaction: hydrogen peroxide, H_2O_2, decomposes to form water, H_2O, and oxygen, O_2. The question states that the reaction is exothermic and that adding a little catalyst produces oxygen more quickly. This change concerns reaction rate.

A catalyst provides a pathway with a lower activation energy and remains unchanged at the end. If completely recovered, its mass is the same before and after the reaction. Therefore B is correct.

A confuses rate with the amount of heat released. A catalyst does not change the energy difference between reactants and products, so it does not make the same reaction more exothermic. C confuses reaching the endpoint sooner with reaching a higher endpoint. Complete decomposition of the same initial amount of hydrogen peroxide gives the same final amount of oxygen under identical gas measurement conditions. D says the catalyst is completely used up, contradicting its remaining unchanged after the reaction. Choose B.

Worked Example 2 (2025·May/June 0620/41 Q3(c)(vi)) · Level 2

(c) Copper(II) sulfate crystals are made by the reaction between copper(II) carbonate and dilute sulfuric acid, using the following steps.

The sulfuric acid has a concentration of $0.100\,mol/dm^3$.

step 1 Powdered copper(II) carbonate is added to dilute sulfuric acid. The mixture is stirred. A reaction occurs.

(vi) **Step 1** is repeated using sulfuric acid of concentration $0.200\,mol/dm^3$ instead of $0.100\,mol/dm^3$.

All other conditions are the same.

The rate of reaction increases.

Explain why the rate of reaction increases. Give your answer in terms of particles.

Solution

Identify the single change: sulfuric acid concentration increases from 0.100 mol/dm³ to 0.200 mol/dm³, while

temperature, copper(II) carbonate powder and other conditions remain the same. The comparison concerns concentration; the solution is not being heated.

A higher concentration means more reacting acid particles in the same volume of solution. More particles near the copper(II) carbonate surface can react with it, so more collisions occur per unit time: collision frequency increases.

With other conditions unchanged, there are also more successful collisions per unit time, so copper(II) carbonate reacts with the acid faster. A complete answer is: there are more reacting particles per unit volume, increasing the frequency of collisions between acid particles and the copper(II) carbonate surface, so the reaction rate increases.

Worked Example 3 (2025·May/June 0620/42 Q6(d)(ii)) · Level 3

(d) C_2H_4 reacts with steam to form ethanol.

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g) \qquad \Delta H = -45\,kJ/mol$$

(ii) Explain, in terms of collision theory, why the rate of the forward reaction increases if the temperature increases.

Solution

The forward reaction proceeds from left to right: ethene, C_2H_4, reacts with steam, H_2O, to form ethanol, C_2H_5OH. This question asks why heating makes that direction faster, rather than how much ethanol is finally present at equilibrium.

First, a higher temperature increases the average kinetic energy of the reacting particles, so they move faster on average.

Second, faster movement causes more collisions per unit time, so collision frequency increases. This explains the increase in the number of collisions.

Third, heating also increases the proportion of particles or collisions with enough energy to reach or exceed the activation energy, E_a. Thus a greater proportion of collisions meets the energy requirement for reaction. Writing only “more collisions” omits this point.

There are therefore more successful collisions per unit time, and the forward reaction rate increases. A complete explanation includes increased average kinetic energy, increased collision frequency and a greater proportion reaching the activation energy. Heating does not itself lower the activation energy of the same reaction pathway.

Worked Example 4 (2024·May/June 0620/12 Q17) · Level 3

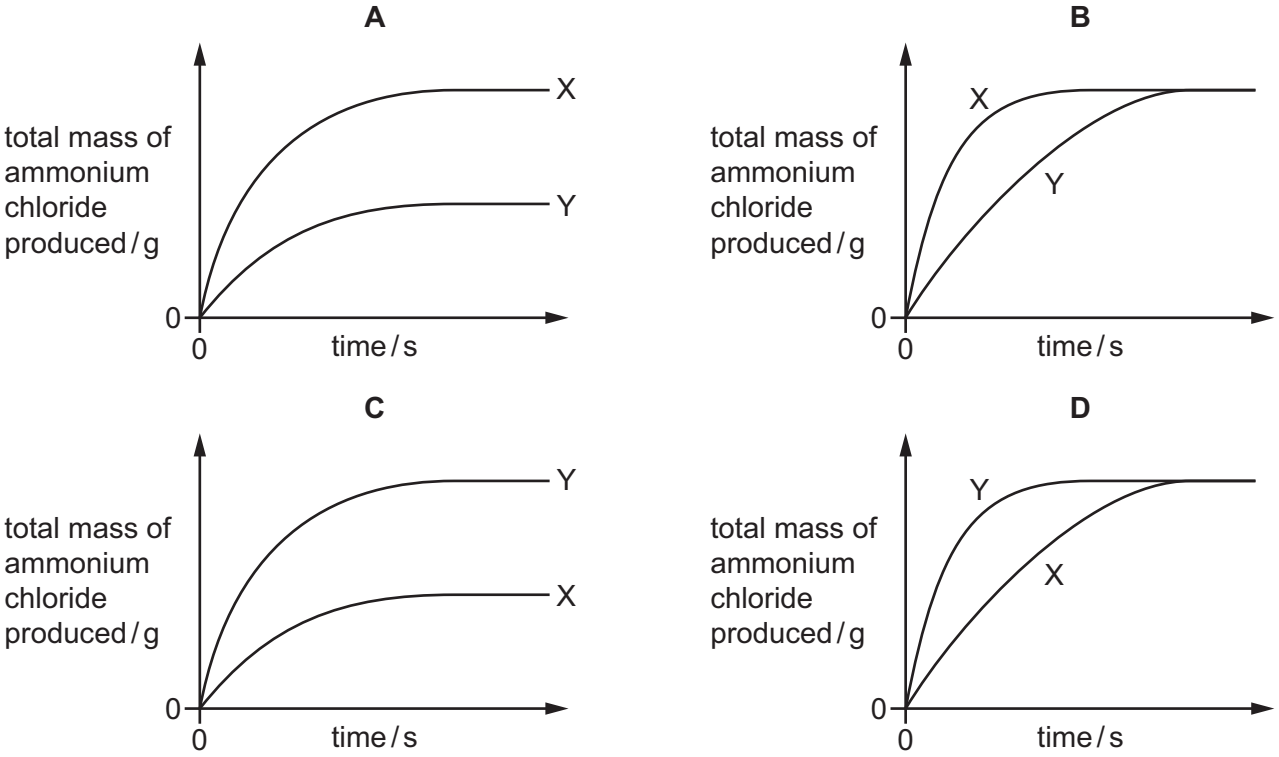
A known mass of gaseous ammonia and excess gaseous hydrogen chloride react together to make solid ammonium chloride.

Line X shows the total mass of ammonium chloride produced over time.

The reaction is repeated at a higher pressure. All other conditions are kept the same.

Line Y shows the total mass of ammonium chloride produced over time at the higher pressure.

Which diagram is correct?



Solution

First assess the rate. Increasing gas pressure at the same temperature and other conditions gives more NH_3 and HCl particles per unit volume and more frequent collisions. The higher-pressure curve Y should therefore rise more steeply at first.

Then assess the final amount. The initial mass of NH_3 is unchanged and HCl remains in excess. Complete reaction of the NH_3 therefore produces the same mass of NH_4Cl, so X and Y must reach plateaus at the same height.

A gives Y a smaller final amount and C gives it a larger one; both are incorrect. B has equal plateaus but shows Y reacting more slowly.

Only D shows Y reacting faster with the same final mass. Choose D.

Practice

1. (2025·May/June 0620/22 Q15) · Level 2

Which statement describes the effect of adding a catalyst to a chemical reaction?

A The activation energy, E_a, of the reaction is increased.

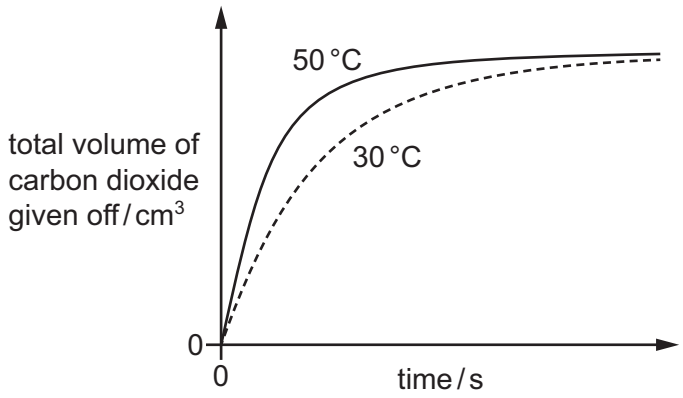
B The enthalpy change, ΔH, of the reaction stays unchanged.

C The frequency of collisions between the particles is decreased.

D The kinetic energy of the particles is increased.

2. (2025·May/June 0620/21 Q15) · Level 2

The graph shows the results of adding excess calcium carbonate to hydrochloric acid at 50 °C and at 30 °C. Only the temperature is changed.



Which row describes the reacting particles at 30 °C compared to those at 50 °C?

	frequency of collisions between particles	average kinetic energy of particles
A	higher	higher
B	higher	lower
C	lower	higher
D	lower	lower

3. (2025·May/June 0620/43 Q3(c)(v)) · Level 2

(c) Zinc sulfate crystals are made by the reaction between zinc carbonate and dilute sulfuric acid, using the following steps.

step 1 An excess of powdered zinc carbonate is added to dilute sulfuric acid.

(v) **Step 1** is repeated using large pieces of zinc carbonate instead of powdered zinc carbonate.

All other conditions are the same.

The rate of reaction decreases.

Explain why the rate of reaction decreases. Give your answer in terms of particles.

4. (2024·October/November 0620/22 Q15) · Level 3

Sulfur dioxide is converted to sulfur trioxide in the Contact process.

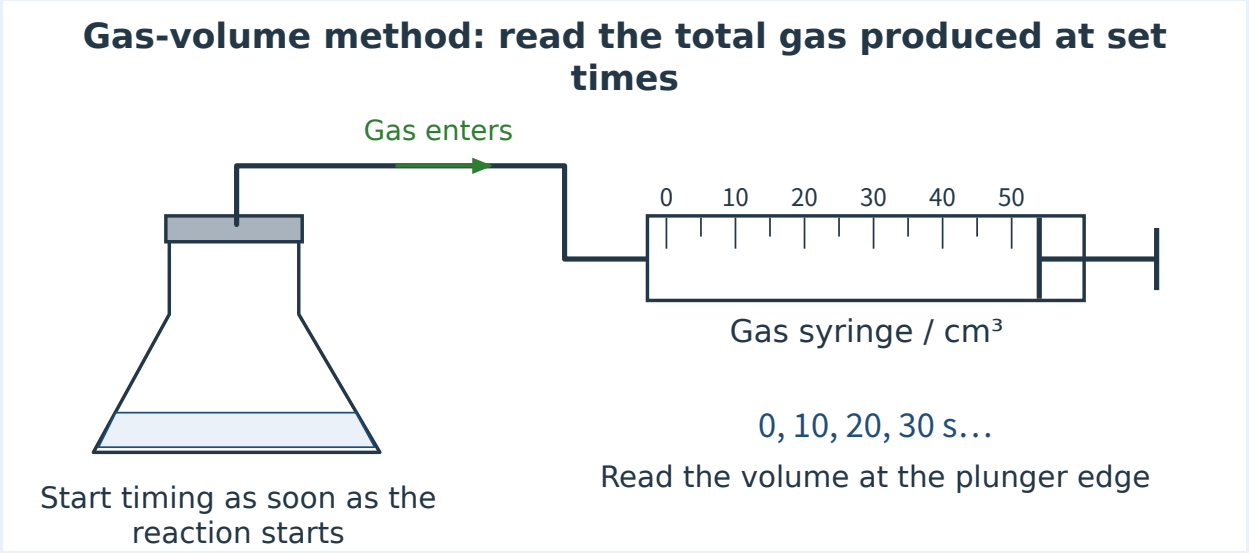
The conditions used are 450 °C and 200 kPa with a vanadium(V) oxide catalyst.

Which row describes and explains the effect of changing conditions on the rate of reaction?

	change in conditions	effect on rate	explanation
A	no catalyst	lower	the activation energy is higher
B	higher pressure	higher	the particles have more kinetic energy
C	lower temperature	lower	the particles collide more frequently
D	lower pressure	higher	there are more particles per unit volume

Type 2: Measuring reaction rates and comparing data and graph gradients

Key Ideas Reaction rate describes how quickly a reaction proceeds. To measure it, choose a quantity that changes during the reaction and can be measured reliably, and record time alongside it. Common choices are the volume of gas produced, the decrease in mass as gas escapes, or the time taken for a fixed length of solid to react completely. Use the same measured quantity and units in each comparison.
If a product is a gas, connect the reaction flask to a gas syringe. Start timing as soon as the reaction begins and read the cumulative volume at fixed intervals. For example, hydrogen peroxide, H_2O_2, decomposes:
$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
Manganese(IV) oxide, MnO_2, can act as a catalyst. The oxygen produced pushes the plunger out, so the volume reading increases as the reaction proceeds.



An airtight reaction flask is connected by a delivery tube to a gas syringe with a freely moving plunger. Readings are taken at equal time intervals. A leak reduces the collected volume. The scale illustrates where to take the reading.

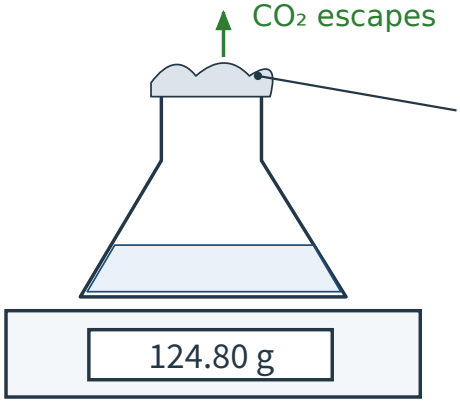
For example, to compare H_2O_2 decomposition under two conditions, read oxygen volumes at the same times, such as 0, 10, 20 and 30 s. Record them in a table and plot cumulative volume against time. Mix the reagents and start the stopwatch as nearly simultaneously as possible. Connections must be airtight and the plunger must move freely. If the reagents are mixed some time before the bung and delivery tube are fitted, early gas escapes and initial volume readings are too low. Gas volume depends on temperature and pressure, so both experiments need the same measurement conditions. Collection over water is suitable for gases with low solubility; significant dissolution in water affects the collected volume.

Another method lets gas escape from the reaction flask while a balance records the mass of the whole apparatus. Calcium carbonate, $CaCO_3$, reacts with dilute hydrochloric acid, HCl, to form calcium chloride, $CaCl_2$, carbon dioxide, CO_2, and water:

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Escaping CO_2 lowers the balance reading. Provided other losses, such as splashing and evaporation, are prevented, the mass decrease represents the mass of gas released.

Mass method: record the apparatus losing mass as gas escapes



Time / s	Apparatus mass / g
0	124.80
20	124.50

Mass lost: 0.30 g

For a carbonate reacting with acid to produce CO_2, readings are taken at set times on the same balance. A loose cotton-wool plug reduces splashing. Do not seal the flask with an airtight bung, since the gas must escape. The mass decrease gives the mass of escaped gas.

For example, the apparatus has a mass of 124.80 g at 0 s and 124.50 g at 20 s. Subtract the final mass from the initial mass to find the loss during these 20 s:

Mass loss = 124.80 − 124.50 = 0.30 g

Then divide by the corresponding time:

$$\text{Mean rate of mass loss} = \frac{0.30}{20} = 0.015 \text{ g/s}$$

A loose cotton-wool plug in the neck reduces splashing while allowing CO_2 to escape. Do not seal it with an airtight bung. Escaping droplets would have their mass incorrectly counted as gas. If the gas mass is very small, the change may be close to the balance resolution, so measuring gas volume may give a clearer reading. When evaluating a method, explain how a specific error affects the reading; “the experiment is inaccurate” is insufficient. Mean rate uses the total change over a time interval. For a gas-producing reaction, divide the increase in gas volume by the corresponding time:

Mean rate = increase in gas volume ÷ time taken

A possible unit is cm^3/s. If readings are cumulative, subtract the values at the two times. Subtract the same two times to obtain the interval. The symbol Δ means “change in”: ΔV is the volume increase and Δt is the elapsed time.

Average rate: change during the interval divided by the elapsed time

Time / s	Total gas / cm³
10	12
30	42

$\Delta V = 42 - 12 = 30\ \text{cm}^3$

$\Delta t = 30 - 10 = 20\ \text{s}$

Average rate = 30 ÷ 20

$= 1.5\ \text{cm}^3/\text{s}$

Over equal time intervals, a larger change means a larger average rate.

This example compares the volume increase from 10–30 s. For both quantities, subtract the earlier value from the later one. The 42 cm³ is the cumulative volume since the reaction began; dividing 42 by 20 does not give the rate over this interval.

For example, 12 cm³ has been collected at 10 s and 42 cm³ at 30 s. To find the mean rate from 10–30 s, first calculate the changes in volume and time:

$\Delta V = 42 - 12 = 30\ \text{cm}^3$

$\Delta t = 30 - 10 = 20\ \text{s}$

Then divide the additional volume by the time for that interval:

$$\text{Mean rate} = \frac{\Delta V}{\Delta t} = \frac{30}{20} = 1.5\ \text{cm}^3/\text{s}$$

For equal time intervals, the increases can be compared directly. For different intervals, divide each change by its actual duration; the larger increase alone does not establish the faster rate.

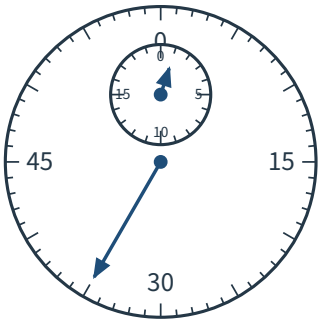
When the endpoint is complete disappearance of a solid, record the time taken to reach the same endpoint. For example, use magnesium ribbon, Mg, of the same length and thickness each time, with sufficient acid for all the magnesium to react. Calculate mean rate using the definition given in the question:

Mean rate = length of magnesium ribbon ÷ time to disappear

The unit is cm/s. When comparing acid concentrations, keep ribbon length and thickness, acid volume, temperature, stirring and the method of judging the endpoint the same. Change only acid concentration. Repeated measurements and checking anomalous results help reduce random differences in readings.

For a stopwatch with two dials, identify which dial shows minutes and which shows seconds, then convert to a single time unit. There are 60 s in 1 min. “1 min 35 s” is neither 1.35 decimal minutes nor 1.35 seconds.

Stopwatch: read minutes and seconds, then convert everything to seconds



Outer dial: 35 s; inner dial: 1 min

$$1\ \text{min} = 60\ \text{s}$$

$$1\ \text{min}\ 35\ \text{s}$$

$$= 1 \times 60 + 35$$

$$= 95\ \text{s}$$

As in the original question, the outer dial shows seconds and the inner dial minutes. This example shows 1 min 35 s. Enter 95 in a table requiring seconds, not 1.35.

For example, a 4.0 cm magnesium ribbon disappears completely in 16 s. Using the given definition:

$$\text{Mean rate} = \frac{4.0}{16} = 0.25\ \text{cm/s}$$

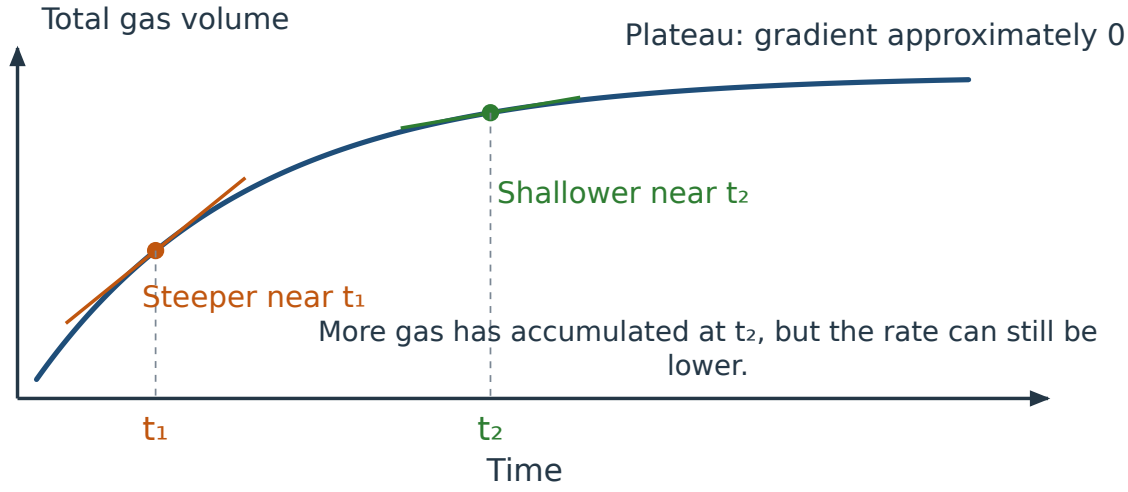
The numerator is the length that actually reacted, and the denominator is the time for it to disappear completely. Do not use cm³/s: this example measures length, not gas volume.

If the stopwatch reads 1 min 35 s, convert to seconds before substituting into the denominator:

$$1 \times 60 + 35 = 95\ \text{s}$$

At a particular instant, use the local gradient: how quickly the cumulative quantity increases over a small time near that point. A steep gradient means fast; a gentle gradient means slow. A horizontal plateau has gradient 0, showing that the plotted product quantity no longer increases. Curve height gives the amount accumulated so far, not directly the amount being produced per second now.

Cumulative curves: local gradient gives rate; a plateau means no further increase

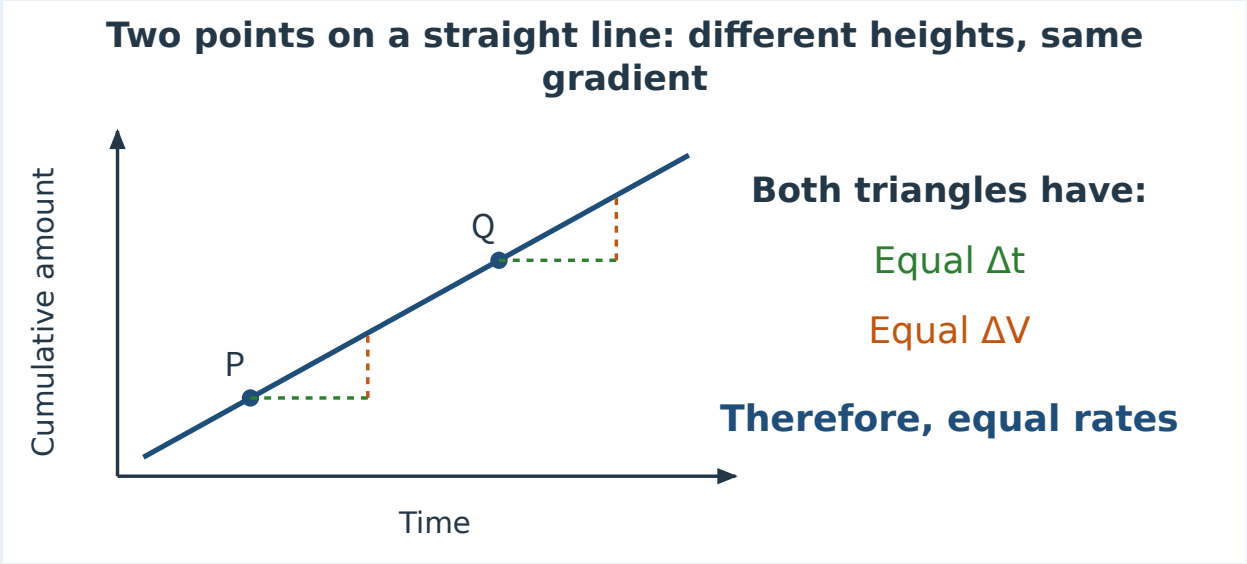


The short straight lines indicate the slope near each point. Curve height represents cumulative amount; gradient represents the production rate at that instant. The point at t_2 is higher, but the curve is flatter there.

For example, locate t_1 and t_2 and examine the curve near each. Near t_1 it is steeper, so volume increases more per unit time. Near t_2 it is flatter, so the rate at t_2 is lower. The larger cumulative volume at t_2 does not alter this

conclusion. A plateau shows that gas production has stopped within the measurement limits. It does not by itself prove that every reactant has been used up, since one may have been in excess.

If cumulative amount against time is a straight line, every point on the line has the same gradient. When comparing different experiments, also use the shared horizontal and vertical scales.



Equal horizontal time intervals on the same straight line give equal vertical increases. Q is higher than P, indicating a greater cumulative amount, not a greater rate.

For example, take equal time intervals near P and Q. The volume increases are equal, so the rates are equal. Q is higher only because more product has accumulated by then. To compare this line with a curve that starts steeply and then flattens, examine the local gradient at each point. A horizontal point has the slowest rate; the highest plateau is not the fastest part.

Answers: Mean rate = change during the interval ÷ actual time interval. Convert minutes to seconds when reading a two-dial stopwatch. A steeper local gradient on a cumulative-amount graph means a faster reaction.

Recognising the Question Check the measured quantity and units, then distinguish an interval mean from the gradient at an instant. Subtract cumulative readings first. For instantaneous rate, compare steepness, not just height.

Worked Example 1 (2025·May/June 0620/61 Q2(d)(i)) · Level 3

A student investigates how the rate of reaction of magnesium ribbon with dilute acid changes as the concentration of the acid is changed. The student uses five solutions of the same acid, **A**, **B**, **C**, **D**, and **E**. Each solution has a different concentration. The acid is in excess in all experiments.

The student does five experiments.

Experiment 1

- Use a 50 cm^3 measuring cylinder to pour 30 cm^3 of acid **A** into a 100 cm^3 conical flask.
- Add a coil of magnesium ribbon to the acid in the conical flask and immediately start a stop-watch.
- Continually swirl the mixture in the conical flask until the magnesium ribbon disappears completely. Immediately stop the stop-watch and record the time in seconds to the nearest second.
- Empty and rinse the conical flask with distilled water.

Experiment 2

- Repeat Experiment 1 using 30 cm^3 of acid **B** instead of acid **A**.

Experiment 3

- Repeat Experiment 1 using 30 cm^3 of acid **C** instead of acid **A**.

Experiment 4

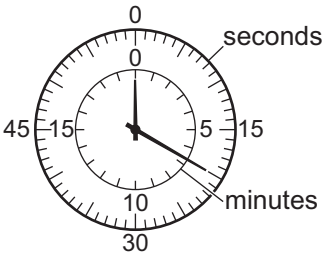
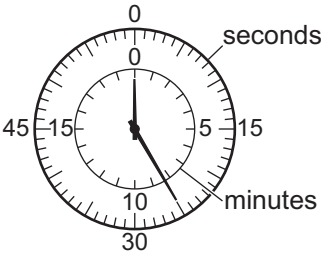
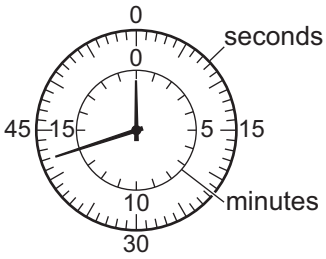
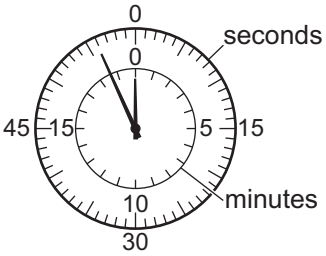
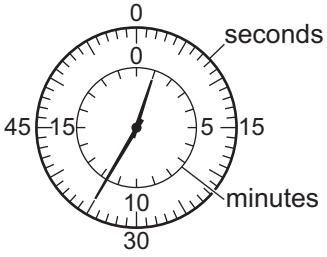
- Repeat Experiment 1 using 30 cm^3 of acid **D** instead of acid **A**.

Experiment 5

- Repeat Experiment 1 using 30 cm^3 of acid **E** instead of acid **A**.

(a) Use the stop-watch diagrams to complete Table 2.1

Table 2.1

experiment	acid	concentration of acid in mol / dm^3	stop-watch diagram	time for magnesium to disappear in s
1	**A**	2.0		
2	**B**	1.5		
3	**C**	1.0		
4	**D**	0.8		
5	**E**	0.5		

(d) The mean rate of reaction is calculated using the equation shown.

$$\text{mean rate of reaction} = \frac{\text{length of magnesium ribbon in cm}}{\text{time for magnesium to disappear in s}}$$

The length of each coil of magnesium ribbon used in all five experiments was 5 cm.

(i) Calculate the mean rate of reaction in Experiment 1. Give units for the rate you have calculated.

units

Solution

(a) The outer dial shows seconds and the inner dial minutes. In experiments 1–4 the inner dial reads 0 min, while the outer dial reads 20, 25, 42 and 56 s. Enter 20, 25, 42 and 56 in order. In experiment 5 the inner dial reads 1 min and the outer dial 35 s. Convert to seconds:

$1 \times 60 + 35 = 95$ s

Enter 95 in the final table cell. The time column requires s, so enter all times as whole seconds.

Original Table 2.1: complete readings of all five stopwatches

Experiment	Acid concentration / mol dm⁻³	Minutes and seconds	Time to enter / s
1	2.0	0 min 20 s	20
2	1.5	0 min 25 s	25
3	1.0	0 min 42 s	42
4	0.8	0 min 56 s	56
5	0.5	1 min 35 s	95

All five rows of s25/61 Q2(a) are complete. The final row converts 1 min 35 s to 95 s. The time column contains whole numbers of seconds only.

(d)(i) In experiment 1, each magnesium ribbon is 5 cm long and takes 20 s to disappear completely. Using the given formula:

$$\text{Mean rate} = \frac{5 \text{ cm}}{20 \text{ s}} = 0.25 \text{ cm/s}$$

The value is 0.25 and the unit is cm/s. Do not replace 5 with the acid volume, 30: the numerator is explicitly the length of magnesium ribbon.

Worked Example 2 (2023·October/November 0620/41 Q4(b)(i)) · Level 2

Aqueous hydrogen peroxide, H_2O_2, slowly forms water and oxygen at room temperature and pressure, r.t.p. This reaction is catalysed by manganese(IV) oxide.

The equation is shown.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

(b) A student investigates the rate of formation of oxygen gas when manganese(IV) oxide is added to aqueous hydrogen peroxide.

The volume of oxygen gas formed is measured at regular time intervals at r.t.p. The results are plotted onto the graph in Fig. 4.1.

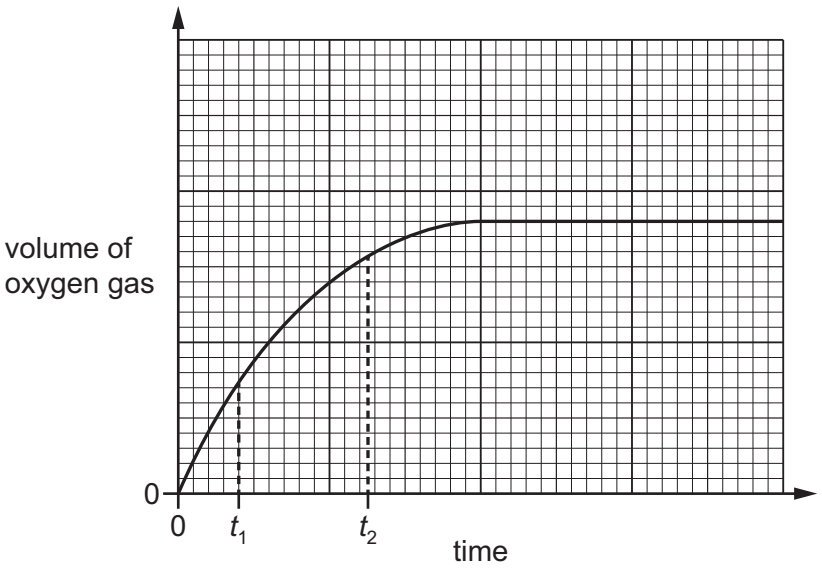


Fig. 4.1

(i) State how the graph in Fig. 4.1 shows the rate of reaction at time t_2, is lower than at time t_1.

Solution

Locate the points on the curve at t_1 and t_2. The curve is steeper near t_1 and flatter near t_2, so the gradient at t_2 is smaller. The gradient of cumulative oxygen volume against time gives reaction rate, so the rate at t_2 is lower than at t_1. Write: "The gradient at t_2 is lower than at t_1." Merely saying that there is more oxygen at t_2 describes cumulative volume, not how quickly oxygen is forming then.

Practice

1. (2025·May/June 0620/13 Q14) · Level 3

Magnesium is reacted with dilute hydrochloric acid. The table shows the total volume of hydrogen produced every 15 seconds.

time / s	0	15	30	45	60	75	90	105	120
total volume of hydrogen / cm^3	0	18	32	48	59	64	68	72	74

During which time period is the rate of reaction the fastest?

A 0–30 seconds

B 30–60 seconds

C 60–90 seconds

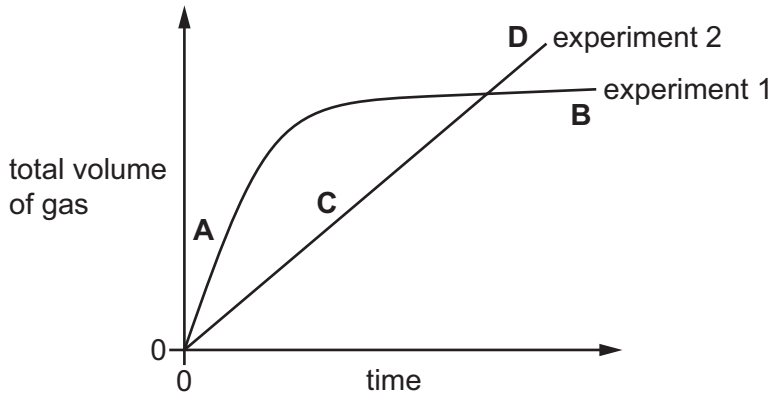
D 90–120 seconds

2. (2025·May/June 0620/12 Q16) · Level 3

The total volume of gas produced in two different experiments is measured against time.

A graph of the results of the two experiments is shown.

Which point shows the fastest rate of reaction?



6.3 Reversible reactions and equilibrium

Type 1: Distinguishing reversible reactions, dynamic equilibrium and equilibrium shifts

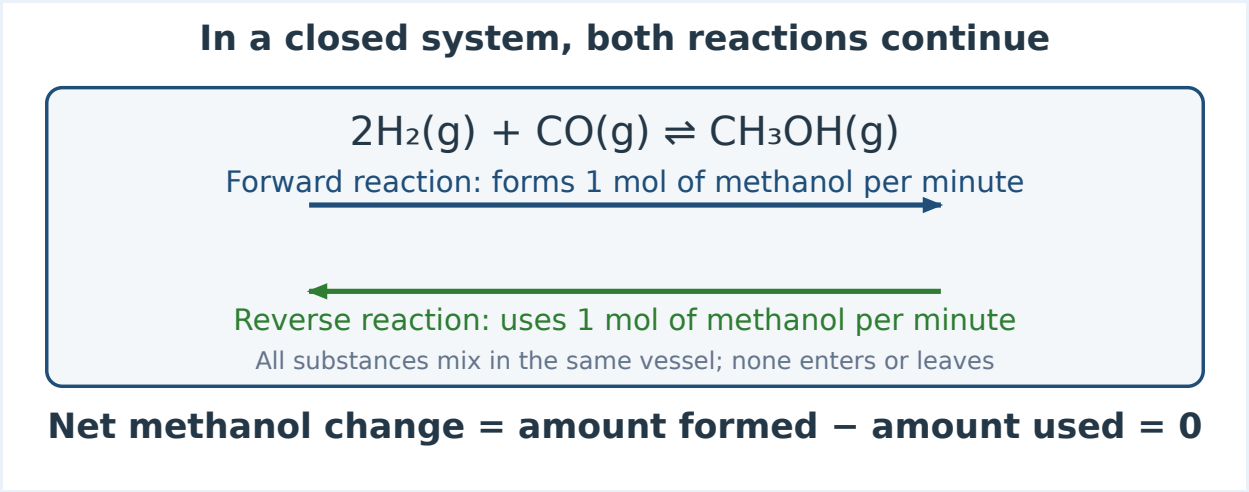
Key Ideas

Consider a specific reaction: hydrogen, H_2, and carbon monoxide, CO, can form methanol, CH_3OH, while methanol can also decompose back into these reactants. The equation is:

$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

The symbol (g) means gaseous. The $\rightleftharpoons$ symbol shows that the reaction is reversible. Left to right is the forward reaction; right to left is the reverse reaction.

In a closed system, substances cannot enter or leave, although heat can be exchanged with the surroundings. When temperature and other conditions remain constant, the forward and reverse reactions can become equally fast: as much methanol is consumed each second as is produced. This is dynamic equilibrium.



The arrows show the two reaction directions. The value 1 mol per minute only illustrates equal rates; actual equilibrium rates depend on the system and its conditions.

For example, if the forward reaction produces 1 mol of methanol in a minute and the reverse reaction consumes exactly 1 mol, the net change is:

Net change in methanol = 1 − 1 = 0 mol

Methanol concentration therefore stays constant, as does the concentration of each other substance. “Equal” describes the two reaction rates; “constant” describes each substance’s own concentration. The concentrations of hydrogen, CO and methanol need not equal one another. The reactions continue, which is why the equilibrium is dynamic.

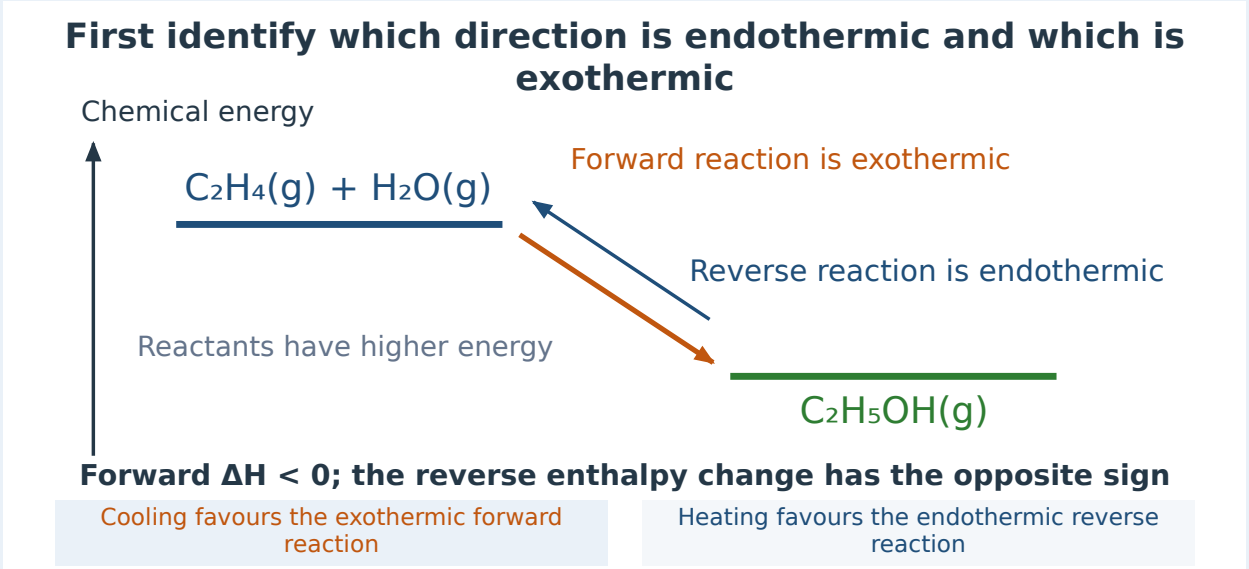
Next, change one condition and allow a new equilibrium to form. “Equilibrium shifts to the right” means that, between the altered state and the new equilibrium, a net forward reaction forms products. A shift left means that a net reverse reaction forms substances on the left. At the new equilibrium, both rates are equal again. Equilibrium yield asks how much target product can form from the same starting materials at equilibrium. Reaction rate asks how quickly the conversion occurs. Assess these separately.

First consider temperature. Heating favours the endothermic direction; cooling favours the exothermic direction. The net reaction proceeds in the direction that partly opposes the temperature change. If the forward reaction is exothermic, the reverse is endothermic, and vice versa. When ΔH is given for the forward reaction, $\Delta H < 0$ means exothermic and $\Delta H > 0$ means endothermic.

For example, ethene has the formula C_2H_4 and ethanol has the formula C_2H_5OH. Their reversible reaction with steam is:

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$

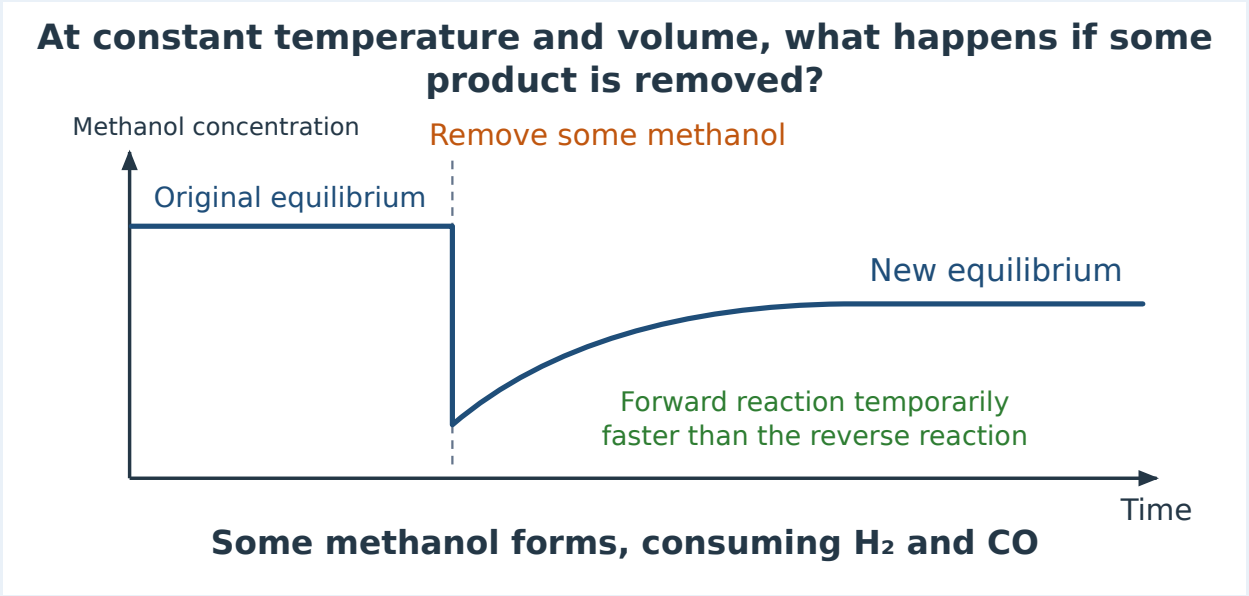
The forward reaction is exothermic.



The energy levels show only that ΔH for the forward reaction is negative. This diagram gives neither activation energy nor reaction rate.

For example, decrease the temperature. Since the forward reaction is exothermic, the exothermic direction is to the right. Cooling favours that direction, so the equilibrium yield of ethanol increases. Heating instead favours the endothermic direction to the left, lowering ethanol's equilibrium yield. At the same time, heating usually increases reaction rate. Reaching equilibrium sooner does not mean that more ethanol is present at equilibrium.

Next consider concentration, the amount of a substance per unit volume. At constant temperature and volume, adding a reactant raises its concentration. The subsequent net reaction tends to consume some of the added substance. Removing a product favours replacing some of that product. After making the addition or removal, close the system again and compare the new equilibrium.



The unnumbered vertical axis shows an immediate drop followed by a partial recovery. Replacing some methanol does not mean restoring its entire original concentration.

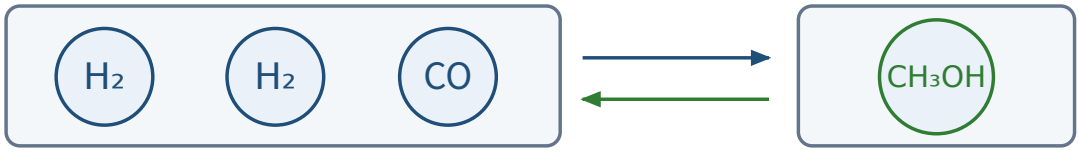
For example, remove some CH_3OH from an equilibrium methanol mixture. Its concentration first drops sharply because of removal. Equilibrium then shifts right, forming some methanol while consuming H_2 and CO. At the same volume, H_2 and CO concentrations therefore fall. Methanol concentration only partly recovers; it need not return fully to its original value.

Reverse the change to reason in the other direction. Adding methanol favours its consumption, shifting equilibrium left. Removing H_2 or CO also favours a leftward shift to replace some of that reactant. Thus adding a substance on the left or removing one on the right favours the right; adding one on the right or removing one on the left favours the left. A substance directly added or removed experiences both the external change and the subsequent reaction, so specify the starting point when comparing its concentration.

For gas pressure, first count the gaseous amounts in the balanced equation. Include only substances labelled (g) and add their coefficients. In this section, pressure is changed by altering container volume. Compressing the same reacting gas mixture at constant temperature favours the side with fewer moles of gas. Increasing volume and lowering pressure favours the side with more moles of gas. If the sums of gaseous coefficients are equal, this pressure change does not favour either side.

Changing gas pressure: first add the gaseous coefficients on each side

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$



Left: 2 + 1 = 3 gas volumes

Right: 1 gas volume

Higher pressure favours the right, with fewer gas molecules; equilibrium methanol yield increases

Lower pressure favours the left, with more gas molecules; equilibrium methanol yield decreases

The circles help count gaseous amounts in the balanced equation. They do not imply that only these molecules are present or that the two sides occupy separate containers.

For example:

$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

Count the gaseous coefficients on each side:

Gas on the left: 2 + 1 = 3 parts

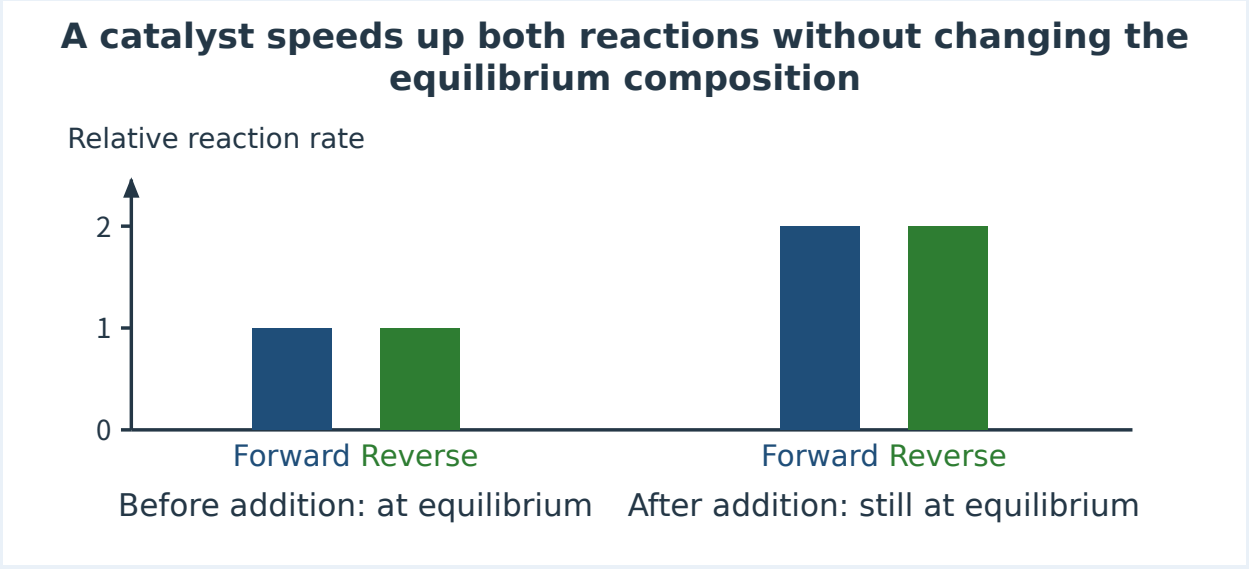
Gas on the right: 1 part

Higher pressure favours methanol formation to the right, increasing its equilibrium yield. Lower pressure favours the left, decreasing the yield. Count coefficients, not formula subscripts, and do not confuse the number of different gas species with the total gaseous amount.

If the same question also asks about rate, apply collision theory separately. At the same temperature, compressed reacting gases have more reacting particles per unit volume and more frequent collisions, so the reaction is usually faster. Lower pressure can therefore have two distinct disadvantages: a slower reaction and a lower equilibrium yield of the desired product.

Another distinction is essential. Compression also changes volume, so “how much reactant is consumed” differs from “how much reactant is present per dm^3”. A shift right means the subsequent net reaction consumes substances on the left. To compare absolute concentration with its value before compression, also account for the volume change. A shift right alone does not prove that the final concentration must fall.

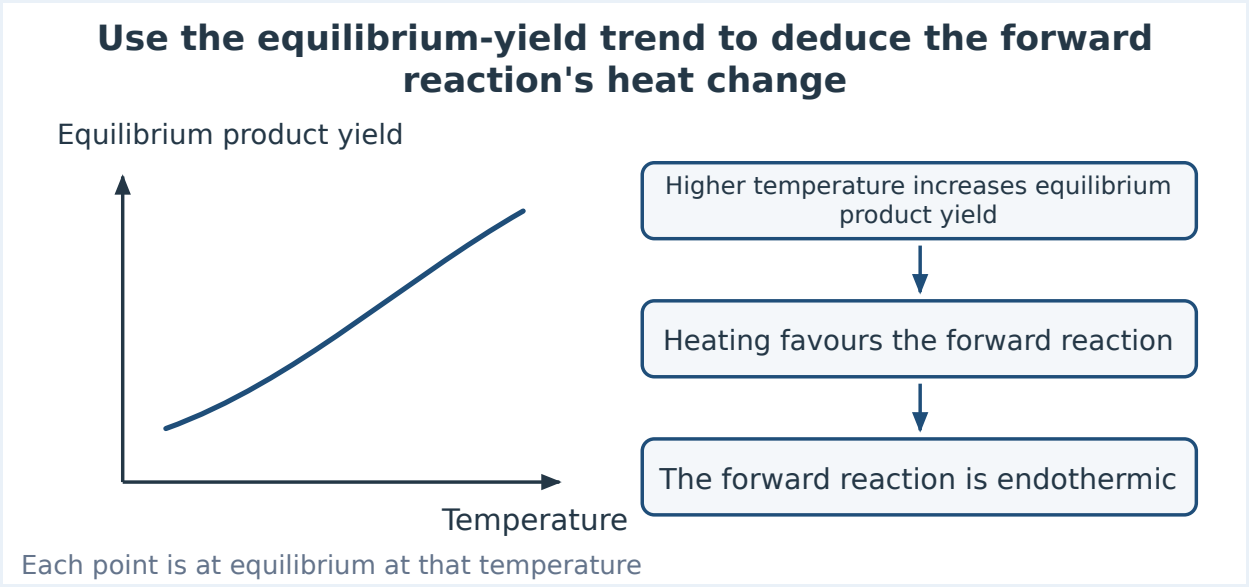
A catalyst provides a lower-activation-energy pathway and speeds up both forward and reverse reactions. At unchanged temperature it alters neither equilibrium composition nor the equilibrium yield of the target substance. If equilibrium has not yet been reached, a catalyst allows it to be reached sooner. If equilibrium already exists, both rates increase and remain equal.



Both sets use the same rate scale. Bar heights illustrate relative rates; they do not mean that a catalyst necessarily doubles the rate.

For example, add a catalyst to a methanol reaction already at equilibrium. If one column asks for the equilibrium concentration of CH_3OH, enter no change. If another asks for the reverse reaction rate, enter increases. Both are correct: formation and consumption are faster, but the net change remains zero. An unchanged equilibrium concentration does not mean an unchanged reverse reaction rate.

Some questions ask you to infer the thermal effect of the forward reaction from the effect of temperature on equilibrium yield. Reverse the temperature reasoning: first decide whether heating favours the forward or reverse direction, then identify that favoured direction as endothermic.



Equilibrium yields at different temperatures are compared using the same starting materials and pressure. The curve illustrates a trend and contains no experimental values.

For example, heating increases the equilibrium yield of the product on the right. More product means heating favours the forward direction. Since heating favours the endothermic direction, the forward reaction is endothermic. If heating instead lowers the product's equilibrium yield, heating favours the reverse reaction. The reverse reaction is then endothermic and the forward reaction exothermic. This reasoning uses yield after equilibrium is reached; a greater amount formed within a few seconds is not an equivalent observation.

For a table with several rows, treat each specified change independently. Recheck whether each column asks about concentration, equilibrium yield or reaction rate. At dynamic equilibrium, the two reaction rates are _____.

Adding a catalyst leaves the equilibrium composition _____. If the forward reaction is exothermic, cooling causes the equilibrium yield of the product on the right to _____.

Heating hydrated copper(II) sulfate or cobalt(II) chloride, and adding water to the anhydrous salts, are reversible changes that can be shown using ⇌. The blue/white and pink/blue colour changes, together with the heating and water-addition directions, are explained with step-by-step diagrams in section 7.3, "Recognising water of crystallisation and anhydrous salts, and explaining heating and rehydration", followed by complete worked examination questions and practice.

Answers: equal; unchanged; increase

Recognising the Question Read the reversible equation, gaseous coefficients and thermal effect of the forward reaction. Then assess the direction favoured by each change. Answer concentration, equilibrium yield and rate separately, specifying volume and the starting point for concentration comparisons.

Worked Example 1 (2025·May/June 0620/41 Q5(a)) · Level 2

Methanol, CH_3OH, is manufactured by the reaction between hydrogen and carbon monoxide.

An equilibrium mixture is produced.

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$

(a) State what happens to the concentration of CH_3OH when the reaction is at equilibrium.

Explain your answer in terms of rate of forward and reverse reactions.

Solution

Read the reversible reaction:

$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

The forward reaction to the right forms methanol, CH_3OH. The reverse reaction to the left consumes methanol and reforms hydrogen and carbon monoxide.

At equilibrium, the forward and reverse reaction rates are equal. Over the same time interval, as much methanol is consumed as is formed. The net change in methanol is zero, so its concentration stays constant.

A complete answer is: the concentration of CH_3OH no longer changes because the forward and reverse reactions continue at equal rates. "Constant" does not mean zero methanol or that both reactions stop. The rates are equal; the concentrations of all the substances need not be equal.

Worked Example 2 (2025·October/November 0620/23 Q19) · Level 3

Ethanol is manufactured by the addition of steam to ethene.

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$$

The reaction conditions are 300 °C and 60 atm.

The forward reaction is exothermic.

Which rows describe the effect on the equilibrium yield of C_2H_5OH when the stated change is made?

	change	equilibrium yield of C_2H_5OH
1	adding a catalyst	no change
2	decreasing the temperature	increases
3	decreasing the pressure	increases

A 1 and 2 **B** 1 and 3 **C** 2 only **D** 3 only

Solution

Identify the quantity being compared: equilibrium yield of C_2H_5OH means ethanol's yield at equilibrium. The forward reaction is:

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$

The question states that the forward reaction is exothermic.

Row 1 adds a catalyst. It speeds up both forward and reverse reactions, reaching equilibrium sooner without changing its composition. Ethanol's equilibrium yield is therefore no change. This row is correct.

Row 2 lowers temperature. Cooling favours the exothermic direction, here the forward reaction to the right. More reactants are converted to ethanol, increasing its equilibrium yield. The table entry increases is correct.

Row 3 lowers pressure. Count only gaseous coefficients: 1 part C_2H_4 plus 1 part H_2O on the left gives 2 parts gas; the right has 1 part C_2H_5OH. Lower pressure favours the left, with more moles of gas. Ethanol's equilibrium yield decreases, so increases is incorrect.

Row	Change	Ethanol equilibrium yield	Original row
1	Add a catalyst	Unchanged	Correct
2	Lower temperature	Increases	Correct
3	Lower pressure	Decreases	Incorrect

Only rows 1 and 2 are correct, so choose A (1 and 2). Each row is assessed separately from the original conditions; do not combine cooling and reduced pressure into one change.

Practice

1. (2025·May/June 0620/12 Q17) · Level 1

The equation for the reaction of sulfur dioxide with oxygen is shown.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

What is meant by the symbol $\rightleftharpoons$?

A redox reaction

B reversible reaction

C displacement reaction

D thermal decomposition reaction

2. (2025·May/June 0620/43 Q5(a)) · Level 2

Ethanoic acid is manufactured in the reaction between methanol and carbon monoxide.

An equilibrium mixture is produced.

$$CH_3OH(g) + CO(g) \rightleftharpoons CH_3COOH(g)$$

(a) State **two** characteristics of an equilibrium.

3. (2025·May/June 0620/41 Q5(b)(ii)) · Level 2

Methanol, CH_3OH, is manufactured by the reaction between hydrogen and carbon monoxide.

An equilibrium mixture is produced.

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$

(b) The reaction is carried out at a pressure of 75 atmospheres and a temperature of 250 °C.

(ii) Complete Table 5.1 using **only** the words **increases**, **decreases** or **no change**.

Table 5.1

	effect on the concentration of $CH_3OH(g)$ at equilibrium	effect on the rate of the reverse reaction
catalyst is added		

4. (2025·May/June 0620/41 Q5(b)(iii)) · Level 2

Methanol, CH_3OH, is manufactured by the reaction between hydrogen and carbon monoxide.

An equilibrium mixture is produced.

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$

(b) The reaction is carried out at a pressure of 75 atmospheres and a temperature of 250 °C.

(iii) If a temperature of more than 250 °C is used, the yield of methanol decreases.

State what can be deduced about the forward reaction.

5. (2025·May/June 0620/43 Q5(b)(ii)) · Level 3

Ethanoic acid is manufactured in the reaction between methanol and carbon monoxide.

An equilibrium mixture is produced.

$$CH_3OH(g) + CO(g) \rightleftharpoons CH_3COOH(g)$$

(b) Both a high yield of ethanoic acid and a high rate of reaction are needed.

The reaction is carried out at 300 °C.

The forward reaction is exothermic.

(ii) Complete Table 5.1 using **only** the words **increases**, **decreases** or **no change**.

Table 5.1

	effect on the equilibrium concentration of $CH_3COOH(g)$	effect on the rate of the forward reaction
catalyst is added		increases
pressure is increased		

6. (2025·May/June 0620/41 Q5(b)(i)) · Level 3

Methanol, CH_3OH, is manufactured by the reaction between hydrogen and carbon monoxide.

An equilibrium mixture is produced.

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$

(b) The reaction is carried out at a pressure of 75 atmospheres and a temperature of 250 °C.

(i) State **two** disadvantages of using a pressure **below** 75 atmospheres.

Question Analysis

1. (2025·May/June 0620/42 Q6(d)(i)) · Level 3

(d) C_2H_4 reacts with steam to form ethanol.

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g) \qquad \Delta H = -45\,kJ/mol$$

The process happens in a closed system and the reaction reaches an equilibrium.
The conditions for this process are 300 °C and 60 atm pressure. H_3PO_4 is used as a catalyst.

(i) Complete Table 6.1 to show the effect, if any, on the concentration of $C_2H_4(g)$ when the following changes to the conditions are applied.

Only use the words **increases**, **decreases** or **no change**.

Table 6.1

change to condition	effect on the concentration of $C_2H_4(g)$ at equilibrium
temperature is decreased	
some $C_2H_5OH(g)$ is removed	
pressure is increased	
a more effective catalyst is used	

Answer: The official mark scheme gives, from top to bottom: decreases; decreases; decreases; no change. Read the pressure row alongside the concentration-comparison discussion below.

Explanation: This question asks for the equilibrium concentration of C_2H_4. The original paper gives $C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$, with forward $\Delta H = -45$ kJ/mol, at 300°C and 60 atm, using phosphoric acid, H_3PO_4, as catalyst.

Change in conditions	Official required entry
Lower temperature	decreases
Remove some ethanol	decreases
Increase pressure	decreases
Use a more effective catalyst	no change

Row 1 lowers temperature. Negative ΔH means an exothermic forward reaction. Cooling favours the forward direction to the right, consuming ethene. Comparing the change caused by this equilibrium shift at the same volume, ethene concentration falls: decreases.

Row 2 removes some ethanol, the product on the right. Removal favours a net forward reaction that replaces some ethanol and consumes ethene. At unchanged temperature and container volume, ethene concentration falls: decreases.

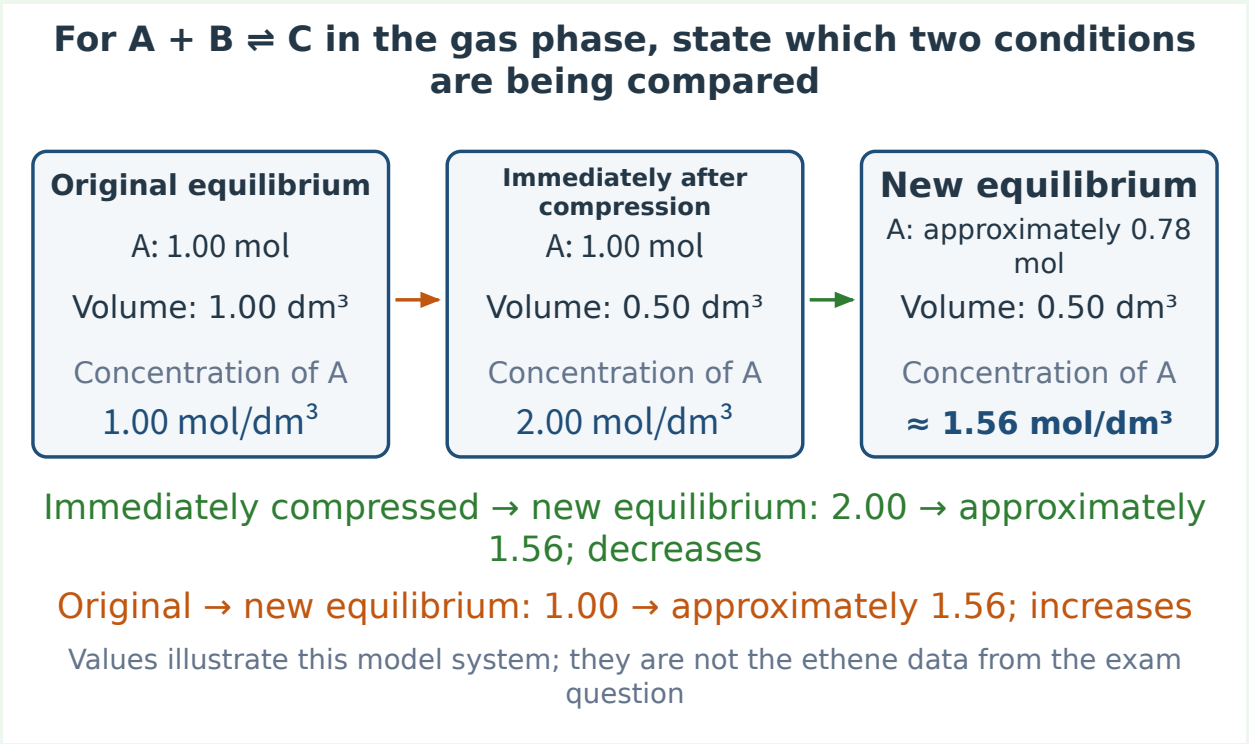
Row 4 uses a more effective catalyst. Both directions become faster without changing equilibrium composition, so ethene's equilibrium concentration shows no change.

Row 3 needs an extra check. The left has 2 parts gas and the right 1. Compression at the same temperature favours the right, so the net reaction during re-equilibration consumes ethene. This establishes a decrease in its amount,

but does not by itself prove that its concentration falls relative to before compression, because concentration also depends on volume.

Distinguish three states: the original equilibrium before compression; the instant after volume is reduced, before reaction can restore equilibrium; and the subsequent new equilibrium. Immediately after compression, no additional ethene has yet reacted, but the same amount occupies a smaller volume, so concentration rises. The following net forward reaction then lowers ethene concentration from that higher value.

A hypothetical A(g) + B(g) $\rightleftharpoons$ C(g) system illustrates the choice of comparison. A concentration equals its amount divided by container volume. The diagram gives both values for all three states, so divide them directly.



Concentration = amount of substance ÷ volume. The net forward reaction reduces the amount of A, but the new volume is also smaller. Reaction direction alone cannot determine concentration relative to the original state.

From "immediately after compression" to "new equilibrium", A concentration falls from 2.00 to about 1.56 mol/dm³. From "original equilibrium" to "new equilibrium", it instead rises from 1.00 to about 1.56 mol/dm³. A can be consumed while its absolute concentration still rises relative to the original state.

The official expected entry for row 3 is therefore decreases, but the question does not specify which of these two starting points is intended. Do not turn it into a general rule that increasing pressure shifts equilibrium right and therefore necessarily lowers the final reactant concentration below its pre-compression value. For a question explicitly asking about equilibrium yield, use gaseous coefficients to identify the direction. For a concentration comparison, also check amount, volume and starting point.

Further reading: OpenStax Chemistry 2e, 13.3: equilibrium shifts and their relationship to concentration and volume.

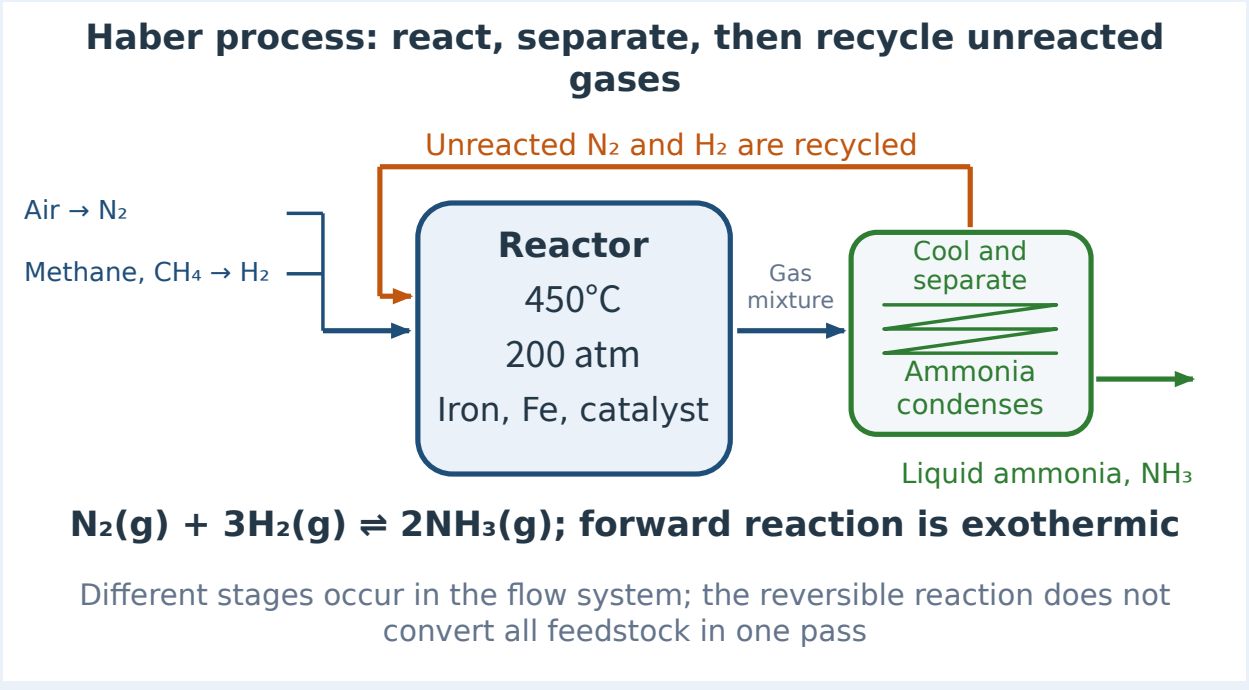
Type 2: Identifying the Haber and Contact processes and explaining compromise conditions

Key Ideas First identify what the Haber process makes. It converts nitrogen, N_2, and hydrogen, H_2, into ammonia, NH_3. Nitrogen comes from air, and hydrogen can be obtained from methane, CH_4. These questions require the sources of the raw materials, not a derivation of every hydrogen-production step.

The reaction is:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

The forward reaction is exothermic. Typical conditions are 450°C, 200 atm (about 20 000 kPa) and an iron, Fe, catalyst. The reactants do not all become ammonia in one pass through the reactor. The mixture is cooled, ammonia condenses to a liquid, and unreacted N_2 and H_2 are returned to the reactor.



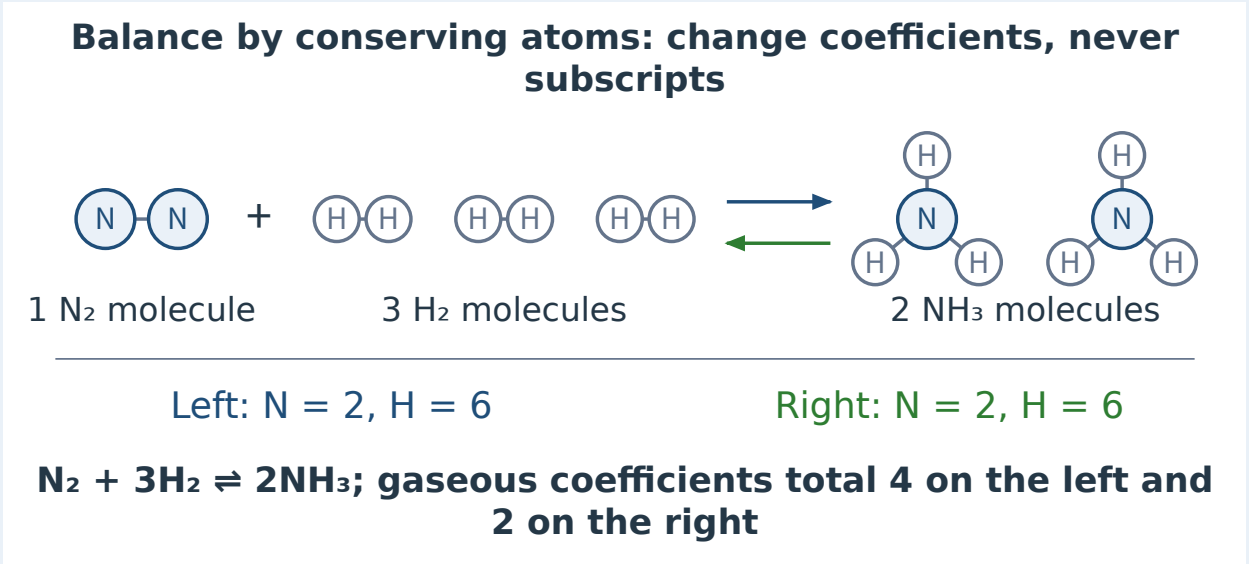
Arrows show material flow. Only the sources of raw materials are listed. The cooler separates ammonia from unreacted nitrogen and hydrogen.

For example, a question asking for "three typical conditions" requires temperature, pressure and catalyst: 450°C; 200 atm; iron catalyst. Cooling and recycling are subsequent separation and raw-material-use steps; neither replaces one of the three reactor conditions.

The equation coefficients have specific meanings. N_2 contains two N atoms, so write 2 before NH_3, giving 2 N and 6 H atoms on the right. Each H_2 molecule contains 2 H atoms, so $3H_2$ is needed on the left to supply 6 H atoms:

$$1N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The coefficient 1 is normally omitted.



The atom circles are for counting. Connecting lines indicate which atoms belong to the same molecule; they do not show actual three-dimensional shapes or every bond type.

For example, check both sides: each has 2 N and 6 H atoms, so the equation is balanced. Do not change NH_3 to NH_2 to make the numbers fit; that changes the substance's formula. For pressure, count gaseous coefficients separately:

Gas on the left: 1 + 3 = 4 parts

Gas on the right: 2 parts

Counting gaseous amounts and counting atoms are different tasks.

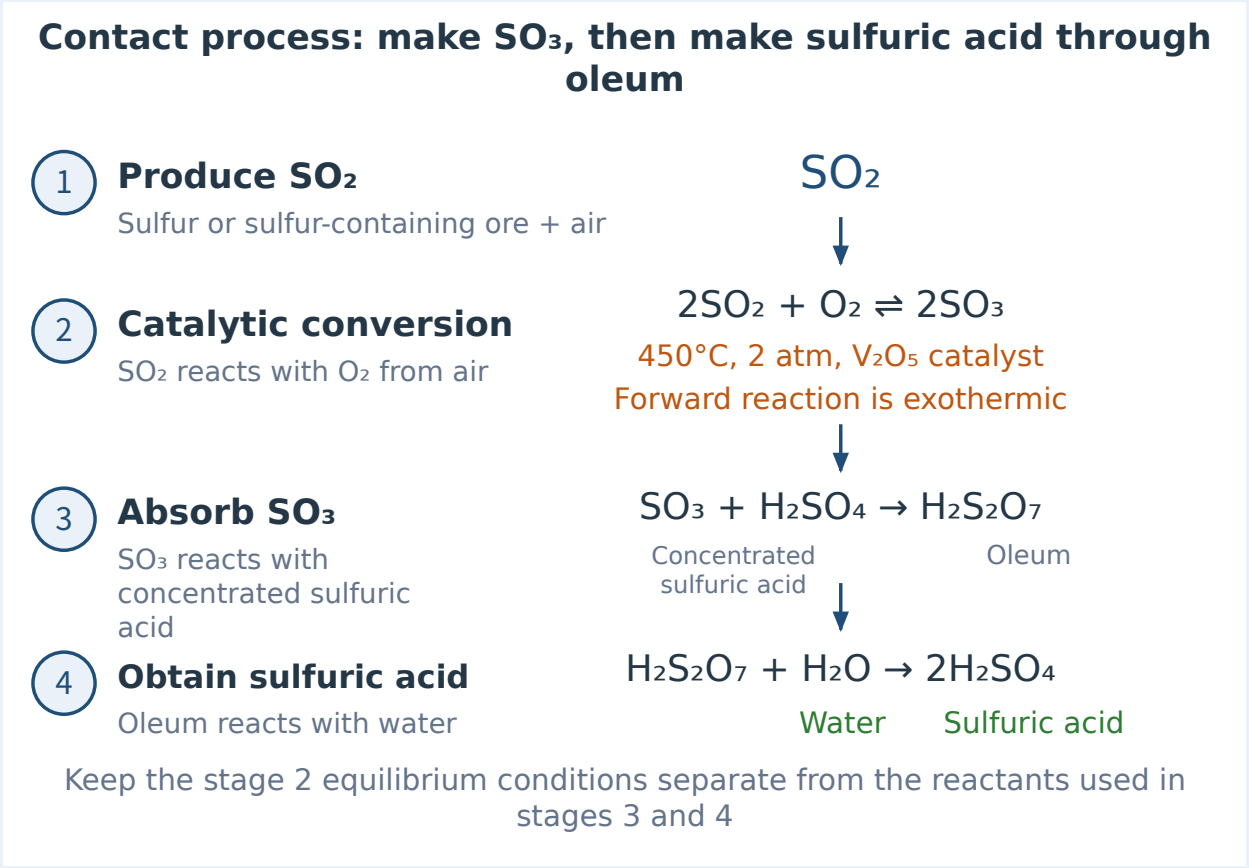
The Contact process produces sulfuric acid, H_2SO_4. Sulfur dioxide, SO_2, is obtained by burning sulfur, S, or roasting sulfur-containing ores in air. Oxygen, O_2, needed in the next stage also comes from air.

The key reversible reaction is:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

This produces sulfur trioxide, SO_3, and the forward reaction is exothermic. Typical conditions for this stage are 450°C, 2 atm (about 200 kPa) and vanadium(V) oxide, V_2O_5, as catalyst. The (V) means vanadium has oxidation number +5; Roman numerals are explained in detail in the next question type.

SO_3 then reacts with concentrated sulfuric acid to produce oleum, represented here by $H_2S_2O_7$. Reaction with water subsequently produces more sulfuric acid. Follow the materials through the four stages specified in the question:



Vertical arrows show sulfur-containing material entering the next stage. Each horizontal equation separately identifies the substance added at that stage.

For example, if asked for the stage 2 equation, write only sulfur dioxide reacting with oxygen to form sulfur trioxide:

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

Each side has 2 S and 6 O atoms, and the reversible arrow must be retained. In stage 3, the reactants instead are SO_3 and H_2SO_4 and the given product is $H_2S_2O_7$. Do not copy O_2 from the previous stage.

For stage 3, use the names to write:

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

Check H, S and O separately. Sulfuric acid is a reactant in this step. Its production again in the next step does not make it the same as the V_2O_5 catalyst in stage 2.

Write the reaction equation, then check all three elements

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

Comparison	H atoms	S atoms	O atoms
Reactant total	0 + 2 = 2	1 + 1 = 2	3 + 4 = 7
Product total	2	2	7

Every element is balanced, so each of the three substances has coefficient 1

Check the stage 3 equation by adding each element's atoms. The subscripts 2, 3, 4 and 7 describe composition and are not altered during balancing.

For example, check stage 3 element by element:

- H: 2 atoms on the left and 2 on the right.
- S: each left-hand substance supplies 1 atom, giving 2 in total; there are 2 on the right.
- O: the left supplies 3 plus 4, giving 7 in total; there are 7 on the right.

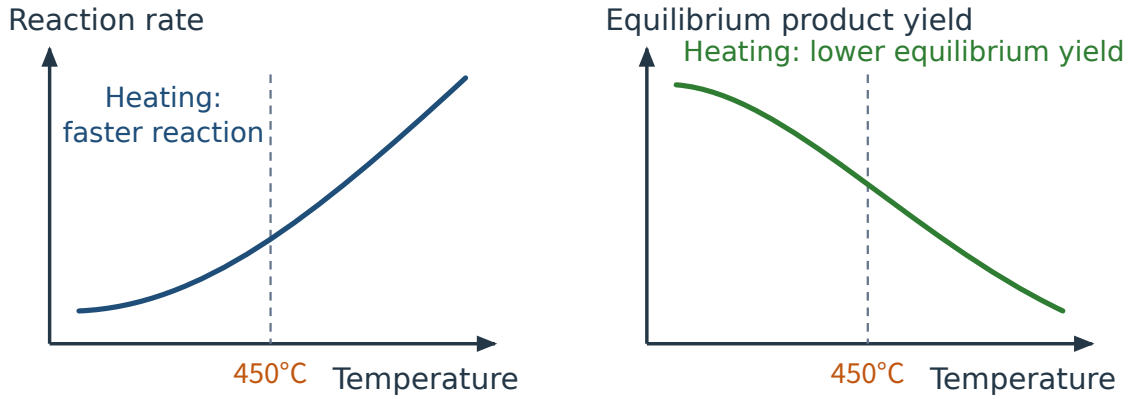
All three atom totals match, so all coefficients are 1. Stage 4 is:

$$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$$

Both sides have 4 H, 2 S and 8 O atoms. Direct reaction of SO_3 with water is violent and can produce an acid mist that is difficult to collect. Industrially, SO_3 is first absorbed in concentrated sulfuric acid, followed by controlled production of sulfuric acid at the required concentration.

Why do both exothermic processes use about 450°C rather than progressively lower temperatures? Separate rate from equilibrium yield. Lower temperature favours the exothermic forward reaction, giving more desired product at equilibrium. However, particles have lower average kinetic energy, so there are fewer successful collisions per unit time and production is slow. Higher temperature speeds the reaction but favours the endothermic reverse direction, lowering the target product's equilibrium yield.

Exothermic industrial reactions: temperature affects rate and equilibrium yield differently



450°C balances an acceptable rate and yield; it is not the intersection of the two curves

The vertical axes show different quantities; only the temperature axes correspond directly

The curves illustrate trends without experimental values. The two graphs address rate and yield after equilibrium is reached separately.

For example, lowering the Haber-process temperature below 450°C decreases rate but increases ammonia's equilibrium yield. Thus "below 450°C, both rate and yield decrease" is only half correct. The typical 450°C condition is a compromise between an acceptable yield and production rate; it does not maximise both simultaneously. Higher temperatures also require more heating energy and more demanding materials and operating conditions.

Apply the same sequence to another exothermic reaction given in a question. For example, methanol, CH_3OH, and carbon monoxide, CO, can form ethanoic acid, CH_3COOH, with a stated typical temperature of 300°C. Below 300°C the disadvantage is slower rate; above 300°C it is a lower equilibrium yield of ethanoic acid. The 300°C belongs to that process as described; do not replace it with the familiar 450°C.

Pressure also brings benefits and costs. Gaseous coefficients decrease from 4 to 2 in the Haber reaction and from 3 to 2 in Contact-process stage 2. Compression at the same temperature favours the product side in both, increasing the desired product's equilibrium yield. Compression also packs reacting gas particles more closely, usually increasing collision frequency and rate.

Pressure affects both processes, but their operating pressures differ

Haber process	Contact process, stage 2
$N_2 + 3H_2 \rightleftharpoons 2NH_3$	$2SO_2 + O_2 \rightleftharpoons 2SO_3$
4 gas volumes → 2 gas volumes	3 gas volumes → 2 gas volumes
200 atm ≈ 20 000 kPa	2 atm ≈ 200 kPa
Iron, Fe, catalyst	Vanadium(V) oxide, V_2O_5, catalyst

Higher pressure improves both equilibrium yields, but compression and pressure-resistant equipment cost more

Both typically use 450°C; remember each process's own pressure and catalyst

All substances in these equations are gaseous; count amounts using coefficients. Typical conditions follow the 2026–2028 syllabus and do not imply identical values in every factory.

For example, raising Haber-process pressure above 200 atm does not lower ammonia yield. Equilibrium still favours ammonia formation. The limitations are the energy for further compression and stronger reactors and pipes, with greater equipment, maintenance and safety costs. The typical high pressure therefore reflects a balance of yield, rate and cost.

The Contact process already achieves high conversion at relatively low pressure. The extra benefit of greatly increased pressure does not justify the equipment and compression costs, so about 2 atm is used, sufficient to move gases through the plant. “Low pressure” is relative to industrial pressures of hundreds of atmospheres. It means neither zero pressure nor that lowering pressure shifts equilibrium towards SO_3. In the multiple-choice description “high temperature, low pressure”, the conditions are 450°C and about 2 atm.

Finally, check the catalysts. Iron and V_2O_5 suit their respective reactions, providing lower-activation-energy pathways and allowing equilibrium to be reached sooner. At the same temperature, a catalyst does not alter equilibrium composition, so an increased equilibrium yield does not explain its use. In practice, a suitable catalyst can allow useful rates under milder conditions and reduce costs. This does not change the equilibrium conclusion for a comparison at the same temperature.

The gaseous raw materials for the Haber process are _____ and _____. The catalyst in Contact-process stage 2 is _____. For an exothermic forward reaction, lowering temperature usually makes the rate _____ and the product’s equilibrium yield _____.

Answers: nitrogen, N_2; hydrogen, H_2; vanadium(V) oxide, V_2O_5; decrease; increase

Recognising the Question Identify the process and stage, then match temperature, pressure and catalyst. Where explanation is required, discuss rate, equilibrium yield, cost and safety separately.

Worked Example 1 (2025·May/June 0620/41 Q4(b)(i)) · Level 2

This question is about nitrogen and its compounds.

Nitrogen contains molecules with the formula N_2.

(b) Ammonia, NH_3, is manufactured by reacting nitrogen with hydrogen, H_2, in the Haber process.

(i) State **three** typical conditions for the reaction between nitrogen and hydrogen in the Haber process.

Solution

The question asks for three typical conditions for nitrogen and hydrogen reacting in the Haber reactor. Organise the answer as temperature, pressure and catalyst.

The temperature is 450°C, the pressure is 200 atm or 20 000 kPa, and the catalyst is iron. Write: 450°C; 200 atm; iron catalyst. Include the appropriate units with numerical answers.

The question specifies ammonia production, so do not give the Contact-process conditions of 2 atm and vanadium(V) oxide. Cooling to liquefy ammonia and recycling unreacted gases are later processing steps and cannot replace these three reaction conditions.

Worked Example 2 (2025·May/June 0620/43 Q4(b)(iii)) · Level 2

This question is about compounds of sulfur.

(a) Sulfuric acid is manufactured in a four-stage process.

stage 1 Sulfur dioxide is produced from ores containing sulfur.

stage 2 Sulfur dioxide reacts with oxygen to form sulfur trioxide.

stage 3 Sulfur trioxide reacts with concentrated sulfuric acid to form oleum, $H_2S_2O_7$.

stage 4 Oleum reacts with water to form concentrated sulfuric acid.

(iii) Write the symbol equation for the reaction in **stage 3**.

Solution

Return to stage 3: sulfur trioxide reacts with concentrated sulfuric acid to form oleum, whose formula is given as $H_2S_2O_7$. This is not stage 2, so do not use the reversible SO_2–O_2 equation.
Sulfur trioxide is SO_3 and sulfuric acid is H_2SO_4. Put both reactants on the left and the product on the right:
$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$
Check each element:

- H: 2 atoms on each side.
- S: the two left-hand substances each contain 1, giving 2 in total; the right also has 2.
- O: the left has 3 plus 4, giving 7; the right also has 7.

The equation is already balanced. The complete answer is:
$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$
The coefficient 1 before each substance can be omitted.

Worked Example 3 (2025·May/June 0620/21 Q16) · Level 3

Industrially, ammonia is made by the Haber process. The equation is shown.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\ \text{kJ/mol}$$

The conditions used are 450 °C and 200 atm.

Which statements about the conditions are correct?

1 Temperatures below 450 °C give a lower rate of reaction.

2 Temperatures below 450 °C give a lower yield of ammonia.

3 Pressures above 200 atm give a lower yield of ammonia.

4 Pressures above 200 atm are too expensive to maintain.

A 1 and 2 **B** 1 and 4 **C** 2 and 3 **D** 3 and 4

Solution

Read the equation and thermal effect:
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
The given enthalpy change is:
$\Delta H = -92\ \text{kJ/mol}$
The negative sign means the forward reaction is exothermic. Count the gaseous coefficients:
Left: 1 + 3 = 4
Right: 2
Assess each statement separately; do not combine rate and yield into one conclusion.

Statement 1: below 450°C, particles have lower average kinetic energy and fewer successful collisions per unit time, so the rate decreases. Correct.

Statement 2: cooling favours the exothermic forward reaction, so ammonia's equilibrium yield increases. The claim that yield decreases is incorrect.

Statement 3: higher pressure favours the right, with fewer parts gas, increasing ammonia's equilibrium yield. The claim that pressure above 200 atm lowers yield is incorrect.

Statement 4: further pressure increases require more compression energy and stronger reactors and pipes, increasing operating costs. The typical stated conditions already balance yield and cost, so this industrial-cost statement is correct.

Only statements 1 and 4 are correct, corresponding to B.

Practice

1. (2025·May/June 0620/42 Q1(b)) · Level 1

Using numbers only, state the:

(b) typical operating temperature of the Contact process in °C

2. (2025·May/June 0620/22 Q16) · Level 2

Which row describes the conditions used in the manufacture of sulfuric acid by the Contact process?

	catalyst	pressure	temperature
A	iron	high	high
B	iron	low	low
C	vanadium(V) oxide	high	low
D	vanadium(V) oxide	low	high

3. (2025·May/June 0620/43 Q4(b)(i)) · Level 2

This question is about compounds of sulfur.

(a) Sulfuric acid is manufactured in a four-stage process.

stage 1 Sulfur dioxide is produced from ores containing sulfur.

stage 2 Sulfur dioxide reacts with oxygen to form sulfur trioxide.

stage 3 Sulfur trioxide reacts with concentrated sulfuric acid to form oleum, $H_2S_2O_7$.

stage 4 Oleum reacts with water to form concentrated sulfuric acid.

(b) (i) State **three** typical conditions for the reaction between sulfur dioxide and oxygen in **stage 2**.

4. (2025·May/June 0620/41 Q4(b)(ii)) · Level 2

This question is about nitrogen and its compounds.

Nitrogen contains molecules with the formula N_2.

(b) Ammonia, NH_3, is manufactured by reacting nitrogen with hydrogen, H_2, in the Haber process.

(ii) Write the symbol equation for the chemical reaction in the Haber process.

5. (2025·May/June 0620/43 Q4(b)(ii)) · Level 2

This question is about compounds of sulfur.

(a) Sulfuric acid is manufactured in a four-stage process.

stage 1 Sulfur dioxide is produced from ores containing sulfur.

stage 2 Sulfur dioxide reacts with oxygen to form sulfur trioxide.

stage 3 Sulfur trioxide reacts with concentrated sulfuric acid to form oleum, $H_2S_2O_7$.

stage 4 Oleum reacts with water to form concentrated sulfuric acid.

(ii) Write a symbol equation for the reaction that occurs in **stage 2**.

6. (2025·May/June 0620/43 Q5(b)(i)) · Level 3

Ethanoic acid is manufactured in the reaction between methanol and carbon monoxide.

An equilibrium mixture is produced.

$$CH_3OH(g) + CO(g) \rightleftharpoons CH_3COOH(g)$$

(b) Both a high yield of ethanoic acid and a high rate of reaction are needed.

The reaction is carried out at 300 °C.

The forward reaction is exothermic.

(i) State the disadvantage of using a temperature:

- below 300 °C
- above 300 °C.

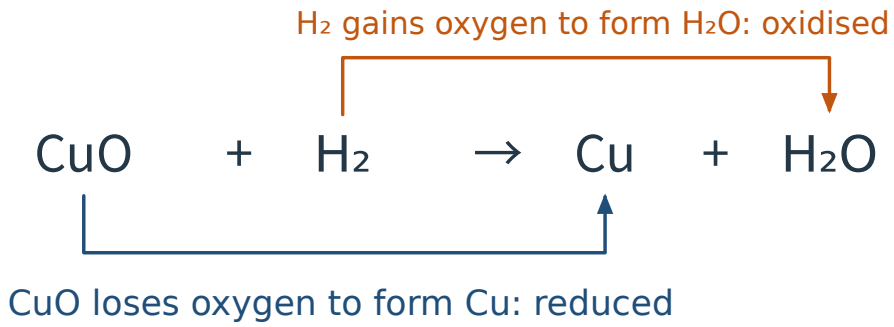
6.4 Oxidation and reduction

Type 1: Tracking oxygen, electrons and oxidation numbers to identify redox and reagent roles

Key Ideas

Begin with a reaction in which oxygen can be tracked:
$CuO + H_2 \rightarrow Cu + H_2O$
Copper(II) oxide, CuO, loses oxygen as it becomes copper, Cu. Hydrogen, H_2, gains oxygen as it becomes water, H_2O. Gain of oxygen is oxidation; loss of oxygen is reduction.

Track oxygen first: which substance loses it, and which gains it?



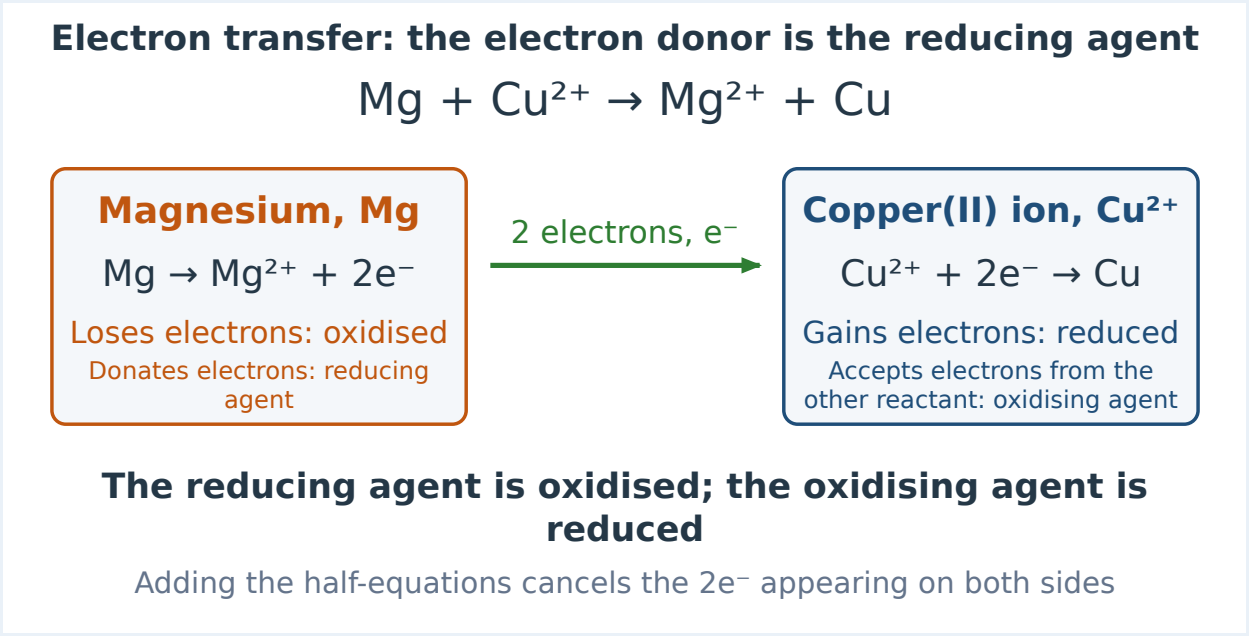
Oxygen from CuO ends up in water; oxidation and reduction occur together

The connecting lines track composition from reactants to products. They do not imply a reaction step involving free oxygen atoms.

For example, follow the copper-containing substance: CuO → Cu. Oxygen disappears from its composition, so CuO is reduced. Follow the hydrogen-containing substance: $H_2 \rightarrow H_2O$. It gains oxygen, so H_2 is oxidised. Both changes occur together in the same reaction, called a redox reaction. When asked which substance is reduced, identify the reactant undergoing change, CuO, rather than merely naming the final Cu product.

Now consider electron transfer. Some redox reactions contain no oxygen but can still be identified using electrons. When magnesium, Mg, reacts with copper(II) ions, Cu^{2+}, Mg loses two electrons to form Mg^{2+}, while Cu^{2+} gains them to form Cu. Electrons carry negative charge and are written e^-. Gaining or losing electrons does not change the number of protons.

Oxidation can also be defined as loss of electrons and reduction as gain of electrons. An equation showing one of these changes separately is a half-equation. The numbers of electrons transferred in the two half-equations must match.



The central arrow shows electrons transferring from magnesium to copper(II) ions. Both half-equations conserve atom numbers and total charge.

For example, first write the half-equation for magnesium losing electrons:

$$Mg \rightarrow Mg^{2+} + 2e^-$$

Electrons on the right show that Mg loses them and is oxidised.

Next write the half-equation for copper(II) ions gaining electrons:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Electrons on the left show that Cu^{2+} gains them, so Cu^{2+} is reduced. Add the equations and cancel $2e^-$ from each side:

$$Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$$

A reducing agent reduces another substance and is itself oxidised. An oxidising agent oxidises another substance and is itself reduced. Here Mg transfers electrons to Cu^{2+}, reducing Cu^{2+}, so Mg is the reducing agent. Cu^{2+} accepts electrons, causing Mg to be oxidised, so Cu^{2+} is the oxidising agent. The name describes its effect on the other substance. Adding what happens to the agent itself helps prevent reversal.

If only formulae are given, without explicit electrons, use oxidation numbers to keep track. An oxidation number is assigned to an atom of an element to track electron allocation. In a covalent molecule it need not represent an actual ionic charge. Use these initial rules: an element in its uncombined state has oxidation number 0, as in Cu, C, H_2, O_2 and N_2. A monatomic ion's oxidation number equals its charge, such as +2 for Cu^{2+} and −1 for Cl^-. In a neutral compound, the sum for all atoms is 0. In a polyatomic ion, the sum equals the overall ionic charge.

For common compounds in this section, H is usually +1 and O usually −2. If a question supplies an oxidation number, use that value. These usual values have exceptions, so do not overlook an explicit value in the question. Multiply each known oxidation number by the number of those atoms, then use the required sum to find the unknown.

Sum of oxidation numbers: first check the particle's overall charge

Neutral molecule	Polyatomic ion
H_2SO_4	SO_4^{2-}
H is usually +1; O is usually −2	The whole ion has charge 2−
$2(+1) + x + 4(-2) = 0$	$x + 4(-2) = -2$
Sulfur: x = +6	**Sulfur: x = +6**

x is the oxidation number of each S atom; the right side is the whole particle's charge

This is oxidation-number accounting, not a structural diagram. The sum is 0 in neutral H_2SO_4 and −2 in SO_4^{2-}.

For example, let x be the oxidation number of each S atom in sulfuric acid, H_2SO_4. The two H atoms are each +1 and the four O atoms each −2. The molecule is neutral:

$$2(+1) + x + 4(-2) = 0$$

Calculate the contributions of H and O, then solve for S:

$$2 + x - 8 = 0$$

$$x = +6$$

For sulfate, SO_4^{2-}, the overall charge is 2−, so the sum must be −2:

$$x + 4(-2) = -2$$

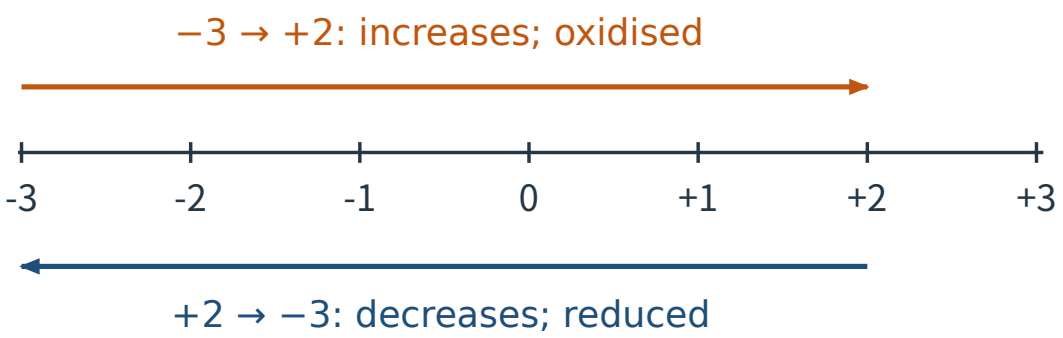
$$x - 8 = -2 \quad \Rightarrow \quad x = +6$$

Each S is still +6. Do not ignore the superscript 2− or report the whole SO_4^{2-} ion's sum, −2, as sulfur's oxidation number.

Roman numerals in chemical names also specify an element's oxidation number. In copper(II) oxide, II means copper is +2. In chlorine(I) oxide, I means chlorine is +1. These are not atom counts. Cl_2O is covalent: saying chlorine is +1 describes its oxidation number, not the presence of free Cl^+ ions in the molecule.

Oxidation and reduction can also be defined through these numbers: an increase is oxidation and a decrease is reduction. Compare signed values correctly. Going from −3 to +2 is an increase; a negative starting value does not make it a decrease.

An increase in oxidation number is oxidation, even from a negative starting value



The number line compares the same element's oxidation number before and after reaction. Its arrow does not show an atom moving through solution.

For example, in $Fe^{2+} \rightarrow Fe^{3+}$, iron's oxidation number increases from +2 to +3, so Fe^{2+} is oxidised. It loses one electron:

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

If the reactant is written $5Fe^{2+}$, the 5 means five amounts of that ion. It does not change each iron's oxidation number to +10.

To decide whether an underlined substance is a reducing agent, identify the element to track, calculate its oxidation number in the reactant and corresponding product, and compare. An increase means that reactant is itself oxidised and is therefore a reducing agent. A decrease identifies the reduced substance. If every element's oxidation number stays unchanged, the substance cannot be identified as this reaction's reducing agent.

For a known polyatomic ion, its overall charge can initially be treated as one unit. This section uses sulfate, SO_4^{2-}, and hydroxide, OH^-, with charges 2− and 1−. For example, iron(II) sulfate, $FeSO_4$, and iron(II) hydroxide, $Fe(OH)_2$, are both neutral.

For a group in brackets, start with the charge of the whole ion

Sulfate, SO_4^{2-}: 2−; hydroxide, OH^-: 1−

$FeSO_4$	$Fe(OH)_2$
$x + (-2) = 0$	$x + 2(-1) = 0$
Fe has oxidation number +2	Fe also has oxidation number +2

The 2 in $Fe(OH)_2$ multiplies the whole OH^- group; it does not make each Fe +4

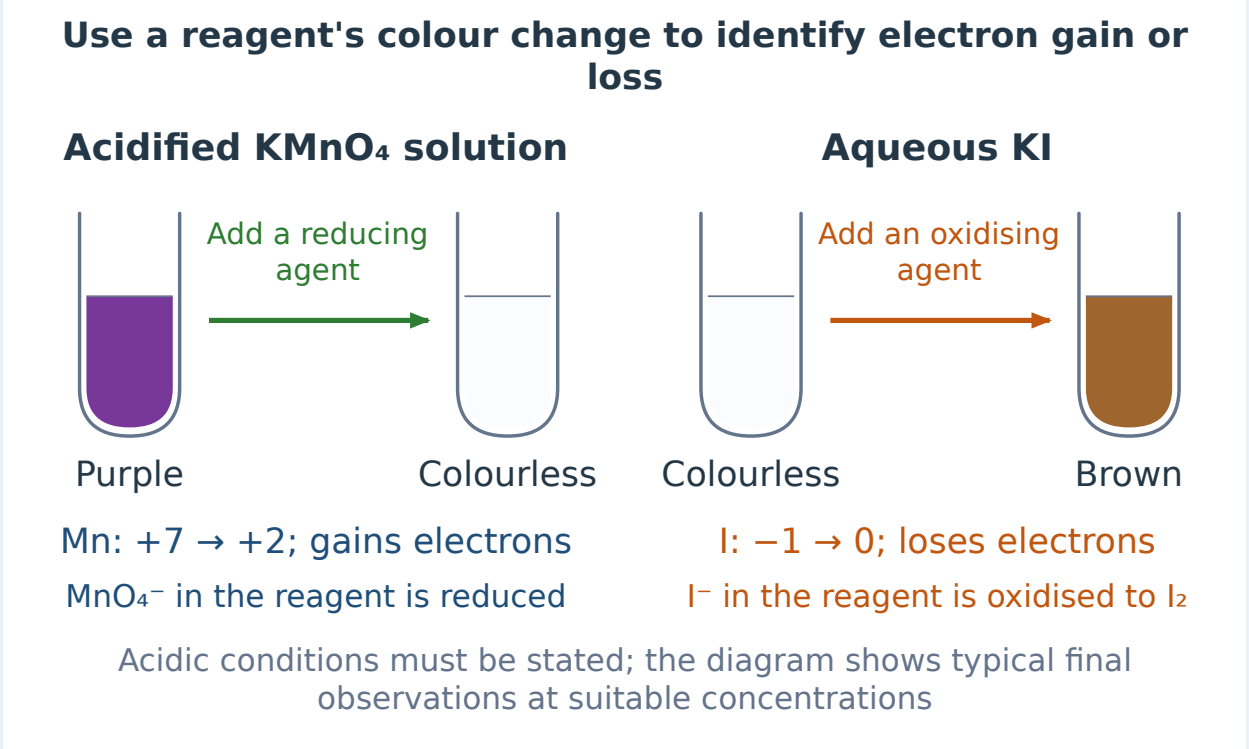
SO_4^{2-} and OH^- are known complete polyatomic ions. The subscript 2 outside the brackets means two OH^- ions.

For example, SO_4^{2-} contributes −2 in $FeSO_4$, so Fe is +2. $Fe(OH)_2$ contains two OH^- ions, again totalling −2, so Fe is also +2. If $Fe(OH)_2$ reacts with sulfuric acid, H_2SO_4, to form $FeSO_4$ and water, iron remains +2. A changed formula alone does not make $Fe(OH)_2$ a reducing agent.

Two common colour tests also require matching the reagent, initial colour, final colour and the species oxidised or reduced.

The first reagent is acidified aqueous potassium manganate(VII), $KMnO_4$. MnO_4^- makes it purple. When a suitable reducing agent reacts with it under sufficiently acidic conditions, MnO_4^- is reduced to Mn^{2+}. The purple colour disappears, and dilute solutions are normally observed as colourless.

The second reagent is colourless aqueous potassium iodide, KI. A suitable oxidising agent makes I^- lose electrons and form iodine, I_2, producing a brown solution.



Transparent regions represent colourless solutions. Mn^{2+} in dilute acidic solution is normally observed as colourless. The brown colour in the KI test comes from iodine produced.

For example, if a substance changes acidified $KMnO_4$ from purple to colourless, the reagent has gained electrons and been reduced. The tested substance supplies electrons and is the reducing agent. If a substance turns colourless KI solution brown, I^- is oxidised. The tested substance causes this electron loss and is the oxidising agent. The iodine half-equation is:

$$2I^- \rightarrow I_2 + 2e^-$$

These are redox changes, not acid–base indicator changes. Acidity, reagent concentration and the extent of reaction all affect the observed result.

When applying a colour test to an unfamiliar salt name, identify the reacting ion. Sodium sulfite, Na_2SO_3, and lithium sulfite, Li_2SO_3, both contain sulfite, SO_3^{2-}. Do not confuse sulfite with sulfate, SO_4^{2-}. Sulfites reduce acidified $KMnO_4$, so the relevant colour change is purple to colourless, not brown.

For example, replacing Na_2SO_3 with Li_2SO_3 changes the accompanying metal cation. The reducing ion under test is still SO_3^{2-} and the reagent is still acidified $KMnO_4$. The initial colour remains purple and the final colour colourless. Complete both the from and to entries.

Another examination question introduces acidified potassium dichromate(VI), $K_2Cr_2O_7$, and states that its reaction with KI produces a dark brown solution. Its role can be established without guessing the new reagent's colour. The known change from colourless KI to brown I_2 shows that I^- is oxidised. $K_2Cr_2O_7$ causes that oxidation and is therefore an oxidising agent. Similarly, when $FeSO_4$ removes the purple colour of acidified $KMnO_4$, Fe^{2+} loses electrons to become Fe^{3+}, so $FeSO_4$ is the reducing agent. Determine each reagent's role from electron transfer in its own reaction.

Finally, consider a question that gives reaction names without equations. Identify the reactants and products represented by each name before tracking oxidation numbers. The following four reactions supply the specific chemistry needed for the practice questions.

Incomplete combustion of propane, C_3H_8, in insufficient oxygen can produce carbon monoxide, CO, and water, H_2O. For example:

$$2C_3H_8 + 7O_2 \rightarrow 6CO + 8H_2O$$

In elemental O_2, oxygen is 0. In CO and H_2O, it is −2. Oxygen is reduced and the fuel is oxidised, so this is redox.

Rusting of iron, Fe, requires oxygen and water. Iron becomes part of hydrated iron(III) oxide, changing from 0 in the element to +3. It is oxidised while oxygen is reduced, so rusting is also redox. Tracking only iron and oxygen, the oxide portion can be represented as:

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$

Actual rust is hydrated. This simplified equation does not imply that iron rusts normally in dry air.

If only a reaction name is given, first identify the chemical change

Incomplete combustion of propane: one route forming CO

$2C_3H_8 + 7O_2 \rightarrow 6CO + 8H_2O$

O: 0 in the element → −2 in CO and H_2O

Rusting iron requires both water and oxygen

Iron: 0 in Fe → +3 in hydrated rust

Iron loses electrons and is oxidised; oxygen is reduced

Both reactions change oxidation numbers, so both are redox reactions

The propane equation shows one possible combination of incomplete-combustion products. Rust is hydrated iron(III) oxide. The diagram tracks iron's state and does not represent rust as dry, pure Fe_2O_3.

For example, incomplete combustion need not produce only CO_2 to count as oxidation. Oxygen in the CO above is −2, already showing its decrease from 0 to −2. In rusting, iron increases from 0 to +3 in iron(III). Both examples involve oxidation-number changes and are redox. Atom counts and oxidation numbers remain distinct: $7O_2$ does not make oxygen's oxidation number 7.

Two other reactions produce new substances without changing oxidation numbers. Adding acidified aqueous silver nitrate, $AgNO_3$, to a solution containing chloride, Cl^-, allows silver ions, Ag^+, and chloride ions to form solid silver chloride, AgCl:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Silver remains +1 and chlorine remains −1. Spectator ions such as nitrate, NO_3^-, do not participate in this precipitation step.

Calcium carbonate, $CaCO_3$, decomposes on heating into calcium oxide, CaO, and carbon dioxide, CO_2:

$$CaCO_3 \rightarrow CaO + CO_2$$

Ca remains +2 and O remains −2. To find carbon's oxidation number in $CaCO_3$, let it be x and use a sum of 0 for the neutral compound:

$$(+2) + x + 3(-2) = 0$$

$$x = +4$$

Carbon is also +4 in CO_2, so no element changes oxidation number.

New substances form, but do oxidation numbers change?

Chloride and silver ions form silver chloride precipitate

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Ag: +1 → +1; Cl: −1 → −1

Calcium carbonate decomposes on heating

$$CaCO_3 \rightarrow CaO + CO_2$$

Ca: +2 → +2; C: +4 → +4; O: −2 → −2

The tracked oxidation numbers do not change, so neither reaction is redox

The symbol (aq) means dissolved in water and (s) means solid. The first equation includes only the ions that form the precipitate; the second shows thermal decomposition.

For example, "a precipitate forms" or "thermal decomposition occurs" alone does not establish whether a reaction is redox. Checking every element in these silver chloride and calcium carbonate examples shows no oxidation-number change, so neither is redox. Do not extend this conclusion to every precipitation reaction or every thermal decomposition.

Electrons on the right of a half-equation indicate electron _____, which is _____. A reducing agent is itself _____. The sum of oxidation numbers in a polyatomic ion equals _____.

Answers: loss; oxidation; oxidised; the overall charge of the ion

Recognising the Question Identify the reactant or underlined substance to track, then compare the same element's oxidation number before and after reaction. For reagent roles, state what it causes in the other substance. If only a reaction name is given, first write the corresponding equation.

Worked Example 1 (2025·May/June 0620/11 Q16) · Level 2

The equation for the reaction between copper(II) oxide and carbon is shown.

$$2CuO + C \rightarrow 2Cu + CO_2$$

Which statement about this reaction is correct?

A CuO is reduced.

B CO_2 is oxidised.

C Cu is oxidised.

D C is reduced.

Solution

The reaction is:
$2CuO + C \rightarrow 2Cu + CO_2$
Track copper first: CuO becomes Cu and loses oxygen, so CuO is reduced. A is correct.

Track carbon next: reactant C becomes CO_2 and gains oxygen, so C is oxidised. CO_2 is the product of that oxidation; do not say CO_2 is oxidised in this reaction. Cu is the product of CuO reduction and is not being oxidised here.

Check with oxidation numbers: Cu decreases from +2 in CuO to 0 in elemental Cu; C increases from 0 to +4 in CO_2. B, C and D do not match the given reaction. Choose A.

Worked Example 2 (2025·May/June 0620/41 Q4(c)(ii)) · Level 2

This question is about nitrogen and its compounds.

Nitrogen contains molecules with the formula N_2.

(c) Ammonia is converted into nitric acid in a two-step process.

In step 1, ammonia and oxygen are passed over a catalyst.

(ii) Give the oxidation number of nitrogen in:

- NH_3
- NO.

Solution

Both substances are neutral molecules, so the sum of all oxidation numbers in each is 0. Let nitrogen's oxidation number be x in each case.
NH_3: each H is usually +1, giving +3 for three H atoms:
$x + 3(+1) = 0$
$x = -3$
Nitrogen in ammonia has oxidation number −3.
NO: oxygen is usually −2:
$x + (-2) = 0$
$x = +2$
Nitrogen in nitrogen monoxide, NO, has oxidation number +2.

Enter −3 and +2 in the given order, with clear signs. A neutral molecule does not require every atom to have oxidation number 0.

Worked Example 3 (2025·May/June 0620/42 Q4(c)(ii)) · Level 3

Chromium is the element with atomic number 24 in the Periodic Table.

(c) $(NH_4)_2Cr_2O_7$ is a compound containing chromium.

The negative ion in $(NH_4)_2Cr_2O_7$ is $Cr_2O_7^{2-}$.

(ii) The oxidation number of each O in $Cr_2O_7^{2-}$ ions is –2.

Determine the oxidation number of each Cr in $Cr_2O_7^{2-}$ ions. Show your working.

Solution

Find the oxidation number of each Cr in $Cr_2O_7^{2-}$. Each O is given as –2, so the total oxygen contribution is:
$7 \times (-2) = -14$

The ion has overall charge 2–, so the sum must be –2. Let each Cr have oxidation number x; there are two Cr atoms:
$2x + (-14) = -2$

Add 14 to both sides to find the total contribution from both Cr atoms:
$2x = +12$

Divide by 2 to find the value for each Cr:
$x = +6$

Check by substitution:
$2(+6) + 7(-2) = 12 - 14 = -2$

The sum matches the overall ionic charge, so each Cr has oxidation number +6.

The question requires working, so +6 alone is insufficient. Show oxygen's total contribution and the steps using the ionic sum to find Cr. Chromium's atomic number, 24, is its proton number and is not used in this oxidation-number calculation.

Worked Example 4 (2025·May/June 0620/22 Q17) · Level 3

Which equations show the underlined species acting as a reducing agent?

1 $\underline{Fe_2O_3} + 3CO \rightarrow 2Fe + 3CO_2$

2 $\underline{Fe} + CuSO_4 \rightarrow FeSO_4 + Cu$

3 $\underline{5Fe^{2+}} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

4 $\underline{Fe(OH)_2} + H_2SO_4 \rightarrow FeSO_4 + 2H_2O$

A 1 and 2 **B** 1 and 4 **C** 2 and 3 **D** 3 and 4

Solution

The underlined substances are, in order, Fe_2O_3, Fe, Fe^{2+} and $Fe(OH)_2$. A reducing agent reduces another substance and is itself oxidised. Track iron's oxidation number in all four.

Statement 1: three O atoms in Fe_2O_3 total –6, so the two Fe atoms total +6 and each is +3. Elemental Fe in the product is 0. The decrease from +3 to 0 means Fe_2O_3 is reduced. It is an oxidising agent, not the reducing agent requested.

Statement 2: elemental Fe is 0. In product $FeSO_4$, SO_4^{2-} contributes –2, so Fe is +2. Iron increases from 0 to +2 and is oxidised, making the underlined Fe a reducing agent.

Statement 3: Fe^{2+} becomes Fe^{3+}, increasing from +2 to +3. It loses electrons and is oxidised, so Fe^{2+} is also a reducing agent. The coefficient 5 counts ions and does not alter each iron's oxidation number.

Statement 4: two OH^- ions in $Fe(OH)_2$ total –2, so Fe is +2. It remains +2 in $FeSO_4$; H, O and S oxidation numbers are also unchanged. This is the given neutralisation reaction. $Fe(OH)_2$ has not lost electrons and been oxidised, so it is not a reducing agent.

Only statements 2 and 3 qualify, corresponding to C.

Worked Example 5 (2024·October/November 0620/22 Q18) · Level 3

Aqueous iron(II) sulfate is added to acidified potassium manganate(VII). The purple colour of the potassium manganate(VII) disappears.

Aqueous potassium iodide is added to acidified potassium dichromate(VI). A dark brown solution forms.

Which row identifies the role of the iron(II) sulfate and the potassium dichromate(VI) in these reactions?

	iron(II) sulfate	potassium dichromate(VI)
A	oxidising agent	oxidising agent
B	oxidising agent	reducing agent
C	reducing agent	reducing agent
D	reducing agent	oxidising agent

Solution

In the first experiment, loss of the purple colour of acidified $KMnO_4$ shows that MnO_4^- is reduced. Added $FeSO_4$ supplies Fe^{2+}, which loses an electron:

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

Fe^{2+} is itself oxidised and reduces the other substance, so $FeSO_4$ is a reducing agent.

In the second experiment, KI initially contains I^-. The dark brown colour shows that I_2 forms:

$2I^- \rightarrow I_2 + 2e^-$

I^- loses electrons and is oxidised. $K_2Cr_2O_7$ causes this oxidation, so it is an oxidising agent.

The combination is $FeSO_4$: reducing agent; $K_2Cr_2O_7$: oxidising agent. Choose D.

Practice

1. (2025·May/June 0620/41 Q4(c)(iii)) · Level 1

This question is about nitrogen and its compounds.

Nitrogen contains molecules with the formula N_2.

(c) Ammonia is converted into nitric acid in a two-step process.

In step 1, ammonia and oxygen are passed over a catalyst.

(iii) Define oxidation in terms of oxidation number.

2. (2024·May/June 0620/42 Q5(d)(ii)) · Level 1

(d) Na_2SO_3 is oxidised by acidified aqueous potassium manganate(VII).

(ii) State the colour change when this reaction happens.

from to

3. (2024·October/November 0620/61 Q3(b)) · Level 1

A student tests two solids: solid **T** and solid **U**.

Tests on solid T

Solid **T** is lithium sulfite.

Complete the expected observations.

The student dissolves the remaining solid **T** in water to form solution **T**.

The student divides solution **T** into two portions.

(b) The student adds the first portion of solution **T** to a test-tube containing acidified aqueous potassium manganate(VII).

State the colour change.

from to

4. (2025·May/June 0620/42 Q2(b)(i)) · Level 2

This question is about covalent compounds.

(b) Chlorine(I) oxide, Cl_2O, is a simple molecule with covalent bonds.

(i) State what is meant by (I) in the name chlorine(I) oxide.

5. (2025·May/June 0620/42 Q4(c)(i)) · Level 2

Chromium is the element with atomic number 24 in the Periodic Table.

(c) $(NH_4)_2Cr_2O_7$ is a compound containing chromium.

The negative ion in $(NH_4)_2Cr_2O_7$ is $Cr_2O_7^{2-}$.

(i) State the sum of the oxidation numbers in the $Cr_2O_7^{2-}$ ion.

6. (2025·May/June 0620/43 Q4(c)) · Level 2

This question is about compounds of sulfur.

(c) Sulfuric acid reacts with carbon. The equation for the reaction is shown.

$$2H_2SO_4 + C \rightarrow CO_2 + 2H_2O + 2SO_2$$

Give the oxidation number of carbon in:

- C
- CO_2.

7. (2025·May/June 0620/13 Q16) · Level 2

Steam reacts with carbon to produce carbon monoxide and hydrogen.

$$C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)$$

Which substance is reduced in the reaction?

A C **B** CO **C** H_2 **D** H_2O

8. (2025·May/June 0620/12 Q18) · Level 4

Which reactions are redox reactions?

1 the incomplete combustion of propane

2 the rusting of iron

3 the reaction of aqueous chloride ions with acidified aqueous silver nitrate

4 the thermal decomposition of calcium carbonate

A 1 and 2 **B** 1 and 3 **C** 2 and 4 **D** 3 and 4

Acids, bases and salts

7.1 Acids and bases

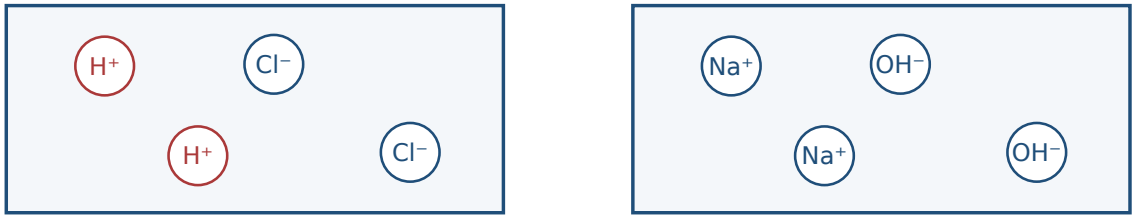
Type 1: Using ions, pH and indicators to interpret acid–base reactions

Key Ideas Compare two familiar aqueous solutions: hydrochloric acid, HCl(aq), and sodium hydroxide, NaOH(aq). Acid solutions contain hydrogen ions, H^+, and alkaline solutions contain hydroxide ions, OH^-. More precisely, H^+ predominates in an acidic solution and OH^- in an alkaline solution. "Contains H^+" does not mean that OH^- is completely absent.

Common bases in this section include metal oxides and hydroxides that react with acids, such as copper(II) oxide, CuO, and sodium hydroxide, NaOH. A base soluble in water is an alkali. NaOH dissolves in water to give Na^+ and OH^-, so it is an alkali. CuO is insoluble: it is a base but not an alkali. Later, proton acceptance will explain the basic behaviour of ammonia.

Aqueous acids and alkalis: identify the reacting ions

Hydrochloric acid, HCl(aq) | Sodium hydroxide, NaOH(aq)



Acidic solutions have relatively more H^+; alkaline solutions have relatively more OH^-.

Only the main solute ions are shown; water and the small amount of the other ion are omitted.

The main solute ions in HCl and NaOH solutions differ. Acidity does not mean OH^- is entirely absent, and alkalinity does not mean H^+ is entirely absent.

For example, the ion associated with hydrochloric acid's acidity is H^+, not Cl^-. The ion associated with NaOH solution's alkalinity is OH^-, not Na^+. Na^+ and Cl^- can coexist in a salt solution; their presence alone does not establish that it is strongly acidic or alkaline.

When an acid and alkali are mixed, H^+ can combine with OH^- to form water. This is neutralisation. The ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

The types and numbers of atoms and the total charge are conserved.

Neutralisation: H^+ and OH^- combine to form water



$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Atom counts: H = 2, O = 1; total charge on each side = 0.

OH^- accepts H^+ to become H_2O. This is proton transfer; the disappearance of ions alone does not make it reduction.

For example, the full equation for hydrochloric acid reacting with NaOH is:

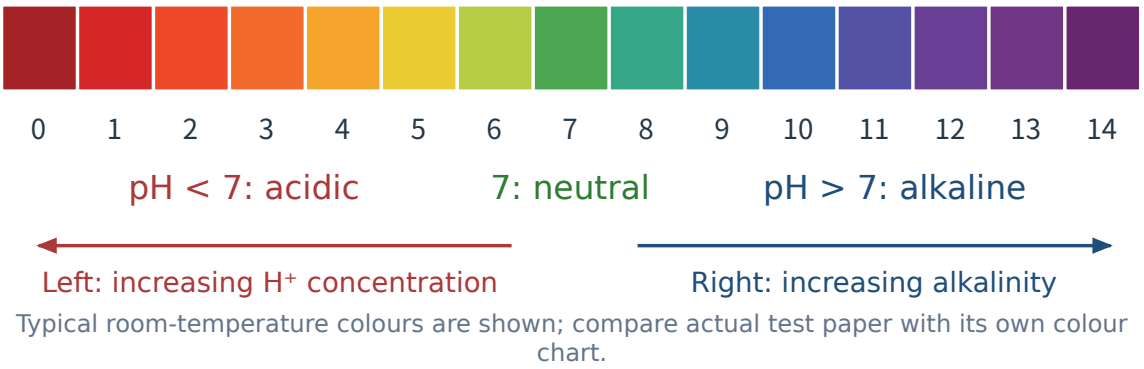
$$HCl + NaOH \rightarrow NaCl + H_2O$$

The salt is sodium chloride, NaCl. Na^+ and Cl^- remain in solution before and after reaction; the net change is H^+ combining with OH^- to form water. Adding NaOH gradually to hydrochloric acid consumes H^+, reduces acidity and raises pH. With insufficient NaOH, the mixture remains acidic. For this acid–alkali pair, exact neutralisation gives an approximately neutral solution. Excess NaOH makes it alkaline. Neutralisation occurring does not mean that every mixture, regardless of amounts, must end neutral.

A more general definition uses protons: an acid is a proton donor and a base is a proton acceptor. When an ordinary hydrogen atom loses its electron to form H^+, only a proton remains. Transferring a proton here means transferring H^+. In the diagram, OH^- accepts H^+ to form water, showing its basic behaviour. Electrons are not transferred between elements, so neutralisation should not simply be called reduction.

The pH expresses the relative acidity or alkalinity of an aqueous solution numerically. At room temperature in this section, pH 7 is neutral, below 7 is acidic and above 7 is alkaline. A lower pH means a higher H^+ concentration. A solution at pH 2 is more acidic than one at pH 5, but pH 5 is still not neutral. This compares solution acidity; pH alone does not establish whether the acid is a strong acid. Strength and concentration are distinguished later.

pH and universal indicator: locate 7, then compare either side



Typical room-temperature colours are shown; compare actual test paper with its own colour chart.

At room temperature, pH 7 is the neutral reference. Colours estimate ranges; this colour scale does not apply to every indicator.

For example, universal indicator is generally red, orange or yellow in clearly acidic solutions, green near neutrality, and increasingly blue to purple in alkaline solutions. A pH of 9 is alkaline, so do not describe it as yellow, which usually indicates an acidic range. In practice, compare the colour of wetted indicator paper with its supplied chart to estimate pH. Do not assign every blue colour the same exact pH.

Different indicators do not share one colour scale. This section distinguishes litmus, thymolphthalein and methyl orange. Their colours in the dilute hydrochloric acid and sodium hydroxide solutions commonly used in these questions are shown below.

Identify the indicator before choosing its colour

Indicator	Dilute HCl used here	NaOH(aq) used here
Litmus	**Red**	**Blue**
thymolphthalein	**Colourless**	**Blue**
methyl orange	**Red**	**Yellow**

Colourless is not white; universal indicator's red–green–purple scale does not apply to every indicator.

Each indicator changes colour over a particular pH range; not all transitions occur at pH 7.

The table compares clearly acidic dilute HCl and clearly alkaline NaOH solutions. For solutions close to an indicator's transition range, use the information for that indicator.

For example, in the same clearly alkaline NaOH solution, thymolphthalein is blue and methyl orange is yellow. The blue colour of one does not make the other blue. In dilute hydrochloric acid, thymolphthalein is colourless and methyl orange red. Colourless means no visible colour, not white. Each indicator has its own pH transition range; they do not all change exactly at pH 7. Near that range, use the supplied information rather than treating "acidic" or "alkaline" as sufficient for every detail.

An acid's products also depend on the other reactant. Suitable reactive metals, such as magnesium, Mg, and zinc, Zn, react with dilute hydrochloric or sulfuric acid to form a salt and H_2. Metal oxides or hydroxides react with acids to form a salt and water. Carbonates form CO_2 as well as a salt and water.

Three common acid reactions do not all produce the same gas

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Reactive metal + dilute acid — Bubbles; H_2 gives a squeaky pop with a lighted splint

$$CuO + 2HCl \rightarrow CuCl_2 + H_2O$$

Metal oxide + acid — Oxide is consumed; gas need not form

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

Carbonate + acid — Bubbles; CO_2 turns limewater milky

Identify products from the reactant category, then distinguish H_2 and CO_2 using the appropriate gas tests. Acids do not release hydrogen with every substance.

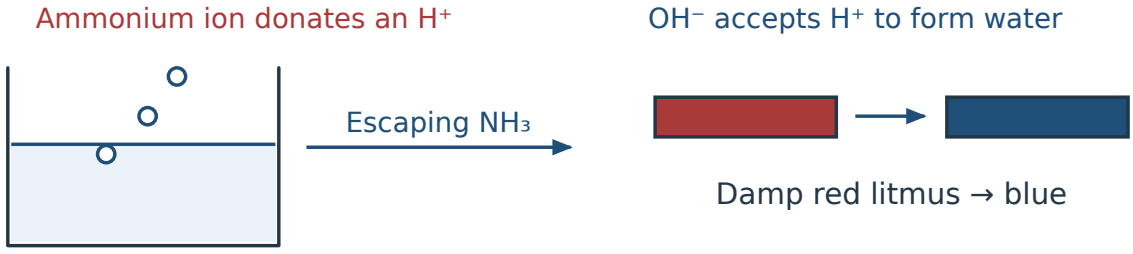
For example, consider magnesium, calcium carbonate and copper(II) oxide separately.
Magnesium is a suitable metal for reaction with dilute hydrochloric acid, producing magnesium chloride and hydrogen:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

The gas is hydrogen, which gives a squeaky pop with a lighted splint.
Calcium carbonate is a carbonate. With dilute hydrochloric acid it forms a salt, water and carbon dioxide:
$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$
The gas is carbon dioxide, which turns clear limewater milky.
Copper(II) oxide is a basic oxide. It reacts with dilute hydrochloric acid to form a salt and water:
$CuO + 2HCl \rightarrow CuCl_2 + H_2O$
This reaction does not require gas production. First distinguish metal, carbonate and base to determine the products. Not every metal releases H_2 with these dilute acids: Cu does not react with dilute HCl or H_2SO_4 in this way. Do not apply the hydrogen-production rule indiscriminately to nitric acid.
A base can also react with an ammonium salt to release ammonia, NH_3; gentle warming usually makes this easier to observe. Ammonium salts contain NH_4^+, as in ammonium chloride, NH_4Cl. The net ionic change is:
$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$
NH_4^+ loses H^+ to form NH_3, while OH^- accepts that H^+ to form water.

Alkali and ammonium salt: released ammonia turns damp red litmus blue

$$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

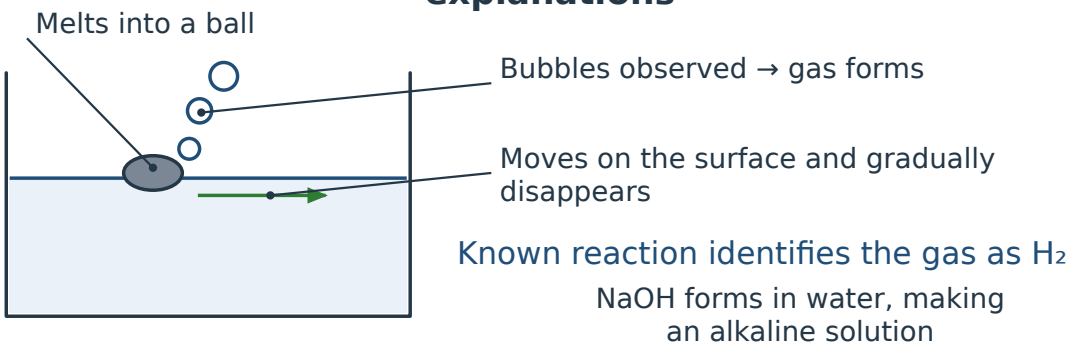


Hydroxides, or basic oxides that form alkalis, can release ammonia from ammonium salts.

First identify NH_3 from the reaction of ammonium ions with OH^-, then test its alkalinity using damp litmus. The red litmus must be damp.

For example, mix NH_4Cl with NaOH solution and warm gently:
$NH_4Cl + NaOH \rightarrow NaCl + NH_3 + H_2O$
The escaping NH_3 turns damp red litmus blue. The paper must be damp because ammonia dissolves in its water before producing alkaline conditions. A potassium salt mixed with a base does not contain NH_4^+ and cannot release NH_3 through this reaction.
In moist soil, basic oxides such as calcium oxide, CaO, can create alkaline conditions that release NH_3 from ammonium salts. For example:
$CaO + 2NH_4Cl \rightarrow CaCl_2 + 2NH_3 + H_2O$
Nitrogen originally present in the ammonium salt is then lost as escaping ammonia. This explains why such alkaline materials should not be added at the same time as ammonium fertilisers. Conversely, NH_3 can accept H^+ from water, producing NH_4^+ and OH^-:
$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
Aqueous ammonia is therefore alkaline, and NH_3 meets the definition of a proton acceptor.
Finally, connect this to a Group I reaction. Sodium, Na, reacts with water to form NaOH and H_2:
$2Na + 2H_2O \rightarrow 2NaOH + H_2$
The NaOH solution left after reaction is alkaline. The fact that the added substance was initially a metal does not make the final solution non-alkaline.

Sodium and water: separate observations from explanations



Bubbles are observed; hydrogen is identified using the reaction or a gas test.

Na reacts with water to form NaOH and H_2. Use this specific reaction to identify the final solution as alkaline before deciding the indicator colour.

For example, if an indicator is added after Na finishes reacting with water, first identify NaOH in the final solution, then the alkalinity caused by OH^-, and finally the colour of the named indicator. Methyl orange is yellow. If the indicator is thymolphthalein, use its alkaline colour instead of copying methyl orange's yellow.

The net ionic equation for neutralisation is H^+ + _____ → _____. A solution at pH 5 is _____ acidic than one at pH 2. Ammonia, which turns damp red litmus blue, is commonly released when a base reacts with an _____ salt.

Answers: OH^-; H_2O; less; ammonium

Recognising the Question Identify the actual solution or final product before assessing acidity, alkalinity, pH and the named indicator. For gases, determine the gas from the reactant category first, then check it with the appropriate test.

Worked Example 1 (2025·May/June 0620/11 Q17) · Level 2

Which row gives the colours observed when thymolphthalein and methyl orange are added separately to the named solution?

	solution	colour with thymolphthalein	colour with methyl orange
A	dilute HC*l*	colourless	yellow
B	dilute HC*l*	blue	red
C	aqueous NaOH	colourless	red
D	aqueous NaOH	blue	yellow

Solution

Classify the solutions first: dilute HCl is acidic and aqueous NaOH is alkaline. Then check each indicator separately; they do not share a colour scale.

In dilute HCl, thymolphthalein is colourless and methyl orange red. A has the correct colourless entry but incorrectly says yellow. B correctly gives red for methyl orange but incorrectly gives blue for thymolphthalein.

In NaOH solution, thymolphthalein is blue and methyl orange yellow. Both entries in C are wrong and both in D are correct. Choose D.

Worked Example 2 (2025·May/June 0620/22 Q18) · Level 2

A farmer incorrectly adds two substances to the soil at the same time.

They react together to form a gas which turns damp red litmus paper blue.

What are the two substances?

A a basic oxide and a potassium salt

B a basic oxide and an ammonium salt

C an acidic oxide and a potassium salt

D an acidic oxide and an ammonium salt

Solution

A gas turning damp red litmus paper blue gives the characteristic test for ammonia, NH_3. Identify the gas from the observation, then find the reaction that produces it.
Ammonium salts contain NH_4^+ and release NH_3 with a base:

$$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

Under damp conditions, basic oxides such as CaO create alkaline conditions, so the combination basic oxide and ammonium salt fits.
For example:

$$CaO + 2NH_4Cl \rightarrow CaCl_2 + 2NH_3 + H_2O$$

A potassium salt supplies no NH_4^+, and an acidic oxide does not provide the alkaline conditions needed. Choose B. This also explains why adding both together causes nitrogen in ammonium fertiliser to escape as ammonia.

Practice

1. (2025·May/June 0620/12 Q19) · Level 2

Which row describes an aqueous solution for the pH shown?

	pH	description
A	1	has a very high alkalinity
B	5	is neutral
C	9	turns universal indicator yellow
D	14	contains OH^- ions

2. (2025·May/June 0620/13 Q18) · Level 2

A solution containing hydroxide ions, $OH^-(aq)$, is added to a solution containing hydrogen ions, $H^+(aq)$. An equation representing the reaction is shown.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Which statement about the reaction is correct?

A The hydrogen ions represent an alkali.

B The reaction is a neutralisation reaction.

C The reaction is a reduction.

D As the OH^- ions are added, the pH of the reaction mixture falls.

3. (2025·May/June 0620/41 Q6(e)(ii)) · Level 4

This question is about the Periodic Table.

(e) Sodium is in Group I of the Periodic Table.

When sodium is added to water a chemical reaction occurs.

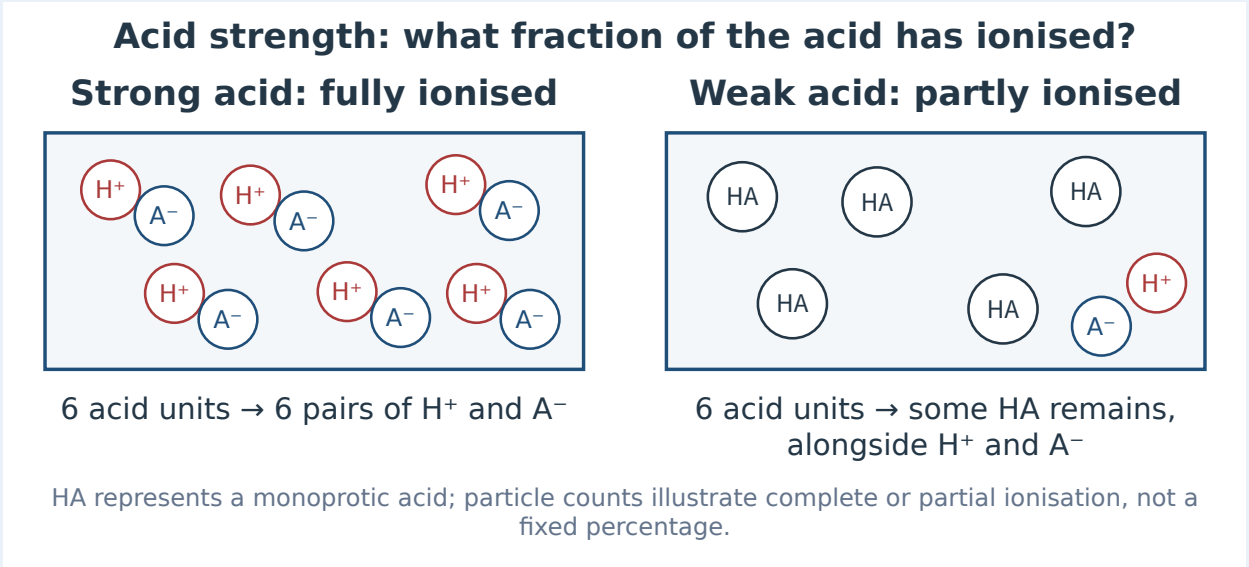
(ii) Thymolphthalein is added to the solution when the reaction has finished.

State the final colour of the thymolphthalein in the solution.

Type 2: Distinguishing acid strength from concentration using dissociation

Key Ideas A sample of acid raises two different questions. First, when the acid enters water, do all its molecules form ions? Second, how much acid is present per unit volume of solution? The first determines acid strength; the second determines concentration.

Use HA as a model monoprotic acid. HA represents one acid molecule. After donating one H^+, its remaining part is written A^-. Formation of H^+ and A^- in water is called dissociation. The diagram uses HA, H^+ and A^- to show whether undissociated molecules remain; water molecules are omitted.



The boxes have equal volumes and equal initial amounts of acid. On the left all acid forms ions; on the right undissociated HA remains. Water molecules are omitted. A^- represents the anion left after loss of H^+.

A strong acid is completely dissociated in aqueous solution; a weak acid is partially dissociated. For example, all 6 original HA units on the left become H^+ and A^-, meeting the strong-acid definition. On the right, HA, H^+ and A^- coexist, showing partial dissociation and therefore a weak acid. Do not say that the right-hand solution contains no H^+ or that a weak acid cannot react with a base.

The course explicitly requires two examples. Hydrochloric acid, aqueous HCl, is a strong acid. Ethanoic acid, CH_3COOH, is weak. A single arrow → represents complete dissociation, while a reversible arrow $\rightleftharpoons$ represents partial dissociation.

Ionisation equations: the ions may be similar, but the arrows have different meanings

Hydrochloric acid: strong; fully ionised

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Ethanoic acid: weak; partly ionised

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

Ethanoic acid loses the acidic H shown; all three H atoms in CH_3 remain in CH_3COO^-.

HCl uses a single arrow for complete dissociation. CH_3COOH uses a reversible arrow to show ions and undissociated molecules coexisting. Both produce H^+; weak does not mean unable to dissociate.

For example, first read:

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

One H and one Cl occur on each side. Charges +1 and −1 total 0, matching the left. The single arrow represents complete dissociation of a strong acid in this chapter.

Next read:

$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

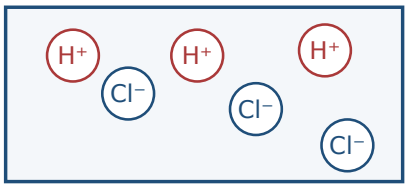
One acidic H is transferred; the three H atoms in CH_3 remain in the anion. The reversible arrow shows that molecules and ions coexist in water, with dissociation and recombination occurring in opposite directions. In a table, check the HCl and CH_3COOH columns separately. One correct column does not make the whole row correct.

Dissolving and dissociation are different processes. Complete dissolving means the acid has entered the solution uniformly; it does not prove that every molecule has formed ions. If the question says completely dissolves, also read the stated dissociated amount.

A concentrated solution contains more total solute per unit volume; a dilute solution contains less. Total acid includes both undissociated molecules and the acid represented by its ions, not merely H^+. Adding water to hydrochloric acid increases volume, reduces H^+ per unit volume and raises pH, but HCl remains completely dissociated in this section's model.

Concentration is amount per unit volume; strength is the extent of ionisation

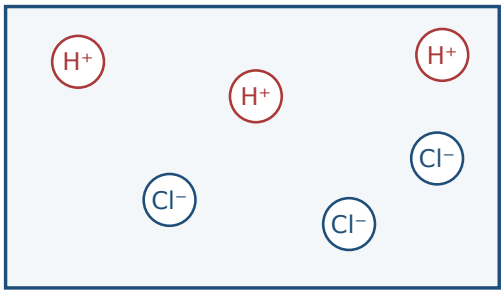
Before dilution: hydrochloric acid



3 ion pairs in the original volume

Add water

After adding water: more dilute hydrochloric acid



The same 3 ion pairs in a larger volume

H^+ concentration falls and the solution is less acidic, but HCl remains fully ionised: dilution does not make a weak acid.

Adding water to the same hydrochloric acid changes solute amount per unit volume. In this chapter's model, HCl remains completely dissociated. The volume change illustrates dilution and is not intended for numerical measurement.

For example, both before and after dilution the diagram contains 3 H^+ and 3 Cl^- ions. Added water distributes them through a larger volume, so the hydrochloric acid becomes more dilute and less acidic. Neither model retains undissociated HCl molecules, so dilution does not mean that hydrochloric acid becomes a weak acid. Similarly, a concentrated weak acid can remain only partially dissociated. A pH value alone, without concentration and other relevant conditions, cannot establish whether an acid is strong or weak.

When amounts in mol are given, identify the initial total amount of acid and the amount dissociated. Compare each acid with its own total. Equality means complete dissociation; a smaller dissociated amount means partial dissociation. There is no need to convert every mol value into particle numbers first.

When moles are given, compare the amount ionised with the original total

Solution	Original total amount of acid	Amount ionised	Compare the amounts	Then conclude
Example A	0.200 mol	0.008 mol	0.008 < 0.200	Partial → weak acid
Example B	0.200 mol	0.200 mol	0.200 = 0.200	Complete → strong acid

A: fraction ionised = 0.008 ÷ 0.200 = 0.04 = 4%, not 100%.

All the acid can be dissolved while only some acid molecules are ionised.

Compare each sample with its own original total. A greater dissociated amount does not automatically mean complete dissociation. Percentages provide further explanation, but checking equality already distinguishes complete from partial dissociation.

For example, sample A starts with 0.200 mol acid and 0.008 mol dissociates. This is less than the total, so dissociation is incomplete and the acid is weak. For the dissociated fraction, divide the dissociated amount by the original total:

$$\text{Dissociated fraction} = \frac{0.008}{0.200} = 0.04 = 4\%$$

Sample B has both an initial amount and a dissociated amount of 0.200 mol. It is completely dissociated and therefore strong.

A third acid might start with 0.200 mol and have 0.050 mol dissociated. Although more dissociates than in sample A, it is still not complete, so this acid is also weak. Always compare with that acid's own total; the larger of two dissociated amounts is not automatically a strong acid.

Answers: Strength depends on complete or partial dissociation; concentration depends on total solute per unit volume. Use $\rightarrow$ for HCl and $\rightleftharpoons$ for CH_3COOH. For numerical data, compare each dissociated amount with that acid's original total.

Recognising the Question Distinguish dissolving from dissociating. For definitions, check complete or partial; for equations, check each column's arrow; for data, compare each acid with its own total.

Worked Example 1 (2024·October/November 0620/22 Q19) · Level 1

Which row shows the difference between a weak acid and a strong acid?

	weak acid	strong acid
A	fully dissociated	partially dissociated
B	concentrated	dilute
C	dilute	concentrated
D	partially dissociated	fully dissociated

Solution

Separate the weak acid and strong acid columns.

A weak acid partially dissociates in water, so its column requires partially dissociated. A strong acid completely dissociates, so its column requires fully dissociated. Only D meets both conditions.

A reverses them. B and C confuse concentration with strength: a dilute solution can contain a strong acid, and a concentrated solution can contain a weak acid.

Choose D.

Practice

1. (2025·May/June 0620/22 Q19) · Level 2

Which row shows the equations for the dissociation of hydrochloric acid and of ethanoic acid in aqueous solution?

	hydrochloric acid	ethanoic acid
A	$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$	$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
B	$HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$	$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
C	$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$	$CH_3COOH(aq) \rightarrow H^+(aq) + CH_3COO^-(aq)$
D	$HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$	$CH_3COOH(aq) \rightarrow H^+(aq) + CH_3COO^-(aq)$

2. (2025·May/June 0620/21 Q18) · Level 3

Two experiments involving acids R or S are done.

- 0.10 mol of acid R completely dissolves in 500 cm^3 of water. 0.0048 mol of acid R is dissociated in the solution.
- 0.10 mol of acid S completely dissolves in 500 cm^3 of water. 0.0179 mol of acid S is dissociated in the solution.

Which statement about acids R and S is correct?

A Acid R is a strong acid, and acid S is a weak acid.

B Acid R is a weak acid, and acid S is a strong acid.

C Acid R and acid S are both strong acids.

D Acid R and acid S are both weak acids.

7.2 Oxides

Type 1: Classifying acidic, basic and amphoteric oxides by their reactions

Key Ideas An oxide is a compound of oxygen with one other element. Names containing oxide or dioxide and formulae containing O help identify candidates, but not every oxygen-containing substance is an oxide. Carbonates, for example, contain carbonate ions; salts and oxides are different categories.

Acidic or basic classification depends on acid–base reactions. A basic oxide reacts with an acid to form a salt and water, as copper(II) oxide, CuO, does with HCl. An acidic oxide reacts with an alkali to form a salt and water, as carbon dioxide, CO_2, does with NaOH.

Classify by reaction: basic and acidic oxides

Basic oxide: CuO	Acidic oxide: CO_2
$CuO + 2HCl \rightarrow CuCl_2 + H_2O$	$CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$
Reacts with acid to form salt and water	Reacts with alkali to form salt and water

Acidic and basic describe oxide reactions; an oxide need not dissolve in water to be classified.

Although insoluble in water, CuO is a basic oxide because it reacts with acids. Acidic oxides react with alkalis.

For example:

$CuO + 2HCl \rightarrow CuCl_2 + H_2O$

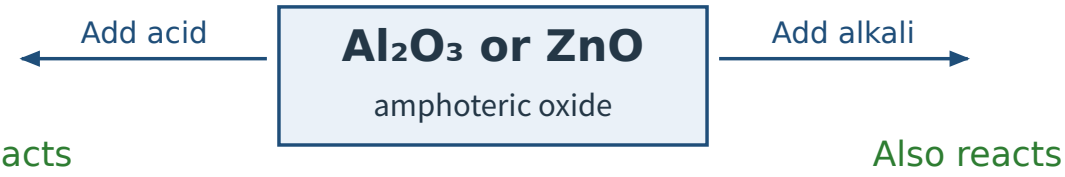
CuO is therefore a basic oxide. Its insolubility in water does not prevent reaction with an acid. A basic oxide need not dissolve in water to form an alkali.

$CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$

This reaction with an alkali identifies CO_2 as an acidic oxide. The product sodium carbonate, Na_2CO_3, is a salt and should not be placed in the same category as CO_2.

Metal oxides are usually basic and non-metal oxides are often acidic. These are useful initial trends, but specific exceptions matter. The two key examples here are aluminium oxide, Al_2O_3, and zinc oxide, ZnO. They react with both acids and alkalis to form the corresponding salts and water and are called amphoteric oxides.

Amphoteric oxides react with both acids and alkalis



The corresponding salt and water form; a metal oxide is not automatically basic.

Amphoteric ≠ neutral; amphoteric does not mean unreactive.

Al_2O_3 and ZnO are the amphoteric oxides to recognise in this section. Both types of reaction must be possible.

For example, although Al is a metal, Al_2O_3 reacts with both acids and alkalis. “Metal plus oxygen” alone does not justify classifying it as an ordinary basic oxide. ZnO also reacts in both ways. Amphoteric means reactive towards both acids and alkalis, not neutral or unreactive towards both.

For questions giving names or formulae, identify the specific substance first and use composition as a check. Basic examples include sodium oxide, Na_2O, magnesium oxide, MgO, calcium oxide, CaO, barium oxide, BaO, and CuO. Acidic examples include CO_2, sulfur dioxide, SO_2, nitrogen dioxide, NO_2, and silicon dioxide, SiO_2. Keep Al_2O_3 and ZnO in the amphoteric category.

Identify the oxide, then check its specific class

Class	Examples needed for this type and its practice questions
Basic	Na_2O, MgO, CaO, BaO, CuO
Acidic	CO_2, SO_2, NO_2, SiO_2
Amphoteric	Al_2O_3, ZnO

For all three acidic, check each oxide; one basic or amphoteric oxide excludes the group.

The specific examples show the usual metal/non-metal trend alongside the essential amphoteric exceptions Al_2O_3 and ZnO, supporting item-by-item decisions.

For example, in a group containing CaO, CO_2 and SO_2, CO_2 and SO_2 are acidic but CaO is basic, so the group is not all acidic. “All” requires every item to qualify; do not stop after finding two familiar acidic oxides. When selecting a basic oxide from a long list, first identify oxides, then look for a known basic example such as CaO or CuO. Chlorine, oxygen and graphite are elements, not oxides.

The list contains specific known examples; it does not turn a trend into an exceptionless rule. Carbon monoxide, CO, is not an acidic oxide, showing why non-metal oxides must still be checked individually.

CuO reacts with _____, so it is a basic oxide. Al_2O_3 and ZnO react with _____ acids and alkalis, so they are _____ oxides.

Answers: acids; both; amphoteric

Recognising the Question First check whether each candidate is an oxide, then classify it by its specific reactions. For all, check every item. Retain Al_2O_3 and ZnO as amphoteric exceptions.

Worked Example 1 (2025·May/June 0620/13 Q19) · Level 1

Which compound is an acidic oxide?

A barium oxide

B carbon dioxide

C copper(II) oxide

D magnesium oxide

Solution

Carbon dioxide, CO_2, is an acidic oxide: it reacts with an alkali to form a salt and water. Choose B.
For example:
$CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$
Barium oxide, BaO, copper(II) oxide, CuO, and magnesium oxide, MgO, are familiar basic oxides that react with acids.

Worked Example 2 (2025·May/June 0620/43 Q1(c)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(c) is an amphoteric oxide

Solution

The two amphoteric oxides to know here are Al_2O_3 and ZnO. The list contains aluminium oxide but not zinc oxide, so select aluminium oxide.
Al_2O_3 reacts with both acids and alkalis, meeting the definition of amphoteric. Do not classify it as an ordinary basic oxide simply because Al is a metal.

Practice

1. (2025·May/June 0620/41 Q1(a)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(a) is a basic oxide

2. (2025·May/June 0620/22 Q20) · Level 2

Which three oxides are **all** acidic?

A CaO, NO_2, SO_2

B CaO, CO_2, Na_2O

C CO_2, NO_2, SO_2

D CO_2, Na_2O, SO_2

7.3 Salt preparation and water of crystallisation

Type 1: Choosing reactants for soluble salts and explaining excess, filtration, titration and crystallisation

Key Ideas Begin salt preparation with two questions: which ions make up the target salt, and is it soluble in water? In a salt, the H^+ of an acid is replaced by metal ions, NH_4^+ or similar cations. Hydrochloric acid forms chlorides, sulfuric acid, H_2SO_4, forms sulfates, and nitric acid, HNO_3, forms nitrates.

Soluble means an appreciable amount dissolves in water under the given conditions. Insoluble here means a very low solubility, allowing a solid to remain or form and be filtered off; it does not mean that not a single particle dissolves. The common rules needed to choose a method are shown below.

Check the salt's water solubility before choosing a separation method

Type of salt or base	Usual solubility and exceptions
Na^+, K^+ and NH_4^+ salts	Soluble
Nitrates, NO_3^-	Soluble
Chlorides, Cl^-	Usually soluble; AgCl and $PbCl_2$ are insoluble
Sulfates, SO_4^{2-}	Usually soluble; exceptions: $BaSO_4$, $CaSO_4$ and $PbSO_4$
Carbonates, CO_3^{2-}	Usually insoluble; exceptions: Na^+, K^+ and NH_4^+ salts
Hydroxides	Usually insoluble; NaOH and KOH dissolve; $Ca(OH)_2$ is sparingly soluble

Aqueous ammonia also contains OH^-; ammonium hydroxide describes the solution, not a solid to filter off.

These are the common solubility rules used in this section. Calcium sulfate is listed as an exception to sulfate solubility in the syllabus; calcium hydroxide is only slightly soluble.

For example, $CuSO_4$ is not one of the Ba^{2+}, Ca^{2+} or Pb^{2+} sulfate exceptions, so crystals can be obtained from its aqueous solution. $CuCO_3$ is not a Na^+, K^+ or NH_4^+ carbonate, so it is insoluble and excess solid can be filtered off. $Cu(OH)_2$ is also insoluble: a base but not a soluble alkali. $CaSO_4$ is a sulfate exception under these rules, while $Ca(OH)_2$ is only slightly soluble. Do not confuse the two calcium compounds.

Ammonia in water forms NH_4^+ and OH^- and is sometimes described as ammonium hydroxide solution. This refers to aqueous ammonia, not a solid NH_4OH that can be filtered from water. Ammonium salts such as NH_4Cl and $(NH_4)_2SO_4$ are treated as soluble.

To choose reactants, split the target salt into its cation and anion. Copper(II) sulfate, $CuSO_4$, requires Cu^{2+} and SO_4^{2-}. The anion calls for H_2SO_4, while the copper source must actually react with that dilute acid.

Split the target salt into ions: where do Cu^{2+} and sulfate come from?

Target: copper(II) sulfate, $CuSO_4$

Cu^{2+}

From a copper-containing reactant

SO_4^{2-}

From sulfuric acid, H_2SO_4

$CuCO_3$, CuO and $Cu(OH)_2$ react with dilute H_2SO_4

Copper metal does not react with dilute H_2SO_4 to release H_2

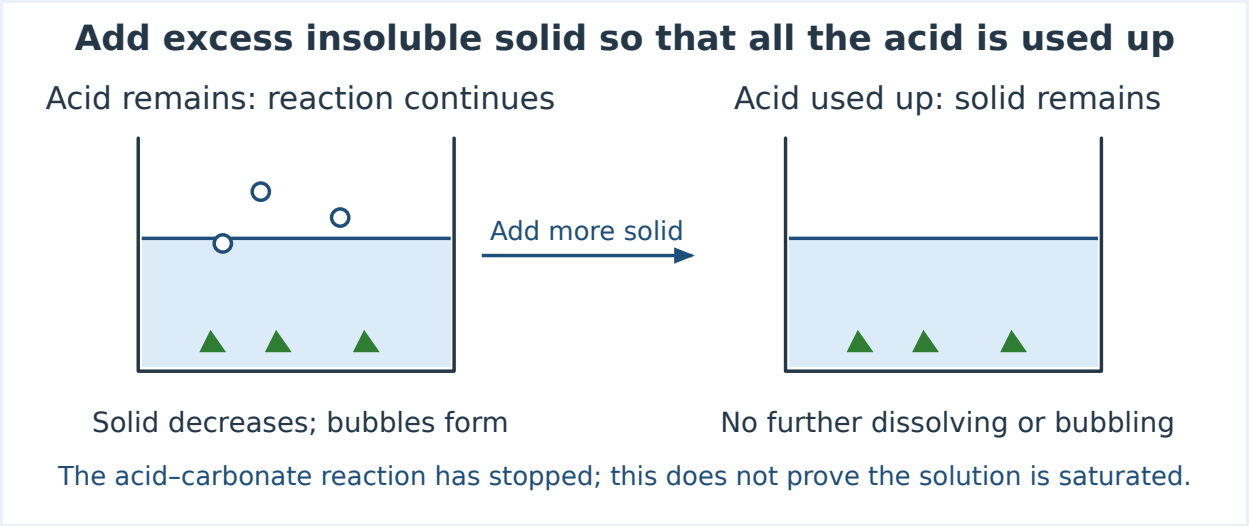
$Cu(OH)_2$ is insoluble: it is a base, but not an alkali.

Identify the target salt's cation and anion, then select reactants that genuinely react. The acid + metal route does not always work for copper.

For example, copper(II) sulfate can be prepared using three copper-containing reactants.

Copper(II) carbonate, $CuCO_3$, reacts with dilute sulfuric acid:
$CuCO_3 + H_2SO_4 \rightarrow CuSO_4 + CO_2 + H_2O$
Copper(II) oxide, CuO, reacts with dilute sulfuric acid:
$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
Copper(II) hydroxide, $Cu(OH)_2$, reacts with dilute sulfuric acid:
$Cu(OH)_2 + H_2SO_4 \rightarrow CuSO_4 + 2H_2O$
All three supply copper, while sulfuric acid supplies sulfate. Copper metal does not react with dilute H_2SO_4 by releasing H_2. A copper-containing target salt does not by itself justify choosing acid plus copper metal.

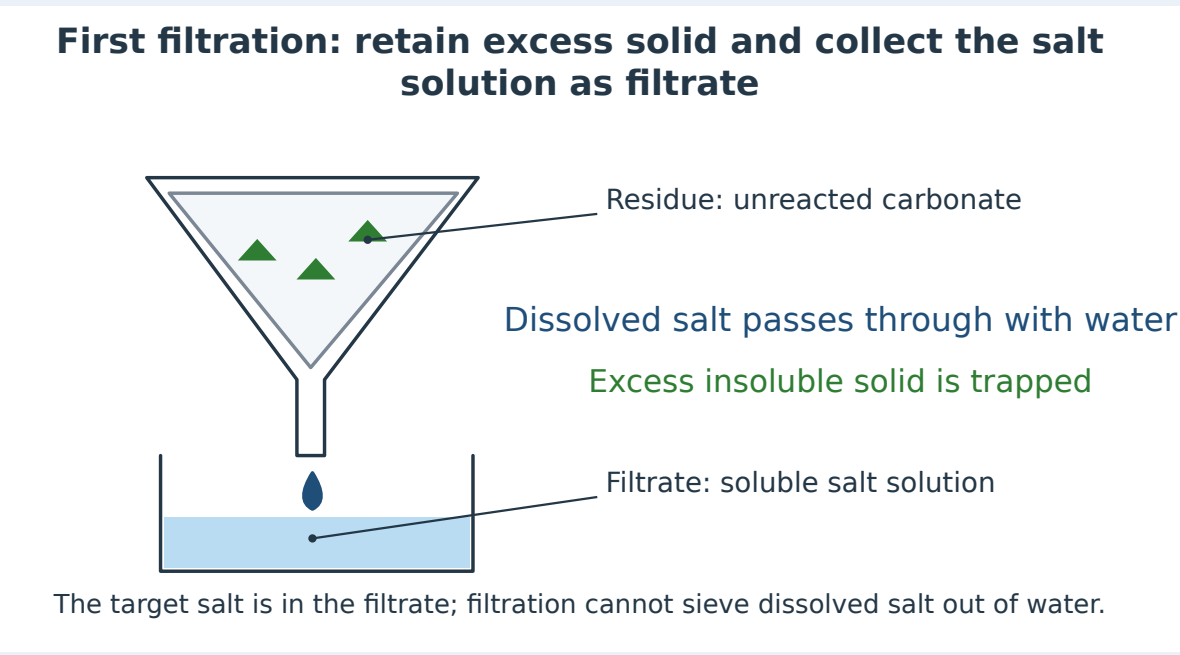
Add an insoluble base or carbonate in portions with stirring until the solid is in excess. Excess does not mean that the initial quantity merely looks large. It means that, after reaction, the acid is exhausted and some solid remains unreacted.



Solid remaining after continued addition and stirring, with no further bubbles, shows that the acid is consumed and the carbonate is in excess.

For example, initial additions of $CuCO_3$ disappear with CO_2 bubbles, forming blue $CuSO_4$ solution. Continue adding and stirring until newly added solid no longer disappears and no more bubbles form. The acid has then been used up. Solid remains, so it is in excess; the reaction has not stopped because the solid ran out. Do not automatically call this saturation. Exhaustion of acid concerns reactant amounts; saturation concerns the maximum salt that can dissolve in water.

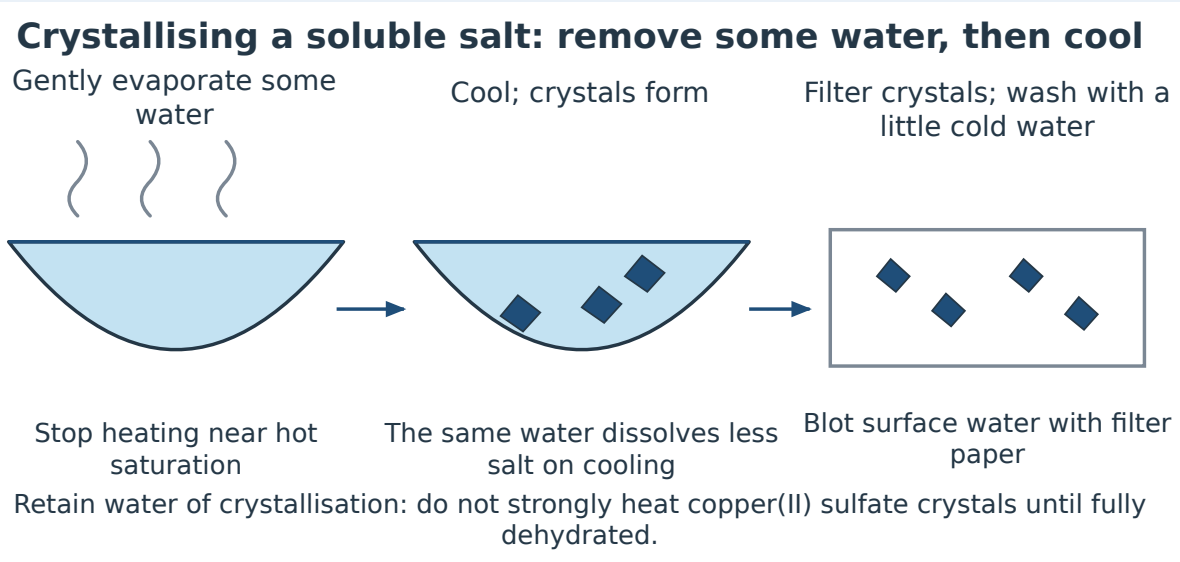
Excess solid would contaminate the product, so filter next. The filter paper retains the insoluble remaining solid while water and dissolved salt pass through and are collected below.



Excess carbonate stays on the paper. The filtrate below contains the target salt, which is subsequently crystallised.

For example, the residue on the paper is excess $CuCO_3$, and the filtrate in the beaker contains $CuSO_4$. Keep the filtrate because the target salt remains dissolved. Discarding it and keeping only the green carbonate would lose the desired copper(II) sulfate. Filtration cannot trap dissolved Cu^{2+} and SO_4^{2-} on paper as it would sand grains.

Next obtain crystals from the filtrate. The dissolved salt is the solute and water is the solvent. A saturated solution contains, at a specified temperature, the maximum amount of solute that can dissolve in that amount of solvent. Include “at a specified temperature”, because solubility may change with temperature.



Copper(II) or zinc sulfate solution is gently concentrated and then cooled to form crystals. Collecting these crystals is a different stage from the earlier collection of filtrate.

For example, gently heat the $CuSO_4$ filtrate to evaporate some water and increase concentration. Stop heating near hot saturation and allow it to cool. Less $CuSO_4$ can remain dissolved at the lower temperature, so the excess crystallises. Filter and keep the crystals, wash with a little cold distilled water to remove surface mother liquor, and blot away surface water with filter paper. Too much wash water, or warm water, dissolves more product and reduces yield.

Copper(II) sulfate crystals may contain water of crystallisation, as in $CuSO_4 \cdot 5H_2O$, explained later. Drying here removes surface water; it does not mean strongly heating the crystals until fully dehydrated. Do not replace

partial evaporation followed by cooling with unconditional heating to dryness. In the first filtration, keep the salt-containing filtrate. In the final filtration, keep the crystals. The same operation retains different parts at different stages.

The same procedure applies to the analogous zinc reaction. Zinc carbonate, $ZnCO_3$, reacts with dilute sulfuric acid:

$$ZnCO_3 + H_2SO_4 \rightarrow ZnSO_4 + CO_2 + H_2O$$

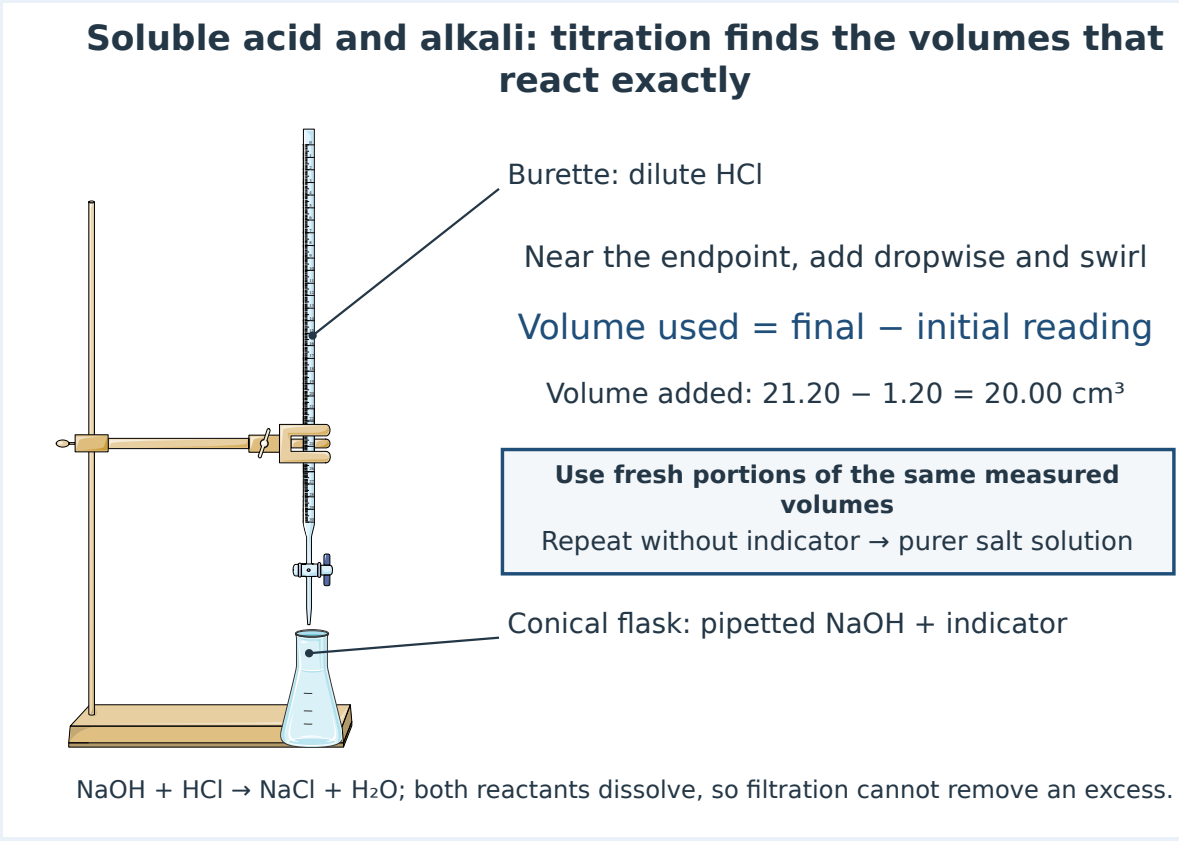
When the acid is exhausted, $ZnCO_3$ remains and bubbling stops. Filter, keep the zinc sulfate, $ZnSO_4$, solution, then concentrate, cool and crystallise. ZnO or $Zn(OH)_2$ can also supply zinc by reacting with sulfuric acid. Recheck the target salt and reactants; the separation reasoning remains the same.

A suitable metal provides another route using excess solid. For example:

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

Filter off excess unreacted Zn, then crystallise $ZnSO_4$ from the filtrate. Both conditions must hold: the metal reacts with that dilute acid, and the salt produced is soluble. Do not simply replace Zn with Cu.

If the other reactant is a soluble alkali such as NaOH, excess NaOH also remains dissolved with the salt and cannot be removed by filtration. Use titration to determine the exact reacting volumes, then use that ratio to prepare a purer salt solution.



Titration first establishes the reacting volumes. Repeat with fresh reagents using the same volumes but no indicator, preventing indicator contamination of the salt.

For example, to prepare NaCl, use a pipette to transfer 25.0 cm³ NaOH solution to a conical flask. Add a few drops of a suitable indicator, such as methyl orange. Fill a burette with dilute HCl and record its initial reading. Add acid while swirling, then dropwise near the endpoint. Stop at a persistent indicator colour change and record the final reading.

Suppose the initial reading is 1.20 cm³ and the final reading 21.20 cm³. Burette readings increase as liquid is delivered downwards, so volume added is final minus initial:

$$\text{Acid volume} = 21.20 - 1.20 = 20.00 \text{ cm}^3$$

Do not use 21.20 directly as the volume added. The endpoint is identified using the chosen indicator's colour change; not every indicator endpoint is fixed at pH 7.

For salt uncontaminated by indicator, take another 25.0 cm³ NaOH and add the established 20.00 cm³ HCl without indicator. The reaction:

$$NaOH + HCl \rightarrow NaCl + H_2O$$

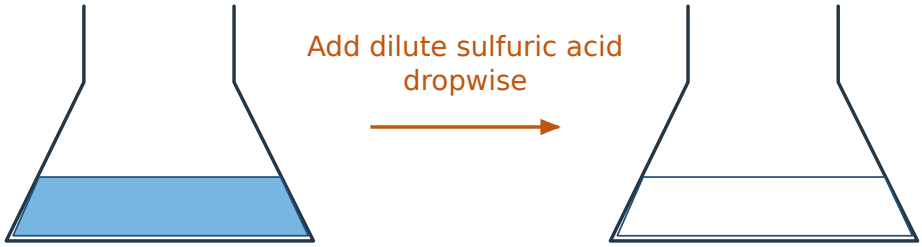
produces a salt solution, which is then concentrated by evaporation, crystallised, separated and dried. Do not automatically use the indicator-containing titration mixture as pure product, or add arbitrary excess soluble NaOH to ensure all the acid reacts.

When preparing Na_2SO_4 from dilute H_2SO_4 and NaOH, preserve the actual direction: acid in the burette and alkali in the flask. The equation is:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

Thymolphthalein is blue in the initial alkali and becomes colourless as acid is added to the endpoint.

Adding sulfuric acid to NaOH: thymolphthalein changes from blue to colourless



Initially: NaOH + indicator

Endpoint: colour disappears permanently

Use the indicator to find the endpoint, then repeat those volumes without indicator to prepare the salt.

Acid is added to a flask containing NaOH, so thymolphthalein changes from blue to colourless. The direction follows from starting with alkali and adding acid. The colourless region represents a transparent solution.

For example, stop adding acid when the blue colour just disappears persistently, then record the volume. Universal indicator passes through several colours, making a single sharp endpoint difficult to identify, so it is unsuitable here. If Na_2SO_4 solution is already supplied, gently heat to near saturation or the crystallisation point, cool to form crystals, filter to collect them, and blot dry with filter paper. “Evaporate” alone does not explain when to stop, how crystals form or how they are dried.

Choosing a salt preparation route: can the excess reagent be filtered off?

Starting material reacting with acid	Can it be added in excess?	How to remove excess material
Suitable metal, e.g. Zn	Yes, when the salt is soluble	Filter off unreacted metal
Insoluble base, e.g. CuO	Yes	Filter off insoluble solid
Insoluble carbonate, e.g. $ZnCO_3$	Yes	Filter off insoluble solid
Soluble alkali, e.g. NaOH	Filtration cannot remove the excess	Titrate, then repeat the measured volumes without indicator

Whether excess reagent can be removed by filtration determines the choice between an excess-solid method and titration.

For example, to choose between the methods, ask whether surplus reagent can be filtered off. Remaining CuO, $ZnCO_3$ or suitable Zn metal can be removed as solids. NaOH and acids are soluble, so any excess remains with the target salt in the filtrate; establish the required amounts first. In either route, identify whether the target salt is still in solution or has become crystals before deciding which part to collect.

When acid is exhausted and carbonate remains, the carbonate is in _____. At the first filtration, keep the salt-containing _____. A saturated solution contains the maximum solute soluble at a specified _____. To prepare a pure salt by titration, establish the volumes and repeat without _____.

Answers: excess; filtrate; temperature; indicator

Recognising the Question Identify the ion sources for the target salt, then check whether the reaction occurs and whether excess reagent can be filtered off. Explain which material is kept at each stage. Do not confuse acid exhaustion with solution saturation.

Worked Example 1 (2025·May/June 0620/41 Q3(c)(ii)) · Level 2

This question is about copper and compounds of copper.

(c) Copper(II) sulfate crystals are made by the reaction between copper(II) carbonate and dilute sulfuric acid, using the following steps.

The sulfuric acid has a concentration of $0.100\,mol/dm^3$.

step 1 Powdered copper(II) carbonate is added to dilute sulfuric acid. The mixture is stirred. A reaction occurs.

step 2 More copper(II) carbonate is added, with stirring, until the reaction stops.

step 3 Unreacted copper(II) carbonate is separated from aqueous copper(II) sulfate by filtration.

step 4 Aqueous copper(II) sulfate is heated until some of the water evaporates.

step 5 The remaining solution is allowed to cool and crystallise.

step 6 The crystals are removed and dried.

(ii) State why the reaction stops in **step 2**.

Solution

Step 2 continues adding $CuCO_3$ with stirring until reaction stops. More solid can still be added but no longer reacts, showing that the other reactant, H_2SO_4, has all been consumed.

Answer: all the sulfuric acid has reacted. $CuCO_3$ is then in excess, and its remaining portion is filtered off in step 3.

Do not say all the copper(II) carbonate has been used up: step 3 explicitly contains unreacted copper(II) carbonate. Nor should the reason be that the copper(II) sulfate solution is saturated; saturation and consumption of acid are different issues.

Worked Example 2 (2025·May/June 0620/41 Q3(c)(iv)) · Level 2

This question is about copper and compounds of copper.

(c) Copper(II) sulfate crystals are made by the reaction between copper(II) carbonate and dilute sulfuric acid, using the following steps.

The sulfuric acid has a concentration of 0.100 mol / dm^3.

step 1 Powdered copper(II) carbonate is added to dilute sulfuric acid. The mixture is stirred. A reaction occurs.

step 2 More copper(II) carbonate is added, with stirring, until the reaction stops.

step 3 Unreacted copper(II) carbonate is separated from aqueous copper(II) sulfate by filtration.

step 4 Aqueous copper(II) sulfate is heated until some of the water evaporates.

step 5 The remaining solution is allowed to cool and crystallise.

step 6 The crystals are removed and dried.

(iv) Name a substance, other than copper(II) carbonate, that can be added to dilute sulfuric acid to produce aqueous copper(II) sulfate.

Solution

The target remains $CuSO_4$. Sulfuric acid supplies SO_4^{2-}, so the other reactant must contain copper and react with dilute H_2SO_4.
Copper(II) oxide, CuO, can be used:
$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
Copper(II) hydroxide, $Cu(OH)_2$, can also be used:
$Cu(OH)_2 + H_2SO_4 \rightarrow CuSO_4 + 2H_2O$
Either answer is sufficient.

Do not substitute Cu metal, which does not react with this dilute sulfuric acid by the hydrogen-producing route. A base containing no copper would form a different metal's salt.

Worked Example 3 (2024·October/November 0620/43 Q4(c)) · Level 3

(a) Dilute sulfuric acid and aqueous sodium hydroxide can be used to prepare sodium sulfate crystals using a method that involves titration.

The apparatus for titration is shown in Fig. 4.1.

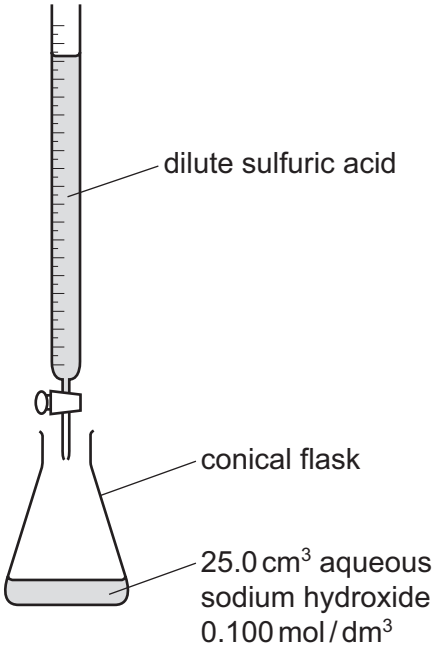


Fig. 4.1

Thymolphthalein is used as an indicator for this titration.

(i) State the colour change of thymolphthalein at the end-point of this titration.

from to

(ii) Suggest why universal indicator is **not** used for this titration.

(c) A student is provided with an aqueous solution of sodium sulfate.

Describe how to prepare a **pure** sample of sodium sulfate crystals from this solution.

Solution

(a)(i) The conical flask initially contains 25.0 cm³ aqueous NaOH; 0.100 mol/dm³ is its concentration. Thymolphthalein is blue in the alkali. Dilute H_2SO_4 added dropwise from the burette neutralises it, and at the endpoint the colour disappears. Enter blue for from and colourless for to.

(a)(ii) Universal indicator undergoes several colour changes, making one clear, sharp endpoint difficult to recognise. It is unsuitable for this titration. A suitable answer is “too many colour changes”.

(c) Gently heat the supplied aqueous Na_2SO_4 in an evaporating basin to evaporate some water until near saturation or the crystallisation point. Stop heating and allow it to cool so Na_2SO_4 crystals form. Filter and keep the crystals. If needed, quickly wash them with a little cold distilled water to remove surface mother liquor, then blot surface water away between clean filter papers to obtain pure, dry crystals. Do not keep heating to dryness: retain the full sequence of concentrating suitably, cooling to crystallise and drying by a stated method.

Practice

1. (2025·May/June 0620/41 Q3(c)(v)) · Level 1

This question is about copper and compounds of copper.

(c) Copper(II) sulfate crystals are made by the reaction between copper(II) carbonate and dilute sulfuric acid, using the following steps.

The sulfuric acid has a concentration of 0.100 mol / dm^3.

step 1 Powdered copper(II) carbonate is added to dilute sulfuric acid. The mixture is stirred. A reaction occurs.

step 2 More copper(II) carbonate is added, with stirring, until the reaction stops.

step 3 Unreacted copper(II) carbonate is separated from aqueous copper(II) sulfate by filtration.

step 4 Aqueous copper(II) sulfate is heated until some of the water evaporates.

step 5 The remaining solution is allowed to cool and crystallise.

step 6 The crystals are removed and dried.

(v) The solution at the end of **step 4** contains the maximum amount of copper(II) sulfate that will dissolve at that temperature.

State the term used to describe this type of solution.

2. (2025·May/June 0620/41 Q3(c)(i)) · Level 2

This question is about copper and compounds of copper.

(c) Copper(II) sulfate crystals are made by the reaction between copper(II) carbonate and dilute sulfuric acid, using the following steps.

The sulfuric acid has a concentration of 0.100 mol / dm^3.

step 1 Powdered copper(II) carbonate is added to dilute sulfuric acid. The mixture is stirred. A reaction occurs.

step 2 More copper(II) carbonate is added, with stirring, until the reaction stops.

step 3 Unreacted copper(II) carbonate is separated from aqueous copper(II) sulfate by filtration.

step 4 Aqueous copper(II) sulfate is heated until some of the water evaporates.

step 5 The remaining solution is allowed to cool and crystallise.

step 6 The crystals are removed and dried.

(i) Give **two** observations in **step 1**.

3. (2025·May/June 0620/11 Q19) · Level 2

Which reactants are used to make the salt copper(II) sulfate?

A dilute acid + alkali

B dilute acid + metal carbonate

C dilute acid + metal

D dilute acid + non-metal oxide

4. (2025·May/June 0620/43 Q3(c)(i)) · Level 2

This question is about zinc and compounds of zinc.

(c) Zinc sulfate crystals are made by the reaction between zinc carbonate and dilute sulfuric acid, using the following steps.

step 1 An excess of powdered zinc carbonate is added to dilute sulfuric acid.

step 2 Excess zinc carbonate is separated from aqueous zinc sulfate by filtration.

step 3 Aqueous zinc sulfate is heated until a saturated solution is formed.

step 4 The saturated solution is allowed to cool and crystallise.

step 5 The crystals are removed and dried.

(i) Give **two** observations which show that the zinc carbonate is in excess in **step 1**.

5. (2025·May/June 0620/43 Q3(c)(iii)) · Level 2

This question is about zinc and compounds of zinc.

(c) Zinc sulfate crystals are made by the reaction between zinc carbonate and dilute sulfuric acid, using the following steps.

step 1 An excess of powdered zinc carbonate is added to dilute sulfuric acid.

step 2 Excess zinc carbonate is separated from aqueous zinc sulfate by filtration.

step 3 Aqueous zinc sulfate is heated until a saturated solution is formed.

step 4 The saturated solution is allowed to cool and crystallise.

step 5 The crystals are removed and dried.

(iii) Name a compound, other than zinc carbonate, that can be added to dilute sulfuric acid to produce aqueous zinc sulfate in **step 1**.

6. (2025·May/June 0620/43 Q3(c)(iv)) · Level 2

This question is about zinc and compounds of zinc.

(c) Zinc sulfate crystals are made by the reaction between zinc carbonate and dilute sulfuric acid, using the following steps.

step 1 An excess of powdered zinc carbonate is added to dilute sulfuric acid.

step 2 Excess zinc carbonate is separated from aqueous zinc sulfate by filtration.

step 3 Aqueous zinc sulfate is heated until a saturated solution is formed.

step 4 The saturated solution is allowed to cool and crystallise.

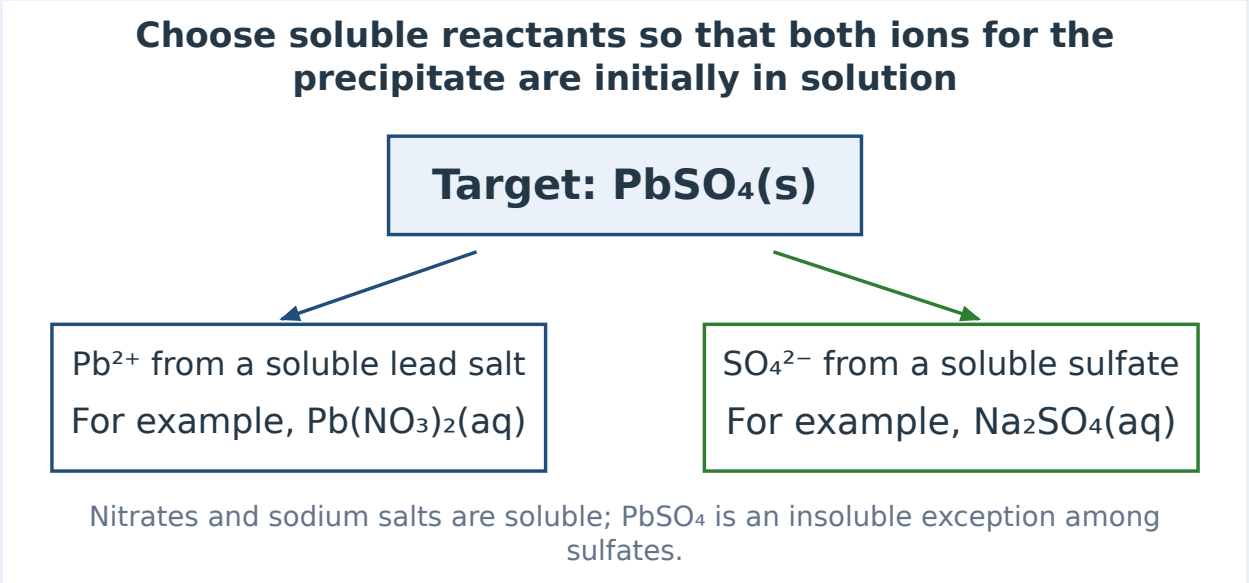
step 5 The crystals are removed and dried.

(iv) State what is meant by the term saturated solution.

Type 2: Preparing, separating and purifying insoluble salts by precipitation

Key Ideas The preceding route first produces a soluble salt solution and then crystallises the salt. If the target salt is insoluble in water, it can form directly as a solid when two solutions are mixed. The process in which dissolved ions combine to form an insoluble solid is precipitation; the solid formed is a precipitate.

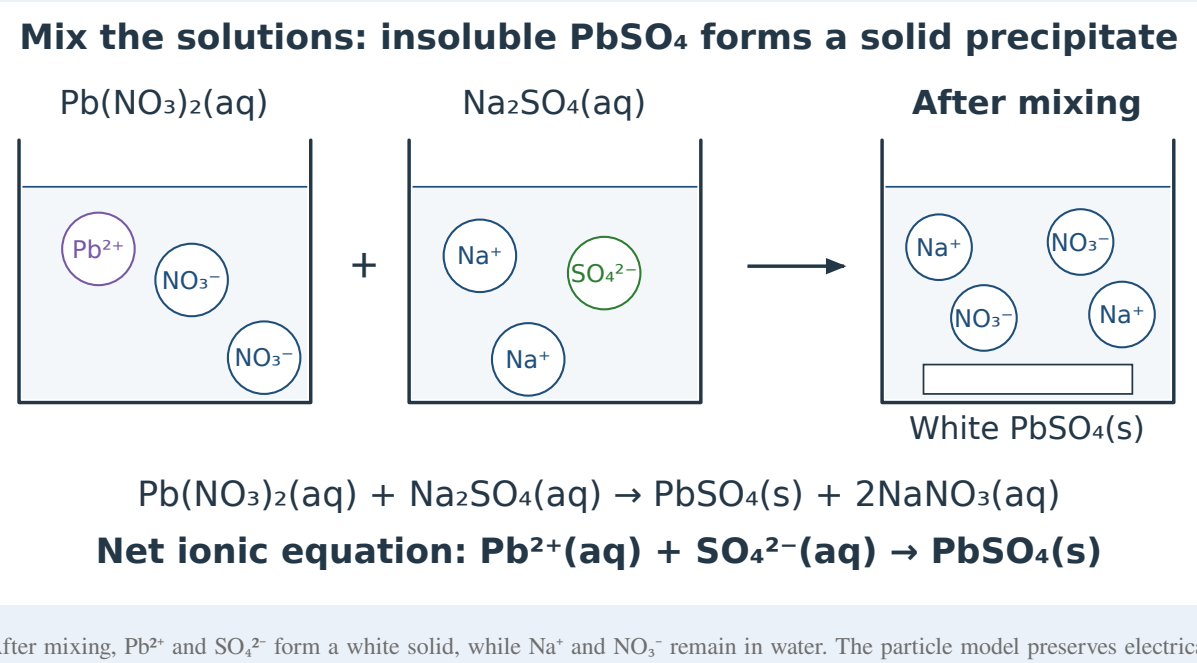
Consider lead(II) sulfate, $PbSO_4$. Most sulfates are soluble, but lead, barium and calcium are the listed exceptions. $PbSO_4$ therefore cannot be recovered by the usual soluble-salt route of crystallising it from a large volume of filtrate. Identify the target ions, Pb^{2+} and SO_4^{2-}, then select soluble reagents that supply them in water.



The target salt is insoluble, but both reactants supplying its ions must form aqueous solutions. Soluble lead(II) nitrate and sodium sulfate provide Pb^{2+} and SO_4^{2-} respectively.

For example, lead(II) nitrate, $Pb(NO_3)_2$, is soluble under the nitrate rule and supplies Pb^{2+}. Sodium sulfate, Na_2SO_4, is soluble under the sodium-salt rule and supplies SO_4^{2-}. Soluble reactants forming an insoluble product are entirely consistent; that difference allows the desired solid to separate.

Mix and stir the solutions. Pb^{2+} and SO_4^{2-} combine into white solid $PbSO_4$, while Na^+ and NO_3^- remain dissolved. Ions unchanged by the reaction are spectator ions. A full equation includes all substances, while the net ionic equation retains only the ions forming the precipitate.



After mixing, Pb^{2+} and SO_4^{2-} form a white solid, while Na^+ and NO_3^- remain in water. The particle model preserves electrical neutrality and matches the balanced full and net ionic equations below.

For example:

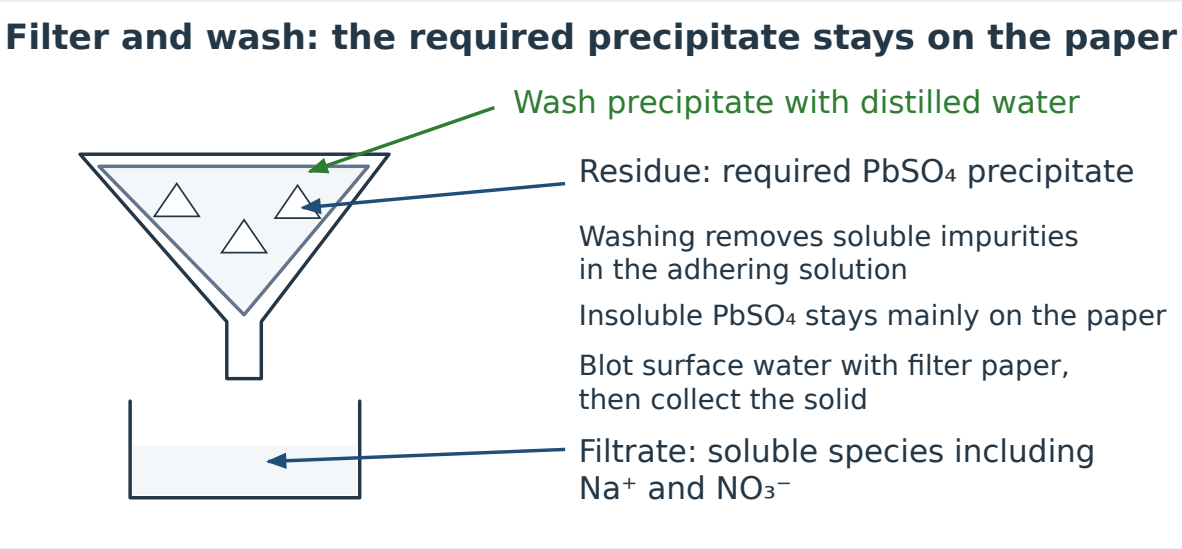
$$Pb(NO_3)_2 + Na_2SO_4 \rightarrow PbSO_4 + 2NaNO_3$$

The left contains 2 Na and 2 NO_3 groups, requiring $2NaNO_3$ on the right. $NaNO_3$ remains aqueous, while $PbSO_4$ is labelled (s) for solid. Cancel the $2Na^+$ and $2NO_3^-$ present on both sides:

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$$

The left-hand charges +2 and −2 total 0, and the salt on the right is also neutral.

The precipitate is already solid but remains mixed with water and other dissolved substances, so it needs separation and purification. Filter through paper: the precipitate remains as residue and the solution passes through as filtrate. Wash the precipitate with distilled water to remove soluble salts or excess reagents in adhering solution. Blot away surface water using filter paper to obtain a purer, dry salt.



The target is now the insoluble salt in the residue. Washing carries soluble substances from adhering solution through the paper. Collect the dried precipitate; evaporating the filtrate does not replace this step.

For example, collect the white residue when preparing $PbSO_4$. Evaporating the filtrate to dryness would mainly recover other dissolved salts, not the $PbSO_4$ just prepared. Washing cannot be omitted on the assumption that

an insoluble precipitate is automatically pure: its wet surface may retain solution containing $NaNO_3$ or excess reagent.

Compare soluble copper(II) sulfate with insoluble lead(II) sulfate. The target's location and purpose of filtration differ.

Filtration is used in both routes; the target's location depends on its solubility

Preparing soluble $CuSO_4$	Preparing insoluble $PbSO_4$
First filter off excess $CuCO_3$	Filter to collect the $PbSO_4$ precipitate
Target salt is in the filtrate	**Target salt is in the residue**
Concentrate and cool the filtrate to crystallise	Wash and dry the precipitate; collect the solid

Ask where the required salt is before choosing to keep the filtrate or residue.

Both soluble- and insoluble-salt preparations may use filtration, but the retained portion differs at this stage. $PbSO_4$ is not obtained by evaporating the filtrate.

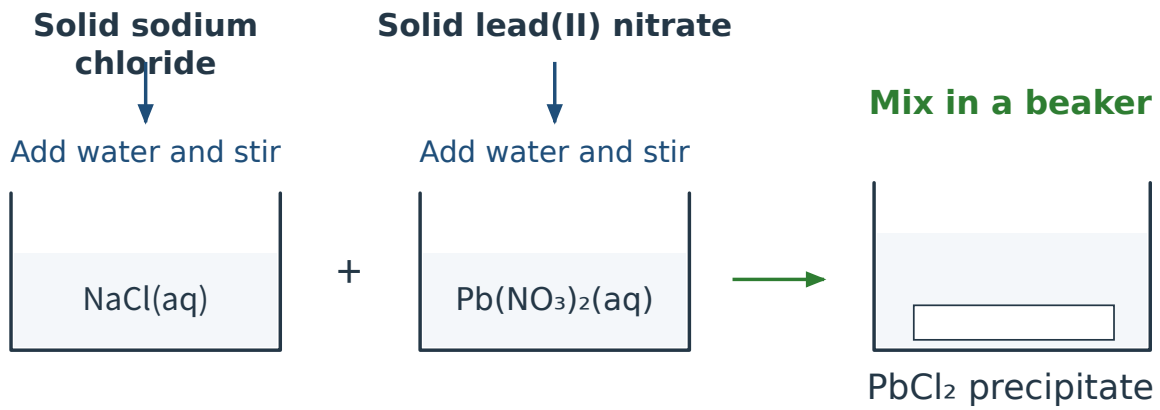
For example, in $CuSO_4$ preparation, excess $CuCO_3$ is an insoluble impurity. The first filtration retains it on paper while $CuSO_4$ remains in the filtrate, which then needs partial evaporation and cooling to crystallise.

In $PbSO_4$ preparation, the white insoluble solid is itself the target. Collect it directly by filtration, then wash and dry it.

Thus, when asked which method is not used to prepare pure $PbSO_4$, exclude crystallisation from solution and retain precipitation, filtration and drying.

If both supplied soluble salts are solids, first make solutions. For example, place solid sodium chloride, NaCl, and solid lead(II) nitrate, $Pb(NO_3)_2$, in separate beakers, add suitable amounts of distilled water, stir until dissolved, then mix the solutions.

Starting with two solids: dissolve each separately, then mix their solutions



Dissolve both salts first so that their ions can mix and react fully.

When NaCl and $Pb(NO_3)_2$ are supplied as solids, dissolve each separately in distilled water with stirring before mixing the solutions to form $PbCl_2$. Mixing dry powders does not allow the dissolution step to be skipped before filtration.

For example, NaCl solution supplies Cl^- and $Pb(NO_3)_2$ solution supplies Pb^{2+}. The target is lead(II) chloride, $PbCl_2$. Like $PbSO_4$, it is classified here as insoluble, but one Pb^{2+} requires two Cl^- to form $PbCl_2$:

$$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$$

Then filter, wash with a little cold distilled water and blot dry with filter paper. Do not replace separate dissolving with simply mixing the two dry powders.

For other insoluble salts, again identify the target ions and choose soluble sources. Mix the solutions, then filter, wash and dry.

For example, silver chloride, AgCl, needs Ag^+ and Cl^-. Mixing $AgNO_3(aq)$ with NaCl(aq) forms AgCl(s).

Barium sulfate, $BaSO_4$, needs Ba^{2+} and SO_4^{2-}. Mixing $BaCl_2(aq)$ with $Na_2SO_4(aq)$ forms $BaSO_4(s)$.

In each case, check the solubility rules: the ion-supplying reagents are soluble and the target product is insoluble.

Answers: Supply the target ions using two soluble reagents and mix to form the insoluble salt. Filter and keep the precipitate, wash away soluble impurities, then dry. The target is in the residue, not obtained by evaporating the filtrate to crystallise it.

Recognising the Question Use solubility to identify an insoluble target salt, then check that both ion-supplying reagents are soluble. For separation, identify whether the target is in the residue or filtrate.

Worked Example 1 (2025·May/June 0620/63 Q4) · Level 3

Lead(II) chloride, $PbCl_2$, is an insoluble salt. Solid lead(II) chloride can be made by reacting aqueous sodium chloride with aqueous lead(II) nitrate.

$$2NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + PbCl_2(s)$$

Describe how to make a pure, dry sample of solid lead(II) chloride.

You are provided with solid sodium chloride, solid lead(II) nitrate, distilled water and common laboratory apparatus.

Solution

Check the starting materials: both are solids. Do not assume two solutions have already been prepared.

Dissolve each separately. Put solid NaCl in one beaker, add a suitable amount of distilled water and stir with a glass rod until dissolved. Prepare aqueous $Pb(NO_3)_2$ in a second beaker in the same way.

Mix to precipitate. Combine the solutions in a beaker and stir, forming a white $PbCl_2$ precipitate:
$2NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + PbCl_2(s)$

Filter and retain the precipitate. Pour the mixture through a funnel containing filter paper and keep the $PbCl_2$ on the paper.

Wash away soluble impurities. Pour a little cold distilled water over the precipitate, allowing $NaNO_3$ and excess soluble reagents in adhering solution to pass through with the washings.

Dry and collect. Remove the precipitate, blot surface water away between clean filter papers and dry it to obtain pure, dry $PbCl_2$.

Include the containers, stirring and dissolving, mixing, filtration with retention of the residue, washing and a specific drying method. Merely listing filtration and drying is insufficient.

Practice

1. (2025·May/June 0620/21 Q19) · Level 2

Lead(II) sulfate is an insoluble salt.

Which process is **not** used to prepare a pure sample of this salt?

A crystallisation

B drying

C filtration

D precipitation

2. (2025·October/November 0620/43 Q5(a)) · Level 3

This question is about metallic compounds.

(a) Insoluble metal salts, such as lead(II) sulfate, are made by precipitation.

When aqueous lead(II) ions react with aqueous sulfate ions, a precipitate of lead(II) sulfate forms.

A sample of pure lead(II) sulfate can be obtained from the precipitate formed.

(i) Write an ionic equation for this precipitation reaction. Include state symbols.

(ii) State the name of an aqueous solution containing lead(II) ions.

(iii) State the name of an aqueous solution containing sulfate ions.

(iv) The lead(II) sulfate precipitate is separated from the rest of the reaction mixture by filtration. The impure lead(II) sulfate remains on the filter paper.

State the general term given to an insoluble substance that remains on the filter paper during filtration.

(v) Describe how the lead(II) sulfate is purified after filtration.

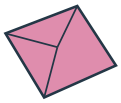
Type 3: Recognising water of crystallisation and anhydrous salts, and explaining heating and rehydration

Key Ideas Earlier, crystals were obtained by cooling salt solutions and blotting liquid from their surfaces. The solid may still contain water internally, because some water molecules form part of the crystal's composition rather than merely wetting its surface.

A hydrated substance is chemically combined with water; an anhydrous substance contains no water. In salt crystals, these molecules are called water of crystallisation: water molecules present within hydrated crystals. Aqueous means dissolved in water, forming a solution. It is not synonymous with hydrated.

Hydrated, anhydrous and aqueous describe different states

Hydrated: contains water of crystallisation



$CoCl_2 \cdot 6H_2O(s)$

Solid

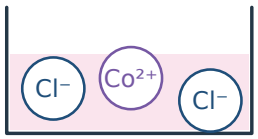
Anhydrous: no water



$CoCl_2(s)$

Solid

Aqueous: dissolved in water



$CoCl_2(aq)$

Dissolved in water

Water of crystallisation is part of the crystal composition; water in a solution is the bulk solvent.

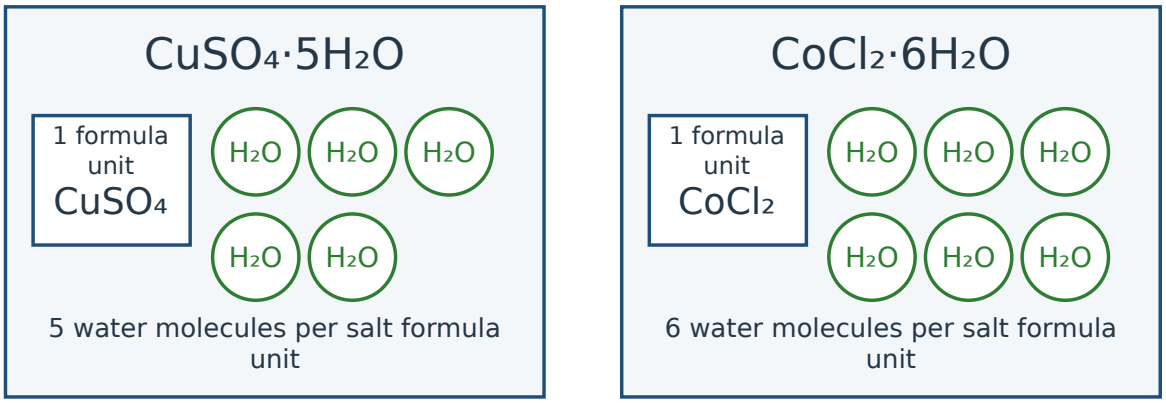
Crystal shapes and solution particle counts are schematic; solvent water molecules are omitted.

Hydrated and anhydrous describe a solid's composition; aqueous means the substance is dissolved in water. Common hydrated cobalt(II) chloride and its dilute aqueous solution are pink, while the anhydrous solid is blue.

For example, match “solid salt chemically combined with water” to hydrated; “solid salt containing no water” to anhydrous; and “salt dissolved in water” to aqueous. A hydrated solid has not thereby become a solution. Decomposition by water is also not the definition of hydrated.

Two hydrated crystals are required here: hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, and hydrated cobalt(II) chloride, $CoCl_2 \cdot 6H_2O$. The dot combines the anhydrous salt’s formula with a fixed ratio of water in one formula.

Water after the dot in a formula belongs to the composition of the hydrated crystal



The water coefficients are 5 and 6 respectively; these waters are part of the crystal composition.

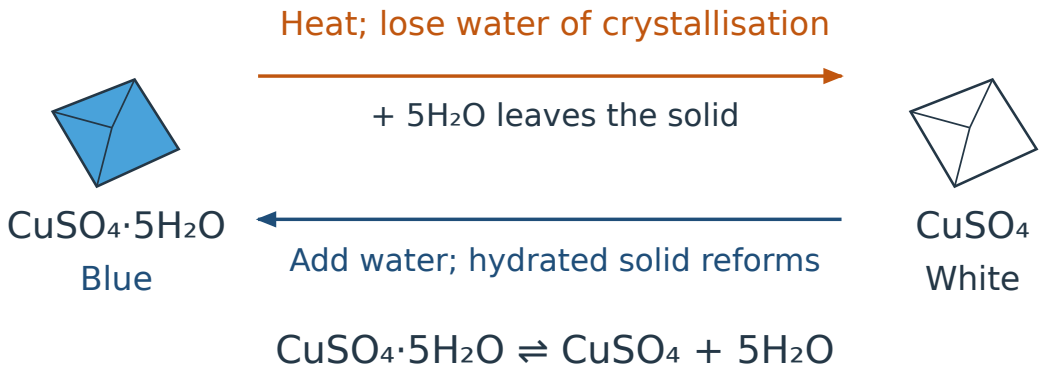
Only composition ratios are shown, not coordinate bonds, ion arrangements or the actual crystal structure.

$CuSO_4 \cdot 5H_2O$ and $CoCl_2 \cdot 6H_2O$ contain 5 and 6 parts water of crystallisation respectively. The number after the dot is not the total water volume in a solution.

For example, in $CuSO_4 \cdot xH_2O$, x is 5; in $CoCl_2 \cdot yH_2O$, y is 6. In mol terms, 1 mol $CuSO_4 \cdot 5H_2O$ contains 5 mol water of crystallisation, and 2 mol contains 10 mol. Not every hydrated salt contains 5 waters, and the word hydrated does not make both coefficients equal. The small circles help count composition; they do not mean that every water molecule is attached to the metal ion by the same type of bond.

Suitable heating can remove water of crystallisation. Blue hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, loses this water to form white anhydrous $CuSO_4$. Adding water can restore the blue hydrated substance. A reaction that can proceed in either direction when conditions change is reversible.

Copper(II) sulfate: heat blue to white; add water to turn white to blue



$$CuSO_4 \cdot 5H_2O \rightleftharpoons CuSO_4 + 5H_2O$$

Follow the arrow, then identify the substance and colour at its destination.

The same substances interconvert under different conditions. Blue hydrated copper(II) sulfate loses water of crystallisation to form the white anhydrous salt; adding water can restore the blue hydrate.

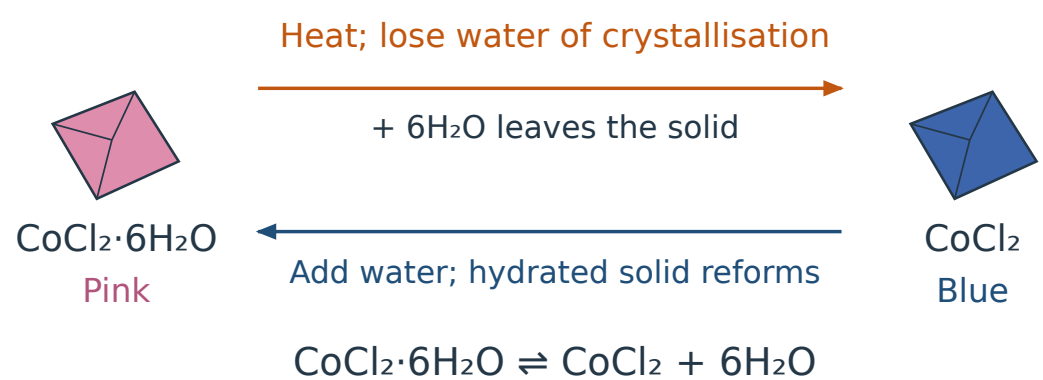
For example, start at blue $CuSO_4 \cdot 5H_2O$ on the left and follow the heating arrow right. Water leaves and the solid becomes white. If water is then added, follow the lower arrow from right to left: the blue colour returns. The equation:

$$CuSO_4 \cdot 5H_2O \rightleftharpoons CuSO_4 + 5H_2O$$

shows both directions together. Different conditions favour different directions; reversible does not mean equal amounts always react in both directions.

Cobalt(II) chloride also dehydrates and rehydrates, but its colours differ: hydrated $CoCl_2 \cdot 6H_2O$ is pink, while anhydrous $CoCl_2$ is blue. Heating changes pink to blue; adding water changes blue to pink.

Cobalt(II) chloride: heat pink to blue; add water to turn blue to pink



$$CoCl_2 \cdot 6H_2O \rightleftharpoons CoCl_2 + 6H_2O$$

Follow the arrow, then identify the substance and colour at its destination.

Cobalt(II) chloride has a different colour sequence from copper(II) sulfate. Identify the salt and hydration state before following the operation. “Adding water makes it blue” is not a universal rule.

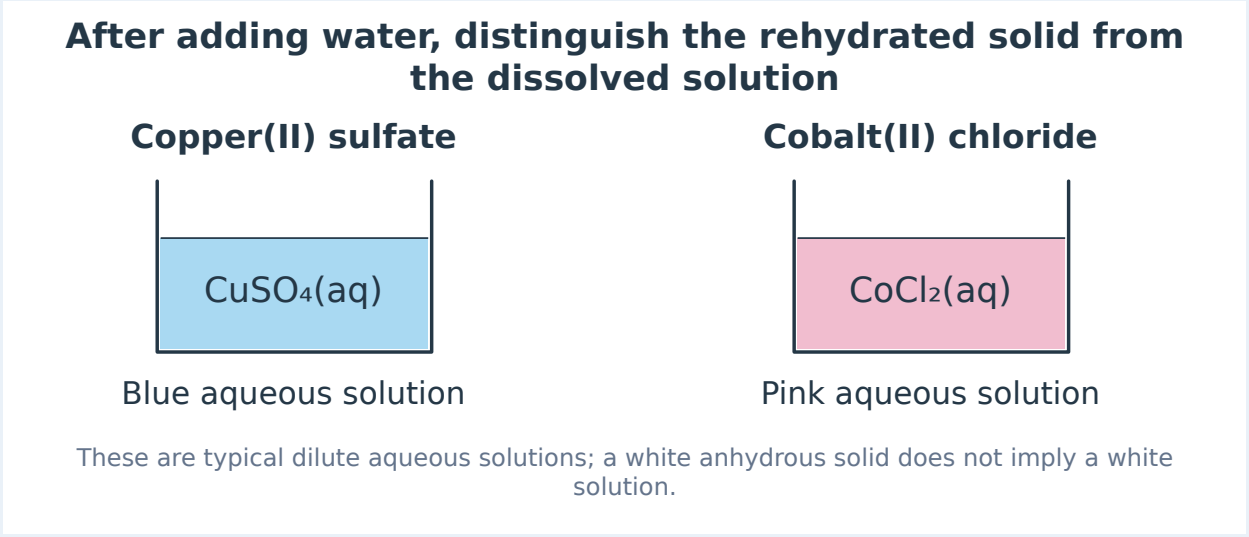
For example, a substance changes from blue to pink when water is added. The operation indicates rehydration of an anhydrous salt. Copper(II) sulfate changes white to blue, so it does not fit. Cobalt(II) chloride changes blue to pink, so the original blue solid is anhydrous cobalt(II) chloride. If the question also states that rehydration is exothermic, that describes energy released to the surroundings in this direction. Reversibility is still established by dehydration and rehydration restoring the starting substances.

A multiple-choice question may mix heating and water addition across options. Identify the salt and operation, then determine the product:

- Heating $CuSO_4 \cdot 5H_2O$ forms anhydrous $CuSO_4$: blue solid becomes white.
- Adding water to anhydrous $CuSO_4$ forms the hydrate: white becomes blue.
- Heating $CoCl_2 \cdot 6H_2O$ forms anhydrous $CoCl_2$: pink solid becomes blue.
- Adding water to anhydrous $CoCl_2$ forms the hydrate: blue becomes pink.

If the product must be white, only dehydration of copper(II) sulfate fits.

Another question may ask for the colour of the solution formed on adding water. This concerns the dissolved salt, not simply the original anhydrous solid. Common copper(II) sulfate solution is blue; common dilute cobalt(II) chloride solution is pink.



Adding water to copper(II) sulfate can produce a blue solution. Common dilute cobalt(II) chloride solution is pink. Distinguish aqueous-solution colours from anhydrous-solid colours.

For example, adding enough water to white anhydrous $CuSO_4$ gives blue $CuSO_4$ solution, so the original white colour is not the answer. Conversely, anhydrous $CoCl_2$ is blue but does not represent a common blue aqueous solution. If indicators appear among the options, recall section 7.1: pure water is near neutral, thymolphthalein is normally colourless and universal indicator near neutral is green. Do not introduce alkaline conditions merely because an indicator can turn blue in alkali.

Distinguish two drying operations. Blotting surface water with filter paper is the drying step in salt preparation. Strong heating that removes water of crystallisation changes the hydrate's composition. When hydrated crystals are required, drying must not be interpreted as removing all their water of crystallisation.

Answers: Hydrated means chemically combined with water; anhydrous means containing no water. $CuSO_4 \cdot 5H_2O$ contains 5 parts water of crystallisation and $CoCl_2 \cdot 6H_2O$ contains 6. Copper(II) sulfate: heating, blue → white; adding water, white → blue. Cobalt(II) chloride: heating, pink → blue; adding water, blue → pink.

Recognising the Question Decide whether the question concerns solid composition, an aqueous solution or colour after an operation. For colour, identify the salt first, then follow the heating or water-addition direction to its endpoint.

Worked Example 1 (2025·May/June 0620/13 Q20) · Level 1

Which statement describes a hydrated salt?

A It is an aqueous solution of a salt.

B It is a solid salt that is chemically combined with water.

C It is a solid salt that contains no water.

D It is a salt that has been broken down by water.

Solution

A hydrated salt is a solid salt chemically combined with water. B includes both solid salt and chemically combined with water. A describes an aqueous solution; C describes an anhydrous substance containing no water; D describes decomposition by water, not water of crystallisation. Choose B.

Worked Example 2 (2025·October/November 0620/11 Q15) · Level 1

Blue hydrated copper(II) sulfate produces a white solid when it is heated.

The white solid turns blue when water is added.

Which word describes the reaction of hydrated copper(II) sulfate?

A neutralisation

B oxidation

C reduction

D reversible

Solution

Track the first change: heating blue hydrated copper(II) sulfate removes water of crystallisation and forms white anhydrous copper(II) sulfate. Then track the second: adding water restores the blue hydrate. Changing conditions allows the products to become the starting substance again, so the reaction is reversible. Choose D. The key is loss and restoration of water of crystallisation; the presence of oxygen in water does not make this cycle oxidation or reduction.

Practice

1. (2025·May/June 0620/13 Q15) · Level 1

Which substance forms a blue solution when it is added to water?

A cobalt(II) chloride

B copper(II) sulfate

C thymolphthalein

D universal indicator

2. (2025·May/June 0620/21 Q20) · Level 1

The formulae of two hydrated salts are shown.

- $CuSO_4 \cdot xH_2O$
- $CoCl_2 \cdot yH_2O$

What are the values of *x* and *y*?

	x	*y*
A	5	5
B	5	6
C	6	5
D	6	6

3. (2025·May/June 0620/41 Q3(c)(vii)) · Level 1

(vii) Copper(II) sulfate crystals have the formula $CuSO_4 \cdot 5H_2O$.

State the term used to describe a substance that is chemically combined with water.

4. (2025·May/June 0620/43 Q3(c)(vi)) · Level 1

step 4 The saturated solution is allowed to cool and crystallise.

(vi) Hydrated crystals form in **step 4**.

State what is meant by the term hydrated.

5. (2025·October/November 0620/12 Q17) · Level 2

When pink cobalt(II) chloride is heated, it forms a blue solid and water.

When water is added to the blue solid, it releases heat and a pink solid is formed.

Which row describes the type of reaction that occurs and the pink cobalt(II) chloride?

	type of reaction	pink cobalt(II) chloride
A	non-reversible	anhydrous
B	non-reversible	hydrated
C	reversible	anhydrous
D	reversible	hydrated

6. (2024·May/June 0620/13 Q17) · Level 2

Which reaction produces a white-coloured substance?

A adding water to anhydrous cobalt(II) chloride

B adding water to anhydrous copper(II) sulfate

C heating hydrated cobalt(II) chloride

D heating hydrated copper(II) sulfate

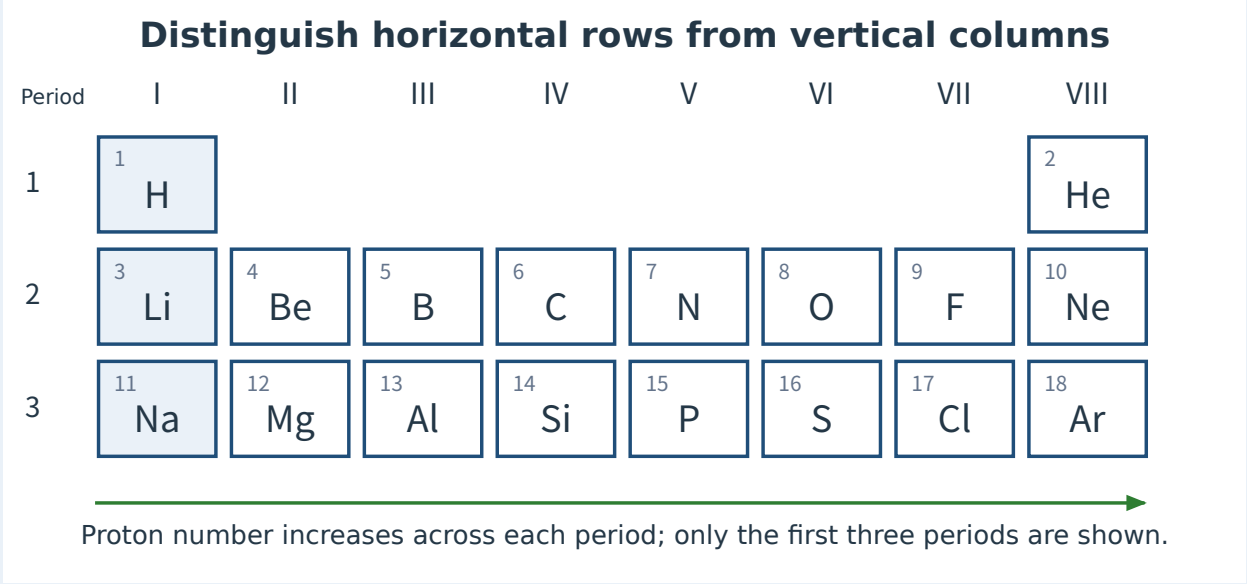
The Periodic Table

8.1 Arrangement and trends in the Periodic Table

Type 1: Using periodic position to predict electron configuration, ions and properties

Key Ideas The Periodic Table arranges elements in increasing proton number, also called atomic number. This determines element identity. The table is not arranged by an atom's nucleon number or strictly by relative atomic mass.

Read position first: a horizontal row is a period and a vertical column is a group. This book follows the questions' main-group labels I, II, III, IV, V, VI, VII and VIII. The final column, VIII, contains the noble gases and is labelled 18 in some tables. The diagram shows the first 3 periods needed in this section.

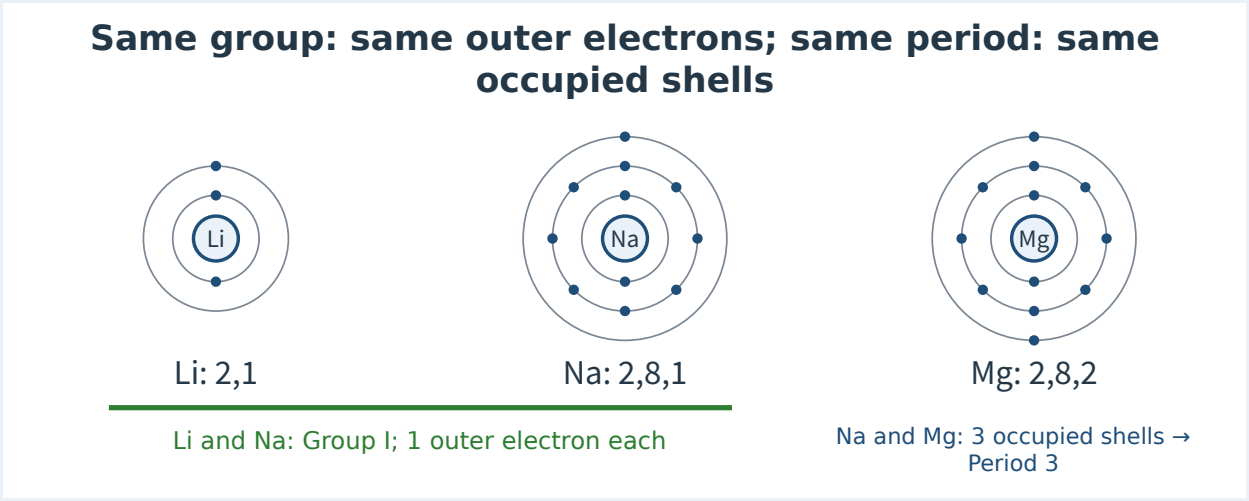


The small number in each box is proton number. Elements in the same horizontal row share a period; those in the same vertical column share a group.

For example, sodium, Na, has 11 in its box, meaning 11 protons in each sodium nucleus. It lies in period 3, Group I. Moving right to magnesium, Mg, in the same period increases proton number to 12. If a statement claims arrangement by nucleon number, distinguish the quantities: nucleon number is protons plus neutrons. Isotopes of one element may have different nucleon numbers without occupying different element boxes.

For the main-group elements here, period number gives the number of occupied electron shells. Group numbers I to VII give the number of outer-shell electrons. Treat helium, He, separately: it is in Group VIII, but only its first shell is occupied, and 2 electrons fill that shell. Group VIII does not mean He has 8 electrons.

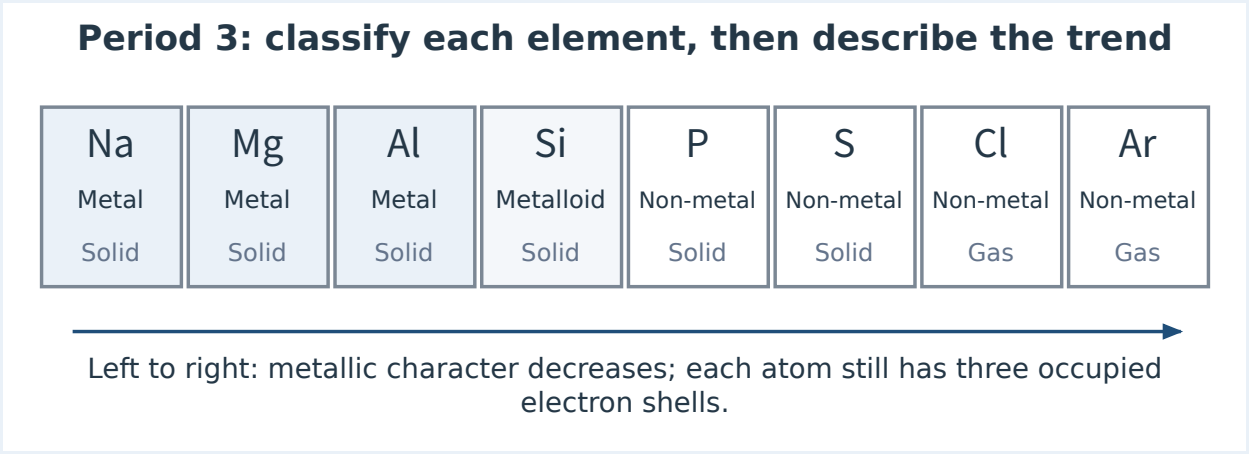
An electronic configuration such as 2,8,1 gives electrons in successive shells from inside out. Commas separate shells; the final 1 is the outer-shell count. Ordinary chemical reactions mainly involve outer electrons, so elements in one group often have similar chemical properties. Their complete configurations can still differ.



Each dot represents an electron. Li and Na have different complete configurations but the same number of outer-shell electrons.

For example, lithium, Li, is 2,1 and sodium, Na, is 2,8,1. Both end in 1 and have one outer electron, placing them in Group I. Li has two shells and Na three, so their complete configurations are not identical. Mg is 2,8,2: like Na it has three shells and lies in period 3, but its outer-shell count of 2 differs from Na's 1, so they are in different groups.

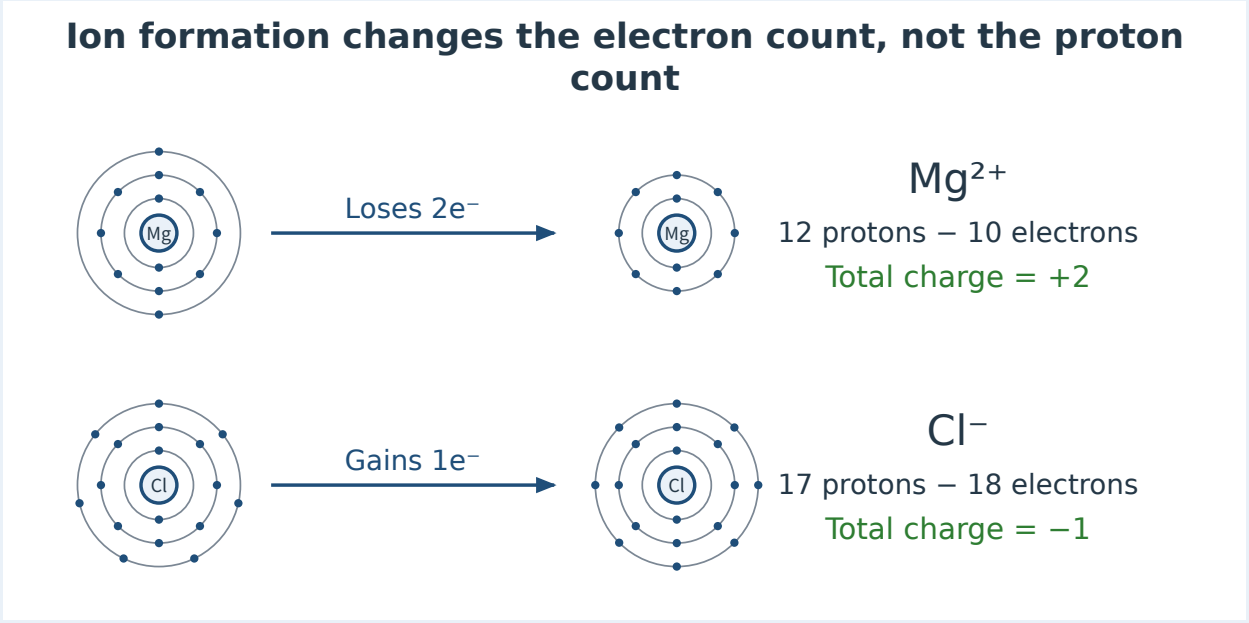
Across a period from left to right, metallic character generally decreases as elements progress from metals to non-metals. Check that the compared boxes are in the same row, then compare positions. Being in the same period does not imply similar chemical properties. It is same-group elements, with the same outer-shell electron count, that often share chemical properties.



The metals in period 3 are Na, Mg and Al. Si is a metalloid and is not counted among these three metals.

For example, period 3 contains Na, Mg, aluminium, Al, silicon, Si, phosphorus, P, sulfur, S, chlorine, Cl, and argon, Ar. Na, Mg and Al are metals; Si is a metalloid; P, S, Cl and Ar are non-metals. The metal count is therefore 3, excluding Si. All atoms in the row retain three occupied shells. Proton number and outer-shell electron count change; shell number does not decrease from box to box. At room temperature, the row goes from solids on the left to gases on the right, not gases to solids.

When an atom forms an ion, proton number stays constant and electron number changes. Losing negatively charged electrons gives a positive ion; gaining them gives a negative ion. Group I metals usually lose 1 electron to form 1+ ions; Group II metals usually lose 2 to form 2+ ions. Aluminium commonly loses 3 to form Al^{3+}. Group VI non-metals usually gain 2 to form 2− ions; Group VII usually gains 1 to form 1− ions.



Mg loses electrons to form a 2+ ion; Cl gains an electron to form a 1− ion. The nuclear proton numbers are unchanged.

For example, a Mg atom has 12 protons and 12 electrons, giving total charge 0. Losing 2 electrons leaves 10. In units of proton charge:

$12 - 10 = +2$

The ion is therefore Mg^{2+}.

A chlorine atom, Cl, has 17 protons and 17 electrons, with 7 in its outer shell. Gaining 1 electron gives 18 electrons in total:

$17 - 18 = -1$

The ion is therefore Cl^-.

For these non-metals gaining electrons to complete an outer shell of 8, charge can also be found as original outer-shell count minus 8. For Group VII:

$7 - 8 = -1$

Do not apply “group number minus 8” to every element. For metals losing electrons, use the actual electron transfer.

Position also helps predict trends within groups, but the direction need not be the same in every group. The key comparison is that Group I reactivity increases from Li to Na to K, whereas Group VII reactivity decreases from Cl to Br to I. Moving down does not make both groups more reactive. The next two sections explain these trends using reactions with water and halogen displacement.

Finally, connect this to chapter 7. A basic oxide reacts with an acid to form a salt and water. An acidic oxide reacts with an alkali. An amphoteric oxide reacts with both. Sodium oxide, Na_2O, and magnesium oxide, MgO, are basic; aluminium oxide, Al_2O_3, is amphoteric; silicon dioxide, SiO_2, phosphorus(V) oxide, P_4O_{10}, and sulfur dioxide, SO_2, are acidic.

Link an element's periodic position to its oxide properties

Element	Example of its common oxide	Reaction with acids and alkalis
Na, Mg	Na_2O, MgO	Basic: reacts with acids
Al	Al_2O_3	Amphoteric: reacts with acids and alkalis
Si, P, S	SiO_2, P_4O_{10}, SO_2	Acidic: reacts with alkalis

A counterexample disproves the claim that all Period 3 oxides are basic.

Group VIII elements are usually unreactive; they are not a group that forms basic oxides.

Comparing Na_2O, MgO, Al_2O_3 and non-metal oxides supplies specific counterexamples needed to connect the chapters.

For example, test the statement "period 3 elements only form basic oxides" by finding counterexamples. Amphoteric Al_2O_3 and acidic SO_2 both disprove the claim for the whole row. Ordinary Group I oxides such as Li_2O and Na_2O are basic. Group VIII elements have complete outer shells, are usually unreactive and do not normally form these common oxides. "Usually do not form oxides" does not mean "form only basic oxides". Combine periodic position with knowledge of specific oxides.

Elements are arranged by _____ number. Atoms in the same period have the _____ number of occupied shells. Elements in the same main group have the same number of _____ electrons. Helium's complete outer shell contains _____ electrons.

Answers: proton; same; outer-shell; 2

Recognising the Question Locate period and group, then choose the relevant relationship: occupied shells, outer electrons, charge after electron transfer or a specific property trend. For all or only form, check for counterexamples.

Worked Example 1 (2025·May/June 0620/13 Q3) · Level 2

Which statement about elements and their atoms is correct?

A Aluminium is in the second period of the Periodic Table and has two occupied electron shells.

B Helium is in Group VIII of the Periodic Table and has eight outer shell electrons.

C Lithium is in Group I of the Periodic Table and has one occupied electron shell.

D Sulfur is in the third period of the Periodic Table and has six outer shell electrons.

Solution

Split each statement into periodic position and electronic configuration, then check both.

A: Al is 2,8,3, with 3 occupied shells, so it is in period 3, not period 2. Incorrect.

B: He is in Group VIII but has only 2 electrons, which complete its first shell. It does not have 8 outer electrons. Incorrect.

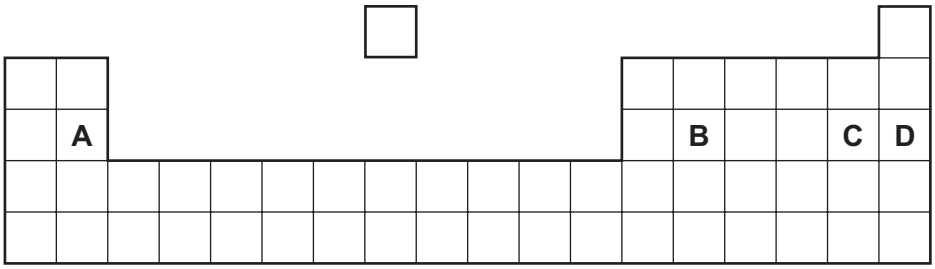
C: Li is in Group I because it has 1 outer electron. Its full configuration is 2,1, so it occupies 2 shells, not 1. Incorrect.

D: S is 2,8,6, with 3 occupied shells and 6 outer electrons. Both period 3 and 6 outer electrons are correct. Choose D.

Worked Example 2 (2025·May/June 0620/11 Q21) · Level 2

Part of the Periodic Table is shown.

Which element shows the most metallic character?



Solution

Read the horizontal row containing the letters. A, B, C and D are all in one period. A is on the left at the Group II position, with B, C and D progressively further right.

Metallic character generally decreases from left to right across a period. A is furthest left and therefore has the greatest metallic character among these candidates. Choose A.

There is no need to complete the whole blank table or guess from alphabetical order. The evidence is the relative position of the actual boxes.

Practice

1. (2025·May/June 0620/13 Q21) · Level 1

In the Periodic Table, how does the metallic character of the elements vary from left to right across a period?

A It decreases.

B It increases.

C It increases then decreases.

D It stays the same.

2. (2025·May/June 0620/42 Q1(c)) · Level 2

Using numbers only, state the:

(c) number of metals in Period 3 of the Periodic Table

3. (2025·May/June 0620/11 Q20) · Level 2

Which statement about the Periodic Table is correct?

A The elements are arranged in order of increasing relative atomic mass.

B The reactivity of the elements in Group I and in Group VII increases as the groups are descended.

C Elements in the same period have similar chemical properties.

D Elements in Group II form ions with a 2+ charge.

4. (2025·May/June 0620/12 Q21) · Level 2

Which statement about elements in the Periodic Table is correct?

A Elements in the same group of the Periodic Table have the same electronic configuration.

B The atoms of the elements in Group VII form ions with a charge equal to the number of outer shell electrons minus 8.

C The metallic character of the elements increases from left to right across a period.

D The Periodic Table is an arrangement of elements in order of increasing nucleon number.

5. (2025·May/June 0620/21 Q21) · Level 2

The properties of the elements change from left to right across the Periodic Table.

Which statement describes how the properties change across Period 3?

A Elements change from gases to solids.

B Elements change from non-conductors to electrical conductors.

C Elements change from metal to non-metal.

D Elements have a decreasing number of electron shells.

6. (2025·May/June 0620/12 Q20) · Level 4

Which parts of the Periodic Table contain elements that **only** form basic oxides?

A Group I and Period 3

B Group I only

C Group VIII and Period 3

D Group VIII only

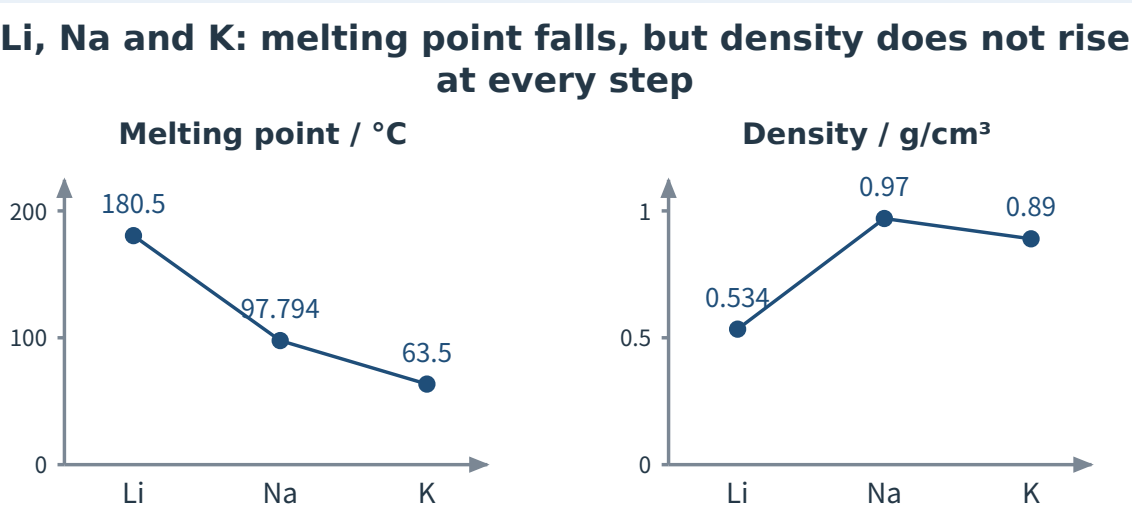
8.2 Group I elements

Type 1: Applying alkali-metal trends, explaining reactions with water and checking exceptions

Key Ideas Lithium, Li, sodium, Na, and potassium, K, are Group I elements called alkali metals. Further down are rubidium, Rb, caesium, Cs, and francium, Fr. Hydrogen, H, is often drawn above Group I but is not an alkali metal.

Distinguish three properties. Melting point is the temperature at which a solid begins to melt. Density is mass per unit volume, often in g/cm³. Reactivity concerns the tendency or behaviour in a particular chemical reaction; neither melting point nor density directly measures it. Li, Na and K are relatively soft metals that conduct heat and electricity and show typical metal properties such as malleability, the ability to be pressed into sheets.

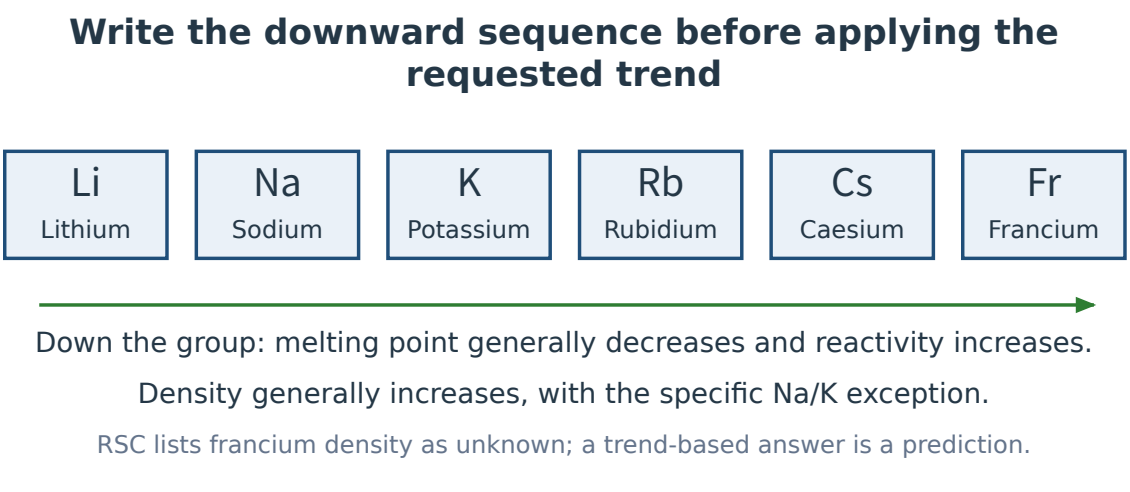
From Li to Na to K, melting point decreases and reaction with water becomes more vigorous. Density generally increases down the group, but a general trend does not mean every adjacent measured pair increases. Na is about 0.97 g/cm³ and K about 0.89 g/cm³: K is actually less dense than Na.



The horizontal axis follows Li, Na and K down the group. The two vertical axes represent melting point and density separately; their scales are not interchangeable.

For example, the left graph gives melting points decreasing from about 180.5°C for Li to 97.794°C for Na and 63.5°C for K. Li has the highest of these three. In the right graph, density first rises from 0.534 to 0.97, then falls to 0.89. For a specific Na–K comparison, answer that K is less dense; the overall trend must not erase this local decrease. The axes describe different quantities, so curve heights cannot identify which element is more reactive. Values come from the RSC element entries for Li, Na and K.

To predict properties of less familiar elements, write the order down the group, then select the property asked about. State that the answer is a trend-based prediction. In particular, the RSC lists francium's density as Unknown. Predicting Fr at the bottom as the high-density endpoint differs from claiming that its density has been measured.



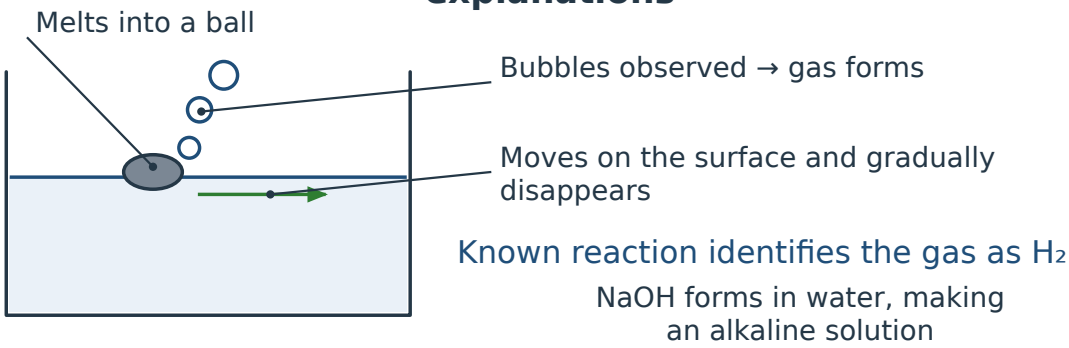
Use group order to apply trends, while distinguishing known data from predictions extrapolated to unmeasured elements.

For example, Li is above K and reactivity generally increases down this group, so Li is less reactive. For melting point, the trend instead decreases downwards, so Li has the higher melting point. Changing the property being asked about means the same "largest is at the bottom" answer cannot simply be reused.

Alkali metals react with water to form a metal hydroxide and hydrogen, H_2. Sodium, for example, forms sodium hydroxide, NaOH, and H_2. Dissolved NaOH supplies OH^-, making the solution alkaline and helping explain the name alkali metal.

For observations, first state what can be seen. Sodium in water produces bubbles, moves on the surface, melts into a ball and gradually disappears. Bubbles alone do not reveal a chemical name. Identifying the gas as H_2 uses reaction knowledge or a lighted-splint test giving a squeaky pop.

Sodium and water: separate observations from explanations



Bubbles are observed; hydrogen is identified using the reaction or a gas test.

The picture shows visible events during the reaction. The gas identity follows from reaction knowledge or a gas test.

For example, “bubbles form” and “the metal moves on the water surface” are two distinct observations. “Hydrogen forms” and “NaOH forms” identify products without expressing the visible observations requested. Lithium also bubbles, floats, moves and gradually disappears, but do not copy sodium’s melting into a ball: ordinary lithium–water reactions do not show this melting.

To write the equation, first establish the formulae Na, H_2O, NaOH and H_2, then balance each atom total:

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

Identify products, then balance atoms using coefficients

Sodium + water → sodium hydroxide + hydrogen

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

Atom	Left-hand total	Right-hand total
Na / O / H	2 / 2 / 4	2 / 2 / 4

Only change coefficients before formulae; never change the subscripts in H_2O or NaOH.

The balanced equation conserves Na, O and H atoms separately. Recheck the equation when replacing the alkali metal with Li.

For example, one H_2 molecule on the right requires two H atoms, while some H remains in NaOH. Use $2H_2O$ to supply 4 H atoms: 2 go into 2NaOH and 2 into H_2. The 2NaOH also requires 2Na on the left. Both sides now have 2 Na, 2 O and 4 H atoms. For Li, the product is lithium hydroxide, LiOH; do not incorrectly write LiO. Check the balance element by element again.

When comparing alkali metals with other metals, name the comparison. Typical transition elements such as iron, Fe, are generally harder and denser; Li, Na and K are softer and less dense. Both types conduct heat and electricity. Low density does not make alkali metals non-conductors.

Metal comparisons need a clear reference: higher or lower than what?

Property	Li, Na, K	Typical transition metal, e.g. Fe
Hardness	Softer; easily cut	Usually harder
Density	Lower	Usually higher
Conducts electricity and heat	Yes	Also yes

Alkali metals are relatively soft and low-density compared with typical transition metals. Both types conduct heat and electricity.

For example, if lower melting and boiling points are already given and two other physical properties are requested, do not repeat them. State that alkali metals are usually softer and less dense than transition metals. These are distinct properties with clear comparison directions.

Alkali metals react with water to form a metal _____ and _____. From Li to K, melting point generally _____ and reactivity _____. In the specific Na–K density comparison, _____ is less dense.

Answers: hydroxide; hydrogen; decreases; increases; K

Recognising the Question Distinguish melting point, density and reactivity, then decide whether to use actual data or an overall group trend. For observations, state visible changes. For equations, identify products before counting atoms.

Worked Example 1 (2025·May/June 0620/42 Q5(b)(i)) · Level 2

Alkali metals are reactive elements.

(b) Identify the alkali metal which:

(i) has the highest melting point

Solution

The order down the alkali metals is Li, Na, K, Rb, Cs, Fr. Melting point generally decreases down the group, so the highest is at the top.

The top alkali metal is lithium; Li is also acceptable. The question concerns melting point, not reactivity. A more vigorous reaction further down does not imply a higher melting point.

Worked Example 2 (2025·May/June 0620/41 Q6(e)(i)) · Level 2

This question is about the Periodic Table.

(e) Sodium is in Group I of the Periodic Table.

When sodium is added to water a chemical reaction occurs.

(i) Give **two** observations when sodium is added to water.

Solution

Give two directly visible observations. For example:

1. Bubbles form: bubbling, fizzing or effervescence.
2. Sodium moves on the water surface.

These describe gas release and metal movement and are distinct. Melting into a ball, floating or gradually disappearing can replace one, provided two different observations remain.

Formation of H_2 and NaOH explains the events, but product names alone do not express the visible observations requested.

Worked Example 3 (2025·May/June 0620/43 Q6(c)(ii)) · Level 2

This question is about the Periodic Table.

(c) Lithium is in Group I of the Periodic Table.

When lithium is added to water a chemical reaction occurs.

(ii) Write a symbol equation for this reaction.

Solution

The reactants are lithium, Li, and water, H_2O. Alkali metals form their metal hydroxide and hydrogen with water, so the products are LiOH and H_2.
Start with the skeleton:
$Li + H_2O \rightarrow LiOH + H_2$
Count atoms. To provide two H atoms for H_2 while retaining H in the hydroxide, use $2H_2O$, supplying 4 H atoms.
On the right, 2LiOH uses 2 H and H_2 uses the other 2. The 2LiOH also requires 2Li.
The complete equation is:
$2Li + 2H_2O \rightarrow 2LiOH + H_2$
Check both sides: 2 Li, 2 O and 4 H atoms each.

Practice

1. (2025·May/June 0620/42 Q5(a)) · Level 1

Alkali metals are reactive elements.

(a) State the group number of the alkali metals.

2. (2025·May/June 0620/43 Q6(a)) · Level 1

This question is about the Periodic Table.

(a) State the name given to Group I elements.

3. (2025·May/June 0620/43 Q6(b)) · Level 1

This question is about the Periodic Table.

(b) State which Group I element is least reactive.

4. (2025·May/June 0620/22 Q22) · Level 1

Sodium is a Group I metal.

What is a physical property of sodium?

A non-conductor of electricity

B high melting point

C high density

D malleable

5. (2025·May/June 0620/42 Q5(b)(ii)) · Level 2

Alkali metals are reactive elements.

(b) Identify the alkali metal which:

(ii) has the highest density

6. (2025·May/June 0620/42 Q5(b)(iii)) · Level 2

Alkali metals are reactive elements.

(b) Identify the alkali metal which:

(iii) has the lowest reactivity

7. (2025·May/June 0620/13 Q22) · Level 2

Lithium, sodium and potassium are elements in Group I of the Periodic Table.

Which statements about these elements are correct?

1 They react with water to produce hydrogen.

2 The melting point increases down the group.

3 The density decreases down the group.

4 The reactivity increases down the group.

A 1 and 2 **B** 1 and 4 **C** 2 and 3 **D** 3 and 4

8. (2025·May/June 0620/43 Q6(c)(i)) · Level 2

This question is about the Periodic Table.

(c) Lithium is in Group I of the Periodic Table.

When lithium is added to water a chemical reaction occurs.

(i) Give **two** observations when lithium is added to water.

9. (2025·May/June 0620/43 Q6(d)) · Level 2

This question is about the Periodic Table.

(d) Group I elements have lower melting points and lower boiling points than transition elements.

Describe **two other** physical properties of Group I elements that are different from transition elements.

Question Analysis

1. (2025·May/June 0620/11 Q22) · Level 3

E is an element from Group I of the Periodic Table. Two properties of E are listed.

- It has a higher melting point than caesium.
- It has a lower density than sodium.

What is the identity of E?

A lithium

B potassium

C rubidium

D francium

Answer: A

Explanation: The official answer is A, lithium. Applying the general trends of decreasing melting point and increasing density down the group, a density below Na suggests Li above it, and Li melts above Cs.

However, actual data do not make the two conditions unique to Li. Li has density about 0.534 g/cm³ and K about 0.89; both are below Na's 0.97. Li melts at about 180.5°C and K at about 63.5°C; both exceed Cs at 28.5°C. Thus both A, Li, and B, K, satisfy the conditions.

Retain both conclusions: A is the original official marking answer, while strict use of the two actual physical-property conditions allows both A and B. The question cannot uniquely establish A because of the Na–K density exception. An overall group trend is not a law for every pair.

The comparison data come from the RSC element entries for Li, K, Na and Cs.

2. (2025·May/June 0620/12 Q22) · Level 3

Which row identifies the trends in properties of the elements lithium to sodium to potassium?

	melting point	density	reactivity
A	decreasing	increasing	increasing
B	increasing	increasing	decreasing
C	increasing	decreasing	increasing
D	decreasing	decreasing	increasing

Answer: A

Explanation: The official answer is A: melting point decreasing, density increasing and reactivity increasing. Melting point decreasing and reactivity increasing match Li → Na → K. Density needs qualification: the actual values are approximately Li 0.534, Na 0.97 and K 0.89 g/cm³, so density first rises and then falls. If "from lithium to sodium to potassium" describes each actual step, density cannot be said to increase in both.

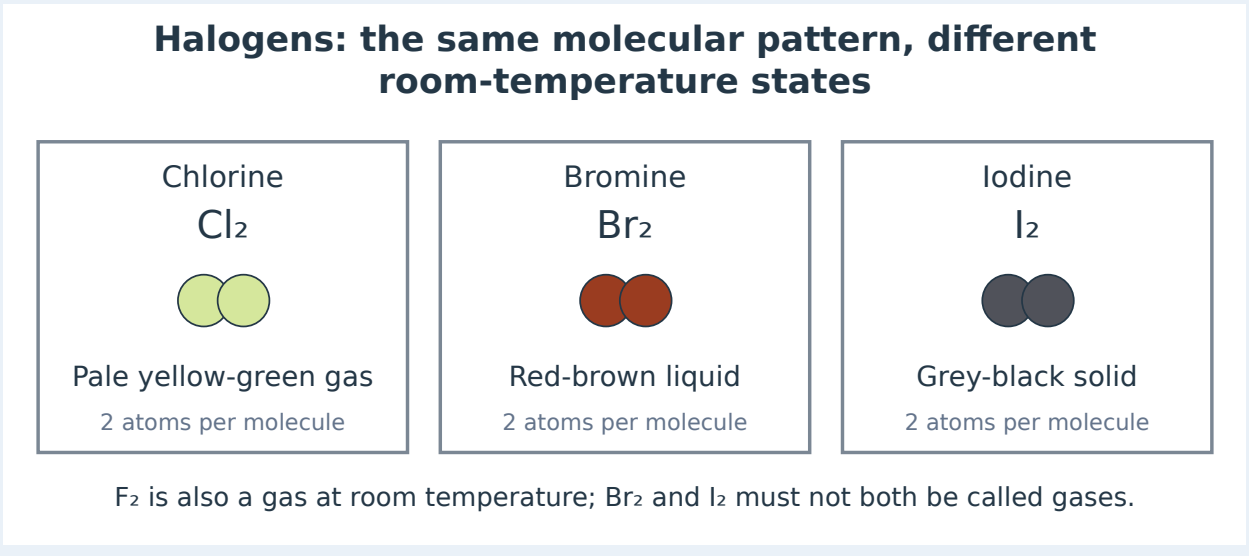
None of the four rows offers density first increasing then decreasing. A records the original answer based on the general trend; it does not prove K is denser than Na. The correct learning point retains both the overall density trend and the measured Na–K exception. Sources: RSC entries for Li, Na and K.

8.3 Group VII elements

Type 1: Identifying halogen states and inferring reactivity from displacement

Key Ideas Group VII elements are called halogens. Familiar members include fluorine, F, chlorine, Cl, bromine, Br, and iodine, I. They are non-metals whose elemental molecules usually contain two atoms of the same element: they are diatomic, as in F_2, Cl_2, Br_2 and I_2. The subscript 2 counts atoms per molecule; it does not mean the substance must be gaseous.

Room temperature and pressure, often abbreviated r.t.p., specifies the observation conditions. At r.t.p., Cl_2 is a pale yellow-green gas, Br_2 a red-brown liquid and I_2 a grey-black solid. F_2 is also a gas. Thus the gaseous halogens at room temperature are F_2 and Cl_2; do not count only chlorine or assume the entire group is gaseous.



Each pair of circles is one diatomic molecule. The adjacent text gives the elemental colour and state at room temperature and pressure.

For example, iodine's state and colour at r.t.p. are solid and grey-black. Purple I_2 vapour formed on heating and brown iodine in solution are different conditions from the room-temperature elemental solid requested. Fix the observation conditions before choosing a colour.

Distinguish elements from ions. Chlorine is elemental Cl_2; chloride means Cl^- and is used in salt names such as sodium chloride, NaCl. Bromine is Br_2 and bromide Br^-; iodine is I_2 and iodide I^-. The different endings describe particles with different charge and composition.

A more reactive halogen displaces a less reactive one from its aqueous halide solution. For Cl, Br and I, the order is Cl > Br > I, where > means more reactive than. For example, Cl_2 added to sodium bromide, NaBr, solution causes Br^- to lose electrons and form Br_2, while Cl_2 gains electrons to form Cl^-.

Chlorine and chloride are different particles



Cl_2: elemental chlorine

$2Br^-$: bromide ions

$2Cl^-$: chloride ions

Br_2: elemental bromine

$$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$$

Chlorine is more reactive than bromine and displaces it from aqueous bromide.

Colours distinguish elements only; spectator ions in solution are omitted.

Elemental Cl_2 becomes Cl^-, while Br^- becomes elemental Br_2. The correct description is chlorine displaces bromine.

For example:

$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$

The reactants contain elemental chlorine and bromide ions; elemental bromine appears as a product. Therefore say "chlorine displaces bromine", not "chloride displaces bromide", which wrongly names both neutral Cl_2 and Br_2 as ions. Each side has two Cl and two Br atoms and total charge 2−. Na^+ undergoes no net change and can be omitted from the ionic equation.

A less reactive added halogen cannot displace a more reactive one. For example, Br_2 gives no such net reaction with chloride solution because Br is less reactive than Cl. A comparison table supports both prediction from known reactivity and inference from observations.

Use reaction or no reaction to compare reactivity

Element added	Cl^- solution	Br^- solution	I^- solution
Cl_2	—	Reaction	Reaction
Br_2	No reaction	—	Reaction
I_2	No reaction	No reaction	—

Cl_2 displaces Br^-, so Cl is more reactive than Br; Br_2 cannot displace Cl^-, so Br is less reactive.

A dash means the same halogen, with no displacement.
Check that unknown elements are different.

Rows identify the added elemental halogen; columns identify the dissolved halide ion. The same halogen and its own ion give no net displacement.

For example, no displacement between Br_2 and Cl^-, two different elements, shows Br is weaker than Cl. No displacement between I_2 and Br^- shows I is weaker than Br. Combining gives Cl > Br > I. For unknowns P, Q and R, turn each observation into one reactivity relation and connect the shared letters; do not guess all identities from one observation. For a halogen with its own ion, such as Cl_2 with Cl^-, no net displacement does not mean chlorine is less reactive than itself. Confirm that different elements are being compared before reversing the reasoning.

Across the full group, F lies above Cl and is the most reactive halogen considered here. Down from Cl through Br to I, reactivity decreases, while density, melting point and boiling point generally increase. This reactivity trend is opposite to Group I.

Predicting within a group: first establish the direction of the trend

F_2	Cl_2	Br_2	I_2
Gas	Gas	Liquid	Solid

Down the group: melting point, boiling point and density generally rise; reactivity falls.

A halogen below iodine is predicted to be solid at room temperature from this trend.

A prediction states an inference from evidence, not a measured property of the new element.

Known states and group trends support predictions for unfamiliar positions. Distinguish a prediction from a measured result.

For example, to predict the r.t.p. state of an unfamiliar halogen below iodine, use Cl_2 gas, Br_2 liquid, I_2 solid and the downward increase in melting and boiling points. A solid is expected. Retain the meaning of expected state as a trend-based prediction rather than claiming an actual measurement.

Halogens usually form _____ molecules. At r.t.p., Br_2 is a _____ _____. Cl_2 converting Br^- to Br_2 shows that Cl is more _____ than Br.

Answers: diatomic; red-brown; liquid; reactive

Recognising the Question Distinguish elemental X_2 from halide X^-. Displacement shows that the added halogen is more reactive. For no displacement, first confirm two different elements, then combine the separate relations into a full order.

Worked Example 1 (2025·May/June 0620/41 Q6(c)) · Level 1

This question is about the Periodic Table.

(c) Give the physical state and colour of iodine at room temperature and pressure.

state

colour

Solution

Fix the conditions: room temperature and pressure. Elemental iodine, I_2, is a grey-black solid under these conditions.

Enter solid for state and grey-black for colour. Purple gas describes heated iodine vapour; brown solution describes dissolved iodine. Neither answers the specified elemental r.t.p. state and colour.

Worked Example 2 (2025·May/June 0620/11 Q23) · Level 2

An equation for the displacement reaction between aqueous halogen, X_2, and aqueous halide ions, Y^-, is shown.

$$X_2(aq) + 2Y^-(aq) \rightarrow 2X^-(aq) + Y_2(aq)$$

Which row identifies X_2 and Y^- and explains why the reaction takes place?

	X_2	Y^-	explanation
A	chlorine	iodide	chlorine is less reactive than iodine
B	chlorine	iodide	chlorine is more reactive than iodine
C	iodine	chloride	chlorine is less reactive than iodine
D	iodine	chloride	chlorine is more reactive than iodine

Solution

On the left, X_2 is an elemental halogen and Y^- is another halogen's ion. The products X^- and Y_2 show X displacing Y from its halide solution, so X must be more reactive than Y.

Chlorine is more reactive than iodine, so X_2 can be chlorine and Y^- iodide:

$$Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$$

A gives these particles but reverses the reason, saying chlorine is less reactive. C and D use weaker iodine as the added halogen, which cannot displace chlorine from Cl^-. Only B has both the correct particles and reason. Choose B.

Practice

1. (2025·May/June 0620/41 Q6(a)) · Level 1

This question is about the Periodic Table.

(a) State the name given to Group VII elements.

2. (2025·May/June 0620/41 Q6(b)) · Level 1

This question is about the Periodic Table.

(b) State which Group VII element is most reactive.

3. (2025·May/June 0620/21 Q22) · Level 1

Which property decreases down Group VII of the Periodic Table, from chlorine to iodine?

A atomic number

B density

C reactivity

D melting point

4. (2025·May/June 0620/43 Q6(e)(i)) · Level 1

This question is about the Periodic Table.

(e) The Group VII elements are known as the halogens.

(i) Give the physical state and colour of chlorine at room temperature and pressure.

state

colour

5. (2025·May/June 0620/42 Q1(d)) · Level 2

Using numbers only, state the:

(d) number of halogens which are gases at r.t.p.

6. (2025·May/June 0620/12 Q23) · Level 2

Which statements about the halogens are correct?

1 They are all diatomic molecules.

2 They are all gases at room temperature and pressure.

3 The density of the elements increases down the group.

4 The reactivity of the elements increases down the group.

A 1 and 2 **B** 1 and 3 **C** 2 and 4 **D** 3 and 4

7. (2025·May/June 0620/22 Q23) · Level 2

Tennessine (atomic number 117) is a manufactured element that is below astatine in Group VII of the Periodic Table.

What is the expected state of tennessine at room temperature and pressure?

A a diatomic gas

B a liquid

C a monatomic gas

D a solid

8. (2025·May/June 0620/22 Q24) · Level 2

When chlorine is bubbled into aqueous sodium bromide, a displacement reaction occurs.

Which description of this reaction is correct?

A Chloride displaces bromide.

B Chlorine displaces bromine.

C Chloride displaces bromine.

D Chlorine displaces bromide.

9. (2025·May/June 0620/13 Q23) · Level 3

X, Y and Z are each one of the elements chlorine, bromine and iodine.

When aqueous X is mixed with aqueous ions of Y, there is no reaction.

When aqueous Y is mixed with aqueous ions of Z, there is no reaction.

Which statement is correct?

A X is more reactive than Y.

B X is a pale yellow-green gas at room temperature and pressure.

C Y is a red-brown liquid at room temperature and pressure.

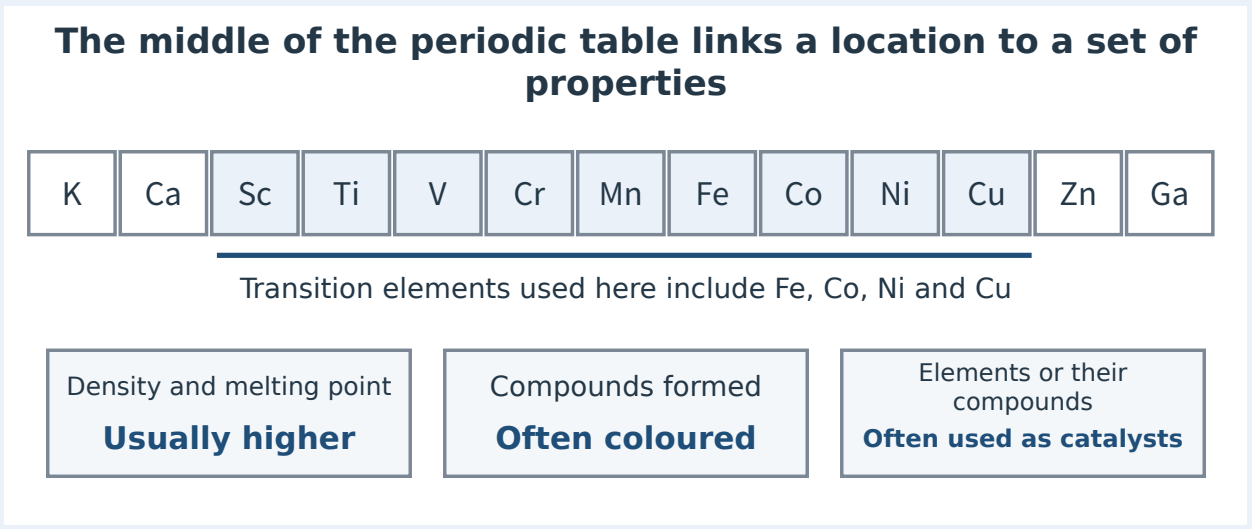
D Z is less reactive than Y.

8.4 Transition elements

Type 1: Recognising transition elements from their properties and choosing catalysts

Key Ideas Many transition elements occupy the centre of the Periodic Table. Familiar examples here are iron, Fe, cobalt, Co, nickel, Ni, and copper, Cu. They are metals and generally conduct electricity and heat. Compared with alkali metals such as Li, Na and K, typical transition metals have higher densities and melting points and are usually harder.

Two further clues are that they often form coloured compounds and that the elements or their compounds are often catalysts. As already explained, a catalyst increases rate and remains at the end of a reaction without changing equilibrium position at the same temperature.

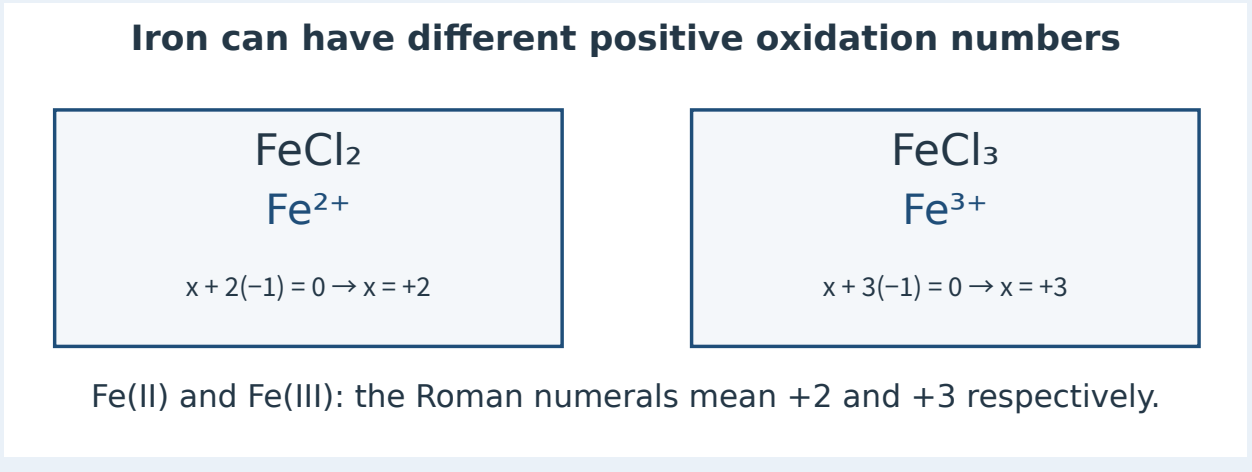


This continuous part of period 4 locates transition elements using familiar examples Fe, Co, Ni and Cu. Often forming coloured compounds does not mean every compound must be coloured.

For example, an unknown metal with a very high melting point, high density and frequent use as a catalyst fits the transition-element region. Use the whole set of clues rather than density alone. Do not replace “often form coloured compounds” with “every compound is coloured”. Some specific transition-element compounds are white or colourless; the question describes a typical combination of properties.

Comparison with Groups I, VII and VIII also eliminates wrong categories. Group I metals are usually softer, with lower melting points and densities. Group VII contains non-metal halogens, while Group VIII contains generally unreactive monatomic gases. A high-melting metal commonly used as a catalyst does not fit either of the latter groups.

Transition elements also often have variable oxidation states. Iron is +2 in iron(II) compounds and +3 in iron(III) compounds. Roman numerals II and III specify oxidation number, not atom count. Check the full formula and charge to understand the difference.



Each Cl^- has charge −1. Neutrality of the whole chloride gives iron +2 and +3 respectively, illustrating variable oxidation states.

For example, neutral iron(II) chloride, $FeCl_2$, contains two Cl^- ions totalling −2, requiring Fe to be +2. Iron(III) chloride, $FeCl_3$, contains three Cl^- ions totalling −3, requiring Fe to be +3. The same element has two oxidation states in different compounds: this is what variable means here.

When asked to suggest a catalyst from a list of elements, classify them by periodic position and select a typical catalytic candidate. Na and K are Group I; Mg and barium, Ba, are Group II; Al is Group III. Carbon, C, and sulfur, S, are non-metals. Fe, Co, Ni and Cu are familiar transition elements.

Candidate catalysts: first identify their periodic-table category

Candidate	Position or category	Select using these properties
Na, K	Group I	Not a transition element
Mg, Ba	Group II	Not a transition element
Fe, Co, Ni, Cu	Transition element in the central block	Common catalyst candidate

Each reaction still needs a suitable catalyst; transition elements are not automatically interchangeable.

Use periodic position to suggest a suitable catalyst category from the options, and state that the selected element is a transition element.

For example, from Na, Mg, Ni and S, classify each. Ni is a transition element and the other three are not, so select nickel and explain “it is a transition element”. If the question asks for the element, identify Ni rather than merely saying “a transition metal”. This is a reasonable choice from the supplied list and categories. Different reactions require suitable catalysts; not every transition metal catalyses every reaction.

Known industrial examples provide a check: Fe is used in ammonia synthesis, while vanadium(V) oxide, V_2O_5, catalyses SO_2 oxidation to SO_3 in the Contact process. The first is a transition element and the second its compound, showing that a catalyst need not be a piece of pure metal. For options accompanying unfamiliar reactions involving methanol, CH_3OH, or ethanoic acid, CH_3COOH, compare the listed element categories first. There is no need to guess product colours from reactant names.

Transition elements typically have _____ densities and melting points, often form _____ compounds and are often used as _____. Iron in Fe(III) has oxidation number _____.

Answers: high; coloured; catalysts; +3

Recognising the Question Use the set of properties together to identify the category. For a catalyst choice, name the specific element and explain that it is a transition element.

Worked Example 1 (2025·May/June 0620/11 Q24) · Level 1

Some information about an element is shown.

melting point / °C	1555
boiling point / °C	2963
colour of oxide	brown
use of element	as a catalyst

What is the position of this element in the Periodic Table?

A Group I

B Group VII

C Group VIII

D transition elements

Solution

The table gives a melting point of 1555°C and boiling point of 2963°C, both high. Its oxide is brown, so it forms a coloured compound, and the element is also used as a catalyst.

High melting and boiling points, coloured compounds and catalytic use together fit transition elements. Group I metals generally melt at lower temperatures. Group VII halogens and Group VIII noble gases also do not fit this high-melting catalytic metal. Choose D.

Practice

1. (2025·May/June 0620/12 Q24) · Level 1

Zirconium is a transition element.

What is a property of zirconium?

A high density

B does **not** conduct heat

C forms white compounds only

D low melting point

2. (2025·May/June 0620/13 Q24) · Level 2

Which row describes the properties of a transition element?

	melting point / °C	density at r.t.p.	colour and state of its oxide at r.t.p.
A	–210	low	brown gas
B	113	high	white solid
C	650	low	white solid
D	1085	high	red solid

3. (2025·May/June 0620/41 Q5(b)(iv)) · Level 2

Methanol, CH_3OH, is manufactured by the reaction between hydrogen and carbon monoxide.

An equilibrium mixture is produced.

$$2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$$

(iv) Suggest which of the elements from the list is a suitable catalyst for the reaction. Give a reason for your answer.

barium **carbon** **copper** **potassium** **sulfur**

catalyst

reason

4. (2025·May/June 0620/43 Q5(b)(iii)) · Level 2

Ethanoic acid is manufactured in the reaction between methanol and carbon monoxide.

An equilibrium mixture is produced.

$$CH_3OH(g) + CO(g) \rightleftharpoons CH_3COOH(g)$$

(iii) Suggest which element from the list is a suitable catalyst for the reaction.

Give a reason for your answer.

aluminium **carbon** **cobalt** **magnesium** **sodium**

catalyst

reason

8.5 Noble gases

Type 1: Explaining noble gases using complete outer shells and predicting from data

Key Ideas Group VIII contains the noble gases, including helium, He, neon, Ne, argon, Ar, krypton, Kr, and xenon, Xe. They are generally unreactive because they have complete outer electron shells. Ordinary reactions cannot readily make these arrangements more stable by simple electron gain or loss.

Distinguish a complete outer shell from an outer shell that always contains 8 electrons. He has only the first shell, filled by 2 electrons. Ne is 2,8 and Ar 2,8,8, each with 8 outer electrons.

Noble gases have full outer shells; helium's full first shell contains only two electrons



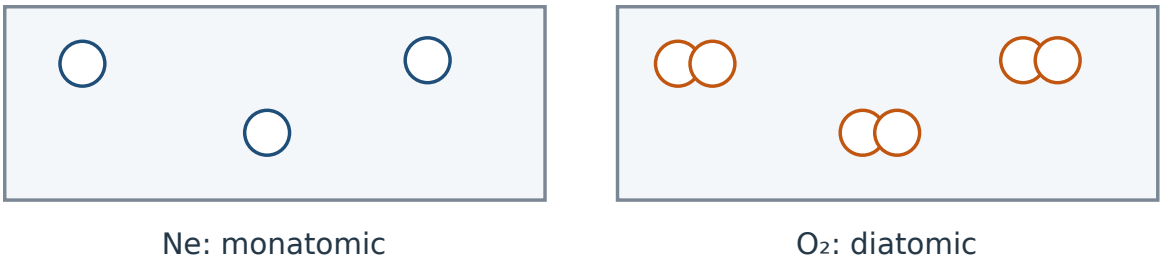
A full outer shell does not always mean eight electrons; check He when a statement says all.

Helium needs only 2 electrons to fill its first shell. Ne and Ar each have 8 outer electrons. All three have complete outer shells.

For example, test “all noble gases have eight outer electrons” using He. It is a noble gas with only 2 outer electrons, so the statement is false. “All have a full outer shell” includes He and has no such counterexample. Heavier common noble gases such as Kr and Xe also have 8 outer electrons. Their greater total electron counts are not their outer-shell counts.

Under ordinary conditions, noble gases exist as separate gaseous atoms and are monatomic. This means one atom per particle, not one atom in an entire container. In contrast, oxygen, O_2, has two atoms per particle and is diatomic. It is a gas but not a noble gas.

Monatomic gas: each particle is a separate atom



Both diagrams represent gases. Each left-hand particle contains one atom; each right-hand particle contains two.

For example, a statement describing gases as both reactive and monatomic requires both parts to be checked. Noble gases are monatomic but not normally reactive, so the complete statement is false. Helium’s lack of 8 outer electrons does not exclude it from noble or monatomic gases; its 2-electron first shell is already complete.

Some questions supply data for predictions rather than requiring memorised densities or boiling points. Read the element order down the group, then assess each column separately. If a missing value lies between adjacent known entries in a consistently increasing or decreasing interval, constrain it between those values. Both columns must satisfy their respective ranges.

Use trends to set a range; compare negative temperatures on a number line

Example: use He and Ar data to set bounds for Ne between them

Element	He	Ne	Ar
Density / g/dm³	0.2	?	1.8
Boiling point / °C	−269	?	−186



0.2 < density < 1.8; −269 < boiling point < −186.

For example, 0.9 and −246 are both within range; adjacent differences need not be equal.

Use the increasing downward trend to bracket the intermediate element. A negative temperature closer to zero is higher; do not compare only the digits after the minus sign.

For example, in the He, Ne, Ar sequence, He has density 0.2 and Ar 1.8 g/dm³, with density increasing down this interval. The predicted Ne value must lie between 0.2 and 1.8. Boiling point also increases from −269 to −186°C. The latter is closer to zero and larger; the larger absolute value 269 does not make −269 hotter. Ne's predicted boiling point lies between −269 and −186°C. The pair 0.9 g/dm³ and −246°C meets both requirements.

A row with a correct density but wrong boiling point is still eliminated. An increasing trend does not imply equal numerical increments between adjacent rows. This question filters values by ranges rather than forcing an arithmetic mean of the neighbouring numbers.

Noble gases usually exist as _____ particles. Their unreactive electronic feature is a _____ He has _____ outer electrons. −186°C is _____ than −269°C.

Answers: monatomic; complete outer shell; 2; higher

Recognising the Question For universal statements, check the essential He exception. Check every condition in a combined statement. For data predictions, establish each column's range separately, then find a row satisfying both.

Worked Example 1 (2025·May/June 0620/12 Q25) · Level 2

Some gases exist as single atoms, are unreactive and have eight electrons in their outer electron shell.

Which gases fit this description?

1 helium

2 argon

3 oxygen

4 krypton

A 1, 2 and 4 **B** 1 and 3 **C** 2 and 4 only **D** 3 only

Solution

All three conditions must hold: monatomic, unreactive and 8 outer electrons.

1, helium, is monatomic and unreactive but has only 2 outer electrons, so it fails the full set.

2, argon, is a noble gas with 8 outer electrons and qualifies.

3, oxygen, normally consists of diatomic O_2 and is not an unreactive noble gas, so it fails.

4, krypton, is a noble gas with 8 outer electrons and qualifies.

Only 2 and 4 fit. Choose C. Helium is excluded by the specific eight-electron condition, not excluded from the noble gases.

Worked Example 2 (2025·May/June 0620/22 Q21) · Level 3

The table shows some properties of noble gases.

noble gas	density in g / dm^3	boiling point in °C
helium	0.2	–269
neon	0.9	–246
argon	1.8	–186
krypton		
xenon	5.9	–108

Which row predicts the density and the boiling point of krypton?

	density in g / dm^3	boiling point in °C
A	1.4	–75
B	2.4	–195
C	3.8	–153
D	8.5	–213

Solution

Kr lies between Ar and Xe. Density generally increases through the listed He, Ne, Ar and Xe data, so Kr's predicted density should lie between 1.8 and 5.9 g/dm³.
Boiling point also increases downwards. Ar is −186°C and Xe −108°C, so Kr's predicted boiling point, in °C, must satisfy:

$-186 < \text{boiling point} < -108$

Compare negative temperatures on a number line. −153 lies in this interval, while −195 and −213 are below −186.

Check each row. A has density 1.4 below 1.8 and boiling point −75 above −108, so both fail. B has a possible density of 2.4 but −195 is outside the boiling-point range. C, 3.8 and −153, satisfies both. D has density 8.5 above 5.9, and −213 also fails.

Only C meets both conditions. Choose C. Do not assume equal differences between adjacent elements or decide using density alone.

Practice

1. (2025·May/June 0620/13 Q25) · Level 2

Which statements about the elements in Group VIII are correct?

1 They all have eight electrons in their outer electron shell.

2 They all react with sodium to form ionic compounds.

3 They are all monatomic gases.

A 1 and 2 **B** 1 only **C** 2 and 3 **D** 3 only

2. (2025·May/June 0620/21 Q24) · Level 2

Two statements about noble gases are listed.

1 Noble gases are reactive, monatomic gases.

2 Noble gases all have a full outer shell of electrons.

Which statements about noble gases are correct?

A Both statements are correct, and statement 2 explains statement 1.

B Both statements are correct but statement 2 does **not** explain statement 1.

C Statement 1 is correct but statement 2 is **not** correct.

D Statement 2 is correct but statement 1 is **not** correct.

Metals

9.1 Metals, uses and alloys

Type 1: Choosing metals and alloys by properties and explaining their atomic layers

Key Ideas **Begin with what the material can do, then name the property.** A metal transferring heat from a hotter end to a cooler one has good thermal conductivity. Allowing charge to flow readily is good electrical conductivity. These concern heat transfer and current, not reaction rate with an acid.

General property comparisons must retain important exceptions

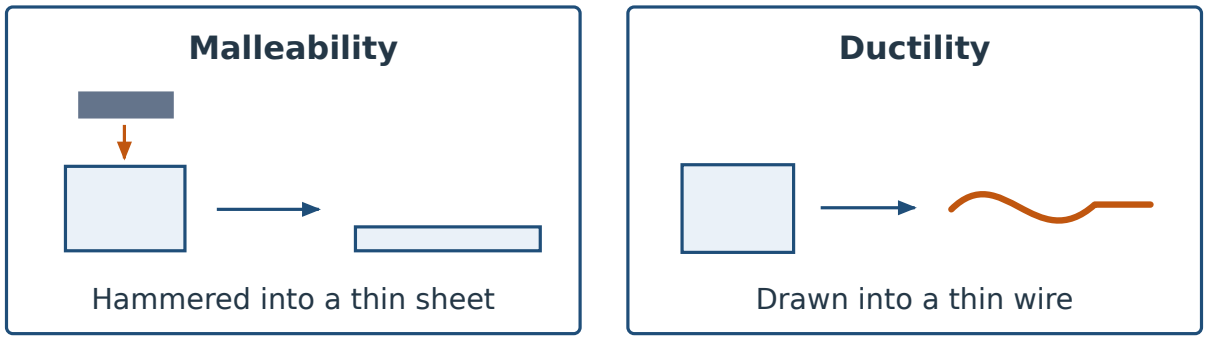
Property	Typical metals	Typical non-metals
Heat and electrical conduction	Good	Poor; graphite conducts electricity
Hammering and stretching	Malleable and ductile	Usually brittle when solid
Melting and boiling points	Usually higher	Usually lower
Important exceptions	Group I metals have low melting points	Diamond and graphite have very high melting points

The diagram compares typical properties, with exceptions considered separately. “All metals have high melting points” is not universally true.

Metals are usually easy to form into sheets or wires and generally have high melting and boiling points. Non-metals generally conduct heat and electricity poorly, are brittle when solid and have lower melting and boiling points. Exceptions matter: Group I metals such as lithium have relatively low melting points; diamond and graphite have very high melting points, and graphite conducts electricity. Do not copy every property of iron directly to lithium when asked which they share.

Distinguish malleability from ductility. Malleability is the ability to be hammered into sheets; ductility is the ability to be drawn into wires. Brittle non-metals such as solid sulfur tend to fracture under force and cannot be formed into continuous sheets or wires like metals.

Hammering into sheets and drawing into wires are different processes

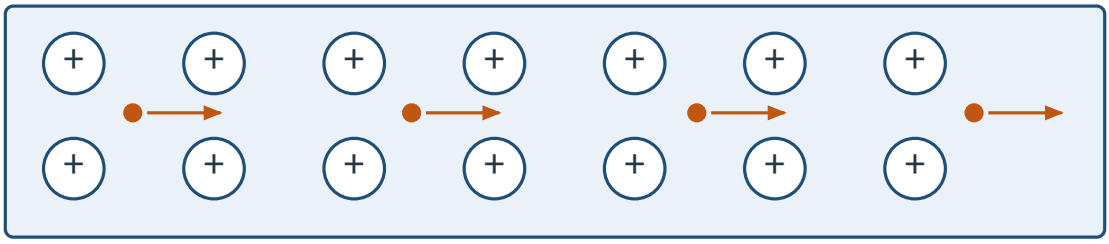


Each blue-arrow sequence shows the same material before and after shaping. The orange arrow on the left shows the hammering direction.

For example, a material drawn into electrical wire is ductile; a material hammered into a sheet is malleable. Comparing potassium and sulfur, potassium is a metal and usually conducts heat better, while non-metal sulfur is less malleable. Read each column carefully: higher and lower ask for opposite comparison directions.

Which particles carry charge in a metal? A metal contains an arrangement of positive metal ions and delocalised electrons that can move through the structure. An applied voltage produces directed electron movement and hence current. Metal atoms do not move freely along the wire, and positive metal ions do not travel through the entire solid wire to carry the current.

Metals conduct because free electrons carry moving charge



Positive metal ions vibrate about fixed positions; free electrons move through the structure.

Large circles are positive metal ions. Small orange dots show only some delocalised electrons and are not for charge counting. Orange arrows illustrate one direction of electron movement under an applied voltage.

For example, “copper conducts because copper atoms can move everywhere” identifies the wrong charge carriers. It should say that copper contains mobile delocalised electrons. Atomic layers sliding under an applied force is a separate phenomenon from electron movement during conduction.

How does a property explain a use? Density is mass per unit volume. At equal volume, lower density means lower mass, helping reduce the weight of aircraft and overhead cables. Aluminium has low density and good electrical conductivity. Copper conducts electricity well and is ductile.

Choosing a material: identify the properties its use requires

Material and use	Required property
Aluminium: aircraft	Low density reduces weight
Aluminium: overhead cables	Low density and good electrical conductivity
Aluminium: food containers	Corrosion resistance
Copper: electrical wires	Good electrical conductivity and ductility

At the same volume, a lower-density material has less mass. A property must both be correct and explain the intended use.

Low density is central to aluminium's use in aircraft. Overhead cables require both low density and good electrical conductivity, while food containers require corrosion resistance. Copper for electrical wire must conduct and be drawn into continuous strands. Aluminium's compact surface oxide layer protects the metal beneath from corrosion. Its apparent resistance does not mean aluminium has very low reactivity.

For example, suppose a question compares aluminium with copper for overhead cables and states that copper conducts better. Even if correct, this does not explain choosing aluminium. A supporting reason is aluminium's lower mass for an equal volume. Explaining a choice is different from selecting every factually correct statement.

An alloy is not a pure compound. An alloy is a mixture of a metal with other elements. At least one component is metallic; others may be metals or a non-metal such as carbon. An alloy is a mixture and need not have a fixed composition.

An alloy is a mixture; its atoms need not all be the same size

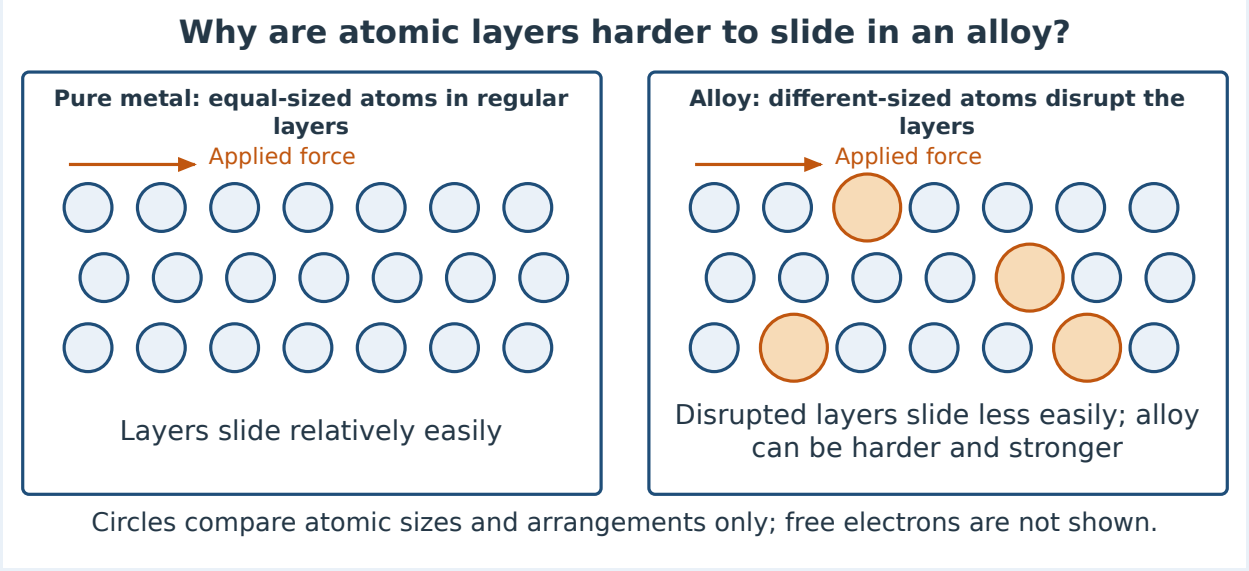
Alloy	Composition	Property or use needed here
Brass	Copper + zinc	Usually harder than pure copper
Stainless steel	Iron + chromium, nickel, carbon, etc.	Hard and rust-resistant; used for cutlery
Bronze	Copper + tin	Different composition from brass

A mixture of a metal with other elements includes adding carbon to make steel.

The + signs indicate mixed components, not reaction equations or fixed proportions. Bronze is included to distinguish a distractor in the examples.

Brass contains copper and zinc. Stainless steel consists mainly of iron with other elements such as chromium, nickel and carbon. Bronze contains copper and tin; do not use its composition for brass. For example, if copper in brass is already given, start with copper + zinc and remove copper: the other metal is zinc. If zinc is given, the other is copper.

Why can adding other elements increase hardness? In a pure metal, atoms have the same size and form relatively regular layers that slide over each other readily under force. Different-sized atoms in an alloy disrupt that regular arrangement, making layers harder to slide.



Both diagrams use the same circular atom model. Orange arrows show applied force. Different-sized atoms in the alloy disrupt the layers.

Alloys **can** therefore be harder and stronger than pure metals and retain their shape better. Harder means more resistant to scratching or indentation; stronger means able to withstand greater force without failure. Do not claim every alloy is stronger than every pure metal, or explain alloys using atoms that are all the same size.

Stainless steel is used for cutlery because it is hard and resists rusting. Knives and forks need to retain their shape while regularly contacting water and food. A steel described as strong but brittle and easily corroded should not be selected on strength alone. For example, of two hard materials, one rusting and the other rust-resistant, only the second meets both cutlery requirements.

The mobile charge carriers in metals are _____. The ability to be drawn into wire is _____.

Brass contains copper and _____. Two properties making stainless steel suitable for cutlery are _____ and _____.

Answers: delocalised electrons; ductility; zinc; hardness; resistance to rusting

Recognising the Question Turn the use into required properties, then check whether the material meets them. For alloys, establish composition and mixture status. Explain hardness through different-sized atoms disrupting layers and making sliding more difficult.

Worked Example 1 (2025·May/June 0620/21 Q25) · Level 1

Which statement about the properties and uses of copper is correct?

A It is used in electrical wiring, and it is ductile.

B It is mixed with tin to make the alloy brass.

C It is used in overhead electrical cables because it has a low density.

D It is used in takeaway food containers because of its resistance to corrosion.

Solution

The copper property and use must match within the same statement. Copper conducts electricity well and is ductile, meaning it can be drawn into wires. A is correct.

B gives the wrong brass composition: brass contains copper and zinc; copper and tin form bronze. C's low-density overhead-cable combination refers to aluminium in this section, not copper. D's corrosion-resistant food-container use also refers to aluminium.

Only A correctly describes copper. Choose A. The word electrical alone is insufficient; check the material, property and specific use together.

Worked Example 2 (2025·May/June 0620/22 Q25) · Level 2

Brass is an alloy of copper.

Which statement about brass or copper is correct?

A Brass does **not** conduct electricity because it is a compound.

B Copper conducts electricity because its atoms are free to move.

C Brass is harder than copper because its layers of atoms **cannot** easily slide over each other.

D Copper is stronger than brass because it is a pure metal.

Solution

Brass is an alloy of copper and zinc. Compared with the regular layers in pure copper, different-sized atoms disrupt brass's layers and make them harder to slide. Brass can therefore be harder than copper. C correctly links structure and property.

A wrongly calls an alloy a compound and concludes it cannot conduct. Brass is a mixture and still conducts. B incorrectly identifies freely moving copper atoms as charge carriers; the carriers are delocalised electrons. D claims copper must be stronger merely because it is pure, which does not fit this comparison. The alloy structure here makes sliding harder. Choose C.

Practice

1. (2025·May/June 0620/41 Q3(b)(i)) · Level 1

This question is about copper and compounds of copper.

(b) Copper is in alloys such as brass.

(i) State **one** reason why alloys are more useful than pure metals.

2. (2025·May/June 0620/41 Q3(b)(ii)) · Level 1

This question is about copper and compounds of copper.

(b) Copper is in alloys such as brass.

(ii) Name the substance that is present in brass, other than copper.

3. (2025·May/June 0620/42 Q4(b)) · Level 1

Chromium is the element with atomic number 24 in the Periodic Table.

(b) Chromium can be mixed with nickel and other elements to form stainless steel.

Name the type of substance formed when a metal is mixed with other elements.

4. (2025·May/June 0620/43 Q3(b)(i)) · Level 1

This question is about zinc and compounds of zinc.

(b) Zinc is present in alloys such as brass.

(i) State the meaning of the term alloy.

5. (2025·May/June 0620/43 Q3(b)(ii)) · Level 1

This question is about zinc and compounds of zinc.

(b) Zinc is present in alloys such as brass.

(ii) Name the substance that is present in brass, other than zinc.

6. (2025·May/June 0620/12 Q29) · Level 2

Pure metals can be combined to make alloys.

Why are the alloys sometimes used rather than the metals they are made from?

A Alloys are easier to extract and cheaper.

B Alloys are easier to hammer into different shapes.

C Alloys are harder and keep their shape better.

D Alloys contain only atoms of the same size but pure metals do **not**.

7. (2025·May/June 0620/13 Q26) · Level 2

Sulfur and potassium are two elements in the Periodic Table.

Which row identifies the element with the higher thermal conductivity and the element with the lower malleability?

	higher thermal conductivity	lower malleability
A	sulfur	potassium
B	sulfur	sulfur
C	potassium	sulfur
D	potassium	potassium

8. (2025·May/June 0620/13 Q27) · Level 2

Three statements about aluminium and copper are listed.

1 The density of aluminium is less than that of copper.

2 The electrical conductivity of copper is higher than that of aluminium.

3 Aluminium is more reactive than copper.

Which statements explain why aluminium is used instead of copper to make overhead power cables?

A 1 and 2 **B** 1 and 3 **C** 1 only **D** 2 and 3

9. (2025·May/June 0620/21 Q23) · Level 2

Some properties of elements are listed.

1 They have good electrical conductivity.

2 They have high melting points.

3 They are malleable.

4 They have ions with variable oxidation numbers.

Which properties are correct for **both** iron and lithium?

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

10. (2025·May/June 0620/11 Q28) · Level 2

Steel is a mixture of iron and one or more other elements.

The table gives some information about three common types of steel.

	type of steel	elements added to iron	properties
1	high-carbon steel	carbon only	strong, brittle and corrodes
2	low-carbon steel	carbon only	soft, easily shaped and corrodes slowly
3	stainless steel	carbon, chromium and nickel	hard and resistant to corrosion

Which rows identify a type of steel that is suitable to make cutlery?

A 1 and 2 **B** 1 only **C** 2 and 3 **D** 3 only

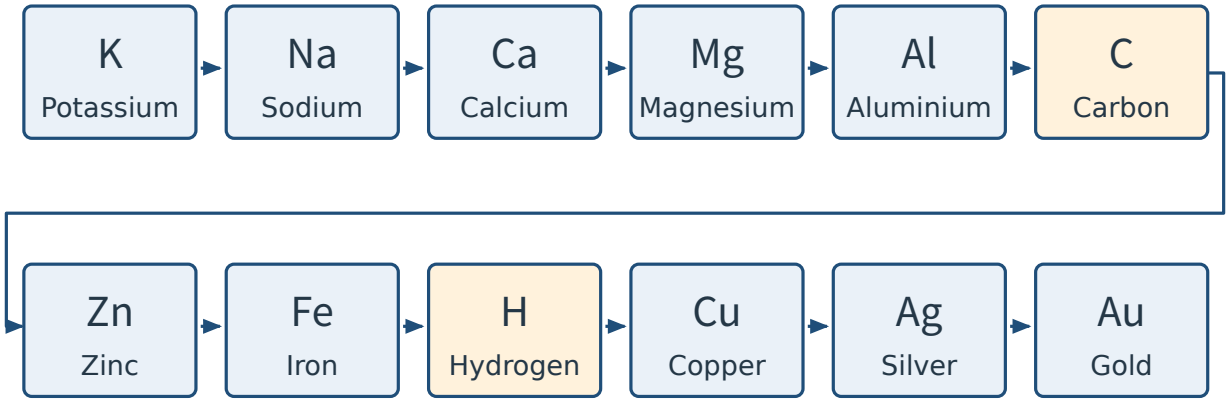
9.4 Metal reactivity

Type 1: Using water, dilute acids and displacement to infer metal reactivity

Key Ideas **What does more reactive mean?** Under the same comparable conditions, some metals react more readily with water, dilute acids or other metals' ions. Arranging them by this behaviour gives the reactivity series. C and H below are non-metals included as reference positions for carbon extraction and reactions with acids.

First locate the metal in the reactivity series

Read each row left to right, then continue on the next row



Potassium is most reactive; gold is least reactive. C and H are comparison points, not metals.

A metal that loses electrons and forms positive ions more readily is generally more reactive.

The arrow shows decreasing reactivity, not conversion of one element into another. C and H are reference positions for extraction and acid reactions.

The order from most to least reactive is potassium, sodium, calcium, magnesium, aluminium, carbon, zinc, iron, hydrogen, copper, silver, gold. Locate both elements before comparing; the earlier is more reactive. Calcium precedes magnesium and reacts more visibly with cold water. Iron precedes copper, so copper cannot be described as reacting faster with dilute hydrochloric acid.

Distinguish reaction conditions first. Cold water and steam differ in temperature and state, so their reaction products may differ. Dilute hydrochloric and sulfuric acids contain different acid anions and form different salts.

For water, steam and dilute acids, check each metal's conditions and products

Metal and conditions	Main products	Observation and conclusion
K, Na, Ca + cold water	Metal hydroxide + H_2	Bubbles; K and Na react more vigorously
Mg + steam	MgO + H_2	Heated magnesium reacts with steam
Mg, Zn, Fe + dilute HCl	Metal chloride + H_2	Bubbles; metal gradually dissolves
Cu, Ag, Au + dilute HCl	Usually no reaction	No H_2 released
Metal + oxygen	Metal oxide	For example, Mg burns to form MgO

Cold water, steam, dilute hydrochloric acid and oxygen are distinct conditions. Products cannot be interchanged merely because reagents contain water or are both acids.

Potassium, sodium and calcium react with cold water to form their hydroxides and hydrogen. Potassium and sodium react vigorously; calcium bubbles steadily. For example:

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

The hydroxides make the solutions alkaline. Magnesium reacts very slowly with cold water but reacts visibly with steam when heated:

$Mg + H_2O(g) \rightarrow MgO + H_2$

This produces magnesium oxide. Do not copy a hydroxide product into the steam reaction.

Magnesium, zinc and iron precede hydrogen and react with dilute hydrochloric acid to form their chlorides and hydrogen. Copper, silver and gold follow hydrogen and do not normally react with dilute hydrochloric acid. For example:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

$Fe + 2HCl \rightarrow FeCl_2 + H_2$

Replacing dilute hydrochloric acid with dilute sulfuric acid gives a sulfate instead:

$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

Do not extend these conclusions unconditionally to every acid.

For example, magnesium produces H_2 with both dilute sulfuric acid and steam. The other product is $MgSO_4$ in the first case and MgO in the second. If asked whether hydrogen forms, both columns require yes; the different other product does not negate hydrogen formation.

Metals reacting with oxygen form metal oxides. Magnesium burns in oxygen with a dazzling white light to form white magnesium oxide:

$2Mg + O_2 \rightarrow 2MgO$

This forms an oxide rather than releasing hydrogen. Identify the added reactant before writing the products.

Use observations to establish order. Under the same acid concentration, temperature and exposed metal surface area, compare bubbling per unit time. A faster reaction usually indicates the more reactive metal. Total gas produced also depends on reactant amounts and does not directly measure speed.

Rank experimental results by comparing the same conditions

Code	Observation per unit time in the same dilute HCl	Relative reactivity
X	Many bubbles	Most reactive
Y	Few bubbles	Intermediate
Z	No gas bubbles	Least reactive

At the same temperature, concentration and metal surface area: X > Y > Z.

X, Y and Z are three metals to compare. The > symbol means more reactive. Use gas production per unit time rather than final total gas volume to compare rate.

For a table containing cold water, steam and dilute-acid results, compare within one column first, then combine columns. A metal reacting clearly with cold water is more reactive; one reacting with steam but not clearly with cold water comes next; one reacting only with dilute acid is weaker; one failing to react even with the stated dilute acid is last. If magnesium is explicitly described as reacting very slowly with cold water, retain that qualification rather than changing it to no reaction.

After establishing the relations, check whether the requested order is most to least or least to most reactive. For example, if X > Y > Z, most to least is X, Y, Z, while least to most is Z, Y, X. Reversing the required presentation does not change the experimental relationship.

Why does displacement reveal reactivity? Metal atoms form positive ions by losing electrons. A zinc atom, for example, loses two electrons to form Zn^{2+}. More reactive metals lose electrons more readily and can displace less reactive metals from their aqueous ions.

Displacement: the more reactive metal loses electrons more readily to form positive ions

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Zinc loses 2 electrons

$$Zn \rightarrow Zn^{2+} + 2e^-$$

Copper ion gains 2 electrons

$$Cu^{2+} + 2e^- \rightarrow Cu$$

$2e^-$ transferred

The reaction shows that Zn is more reactive than Cu.

The equations conserve atoms, charge and electron numbers. The central orange arrow shows electrons transferring from zinc to copper ions, not zinc atoms turning into copper atoms.

Zinc reacts with copper ions:

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

Zinc loses electrons and copper ions gain them, establishing Zn as more reactive than Cu. Magnesium, zinc, iron, copper and silver follow the same rule: an earlier metal usually displaces a later one from its aqueous ions. In ionic equations, aq means aqueous and s means solid. An atom and ion of the same named metal are not the same reactant.

How do you read a displacement matrix? The row heading gives the added metal; the column heading gives the dissolved metal ion. A tick means the row metal displaces the column's ions and is more reactive than the corresponding metal. A cross means no displacement under the given conditions. A metal with its own ions also gives no such net displacement.

For example, a tick at magnesium row/Zn^{2+} column gives Mg > Zn. A tick at zinc row/Cu^{2+} column gives Zn > Cu. Together, Mg > Zn > Cu. If asked which forms positive ions more readily, choose the more reactive metal, not the ion that gains electrons more readily.

If information is incomplete, test the missing pair. Several reactions may establish only a partial order. Two metals both displacing the same less reactive metal does not compare those two with each other.

The available comparisons may not establish a complete order

Known experiment	What can be deduced?	What is still unknown?
X displaces Y^{2+}	X > Y	Is X or Z more reactive?
Z displaces Y^{2+}	Z > Y	These two results alone cannot decide

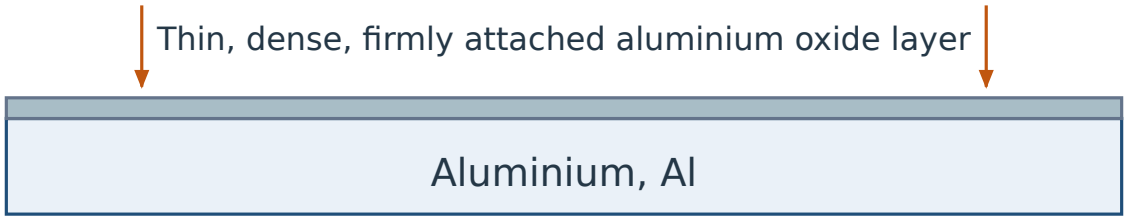
Testing X with Z^{2+} directly checks the missing X–Z relationship.

X, Y and Z represent metals. Both X and Z being more reactive than Y does not establish an order between X and Z.

The diagram gives X > Y and Z > Y but no X–Z comparison. Confirming X displaces Y again adds nothing. Testing X with Z^{2+} can supply the missing relation. If displacement occurs, X > Z, which combines with Z > Y to give X > Z > Y.

A reactive metal can sometimes appear slow to react. Aluminium precedes zinc, iron and copper, but forms a thin, compact, strongly adherent aluminium oxide layer that prevents further water or air contact with the metal beneath. Its apparent lack of reactivity in some observations is surface protection, not a low position in the series.

Aluminium's surface oxide layer prevents further reaction



A slow surface reaction does not mean aluminium is less reactive than copper.

Arrows show external substances approaching and stopping at the protective layer. This is a schematic cross-section; oxide thickness is not to scale.

Reactivity also affects extraction. More reactive metals more readily combine with other elements as positive ions, so recovering metal from an ore compound is generally harder and needs stronger reducing conditions or electrolysis. Aluminium being more reactive than iron does not make it easier to extract. The extraction section applies this relationship to the blast furnace and electrolysis.

Metals losing electrons form _____. Zinc displacing copper from Cu^{2+} solution is more _____ than copper. The solid formed by Mg with steam is _____. Aluminium's apparent lack of reactivity is due to a surface _____.

Answers: positive ions; reactive; magnesium oxide, MgO; compact aluminium oxide layer

Recognising the Question Read reagents and conditions first. Compare observation columns separately before combining them; identify row metals and column ions in displacement tables. Convert results into reactivity relations, check the requested direction and identify any missing pair needed for a complete order.

Worked Example 1 (2025·May/June 0620/11 Q25) · Level 2

The table shows the observations when four metals, Q, R, S and T, are added separately to cold water and to dilute hydrochloric acid.

metal	observation with cold water	observation with dilute hydrochloric acid
Q	slow fizzing	fizzing
R	no reaction	fizzing
S	no reaction	no reaction
T	vigorous fizzing	vigorous fizzing

Which row gives the order of reactivity of the metals?

	least reactive →			most reactive
A	S	Q	R	T
B	T	R	Q	S
C	S	R	Q	T
D	T	Q	R	S

Solution

Check the requested direction. The table runs from least reactive to most reactive.

Compare cold water first. T bubbles vigorously and Q slowly, so T is more reactive than Q. R and S do not react, so this column alone cannot distinguish them. Q reacts with cold water and therefore precedes both.

Use dilute hydrochloric acid for the remaining pair. R bubbles with the acid while S still does not react, so R is more reactive than S. Together, most to least is T > Q > R > S.

The question requests the reverse: S, R, Q, T, row C. Choose C. Guessing the R–S order from cold water alone would omit the necessary acid evidence.

Worked Example 2 (2025·May/June 0620/21 Q27) · Level 3

Four metals J, K, L and M are added separately to aqueous solutions of their ions.

The results are shown.

	aqueous ions			
metal	J^{2+}	K^{2+}	L^{2+}	M^{2+}
J	✗	✓	✓	✓
K	✗	✗	✗	✓
L	✗	✓	✗	✓
M	✗	✗	✗	✗

key

✓ = reaction

✗ = no reaction

Which statement about the metals is correct?

A J has a greater tendency to form positive ions than K.

B L has a greater tendency to form positive ions than J.

C K has a greater tendency to form positive ions than L.

D M has a greater tendency to form positive ions than K.

Solution

Identify rows, columns and the key. J, K, L and M at the left are added metals. J^{2+}, K^{2+}, L^{2+} and M^{2+} across the top are dissolved ions. A tick means reaction and a cross no reaction. A row metal that displaces a column ion is more reactive and loses electrons to form positive ions more readily.

J reacts with K^{2+}, L^{2+} and M^{2+}, so it precedes all three. L displaces K^{2+} and M^{2+} but not J^{2+}, giving:

$J > L > K$

L is also more reactive than M. K displaces only M^{2+}, giving $K > M$. Combining:

$J > L > K > M$

A states J forms positive ions more readily than K, matching this order. B reverses L and J; C reverses K and L; D reverses M and K. Choose A.

The comparison concerns metal atoms **losing** electrons. Positive ions gaining electrons to become metal is the other half-reaction, not the definition of metal reactivity.

Practice

1. (2025·May/June 0620/11 Q26) · Level 2

Magnesium is reacted separately with dilute sulfuric acid and with steam.

Which row correctly identifies if hydrogen is formed as a product in each reaction?

	reaction with dilute sulfuric acid	reaction with steam
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

key

✓ = hydrogen is formed

✗ = hydrogen is **not** formed

2. (2025·May/June 0620/12 Q27) · Level 2

Four metals, P, Q, R and S, are separately reacted with dilute hydrochloric acid.

The table shows the observation for each metal.

metal	observation
P	lots of bubbles produced
Q	few bubbles produced
R	no bubbles produced
S	very few bubbles produced

What is the order of reactivity of these four metals, from most reactive to least reactive?

A R → S → Q → P

B P → S → Q → R

C P → Q → S → R

D Q → R → P → S

3. (2025·May/June 0620/21 Q26) · Level 2

Three statements about solid Z are listed.

- It reacts with dilute hydrochloric acid to produce hydrogen.
- It reacts very slowly with cold water and quickly with steam.
- It is a good thermal conductor.

What is solid Z?

A gold

B magnesium

C potassium

D silver

4. (2025·May/June 0620/12 Q26) · Level 3

Which statements about metals are correct?

1 Calcium reacts faster than magnesium with cold water.

2 Aluminium is more easily extracted than iron from its ore.

3 Copper reacts faster than iron with dilute hydrochloric acid.

4 Potassium and silver are both good conductors of electricity.

A 1 and 2 **B** 1 and 4 **C** 2 and 3 **D** 3 and 4

5. (2025·May/June 0620/13 Q28) · Level 3

The table gives some information about the reactions of four metals, J, K, L and M.

metal	reactions
J	reacts with steam and hydrochloric acid but **not** with cold water
K	reacts with hydrochloric acid but **not** with steam or cold water
L	reacts with hydrochloric acid and with cold water
M	does **not** react with hydrochloric acid

What is the order of reactivity of metals J, K, L and M?

	most reactive ⟶ least reactive			
A	L	J	K	M
B	L	K	J	M
C	M	J	K	L
D	M	K	J	L

6. (2025·May/June 0620/22 Q26) · Level 3

The ionic equations for some reactions between metals and aqueous metal ions are shown.

$$Cu + 2Ag^{+} \rightarrow 2Ag + Cu^{2+}$$

$$Mg + Zn^{2+} \rightarrow Mg^{2+} + Zn$$

$$Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$$

$$Sn + Cu^{2+} \rightarrow Sn^{2+} + Cu$$

Which further information can be used to place copper, magnesium, silver, tin and zinc in order of their reactivity?

A Zinc reacts with Sn^{2+} ions.

B Copper does **not** react with Sn^{2+} ions.

C Magnesium reacts with Cu^{2+} ions.

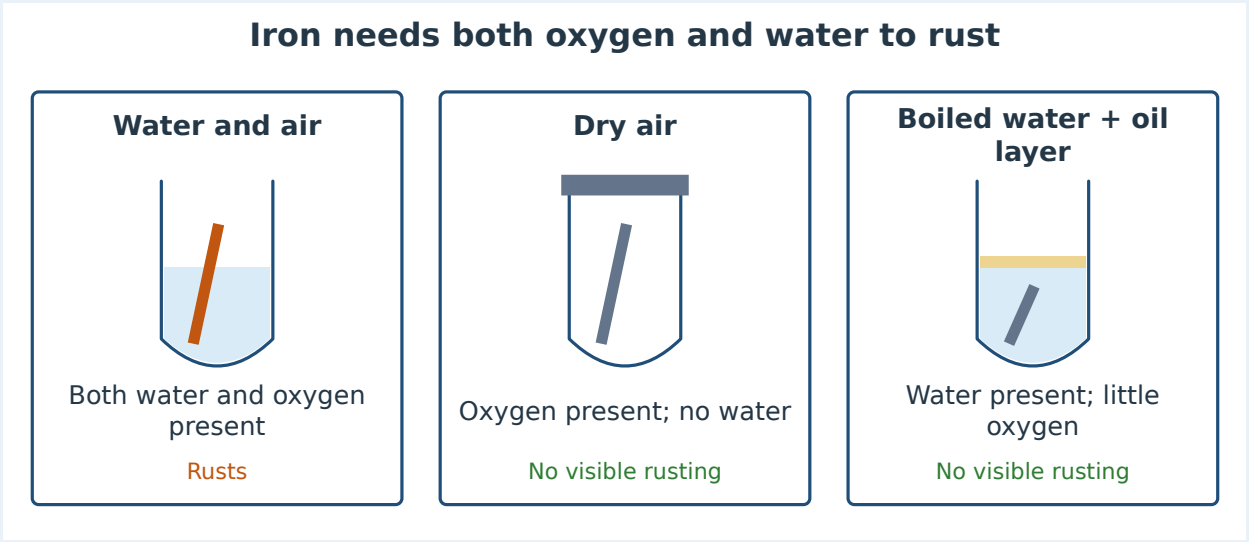
D Magnesium reacts with Ag^{+} ions.

9.5 Metal corrosion

Type 1: Explaining rust prevention using water, oxygen, barriers and electron loss

Key Ideas **Identify rust and its requirements first.** Corrosion is gradual damage as a material reacts with its environment; rusting of iron and steel is one example. In this course, rust is hydrated iron(III) oxide. Hydrated indicates water content and III specifies iron's oxidation number of +3.

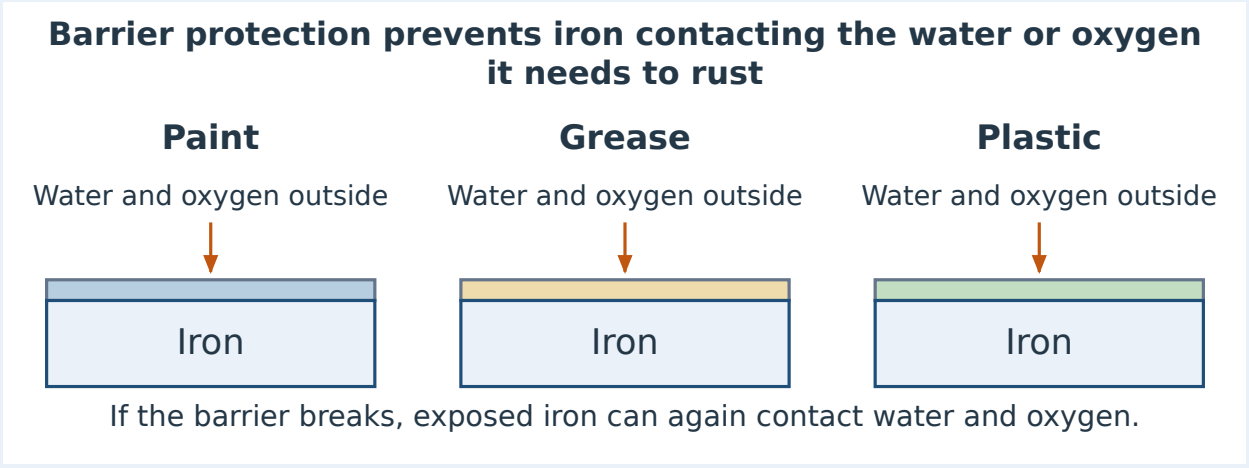
Iron needs access to both oxygen and water to rust. Dry air lacks water. Water with its dissolved oxygen removed and air excluded lacks oxygen. Neither supplies both necessary conditions.



Similar iron samples are compared under different water and oxygen conditions. Boiling removes dissolved oxygen; an oil layer prevents atmospheric oxygen re-entering the water.

For example, a nail in water exposed to air has both water and oxygen and rusts. A nail in dry air lacks water. A nail in boiled water covered with oil lacks dissolved oxygen, while the oil prevents oxygen being resupplied from air. Air alone or water alone is not enough to classify the latter cases as rusting.

Barrier protection prevents necessary contact. Paint, grease and plastic coatings form a barrier between iron and its surroundings. Excluding water or oxygen prevents access to the reactants required for rusting. The coating does not turn iron into a less reactive metal.

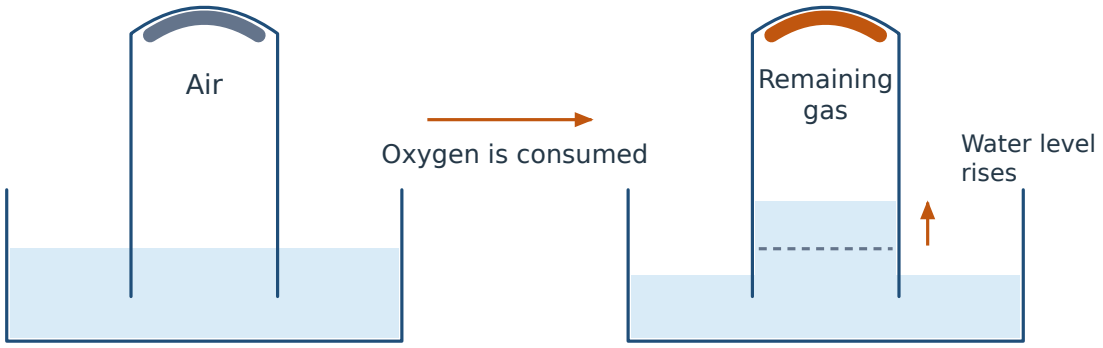


Vertical arrows show water and oxygen approaching each protective coating and stopping outside it. The comparison is between coatings, not conversion of paint into grease or plastic.

For example, a complete grease coating prevents easy air and water contact with an iron strip, reducing rusting. Scratching ordinary paint exposes iron to water and oxygen, allowing rust at the scratch. Coating integrity is part of the conditions to check.

Some experiments use water level to indicate gas change. Place damp iron in an inverted tube with its opening under water. Rusting consumes oxygen from the trapped air. Keep temperature, external atmospheric pressure and other conditions unchanged, then track the gas: oxygen enters solid rust products and no longer remains in the original gas space.

Why does water rise when oxygen is used up in an inverted tube?



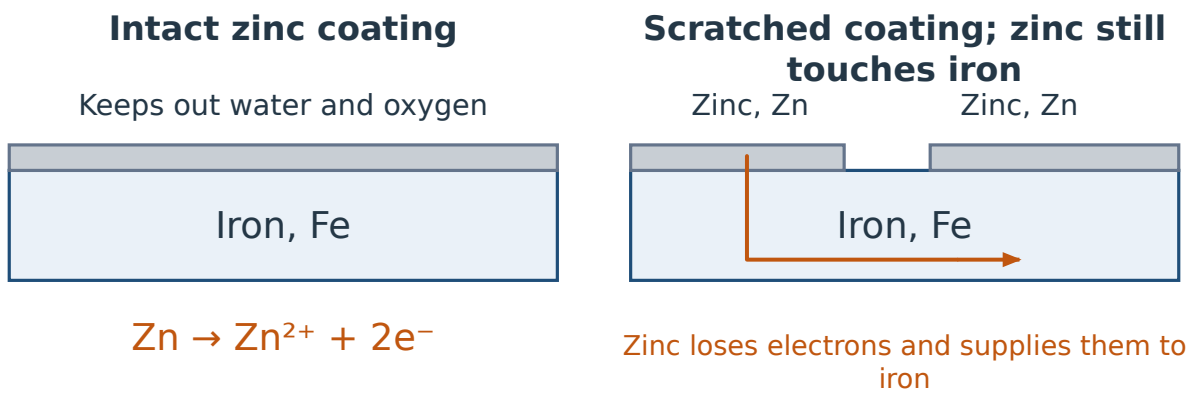
The central arrow shows the sequence during rusting; the short arrow on the right shows water rising. The remaining gas volume decreases. No volume fraction is specified.

Oxygen consumption initially lowers gas pressure inside the tube. Atmospheric pressure then pushes water in until a new balance is reached. The remaining gas occupies less volume and the water level rises. Less gas does not make the water fall; water enters the space previously occupied by gas.

For example, two tubes have equal iron surface area, temperature and other conditions. One iron sample has a complete grease coating, while the other is bare and damp. The coated sample hardly rusts, consumes no appreciable oxygen and gives little water-level change. The bare sample consumes oxygen and the water rises. First decide which rusts, then which gas volume decreases.

Galvanising adds another form of protection. Coating iron with zinc is galvanising. An intact zinc coating first acts as a barrier against water and oxygen. Even with a small scratch, zinc can lose electrons preferentially and protect exposed iron, provided zinc remains in contact with it and has not been used up.

Galvanising provides both a barrier and sacrificial protection



Cross-sections show intact and scratched zinc coatings. Orange arrows represent zinc supplying protective electrons to iron. Zinc must remain in contact with the iron and not be exhausted.

Zinc precedes iron in the reactivity series and loses electrons more readily:

$$Zn \rightarrow Zn^{2+} + 2e^-$$

It supplies electrons to iron and suppresses iron's own oxidation. Zinc is consumed preferentially, so this is sacrificial protection. Sacrificial means that the protective metal oxidises first, not that iron rusts faster.

For example, "zinc is more reactive, so it accepts electrons more readily" reverses the direction. A reactive metal more readily **loses** electrons to form positive ions. The wording must be loses electrons more readily to explain sacrificial protection.

Rust is called _____. Grease gives _____ protection; galvanising also gives _____ protection. During sacrificial protection, zinc more readily _____ electrons.

Answers: hydrated iron(III) oxide; barrier; sacrificial; loses

Recognising the Question Check whether iron contacts both water and oxygen. For an inverted tube, connect oxygen consumption to gas volume and water level. For galvanising, check the barrier, zinc–iron contact and the direction of zinc's electron loss.

Worked Example 1 (2025·May/June 0620/21 Q29) · Level 2

Galvanising prevents iron from rusting.

Which statements about galvanising are correct?

1 Galvanising involves coating the iron with silver.

2 Galvanising prevents oxygen and water from coming into contact with the iron.

3 Galvanising works because the metal coating loses electrons more readily than iron.

A 1, 2 and 3 **B** 1 and 3 only **C** 2 and 3 only **D** 2 only

Solution

Galvanising coats iron with zinc, not silver, so statement 1 is wrong. An intact zinc layer separates iron from oxygen and water and provides barrier protection, so statement 2 is correct.

Zinc is more reactive than iron and loses electrons more readily:

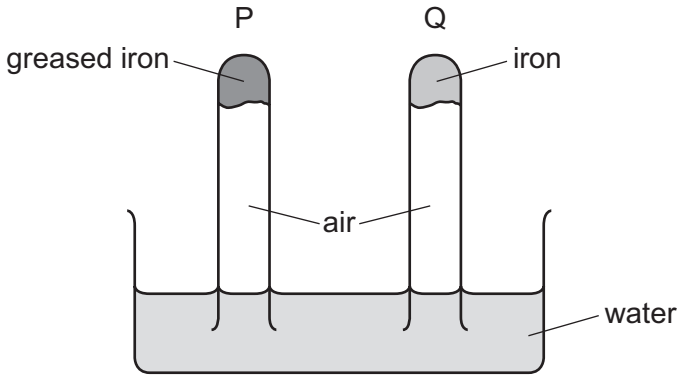
$Zn \rightarrow Zn^{2+} + 2e^-$

While zinc remains in contact with iron, its preferential oxidation gives sacrificial protection. Statement 3, loses electrons more readily, is correct.

Statements 2 and 3 are correct. Choose C. Identify the coating material, barrier action and electron-loss mechanism together; excluding air alone is not the whole explanation.

Worked Example 2 (2025·May/June 0620/11 Q29) · Level 3

The diagram shows an experiment to investigate the corrosion of iron.



What happens to the water level in tubes P and Q?

	tube P	tube Q
A	rises	falls
B	no change	rises
C	falls	rises
D	no change	no change

Solution

Identify the iron in each tube. P contains greased iron at the top; Q contains bare iron. Both tubes are inverted in water and contain air.

Assess oxygen consumption. Grease on P blocks water and oxygen contact, so the iron hardly rusts. No appreciable oxygen is consumed by rusting and its water level stays nearly unchanged. Iron in Q contacts damp air, rusts and consumes oxygen.

As oxygen decreases in Q, the gas occupies less volume and atmospheric pressure pushes water into the tube. The level rises. The combination P no change, Q rises is B.

The iron does not pull water upwards, nor is all the air used up. The explanation is rusting, oxygen consumption, reduced gas amount and water entering the tube.

Practice

1. (2025·May/June 0620/13 Q29) · Level 1

What is the chemical name for rust?

A anhydrous iron(II) oxide

B anhydrous iron(III) oxide

C hydrated iron(II) oxide

D hydrated iron(III) oxide

2. (2025·May/June 0620/22 Q27) · Level 2

Zinc is used to galvanise iron to prevent it from rusting.

Which statements are correct?

1 Galvanising is an example of a barrier method.

2 If the zinc is scratched, the iron will rust very quickly.

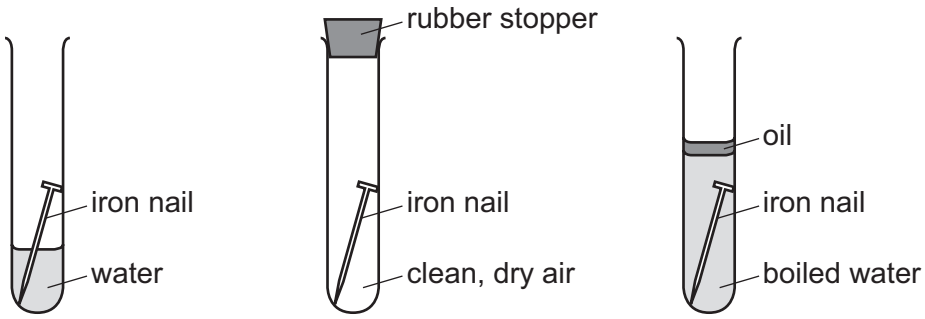
3 Galvanising is an example of sacrificial protection.

4 Zinc is more reactive than iron and so accepts electrons more readily.

A 1, 2 and 4 **B** 1 and 3 **C** 2, 3 and 4 **D** 2 and 3 only

3. (2025·May/June 0620/12 Q28) · Level 3

The diagram shows iron nails which are left for one week in separate test-tubes with the conditions shown.



Which test-tubes show evidence of rusting after one week?

A 1 and 3 **B** 1 only **C** 2 and 3 **D** 2 only

9.6 Extraction of metals

Type 1: Tracking heat supply, reduction and impurity removal in a blast furnace

Key Ideas **First identify the metal's original form.** An ore is a mineral raw material used to extract a metal, usually containing metal compounds mixed with other substances. If iron is present as an oxide, extraction must convert iron in that compound into metallic iron. Grinding or filtering the ore does not carry out this chemical change.

More reactive metals are generally harder to obtain from their compounds. Locate them relative to carbon: aluminium precedes carbon, so carbon cannot reduce aluminium oxide to aluminium. Iron follows carbon and can be extracted by reduction. This difficulty concerns chemical extraction, not how easy the ore is to mine.

Use reactivity first: can carbon reduce the metal compound?

Position	Typical extraction rule in this course
Aluminium is above carbon and more reactive	Carbon cannot reduce aluminium oxide; use electrolysis
Iron is below carbon and less reactive	CO made from carbon can reduce iron oxide
Higher reactivity	Usually harder to extract the metal from its compounds

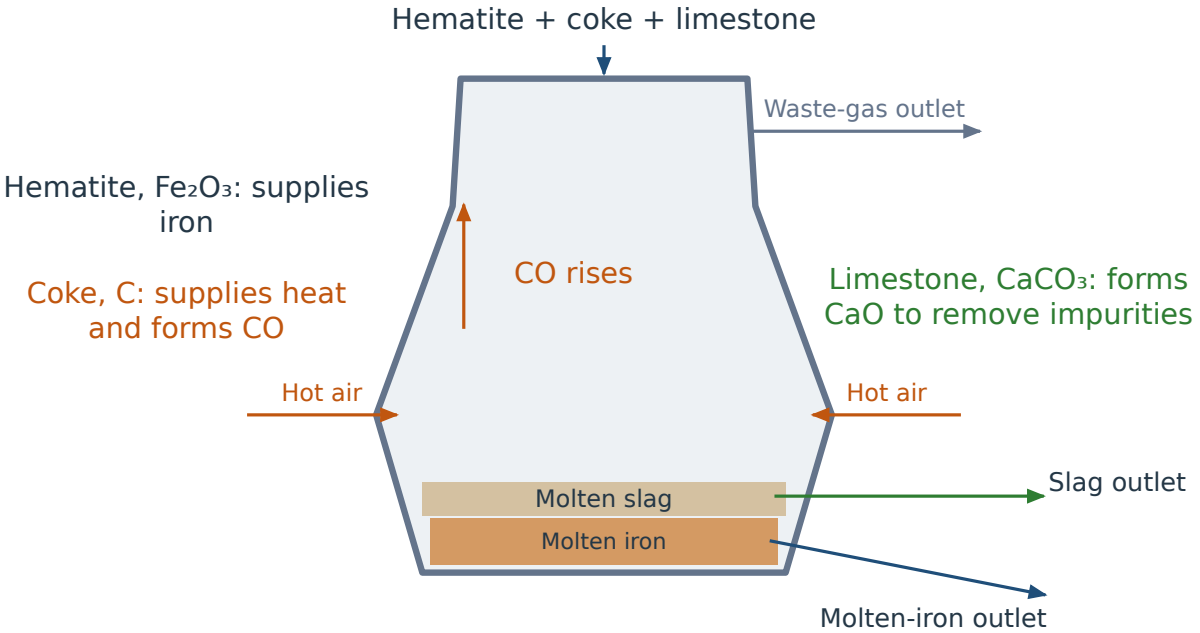
Part of the series: Mg → Al → C → Zn → Fe

The comparison concerns typical oxide extraction in this course. The arrow follows decreasing reactivity, not a reaction direction.

For example, iron and aluminium cannot simply use the same extraction method because Al is above C and Fe below C. Aluminium oxide requires electrolysis; iron oxide can be reduced by CO generated in a blast furnace. Aluminium's resistance to visible corrosion does not place it below iron; its protective surface layer was explained previously.

Identify the three solid blast-furnace inputs. Hematite supplies iron(III) oxide, Fe_2O_3. Coke is mainly carbon, C, and limestone mainly calcium carbonate, $CaCO_3$. They enter at the top while hot air is blown in below. Molten iron and slag collect near the bottom and are tapped off separately.

Blast furnace: each of the three solid raw materials has a role



Slag floats above molten iron; positions explain the process and are not drawn to industrial scale.

Blue arrows show charging materials or removing molten iron; orange arrows show hot air or CO-containing gases; green shows slag removal. Arrow thickness does not represent flow rate.

For example, the input supplying iron is hematite. The substance supplying CaO to remove SiO_2 impurities is limestone. Coke supplies heat and generates the reducing agent. Do not describe all three materials as reducing iron.

Supply heat, then produce the reducing agent CO. Carbon in coke burns in oxygen from the hot air to form carbon dioxide:

$C + O_2 \rightarrow CO_2$

This exothermic step supplies the furnace's high temperature.
The carbon dioxide then reacts with hot coke to form carbon monoxide:

$C + CO_2 \rightarrow 2CO$

CO_2 loses some oxygen as it becomes CO, so CO_2 is reduced in this step.

From heat-producing combustion to the CO that reduces iron oxide

1. Coke burns and supplies heat
$C + O_2 \rightarrow CO_2$

↓

2. CO_2 reacts with hot coke to make CO
$C + CO_2 \rightarrow 2CO$

↓

3. CO removes oxygen from iron oxide, forming iron
$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Vertical arrows show one step supplying the substance needed by the next. Step ① releases heat; ② and ③ are different reactions.

Carbon monoxide then reacts with iron(III) oxide to form iron and carbon dioxide:

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Fe_2O_3 loses oxygen and is reduced. CO gains oxygen to become CO_2 and is oxidised. CO causes another substance to be reduced, so it is the reducing agent.
For example, if coke combustion, CO reduction of iron oxide, and CO_2 reacting with coke are numbered ①, ② and ③ respectively, the dependency order is ①→③→②. CO_2 must form before it can produce CO, and CO must form before reducing Fe_2O_3. If asked specifically for the heat-supplying reaction, choose:

$C + O_2 \rightarrow CO_2$

Not every equation containing carbon is the heat-supplying reaction.
Impurity removal follows a separate sequence. Calcium carbonate in limestone decomposes on heating into calcium oxide and carbon dioxide:

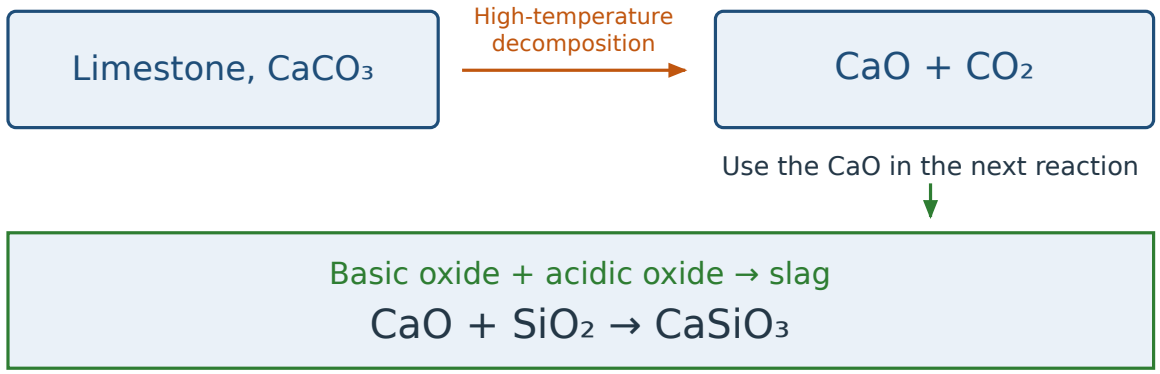
$CaCO_3 \rightarrow CaO + CO_2$

This is thermal decomposition: heating breaks one substance into two products.
Basic CaO then reacts with acidic SiO_2 in the ore to form calcium silicate:

$CaO + SiO_2 \rightarrow CaSiO_3$

Calcium silicate forms slag, which is removed separately from molten iron.

Limestone decomposes first; the calcium oxide then removes an impurity



The slag is calcium silicate; this removes SiO_2 without reducing it to silicon.

$CaCO_3$ has two decomposition products. The following arrow tracks CaO reacting with the ore's SiO_2 impurity.

For example, the two reactants forming slag are CaO and SiO_2. Only if asked for the source of that CaO should you trace back to $CaCO_3$. Keep $CaCO_3$, CaO and $CaSiO_3$ distinct: they are the main input component, intermediate product and main slag component respectively. Silicon(IV) oxide is SiO_2 and calcium oxide is CaO.

Check whether the extracted metal is pure. An ore may contain more than one metal. If reduction produces two metals together, further separation is required to obtain either one pure. This disadvantage follows from the stated products; no invented substances are needed.

Obtaining a metal by reduction does not automatically give a pure product



Identify all products of reduction before deciding whether further separation is needed.

M and N are two metals produced together. Obtaining the target M requires separating it from N.

For example, if M is the target but both M and N form, a disadvantage is that M must be separated from N. Producing a metal does not eliminate all subsequent steps. In the later ore question, likewise read the metal products already given in the equation.

The heat-supplying blast-furnace reaction is _____. The substance directly reducing Fe_2O_3 is _____. $CaCO_3$ decomposes into _____ and _____. The reactants forming slag are _____ and _____.

Answers: $C + O_2 \rightarrow CO_2$; CO; CaO; CO_2; CaO; SiO_2

Recognising the Question Distinguish iron supply, heat supply, CO production, iron reduction and impurity removal before selecting the reaction. For sequence questions, identify which substance each step supplies to the next. For purity, check whether the stated products include another metal.

Worked Example 1 (2025·May/June 0620/43 Q1(f)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(f) react together to form slag in the blast furnace

Solution

The question asks which listed substances react together to form blast-furnace slag. React together requires two reactants.
Use the impurity-removal equation:
$CaO + SiO_2 \rightarrow CaSiO_3$
CaO is calcium oxide and SiO_2 silicon(IV) oxide. $CaSiO_3$ is the calcium silicate in slag. Write **calcium oxide and silicon(IV) oxide**.

The two names correspond to two marking points. Iron oxide in the ore and carbon involved in reduction have other furnace roles and do not belong in this impurity-removal reaction.

Worked Example 2 (2025·May/June 0620/22 Q28) · Level 2

Which equation represents a reaction that occurs in a blast furnace during the extraction of iron from hematite?

A $Ca + CO_2 \rightarrow CaO + CO$

B $Fe_2O_3 + 3Ca \rightarrow 2Fe + 3CaO$

C $Ca(OH)_2 \rightarrow CaO + H_2O$

D $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Solution

Hematite supplies Fe_2O_3. The question asks for the blast-furnace reaction extracting iron, so iron in that compound must become metallic Fe.
CO removes oxygen from Fe_2O_3:
$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
This is D. Check atoms: each side has 2 Fe, 3 C and 6 O.

A introduces calcium metal into this furnace step. B may be a balanceable reduction equation, but it is not the taught blast-furnace process, which does not use calcium metal as reductant. C gives $Ca(OH)_2$, whereas limestone is mainly $CaCO_3$. Choose D.

Worked Example 3 (2025·October/November 0620/23 Q32) · Level 2

Iron is extracted from iron(III) oxide in the blast furnace.

Three steps in the process are listed.

1 Carbon (coke) is completely oxidised.

2 Iron(III) oxide is reduced.

3 Carbon dioxide is reduced to carbon monoxide.

What is the correct order of the steps in the extraction of iron?

A 1 → 2 → 3 **B** 1 → 3 → 2 **C** 2 → 1 → 3 **D** 3 → 1 → 2

Solution

Order the reactions by where their required substances come from, not by their listed numbers.
Step 1: complete oxidation of coke:
$C + O_2 \rightarrow CO_2$
This first supplies CO_2 and releases heat.
Step 3: CO_2 reacts with more hot coke:
$CO_2 + C \rightarrow 2CO$
This forms the CO needed next.
Step 2: CO reduces Fe_2O_3:
$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
The order is therefore 1, 3, 2, option B. Placing 2 before 3 would leave the source of its reducing CO unexplained.

Practice

1. (2025·May/June 0620/21 Q28) · Level 1

Which process is used in the extraction of iron from hematite?

A electrolysis

B precipitation

C reduction

D fractional distillation

2. (2025·October/November 0620/22 Q28) · Level 1

Iron is extracted from hematite in the blast furnace.

Which reaction increases the temperature in the blast furnace to over 1500 °C?

A calcium carbonate → calcium oxide + carbon dioxide

B calcium oxide + silicon(IV) oxide → calcium silicate

C carbon + oxygen → carbon dioxide

D carbon dioxide + carbon → carbon monoxide

3. (2025·May/June 0620/41 Q1(h)) · Level 2

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(h) is produced when limestone thermally decomposes in the blast furnace

4. (2025·February/March 0620/22 Q30) · Level 2

Which row gives the symbol equation for the formation of carbon monoxide and for the reduction of iron(III) oxide in a blast furnace?

	equation for the formation of carbon monoxide	equation for the reduction of iron(III) oxide
A	$2C + O_2 \rightarrow 2CO$	$FeO + CO \rightarrow Fe + CO_2$
B	$CO_2 + C \rightarrow 2CO$	$FeO + CO \rightarrow Fe + CO_2$
C	$C + O_2 \rightarrow CO_2$	$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
D	$CO_2 + C \rightarrow 2CO$	$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

5. (2025·May/June 0620/42 Q4(a)(ii)) · Level 2

Chromium is the element with atomic number 24 in the Periodic Table.

(a) The main ore of chromium is chromite. Chromite contains $FeCr_2O_4$.

$FeCr_2O_4$ reacts with carbon.

$$FeCr_2O_4 + 4C \rightarrow Fe + 2Cr + ...$$

(ii) Suggest **one** disadvantage of extracting chromium by reacting chromite with carbon.

Type 2: Explaining aluminium extraction using reactivity and electrolysis

Key Ideas **Aluminium is above carbon, so extraction needs electrical energy.** Aluminium is more reactive than carbon and cannot be extracted by carbon reduction of its oxide as iron is in a blast furnace. Its main ore is bauxite, which contains aluminium oxide, Al_2O_3. Industry first obtains relatively pure aluminium oxide, then electrolyses it. Detailed industrial purification of bauxite is beyond this course.

Keep aluminium's ore, feedstock, solvent and product distinct

Substance	Role in extraction
Bauxite	Main ore containing aluminium oxide
Aluminium oxide, Al_2O_3	Purified feedstock supplying aluminium ions
Molten cryolite	Dissolves aluminium oxide and lowers operating temperature
Aluminium, Al	Metal obtained at the cathode after electrolysis

Cryolite is neither aluminium's main ore nor aluminium metal. Detailed industrial bauxite purification is not required here.

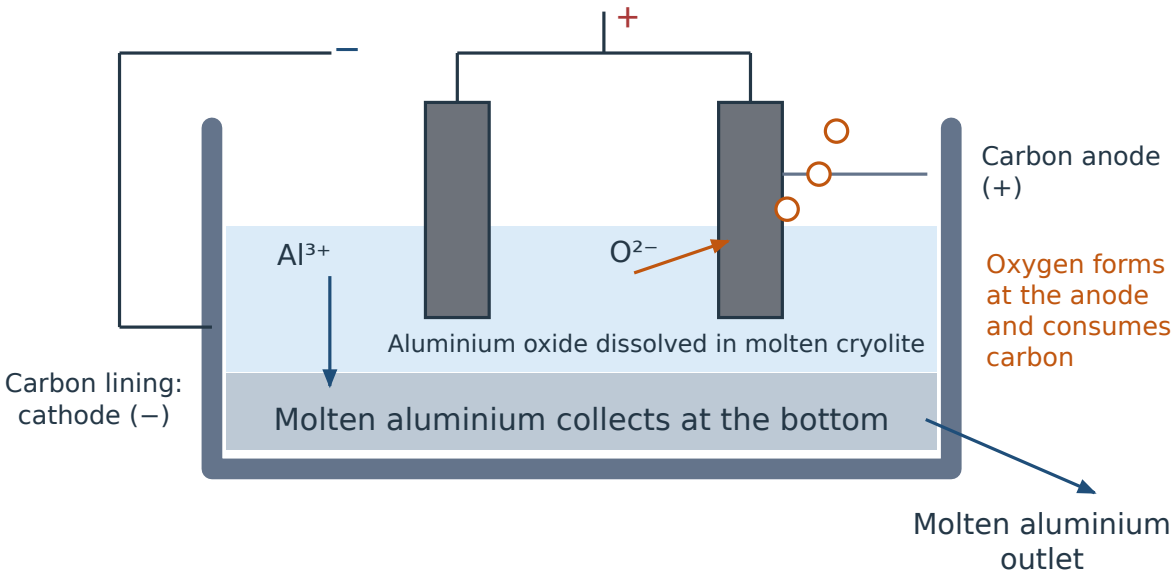
Cryolite reduces operating temperature. Melting aluminium oxide directly requires a very high temperature. Dissolving it in molten cryolite, Na_3AlF_6, allows electrolysis at a lower temperature and reduces the energy needed to maintain heating. Retain the complete explanation: aluminium oxide dissolves in molten cryolite, the operating temperature falls, and heating energy decreases. Cryolite is not a catalyst and does not remove the need for electrical energy.

For example, the main ore is bauxite; the added substance lowering operating temperature is cryolite; the purified substance directly supplying aluminium ions is aluminium oxide. These are different roles in the same process.

Label electrode charges first. Carbon anodes connect to the positive terminal; the carbon-lined cathode connects to the negative terminal. Mobile Al^{3+} ions move to the cathode and gain electrons to form aluminium. O^{2-} ions move to the anode and lose electrons to form oxygen. Molten aluminium collects at the bottom and is tapped off. Ions carry charge in the electrolyte; electrons carry it through the external wires.

Electrolytic cell: oxygen at the positive anode, aluminium at the negative cathode

D.c. supply: positive terminal to carbon anodes; negative terminal to carbon lining



Moving ions conduct in the melt; electrons travel through external wires, not freely through the electrolyte.

Blue arrows show Al^{3+} moving to the cathode and molten aluminium being removed. Orange arrows show O^{2-} moving to the anode. Circles indicate gas production and are not for counting.

For example, a statement placing aluminium at the positive anode and oxygen at the negative cathode reverses both. Positive Al^{3+} moves to the negative cathode and gains electrons to become Al; negative oxide ions move to the positive anode. Aluminium being a metal does not associate its production with the positive electrode.

Check both atoms and charge in half-equations. At the cathode, 1 Al^{3+} ion accepts 3 electrons to become neutral Al:

$$Al^{3+} + 3e^- \rightarrow Al$$

Electrons on the left are gained, so this is reduction.

At the anode, the gas is O_2, not single O atoms. Two O^{2-} supply two oxygen atoms and lose 4 electrons, written on the right:

$$2O^{2-} \rightarrow O_2 + 4e^-$$

The left has total charge −4. On the right, O_2 is neutral and $4e^-$ totals −4, conserving charge. This is oxidation.

Half-equations must balance both atom counts and total charge

Cathode: aluminium ions gain electrons	Anode: oxide ions lose electrons
$Al^{3+} + 3e^- \rightarrow Al$	$2O^{2-} \rightarrow O_2 + 4e^-$
Charge: $+3 - 3 = 0$	Charge: $2 \times (-2) = -4$
Reduction	Oxidation

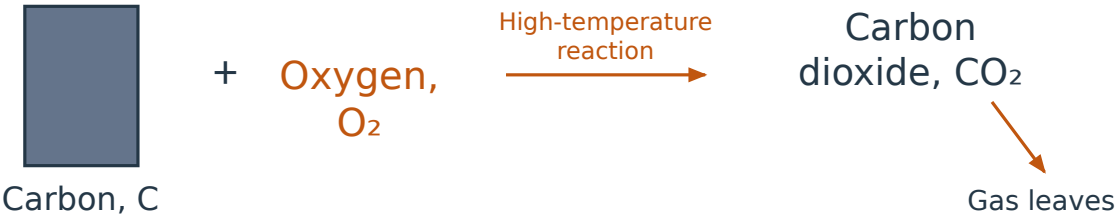
O_2 contains two oxygen atoms, so start with $2O^{2-}$; together they lose four electrons.

The columns show different electrode reactions. Electron numbers must be equalised before adding them; each is checked separately here.

For example, consider:
$O^{2-} \rightarrow O_2 + 2e^-$
Both charges are −2, but one oxygen atom on the left becomes two on the right, so the equation is still wrong. Change the left to $2O^{2-}$, then the electrons to 4. Checking charge alone misses the atom imbalance. If state symbols are included, aluminium is Al(l) and oxygen O_2(g).

Why do the carbon anodes shrink? At high temperature, oxygen formed at the anode reacts with its carbon to produce carbon dioxide:
$C + O_2 \rightarrow CO_2$
Carbon moves from the solid anode into a gaseous product, consuming the anode and requiring regular replacement. This is a chemical reaction, not melting. High temperature alone is an incomplete explanation without oxygen and CO_2.

The carbon anode does more than heat up: oxygen consumes it



$C + O_2 \rightarrow CO_2$; carbon becomes a gas product, so the anode wears away and must be replaced.

The thick grey column is solid carbon anode material. The product arrow shows CO_2 leaving as a gas, explaining the loss of solid material.

For example, “the anode needs replacement because gas forms” omits the key link. Oxygen produced at the anode reacts with carbon to form CO_2 and consumes the electrode material. The O_2 is first formed and then reacts with carbon; the carbon anode is not inert.

Aluminium’s main ore is _____. The added substance lowering operating temperature is _____. The cathode half-equation is _____ the anode half-equation is _____. Carbon anodes react with _____ to form _____, so they require replacement.

Answers: bauxite; cryolite; $Al^{3+} + 3e^- \rightarrow Al$; $2O^{2-} \rightarrow O_2 + 4e^-$; oxygen; CO_2

Recognising the Question Separate the ore name from the substance lowering operating temperature. Label electrode charges and read ionic charges before placing electrons on the left or right. Explain anode replacement through oxygen reacting with carbon to form CO_2 and consume carbon.

Worked Example 1 (2025·May/June 0620/42 Q3(d)(iii)) · Level 2

This question is about electrolysis.

(d) Aluminium is extracted from its purified ore by electrolysis.

(iii) Explain why the anodes need to be continually replaced during this process.

Solution

Explain why carbon anodes need continual replacement. First track the anode product: oxide ions lose electrons at the positive anode and form oxygen.
At high temperature, oxygen reacts with the carbon or graphite anode to form carbon dioxide:
$C + O_2 \rightarrow CO_2$
Carbon atoms enter the gaseous product, continuously consuming anode material and requiring replacement.

A complete answer is: **The carbon anodes react with oxygen to form carbon dioxide, so the anodes are consumed and need replacing.** The two marking points are carbon anodes reacting with oxygen and formation of CO_2. Merely saying hot, oxidation or gas forms does not complete the explanation of material loss.

Worked Example 2 (2025·October/November 0620/22 Q29) · Level 3

Aluminium is extracted by the electrolysis of purified bauxite.

Which row shows the half-equations at the electrodes and explains why one of the electrodes has to be replaced regularly?

	half-equation at anode	half-equation at cathode	electrode replacement
A	$2O^{2-} \rightarrow O_2 + 4e^-$	$Al^{3+} + 3e^- \rightarrow Al$	cathode replaced as it gets coated with aluminium
B	$2O^{2-} \rightarrow O_2 + 4e^-$	$Al^{3+} + 3e^- \rightarrow Al$	anode replaced as it burns away in oxygen
C	$Al^{3+} + 3e^- \rightarrow Al$	$2O^{2-} \rightarrow O_2 + 4e^-$	cathode replaced as it burns away in oxygen
D	$Al^{3+} + 3e^- \rightarrow Al$	$2O^{2-} \rightarrow O_2 + 4e^-$	anode replaced as it gets coated with aluminium

Solution

Check each column heading. At the anode, oxide ions lose electrons:
$2O^{2-} \rightarrow O_2 + 4e^-$
At the cathode, aluminium ions gain electrons:
$Al^{3+} + 3e^- \rightarrow Al$
The first two columns therefore allow A or B. C and D swap the electrode reactions.

Now check electrode replacement. Oxygen forms at the anode and reacts with its carbon to produce CO_2. The anode consumed by oxygen is replaced. B, anode burns away in oxygen, fits.

A has the first columns correct but describes replacing a cathode covered in aluminium, which does not explain the actual carbon-anode consumption. Read the final column as well as recognising the half-equations. Choose B.

Practice

1. (2025·May/June 0620/41 Q1(e)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(e) contains aluminium oxide

2. (2025·May/June 0620/42 Q3(d)(i)) · Level 1

This question is about electrolysis.

(d) Aluminium is extracted from its purified ore by electrolysis.

(i) Name this ore of aluminium.

3. (2025·May/June 0620/42 Q3(d)(ii)) · Level 1

This question is about electrolysis.

(d) Aluminium is extracted from its purified ore by electrolysis.

(ii) Name the substance added to this process to reduce the operating temperature.

4. (2025·October/November 0620/43 Q3(a)(iv)) · Level 2

This question is about aluminium and other metallic elements.

(a) Aluminium is extracted from purified bauxite by electrolysis. Fig. 3.1 shows a diagram of the electrolytic cell used.

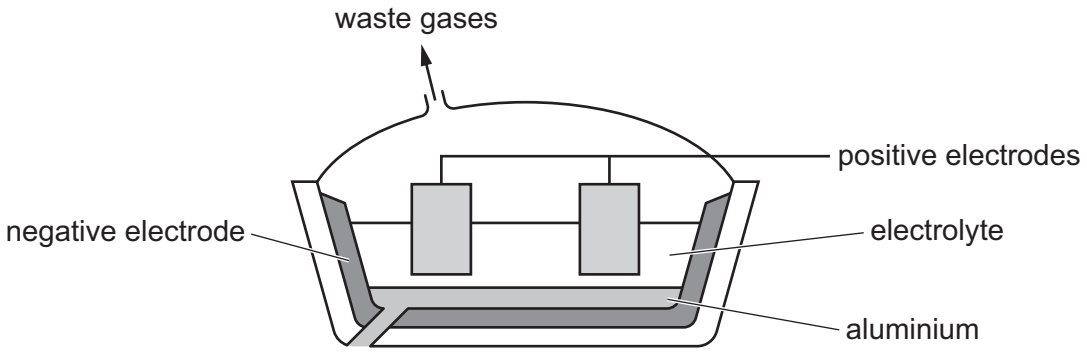


Fig. 3.1

(iv) Write the ionic half-equation for the production of aluminium at the negative electrode.

5. (2025·May/June 0620/11 Q27) · Level 2

Which statements about aluminium are correct?

1 It is more reactive than calcium.

2 The main ore of aluminium is bauxite.

3 It can be extracted from its oxide using carbon.

4 Brass is an alloy of aluminium and copper.

A 1 and 2 **B** 1 and 3 **C** 2 only **D** 3 and 4

Environmental chemistry

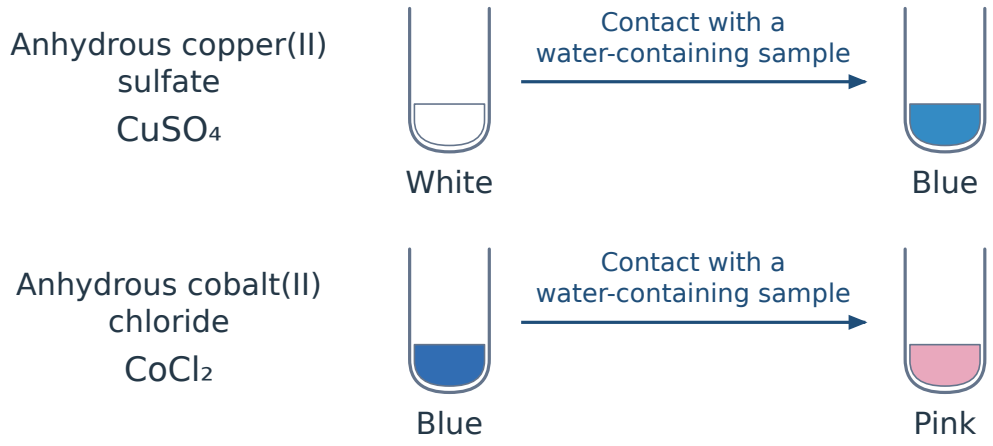
10.1 Water

Type 1: Distinguishing water detection, purity testing and domestic water treatment

Key Ideas **Contains water and is pure water are different conclusions.** Water has formula H_2O. Salt water contains H_2O and dissolved salt, so it gives a positive chemical water test but is not pure water. Identify whether the question asks for presence of water or purity before choosing a method.

Pair each water-test reagent with its colour change. Anhydrous means the reagent initially contains no water of crystallisation. Anhydrous copper(II) sulfate, $CuSO_4$, is white and turns blue with water. Anhydrous cobalt(II) chloride, $CoCl_2$, is blue and turns pink with water.

Two chemical tests for water: remember the reagent and direction of colour change



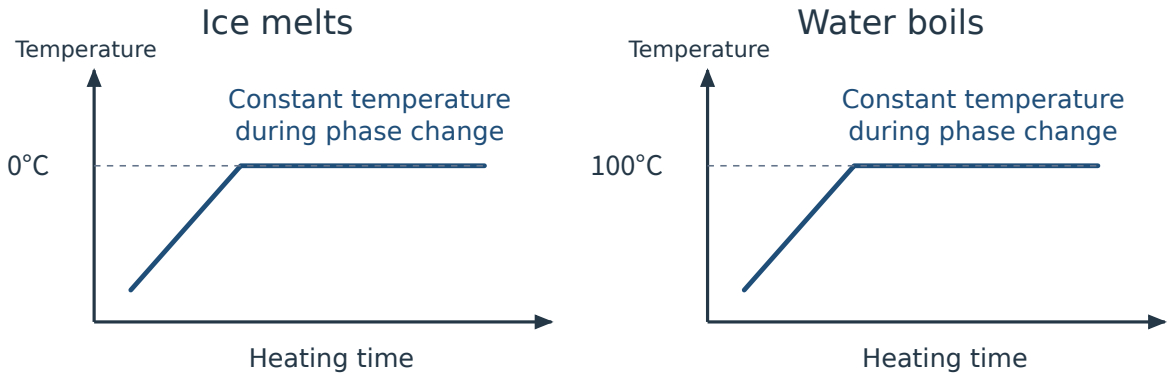
Anhydrous reagents give the correct starting colours; a positive test shows water is present, not that it is pure.

Grey outlines make the white powder visible. Each row compares the same reagent before and after contact with a water-containing sample. Do not interchange their starting colours.

If a drop of sample turns white anhydrous $CuSO_4$ blue, water is present. With blue anhydrous $CoCl_2$, a positive result is pink. For test and observation, name the reagent and its corresponding colour change. "Changes colour" is insufficient. Litmus and methyl orange indicate acidity or alkalinity; they do not replace these chemical tests for water.

To assess purity, measure a phase-change temperature. Melting changes solid to liquid, and melting point is the temperature at which this occurs. Boiling forms gas bubbles throughout a liquid, and boiling point is the boiling temperature. At ordinary atmospheric pressure, pure water melts at 0°C and boils at 100°C. A test description needs more than "100°C": heat the sample, measure its boiling point with a thermometer and compare with pure water's known value.

Testing purity by phase-change temperature: measure, then compare with pure water



Compare at normal atmospheric pressure; only the phase-change region is shown, not measured heating times.

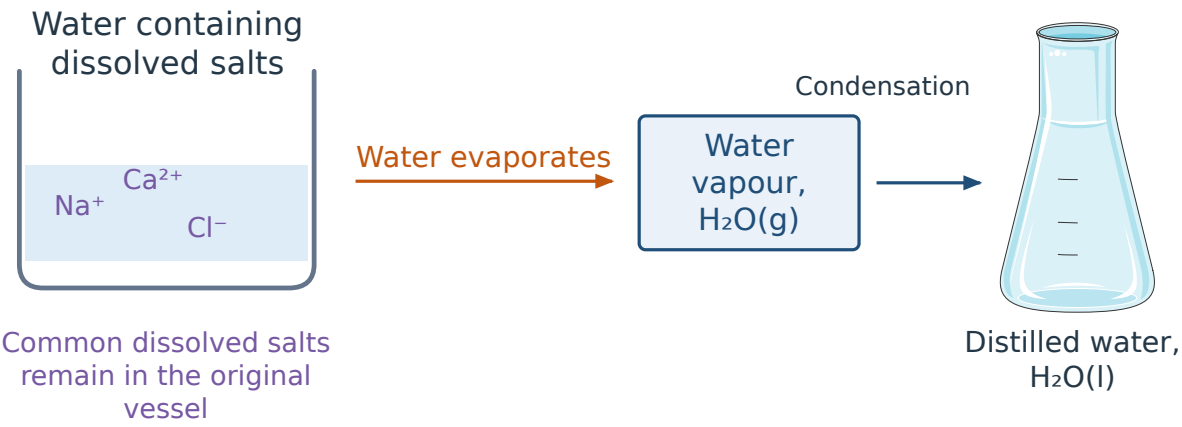
On the left, freeze the sample before warming it to measure melting point. On the right, heat the liquid to measure boiling point. Neither horizontal axis is numbered, so heating times are not compared.

If melting point is specified for an initially liquid sample, first cool it to form ice, then warm slowly and measure the melting temperature. Pure ice has a sharp melting point of 0°C. Impure samples may melt at a lower temperature or over a wider range. Compare at the same pressure: water's lower boiling point at high altitude should not automatically be attributed to impurities.

At ordinary atmospheric pressure, a water sample boiling appreciably above 100°C does not match pure water's reference boiling point. An ice sample melting gradually over a temperature range does not match a pure substance's sharp melting point. If the question asks only whether water is present, these measurements do not replace the specified chemical reagent. Matching temperatures is the course's purity criterion; it does not imply that the equipment detects every trace impurity.

Why is distilled water often used in experiments? Tap water usually contains dissolved minerals and other impurities, which may react, affect observations or alter measurements. Distilled water is produced by evaporating water and condensing the vapour and contains fewer such impurities. It is more suitable for preparing solutions and experiments requiring controlled composition.

Distillation: water leaves dissolved impurities, then condenses to a liquid



Process diagram: laboratory distilled water contains fewer dissolved impurities; tap water is not pure water.

The flask artwork is from Bioicons (Servier). Arrows show evaporation and condensation material flow, not connections in a distillation apparatus. Salt ions are illustrative examples.

The central reason for using distilled rather than tap water is fewer impurities. This reduces reactions or concentration changes caused by unknown minerals. Do not say tap water has no water molecules: both contain H_2O, and the difference is the impurities accompanying it.

Natural water contains more than H_2O: distinguish beneficial and harmful components. Dissolved oxygen, O_2, is generally beneficial because aquatic animals and other organisms use it for respiration. Some dissolved metal compounds supply essential minerals; others may be toxic. Neither contains metal nor natural automatically means beneficial.

Natural water may also contain plastics, sewage, harmful microorganisms, nitrates from fertilisers, and phosphates from fertilisers and detergents. Plastics harm aquatic life; sewage and harmful microorganisms can spread disease. Excess nitrates and phosphates can lower oxygen levels in water and damage aquatic organisms. This course does not require every biological step in that oxygen-depletion process.

Natural water: assess each component's benefits and risks

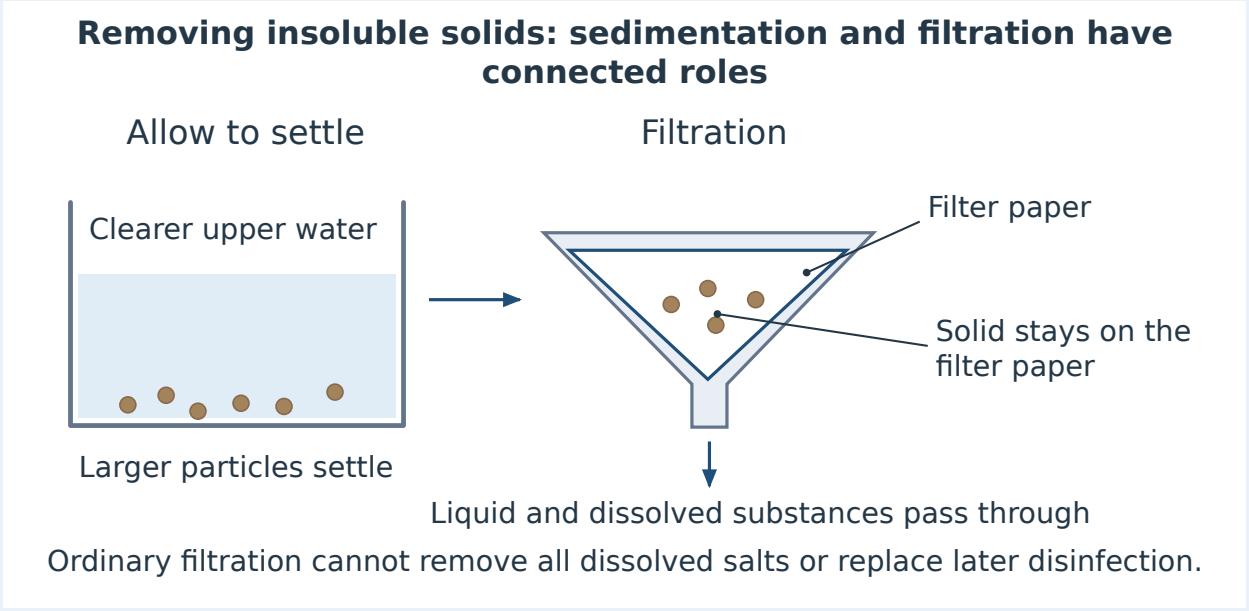
Component in natural water	Possible benefit or risk
Dissolved oxygen, O_2	Supports aquatic respiration; usually beneficial
Some metal compounds	Supply minerals needed for life
Other metal compounds	May be toxic; not all minerals are beneficial
Plastics, sewage and harmful microorganisms	Plastics harm aquatic life; sewage and pathogens spread disease
Excess nitrates and phosphates	Can lower dissolved oxygen and harm aquatic life
Common nutrient sources	Nitrates: fertilisers; phosphates: fertilisers and detergents

A dissolved substance is not automatically helpful or harmful; its identity and amount matter.

The table distinguishes generally beneficial from harmful components. Excess nutrients are linked to harm through reduced dissolved oxygen.

For example, among dissolved oxygen, excess nitrates and excess phosphates, dissolved oxygen itself is usually useful, while excessive nutrients can cause oxygen depletion. Clear-looking water may still contain dissolved substances and microorganisms. No visible solids does not establish safety or purity.

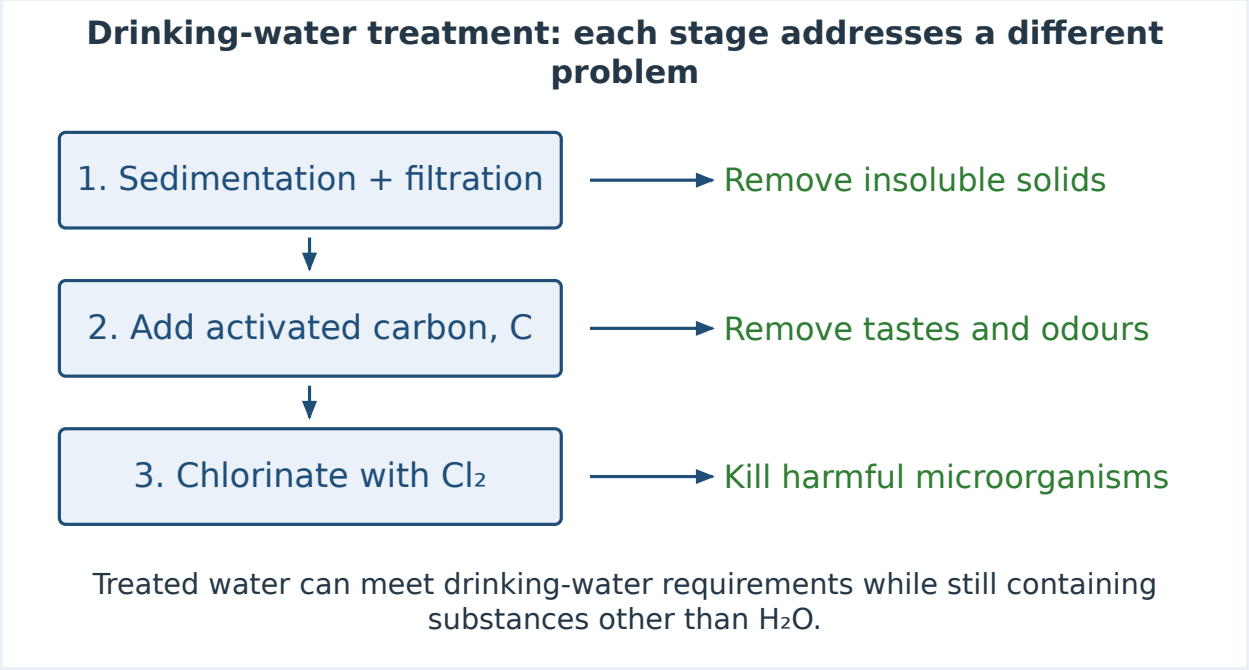
Domestic water treatment first removes solids, then addresses other problems. Sedimentation lets larger suspended particles settle. Filtration passes water through filter paper or a filtering layer that retains insoluble solids. Water and many dissolved substances pass through, so clear filtrate may still contain salts or microorganisms.



Brown dots show insoluble particles. Arrows indicate further treatment of the upper water and filtrate leaving, not solids passing through filter paper.

If water contains sand and dissolved salt, settling and filtering remove the sand, not all the salt. To remove solids corresponds to sedimentation or filtration; this does not yet answer to kill microbes.

Carbon and chlorine address taste and microorganisms separately. Activated carbon, mainly C, removes unwanted tastes and odours by adsorbing substances onto its surface. Chlorination adds chlorine, Cl_2, to kill harmful microorganisms and reduce pathogen transmission. Do not exchange these roles or turn killing microbes into removing all dissolved salts.



Vertical arrows show treatment order; horizontal arrows link each step to its purpose. Sedimentation and filtration are grouped as solid-removal stages here.

In a sequence of solid removal, carbon addition and a final treatment, the last box is chlorination. Do not repeat filtration if it is already present. For matching, connect carbon with tastes and odours, sedimentation with solids, and chlorination with microbes. Water treated adequately for domestic use still differs from chemically pure H_2O.

Identify what the question asks before choosing a method

Question	Main method	Supported conclusion
Does it contain water?	Anhydrous $CuSO_4$ or $CoCl_2$	Expected colour change shows water is present
Is it pure?	Measure melting or boiling point; compare with pure water	Does the phase-change temperature match?
How is drinking water treated?	Sedimentation, filtration, activated carbon and chlorine	Remove solids, remove tastes and odours, and disinfect

The same sample may contain water, retain impurities and have been disinfected. These three conclusions answer different questions and do not replace one another.

A disinfected salt solution turns anhydrous $CuSO_4$ blue, showing water is present; it does not show that salt has been removed. If asked about purity, measure melting or boiling point instead. If asked why carbon is added during treatment, explain taste and odour removal rather than a boiling-point test. Reading the requested conclusion first keeps the method appropriate.

Anhydrous $CuSO_4$ changes from _____ to _____ with water; anhydrous $CoCl_2$ changes from _____ to _____. At ordinary atmospheric pressure, pure water melts at _____ and boils at _____. Carbon removes _____ during treatment; chlorine _____.

Answers: white; blue; blue; pink; 0°C; 100°C; tastes and odours; kills harmful microorganisms

Recognising the Question Identify presence, pure or treatment. Water detection needs an anhydrous reagent and colour change; purity needs a measurement and reference temperature. Link sedimentation/filtration to solids, carbon to tastes and odours, and chlorine to microorganisms.

Worked Example 1 (2025·May/June 0620/31 Q8(a)) · Level 1

This question is about water and air.

(a) River water needs to be treated to make it safe to drink.

Two of the stages used in the treatment of domestic water are the addition of carbon and chlorination.

Describe the reason for each of these stages.

addition of carbon

chlorination

Solution

The two stages are addition of carbon and chlorination. Give each its own reason rather than merely saying both purify water.

Carbon is added **to remove tastes and odours**. It removes the responsible impurities and is not being used here as a disinfectant.

Chlorination is used **to kill microbes**. Write each purpose on its corresponding line: carbon—remove tastes and odours; chlorination—kill microbes. Each treatment then has its specific function.

Worked Example 2 (2025·May/June 0620/31 Q8(c)) · Level 1

This question is about water and air.

(c) Explain why distilled water is used in practical chemistry rather than tap water.

Solution

Compare distilled and tap water for experimental use. Both contain water, so "distilled water contains H_2O but tap water does not" is wrong.

Distilled water contains fewer impurities. Tap water generally retains dissolved minerals that may react or alter solution composition, making experimental conditions harder to control.

The core answer is: **Distilled water contains fewer impurities than tap water.** A full distillation apparatus description is not required; focus on the relevant property.

Worked Example 3 (2025·May/June 0620/31 Q8(b)) · Level 2

This question is about water and air.

(b) Describe how to test whether a sample of water is pure using boiling point.

Solution

The question specifies using boiling point, so describe measurement and the criterion. "100°C" alone does not explain how to test the sample.

First heat the sample to boiling and measure its boiling point with a thermometer. Then compare it with pure water at the same pressure. At ordinary atmospheric pressure, pure water boils at 100°C. A substantially different value does not match pure water's boiling-point behaviour.

A complete answer is: **Heat the sample until it boils and measure its boiling point. Compare this with 100°C, the boiling point of pure water at ordinary atmospheric pressure.** The first sentence gives the method and measurement; the second gives the reference and purity criterion. Anhydrous $CuSO_4$ instead detects water and does not perform the specified purity test.

Worked Example 4 (2025·May/June 0620/22 Q29) · Level 2

Water is added separately to anhydrous copper(II) sulfate and to anhydrous cobalt(II) chloride.

Which row shows the colour changes?

	anhydrous copper(II) sulfate	anhydrous cobalt(II) chloride
A	blue to white	blue to pink
B	blue to white	pink to blue
C	white to blue	blue to pink
D	white to blue	pink to blue

Solution

Water is added separately to two anhydrous reagents. Find a row with both colour changes correct.

The first is anhydrous copper(II) sulfate, $CuSO_4$: white to blue. A and B reverse this as blue to white, so eliminate them.

The second is anhydrous cobalt(II) chloride, $CoCl_2$: blue to pink. Of C and D, only C matches. Check both columns and choose **C**.

Blue occurs twice with different roles: after water reaches $CuSO_4$, and before water reaches anhydrous $CoCl_2$. Pair each colour with its reagent.

Practice

1. (2025·May/June 0620/11 Q31) · Level 1

Which substances can be used to detect the presence of water?

1 anhydrous cobalt(II) chloride

2 anhydrous copper(II) sulfate

3 litmus

4 methyl orange

A 1 and 2 **B** 1 and 3 **C** 2 and 4 **D** 3 and 4

2. (2025·May/June 0620/32 Q8(a)) · Level 2

This question is about water and air.

(a) Water needs to be treated to make it safe to drink.

Draw a line from each stage in the treatment of domestic water to the reason why it is carried out.

stage	reason
addition of carbon	to kill microbes
sedimentation	to remove tastes and odours
chlorination	to remove solids

3. (2025·May/June 0620/33 Q7(a)) · Level 2

This question is about air and water.

(a) Water needs to be treated to make it safe to drink.

Complete Fig. 7.1 to show the three stages needed in the treatment of domestic water. The second stage has been completed.

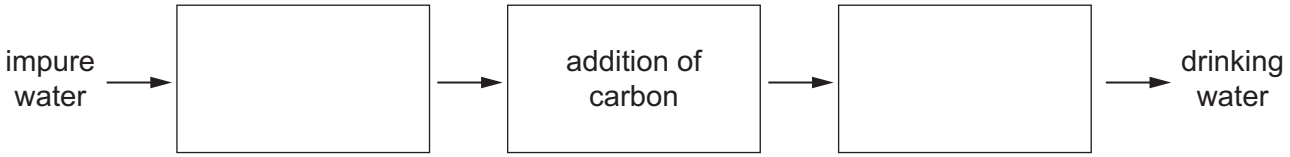


Fig. 7.1

4. (2025·May/June 0620/32 Q8(b)) · Level 2

This question is about water and air.

(b) Describe how to test whether a sample of water is pure using melting point.

5. (2025·May/June 0620/33 Q7(b)) · Level 2

This question is about air and water.

(b) Describe a chemical test for the presence of water.

test

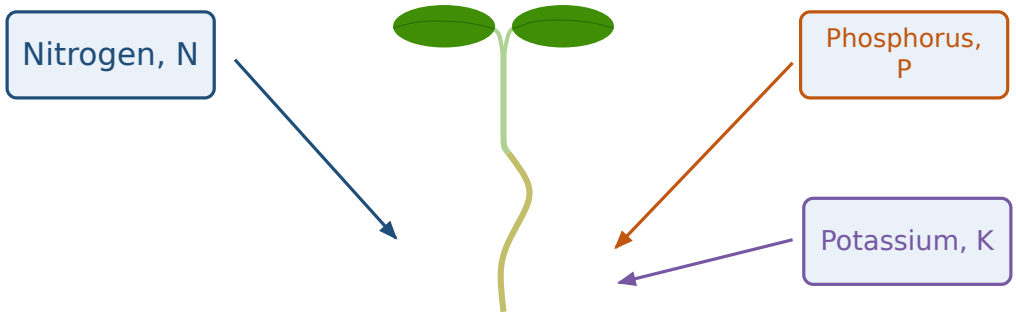
observations

10.2 Fertilisers

Type 1: Identifying NPK from names and formulae and explaining effects on water

Key Ideas **Fertilisers replenish elements, not just a set of letters.** Harvesting crops removes some nutrients from soil. Fertilisers replace nutrients needed for growth. NPK means nitrogen, N, phosphorus, P, and potassium, K. K is potassium, not calcium, Ca.

NPK names three elements; plants obtain them in nutrient compounds



Arrows show absorbable nutrients containing N, P and K reaching roots, not application of the pure elements.

The plant outline is from Bioicons. N, P and K are symbols of nutrient elements; the diagram gives neither concentration nor fertiliser dose.

Ammonium salts and nitrates can be fertilisers. Both ammonium, NH_4^+, and nitrate, NO_3^-, contain nitrogen. Common fertilisers do not simply place nitrogen gas, elemental phosphorus or potassium metal beside roots. Read the element symbols in a fertiliser's name or formula to identify what it supplies.

Potassium nitrate, KNO_3, supplies K and N but no P. Ammonium nitrate, NH_4NO_3, contains N in both parts but does not thereby supply P or K. When asked which fertiliser nutrient is an alkali metal, N and P are non-metals; the answer is potassium, K.

To check for N, P and K, read each element in the formula

Substance	Formula	N / P / K
Potassium nitrate	KNO_3	Yes / No / Yes
Ammonium nitrate	NH_4NO_3	Yes / No / No
Ammonium phosphate	$(NH_4)_3PO_4$	Yes / Yes / No
Potassium chloride	KCl	No / No / Yes
Calcium phosphate	$Ca_3(PO_4)_2$	No / Yes / No
Potassium sulfate	K_2SO_4	No / No / Yes
Ammonium chloride	NH_4Cl	Yes / No / No
Potassium phosphate	K_3PO_4	No / Yes / Yes

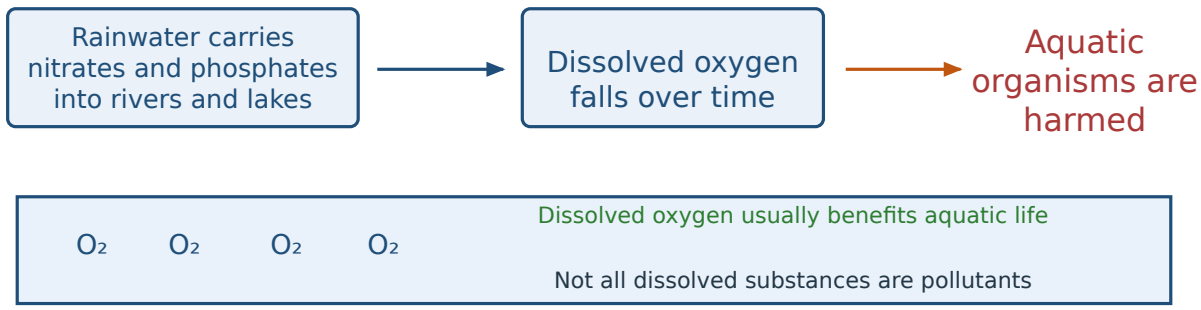
The table checks element presence, not nutrient quantities. Two N atoms still only establish nitrogen content; they cannot replace missing P or K.

Combine the elements supplied by two fertilisers. List which of N, P and K occurs in the first, add those from the second and check whether all three are present. “Provides all three elements” does not ask for mass percentages. Repeated nitrogen cannot make up for absent potassium.

Mixing KNO_3 with $Ca_3(PO_4)_2$ gives K and N from the first and P from the second: all three. NH_4NO_3 with K_2SO_4 gives N and K but still lacks P. S in sulfate is sulfur; a complex-looking formula does not automatically supply every nutrient.

Useful nutrients can harm rivers and lakes in excess. Dissolved oxygen generally benefits aquatic organisms, and some metal compounds provide essential minerals. However, excess nitrates and phosphates can reduce oxygen levels and harm aquatic life. Nitrates may come from fertilisers; phosphates from fertilisers and detergents. Learn the syllabus-required oxygen-depletion result without needing every extra stage of eutrophication.

Nutrients that help plants can harm water ecosystems when present in excess



Arrows represent several environmental steps, not a direct chemical reaction between nitrate and oxygen.

The diagram explains the required oxygen-depletion result. Arrows summarise a process over time, not an invented single-step chemical reaction.

If asked to identify usually harmful components from dissolved oxygen, nitrates and phosphates, retain oxygen's beneficial role and recognise the depletion risk from excess nutrients. Being dissolved does not make all three pollutants.

NPK means _____, _____ and _____. A mixture of NH_4NO_3 and KCl still lacks _____. Excess nitrates and phosphates can reduce _____ in water and harm aquatic life.

Answers: nitrogen; phosphorus; potassium; phosphorus, P; dissolved oxygen

Recognising the Question Combine the N, P and K supplied by both fertilisers and check for a missing element. In natural water, distinguish dissolved oxygen's benefit from oxygen depletion caused by excess nitrates and phosphates.

Worked Example 1 (2025·May/June 0620/21 Q30) · Level 2

Nitrogen, phosphorus and potassium are essential elements for improved plant growth.

Which mixture provides all **three** essential elements?

	mixture	formulae
A	ammonium phosphate + potassium chloride	$(NH_4)_3PO_4$ + KCl
B	ammonium phosphate + ammonium nitrate	$(NH_4)_3PO_4$ + NH_4NO_3
C	ammonium phosphate + ammonium chloride	$(NH_4)_3PO_4$ + NH_4Cl
D	ammonium nitrate + potassium chloride	NH_4NO_3 + KCl

Solution

All N, P and K are required. Read both formulae in each row and combine their target elements.

A: $(NH_4)_3PO_4$ contains N and P; KCl supplies K, so all three are present. B supplies only N and P and lacks K. C also lacks K. D supplies N from NH_4NO_3 and K from KCl but no P.

Choose A. This tests element presence, not mass fractions. Repeated N does not supply absent K, and O cannot replace P.

Worked Example 2 (2025·May/June 0620/13 Q30) · Level 2

Water from natural sources can contain many dissolved substances.

Which substances are often harmful to aquatic life?

1 dissolved oxygen

2 nitrates

3 phosphates

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

Solution

Distinguish beneficial oxygen from harmful excess nutrients. Dissolved oxygen in 1 supports aquatic life and is generally beneficial, so it does not belong in the usually harmful combination.

Excess nitrates, 2, and phosphates, 3, can lower oxygen levels and harm aquatic organisms. Both qualify. Choose 2 and 3 only, option D.

Assess each component's effect rather than assuming every dissolved substance is harmful.

Practice

1. (2025·May/June 0620/13 Q32) · Level 1

Farmers use fertilisers to replace elements in the soil that have been removed by the crops they grow.

Which elements in the soil are replaced by adding fertilisers?

A Ca, P, O **B** K, O, S **C** N, O, S **D** N, K, P

2. (2025·May/June 0620/42 Q5(b)(v)) · Level 1

Alkali metals are reactive elements.

(b) Identify the alkali metal which:

(v) is found in fertilisers to improve plant growth.

3. (2025·May/June 0620/11 Q30) · Level 2

Fertilisers are used to provide the three elements needed for improved plant growth.

Which two compounds would provide **all three** of these elements?

A ammonium nitrate and calcium phosphate

B ammonium nitrate and potassium sulfate

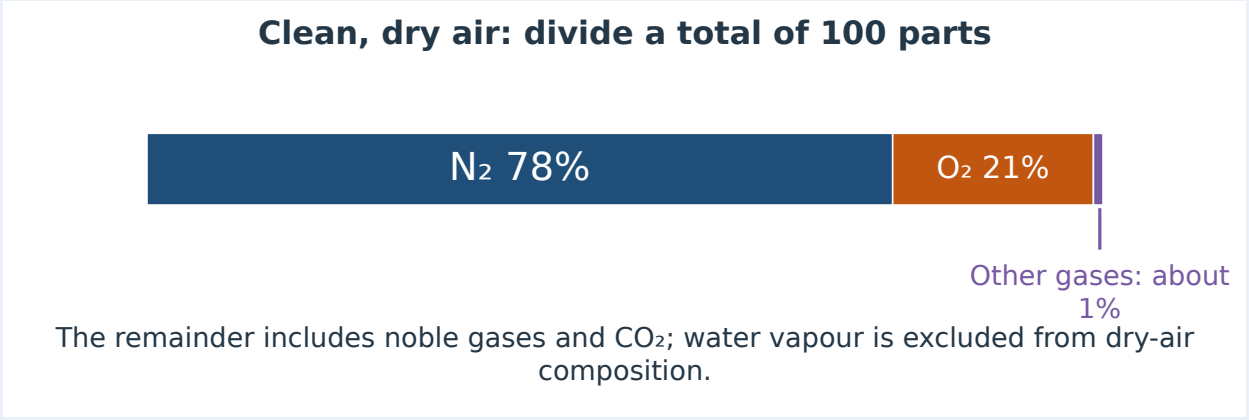
C potassium nitrate and calcium phosphate

D potassium phosphate and potassium sulfate

10.3 Air and climate

Type 1: Using gas composition and volume changes to determine air proportions

Key Ideas **Air is a mixture: identify what the percentages measure.** By volume, clean, dry air contains approximately 78% nitrogen, N_2, 21% oxygen, O_2, and 1% other gases, mainly noble gases and carbon dioxide. Dry excludes water vapour from these proportions. Clean excludes smoke particles and additional pollutants from this standard composition.



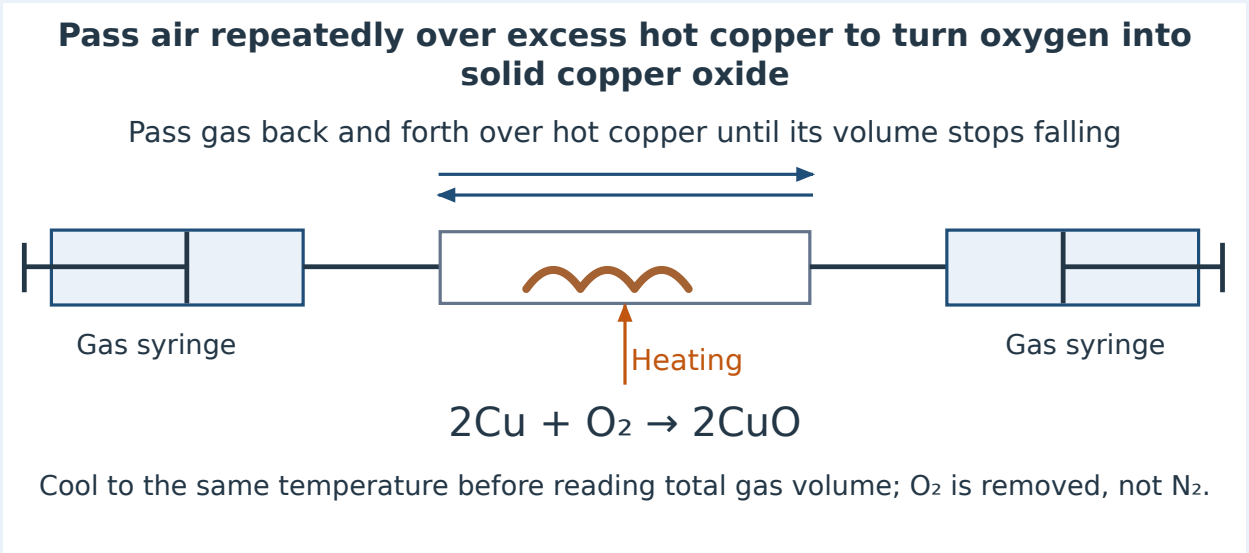
Bar lengths follow 78:21:1. The 1% is not all CO2. The approximation of oxygen as one-fifth cannot replace 21% in every question.

Thus 100 cm³ clean, dry air contains about 78 cm³ N_2, 21 cm³ O_2 and 1 cm³ other gases. Only N_2 exceeds 30%; 21% does not. Removing all O_2 leaves 79 cm³ containing both N_2 and other gases. The entire 79% must not be labelled nitrogen.

How is the oxygen consumed from air measured? Pass an air sample repeatedly over excess heated copper. Copper reacts with oxygen to form solid copper(II) oxide, usually changing from red-brown to black:

$$2Cu + O_2 \rightarrow 2CuO$$

Oxygen enters the solid product and no longer occupies gas-syringe volume. Nitrogen and most other gases remain.



This unscaled apparatus diagram illustrates the principle; numerical volumes come from the text. Keep the apparatus airtight, use excess copper and compare readings at the same temperature and pressure.

Pass the gas repeatedly until total volume stops falling, allowing O_2 to react fully. Cool back to the initial temperature before reading at the same pressure, avoiding thermal expansion effects. Airtight apparatus and enough copper are necessary to attribute volume loss to oxygen consumption.

Record the initial total volume and the final total after reaction and cooling. The decrease represents oxygen consumed:

$$O_2 \text{ volume consumed} = \text{initial volume} - \text{final volume}$$

For its percentage in air, divide by the original total air volume:

$$\text{Oxygen percentage} = \frac{O_2 \text{ volume consumed}}{\text{initial air volume}} \times 100\%$$

The denominator is the original air, not the remaining gas.

For example, 150 cm³ air leaves 118.5 cm³ after complete reaction and cooling. Subtract the remainder from the initial volume:

$$150 - 118.5 = 31.5 \quad \text{cm}^3$$

Compare that 31.5 cm³ with the original 150 cm³:

$$\text{Oxygen percentage} = \frac{31.5}{150} \times 100\% = 21\%$$

If asked for the final reading, give 118.5. If asked for oxygen consumed, give 31.5. Both come from the same experiment but answer different questions.

How much air supplies a known oxygen requirement? As a decimal, 21% is 0.21. Oxygen is part of the air, so:

$$\text{Oxygen volume} = \text{total air volume} \times 0.21$$

For a known oxygen requirement, rearrange:

$$\text{Total air volume} = \frac{\text{required oxygen volume}}{0.21}$$

Use the same volume units and temperature and pressure conditions for the gases compared.

Air to oxygen: multiply by 0.21; oxygen to air: divide by 0.21



For 42 dm³ of O_2, minimum air volume = 42 ÷ 0.21 = 200 dm³.

Check: 200 × 0.21 = 42; the air volume must exceed its oxygen volume.

Both sides use the same volume units and temperature and pressure. Reversing the calculation requires division, not another multiplication by 21%.

For example, if 42 dm³ oxygen is required and all oxygen can be extracted and used, the minimum air volume is:

$$\frac{42}{0.21} = 200 \quad \text{dm}^3$$

Check that 21% of this air supplies the required oxygen:

$$200 \times 0.21 = 42 \quad \text{dm}^3$$

Multiplying 42 by 0.21 instead gives 8.82 dm³, making total air smaller than the oxygen it contains and revealing the reversed calculation. Minimum assumes complete oxygen recovery and use; losses require more air.

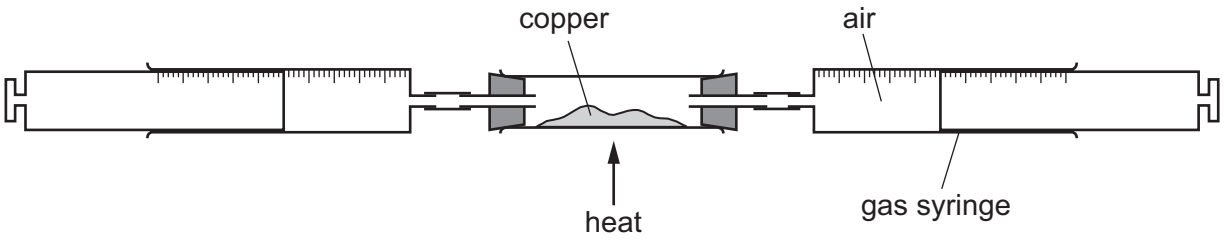
Clean, dry air contains about _____% N_2 and _____% O_2. In the heated-copper experiment, oxygen percentage = (_____ − _____) ÷ _____ × 100%. To find total air from a known O_2 volume, _____ by 0.21.

Answers: 78; 21; initial total volume; final total volume; initial total volume; divide

Recognising the Question Identify whether the target is oxygen, remaining gas or original air. Heated-copper experiments use the volume difference, with original volume as the percentage denominator. To infer air from required oxygen, divide by 0.21 and check the scale.

Worked Example 1 (2025·May/June 0620/12 Q30) · Level 2

In the experiment shown, 100 cm^3 of clean, dry air is passed backwards and forwards through a tube containing heated copper until there is no further decrease in the volume of the air.



All of the gas is pushed into **one** syringe.

What is the final reading on this gas syringe?

A 0 cm^3 **B** 21 cm^3 **C** 79 cm^3 **D** 100 cm^3

Solution

Initially there is 100 cm^3 clean, dry air containing about 21% oxygen. Its oxygen volume is:

$$100 \times 0.21 = 21 \quad cm^3$$

Repeated passage over hot copper until the volume stops decreasing indicates complete oxygen consumption into solid CuO. Other gases remain. Pushing all gas into one syringe gives the total remaining volume:

$$100 - 21 = 79 \quad cm^3$$

Choose C.

The 21 cm^3 is removed oxygen, not the final reading. The remaining 79 cm^3 is not entirely nitrogen; it also includes noble gases and others. Make this comparison under temperature and pressure conditions comparable with the initial reading.

Practice

1. (2025·May/June 0620/13 Q31) · Level 1

Which gas is over 30% of clean, dry air?

A argon

B carbon dioxide

C nitrogen

D oxygen

2. (2025·May/June 0620/43 Q1(i)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(i) is a gas that is approximately 78% of clean, dry air.

3. (2025·May/June 0620/42 Q1(a)) · Level 1

Using numbers only, state the:

(a) percentage of nitrogen in clean, dry air

4. (2025·May/June 0620/41 Q1(i)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(i) is a gas that is approximately 21% of clean, dry air

5. (2025·May/June 0620/11 Q12) · Level 2

A hydrogen–oxygen fuel cell uses 630 dm^3 of oxygen.

The oxygen for the reaction is extracted from clean, dry air.

What is the minimum volume of clean, dry air needed to provide this volume of oxygen?

A 788 dm^3 **B** 808 dm^3 **C** 3000 dm^3 **D** 3316 dm^3

Type 2: Matching pollutant sources, effects and treatment reactions

Key Ideas

For carbon-containing fuels, check oxygen supply first. With sufficient oxygen, fuels such as methane, CH_4, burn completely to form CO_2. Hydrogen-containing fuels also form water. Methane's complete combustion is:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Insufficient oxygen can cause incomplete combustion, producing toxic carbon monoxide, CO, and possibly carbon particles and other particulates.

Products of carbon-fuel combustion depend on oxygen supply

Conditions	Representative methane reaction
Complete combustion: sufficient oxygen	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
Incomplete combustion: CO may form	$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$
Incomplete combustion: carbon particles may also form	$CH_4 + O_2 \rightarrow C + 2H_2O$

Actual incomplete combustion can give mixed products; each row shows one atom-balanced case.

The two incomplete-combustion equations show possible pathways, not a fixed product ratio from one sample.

Why can incomplete combustion of methane not produce SO_2? CH_4 contains C and H, while oxygen supplies O. There is no S in the reactants, so SO_2 cannot appear. CO can form. SO_2 requires a sulfur source, such as explicitly stated sulfur compounds in the fuel. Check element conservation before applying combustion conditions.

Keep pollutant sources and effects matched. CO_2 is associated with complete combustion of carbon fuels; CO and particulates with incomplete combustion. CO is toxic. Particulates increase the risk of respiratory problems and cancer. Methane, CH_4, can come from vegetation decomposing without oxygen and waste gases from animal digestion. Increased CO_2 and CH_4 strengthen global warming and affect climate.

Oxides of nitrogen, often collectively NO_x, can cause acid rain, photochemical smog and respiratory problems. Burning fossil fuels containing sulfur compounds produces SO_2, which also contributes to acid rain. NO_x describes a group of nitrogen oxides, not one fixed molecular formula.

Pollutants: connect each source with its effects

Pollutant	Main source	Effects required in this course
CO_2	Complete combustion of carbon-containing fuels	Increases global warming
CO	Incomplete combustion of carbon-containing fuels	Toxic
Particulates	Incomplete combustion of carbon-containing fuels	Respiratory problems and cancer risk
CH_4	Decomposing vegetation and animal digestion	Increases global warming
Nitrogen oxides, NO_x	High-temperature reactions in vehicle engines	Acid rain, photochemical smog and respiratory problems
SO_2	Combustion of sulfur-containing fossil fuels	Acid rain

NO_x covers several nitrogen oxides; it is not the formula of one fixed molecule.

The table follows the syllabus scope. Do not interchange CO toxicity, SO2/NOx acid-rain effects and CO2/CH4 greenhouse effects.

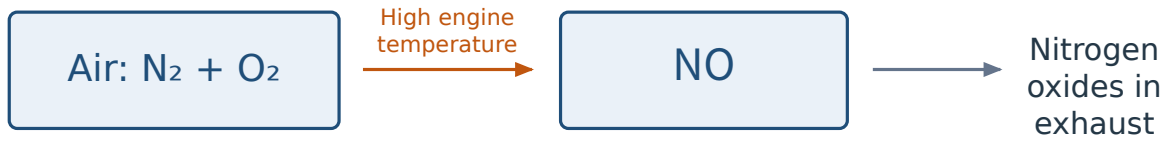
From CO, SO_2 and NO_x, select SO_2 and NO_x for acid rain. CO is toxic, but harmful does not automatically make it a required acid-rain source. Conversely, global warming points to CO_2 and CH_4. Acid rain, toxicity and the greenhouse effect are different outcomes.

Where do engine nitrogen oxides come from? Air entering a car engine contains N_2 and O_2. High engine temperatures cause them to react and form nitrogen oxides. For NO:

$$N_2 + O_2 \rightarrow 2NO$$

Both the air source of nitrogen and oxygen and the high engine temperature are essential. Do not arbitrarily substitute petrol as the nitrogen source.

At high engine temperatures, N_2 and O_2 from air can react



$$N_2 + O_2 \rightarrow 2NO$$

Nitrogen and oxygen come from intake air; the catalytic converter treats exhaust and is not where this formation occurs.

Arrows connect the reactant source, formation condition and product. NO is shown as one nitrogen-oxide example.

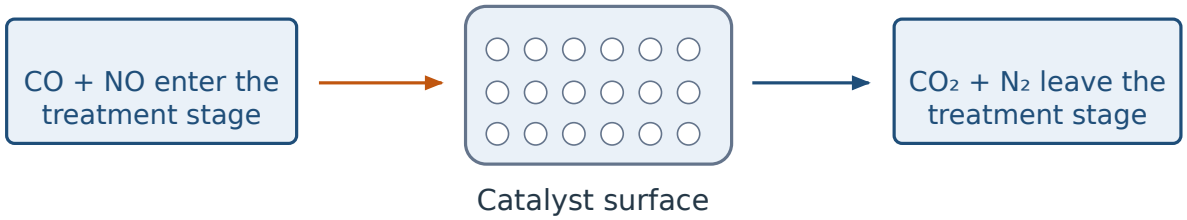
A statement that N_2 and O_2 react at high engine temperature correctly links source, conditions and result. A statement that NO mainly forms in the catalytic converter misidentifies the treatment device as the formation site. The converter treats exhaust to reduce these pollutants.

For treatment, check what the pollutant becomes. A catalytic converter causes exhaust CO and NO to react:

$$2CO + 2NO \rightarrow 2CO_2 + N_2$$

CO gains oxygen and is oxidised to CO_2; NO loses oxygen and is reduced to N_2. Both toxic CO and nitrogen oxides decrease. The CO_2 produced still has a greenhouse effect, so the exhaust is not free of environmental impact.

Catalytic converter: CO reacts with NO to form CO_2 and N_2



$$2CO + 2NO \rightarrow 2CO_2 + N_2$$

CO is oxidised and NO is reduced; CO_2 remains a greenhouse gas.

The honeycomb illustrates catalyst contact surface, not pore size or a specific metal formulation.

Compare another equation:

$$2NO + O_2 \rightarrow 2NO_2$$

It conserves atoms but merely converts one nitrogen oxide to another. It therefore does not answer the required removal of nitrogen oxides. Check more than balance: NO must become a non-nitrogen-oxide product, here N_2 on the right.

SO_2 can be reduced at its source or in flue gases. Low-sulfur fuels reduce the available sulfur. Basic CaO can also react with acidic SO_2 in exhaust gases and remove it: flue gas desulfurisation. Match the measure to the pollutant: catalytic converters for vehicle nitrogen oxides; low-sulfur fuel and CaO treatment for SO_2.

Reducing acid rain: treat nitrogen oxides and sulfur dioxide separately

Pollutant	Measure	Reason
Nitrogen oxides	Use vehicle catalytic converters	Convert NO and related oxides to N_2
Sulfur dioxide	Use low-sulfur fuels	Reduce the sulfur available for oxidation
Sulfur dioxide	Pass flue gas over calcium oxide	CaO reacts with acidic SO_2 to remove it

CaO is a basic oxide; this resembles its removal of acidic SiO_2 in the blast furnace.

Each control measure addresses a pollutant source or reaction property. The catalytic converter is not presented as removing every greenhouse gas.

For SO_2-containing factory gases, proposing contact with CaO needs the explanation that basic CaO reacts with acidic SO_2 and lowers its emissions. “Add a chemical” omits both identity and action; “plant trees” does not directly explain this acid-rain control.

Incomplete combustion of methane can produce toxic _____. The two acid-rain pollutant groups are _____ and _____. Engine NO forms when _____ and _____ from air react at high temperature. A catalytic converter changes NO into _____.

Answers: CO; SO_2; NO_x; N_2; O_2; N_2

Recognising the Question For sources, check whether the reactants contain the required element, then assess combustion conditions. Separate toxicity, acid rain and greenhouse effects. A treatment equation must conserve atoms and actually convert the target pollutant into the required product.

Worked Example 1 (2025·May/June 0620/21 Q32) · Level 2

The exhaust gases from cars contain oxides of nitrogen.

How are these oxides of nitrogen formed?

A Nitrogen and oxygen from the air react together at the high temperatures in the engine.

B Nitrogen and oxygen from the petrol react together in the car exhaust.

C Nitrogen from the petrol reacts with oxygen at the high temperatures in the engine.

D Nitrogen reacts with oxygen from the air in the catalytic converter.

Solution

Check the source: air entering the engine already contains N_2 and O_2. Check the conditions: high engine temperature makes them react to form nitrogen oxides. A satisfies both, so choose A.

B and C replace the nitrogen source with petrol, contrary to this mechanism. D places formation in the catalytic converter, which treats and removes nitrogen oxides from exhaust rather than being their main formation site.

NO formation can be written:

$$N_2 + O_2 \rightarrow 2NO$$

Retain both key facts: air and high engine temperature.

Worked Example 2 (2025·May/June 0620/22 Q30) · Level 2

Gases which cause acid rain form in car engines.

These gases can be removed using a catalytic converter.

Which equation represents the removal of one of these gases using a catalytic converter?

A $CO_2 + C \rightarrow 2CO$

B $2CO + 2NO \rightarrow 2CO_2 + N_2$

C $N_2 + 2O_2 \rightarrow 2NO_2$

D $2NO + O_2 \rightarrow 2NO_2$

Solution

Select the catalytic reaction removing an acid-rain-related gas. NO in vehicle exhaust should become N_2, while CO can accept its oxygen.
B gives:

$$2CO + 2NO \rightarrow 2CO_2 + N_2$$

NO is consumed and becomes N_2 while CO becomes CO_2, meeting the requirement. Choose B. Check atoms: both sides have 2 C, 2 N and 4 O.

A does not remove NO and instead produces toxic CO. C forms NO_2 from N_2 and O_2. D converts NO to NO_2, retaining a nitrogen oxide and failing the removal purpose. CO_2 still has a possible greenhouse impact, but B nevertheless lowers NO emissions.

Practice

1. (2025·May/June 0620/11 Q32) · Level 1

What is produced by the incomplete combustion of methane?

A carbon monoxide

B hydrogen

C lead compounds

D sulfur dioxide

2. (2025·May/June 0620/41 Q1(f)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(f) causes acid rain

3. (2025·May/June 0620/12 Q31) · Level 2

Sulfur dioxide, carbon monoxide and oxides of nitrogen are gaseous pollutants found in the air.

Which pollutants contribute to acid rain?

A carbon monoxide and sulfur dioxide

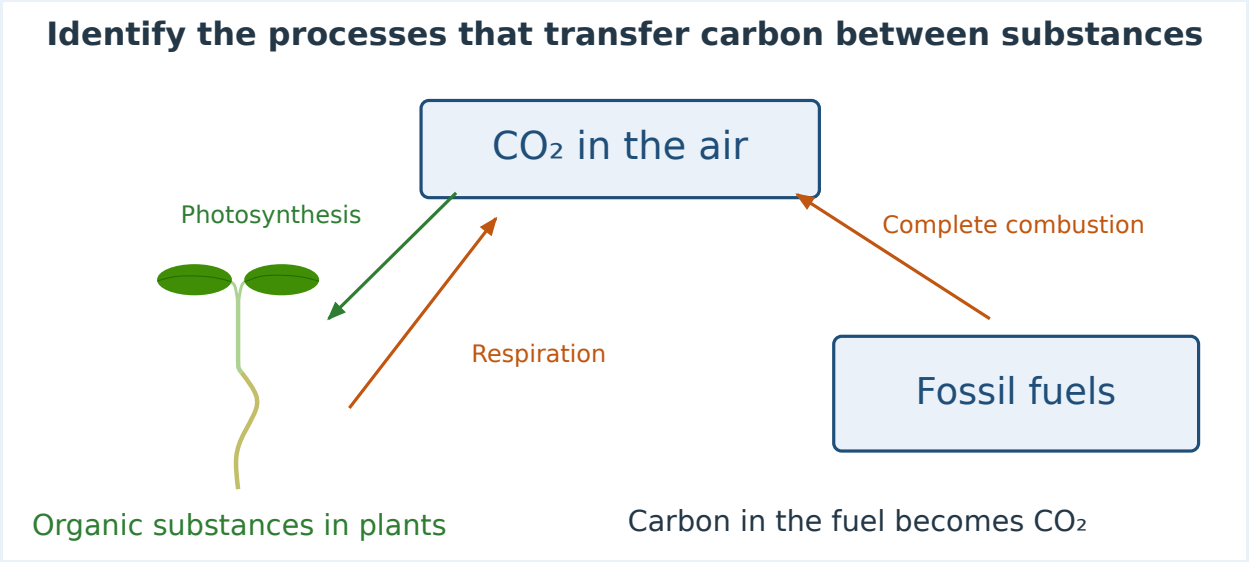
B oxides of nitrogen and sulfur dioxide

C oxides of nitrogen only

D sulfur dioxide only

Type 3: Explaining photosynthesis and climate change through carbon transfer and heat radiation

Key Ideas **Connect carbon's sources and destinations.** Complete combustion transfers carbon from fuel into atmospheric CO_2. Photosynthesis consumes atmospheric CO_2 and incorporates its carbon into glucose and other organic substances. Plant and animal respiration also releases CO_2; plants do not only absorb CO_2 without ever releasing any.

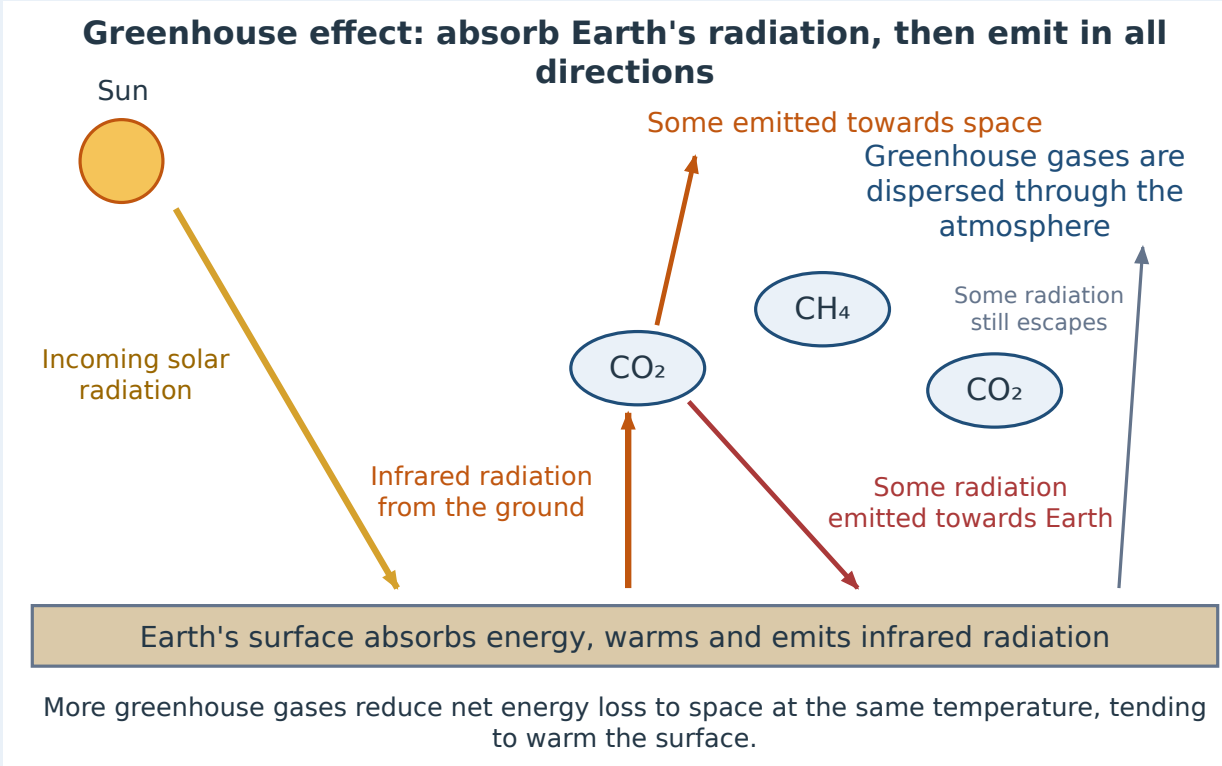


Only the three carbon-transfer processes needed here are shown. Plants also respire. The diagram does not suggest that fossil fuels form again over a short timescale.

To explain increased atmospheric CO_2 from fossil fuels, state that combustion converts fuel carbon into CO_2 and releases it. To explain how growing trees help, state that photosynthesis absorbs CO_2. These arrows transfer carbon in opposite directions.

Methane, CH_4, is another important greenhouse gas. It can form during vegetation decomposition without oxygen and from animal digestive waste gases. When a question names decomposition of vegetation, recognise the methane source. Not every change involving carbon is combustion or produces only CO_2.

How does Earth gain and lose energy? Earth's surface absorbs solar radiation, warms and emits infrared radiation, often described in questions as thermal energy from the Earth. Atmospheric CO_2 and CH_4 absorb some outgoing infrared radiation and emit energy in all directions. Some travels towards the surface and lower atmosphere, and some towards space.



Arrows show energy-transfer directions, not quantitative proportions. Greenhouse gases are dispersed particles, not a visible mirror or opaque shell.

With other conditions unchanged and surface temperature initially the same, increasing greenhouse gases reduces net energy loss to space. The surface then tends to warm. Sustained global warming can cause climate change. Distinguish the natural greenhouse effect from its enhancement by human-caused gas increases; the focus here is higher concentrations strengthening an existing warming effect.

"More CO_2 makes Earth lose energy to space faster" reverses the direction: energy loss decreases. "Only CO_2 causes global warming" is also wrong because CH_4 contributes too.

Older questions sometimes describe returning energy as reflects or reflected back to Earth. Retain the relationship they test: some terrestrial heat radiation returns and energy loss decreases. For the mechanism, use **absorption and re-emission**, not gas molecules acting as mirrors. Do not replace radiation from Earth with all incoming solar radiation. NASA's "Global Warming" explains this distinction in its discussion of the natural greenhouse effect.

Connect measures to a gas source or removal process. Planting trees and protecting vegetation allow growing plants to absorb CO_2 by photosynthesis. Using less fossil fuel and more wind and solar energy reduces combustion emissions. Reducing livestock-related emissions lowers CH_4 from animal digestion. Hydrogen fuel produces water in use, but total emissions also depend on how the hydrogen is made.

Reducing greenhouse gases: explain what each measure reduces

Measure	Explanation
Plant trees and protect vegetation	Growing plants absorb CO_2 by photosynthesis
Use less fossil fuel	Reduce CO_2 from burning carbon-containing fuels
Use more wind and solar energy	Generate electricity without burning carbon-containing fuels
Reduce livestock-related emissions appropriately	Reduce CH_4 emissions associated with animal digestion
Expand suitable uses of hydrogen energy	Using H_2 forms water; total emissions also depend on hydrogen production

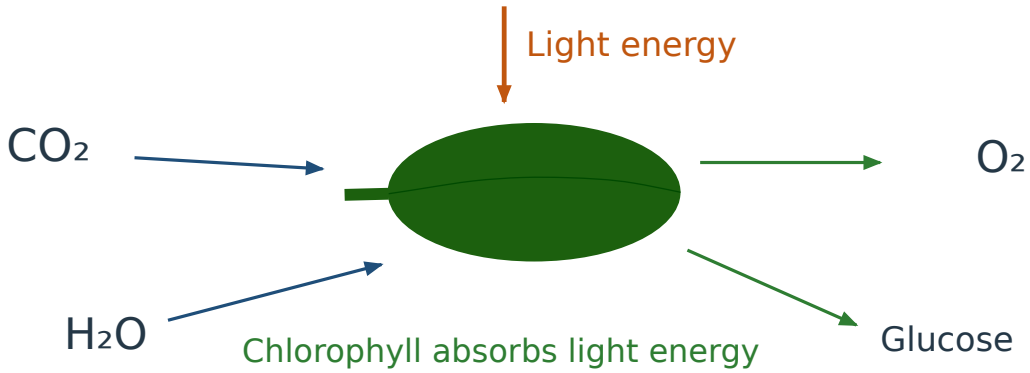
Anaerobic vegetation decay and animal digestion can both produce CH_4; CO_2 is not the only greenhouse gas.

The table explains where syllabus-listed measures act. It does not assume every hydrogen-production method emits no CO2 or give unsupported numerical emission reductions.

For example, reducing livestock emissions should be linked to the CH_4 source. Replacing coal-fired electricity with wind reduces CO_2 from burning carbon-containing fuel. Both can address climate change, but their reasons cannot be exchanged into an unsupported claim such as wind turbines absorbing methane.

Separate photosynthesis substances from conditions. Carbon dioxide and water are reactants; glucose and oxygen are products. Chlorophyll and light energy are required conditions. Light supplies energy, not carbon atoms as a chemical reactant.

Photosynthesis uses CO_2 and water as matter, and light as energy



Carbon dioxide + water → glucose + oxygen

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Blue arrows show reactants entering, green arrows products and orange energy input. Chlorophyll is a condition, not a reactant in the material-balance equation.

Write the word equation first:
carbon dioxide + water → glucose + oxygen

Then the balanced symbol equation:
$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
The conditions are chlorophyll and light. Check atoms: each side has 6 C, 12 H and 18 O.

A question describing CO_2 and water forming glucose and oxygen in chlorophyll and sunlight identifies photosynthesis through reactants, products and conditions. It is not combustion, since the energy input and substance changes differ, or displacement, since one element does not replace another in a compound.

The two greenhouse gases here are _____ and _____. They absorb _____ radiation from Earth and _____ net energy loss to space. Photosynthesis uses _____ and _____ to form _____ and _____ the energy comes from _____.

Answers: CO_2; CH_4; infrared; reduce; CO_2; water; glucose; O_2; light

Recognising the Question Separate substance and energy changes. Combustion, decomposition and photosynthesis determine carbon transfer. Greenhouse questions track radiation emitted by Earth and net loss to space; attend to Earth versus Sun and increase versus decrease in the original wording.

Worked Example 1 (2025·May/June 0620/21 Q31) · Level 2

Which statements help explain why an increasing amount of carbon dioxide or methane in the atmosphere causes global warming?

1 Carbon dioxide absorbs thermal energy from the Earth.

2 Methane does **not** absorb thermal energy from the Sun.

3 Carbon dioxide does **not** reflect thermal energy from the Sun to the Earth.

4 Methane reflects thermal energy from the Earth back to the Earth.

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

Solution

The question concerns heat leaving Earth's surface and energy returning. Statement 1 says CO_2 absorbs thermal radiation from Earth, matching the greenhouse effect. Statement 4 uses reflects for methane causing some terrestrial thermal energy to return. Under this question's wording and official answer, 1 and 4 give the required explanation. Choose B.

Understand statement 4 as the outcome of some radiation energy emitted by Earth returning to it. More precisely, CH_4 absorbs infrared radiation and re-emits in all directions, partly towards Earth. Gas is not a mirror bouncing each ray back unchanged.

Statements 2 and 3 instead address solar radiation and merely say not absorbed or not reflected. They do not establish the required reduction in Earth's net energy loss to space and are not the official choices. Preserve the original wording when discussing the question, but use absorption and re-emission in your own physical explanation.

Practice

1. (2025·May/June 0620/41 Q1(d)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(d) is produced by the decomposition of vegetation

2. (2025·May/June 0620/13 Q33) · Level 1

The equation for the reaction between carbon dioxide and water in the presence of chlorophyll and sunlight is shown.

$$\text{carbon dioxide + water} \xrightarrow[\text{sunlight}]{\text{chlorophyll}} \text{glucose + oxygen}$$

Which process does this equation represent?

A combustion

B decomposition

C displacement

D photosynthesis

3. (2025·May/June 0620/22 Q31) · Level 2

Which statements about carbon dioxide are correct?

1 Carbon dioxide absorbs thermal energy from the Earth.

2 Carbon dioxide increases the thermal energy loss to space.

3 The level of carbon dioxide in the atmosphere has increased due to the burning of fossil fuels.

4 Carbon dioxide is the only gas that causes global warming.

A 1 and 2 **B** 1 and 3 **C** 2 and 3 **D** 2 and 4

Organic chemistry

11.1 Molecular formulae, structural formulae and functional groups

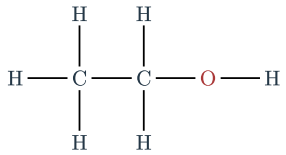
Type 1: Identify organic families from atoms, functional groups and general formulae

Key Ideas **Start by asking which atoms a molecule contains and how they are connected.** In organic chemistry, carbon is represented by C, hydrogen by H and oxygen by O. We begin with some simple molecules made from these elements. Carbon usually forms 4 covalent bonds, hydrogen forms 1 and oxygen forms 2. When drawing a displayed formula, check each atom individually; a plausible overall total is not enough.

Displayed, structural and molecular formulae each answer a different question. A displayed formula shows every atom and every bond. A structural formula groups some C—H bonds into units such as CH_3 and CH_2, while still making the order of connections clear. A molecular formula gives only the total number of atoms of each element. For example, the structural formula CH_3CH_2OH for ethanol shows a CH_3 group joined to a CH_2 group, which is then joined to O—H.

Three formulae can represent the same ethanol molecule

Displayed formula: all atoms and bonds



Structural formula: keeps connectivity

CH_3CH_2OH

Left C—middle C—O—H

Molecular formula: total atom counts only

C_2H_6O

C: 2; H: 3+2+1; O: 1

Short lines represent covalent bonds. The structural diagrams show connections, rather than the actual shape in space. All three boxes represent the same ethanol molecule.

To obtain the molecular formula from CH_3CH_2OH, count each element separately. CH_3 and CH_2 each contain 1 C atom, giving 2 C atoms in total; OH contains 1 O atom. Include the H in OH when counting hydrogen:

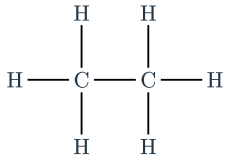
$$\text{Total number of H atoms} = 3 + 2 + 1 = 6$$

The molecular formula is therefore C_2H_6O. Conversely, knowing only C_2H_6O does not uniquely determine how the atoms are connected, so a molecular formula cannot automatically replace a displayed formula.

A functional group is an atom or group of atoms that determines the chemical properties of a particular homologous family. Alkenes contain C=C. The monohydric alcohols in this section contain —OH attached to carbon. Carboxylic acids contain —COOH, which includes one C=O and one O—H. Recognise —COOH as a complete carboxyl group: the OH within it does not make a carboxylic acid an alcohol. Alkanes are identified by a saturated hydrocarbon framework containing only C and H, with single bonds between all carbon atoms.

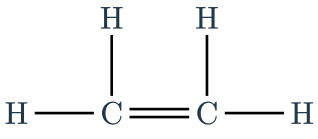
Find the functional group that controls the family's characteristic reactions

Ethane: alkane



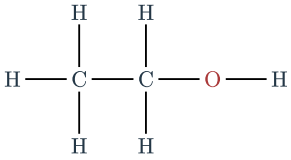
All carbon–carbon bonds are single

Ethene: alkene



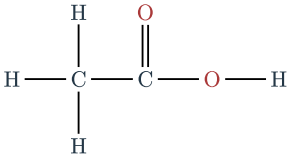
Functional group: C=C

Ethanol: alcohol



Functional group: —OH

Ethanoic acid: carboxylic acid



Functional group: —COOH

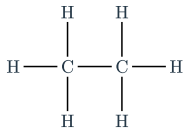
C=C is a carbon–carbon double bond. —OH must be attached to a suitable carbon framework. The carboxyl group —COOH contains one C=O and one O—H. Alkanes are identified by their saturated hydrocarbon framework.

In CH_3CH_2OH, O is joined by single bonds to C and H, so the compound is an alcohol. In CH_3COOH, the terminal C is joined to one O by a double bond and to an O—H group, so the compound is a carboxylic acid. Although both contain C, H and O, they can show different reactions because their atoms form different functional groups.

Decide whether a compound is a hydrocarbon and whether it is saturated in two separate steps. A hydrocarbon contains carbon and hydrogen only; ethanol contains oxygen and is therefore not a hydrocarbon. Saturated means that all carbon–carbon bonds in the molecule are single bonds. Unsaturated means that at least one carbon–carbon bond is not a single bond; the most common example in this section is C=C. Check the bonds between two carbon atoms: C=O is not C=C.

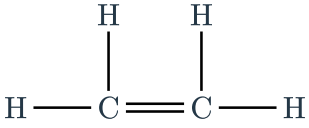
To check saturation, examine only bonds between carbon atoms

Ethane



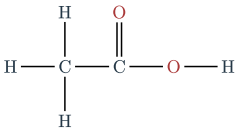
All C—C bonds single: saturated and a hydrocarbon

Ethene



Contains C=C: unsaturated and a hydrocarbon

Ethanoic acid



C—C is single.
C=O is not C=C: saturated.
It contains oxygen, so it is not a hydrocarbon.

Use the syllabus definition based on carbon–carbon bonds to decide whether a compound is saturated or unsaturated. Separately, check whether it contains only C and H to decide whether it is a hydrocarbon.

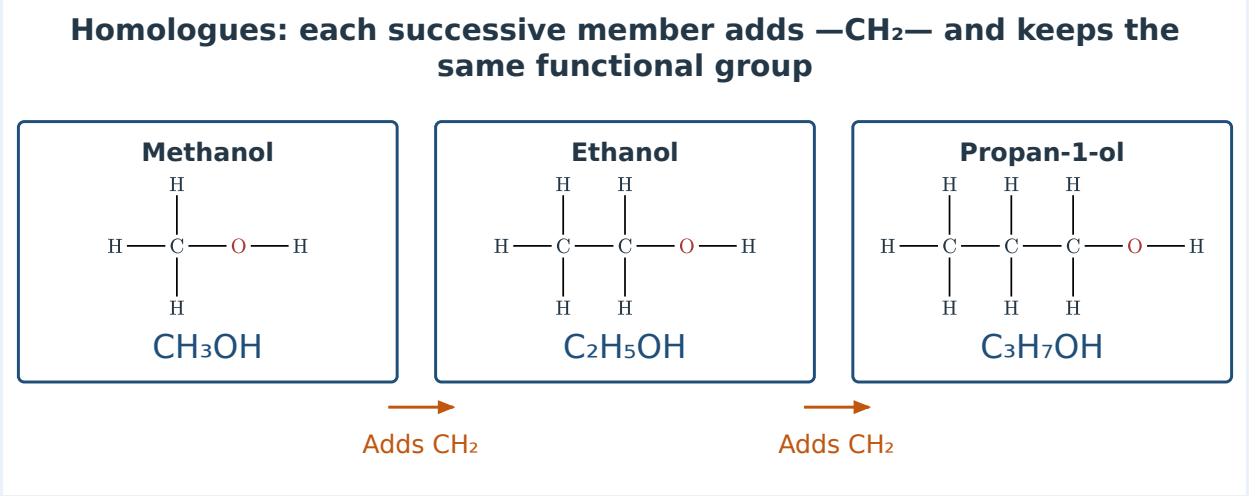
CH_3CH_3 contains only C and H, with a single carbon–carbon bond, so it is both a hydrocarbon and saturated.

$CH_2{=}CH_2$ also contains only C and H, but has C=C, so it is an unsaturated hydrocarbon.

Although CH_3COOH contains C=O, its carbon–carbon bond is single. It is therefore saturated according to the carbon–carbon bond definition, but it contains oxygen and is not a hydrocarbon.

A homologous series is a family of compounds with shared structural features and similar chemical properties. Members with the same functional group have similar chemical properties and follow the same

general formula. Moving from one member to the next adds one —CH_2— to the composition. Physical properties show a gradual trend; the members do not all have the same boiling point.



The arrows show that consecutive members differ by CH_2; they do not represent a direct reaction converting methanol into ethanol. The difference in relative molecular mass between consecutive members is 12+2=14.

Methanol, CH_3OH, and ethanol, C_2H_5OH, both contain the OH group of an alcohol. Ethanol has 1 more C and 2 more H atoms than methanol: an extra CH_2. This CH_2 consists of 1 C and 2 H atoms, so the increase in relative molecular mass is:

$$12 + 2 \times 1 = 14$$

Along the straight-chain alcohol series studied here, the boiling point generally increases as the carbon chain becomes longer, while the members still undergo similar alcohol reactions. If asked for two characteristics, you could give the same functional group and a difference of CH_2 between consecutive members. Do not claim that they have the same boiling point.

Define the letters in a general formula before using it. The general formulae for the alkanes, open-chain monoalkenes and saturated monohydric alcohols in this section are C_nH_{2n+2}, C_nH_{2n} and $C_nH_{2n+1}OH$ respectively. Here, n is the total number of carbon atoms in the molecule. An alkene needs at least two carbon atoms to form C=C. In the alcohol formula, the number of H atoms before OH is 2n+1; adding the 1 H in OH gives a total of 2n+2.

The syllabus gives the general formula for carboxylic acids as $C_nH_{2n+1}COOH$. Here, n counts the carbon atoms in the hydrocarbon part beside COOH; COOH itself contains one further carbon atom. For example, n=1 gives CH_3COOH, which has two carbon atoms in total and is called ethanoic acid. In methanoic acid, HCOOH, the hydrocarbon part is just H, corresponding to n=0. If N instead represents the total number of carbon atoms in the whole molecule, these saturated monocarboxylic acids contain N C atoms, 2N H atoms and 2 O atoms.

State which carbon atoms n counts before using a general formula

Alkanes	C_nH_{2n+2}	$n=3 \rightarrow C_3H_8$
n: all carbon atoms in the molecule		
Alkenes	C_nH_{2n}	$n=3 \rightarrow C_3H_6$
Here: open-chain compounds with one C=C		
Saturated monohydric alcohols	$C_nH_{2n+1}OH$	$n=3 \rightarrow C_3H_7OH \rightarrow C_3H_8O$
Include the extra H in OH in the total		
Saturated monocarboxylic acids	$C_nH_{2n+1}COOH$	$n=3 \rightarrow C_3H_7COOH \rightarrow C_4H_8O_2$
n counts the side-chain carbons; COOH contains one more C		

OH in an alcohol and COOH in a carboxylic acid are part of the general formula and must be included when finding the molecular formula. If using the total carbon count for a carboxylic acid, include the carbon in the carboxyl group and recalculate the hydrogen count.

First consider a saturated monohydric alcohol with three carbon atoms. Substituting n=3 into the alcohol general formula gives C_3H_7OH. Including the H in OH gives the molecular formula C_3H_8O.

Now consider a saturated monocarboxylic acid with three carbon atoms. With a total carbon count of N=3, it contains 3 C, 6 H and 2 O atoms, giving $C_3H_6O_2$. Alternatively, in the syllabus formula based on the side chain, n=2 gives C_2H_5COOH; counting its atoms again gives $C_3H_6O_2$.

Do not force more complex structures with several functional groups or C=C into these simple family formulae. Count the atoms directly from the diagram.

When reading a name, use the prefix to identify the carbon count and the suffix to identify the family. meth-, eth-, prop- and but- correspond to 1, 2, 3 and 4 carbon atoms respectively. -ane indicates an alkane, -ene an alkene, -ol an alcohol and -oic acid a carboxylic acid. The carbon count of a carboxylic acid includes the C in COOH. For example, methanoic acid, HCOOH, has one carbon atom in total; ethanoic acid, CH_3COOH, has two.

Read the carbon-count prefix, then the family suffix

Carbon atoms	Prefix	Alkane: -ane	Alkene: -ene	Carboxylic acid: -oic acid
1	meth-	methane CH_4	At least 2 C atoms	methanoic acid HCOOH
2	eth-	ethane C_2H_6	ethene C_2H_4	ethanoic acid CH_3COOH
3	prop-	propane C_3H_8	propene C_3H_6	propanoic acid C_2H_5COOH
4	but-	butane C_4H_{10}	butene C_4H_8	butanoic acid C_3H_7COOH

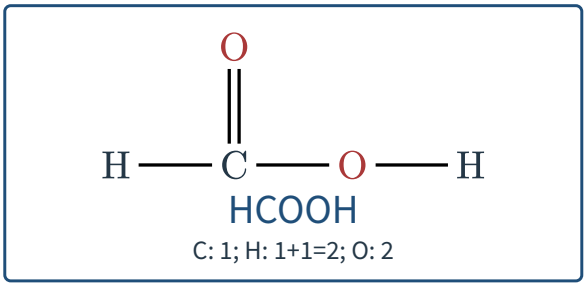
Alcohols end in -ol: methanol, ethanol, propanol and butanol have 1–4 carbons; state the OH position when needed.

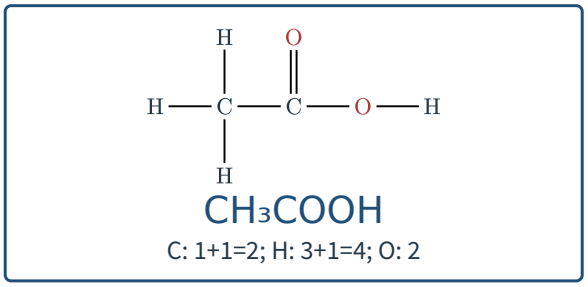
The same prefix indicates the same total number of carbon atoms. Butene, propanol and butanol also need a position number to specify the connections. Their displayed formulae are given in the next question type.

In propene, prop- means 3 carbon atoms and -ene indicates an alkene; its molecular formula in this family is C_3H_6. Propane also has 3 carbon atoms, but -ane indicates an alkane, giving C_3H_8. A difference of just one letter corresponds to different bond types and hydrogen counts. Ethanol is C_2H_6O: the shared eth- prefix does not make it the same compound as ethene, C_2H_4.

To find a molecular formula from a complex displayed formula, count each element in a separate pass. Count all the C atoms first, then all the H atoms, then O and any other elements. Count each explicitly drawn element symbol once. Bond lines are not atoms, and the two lines of a double bond do not add any atoms. For COOH, include its C, two O atoms and one H atom.

Count atoms: each element symbol represents an atom; bond lines are not atoms





Two parallel lines represent a double bond; they do not add two oxygen atoms. The H in O—H must also be included in the molecular formula.

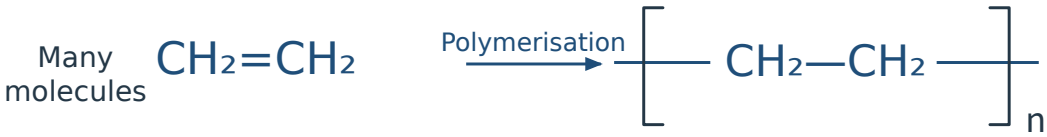
In HCOOH, the first H and the H in the final OH each contribute one H, giving H_2 in total. There is only one C in the middle, but two O atoms, so the molecular formula is CH_2O_2.

CH_3COOH has two C atoms. The 3 H atoms in CH_3 and the 1 H in OH give 4 H atoms, with two O atoms as well, so its molecular formula is $C_2H_4O_2$.

If asked how many atoms a molecule contains, add all the subscripts in its molecular formula. Giving the relative molecular mass does not answer that question.

The word monomer describes what a molecule can do. A monomer is a small molecule that can join to others to form a large molecule. A polymer is a long chain made from many such units. Ethene can undergo addition polymerisation to form poly(ethene). The same ethene molecule can therefore be described as both a hydrocarbon and a monomer; these terms are compatible.

Ethene can act as a monomer: small molecules join into a long chain



Monomer: a small molecule that joins others to form a long chain.

The brackets contain the repeat unit; bonds extending at both ends connect to neighbouring units.

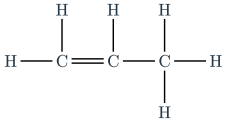
The reaction arrow represents addition polymerisation of ethene. CH_2 is a condensed structural group; n indicates many repeating units, not an extra atom.

When assessing descriptions of ethene as a hydrocarbon, a monomer, saturated or unreactive, check its composition, use, carbon–carbon bonding and reactivity separately. Ethene contains only C and H and can join into long chains. It has C=C, so it is unsaturated. Because C=C undergoes addition reactions, ethene cannot be described as generally unreactive in the way that alkanes are in this section.

Matching a general formula does not uniquely determine a structure. Within the simple open-chain families specified in this syllabus, C_2H_4 and C_4H_8 can be assigned to the alkene series using C_nH_{2n}. If other arrangements of atoms are allowed, however, C_nH_{2n} alone does not always prove the presence of C=C. For a statement that a compound **could be** a member of a family, one possible structure satisfying the conditions is enough. To say it **must be** a member, other possible connections must also be excluded.

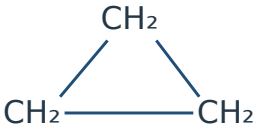
A molecular formula alone does not determine how the carbon atoms connect

Propene: $CH_2=CH—CH_3$



C_3H_6, with C=C

Another possible ring structure



Also C_3H_6, but all C—C bonds are single

The diagram on the right shows the limits of conclusions drawn from a molecular formula. Naming cyclic structures is not required in this section. Each CH_2 contains 1 C and 2 H atoms.

$C_4H_8O_2$ has the atom ratio of the saturated monocarboxylic acids studied here, so it could be a carboxylic acid. The molecular formula alone, however, does not prove that it contains COOH. Although C_3H_8O contains oxygen, it has only one O atom, whereas a carboxylic acid requires at least two. C_6H_{14} contains no O atoms at all, so it cannot be one of the alcohols or carboxylic acids in this section.

Members of a homologous series have the _____ functional group and _____ chemical properties; consecutive members differ by _____. Hydrocarbons contain only the two elements _____. To identify unsaturation, look for _____. The molecular formula obtained by combining the groups in C_3H_7OH is _____ HCOOH contains _____ atoms in total.

Answers: same; similar; CH_2; carbon and hydrogen; a carbon–carbon bond that is not single, such as C=C; C_3H_8O; 5

Recognising the Question First identify whether the question asks for a family, a molecular formula or an atom count. Look for the functional group to identify a family, and count each element separately for a molecular formula. Define which carbon atoms a general formula counts before using it. Distinguish a possible structure from a uniquely determined structure.

Worked Example 1 (2025·May/June 0620/11 Q33) · Level 1

Which row identifies compounds in the same homologous series?

	chemical properties	functional group
A	different	different
B	different	same
C	similar	different
D	similar	same

Solution

The table requires both the chemical properties and the functional group to be described correctly in the same row. Members of a homologous series have the same functional group and therefore undergo similar chemical reactions. The first column should be similar and the second same, which gives D.

A and B describe the chemical properties as different. C gives similar properties but different functional groups. Checking both columns leaves only D correct, so choose D. Similar chemical properties do not require all physical properties, such as boiling point, to be identical.

Worked Example 2 (2025·May/June 0620/41 Q5(c)(ii)) · Level 2

(c) Methanol is a member of the homologous series of alcohols.

(ii) State the molecular formula of an alcohol that contains five carbon atoms.

Solution

The question asks for the molecular formula of an alcohol with five carbon atoms. Use the syllabus general formula for saturated monohydric alcohols, $C_nH_{2n+1}OH$, where n is the total number of carbon atoms in the molecule. Substituting n=5 gives $C_5H_{11}OH$. Include the H in OH in the total hydrogen count:

$$\text{Total number of H atoms} = 11 + 1 = 12$$

There is 1 O atom, so the molecular formula is $\mathbf{C_5H_{12}O}$.

The question does not ask you to name a five-carbon alcohol or draw all its isomers. Only the correct molecular formula is needed. $C_5H_{11}O$ would omit the H in OH.

Worked Example 3 (2025·May/June 0620/42 Q6(e)(i)) · Level 2

(e) Compound **B** has the displayed formula shown in Fig. 6.1.

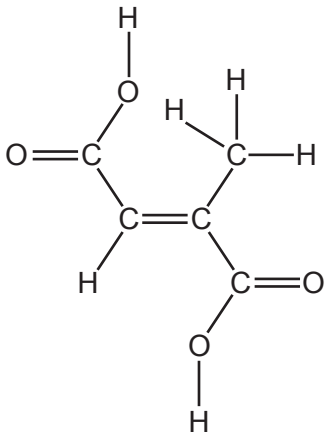


Fig. 6.1

(i) Deduce the molecular formula of compound **B**.

Solution

Count each element separately in Fig.6.1. Do not apply the general formula for saturated monocarboxylic acids just because two COOH groups are visible: this molecule also contains C=C and has a more complex structure than that family.

Count C first: 1 in the carboxyl group at the upper left, 2 in the central C=C, 1 in the CH_3 group at the upper right and 1 in the carboxyl group at the lower right. Add these four parts:

$$\text{Total number of C atoms} = 1 + 2 + 1 + 1 = 5$$

Now count H: the two O—H groups contribute 1 each, giving 2; there is 1 below the left-hand carbon of C=C; and the CH_3 group at the upper right contains 3. Add these three parts:

$$\text{Total number of H atoms} = 2 + 1 + 3 = 6$$

Finally, count O: each carboxyl group contains 2 O atoms, so the two groups contain 4. The molecular formula is therefore $\mathbf{C_5H_6O_4}$. Bond lines do not count as atoms, and a double bond does not double the number of C or O atoms.

Practice

1. (2025·May/June 0620/13 Q34) · Level 1

Which statement about members of a homologous series is correct?

A They are elements with the same chemical properties.

B They are compounds with the same functional group.

C They are atoms with the same number of outer shell electrons.

D They are molecules with the same boiling point.

2. (2025·May/June 0620/42 Q6(a)) · Level 1

Ethene, C_2H_4, is the first member of a family of similar compounds which contains the alkene functional group.

(a) State the term for a family of similar compounds which contain the same functional group.

3. (2025·May/June 0620/42 Q6(e)(ii)) · Level 1

(e) Compound **B** has the displayed formula shown in Fig. 6.1.

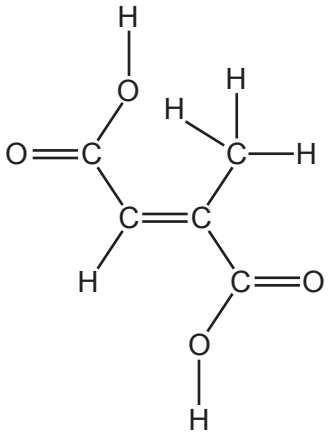


Fig. 6.1

(ii) State why compound **B** is unsaturated.

4. (2025·May/June 0620/43 Q1(b)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(b) is an unsaturated hydrocarbon

5. (2025·May/June 0620/43 Q1(g)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(g) belongs to a homologous series whose general formula is C_nH_{2n+2}

6. (2025·May/June 0620/43 Q5(c)(i)) · Level 1

(c) Ethanoic acid is a member of the carboxylic acid homologous series.

(i) Name the carboxylic acid that contains only one carbon atom.

7. (2025·May/June 0620/12 Q32) · Level 2

The formulae for three organic compounds are listed.

1 C_2H_4

2 C_4H_8

3 C_4H_{10}

Which compounds belong to the same homologous series?

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

8. (2025·May/June 0620/41 Q1(g)) · Level 2

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(g) is a simple molecule with 11 atoms

9. (2025·May/June 0620/41 Q5(c)(i)) · Level 2

(c) Methanol is a member of the homologous series of alcohols.

(i) State **two** characteristics of all members of a homologous series.

10. (2025·May/June 0620/42 Q6(b)) · Level 2

Ethene, C_2H_4, is the first member of a family of similar compounds which contains the alkene functional group.

(b) Determine the difference in relative molecular mass between C_2H_4 and the next member in this family of similar compounds.

11. (2025·May/June 0620/43 Q1(e)) · Level 2

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(e) contains simple molecules with 9 atoms

12. (2025·May/June 0620/43 Q5(c)(ii)) · Level 2

(c) Ethanoic acid is a member of the carboxylic acid homologous series.

(ii) State the molecular formula of a carboxylic acid that contains four carbon atoms.

13. (2025·May/June 0620/11 Q37) · Level 2

Some words used to describe organic compounds are listed.

1 hydrocarbon

2 monomer

3 saturated

4 unreactive

Which words describe ethene?

A 1 and 2 **B** 1 and 3 **C** 2 and 4 **D** 3 and 4

14. (2025·May/June 0620/13 Q35) · Level 3

The molecular formulae of four organic compounds, W, X, Y and Z, are shown.

W	X	Y	Z
C_3H_8O	C_3H_6	$C_4H_8O_2$	C_6H_{14}

Which statement about compounds W, X, Y and Z is correct?

A W and Z could both be alcohols.

B Y is the only compound that could be a carboxylic acid.

C W is a hydrocarbon.

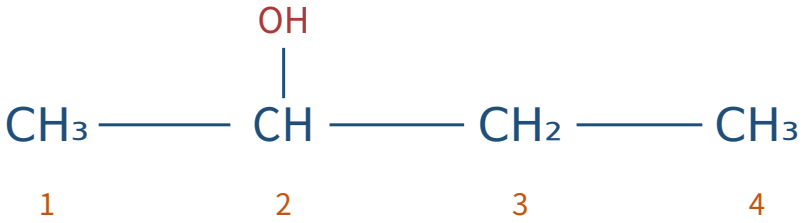
D X is ethene.

11.2 Organic nomenclature and structural isomerism

Type 1: Draw structures from names, count bonds and remove duplicate isomers

Key Ideas **Break a name into three pieces of information: carbon count, family and position.** In butan-2-ol, but- requires four carbon atoms, -ol identifies an alcohol, and 2 means that OH is attached to carbon 2. Draw the four-carbon chain first, put OH on the specified carbon, then add hydrogen so that each C forms 4 bonds, each O forms 2 and each H forms 1. The number is neither a count of molecules nor a charge on carbon.

Butan-2-ol: four carbons, with OH on carbon 2



Number from the end nearer OH; numbering it 3 from the other end must be changed to 2.

$CH_3CH(OH)CH_2CH_3$ and $CH_3CH_2CH(OH)CH_3$ show the same molecule written in opposite directions.

Horizontal and vertical lines show connections. The numbers identify carbon positions; they are not coefficients or charges on carbon atoms.

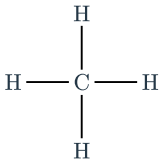
Propan-2-ol requires three carbon atoms with OH on the middle carbon, giving $CH_3CH(OH)CH_3$. The middle C is already bonded to the C atoms on either side and to O, leaving just one bond for H, so write CH rather than CH_2. In propan-1-ol, OH is at the end, giving $CH_3CH_2CH_2OH$. These compounds have the same molecular formula but different connections.

For a straight-chain alkane, join the carbon atoms first, then complete four bonds around each carbon. Methane, with one carbon atom, is CH_4. When two carbon atoms are joined, C—C uses one bond from each, so each end carbon is joined to three H atoms: ethane is CH_3CH_3. A C atom inside a straight chain is joined to two other C atoms, leaving two bonds for H, so it is written CH_2.

Straight-chain alkanes with 1–4 carbon atoms

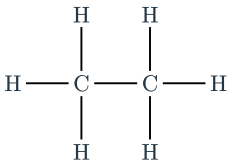
Methane

CH_4



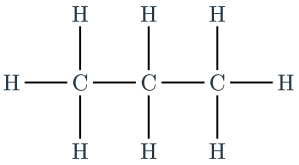
Ethane

CH_3CH_3



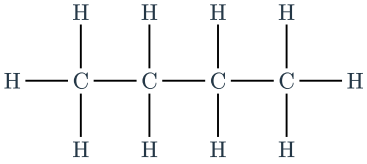
Propane

$CH_3CH_2CH_3$



Butane

$CH_3CH_2CH_2CH_3$



Every C, H and O atom and every bond is shown on the right. Changing a molecule's orientation on the page does not change its connections.

To draw butane, begin with C—C—C—C. Add three H atoms to each end carbon and two to each middle carbon, giving $CH_3CH_2CH_2CH_3$. Add the H atoms attached to the individual carbons:

$$\text{Total number of H atoms} = 3 + 2 + 2 + 3 = 10$$

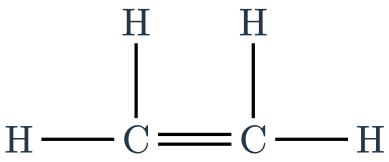
This agrees with C_4H_{10}, obtained from the alkane general formula for four carbon atoms. Writing only C_4H_{10} does not show how the carbon atoms are connected, so it cannot replace a displayed formula when one is requested.

The position number in an alkene name identifies the carbon where the double bond begins. In C=C, each participating carbon uses two of its four bonds to join the other carbon, so the alkene has two fewer H atoms than the corresponding alkane. Number from the end that gives the double bond the lower number. In but-1-ene, the double bond is between carbons 1 and 2; in but-2-ene, it is between carbons 2 and 3.

Straight-chain alkenes with one C=C and 2–4 carbon atoms

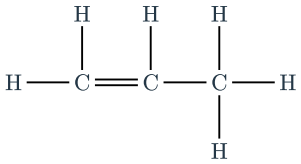
Ethene

$CH_2=CH_2$



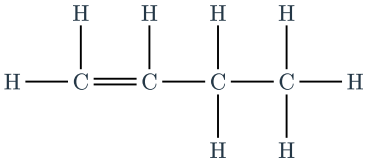
Propene

$CH_2=CHCH_3$



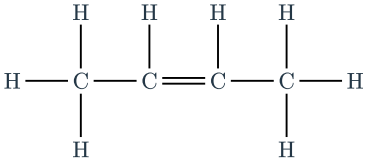
But-1-ene

$CH_2=CHCH_2CH_3$



But-2-ene

$CH_3CH=CHCH_3$



Every C, H and O atom and every bond is shown on the right. Changing a molecule's orientation on the page does not change its connections.

In $CH_3CH_2CH=CH_2$, counting from the right places the start of the double bond at carbon 1, so the name is but-1-ene. Do not call it but-3-ene simply because the formula was written from left to right.

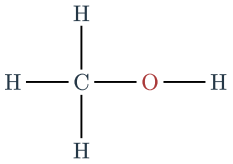
In $CH_3CH=CHCH_3$, the double bond is in the middle and has position number 2 from either end, so the name is but-2-ene. Both compounds are C_4H_8; the position of the double bond differs.

The OH position in an alcohol identifies which carbon has had an H replaced by OH. The simple straight-chain structures of methanol and ethanol do not need a separate OH position number. From a three-carbon chain onwards, placing OH at an end or in the middle gives different connections. In a displayed formula, show O and H separately. O cannot form three ordinary covalent bonds at once.

Straight-chain alcohols with 1–3 carbons: also locate OH

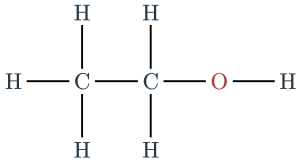
Methanol

CH_3OH



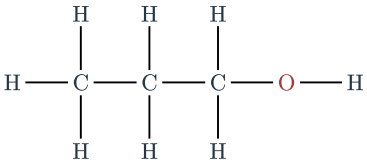
Ethanol

CH_3CH_2OH



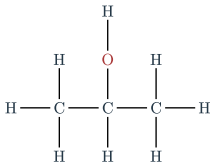
Propan-1-ol

$CH_3CH_2CH_2OH$



Propan-2-ol

$CH_3CH(OH)CH_3$



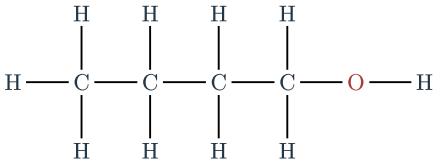
Every C, H and O atom and every bond is shown on the right. Changing a molecule's orientation on the page does not change its connections.

In $CH_3CH_2CH_2OH$, OH is at the end. Writing it backwards as $HOCH_2CH_2CH_3$ still gives propan-1-ol. In $CH_3CH(OH)CH_3$, OH is on the middle carbon, giving propan-2-ol. Both contain 3 C, 8 H and 1 O atoms, so their molecular formula, C_3H_8O, cannot distinguish them.

Straight-chain four-carbon alcohols: reverse numbering to exclude duplicates

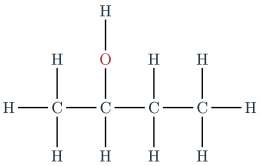
Butan-1-ol

$CH_3CH_2CH_2CH_2OH$



Butan-2-ol

$CH_3CH(OH)CH_2CH_3$



Every C, H and O atom and every bond is shown on the right. Changing a molecule's orientation on the page does not change its connections.

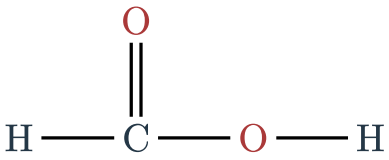
On a straight four-carbon chain, OH on what appears to be carbon 3 is on carbon 2 when numbered from the other end, so this is still butan-2-ol. Similarly, carbon 4 becomes carbon 1. After removing these duplicates caused by reversing the numbering, the OH positions for straight-chain butanols required here are 1 and 2.

The carbon count in a carboxylic acid name includes the carbon in the carboxyl group. To draw the group, join the terminal carbon to one O by a double bond and to O—H by a single bond. Use its remaining bond to join H or a carbon chain. Methanoic acid contains only this carboxyl carbon, so it has H on the left; ethanoic acid contains two carbon atoms, so CH_3 is attached on the left of the carboxyl group.

Straight-chain carboxylic acids with 1–4 carbons: include the COOH carbon in the name

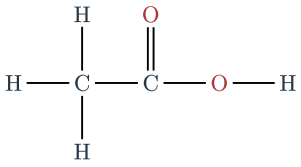
Methanoic acid

$HCOOH$



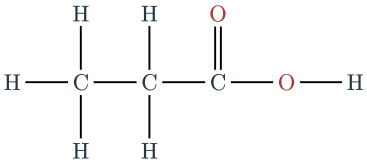
Ethanoic acid

CH_3COOH



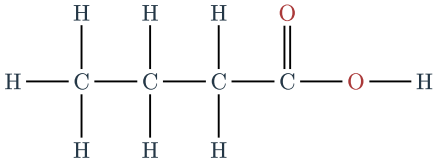
Propanoic acid

CH_3CH_2COOH



Butanoic acid

$CH_3CH_2CH_2COOH$



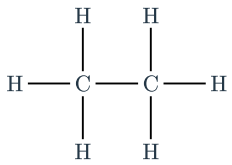
Every C, H and O atom and every bond is shown on the right. Changing a molecule's orientation on the page does not change its connections.

To draw propanoic acid with three carbon atoms, reserve one C for COOH. The other two form CH_3CH_2—, giving CH_3CH_2COOH. Writing $CH_3CH_2CH_2COOH$ instead gives four carbon atoms and therefore butanoic acid. Count the total carbons first, then assign one to the carboxyl group, to avoid a one-carbon mismatch between name and structure.

A total bond count must include every type of connection. C—H, C—C, C—O and O—H in a displayed formula are all covalent bonds. Where only single bonds are present, each line represents one covalent bond. Count each type separately, add the totals, then check every bond against the structure. Atom count and bond count are different: adding the subscripts in a molecular formula does not give the number of bonds.

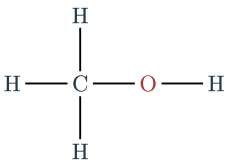
Count each type of bond line; atom count is not bond count

Ethane, C_2H_6



6 C—H + 1 C—C = 7 covalent bonds

Methanol, CH_3OH



3 C—H + 1 C—O + 1 O—H = 5 covalent bonds

Both molecules here contain only single bonds, so each connecting line counts as 1 covalent bond. If double bonds are present, first check whether the question asks for bonds or electron pairs.

Methanol, CH_3OH, has 3 C—H bonds, 1 C—O bond and 1 O—H bond, giving 5 bonds in total. To find its total number of atoms instead, count C, H and O:

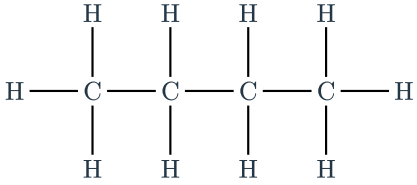
$$\text{Total number of atoms} = 1 + 4 + 1 = 6$$

This shows why total atoms and total covalent bonds must be treated separately. If the original structure contains a double bond, check whether the question asks about connections, covalent bonds or bonding electrons. Do not automatically apply the counting method used for single bonds alone.

For structural isomerism, check the totals before comparing the connections. Structural isomers are compounds with the same molecular formula but different structural formulae. First count every element in both structures. If their atom counts differ, they cannot be structural isomers. If the counts match, examine which atoms are joined to which. Reversing or rotating a drawing does not change these neighbours and does not create a new isomer.

Structural isomers have the same molecular formula but different atom connections

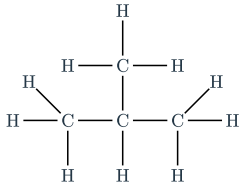
Four carbons in one continuous chain



$CH_3CH_2CH_2CH_3$

Both C_4H_{10}, with different connections

Central carbon joins three carbon branches



$CH_3CH(CH_3)CH_3$

Both C_4H_{10}, with different connections

Structural isomers must have both the same molecular formula and different connections. Reversing or rotating a drawing does not produce a new structural isomer.

$CH_3CH_2CH_2CH_3$ and $CH_3CH(CH_3)CH_3$ are both C_4H_{10}. In the second structure, the middle carbon is joined to three other carbon atoms, whereas no carbon in the first structure has this arrangement. Their connections therefore differ, making them structural isomers. The CH_3 in brackets is attached to the CH immediately before the brackets; it does not multiply the whole chain by 3.

Expand condensed oxygen-containing formulae according to the actual atoms present. In CH_3COOCH_3, COO means 1 C and 2 O atoms; it is not a new element symbol. A question may directly give the structure of an ester, such as methyl ethanoate, CH_3COOCH_3. To compare molecular formulae, count directly from the given

structure; you do not need to derive how the ester is made first. Ester naming and synthesis are covered in detail in a later question type.

Expand a structural formula into atom counts before comparing molecular formulae

CH_3COOCH_3	C: 1+1+1=3	H: 3+3=6	O: 2
$HCOOCH_2CH_3$	C: 1+1+1=3	H: 1+2+3=6	O: 2
CH_3CH_2COOH	C: 1+1+1=3	H: 3+2+1=6	O: 2
$HOCH_2CH_2CH_2OH$	C: 1+1+1=3	H: 1+2+2+2+1=8	O: 2

Count H outside brackets, in OH and at the start of HCOO; equal molecular formulae still require a connectivity comparison.

COO means 1 C and 2 O atoms, not a new element symbol. The table shows how the atom counts build up one group at a time.

$HOCH_2CH_2CH_2OH$ also appears to have three carbon atoms and two oxygen atoms, but each terminal OH contains one H. Including the H atoms in all three CH_2 groups gives:

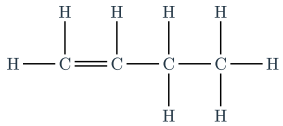
$$1 + 2 + 2 + 2 + 1 = 8$$

If the reference molecule has only 6 H atoms, this cannot be its structural isomer. Do not overlook hydrogen simply because both structures have three carbon atoms and two oxygen atoms.

When several restrictions apply, check each one. Unbranched means that the carbon chain has no side branches. Decolourise aqueous bromine means that the colour of bromine water disappears. In the usual hydrocarbon test studied here, alkenes containing C=C undergo addition and turn bromine water from orange-brown to colourless. Alkanes do not undergo this addition. Use this observation to identify candidates containing C=C, then draw structures that also meet the carbon-count and straight-chain requirements.

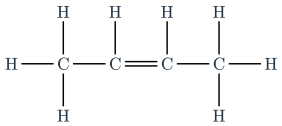
C_4H_8: keep a straight chain and place C=C in distinct positions

Position 1: but-1-ene



$CH_2=CHCH_2CH_3$

Position 2: but-2-ene



$CH_3CH=CHCH_3$

A double bond at position 3 becomes position 1 when numbered from the other end; do not count it again.

Both structures contain C=C and decolourise bromine water. This diagram counts structural isomers only, not different spatial arrangements with the same connections.

Arrange four carbon atoms in a straight chain and place C=C at the left, in the middle and at the right in turn. The right-hand position becomes the left-hand position when read from the other end, so remove that duplicate. Exclude branched structures because they fail the unbranched condition. Finally, add H so that every C has four bonds and check that the molecular formula remains C_4H_8. If asked for structural isomers, count changes in connectivity, not different spatial arrangements of the same connections.

but- indicates _____ carbon atoms; -ol indicates an _____. Number the chain from the end that gives OH the _____ position number. Structural isomers must have the same _____ but different _____. When counting bonds in ethane, include the _____ as well as the C—H bonds.

Answers: 4; alcohol; lower; molecular formula; structural formulae, or atom connections; C—C bond

Recognising the Question Use the prefix to draw the carbon chain, then place the functional group at the specified position and add H. For isomers, first check that the molecular formulae match, then compare connections. Remove duplicates caused only by rotation or reversed numbering. Bond counts must include more than just C—H bonds.

Worked Example 1 (2025·May/June 0620/22 Q32) · Level 2

What is the structural formula of butan-2-ol?

A $CH_3CH_2CH_2OH$

B $CH_3CH(OH)CH_3$

C $CH_3CH_2CH_2CH_2OH$

D $CH_3CH_2CH(OH)CH_3$

Solution

Break down butan-2-ol: but- requires 4 carbon atoms, -ol requires an alcohol OH group, and 2 places OH on the second carbon.

A, $CH_3CH_2CH_2OH$, has only 3 carbon atoms. B, $CH_3CH(OH)CH_3$, also has only 3, so both are excluded. C has 4 carbon atoms but OH is at the end, making it butan-1-ol.

D is $CH_3CH_2CH(OH)CH_3$. Numbering from right to left places OH on carbon 2, and the chain has exactly 4 carbon atoms, so choose **D**. Reversing the formula to $CH_3CH(OH)CH_2CH_3$ gives the same structure. Do not always number from the left-hand side of the page.

Worked Example 2 (2025·May/June 0620/21 Q34) · Level 2

What is the **total** number of covalent bonds in one molecule of propane?

A 8 **B** 9 **C** 10 **D** 11

Solution

Propane is a three-carbon alkane. First turn its name into the structure CH_3—CH_2—CH_3.
Count the C—H bonds on each carbon:

$$\text{Number of C—H bonds} = 3 + 2 + 3 = 8$$

Joining three carbon atoms into one chain also requires 2 C—C bonds. Add the two types of bond:

$$\text{Total number of covalent bonds} = 8 + 2 = 10$$

Choose **C**.

A counts only the 8 C—H bonds and omits C—C. D gives 11, which is the atom count, 3+8, rather than the bond count. Count every connection once and make sure the quantity counted is the one requested.

Worked Example 3 (2025·May/June 0620/21 Q33) · Level 3

The structural formula of methyl ethanoate is CH_3COOCH_3.

Which compounds are structural isomers of methyl ethanoate?

1 $HCOOCH_2CH_3$

2 CH_3CH_2COOH

3 $HOCH_2CH_2CH_2OH$

A 1, 2 and 3 **B** 1 and 2 only **C** 2 and 3 only **D** 3 only

Solution

First count the atoms in the reference compound, methyl ethanoate, CH_3COOCH_3. It has 3 C atoms, 3+3=6 H atoms and 2 O atoms, so its molecular formula is $C_3H_6O_2$. A structural isomer must have the same molecular formula but different connections.

1 is $HCOOCH_2CH_3$. There is 1 C atom in each of three positions, giving 3 C atoms, and there are 2 O atoms. The hydrogen count is:

$$1 + 2 + 3 = 6$$

Its molecular formula is also $C_3H_6O_2$, but the carbon atoms are distributed differently on the two sides of the ester linkage, so it is a structural isomer.

2 is CH_3CH_2COOH. It contains 3 C and 2 O atoms. The hydrogen count is:

$$3 + 2 + 1 = 6$$

Its molecular formula is again $C_3H_6O_2$, but the two O atoms form COOH, which differs from the reference compound's ester linkage. It is also a structural isomer.

3 is $HOCH_2CH_2CH_2OH$. It has 3 C and 2 O atoms. Include the H in both terminal OH groups:

$$\text{Total number of H atoms} = 1 + 2 + 2 + 2 + 1 = 8$$

Its molecular formula is $C_3H_8O_2$. The hydrogen count differs, so exclude it. Only 1 and 2 qualify; choose **B**. Do not omit the H in OH. Matching 3 C and 2 O atoms alone does not establish isomerism.

Practice

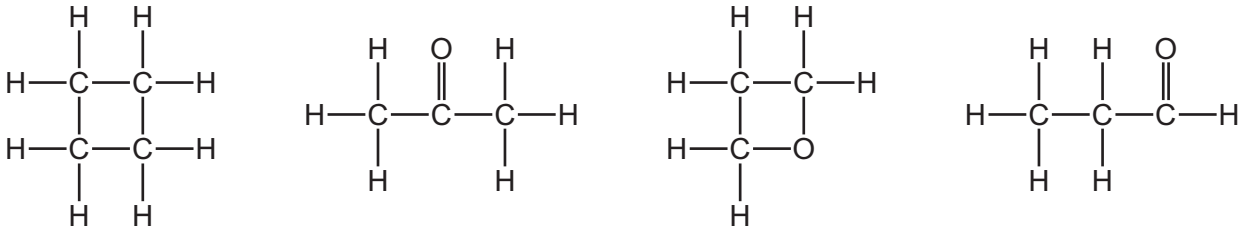
1. (2025·May/June 0620/42 Q1(g)) · Level 2

Using numbers only, state the:

(g) number of covalent bonds in one molecule of ethanol

2. (2025·May/June 0620/12 Q10) · Level 3

The diagrams show the structures of four molecules.



Which statement about these molecules is correct?

A All four molecules have different molecular formulae.

B All four molecules have the same molecular formula.

C Only two molecules have the same molecular formula.

D Only three molecules have the same molecular formula.

3. (2025·May/June 0620/42 Q1(f)) · Level 3

Using numbers only, state the:

(f) number of unbranched structural isomers of C_4H_8 which decolourise aqueous bromine

11.3 Fuels and petroleum

Type 1: Use the fractionating column to explain petroleum fractions, their properties and uses

Key Ideas **First identify what petroleum is.** Coal, natural gas and petroleum are fossil fuels. The main constituent of natural gas is methane, CH_4; main does not mean that the entire sample contains methane only. Petroleum is a mixture of many different hydrocarbons. No single chemical formula represents all of petroleum.

Distinguish the three fossil fuels

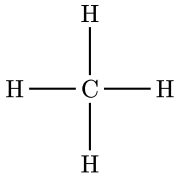
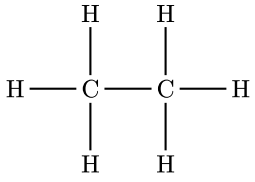
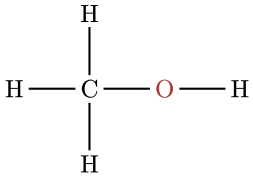
Coal	Natural gas	Petroleum
coal	natural gas	petroleum
Mainly carbon, with other substances	Main component: methane, CH_4	Mixture of many hydrocarbons

The three separate boxes list the fossil fuels and the composition details required in this section. There are no conversion arrows between the boxes.

If asked for the main constituent of natural gas, write methane. If asked for a fossil fuel made from many hydrocarbons, write petroleum. Establish whether the question wants the name of one substance or the name of a fuel mixture.

The word only is essential to the definition of a hydrocarbon. A hydrocarbon is a compound containing carbon and hydrogen only. First check that it is a compound, then examine every element symbol: it must contain no elements other than C and H, such as O, N or Cl.

A hydrocarbon contains only C and H

Methane	Ethane	Methanol
		
CH_4	C_2H_6	CH_3OH
Only C and H: a hydrocarbon	Only C and H: a hydrocarbon	Also contains O: not a hydrocarbon

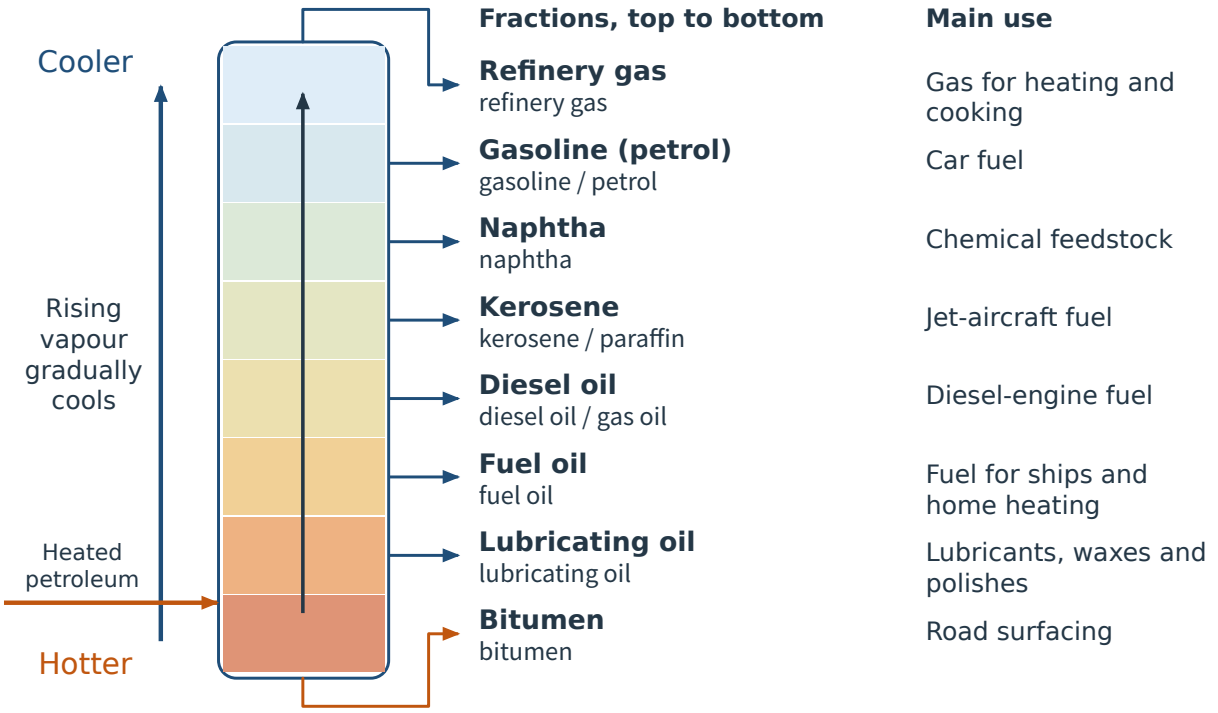
Short lines in the structures represent covalent bonds. C is carbon, H is hydrogen and O is oxygen. The diagrams show connections, rather than the molecules' actual shapes.

Methane, CH_4, and ethane, C_2H_6, contain only C and H, so they are hydrocarbons. Methanol, CH_3OH, contains carbon and hydrogen but also oxygen, so it is not a hydrocarbon. C_3H_8 lists only C and H and is a hydrocarbon; C_2H_5Cl also contains Cl and is not. Seeing a C alone is not enough.

How can the substances mixed together in petroleum be separated? Hydrocarbon molecules of different sizes have different boiling points. Heat petroleum so that most of it vaporises, then feed the hot mixture of vapours into a fractionating column. The bottom is hotter and the top is cooler, so the vapours cool as they rise.

When vapours reach a cooler region suitable for condensation, hydrocarbons within the corresponding boiling-point range become liquid and are collected separately. Higher-boiling components condense lower down. Lower-boiling components rise further before condensing higher up. The refinery gases, with the lowest boiling points, leave at the top. The heaviest components, which are difficult to vaporise, remain at the bottom as bitumen. This separation using differences in boiling point is called fractional distillation.

Fractional distillation of petroleum: cooler at the top, hotter at the bottom



Levels show collection positions; the column does not cut long molecules into shorter ones.

Arrows show vapours rising or fractions leaving. Refinery gases leave at the top; bitumen is the heavy material remaining at the bottom. The spacing between levels does not represent actual temperature differences.

Each collected part is called a **fraction**. A fraction still contains several hydrocarbons with similar boiling points; it is not a bottle of one pure substance. Fractional distillation separates existing substances into different locations without breaking their molecules into new molecules, so it is a physical separation. Cracking, studied later, breaks longer hydrocarbon molecules into shorter ones.

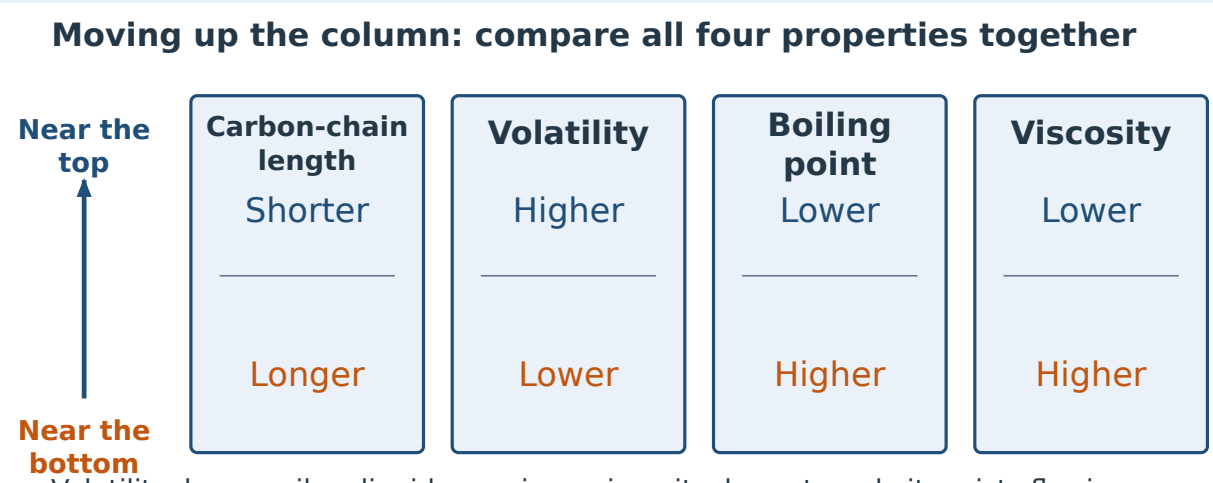
Consider two hydrocarbon vapours: A has a higher boiling point and B has a lower boiling point. As they rise into cooler parts of the column, A condenses lower down while B continues to rise. Explain their collection positions using different boiling points and different condensation levels, rather than claiming that A turns into B inside the column.

Learn the names, order and uses together. From the top down, the order is refinery gas, gasoline, naphtha, kerosene, diesel oil, fuel oil, lubricating oil and bitumen. Each row in the diagram gives the fraction's name and use. Gasoline and petrol mean the same fraction; kerosene and paraffin both refer to the same fraction in this section; gas oil is diesel oil. Do not confuse gas oil with refinery gas at the very top.

Connect each use to a specific situation: cars use gasoline/petrol, jet aircraft use kerosene/paraffin, and diesel engines use diesel oil/gas oil. Domestic heating may use fuel oil or refinery gas, so when a question specifies a fraction, check the complete name–use pairing in that row. Naphtha is a chemical feedstock. Lubricating oil is

used for lubricants, waxes and polishes, and bitumen for road surfacing. Refinery gas is also used for cooking, and fuel oil for ships.

Keep the directions of the four property trends distinct. Carbon-chain length describes how long the linked chain of carbon atoms is in a hydrocarbon molecule. Moving upwards through the column, molecules in the fractions generally become smaller, carbon chains shorter and boiling points lower. Volatility describes how readily a substance becomes vapour: easier vaporisation means higher volatility. Viscosity is resistance to flow: a substance that flows less readily has a higher viscosity.



Compare fractions collected at different positions in the same column, interpreting volatility and viscosity under similar external conditions. These are trends, not fixed numerical values.

The four trends for fractions higher in the column are therefore shorter carbon chains, higher volatility, lower boiling points and lower viscosity. The trends reverse towards the bottom. These are overall trends between fractions; you do not need to assign each fraction one fixed carbon count or one fixed boiling point.

Kerosene is below naphtha, so it generally has longer carbon chains, a higher boiling point, higher viscosity and lower volatility. If asked about just one property, first locate the two fractions relative to each other in the column, then state that property's trend. For example, saying that kerosene is more volatile than naphtha reverses the correct direction.

The main constituent of natural gas is _____. Hydrocarbons contain only the two elements _____ and _____. From the bottom to the top of the column, boiling points gradually _____ and volatility gradually _____. The fraction used as jet aircraft fuel is _____.

Answers: methane; carbon; hydrogen; decrease; increases; kerosene/paraffin

Recognising the Question Locate each petroleum fraction relative to the others in the column. For properties, check the four trends; for uses, check the complete name–use pairing. In a flow chart, follow every yes/no condition. In an ordering question, check the whole sequence.

Worked Example 1 (2025·May/June 0620/22 Q34) · Level 1

Petroleum can be separated into useful fractions by fractional distillation.

Which row shows a correct use of the named fraction?

	fraction	use
A	bitumen	lubricant
B	naphtha	fuel for home heating
C	kerosene	jet fuel
D	refinery gas	making chemicals

Solution

The question asks you to match each fraction to its use. Begin with the correct pair: kerosene is used as jet fuel, so the name and use in C match.

Check the other three rows. In A, bitumen is mainly used for road surfacing; lubricant corresponds to the lubricating-oil fraction. In B, naphtha is a chemical feedstock, rather than the domestic heating fuel specified in that row; fractions used for heating include fuel oil. In D, the syllabus uses of refinery gas are heating and cooking; making chemicals corresponds to naphtha as a chemical feedstock.

Only C correctly matches the specified fraction with its use in the same row, so choose C. Do not decide merely that several fractions can burn: check the particular use given.

Worked Example 2 (2025·May/June 0620/13 Q36) · Level 3

Fractional distillation is used to separate petroleum into useful fractions.

Which statement about the properties of the fractions of petroleum is correct?

A The refinery gas fraction has a lower volatility than the gasoline fraction.

B The molecules in the gasoline fraction have a longer chain length than the molecules in the naphtha fraction.

C The naphtha fraction has a higher boiling point than the kerosene fraction.

D The fuel oil fraction has a higher viscosity than the diesel oil fraction.

Solution

First restore the fractionating-column order. The relevant positions, from top to bottom, are refinery gas, gasoline, naphtha, kerosene, diesel oil and fuel oil. Fractions higher up generally have shorter carbon chains, higher volatility, lower boiling points and lower viscosity.

Check the direction of each comparison. In A, refinery gas is above gasoline and should have higher volatility, but the statement says lower volatility, so A is wrong. In B, gasoline is above naphtha and should have shorter carbon chains, but the statement says longer chain length, so B is wrong.

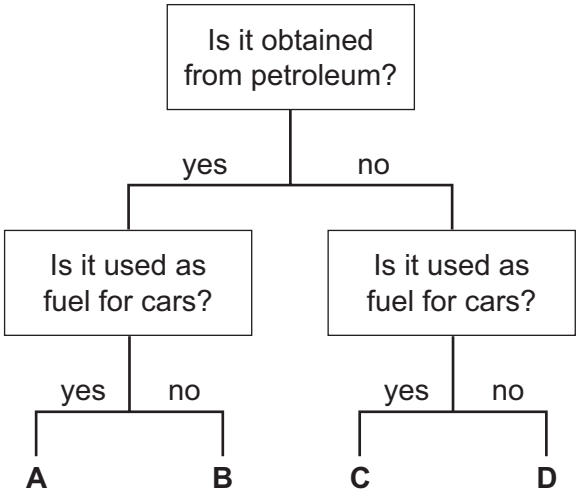
In C, naphtha is above kerosene and should have a lower boiling point, but the statement says higher boiling point, so C is wrong. In D, fuel oil is below diesel oil and generally resists flow more, giving higher viscosity. This agrees with the trend, so D is correct.

Choose D. The four statements compare four different properties. Each time, identify which fraction is higher before deciding the direction of the particular property change.

Practice

1. (2025·May/June 0620/12 Q33) · Level 2

Which fuel could be gasoline?



2. (2025·May/June 0620/11 Q35) · Level 2

Which list shows the fractions obtained from the fractional distillation of petroleum, in order of increasing boiling point?

A bitumen → diesel oil → fuel oil → lubricating oil

B diesel oil → gasoline → naphtha → kerosene

C gasoline → naphtha → kerosene → diesel oil

D kerosene → lubricating oil → naphtha → refinery gas

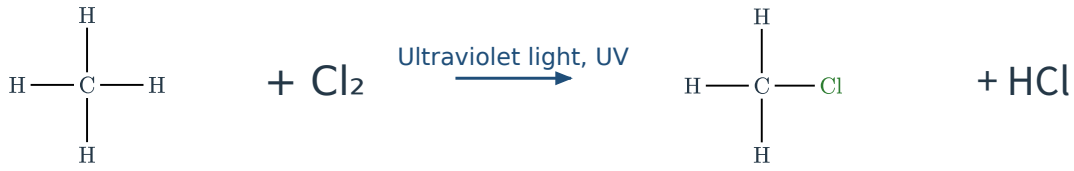
11.4 Alkanes

Type 1: Write alkane monosubstitution products while retaining the carbon framework

Key Ideas **Begin alkane reactions with their existing single-bonded structures.** Alkanes contain C and H only, and all carbon–carbon bonds are single covalent bonds, so they are saturated hydrocarbons. They are generally unreactive, but the syllabus explicitly requires two types of reaction: combustion and substitution with chlorine in ultraviolet light. Generally unreactive does not mean unable to react under any conditions.

Substitution replaces one atom or group of atoms with another. For methane and chlorine, one H in methane is replaced by one Cl. The carbon framework is retained, forming chloromethane, CH_3Cl. The displaced H joins the other Cl to form hydrogen chloride, HCl. The reaction needs ultraviolet light, abbreviated to UV. Light supplies the activation energy, E_a, needed to start the reaction, so this is a photochemical reaction.

One substitution: Cl replaces one H; the other Cl joins the displaced H to form HCl



CH_4: starting methane — CH_3Cl: chloromethane

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$; both sides contain 1 C, 4 H and 2 Cl.

The arrow represents the reaction. Ultraviolet light supplies the activation energy needed to initiate it. Only one H is replaced by Cl here, and the carbon framework is retained.

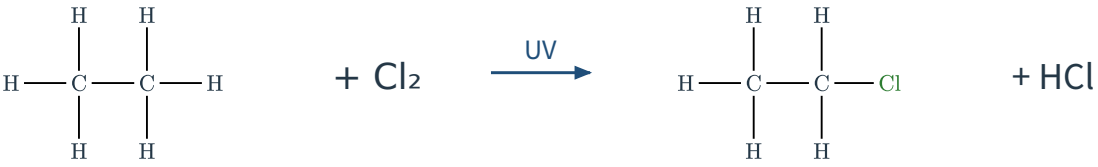
Start with CH_4 and replace only one H. Keep C_1, reduce the H count from 4 to 3 and add 1 Cl to obtain CH_3Cl. The second product, HCl, accounts for the displaced 1 H and the remaining 1 Cl. Include both products in the equation:

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

Both sides contain 1 C, 4 H and 2 Cl atoms, so atoms are conserved. Only monosubstitution is required here; do not replace several H atoms with Cl without being asked.

For another alkane, retain the carbon framework in the same way. Ethane is C_2H_6. In one chlorine substitution, the two carbon atoms remain joined and one of the six H atoms is replaced. The main product is C_2H_5Cl, chloroethane, and the other product is again HCl.

Apply substitution to ethane: keep the two-carbon skeleton, then replace one H



C_2H_6 → one fewer H, one more Cl — C_2H_5Cl: chloroethane

Check: C unchanged; 6 H split as 5 + 1; 2 Cl split as 1 + 1.

The complete displayed formula shows that the carbon–carbon single bond is retained. Only one H is substituted: the carbon chain is not broken, and Cl_2 does not mean two chlorine molecules.

If a question gives butane, C_4H_{10}, and asks for the reagent and reaction type, first identify it as an alkane using C_nH_{2n+2}, then connect it to chlorine, UV and substitution. There is no need to discuss products from different hydrogen positions merely to identify the reaction type. If only the molecular formula after one substitution is requested, it is C_4H_9Cl, with HCl also formed. Addition of steam involves C=C; do not apply an alkene addition reaction to a saturated alkane.

Complete combustion transfers carbon into CO_2 and hydrogen into H_2O. Complete combustion requires sufficient oxygen. To balance the equation, use the fuel's C count to set the number of CO_2 molecules, then its H count to set the number of H_2O molecules. Finally, calculate the total O atoms required in the products to set the O_2 coefficient. The subscript 2 in O_2 means that each oxygen molecule contains two oxygen atoms and must not be changed.

Complete combustion: use carbon to find CO_2, then hydrogen to find H_2O

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Count C first	Then count H	Count O last
1 C atom → 1 CO_2	4 H atoms → 2 H_2O	4 O atoms needed → 2 O_2

Coefficients change the numbers of molecules, not the formulae of CH_4, CO_2 or H_2O. The arrow represents complete combustion.

For complete combustion of ethane, C_2H_6, balance C and H first: two C atoms require $2CO_2$, and six H atoms require $3H_2O$. On the right, $2CO_2$ contains 4 O atoms and $3H_2O$ contains 3, giving:

$$4 + 3 = 7$$

Each O_2 supplies 2 O atoms, so 3.5 O_2 are needed per ethane molecule in this proportional calculation. Multiply the whole equation by 2 to obtain integer coefficients:

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

Check that both sides contain 4 C, 12 H and 14 O atoms. Insufficient oxygen may produce CO or soot; these are not products of the complete combustion specified in the question.

Connect the difference between substitution and addition to the starting structure and products. Alkanes have no C=C and cannot undergo the double-bond addition studied for alkenes here. Under normal test conditions, aqueous bromine remains orange-brown with an alkane. Alkenes contain C=C and undergo addition with bromine, removing its colour. Substitution retains the carbon framework while replacing atoms and often forms another small-molecule product. The simple additions studied here combine the reactants to form just one product.

Read the reagents and conditions before identifying the reaction

Alkane + chlorine, UV	One H replaced by Cl	Substitution; HCl also forms
Alkene + bromine water	Br adds to both carbons of the original C=C	Addition; orange-brown bromine water decolourises
Alkane + bromine water, usual test conditions	No C=C for this addition	Bromine water does not decolourise

The bromine-water comparison assumes normal test conditions. Keep the UV condition when discussing photochemical substitution, and distinguish it from the bromine-water addition test.

When you see chlorine+UV, check for photochemical substitution of an alkane. When aqueous bromine is decolourised, check for C=C that can undergo addition. The statement that alkanes do not decolourise bromine water assumes normal test conditions; do not discard the conditions and combine all halogen reactions into one rule. In a multiple-statement question, check the general formula, bond type, reactivity and experimental observation separately.

Alkanes contain _____ carbon–carbon bonds. Monosubstitution with chlorine requires _____. Replacing one H in methane gives the organic product _____ and the other product _____. The products of complete combustion are _____ and _____.

Answers: single; ultraviolet light, UV; CH_3Cl; HCl; CO_2; H_2O

Recognising the Question For alkane+chlorine+UV, identify monosubstitution: retain all C atoms, replace only one H with Cl, add HCl and check the atom counts. Balance complete combustion in the order C, H, then O. Link the bromine-water observation to C=C.

Worked Example 1 (2025·May/June 0620/22 Q35) · Level 2

Ethane reacts with chlorine in the presence of ultraviolet light.

Which products are formed?

A $C\mathit{l}CH_2CH_2C\mathit{l}$ and H_2

B $2CH_3C\mathit{l}$ and H_2

C $C_2H_5C\mathit{l}$ and $HC\mathit{l}$

D $CH_2{=}CH_2$ and $2HC\mathit{l}$

Solution

Ethane is C_2H_6. Chlorine and ultraviolet light in the question indicate photochemical substitution of an alkane.

Keep the two-carbon framework and replace only one H with Cl to obtain C_2H_5Cl. The displaced H joins the other Cl from Cl_2 to form HCl. The products are therefore **C_2H_5Cl and HCl**, corresponding to C.

Check the atoms: before reaction there are 2 C, 6 H and 2 Cl. C_2H_5Cl contributes 2 C, 5 H and 1 Cl, while HCl adds 1 H and 1 Cl, conserving the totals. A replaces two H atoms and gives H_2, which does not match monosubstitution here. B breaks the two-carbon framework. D forms C=C, which is also not the substitution specified. Choose **C**.

Practice

1. (2025·May/June 0620/11 Q34) · Level 2

The molecular formula of compound Z is C_4H_{10}.

Which row identifies the substance that reacts with Z and describes the type of reaction that occurs?

	substance	type of reaction
A	chlorine	addition
B	chlorine	substitution
C	steam	addition
D	steam	substitution

2. (2025·May/June 0620/13 Q37) · Level 2

Which statements about the alkanes are correct?

1 They are generally unreactive except in terms of combustion and substitution by chlorine.

2 They have the general formula C_nH_{2n+2}.

3 They contain double carbon–carbon covalent bonds.

4 They decolourise aqueous bromine.

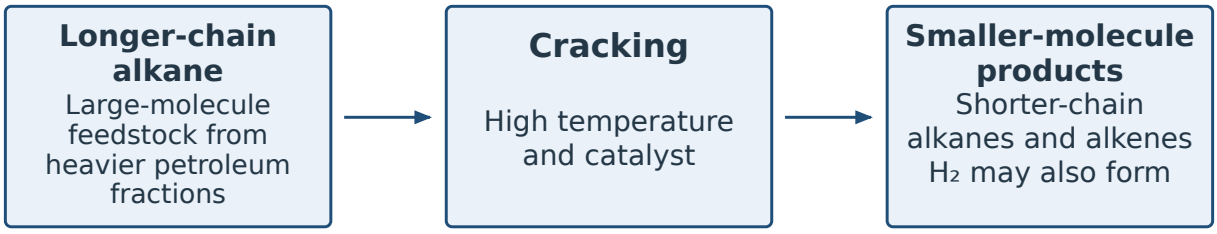
A 1, 2 and 3 **B** 1 and 2 only **C** 1, 3 and 4 **D** 2 and 4

11.5 Alkenes and cracking

Type 1: Use cracking conditions, uses and conservation of atoms to assess products

Key Ideas **Cracking turns larger molecules into smaller, useful molecules.** In cracking, longer-chain alkanes undergo bond breaking and rearrangement at a high temperature and in the presence of a catalyst, producing smaller molecules. When asked for conditions, state high temperature and a catalyst separately; simply saying heat it is insufficient.

Cracking uses high temperature and a catalyst to turn large molecules into smaller ones



Balanced equation: $C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$

Bonds break and rearrange to form new substances; this is not separation of a mixture by boiling point.

The arrow represents cracking. This is one possible combination of products; not every cracking reaction follows this equation.

$C_{10}H_{22}$ can form C_8H_{18} and C_2H_4 as shown. One type of molecule becomes two new substances, so this is a chemical change. Fractional distillation separates molecules already present according to their boiling points, without changing their chemical formulae. The phrase long-chain molecules made into short-chain molecules points to cracking.

Products often include alkenes, but they need not be alkenes only. Cracking can produce shorter-chain alkanes and alkenes, and may also produce hydrogen, H_2. The specific combination must conserve both carbon and hydrogen atoms. Do not restrict all possible products to one family in advance.

Cracking can give different products, but both C and H must be conserved

One combination

$C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$

Alkane + alkene; all 10 C and 22 H retained

Another combination

$C_{10}H_{22} \rightarrow C_8H_{16} + C_2H_4 + H_2$

Alkene + alkene + hydrogen; H_2 can also be a product

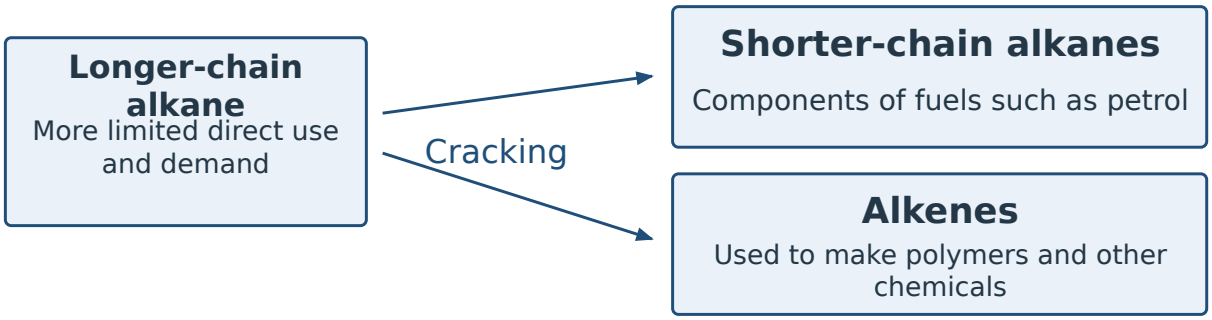
Only alkenes excludes other possible products and is too restrictive.

The same large alkane is used as the starting material in each case. The equations show possible combinations that conserve atoms; the actual distribution of products depends on the conditions.

On the right of the first equation, C_8H_{18} is an alkane and C_2H_4 is an alkene. Thus, cracking produces alkenes can be correct, whereas cracking only produces alkenes is false. The second equation contains 16+4+2=22 H atoms on the right, including H_2. The word only restricts the claim and must not be overlooked.

Explain the value of shortening chains by connecting it to uses and demand. Shorter-chain alkanes can be components of fuels such as petrol. Alkenes are raw materials for polymers and other chemicals. Cracking therefore converts heavier fractions with more limited demand into more useful products in greater demand. Petrol is itself a mixture, so it has no single molecular formula representing the whole fuel.

Why crack? Convert plentiful large molecules into products with greater demand



The arrows show two types of use for products made from the feedstock. They do not mean that every molecule becomes all the products shown at once.

If asked for one reason, give a complete answer such as the products are more useful or the products are in greater demand. Saying only to make small molecules repeats the change without explaining its value. More specifically, you could mention producing more petrol components or alkenes for making polymers.

To assess a product combination, total the atoms in every molecule. Consider this cracking reaction:

$C_8H_{18} \rightarrow C_5H_{12}$ + unknown product

The reactant contains 8 C and 18 H atoms. The known product accounts for 5 C and 12 H. All remaining atoms must enter the unknown product, so subtract separately:

$$\text{Remaining C atoms} = 8 - 5 = 3$$

$$\text{Remaining H atoms} = 18 - 12 = 6$$

The unknown product can therefore be C_3H_6. If a product has a coefficient of 2, count every type of atom in that molecule twice.

Find a missing product: subtract C, then H, then check its plausibility

$$C_8H_{18} \rightarrow C_5H_{12} + ?$$

Carbon conservation

$8 - 5 = 3$

3 more C atoms needed

Hydrogen conservation

$18 - 12 = 6$

6 more H atoms needed

Combine the counts

C_3H_6

Fits the alkene general formula

Check: 5 + 3 = 8 C; 12 + 6 = 18 H. Coefficients are not atomic subscripts.

All counts refer to the total atoms on each side of the equation. Subtract the atoms in the known product from those in the reactant to identify the unknown product.

Insert C_3H_6 and check the complete reaction:

$C_8H_{18} \rightarrow C_5H_{12} + C_3H_6$

The right-hand side has 5+3=8 C atoms and 12+6=18 H atoms, matching the left-hand side. Using C_3H_8 instead would keep the carbon count correct but raise the hydrogen count to 20, so it would be unacceptable. In a multiple-choice question, counting C and H in each product combination rules out smaller-looking molecules that fail to conserve atoms.

The two conditions for cracking are _____ and _____. The alkene products can be used to make _____. Of cracking and fractional distillation, the process that forms new substances is _____. A product combination must conserve both _____ and _____.

Answers: high temperature; a catalyst; polymers; cracking; carbon; hydrogen

Recognising the Question Identify the chemical conversion of long chains into shorter chains, then state high temperature and a catalyst. Link uses to fuel components or chemical feedstocks in greater demand. Total C and H for every product option, and pay attention to only.

Worked Example 1 (2025·May/June 0620/31 Q4(b)) · Level 2

(b) (i) Describe how long chain hydrocarbon molecules can be made into short chain molecules.

Include in your answer:

- the name of this process
- the conditions needed for this process to take place
- the products of this process.

(ii) State **one** reason why long chain hydrocarbon molecules are made into short chain hydrocarbon molecules.

Solution

(i) The question describes long-chain hydrocarbon molecules being converted into shorter-chain molecules, so the process is **cracking**. State two separate conditions: **high temperature** and **a catalyst**.

Cracking produces smaller molecules, which may include shorter-chain alkanes, alkenes and hydrogen. For this question's product mark, explicitly write **alkenes** or **hydrogen**. Merely repeating short-chain molecules does not identify a product type.

A complete answer is: Long-chain hydrocarbons are cracked using a high temperature and a catalyst, producing smaller molecules including alkenes.

(ii) The purpose concerns product uses and demand: **The products are more useful / in greater demand.** Examples include obtaining more fuel components and alkenes for making polymers.

The official marks for (i) are awarded separately for the process name, a catalyst, high temperature, and alkenes or hydrogen. Part (ii) separately credits products being more useful or in greater demand. Saying only that the chains become shorter does not explain why.

Practice

1. (2025·May/June 0620/11 Q36) · Level 2

Three statements about cracking of larger alkane molecules are listed.

1 Cracking produces petrol for cars.

2 Cracking only produces short-chain alkenes.

3 Cracking produces alkenes used to make polymers.

Which statements are correct?

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

2. (2025·May/June 0620/13 Q38) · Level 3

Which products can be formed by the cracking of one molecule of hexane, C_6H_{14}?

A C_4H_{10} and C_2H_4 only

B $C_{12}H_{26}$ and H_2 only

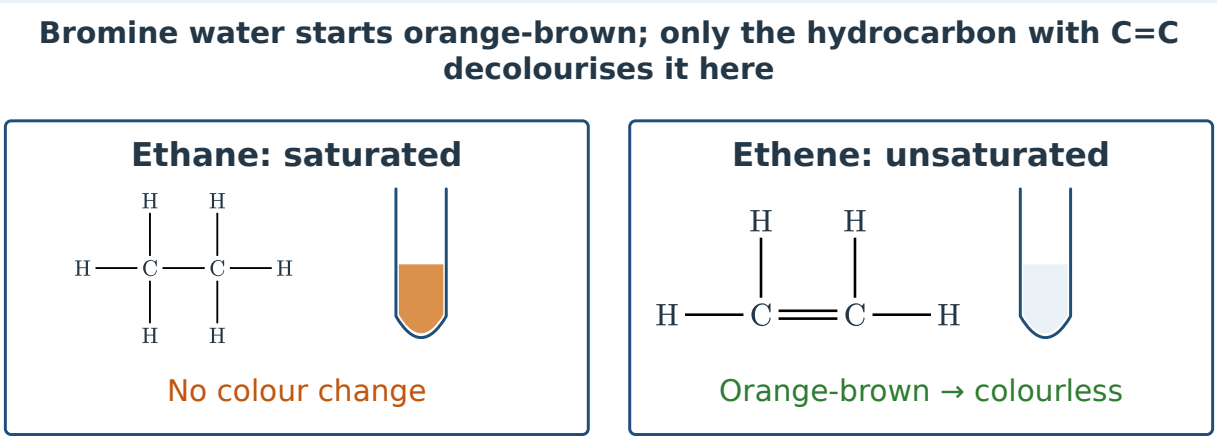
C C_3H_7 only

D C_2H_6 and C_4H_{10} only

Type 2: Draw addition products from C=C and predict the bromine-water observation

Key Ideas **Identify C=C before working out an alkene addition reaction.** Alkenes contain a carbon–carbon double covalent bond and are unsaturated hydrocarbons. An addition reaction adds atoms from the reagent to the original molecule. For each particular arrangement of connections, the reactants combine into one organic molecule without eliminating a small molecule such as HCl or water. Different arrangements can sometimes give products with different structures, as shown below. To draw a product, change the original C=C to C—C. Each carbon then has one available bond for a new atom or group.

Bromine water helps to detect this double bond. Aqueous bromine contains Br_2 and is orange-brown. Under normal test conditions, an alkene undergoes addition with bromine. Bromine is consumed and the solution changes from orange-brown to colourless. Saturated alkanes do not undergo this addition, so the colour remains. Colourless means having no colour; it does not mean white.

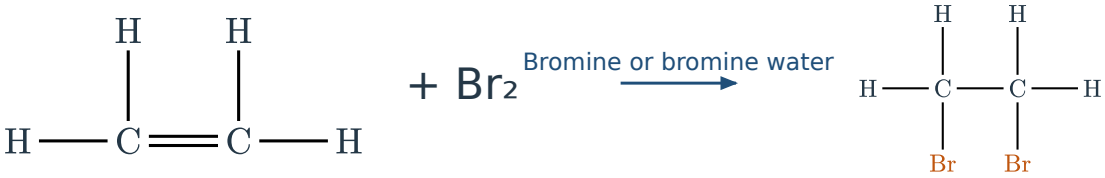


Compare the same bromine water with similar amounts of sample under normal test conditions. The test tubes indicate colour only, not particle counts or exact volumes.

Add the same bromine water to separate samples of CH_3CH_3 and $CH_2{=}CH_2$. First examine the carbon–carbon bonds: the former has only a single bond, while the latter has a double bond. The former leaves the colour unchanged; the latter decolourises the bromine water. Their shared C-and-H-only composition does not imply the same observation: this test distinguishes carbon–carbon bond types.

In bromine addition, one Br attaches to each end of the original double bond. Start with ethene, $CH_2{=}CH_2$. Retain every C and H atom and change the double bond to a single bond. Add one Br to each carbon to obtain CH_2BrCH_2Br, called 1,2-dibromoethane. Both Br atoms come from one Br_2 molecule, and the original two carbon atoms remain joined.

Bromine addition: one Br joins each carbon of the original double bond



$CH_2{=}CH_2 + Br_2 \rightarrow CH_2BrCH_2Br$

The double bond becomes single; each carbon gains a C—Br bond.

The complete displayed formulae on both sides retain the carbon framework. Reaction conditions appear above the arrow; new atoms in the product come from the added reagent.

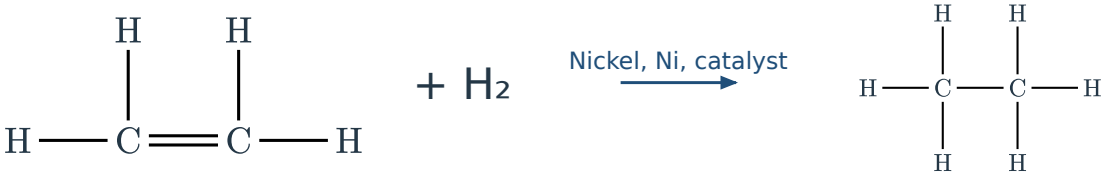
Check each C in the ethene–bromine product: it is joined to 2 H atoms, 1 Br and the other C, giving 4 bonds. The whole molecule, $C_2H_4Br_2$, has the same C and H counts as ethene, with Br_2 added. Keeping C=C while adding Br to both carbons would give carbon more than 4 ordinary covalent bonds. Removing H would incorrectly turn the drawing into substitution.

Hydrogen addition requires a nickel catalyst and produces an alkane. Hydrogen is H_2 and nickel has the symbol Ni. When ethene reacts with hydrogen over a nickel catalyst, the double bond becomes single and each carbon gains one H:

$CH_2{=}CH_2 + H_2 \rightarrow CH_3CH_3$

The product is ethane. Write nickel among the conditions; it does not become part of the organic product.

Hydrogen addition: C=C becomes C—C and each carbon gains one H



$CH_2{=}CH_2 + H_2 \rightarrow CH_3CH_3$

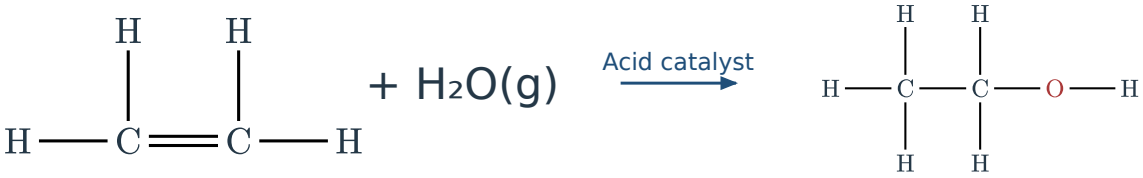
The product is ethane; no carbon or oxygen is added.

The complete displayed formulae on both sides retain the carbon framework. Reaction conditions appear above the arrow; new atoms in the product come from the added reagent.

When propene, $CH_2{=}CHCH_3$, adds hydrogen, change only the two double-bonded carbons. The left-hand CH_2 gains one H to become CH_3; the middle CH gains one H to become CH_2; the terminal CH_3 remains unchanged. The product, $CH_3CH_2CH_3$, is propane. The carbon count remains 3, while the hydrogen count increases from 6 to 8, matching the addition of H_2.

In steam addition, H and OH from water enter the molecule to form an alcohol. Steam is water vapour, $H_2O(g)$. Ethene reacts with steam in the presence of an acid catalyst to form ethanol, CH_3CH_2OH. When drawing the overall product, treat water as supplying one H and one OH. The CH_2 at one end of the double bond gains H to become CH_3; the other CH_2 gains OH; and the carbon–carbon double bond becomes single.

Steam addition: H and OH join the two carbons of the original C=C



$$CH_2{=}CH_2 + H_2O \rightleftharpoons CH_3CH_2OH$$

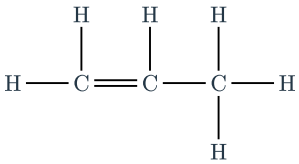
The product is ethanol; this tracks the overall atoms, not the detailed reaction mechanism.

The complete displayed formulae on both sides retain the carbon framework. Reaction conditions appear above the arrow; new atoms in the product come from the added reagent.

Ethene begins with 2 C and 4 H atoms. Adding one water molecule supplies 2 H and 1 O, so the product should contain 2 C, 6 H and 1 O atoms. CH_3CH_2OH matches exactly: each C has four bonds, and O has two, joining C and H. Ethane contains no O, ethanoic acid has one more O than this product, and poly(ethene) is not the product of this hydration step. Similar names do not make any of them substitutes for ethanol.

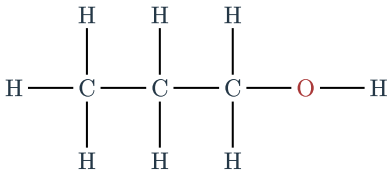
If the two ends of the original double bond differ, reversing H and OH may give different products. Propene is $CH_2{=}CHCH_3$. One end of its double bond is a terminal CH_2; the other is CH attached to CH_3. Putting OH on the terminal carbon and H on the middle carbon gives $CH_2OHCH_2CH_3$, propan-1-ol. Swapping H and OH gives $CH_3CH(OH)CH_3$, propan-2-ol. Number from the end that gives OH the lower position number.

Propene's ends differ: OH can join either carbon of the original C=C



+ H_2O(g), acid catalyst; original C=C becomes C—C

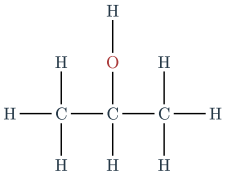
OH on the terminal C; H on the other C



Propan-1-ol

$CH_3CH_2CH_2OH$

H on the terminal C; OH on the other C



Propan-2-ol

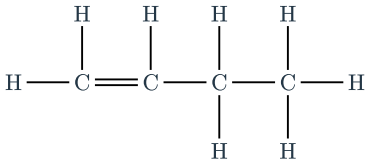
$CH_3CH(OH)CH_3$

The two branches show two possible arrangements of atom connections. They do not imply equal amounts of products, and the syllabus does not require a mechanism to predict their proportions. Every atom and bond in each product is shown.

Check the middle C in the right-hand structure: it is joined to the left C, the right C, H and O, giving four bonds; O is also joined to H. Both structures have the molecular formula C_3H_8O, but OH is attached to a different carbon. Flipping the whole drawing of propan-1-ol left to right leaves its connections unchanged and does not provide a second answer. In the second structure, OH must be attached to the middle carbon. If asked for two possible products, give these two; this does not claim that they necessarily form in equal amounts.

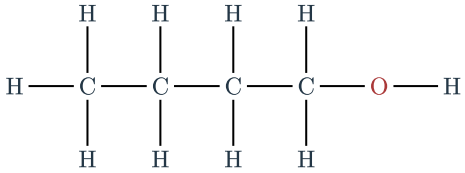
For a longer carbon chain, change only the two ends of the original double bond. But-1-ene is $CH_2=CHCH_2CH_3$. Retain the connections between all four carbon atoms, change the double bond between the leftmost two carbons to a single bond, and then place H and OH in the two possible ways. Keep the original CH_2CH_3 group on the right intact. The products are butan-1-ol and butan-2-ol.

But-1-ene plus steam: keep the carbon skeleton and place H and OH



+ $H_2O(g)$, acid catalyst; original C=C becomes C—C

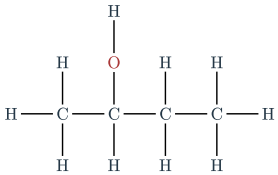
OH on the terminal C; H on the other C



Butan-1-ol

$CH_3CH_2CH_2CH_2OH$

H on the terminal C; OH on the other C



Butan-2-ol

$CH_3CH(OH)CH_2CH_3$

The two branches show two possible arrangements of atom connections. They do not imply equal amounts of products, and the syllabus does not require a mechanism to predict their proportions. Every atom and bond in each product is shown.

When OH attaches to carbon 1, the end becomes CH_2OH and carbon 2 becomes CH_2. When OH attaches to carbon 2, carbon 1 becomes CH_3 and carbon 2 becomes CH(OH). Each product contains 4 C, 10 H and 1 O atoms, agreeing with C_4H_8 plus H_2O. Do not place OH on the original carbon 3 without accounting for a change in chain numbering. Track the original double bond while drawing, then renumber from the appropriate end when naming the product.

All three reactions retain the carbon framework, but add different atoms or groups. Bromine supplies two Br atoms, hydrogen supplies two H atoms, and steam supplies H and OH. The two Br atoms must attach to the two neighbouring carbons of the original double bond. Adding one Br and one H represents a different reagent combination, not Br_2. After completing a displayed formula, check the bond count of each C, H, O and Br atom, then the total atom counts. Do not swap the nickel catalyst for hydrogen addition with the acid catalyst for steam addition.

Alkenes contain _____. Bromine water changes from _____ to _____. Ethene adds H_2 using a _____ catalyst to form _____. Ethene adds steam using an _____ catalyst to form _____.

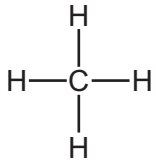
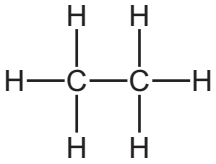
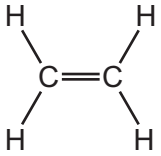
Answers: C=C; orange-brown; colourless; nickel, Ni; ethane; acid; ethanol

Recognising the Question Locate the original C=C and change it to C—C. Add Br and Br, H and H, or H and OH to its two carbon atoms. Retain the original carbon and hydrogen framework and check four bonds around every C. For an observation question, also state the corresponding bromine-water colour change.

Worked Example 1 (2025·May/June 0620/12 Q34) · Level 2

Three hydrocarbons are separately mixed with aqueous bromine.

Which row gives the correct observation for each hydrocarbon?

			
A	colour changes from orange to colourless	colour stays unchanged	colour changes from orange to colourless
B	colour changes from orange to colourless	colour changes from orange to colourless	colour stays unchanged
C	colour stays unchanged	colour stays unchanged	colour changes from orange to colourless
D	colour stays unchanged	colour changes from orange to colourless	colour stays unchanged

Solution

Read the three structures by column. The first has one C joined to four H atoms and is methane. The second has two C atoms joined by a single bond and is ethane. The third has a double bond between its two C atoms and is ethene.

Methane and ethane are saturated hydrocarbons and do not decolourise bromine water under normal test conditions. Ethene contains C=C and adds bromine, changing the colour from orange to colourless.

The three columns should therefore read no change, no change, and orange to colourless. Only the whole of row C matches, so choose **C**. One correct entry is not enough to establish a correct row, and colourless does not mean white.

Worked Example 2 (2023·May/June 0620/42 Q5(c)) · Level 3

(c) Propene undergoes addition reactions with steam.
There are two possible products, **A** and **B**.

Draw the displayed formula and name each product.

displayed formula of product **A**

name of product **A**

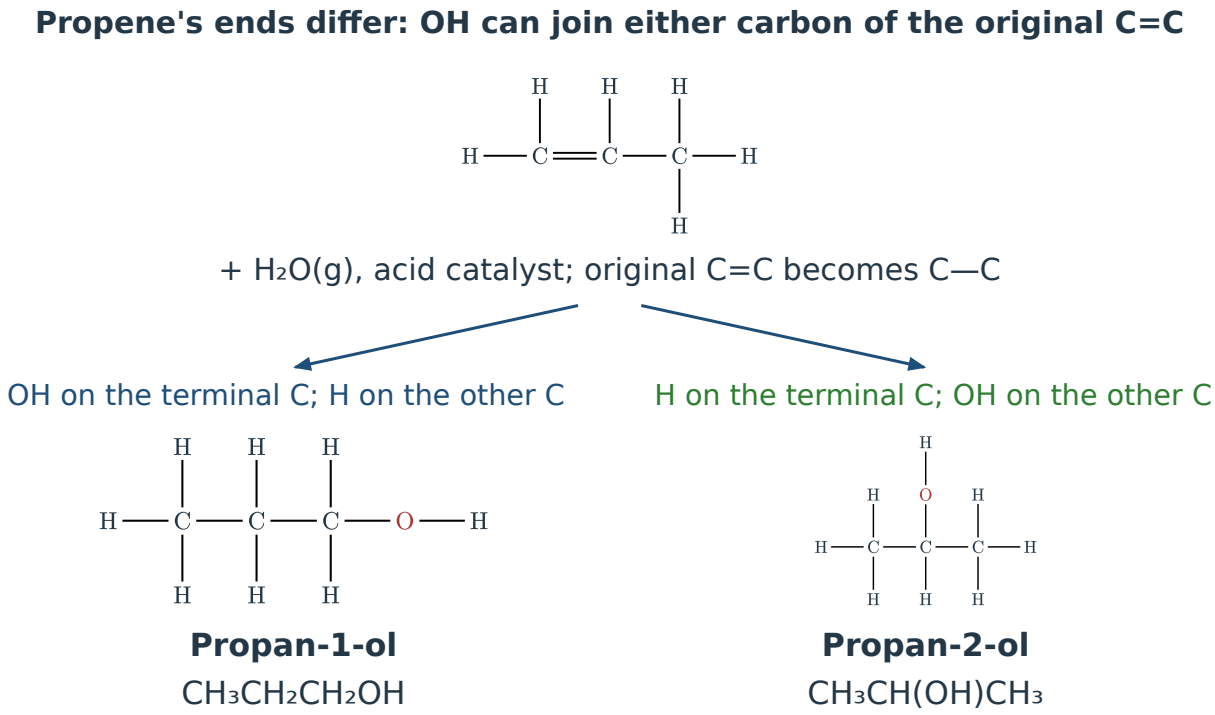
displayed formula of product **B**

name of product **B**

Solution

Propene has the structure $CH_2{=}CHCH_3$. Steam supplies H and OH. First change C=C to C—C while keeping the original three carbon atoms joined in order. The ends of the original double bond differ, so there are two possible placements.

First, place OH on the terminal C and H on the middle C, giving $CH_2OHCH_2CH_3$. Numbering from the end nearer OH gives **propan-1-ol**. Second, place H on the terminal C and OH on the middle C, giving $CH_3CH(OH)CH_3$, **propan-2-ol**. The complete displayed formulae are:



The two branches show two possible arrangements of atom connections. They do not imply equal amounts of products, and the syllabus does not require a mechanism to predict their proportions. Every atom and bond in each product is shown.

Draw one structure in each box and write the matching name below it. Their order in A and B may be reversed. Every C should have four bonds, O two and H one; both formulae are C_3H_8O. The official 4 marks cover the two complete displayed formulae and the two correctly matched names. Condensed structural formulae do not satisfy a request for displayed formulae. Drawing the same alcohol reversed does not give two products.

Practice

1. (2025·May/June 0620/12 Q35) · Level 1

Which compound is formed when ethene reacts with steam?

A ethane

B ethanoic acid

C ethanol

D poly(ethene)

2. (2025·May/June 0620/43 Q1(h)) · Level 1

A list of substances is shown.

aluminium oxide **calcium oxide** **chlorine** **ethanol** **graphite**

nitrogen **oxygen** **propane** **propene** **silicon(IV) oxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(h) is manufactured by the catalytic addition of steam to ethene

3. (2024·February/March 0620/42 Q5(d)(ii)) · Level 3

(d) When but-1-ene reacts with steam, **two** possible products form.

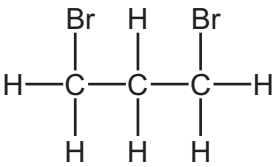
(ii) Name and draw the displayed formulae of the **two** possible products.

product 1	product 2
name	name .....
displayed formula	displayed formula

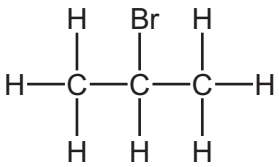
4. (2025·October/November 0620/23 Q36) · Level 3

Propene undergoes addition reactions with bromine, hydrogen and steam.

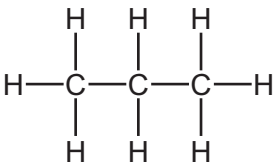
Which structures represent a product of these reactions?



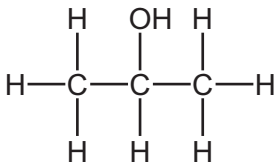
1



2



3



4

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 3 and 4

11.6 Alcohols

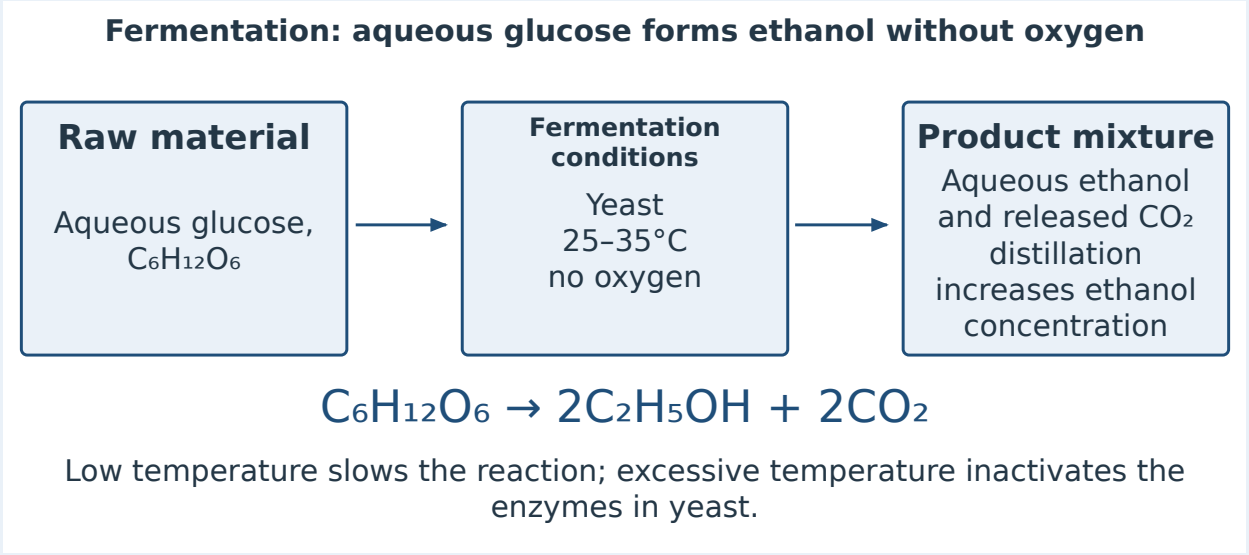
Type 1: Compare fermentation and ethene hydration: conditions, advantages, disadvantages and uses

Key Ideas **Both routes make ethanol; first distinguish their raw materials and conditions.** Ethanol has the structure CH_3CH_2OH, also commonly written C_2H_5OH. The first route is fermentation. An aqueous solution of glucose is kept at 25—35°C with yeast and in the absence of oxygen, converting glucose into ethanol and carbon dioxide.

Glucose has the molecular formula $C_6H_{12}O_6$. The overall fermentation equation is:

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

Of the 6 C atoms in each glucose molecule, 4 enter the ethanol molecules and 2 enter CO_2. All 12 H atoms are accounted for in the two ethanol molecules, and O is also conserved. Yeast provides enzymes that act as catalysts; do not treat yeast as a reactant consumed in a fixed ratio in the equation.



Arrows represent production steps or the overall reaction. Glucose comes from renewable plant material. Yeast is not shown as a reactant that is consumed.

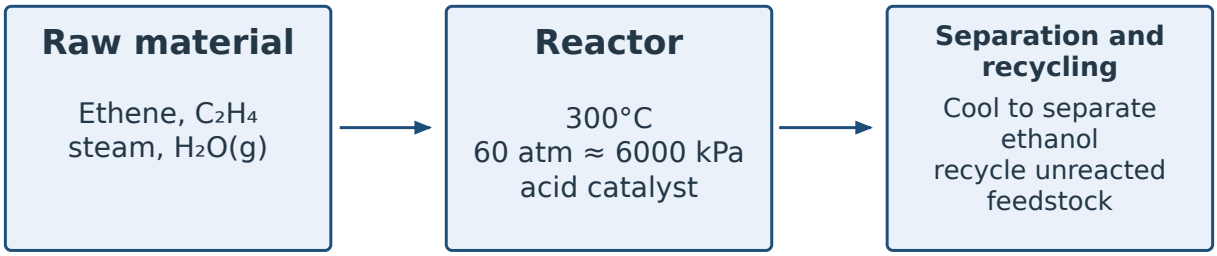
A typical fermentation temperature is 30°C, within the range 25—35°C. A temperature of 100°C is unsuitable: excessive heat inactivates the enzymes, preventing the yeast from fermenting effectively by this route. If the temperature is too low, the reaction is slow. Also state absence of oxygen. Introducing oxygen may change the process carried out by the microorganisms, so this condition must not be omitted.

The second route is catalytic hydration of ethene. Typical industrial conditions for ethene reacting with steam are 300°C, 6000 kPa or 60 atm, and an acid catalyst, commonly exemplified by phosphoric acid, H_3PO_4. The overall equation is:

$$C_2H_4 + H_2O \rightleftharpoons C_2H_5OH$$

Ethene comes from the petrochemical route studied in the syllabus. Steam is a reactant; yeast is not used in this reactor.

Ethene and steam: high temperature, elevated pressure and an acid catalyst



$$C_2H_4 + H_2O \rightleftharpoons C_2H_5OH$$

Phosphoric acid, H_3PO_4, is commonly used; steam is a reactant and yeast is not used in this route.

Arrows show where materials flow in the process. The reversible symbol refers to the reaction itself; raw materials not converted in one pass can be recycled.

For catalytic steam-addition conditions, give 300°C, 60 atm and an acid catalyst separately. Do not substitute fermentation's 30°C and yeast, or the Haber process conditions of 450°C and 200 atm. The reaction is reversible, so the reactants need not be completely converted in one pass through the reactor. After separating ethanol, unreacted materials can be recycled.

Compare advantages and disadvantages one criterion at a time. Fermentation uses sugars from plants, so its raw material is renewable, and its reaction temperature is lower. However, the reaction is slower, production is usually in batches, and the resulting aqueous ethanol is dilute. Ethene hydration is faster, can operate continuously and gives a purer product. Its costs include the energy and equipment required for higher temperature and pressure. Ethene from the petroleum route specified in this syllabus is also non-renewable.

Compare the two ethanol routes using matching criteria

	Fermentation	Ethene + steam
Raw material	Glucose: renewable plant feedstock	Ethene: non-renewable petroleum route in this course
Rate and production mode	Slower; usually batch	Faster; can be continuous
Reaction conditions	25–35°C; yeast; no oxygen	300°C; 60 atm; acid catalyst
Ethanol obtained	Dilute aqueous solution; further separation needed	Purer product; separation still needed
Energy and equipment	Lower reaction temperature; distillation still needs energy	Higher temperature and pressure require more energy and demanding equipment

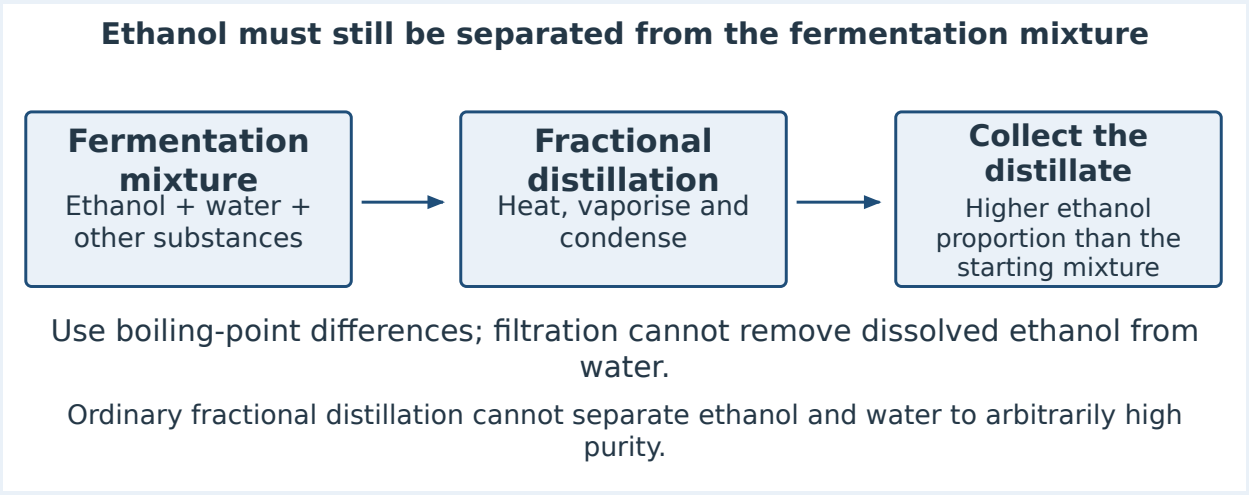
Purer compares the two routes; it does not mean 100% ethanol is obtained without separation.

Each row compares the same criterion. Raw-material renewability is assessed for the syllabus routes using plant-derived glucose and petroleum-derived ethene.

If asked for two advantages of ethene hydration, give a faster reaction and a purer product. Uses more energy describes a cost, so it is not an advantage even if the statement is true. Renewable raw material belongs to the

glucose route in this comparison. Purity and yield also differ: purity concerns the proportion of impurities, whereas yield concerns how much desired product is obtained. They are not interchangeable.

Fermentation leaves a mixture that still needs separation. Ethanol and water are miscible. Filtration cannot trap dissolved ethanol as it would sand. Fractional distillation uses their different boiling points: heat the mixture to produce vapour, then condense and collect a fraction richer in ethanol. Separation also requires energy, so a lower fermentation temperature does not mean that the whole production chain uses no energy.

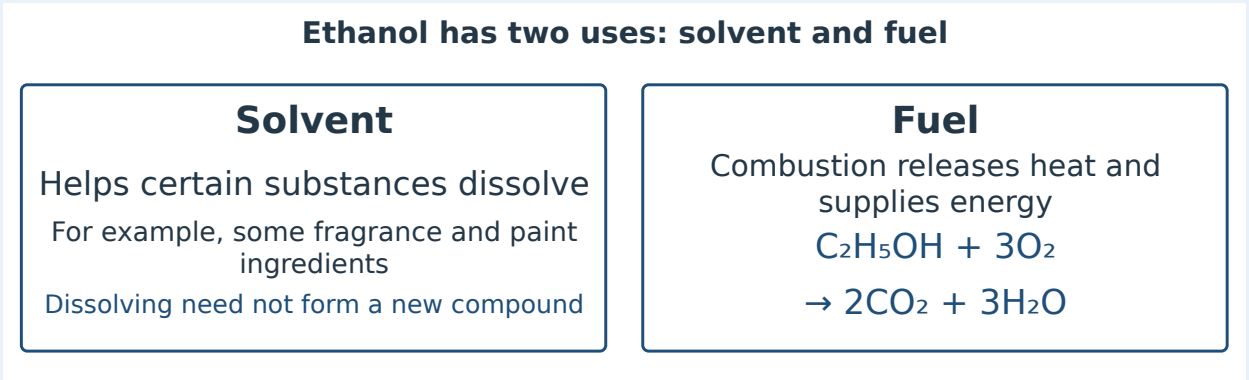


Arrows represent separation steps. Ethanol and water are miscible and cannot be separated by filter paper as though they were an insoluble solid and a liquid.

A fermentation mixture contains water, ethanol and other substances. Filtering insoluble material removes those solids, but does not also produce pure ethanol; further separation by boiling point is required. In a comparison, purer product is sufficient. Do not exaggerate this to 100% ethanol without treatment: ordinary fractional distillation also has limits on the purity it can achieve.

Ethanol is used as both a solvent and a fuel. A solvent is a liquid that helps other substances dissolve. Ethanol can dissolve ingredients in some perfumes, paints and related products. A fuel releases useful energy when burned. Complete combustion of ethanol produces CO_2 and H_2O:

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$



The left-hand side describes a use in dissolving substances. The right-hand side shows the chemical change of complete combustion, which requires sufficient oxygen.

To balance ethanol's complete combustion, 2 C atoms first give $2CO_2$ and 6 H atoms give $3H_2O$. The products contain 7 O atoms in total. The fuel already contains 1 O, leaving 6 to be supplied by $3O_2$, which balances the equation. As a solvent, ethanol helps substances disperse and dissolve without requiring this combustion reaction. Different uses of the same substance depend on different properties.

Fermentation uses an aqueous solution of _____ at _____, with _____ and excluding _____. Typical ethene hydration conditions are a temperature of _____, a pressure of _____ and an _____ catalyst. Two uses of ethanol are as a _____ and a _____.

Answers: glucose; 25—35°C; yeast; oxygen; 300°C; 60 atm or 6000 kPa; acid; solvent; fuel

Recognising the Question Identify fermentation or catalytic steam addition, then match its temperature, raw materials, catalyst and oxygen conditions. Assess each proposed advantage separately; purity, speed, renewability and energy use describe different criteria.

Worked Example 1 (2025·May/June 0620/21 Q35) · Level 2

Industrially, ethanol is produced by fermentation or by the reaction of ethene with steam.

Which row is correct?

	fermentation	ethene + steam
A	uses a temperature of 100 °C	uses a temperature of 350 °C
B	needs the presence of yeast	does **not** need a catalyst
C	slow reaction	fast reaction
D	high yield of ethanol	low yield of ethanol

Solution

Check both columns against the same comparison criterion. A gives fermentation at 100°C, outside the typical 25—35°C range. It also gives ethene steam addition at 350°C instead of the syllabus value of 300°C, so exclude A.

In B, fermentation needing yeast is correct, but ethene hydration also needs an acid catalyst, so does not need a catalyst is false. C describes fermentation as slower and ethene hydration as faster, correctly matching the two processes.

D is not a valid general comparison of the overall yields of these industrial routes. Although reversible hydration does not give complete conversion in one pass, unreacted materials can be recycled. This does not directly imply a low final ethanol yield. Nor should the low ethanol concentration in fermentation mixtures be confused with yield. Choose **C**, with slow reaction and fast reaction in the respective columns.

Practice

1. (2025·May/June 0620/41 Q1(c)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(c) is manufactured by fermentation

2. (2025·May/June 0620/42 Q1(e)) · Level 1

Using numbers only, state the:

(e) typical temperature for the fermentation of aqueous glucose in °C

3. (2025·May/June 0620/22 Q36) · Level 2

Ethanol can be made by the fermentation of aqueous glucose and by the catalytic addition of steam to ethene.

What are two advantages of making ethanol by the catalytic addition of steam to ethene rather than by the fermentation of aqueous glucose?

A faster reaction and renewable raw materials

B purer product and faster reaction

C renewable raw materials and continuous process

D uses more energy and forms a purer product

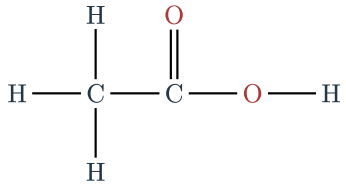
11.7 Carboxylic acids and esters

Type 1: Determine products, observations and quantities in organic reactions

Key Ideas **Organic reactions still follow the acid chemistry and conservation of atoms already studied.**

Ethanoic acid has the structural formula CH_3COOH and the functional group COOH. It is a weak acid: only some molecules ionise in water to form H^+ and CH_3COO^-, the ethanoate ion. A weak acid can still be clearly acidic. An acidic solution has a pH below 7 and turns blue litmus red.

Ethanoic acid solution is acidic; one H in COOH can be neutralised



$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

Weak acid: only some molecules ionise

Acidic solution: pH < 7; blue litmus turns red

Strength describes ionisation; concentration describes solute per unit volume. Keep these separate.

COOH in the complete displayed formula on the left corresponds to the source of acidity on the right. The reversible arrow indicates partial ionisation, not complete conversion of all molecules into ions.

An aqueous solution of an organic compound with pH 3 that releases CO_2 with a carbonate shows acid reactions. Of the alkane, alkene, alcohol and carboxylic acid families in this section, a carboxylic acid fits. Being liquid and soluble in water alone does not uniquely identify the family; use the more direct evidence of acidity. Weak refers to partial ionisation and does not justify a pH of 10.

Acids give different products with metals, alkalis and basic oxides. Ethanoic acid reacts with magnesium, Mg, to form magnesium ethanoate and hydrogen. With sodium hydroxide, NaOH, it forms sodium ethanoate and water. With copper(II) oxide, CuO, it forms copper(II) ethanoate and water. In the salt name, ethanoate comes from ethanoic acid and the metal name comes from the other reactant.

The ethanoate ion, CH_3COO^-, has one negative charge. Na^+ therefore pairs with one ethanoate ion to give CH_3COONa. Mg^{2+} or Cu^{2+} needs two ethanoate ions, giving $(CH_3COO)_2Mg$ or $(CH_3COO)_2Cu$. The 2 outside the brackets counts the whole ethanoate ion twice; do not apply it only to the final O or to the metal.

Identify the other reactant to decide what forms alongside the salt

Magnesium, Mg: reactive metal

$$2CH_3COOH + Mg \rightarrow (CH_3COO)_2Mg + H_2$$

Salt: magnesium ethanoate; hydrogen also forms

Sodium hydroxide, NaOH: alkali

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

Salt: sodium ethanoate; water also forms

Copper(II) oxide, CuO: basic oxide

$$2CH_3COOH + CuO \rightarrow (CH_3COO)_2Cu + H_2O$$

Salt: copper(II) ethanoate; water also forms

The 2 outside the brackets means two ethanoate ions, CH_3COO^-. Mg^{2+} or Cu^{2+} needs two ions of charge −1 to balance its charge.

First write the reaction with Mg. Each Mg forms Mg^{2+} and requires two CH_3COO^- ions, so use $2CH_3COOH$. The two acidic H atoms form one H_2 molecule. The complete equation is:

$2CH_3COOH + Mg \rightarrow (CH_3COO)_2Mg + H_2$

If the other reactant is changed to NaOH, determine the products again. OH^- combines with the acidic H to form H_2O:

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

You must therefore not retain H_2 as a gaseous product.

An acid reacting with a carbonate also produces carbon dioxide. Examples of carbonates include sodium carbonate, Na_2CO_3, and calcium carbonate, $CaCO_3$. With ethanoic acid, they produce an ethanoate salt, water and CO_2. Bubbles are observed. Passing the gas through clear limewater and observing cloudiness provides further evidence that it is CO_2.

Acid and carbonate: carbon dioxide is a key identifying clue

$$2CH_3COOH + Na_2CO_3$$

$$\rightarrow 2CH_3COONa + H_2O + CO_2$$

Ethanoic acid + sodium carbonate → sodium ethanoate + water + carbon dioxide

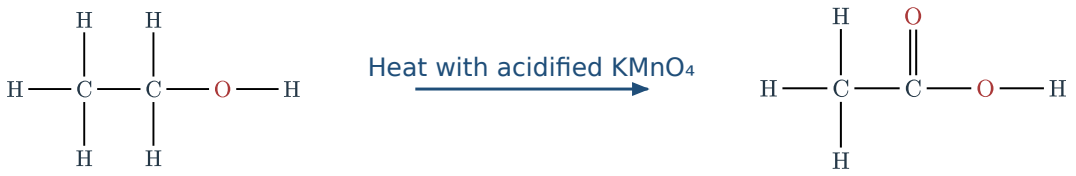
Bubbles are seen; the CO_2 turns limewater milky.

The arrow represents an acid–carbonate reaction. Bubbles show that gas is escaping; the limewater test provides further evidence that the gas is CO_2.

For Na_2CO_3, the two Na^+ ions require two ethanoate ions, so $2CH_3COOH$ gives $2CH_3COONa$. The two acidic H atoms and one O from carbonate account for H_2O; the remaining C and O account for CO_2. With $CaCO_3$, the salt becomes $(CH_3COO)_2Ca$, while the other products remain H_2O and CO_2. Do not apply the acid–metal reaction that produces H_2 directly to a metal carbonate.

Oxidising ethanol can produce ethanoic acid; match the reagent to its colour change. Heating ethanol with acidified aqueous potassium manganate(VII), $KMnO_4$, oxidises ethanol to ethanoic acid. Ethanol undergoes oxidation, while the oxidising agent's purple colour fades to colourless. A decrease in reagent colour does not mean that ethanol undergoes reduction.

Ethanol oxidises to ethanoic acid as the purple oxidising agent decolourises



Ethanol, CH_3CH_2OH Ethanoic acid, CH_3COOH

$$CH_3CH_2OH + 2[O] \rightarrow CH_3COOH + H_2O$$

Reaction: ethanol oxidation; observation: purple → colourless solution.

[O] is the syllabus shorthand for the oxidising action supplied by the oxidising agent; it does not mean that the reagent bottle contains free oxygen atoms. $KMnO_4$ is potassium manganate(VII).

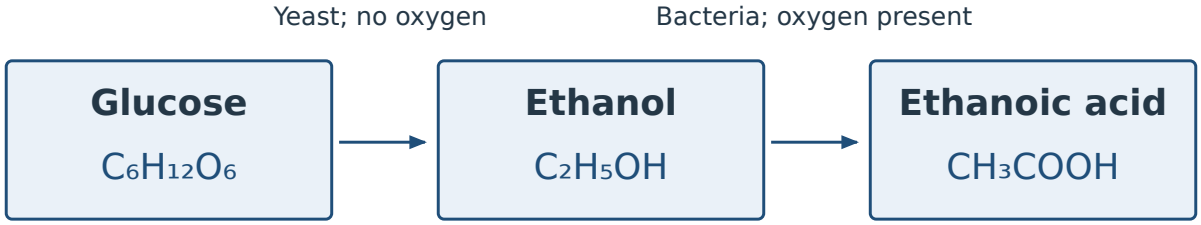
Ethanol, CH_3CH_2OH, becomes CH_3COOH. Its two-carbon framework remains, while the structure gains bonding to oxygen, showing that ethanol is oxidised. For reaction type and observation, give oxidation; purple to colourless. [O] is shorthand for the action supplied by the oxidising agent. You do not need to expand every ionic reaction of $KMnO_4$ to answer this question.

Vinegar production is another oxidation of ethanol and requires oxygen. Certain bacteria oxidise ethanol to ethanoic acid in the presence of oxygen, a process used to make vinegar. Oxygen, O_2, is a reactant here. The overall change can be written as:

$$C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$$

This is a different process from the anaerobic fermentation in which yeast converts glucose into ethanol.

Fermentation and vinegar production use different microorganisms and oxygen conditions



Vinegar production: $C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$

The arrows on the left and right represent two different processes. Bacterial oxidation in vinegar production needs oxygen; do not transfer the oxygen-free condition from fermentation.

For sugar becoming ethanol, give fermentation with yeast at 25—35°C without oxygen. For ethanol becoming ethanoic acid in vinegar production, give bacterial oxidation with oxygen. Both involve microorganisms, but their oxygen conditions differ.

For complete combustion of an organic compound, balance carbon first, hydrogen next and oxygen last. Ethene, C_2H_4, and ethanol, C_2H_5OH, both burn, but only ethanol contains oxygen in the fuel itself. With sufficient oxygen, C enters CO_2 and H enters H_2O. When calculating the required O_2, subtract the O atoms already present in the fuel molecules.

Organic combustion: when balancing oxygen, include the O already in the fuel

Ethene, C_2H_4 $2CO_2 + 2H_2O$ Needs $3O_2$

Products need 6 O atoms; fuel contains 0

Ethanol, C_2H_5OH $2CO_2 + 3H_2O$ Also needs $3O_2$

Products need 7 O atoms; fuel already contains 1

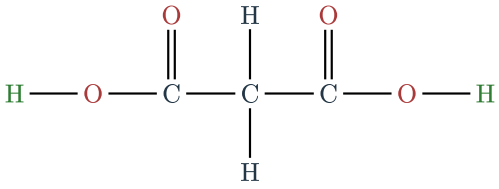
Use C and H to set the products, then subtract the fuel's O atoms from the product total to find O_2.

Both equations compare complete combustion per 1 mol of fuel. The product formulae CO_2 and H_2O remain fixed; only coefficients are adjusted.

One ethene molecule has 2 C and 4 H atoms, giving $2CO_2$ and $2H_2O$, which require 6 O atoms. Ethene contains no O, so use $3O_2$. Ethanol has 2 C and 6 H atoms, giving $2CO_2$ and $3H_2O$, which require 7 O atoms. Ethanol already supplies 1, so again use $3O_2$. Both fuels happen to require $3O_2$, but their water coefficients differ; calculate each separately.

Each COOH group in a molecule provides one acidic H that NaOH can neutralise. The O—H in one COOH supplies one acidic H. The OH^- from one NaOH neutralises it to form one H_2O. If a molecule contains two COOH groups, complete neutralisation requires two NaOH for each acid molecule. Ordinary H atoms attached to C do not count in this carboxylic-acid neutralisation ratio.

Two COOH groups: one acid molecule needs two OH^- for neutralisation



$HOOC—CH_2—COOH + 2NaOH$

$\rightarrow NaOOC—CH_2—COONa + 2H_2O$

Green H atoms are the acidic O—H hydrogens; NaOH does not neutralise the central C—H hydrogens this way.

1 mol of this dicarboxylic acid reacts with 2 mol of NaOH; the ratio comes from the number of COOH groups.

$HOOC—CH_2—COOH$ contains two carboxyl groups. The green H at each end marks one of the two acidic hydrogens that can be neutralised. For compound B, count the carboxyl groups again using the original question's structure.

Suppose there are 0.020 mol of the dicarboxylic acid shown. Each molecule contains two COOH groups, so complete neutralisation requires twice as many moles of NaOH:

$$n(\text{NaOH}) = 2 \times 0.020 = 0.040 \quad \text{mol}$$

If the NaOH concentration is 0.100 mol/dm³, divide the amount in moles by the concentration to find the solution volume. Both n and c here refer to NaOH:

$$V = \frac{n}{c} = \frac{0.040}{0.100} = 0.400 \quad \text{dm}^3$$

Finally, convert to cm³:
$0.400 \times 1000 = 400 \quad \text{cm}^3$
Count COOH groups first, calculate the amount of NaOH, then divide by concentration and convert the volume unit. Dividing the acid amount directly by the NaOH concentration would omit the factor of 2.

Ethanoic acid produces _____ with Mg and _____ with a carbonate. With NaOH, it forms _____ and _____. When acidified $KMnO_4$ oxidises ethanol, the colour changes from _____ to _____. Complete neutralisation of 1 mol of an acid containing two COOH groups requires _____ mol of NaOH.

Answers: H_2; CO_2; sodium ethanoate, CH_3COONa; water, H_2O; purple; colourless; 2

Recognising the Question Identify whether the other reactant is a metal, base or carbonate. In oxidation questions, track ethanol's change separately from the reagent's colour. For calculations, count COOH groups to establish the NaOH ratio, then use n=cV and convert units. Balance combustion by conserving C, H and O.

Worked Example 1 (2025·May/June 0620/12 Q36) · Level 2

Which statements about ethanoic acid are correct?

1 Its aqueous solution has a pH value of 10.

2 It reacts with metal carbonates to produce carbon dioxide gas.

3 It reacts with magnesium metal to produce hydrogen gas.

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

Solution

1: Aqueous ethanoic acid is acidic and should have a pH below 7. A pH of 10 is alkaline, so this is false.

2: An acid reacts with a metal carbonate to produce a salt, water and CO_2, so this is true. 3: Ethanoic acid reacts with a reactive metal such as Mg to form magnesium ethanoate and H_2, so this is true.

Statements 2 and 3 are correct, giving **D**. The two reactions produce different gases: the carbonate gives CO_2, while Mg metal gives H_2. Do not treat them as the same gas.

Worked Example 2 (2025·May/June 0620/42 Q6(e)(v)) · Level 3

(e) Compound **B** has the displayed formula shown in Fig. 6.1.

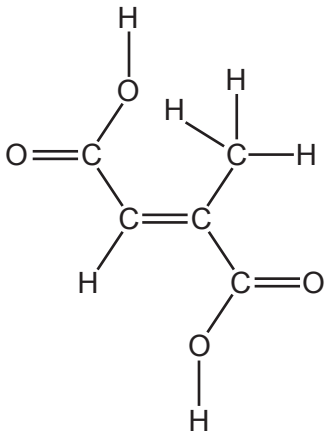


Fig. 6.1

(v) Calculate the volume, in cm^3, of $0.250\,mol/dm^3$ NaOH that reacts with 0.100 mol of compound **B**.

Solution

Locate the two COOH groups in Fig.6.1. Each contains one acidic H that needs one OH^- for neutralisation, so 1 mol of B requires 2 mol of NaOH. Other H atoms attached to C do not increase this neutralisation ratio.
First calculate the amount of NaOH, using two NaOH for the two COOH groups:

$$n(\text{NaOH}) = 2 \times 0.100 = 0.200 \quad \text{mol}$$

The given 0.100 mol is already the amount of B. Do not divide it by B's relative molecular mass again.
Second, use NaOH's own concentration to find its volume. Concentration is amount divided by volume, so volume is amount divided by concentration:

$$V = \frac{n}{c} = \frac{0.200}{0.250} = 0.800 \quad \text{dm}^3$$

Third, the question requires cm^3. Since 1 dm^3 equals 1000 cm^3:

$$0.800 \times 1000 = 800 \quad \text{cm}^3$$

The final answer is **800 cm^3**. The official two stages credit the correct 0.200 mol and the final 800 cm^3. Ignoring the two COOH groups and dividing 0.100 directly by the concentration would give half the required volume.

Practice

1. (2025·May/June 0620/21 Q36) · Level 2

Some properties of an organic compound, J, are listed.

- It is a liquid at room temperature.
- It is soluble in water.
- A solution of J reacts with calcium carbonate to form carbon dioxide.
- A solution of J has a pH of 3.

In which homologous series does J belong?

A alkane

B alkene

C alcohol

D carboxylic acid

2. (2025·May/June 0620/21 Q17) · Level 2

Ethanol is warmed with acidified aqueous potassium manganate(VII).

Which row describes the type of reaction of the ethanol and the colour change observed in the reaction mixture?

	type of reaction	colour change
A	oxidation	colourless to purple
B	oxidation	purple to colourless
C	reduction	colourless to purple
D	reduction	purple to colourless

3. (2025·May/June 0620/42 Q6(c)) · Level 2

Ethene, C_2H_4, is the first member of a family of similar compounds which contains the alkene functional group.

(c) Write the symbol equation for the complete combustion of C_2H_4.

4. (2025·May/June 0620/42 Q6(e)(iv)) · Level 2

(e) Compound **B** has the displayed formula shown in Fig. 6.1.

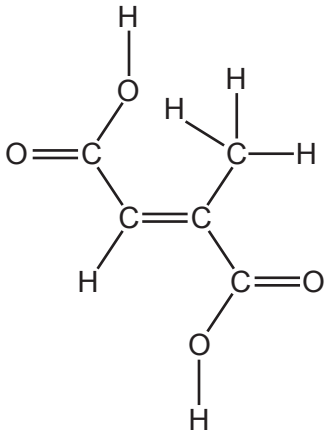


Fig. 6.1

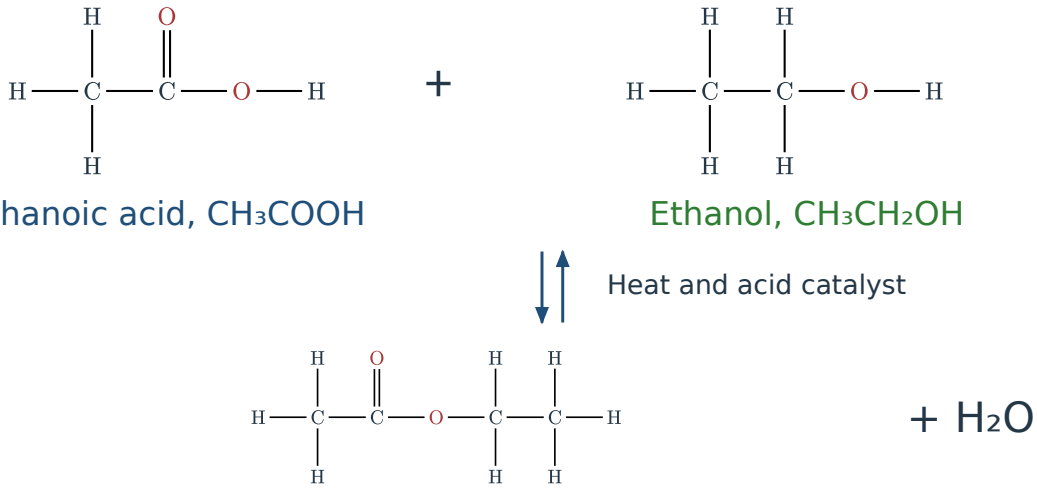
(iv) Explain why 1 mol of compound **B** reacts with 2 mol of sodium hydroxide, NaOH.

Type 2: Split ester names and draw every atom and the ester linkage

Key Ideas **Find the ester linkage, then connect it to the parent acid and alcohol.** The ester linkage studied here contains C(=O)—O—C: one O is attached beside C=O and is also joined to another carbon. A carboxylic acid ends in C(=O)—O—H. The distinction is whether the atom on the other side of O is C or H; two O atoms alone do not establish that a compound is an ester.

A carboxylic acid and an alcohol undergo esterification when heated with an acid catalyst, forming an ester and water. Concentrated sulfuric acid, H_2SO_4, is commonly used as the catalyst. The reaction is reversible, so use $\rightleftharpoons$. When constructing the product, retain the acid-side part containing C=O and the alcohol-side carbon chain, join them through C(=O)—O—C, and include H_2O. This drawing method checks the overall structure without requiring a detailed reaction mechanism.

Acid + alcohol form an ester: keep both carbon chains, form the ester linkage and release water



The two reactants above and the ester and water below conserve the total number of atoms. The two arrows pointing in opposite directions indicate a reversible reaction.

For example, ethanoic acid and ethanol undergo this esterification:

$$CH_3COOH + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + H_2O$$

The acid side retains $CH_3C(=O)—$, and the alcohol side supplies CH_2CH_3 after O, giving 2+2=4 carbon atoms in total. The product contains an ester linkage and is called ethyl ethanoate. The reaction produces both an ester and water, so it is not one of the simple addition reactions studied here, which form only one product.

An ester name puts the alcohol-derived group first and the acid-derived part second. The alcohol gives the alkyl name: methanol → methyl, ethanol → ethyl, propan-1-ol → propyl, and butan-1-ol → butyl. The carboxylic acid gives the second part: methanoic acid → methanoate, ethanoic acid → ethanoate, propanoic acid → propanoate, and butanoic acid → butanoate. Keep this order when constructing or interpreting the name.

Ester names: the alcohol-derived alkyl group comes before the acid-derived part



Butyl comes from butan-1-ol; ethanoate comes from ethanoic acid. Keep the two sources distinct.

The arrows trace each part of an ester name back to the corresponding reactant name. They do not represent experimental steps that separately convert the ester into these substances.

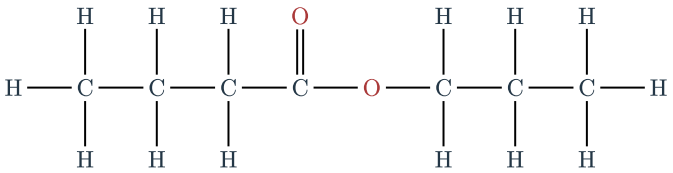
In ethyl propanoate, ethyl comes from ethanol and propanoate comes from propanoic acid, so the reactants are ethanol and propanoic acid. If asked for both reactants, enter each under its correct acid or alcohol heading. Do not assume that the first word of the ester name belongs in the first answer field.

The carbon counts in the two parts of an ester name add directly. The acid's carbon count includes the C in its carboxyl group. The alcohol-derived alkyl group contributes all of its carbon atoms separately. Esterification forms H_2O, but water contains no C, so losing water does not reduce the total carbon count.

Count each side's carbons separately: propyl butanoate has 3 + 4 carbons

Butanoate: 4 C from the acid Propyl: 3 C from the alcohol

$$CH_3CH_2CH_2C(=O)—O—CH_2CH_2CH_3$$



Water formation removes only H and O, not C; the total is 7 carbons.

The condensed structure above explicitly identifies C=O. The complete displayed formula below allows every atom to be checked. Coloured underlines show the origins of the two carbon-containing parts.

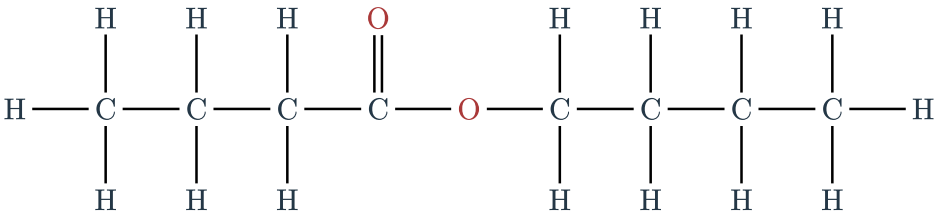
In methyl propanoate, methyl contributes 1 C and propanoate contributes 3 C, giving 4 C in total. Its structure is $CH_3CH_2COOCH_3$. The acid-derived part on the left contains three carbon atoms, including the C in COO; the alcohol-derived part on the right contains one. Do not count only the acid's hydrocarbon chain, CH_3CH_2, and omit its carboxyl carbon.

For a new combination, build the structure instead of memorising every complete formula. Choose one fragment from each of the two sets below. Form a single bond between the acid-side O and the first C of the alcohol-side fragment. Then check four bonds around every C, two around every O and one around every H. If a displayed formula is requested, expand CH_3, CH_2 and COO into all their atoms and bonds.

Combine acid and alcohol fragments with 1–4 carbons to form straight-chain esters

Acid side: retain COO—		Alcohol side: retain the carbon chain attached to O	
methanoate	HCOO—	methyl	—CH_3
ethanoate	CH_3COO—	ethyl	—CH_2CH_3
propanoate	CH_3CH_2COO—	propyl	—$CH_2CH_2CH_3$
butanoate	$CH_3CH_2CH_2COO$—	butyl	—$CH_2CH_2CH_2CH_3$

Choose any acid row and any alcohol row, then join with an O—C single bond; they need not be matching rows.



Example: butyl butanoate has 8 C atoms; the displayed formula above shows every atom and bond.

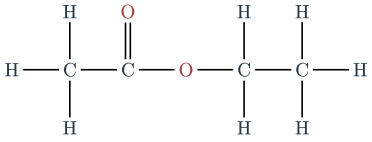
A line on a fragment represents a bond to be connected, not an extra atom. The complete structure uses a four-carbon acid and a four-carbon primary alcohol, covering the longest carbon counts in these combinations.

For butanoic acid and butan-1-ol, take $CH_3CH_2CH_2COO$— from the acid side and —$CH_2CH_2CH_2CH_3$ from the alcohol side. Joining them gives $CH_3CH_2CH_2COOCH_2CH_2CH_2CH_3$, named butyl butanoate. The acid and alcohol each supply 4 C atoms, giving 8 altogether; the complete displayed formula appears above. Use the same method for other combinations of 1—4 carbon atoms. The table does not restrict you to pairing fragments in the same row.

To work backwards from a total carbon count, first allow the minimum needed on each side of the ester linkage. The acid needs at least one carboxyl C, and the alcohol side also needs at least one C, so the simplest ester here has at least two carbon atoms. For three carbon atoms, distribute them between the acid and alcohol sides as 2+1 or 1+2, then construct and check each structure. A correct ester linkage alone is insufficient: draw all remaining C, H and O atoms and all bonds too.

Work backwards from an ester's carbon total: allow at least one C on each side

2 C in total	3 C in total	One option with 4 C in total
Acid: 1 C + alcohol: 1 C	Acid 1 C + alcohol 2 C, or acid 2 C + alcohol 1 C	Acid: 2 C + alcohol: 2 C



Four-carbon example: $CH_3COOCH_2CH_3$; the acid's two carbons already include the carboxyl carbon.

The upper part lists carbon-count distributions. Below, every atom and bond is shown for a four-carbon ester with a 2+2 distribution. This does not imply that only one four-carbon ester structure is possible.

To construct one four-carbon ester, choose 2 C on the acid side and 2 C on the alcohol side, giving the $CH_3COOCH_2CH_3$ already discussed. Check that the C in the acid-side COO is included in its 2 C, while CH_2CH_3 after O contributes another 2. Attaching H after O instead would produce a carboxylic acid, not an ester. Omitting C=O would leave the ester linkage incomplete.

Esterification requires a carboxylic acid and an _____, forming an ester and _____, usually with an _____ catalyst. Ethyl corresponds to _____, and butanoate to _____. Propyl butanoate contains _____ carbon atoms. The linking O must join the C of C=O to a _____ on the alcohol side.

Answers: alcohol; water; acid, such as concentrated sulfuric acid; ethanol; butanoic acid; 7; carbon atom

Recognising the Question Use the first part of an ester name to identify the alcohol and the second to identify the acid; add their carbon counts. Draw C(=O)—O—C first, attach the carbon chains and complete the H atoms. Check every atom, and do not substitute a condensed structural formula for a displayed formula.

Worked Example 1 (2025·May/June 0620/41 Q5(d)(ii)) · Level 2

(d) Carboxylic acids react with alcohols to form esters.

(ii) Butyl ethanoate is an ester.

Name the alcohol and the carboxylic acid that react to produce butyl ethanoate.

alcohol

carboxylic acid

Solution

Split butyl ethanoate into two parts. Butyl comes from butan-1-ol and specifies the alcohol-side straight four-carbon chain. Ethanoate comes from ethanoic acid.

The first field asks for the alcohol, so write **butan-1-ol**. The second asks for the carboxylic acid, so write **ethanoic acid**. Follow the field headings when ordering your answers.

Check with the structures: CH_3COOH and $CH_3CH_2CH_2CH_2OH$ form $CH_3COOCH_2CH_2CH_2CH_3$ and H_2O. Butanoic acid with ethanol would instead give ethyl butanoate. Although the total carbon count is the same, the carbon chains on the two sides of the linkage come from different reactants.

Worked Example 2 (2025·May/June 0620/43 Q5(d)(i)) · Level 3

(d) Carboxylic acids react with alcohols to form esters.

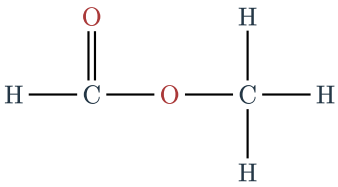
(i) Draw the displayed formula of the ester which contains two carbon atoms.

Solution

The question specifies 2 carbon atoms in total and asks for a displayed formula. An ester requires at least 1 carboxyl C on the acid side and at least 1 C on the alcohol side, so only a 1+1 distribution is possible.

Choose methanoic acid, HCOOH, and methanol, CH_3OH. Their ester is $HCOOCH_3$, methyl methanoate. Begin with H—C(=O)—O—C, then add three H atoms to the right-hand C:

The smallest ester has two carbons: methyl methanoate



$HCOOCH_3$; 1 C on the acid side and 1 C on the alcohol side; all atoms and bonds are shown.

C=O and C—O—C form the ester linkage. H is attached at the left end and CH_3 is on the right, giving two carbon atoms in total.

Check the acid-side C: it has a bond to H, a double bond to O and a single bond to O, giving four bonds. The middle O joins the two C atoms. The right-hand C joins O and three H atoms. All two C, two O and four H atoms are displayed. The official marks require the correct ester linkage and then the complete methyl methanoate structure. $HCOOCH_3$ alone does not meet the request for a displayed formula.

Practice

1. (2025·May/June 0620/21 Q14) · Level 2

The equation represents the reaction between ethanoic acid and ethanol.

$$CH_3COOH + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + H_2O$$

Which type of reaction does this show?

A addition

B neutralisation

C reduction

D reversible

2. (2025·May/June 0620/42 Q1(h)) · Level 2

Using numbers only, state the:

(h) number of carbon atoms in one molecule of propyl butanoate.

3. (2025·May/June 0620/43 Q5(d)(ii)) · Level 2

(d) Carboxylic acids react with alcohols to form esters.

(ii) Propyl butanoate is an ester.

Name the carboxylic acid and alcohol that react to produce propyl butanoate.

carboxylic acid

alcohol

4. (2025·May/June 0620/41 Q5(d)(i)) · Level 3

(d) Carboxylic acids react with alcohols to form esters.

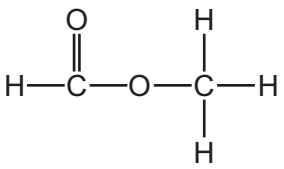
(i) Draw the displayed formula of an ester which contains three carbon atoms.

5. (2025·May/June 0620/22 Q33) · Level 3

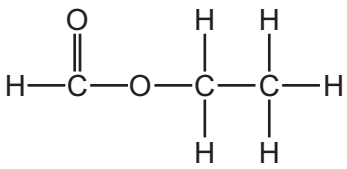
Methanol reacts with ethanoic acid to form an ester.

What is the displayed formula for this ester?

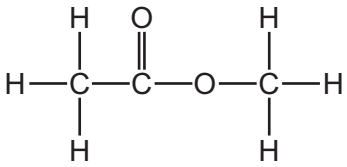
A



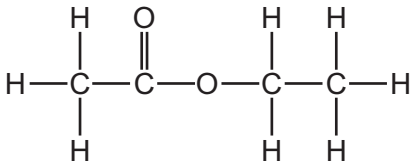
B



C



D

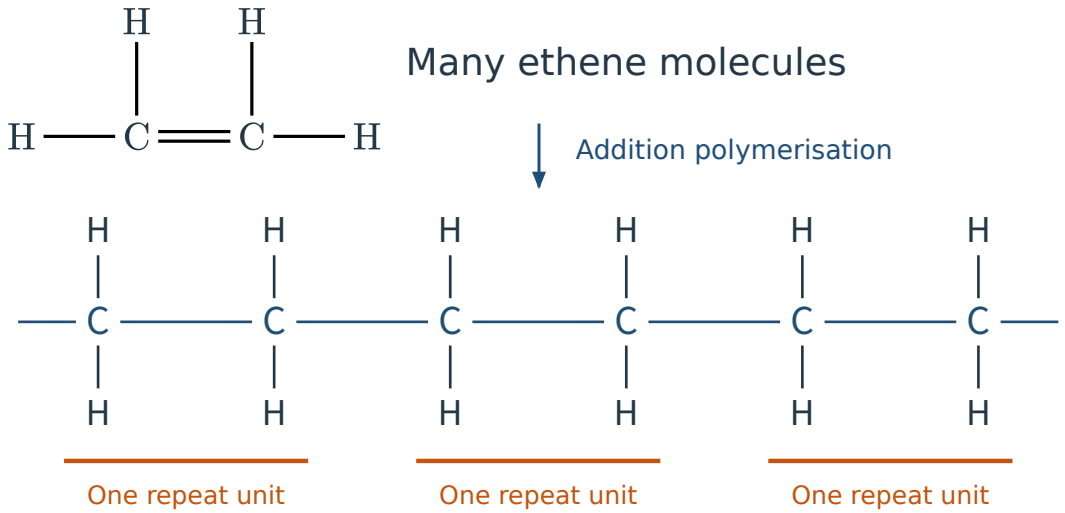


11.8 Polymers

Type 1: Draw repeating units from monomers and recover complete monomers from polymers

Key Ideas **First see how small molecules become a long molecule, then learn how to draw it.** A polymer is a large molecule formed by joining many small molecules called monomers. Ethene contains C=C and undergoes addition polymerisation: the double bonds in many ethene molecules become carbon–carbon single bonds, and the two original double-bonded carbons join neighbouring units to form poly(ethene).

Monomers join into a long chain: a poly(ethene) segment contains several repeat units



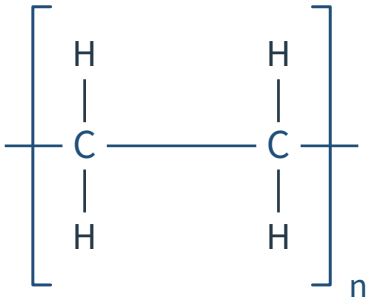
A monomer is the starting small molecule; a repeat unit is the recurring connected section of the chain.

Only three repeating units of the long chain are shown below. The short bonds at the ends show that the chain continues; this is not a complete, separate six-carbon molecule.

One ethene molecule contains 2 C and 4 H atoms. The three ethene units shown contribute 6 main-chain C atoms and 12 H atoms. All remain in the chain; no extra H_2O is released. Bonds extend at both ends to show that this is only part of a long chain. Do not add an H at each end and turn it into a separate alkane molecule.

A repeating unit is a complete set of connections that occurs repeatedly. A repeat unit is neither an arbitrarily chosen atom nor the whole polymer molecule. One repeating unit of poly(ethene) is —CH_2—CH_2—. Put this group inside square brackets and extend the connecting bonds through both sides of the brackets. The n outside indicates that the group repeats many times.

One repeat unit in brackets with n represents many repetitions of that structure



n: number of repeats

Not an extra N atom

End bonds extend beyond the brackets

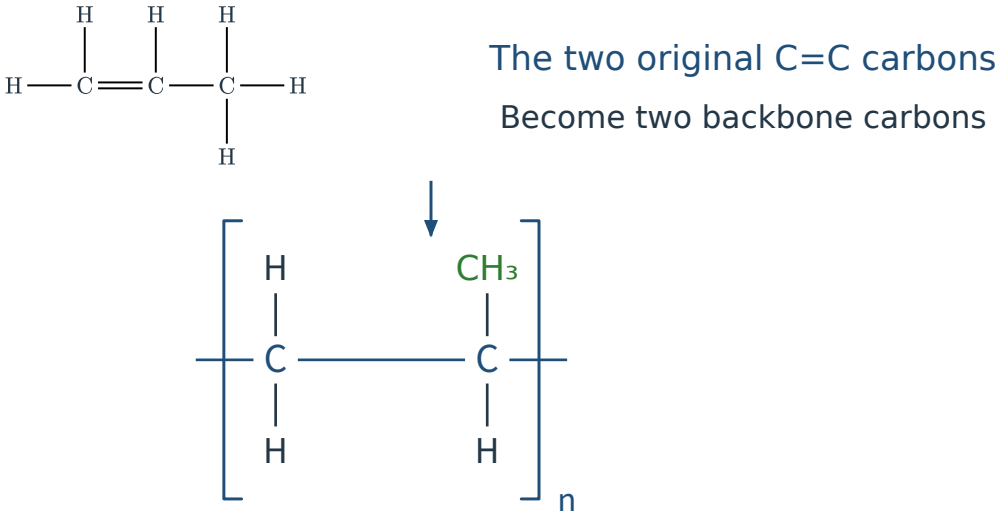
$nCH_2{=}CH_2 \rightarrow [—CH_2—CH_2—]_n$; repeat-unit formula C_2H_4; n is not the subscript 4.

All H atoms are explicitly shown. The horizontal bonds extending through the brackets are covalent bonds connecting to the next units.

If asked for one repeat unit, draw two joined main-chain C atoms, each bonded to two H atoms, with continuation bonds on the left and right. For three repeat units, copy the whole group three times, giving six main-chain C atoms. n indicates repetition, not the element symbol N for nitrogen. Adding n also does not turn the C_2H_4 composition of a unit into C_2H_{4n}.

When converting a monomer to a polymer, identify the two carbons of the original C=C first. Follow four steps: locate C=C; change it to C—C; retain the original atoms or groups attached to each double-bonded carbon in their positions; add bonds to neighbouring units on the left and right. A carbon chain outside the original double bond becomes a side chain here. Do not force it into the polymer's main chain merely to make the drawing look straight.

Propene polymerisation: the original CH_3 stays on its original carbon

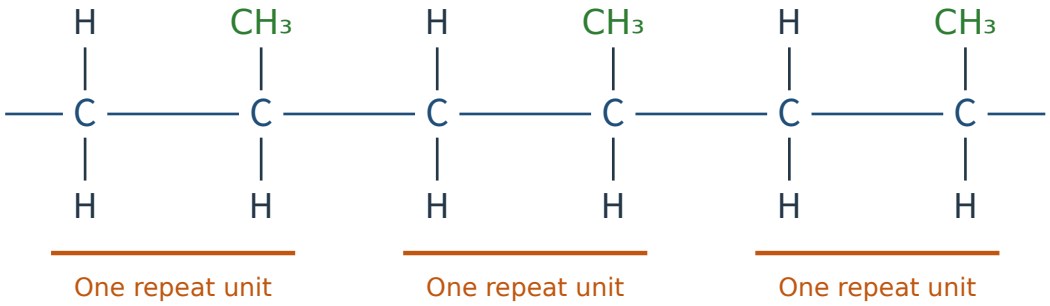


The displayed formula above shows every atom and bond in propene. Below, the side chain is abbreviated as CH_3. The connections of each carbon remain traceable.

In propene, $CH_2{=}CHCH_3$, the left double-bonded C has two H atoms; the right one has one H and one CH_3. After polymerisation, the left C still carries two H atoms and the right C still carries H and CH_3, giving —CH_2—$CH(CH_3)$—. Each main-chain C has a single bond to the C on either side and bonds to its two original attached atoms or groups, giving four bonds altogether. The product is poly(propene).

When drawing several units, copy the entire repeating unit, including the side-chain positions. Three units of poly(propene) have six main-chain C atoms, with CH_3 on carbons 2, 4 and 6. Reversing the whole segment places them on carbons 1, 3 and 5 and represents the same connections. Do not draw only three main-chain C atoms or put all three CH_3 groups on three consecutive carbons.

Three repeat units means copying the whole unit, not drawing only three carbons



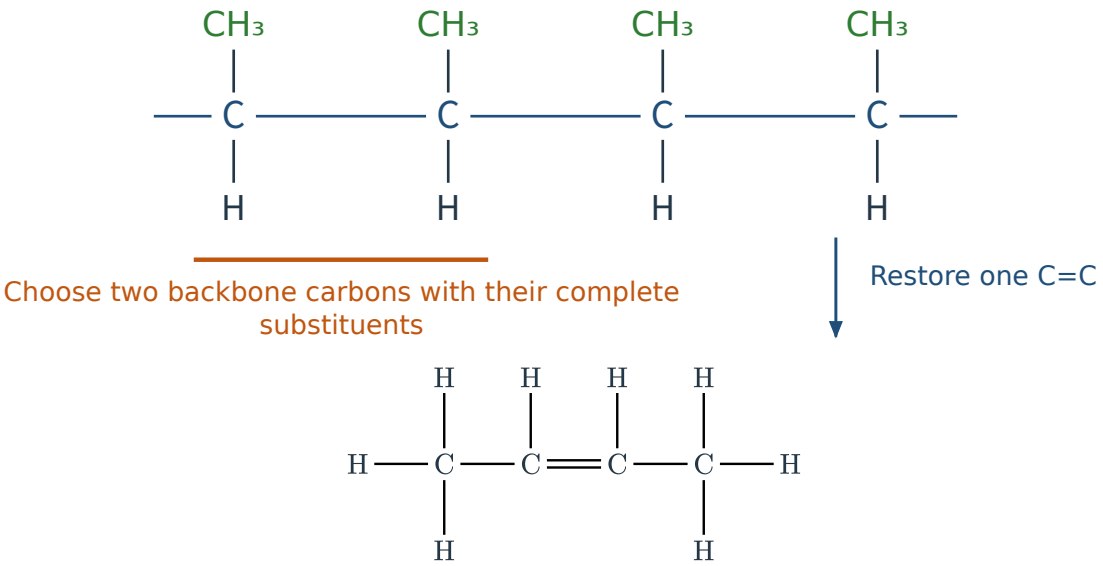
6 backbone carbons, 3 CH_3 side chains, single bonds throughout the backbone and extending bonds at both ends.

Each orange outline identifies the two main-chain C atoms and their side chain contributed by one propene monomer. CH_3 is an accepted abbreviation for the side chain.

Check from left to right in complete groups: CH_2, then CH bearing CH_3; repeat that pair, then repeat it once more. First check that all six main-chain C atoms are joined by single bonds, then check the three CH_3 groups, all main-chain H atoms and both continuation bonds. Where the question allows CH_3 as a side-chain abbreviation, the whole CH_3 group means one C and three H atoms, not three C atoms.

To recover a monomer, take two main-chain carbons together with their attached groups. Select two neighbouring main-chain C atoms that represent the complete repeating pattern. Remove the two continuation bonds that join them to external units, then restore a double bond between the pair. Keep all four original attached atoms or groups on their corresponding carbons.

Recover a monomer: identify the full repeat structure and restore C=C between its two backbone carbons



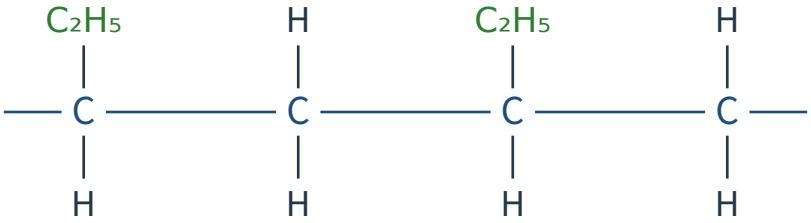
Each C=C carbon carries CH_3: $CH_3CH{=}CHCH_3$, named but-2-ene.

The upper diagram shows a long-chain segment; the lower diagram gives the monomer's complete displayed formula. Four main-chain C atoms in the upper segment do not mean that the monomer needs four double-bonded carbon atoms.

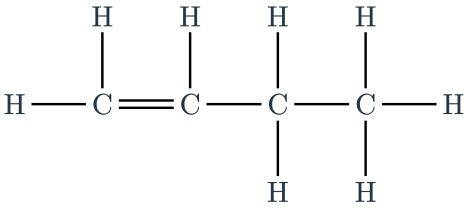
If both main-chain C atoms carry H and CH_3, restoring the double bond gives $CH_3CH{=}CHCH_3$. There are four carbon atoms, so the prefix is but-. From either end, the double bond lies between carbons 2 and 3, giving but-2-ene. This is not propene, which has CH_3 on only one side and CH_2 on the other. Even if the original question shows a long polymer segment, you do not need to convert the whole segment into one enormous monomer.

Expand abbreviated side chains fully when a displayed formula of the monomer is required. C_2H_5 is an ethyl group, with structure CH_2CH_3. If one main-chain C carries H and C_2H_5 while the other carries two H atoms, restoring the double bond gives $CH_2{=}CHCH_2CH_3$. The molecule contains four carbon atoms and the double bond begins at carbon 1 when numbered from the nearest end, so it is but-1-ene.

Recovering a monomer: keep CH_3 and C_2H_5 distinct, preserving each side chain's carbon count



C_2H_5 is CH_2CH_3, with two C atoms; the other backbone carbon carries two H atoms.



Monomer $CH_2{=}CHCH_2CH_3$: number from the end nearer C=C to obtain but-1-ene.

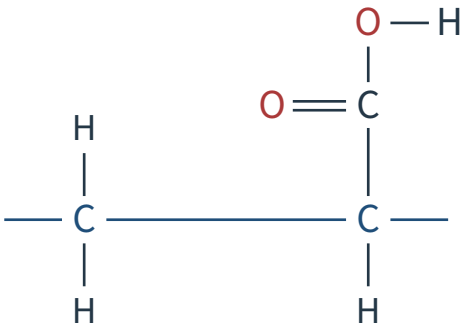
A polymer side chain may be abbreviated as C_2H_5 in the question diagram. When a displayed monomer is required, expand its C atoms and all its H atoms.

First expand C_2H_5 into CH_2—CH_3, attached beside the original double-bonded C. Then draw each H in every CH_2 and CH_3 group separately. The finished molecule should have 4 C and 8 H atoms, one C=C and only single covalent bonds elsewhere. If asked to draw and name, give both the correct structure and the name but-1-ene; either alone is incomplete.

Other functional groups remain attached during this addition-polymerisation step. For example, $CH_2{=}CHCOOH$ contains both C=C and COOH. It contains O, so it is not a hydrocarbon made only of C and H. If addition polymerisation is specified, change C=C to obtain —CH_2—CH(COOH)—. The carboxyl group stays on its original carbon; it is neither removed nor used to eliminate water in this step.

With both C=C and COOH present, addition polymerisation still acts only on C=C

Monomer: $CH_2=CHCOOH$



Repeat unit: —CH_2—CH(COOH)—

COOH is retained; this reaction does not use it in water-forming condensation.

C: 3; H: 4; O: 2, as in the monomer. End bonds continue into the long chain.

The carboxylic acid contains two recognisable structural features, including C=C. Every atom and bond in its carboxyl group is shown. Because it contains oxygen, it is not a hydrocarbon.

The right-hand main-chain C still carries one H and one COOH; the left-hand C carries two H atoms. Counting all atoms gives 3 C, 4 H and 2 O per unit, the same as the monomer. For the more complex molecule in the original question, with two COOH groups and one CH_3, use the same method: record the four groups on the two ends of the original double bond and transfer them to the new C—C. No new drawing method is needed.

In addition polymerisation, C=C becomes _____ and the monomer's atoms are _____. One propene repeating unit contains _____ main-chain C atoms; three units contain _____ main-chain C atoms and _____ CH_3 groups. When recovering a monomer, expand the original side chain C_2H_5 to _____, then draw every H and bond.

Answers: C—C; all retained; 2; 6; 3; CH_2CH_3

Recognising the Question From monomer to polymer, change C=C to C—C, retain all four attached atoms or groups in place and leave continuation bonds at both ends. Copy whole units for the required number of repeats. In reverse, select two representative main-chain C atoms, restore C=C, expand CH_3/C_2H_5 and name the monomer from the double-bond position.

Worked Example 1 (2024·October/November 0620/42 Q5(d)(ii)) · Level 3

(d) Propene will undergo polymerisation.

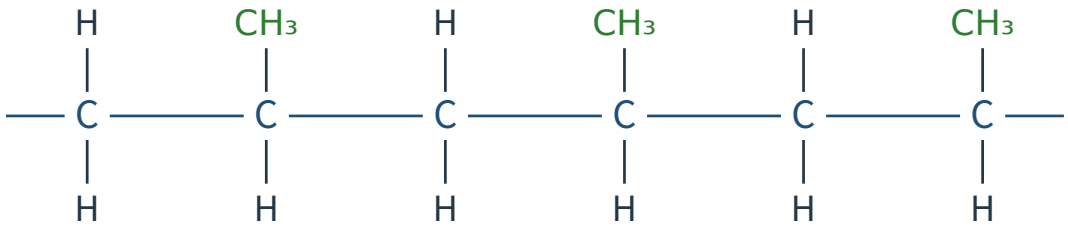
(ii) Draw part of this polymer molecule to show **three** repeat units.

Solution

The monomer is propene, $CH_2=CHCH_3$. Its left double-bonded C carries two H atoms, while the right carries one H and one CH_3. Polymerisation changes the double bond between these carbons to a single bond and joins them into the main chain.

One repeating unit is —CH_2—CH(CH_3)—. The question asks for three repeat units, so copy this whole group three times, giving six consecutive main-chain C atoms with one CH_3 on every other carbon:

Three propene repeat units: six backbone carbons and three CH_3 side chains



—CH_2—$CH(CH_3)$— repeated 3 times, with extending bonds at both ends.

For w24/42/5(d)(ii). All main-chain bonds are single. The side chains use CH_3, as allowed by the original mark scheme, and all main-chain H atoms are explicitly shown.

Check three features against the diagram: all six main-chain C atoms are joined by single bonds; there are three correctly spaced CH_3 groups; and every required H and both continuation bonds are present. The official three marks correspond to these checks. CH_3 may be abbreviated as shown, but main-chain H atoms must not be omitted. Do not cap the ends with H and remove the chain's continuation.

Worked Example 2 (2024·May/June 0620/41 Q7(b)) · Level 3

(b) Fig. 7.1 shows a section of polymer **Q**.

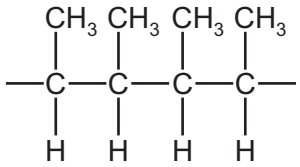


Fig. 7.1

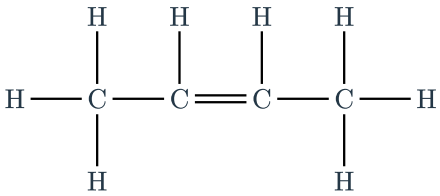
- Draw the displayed formula of the monomer that forms polymer **Q**.
- Name the monomer used to form polymer **Q**.

Solution

Read Fig.7.1: every main-chain C carries one CH_3 and one H. Selecting two neighbouring main-chain C atoms as a unit therefore gives CH_3 and H attached to each.

Remove the two bonds extending out of this unit, then restore C=C between its carbons to obtain $CH_3CH{=}CHCH_3$. Expand both CH_3 groups into all their C—H bonds. The complete monomer is:

But-2-ene: each double-bonded carbon carries one CH_3



$CH_3CH=CHCH_3$; but-2-ene

The complete displayed formula shows four C atoms, eight H atoms and every bond. Rotating the whole diagram does not change the connections.

The molecule has 4 C atoms, with the double bond between carbons 2 and 3, so its name is **but-2-ene**. The original question asks for both a drawing and a name; give both.

The official marks first credit C=C with four bonds around each of its carbons, then the correct complete displayed formula of but-2-ene, and finally its name. Either permitted spatial arrangement is accepted; the connections are what matter. Propene is incorrect because every main-chain carbon in the given polymer has CH_3, whereas poly(propene) has CH_3 on alternate main-chain carbons.

Practice

1. (2025·May/June 0620/41 Q1(j)) · Level 1

A list of substances is shown.

bauxite **calcium oxide** **ethanol** **graphite** **methane**

nitrogen **oxygen** **propane** **propene** **sulfur dioxide**

Answer the questions using the list of substances.

Each substance may be used once, more than once, or not at all.

State which of the substances:

(j) is a monomer in addition polymerisation.

2. (2024·February/March 0620/42 Q5(e)(ii)) · Level 3

(e) But-1-ene undergoes polymerisation.

(ii) Draw part of the polymer molecule to show **three** repeat units.

3. (2025·May/June 0620/42 Q6(e)(iii)) · Level 3

(e) Compound **B** has the displayed formula shown in Fig. 6.1.

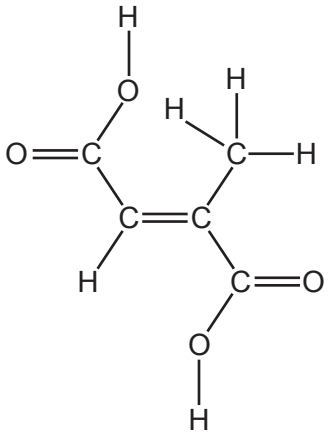


Fig. 6.1

(iii) Draw the structure of **one** repeat unit of the polymer formed when compound **B** undergoes addition polymerisation.

4. (2024·May/June 0620/43 Q7(b)(ii)) · Level 3

(b) Fig. 7.1 shows a section of a polymer formed from an alkene.

(ii) A section of a polymer is shown in Fig. 7.1.

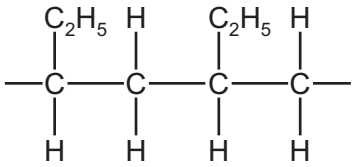


Fig. 7.1

- Draw the displayed formula of the monomer that forms this polymer.
- Name the monomer used to form this polymer.

5. (2025·October/November 0620/43 Q6(d)(iv)) · Level 3

(d) Fig. 6.1 shows part of a polymer structure.

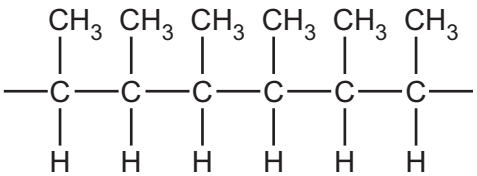


Fig. 6.1

(iv) Draw the displayed formula of the monomer used to make this polymer.

Type 2: Draw polyesters and polyamides, then recover monomers and amino acid structures

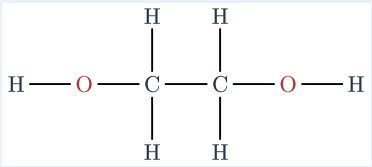
Key Ideas **First identify two ends that can continue joining to other molecules.** Addition polymerisation in the previous question type uses C=C. Condensation polymerisation instead joins functional groups, forming a large molecule and a small molecule such as water. To extend repeatedly from both ends, a monomer needs two suitable functional groups, rather than just one end that can react once.

A diol contains two alcohol OH groups, a dicarboxylic acid contains two COOH groups, and a diamine contains two NH_2 groups. NH_2 is an amino group: N is joined to two H atoms and, by a third bond, to the carbon framework. The prefix di- means two; it does not mean that the whole molecule contains only two carbon atoms.

To join at both ends into a long chain, a monomer needs two suitable functional groups

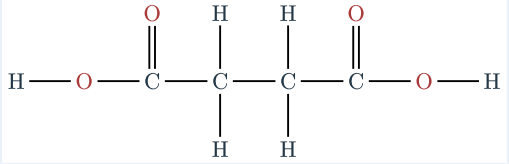
Diol

Two alcohol OH groups



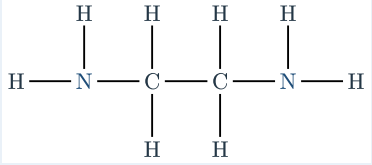
Dicarboxylic acid

Two COOH groups



Diamine

Two NH_2 groups



On the left, di- indicates two functional groups. The displayed formulae on the right show every atom and bond at each end. These specific molecules are examples, not the only possible starting materials for all polymers.

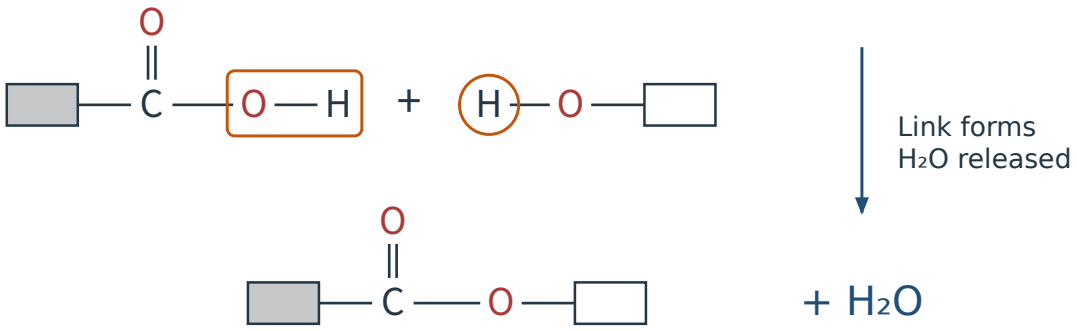
$HOCH_2CH_2OH$ has OH at both ends and can join from either end. $HOOCCH_2CH_2COOH$ has COOH at both ends, also giving two reaction sites. By contrast, CH_3OH has only one OH and cannot form the same long chain through this two-ended joining process alone. Count functional groups first, and do not count the OH inside COOH as a separate alcohol group.

To understand an ester linkage, examine one joining step closely. The C=O in a COOH group is retained. The acid-side OH and the H from the alcohol-side OH together account for H_2O, leaving the acid-side C joined to the alcohol-side O as C(=O)—O. This tracks the overall structure and atom balance to construct the product accurately; it does not require a reaction mechanism.

Coloured blocks in the diagram represent the remaining carbon frameworks, with their internal atoms temporarily omitted. A block is not a new element. Different questions may define the blocks differently, so read each diagram and keep the identity of each framework consistent throughout that question.

One ester linkage: acid OH and alcohol H form one water molecule

Enlarged functional groups at a single joining point



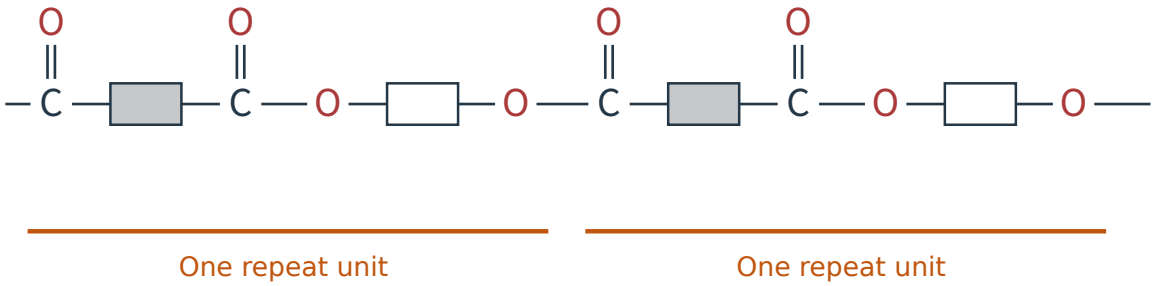
The link is —C(=O)—O—; the retained O connects to the alcohol's carbon skeleton.

Coloured blocks represent the remaining carbon frameworks, whose internal structures are omitted here. The orange box and circle help check the overall atom balance; they do not prescribe a detailed one-step mechanism.

The acid side starts with two O atoms: one remains in C=O and the other enters water as part of OH. The alcohol-side O remains and joins the acid-side C to the alcohol-side framework, still forming two bonds. Do not remove both O atoms or join the acid-side C directly to the alcohol carbon chain while omitting the linking O.

Repeated joining of a dicarboxylic acid and a diol forms a polyester. It is called a polyester because C(=O)—O ester linkages recur along the chain. One repeating unit contains one acid residue and one alcohol residue. A residue is the part of a monomer remaining in the chain after the water-forming atoms have been removed. For two repeating units, draw two acid-side blocks and two alcohol-side blocks, alternating with the correct original connections.

Dicarboxylic acid + diol: repeated ester-link formation produces a polyester



Grey: acid carbon skeleton; white: alcohol carbon skeleton. Two repeat units contain four blocks.

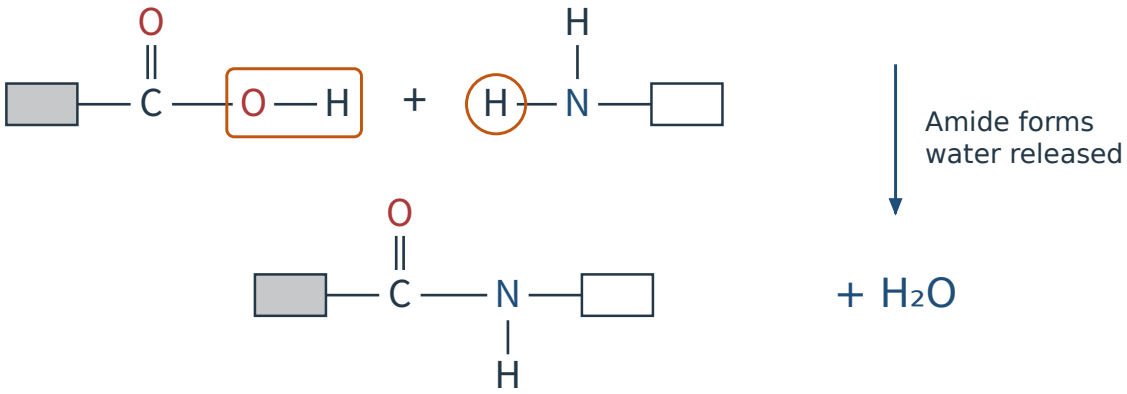
Keep every C=O; each single-bonded O joins both sides; the outermost bonds extend onwards.

Two complete repeating units are shown, with every atom and bond in the ester linkages displayed. The internal carbon frameworks of the blocks are omitted as allowed by the question.

Read from the left: C(=O)—acid-side block—C(=O)—O—alcohol-side block—O forms one unit. Copy the whole unit again to obtain two. The O at the end of one unit joins C(=O) in the next, forming the same ester linkage. Finally, check every C=O, every singly bonded O and both continuation bonds. Four boxes joined only by plain lines do not show the required linkages.

An amide linkage differs from an ester linkage because N is attached beside C=O. The amide linkage here is C(=O)—NH. A COOH end of a dicarboxylic acid joins an NH_2 end of a diamine. The acid's OH and one H from the amine form water. N retains its other H and joins the structures on both sides, giving three covalent bonds.

One amide linkage: acid OH and one amine H form water



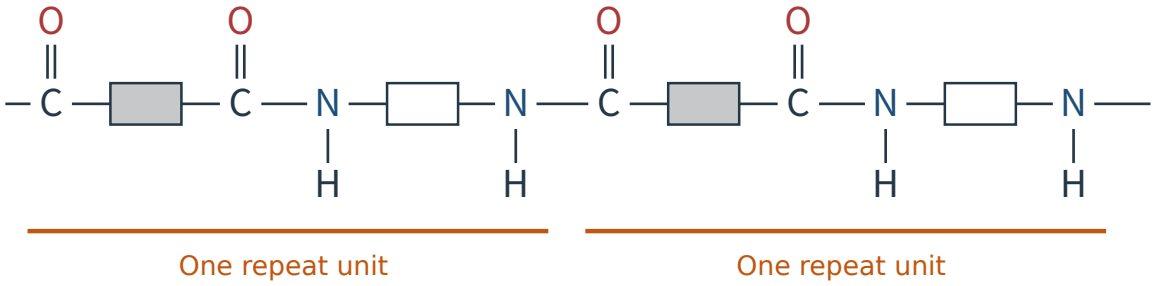
The link is —C(=O)—NH—; N retains one H, so do not remove both.

The coloured blocks represent the remaining carbon frameworks. The diagram tracks the overall atom balance for one joining step. C=O is retained, and N still forms three covalent bonds after joining.

Retain the acid's C=O, then join that C to the N of the original NH_2. N is now bonded to the acid-side C, its original carbon framework and one H, giving exactly three bonds. Keeping both original NH_2 hydrogens while adding C—N would give one ordinary covalent bond too many here. Replacing NH with O throughout would give a polyester instead.

A dicarboxylic acid and a diamine form a polyamide; nylon is a typical example. A polyamide such as nylon has recurring amide linkages in its main chain. The drawing method resembles that for a polyester, but N bearing one H joins the amine-side carbon framework. Do not copy the polyester's O atoms unchanged.

Nylon is a polyamide formed from a dicarboxylic acid and a diamine



Each N has three bonds: two to neighbouring sections and one to H; each amide carbon retains C=O.

Two repeat units are shown; grey blocks are acid-derived and white blocks are amine-derived.

The nylon diagram uses blocks for the internal carbon chains, following the syllabus convention. All amide linkages, both repeating units and both continuation bonds are complete.

Draw one unit as C(=O)—acid-side block—C(=O)—NH—amine-side block—NH, then copy it in full. The result should contain four framework blocks, four N atoms each bearing H, four C=O groups and continuation

bonds at both ends. Follow each connection along the same chain. Do not arbitrarily reverse one unit partway through and create incorrect C-to-C or N-to-N joins.

Both addition and condensation polymerisation form long chains, but they form them differently. In the addition reactions studied here, monomer C=C becomes C—C and all atoms enter the polymer without additional loss of water. The condensation examples use functional groups at both ends, producing water when the corresponding new linkages form. Saying only that a large molecule forms does not distinguish them; add the structural change or the small-molecule by-product.

Addition versus condensation: how do monomers join, and is a small by-product released?

	Addition polymerisation	Condensation polymerisation
Key monomer structure	Look for C=C here	Here, two end groups: COOH, OH or NH_2
When the backbone forms	C=C becomes C—C; all atoms retained	Functional groups join and release small molecules such as water
Typical polymers	Poly(ethene)	Polyesters and polyamides

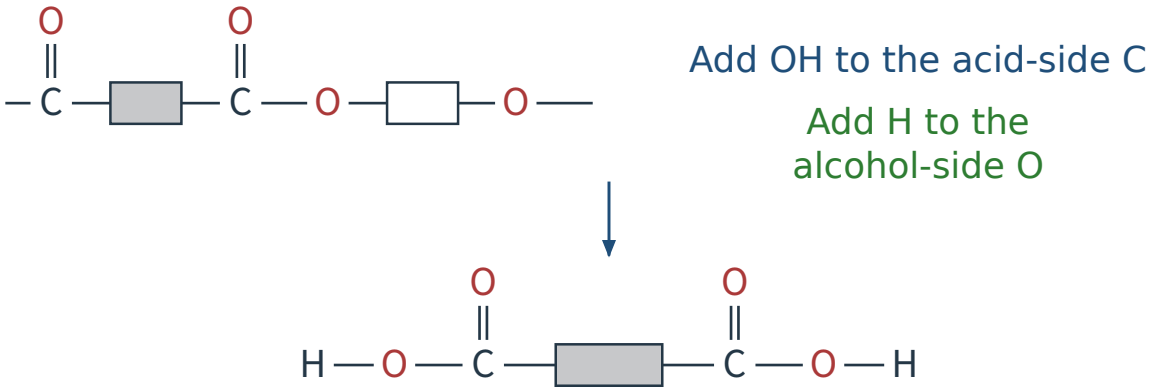
Long chain or large molecule alone does not distinguish them; examine how the links form.

This comparison is limited to the syllabus examples of addition and condensation polymerisation. Water is the small molecule produced in these condensation reactions; not every polymerisation eliminates water.

Ethene forming poly(ethene) is addition polymerisation. A dicarboxylic acid and a diol forming a polyester is condensation polymerisation. The first changes C=C; the second joins COOH and OH while producing water. A polyester being made from monomers does not make its formation addition polymerisation.

To recover polyester monomers, first separate the C(=O)—O linkage. Add OH back to the acid-side C to restore COOH. Add H back to the alcohol-side O to restore OH. Restore both ends of each fragment to obtain the complete dicarboxylic acid and diol. Repairing only the most obvious central linkage leaves the other end incorrectly missing H or OH.

Recover polyester monomers: split every C(=O)—O linkage



Restore dicarboxylic acid: complete COOH at both ends



Restore diol: O—H at both ends

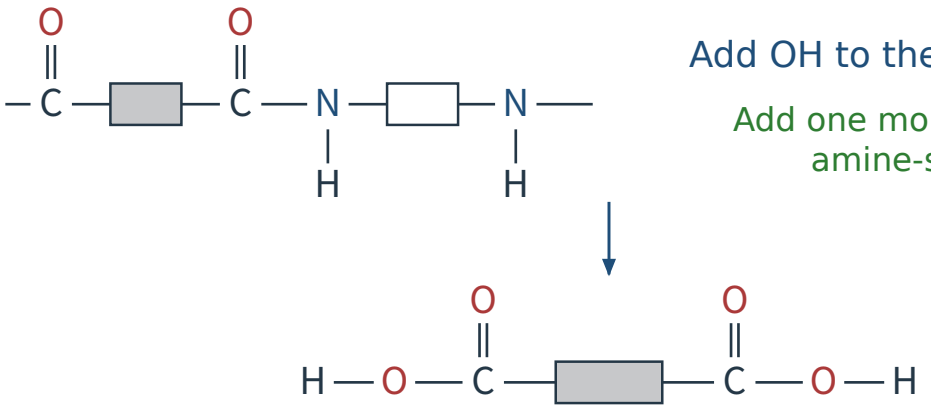
Restore both ends of every monomer, not just the one connection drawn in the middle.

The upper diagram shows only one unit of a polyester-chain segment. Both complete monomers below have their two terminal functional groups restored. Colours track the framework identities within this diagram.

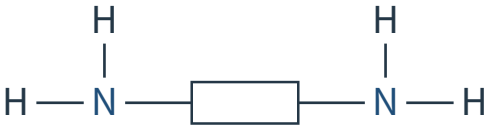
For —C(=O)—grey block—C(=O)—O—white block—O—, restore the acid side to HOOC—grey block—COOH and the alcohol side to HO—white block—OH. Retain the full carbon chain inside each block. This separation is a way of working backwards from the structure in an answer. If experimental conditions are not requested, there is no need to add a temperature or reagent.

Recover polyamide monomers similarly, but restore NH_2 on the amine side. Separate C(=O)—N. Add OH to the acid-side C. N already carries one H, so add one more to restore NH_2. After identifying the dicarboxylic acid and diamine, fully display the functional groups at both ends. In particular, do not omit either H on N or the acid's C=O.

Recover polyamide monomers: split C(=O)—N, then add OH and H respectively



Dicarboxylic acid: restore COOH at both ends



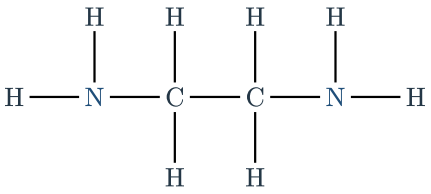
Diamine: each N ends with two H atoms and one skeleton C, making three bonds.

Amide N already has one H, so restoring the monomer requires just one additional H. The O in C=O is retained throughout.

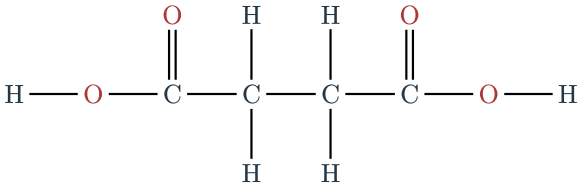
If the grey block joins two C=O groups, it belongs to the acid side and becomes HOOC—grey block—COOH. If the white block joins two NH groups, it belongs to the amine side and becomes H_2N—white block—NH_2. Their identities do not change because their left–right positions differ. For Complete Fig. questions, add the correct functional groups directly to both ends of the original boxes, retaining their positions and identities.

Retain each framework's length exactly. Correct functional groups alone are not enough: questions may use dicarboxylic acids or diamines with different carbon counts as distractors. On the acid side, both COOH carbons were already present and must be counted. On the amine side, count only the carbons in the segment between the two N atoms; do not include carbons from the other segment.

After restoring both monomers, count each original carbon skeleton separately



Diamine example: $H_2NCH_2CH_2NH_2$ has only 2 C atoms between its N atoms.



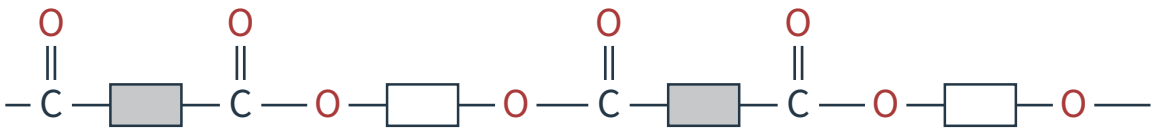
Dicarboxylic acid example: $HOOCCH_2CH_2COOH$ has 4 C atoms, including both carboxyl carbons.

The displayed formulae allow the carbons in both monomers to be counted individually. A carboxyl carbon is part of a functional group and is also an actual carbon atom.

$H_2NCH_2CH_2NH_2$ has 2 C atoms between its N atoms. $HOOCCH_2CH_2COOH$ has 2 C atoms in the middle and 4 in total when the two carboxyl carbons are included. Replacing the middle CH_2CH_2 with one CH_2 would give the wrong monomer even if both COOH groups were drawn correctly. Comparing only the total carbon count of the whole chain cannot detect an error that swaps the lengths of the two segments.

For PET, recognise the polymer and draw its linkages. PET is a polyester with two frameworks contributed by a dicarboxylic acid and a diol. Its chain contains recurring C(=O)—O linkages. If the syllabus diagram uses coloured blocks to omit internal structures, you may use that representation, but every C and O at a linkage, each C=O double bond, each C—O single bond and the continuation bonds must be present. Writing PET alone does not replace a drawing.

PET is a polyester: show ester linkages and alternating types of skeleton



Grey block: internal dicarboxylic-acid skeleton; white block: internal diol skeleton.

Only the blocks' internal structures are omitted; every linkage C and O and every single and double bond is retained.

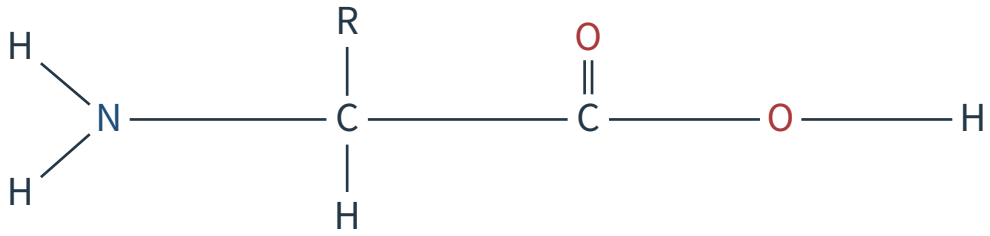
Two repeating units are represented using the syllabus block diagram for PET. Memorising PET's full name is not required.

If asked to name the two types of monomer molecules, give **dicarboxylic acids** and **diols**. Acid and alcohol alone omit the two-functional-group requirement. For two repeating units, draw two acid-side blocks, two alcohol-side

blocks and every ester linkage. If another question instead defines the white block as the acid side, follow that question's definition. Colour is not an intrinsic property of a functional group.

An amino acid has NH_2 and COOH in the same molecule. The general amino acid structure in this syllabus is $H_2N—CH(R)—COOH$. The central C joins NH_2, COOH, H and R, giving four bonds. R represents a variable side chain. It may be H or another, carbon-containing structure; it is not a new element symbol.

An amino acid has NH_2, COOH, H and R attached to the same central carbon



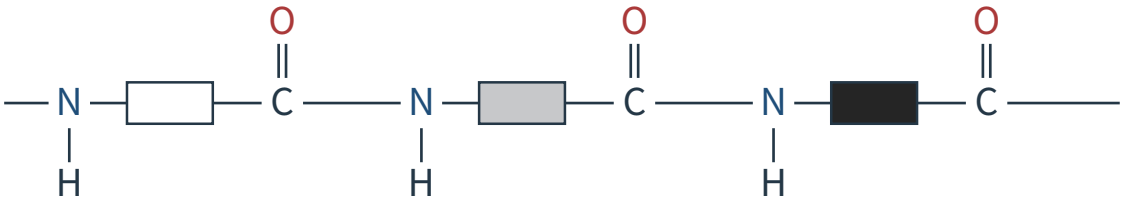
R represents a side chain that varies between amino acids; it is not an element symbol.

Every atom and bond in NH_2 and COOH is displayed. R represents an unexpanded side chain. The diagram shows connections, not actual three-dimensional bond angles.

If the original diagram gives a central C, R above, H below and short bonds to the left and right, complete the ends. Add N with two H atoms on one side and C=O with O—H on the other. NH_2 and COOH may swap left and right provided both join the same central C and every atom has the correct bond count. Do not omit the C in COOH or replace NH_2 with an oxygen-containing group.

Proteins are natural polyamides formed from amino acids. The COOH group of one amino acid joins the NH_2 of another, forming an amide linkage and water. The original central C and side chain remain in the polymer. Different amino acids have different R groups, which can be represented by different coloured blocks, but the C(=O)—NH linkage is the same.

Amino acids form proteins: each monomer has an NH_2 end and a COOH end



Original amino acid A Original amino acid B Original amino acid C

The C(=O)—NH link is an amide; coloured blocks retain each amino acid's internal structure.

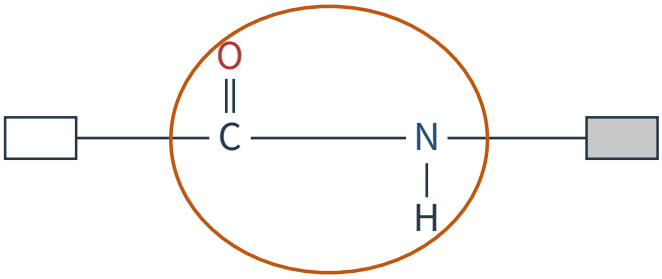
Only a chain segment is shown; keep extending end bonds instead of closing three separate molecules.

The three differently coloured blocks represent the central C and side-chain portions of three amino acids. Colours do not represent elements. All amide linkages and bonds continuing the chain are shown.

If amino acids A, B and C are given with three specified box positions, retain their order as A's white block, B's grey block and C's black block. Put NH before each block and C=O after it, then join the carboxyl C after one block to the N before the next. Each of the three N atoms retains one H, every C=O remains and the chain continues at both ends. A, B and C identify the three amino acids; use the original coloured blocks in the actual answer. Nylon also has amide linkages, but it is not the amino-acid-derived protein studied here.

If asked to circle one amide linkage, include one complete linkage. Find N directly attached beside C=O and include the H on that N, identifying C(=O)—NH. Do not circle only O, the next framework block or two separate linkages in one large circle.

Circle an amide linkage: include N and its H beside C=O, not just one O



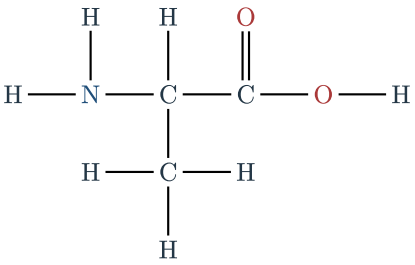
C(=O)—NH is an amide; C(=O)—O is an ester. The atom attached to the carbonyl carbon differs.

The orange circle identifies just one amide linkage. The two neighbouring framework blocks remain as reference points.

In a chain with several amide linkages, any one may be selected. If O rather than N is attached beside C=O, the linkage is an ester and does not answer amide. The original diagram may alternate upwards and downwards. Follow the actual C—N bond instead of looking mechanically for letters at one height or in one direction.

When completing an amino acid with a specified total carbon count, allocate the essential carbons first. The central CH contains one C, and COOH contains another. If the whole molecule has three carbon atoms, the remaining C must be a CH_3 side chain attached to the central carbon. Complete NH_2, COOH and every H, then check each atom's bonding.

For a three-carbon amino acid, reserve one C for the centre and one for COOH



Central C: 1; COOH C: 1; remaining CH_3 side-chain C: 1.

3 C, 7 H, 1 N and 2 O; NH_2 and COOH attach to the same central carbon.

This is the displayed formula of 2-aminopropanoic acid, which can be used to check the completed atoms and bonds. Do not assume that R always represents a single carbon atom.

The two essential carbons account for 2 of the total: 3−2=1 remains, so R is CH_3. The central C already joins H, CH_3, NH_2 and COOH and cannot have a fifth bond. The three H atoms of CH_3 must still be drawn. If the side

chain were CH_2CH_3, the whole molecule would have four carbon atoms. It would retain both correct functional groups but fail the specified total carbon count.

A dicarboxylic acid and a diol form a _____ with _____ linkages. A dicarboxylic acid and a diamine form a _____ with _____ linkages. Both reactions are _____ and produce _____. Protein monomers are called _____, and R in their general structure represents _____. When recovering an amine monomer, each N must ultimately carry _____ H atoms.

Answers: polyester; C(=O)—O; polyamide; C(=O)—NH; condensation polymerisation; water, H_2O; amino acids; variable side chains; 2

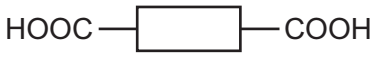
Recognising the Question Identify COOH+OH, COOH+NH_2, or the two ends of the same amino acid. Draw the appropriate ester or amide linkage with all atoms, bonds and the required number of repeats. To recover monomers, separate C(=O)—O or C(=O)—N, add OH to the acid side and H to the alcohol O or amine N, and retain each original carbon framework.

Worked Example 1 (2025·October/November 0620/41 Q5(e)) · Level 3

(e) Butane-1,4-diol can be represented as shown.



Butanedioic acid can be represented as shown.



Butane-1,4-diol reacts with butanedioic acid to form a polymer.

(i) Draw **two** repeat units of the polymer formed from the reaction of butane-1,4-diol with butanedioic acid.

Show all the atoms and all the bonds in the ester linkages.

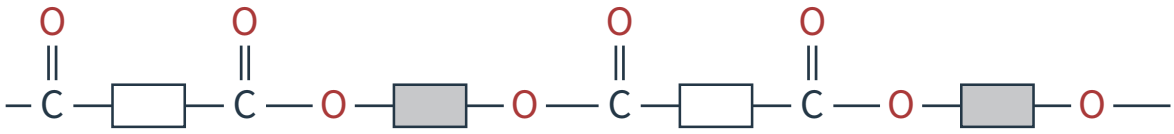
(ii) State the type of polymerisation when butane-1,4-diol reacts with butanedioic acid.

Solution

(i) Read the block definitions in this question first. The framework of butane-1,4-diol is the **grey block**, with OH at each end. The framework of butanedioic acid is the **white block**, with COOH at each end. Use these definitions; do not assume that colours always match the same monomer type in other diagrams.

When an ester linkage forms, the acid-side C=O is retained. The acid's OH and the alcohol's H form water, leaving C(=O)—O joined. One repeating unit needs one white acid-side block and one grey alcohol-side block. The question asks for two repeat units, so draw two complete groups:

Two polyester repeat units: keep this question's white acid and grey alcohol blocks



White blocks come from butanedioic acid; grey blocks come from butane-1,4-diol.

Two acid residues + two alcohol residues, with all C=O and C—O bonds and both extending end bonds.

For w25/41/5(e)(i), using the block identities in the original question: white for the acid and grey for the alcohol. All ester linkages in both repeating units are complete.

Count the structure: two white and two grey blocks alternate. C=O is retained at both ends of each acid-side block, while each alcohol-side block connects through O at both ends. Continuation bonds are present at the far left and right. The question explicitly requires all atoms and all bonds in the ester linkages, so draw each C, both associated O atoms and the C=O double bond.

The official three drawing points check at least one correct complete ester linkage; all three connections between the four blocks being correct and consistently oriented; and two complete repeating units with continuation bonds. Shading in the mark-scheme diagram identifies frameworks. This answer retains the original question paper's white-acid and grey-alcohol identities.

(ii) The functional groups of a diol and a dicarboxylic acid join while producing water, so the reaction is **condensation polymerisation**. It is not addition polymerisation.

Worked Example 2 (2024·October/November 0620/42 Q6(a)(ii)) · Level 3

(a) Part of the structure of a polyamide is shown in Fig. 6.1.

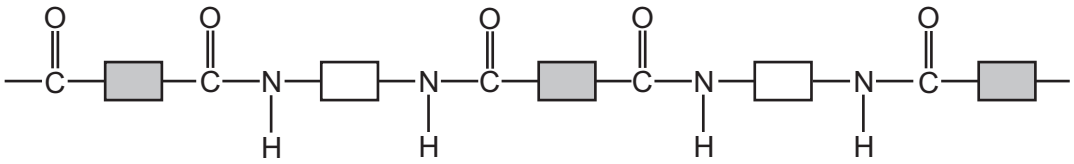


Fig. 6.1

(ii) Complete Fig. 6.2 to show the structures of the **two** monomers needed to make the polymer in Fig. 6.1.
Show all of the atoms and all of the bonds in the functional groups.





Fig. 6.2

Solution

In Fig.6.1, the grey block is joined to C=O at both ends, identifying it as the dicarboxylic-acid part. The white block joins NH at both ends, identifying it as the diamine part.

Separate each C(=O)—N linkage. Add OH to both carboxyl carbons on the grey-block side, restoring COOH at both ends. Each N on the white-block side already has one H; add one more to each, restoring NH_2 at both ends. The completed original Fig.6.2 is:

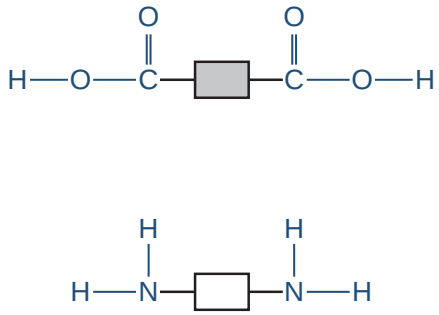


Fig. 6.2

The original grey and white boxes and short bonds in Fig.6.2 remain in place. Blue additions complete COOH at both ends of the grey block and NH_2 at both ends of the white block.

Check every end: the acid has two C=O groups and two O—H groups, while the amine has two N atoms, each bearing two H atoms. The official two points credit complete COOH groups at both ends of the grey block and complete NH_2 groups at both ends of the white block. COOH or NH_2 abbreviations alone are insufficient because all atoms and bonds within the functional groups are requested. Do not swap the block identities.

Worked Example 3 (2025·October/November 0620/43 Q6(e)) · Level 3

(e) Fig. 6.2 shows part of the general structure of an amino acid.

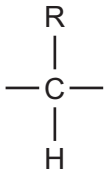


Fig. 6.2

(i) Complete Fig. 6.2 to show all the atoms and all the bonds in the two functional groups of the amino acid.

(ii) Part of a natural polyamide structure is shown in Fig. 6.3.

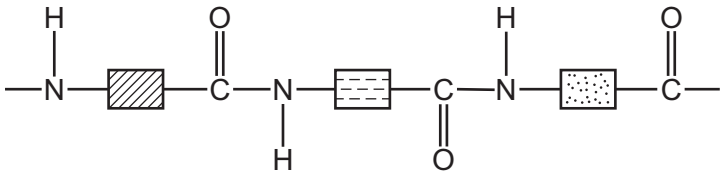


Fig. 6.3

On Fig. 6.3, draw a circle around **one** amide linkage.

(iii) State the name given to natural polyamides made from amino acids.

Solution

(i) The central C in the original diagram already joins R above and H below. Its other two short bonds must join NH_2 and COOH. Add N with two H atoms on the left, and C=O with O—H on the right, giving the complete general amino acid structure:

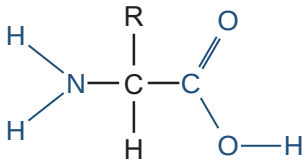


Fig. 6.2

The original central C, R, H and four starting bonds are retained. Blue additions complete NH_2 and COOH on the left and right.

N has three bonds in total. The carboxyl C joins the central C, one O by a double bond and another O by a single bond; the singly bonded O also joins H. The official two points credit complete NH_2 and complete COOH. These functional groups may swap sides if their connections remain correct.

(ii) An amide linkage is C(=O)—NH. Find a carbonyl C directly joined to the next N in the original diagram, and circle this linkage together with its corresponding O and H:

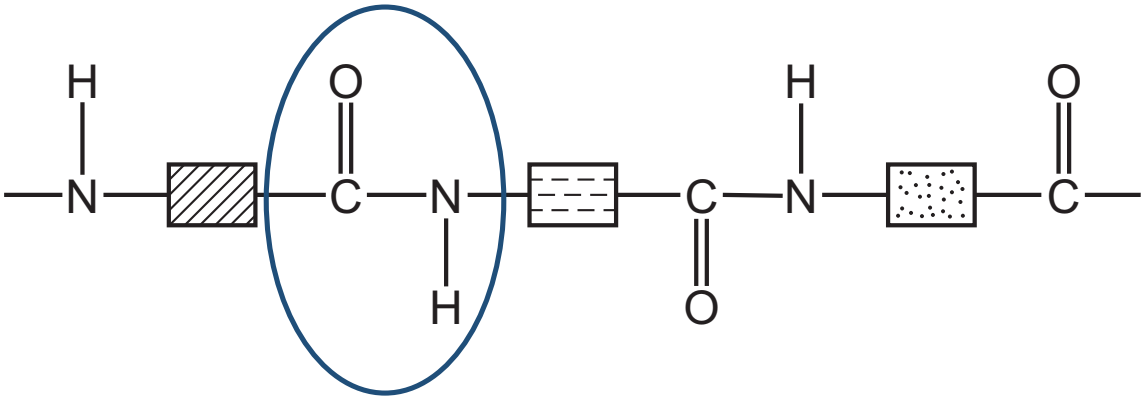


Fig. 6.3

The original Fig.6.3 is retained. The blue circle identifies one C(=O)—NH amide linkage, including its associated O and H.

The official mark scheme accepts any one amide linkage in the diagram. Do not circle only O or treat a whole amino acid framework block as the linkage.

(iii) These natural polyamides formed from amino acids are called **proteins**. Write proteins rather than nylon. Both have amide linkages, but their monomer origins differ, and natural is a condition of this question.

Worked Example 4 (2024·October/November 0620/41 Q6(d)) · Level 3

(d) Three different amino acids are represented as shown in Fig. 6.3.

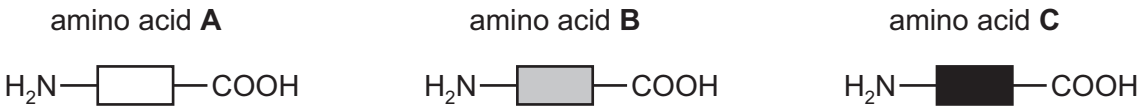


Fig. 6.3

Complete the diagram in Fig. 6.4 to show the part of the structure of the natural polyamide that forms when the three amino acids, **A**, **B** and **C**, combine.
Show all of the atoms and all of the bonds in the linkages.



Fig. 6.4

Solution

The original diagram gives A as a white block, B as a grey block and C as a black block, in that order. Each amino acid has NH_2 on the left and COOH on the right. Retain all three blocks and their order. The C of the right-hand COOH joins the N of the next molecule's left-hand NH_2, forming C(=O)—NH and water.

In the original Fig.6.4, add NH to the left of each block and C=O to the right. Join the C after one block to the N before the next:

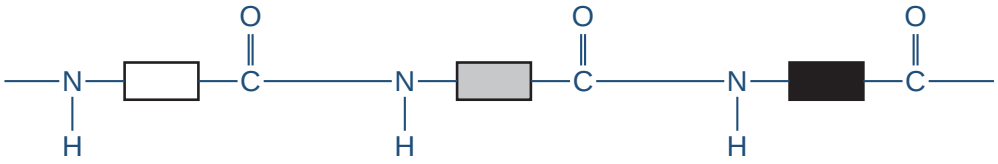


Fig. 6.4

The original Fig.6.4 order and positions of A's white block, B's grey block and C's black block are all retained. Blue additions complete the linkages and terminal continuation bonds.

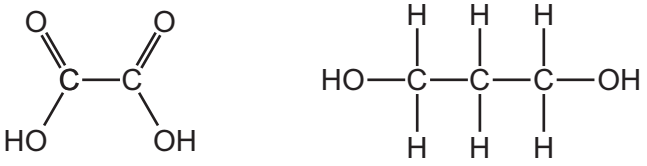
Check both internal connections, A—B and B—C: each must be C(=O)—NH with consistent orientation. Each N retains one H and each C retains its double-bonded O. The NH at the far left and C=O at the far right connect onwards in the same direction, leaving continuation bonds.

The official three points credit complete amide linkages, consistent orientation of the linkages and ends, and polymer continuation bonds. The question asks for a segment of a long chain. Do not cap it with NH_2 and COOH to make an isolated tripeptide.

Practice

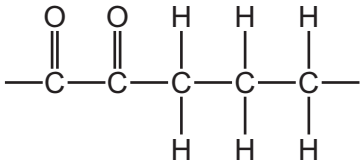
1. (2025·May/June 0620/21 Q37) · Level 3

The structures of two monomers that are used to make a condensation polymer are shown.

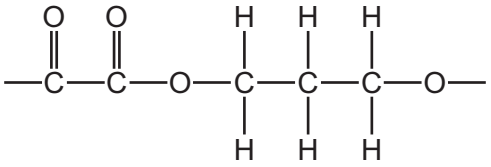


Which structure represents one repeat unit of the condensation polymer formed?

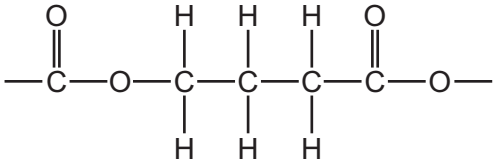
A



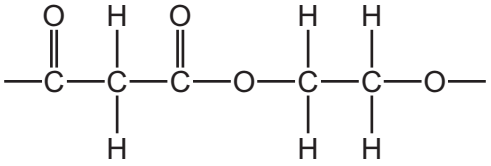
B



C

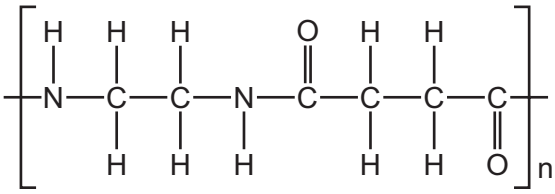


D

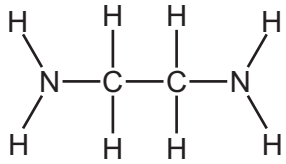
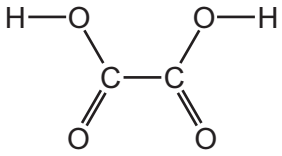
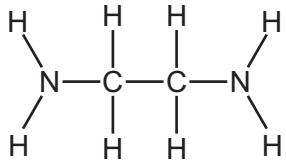
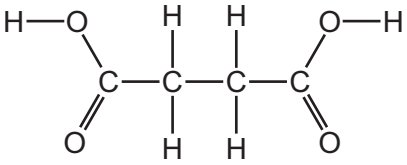
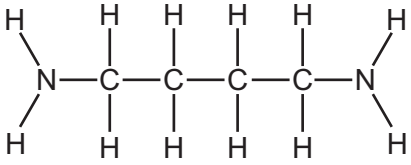
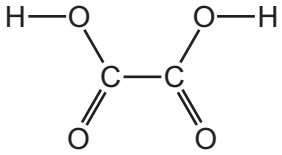
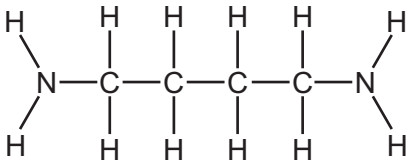
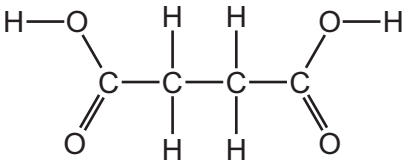


2. (2025·May/June 0620/22 Q37) · Level 3

The structure of a polymer repeat unit is shown.



Which pair of monomers is used to make this polymer?

	monomer 1	monomer 2
A		
B		
C		
D		

3. (2024·October/November 0620/42 Q6(b)) · Level 3

(b) PET is a polyester.

(i) Name the **two** types of monomer molecules needed to make polyesters.

(ii) Draw part of the structure of PET which shows **two** repeat units.

Show all of the atoms and all of the bonds in the linkages.

4. (2024·October/November 0620/42 Q6(a)(vi)) · Level 3

(vi) One of the monomers which forms part of a natural polyamide has **three** carbon atoms.

Complete Fig. 6.3 to show the displayed formula of this monomer.

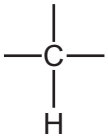


Fig. 6.3

5. (2023·October/November 0620/43 Q6(e)(ii)) · Level 3

(e) Polymers containing ester linkages are known as polyesters.

Polyamides are another type of polymer. Nylon is a polyamide.

The structure of nylon is shown in Fig. 6.2.

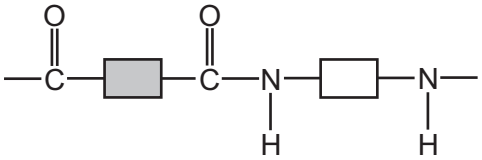


Fig. 6.2

(ii) Complete Fig. 6.3 to show the structures of the monomers used to produce nylon. Show all of the atoms and all of the bonds.

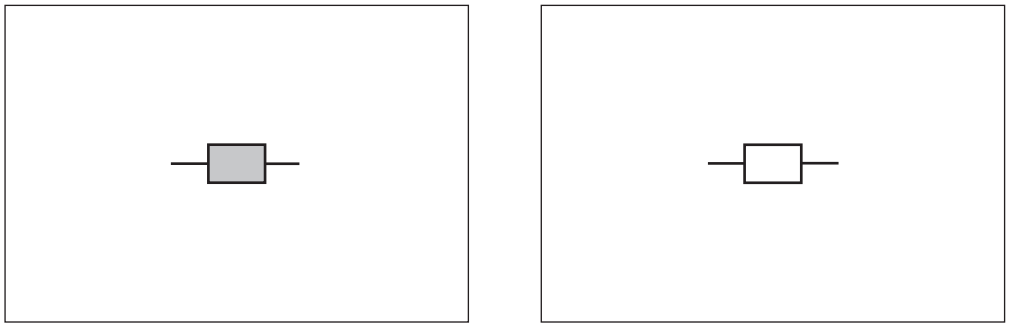


Fig. 6.3

6. (2023·May/June 0620/43 Q7(c)) · Level 3

This question is about organic compounds.

(c) A repeat unit of a condensation polymer is shown in Fig. 7.2.
The polymer is made from two monomers.

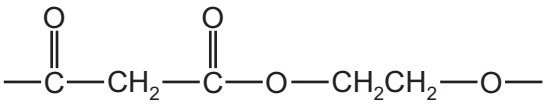


Fig. 7.2

(i) Draw the structures of the monomers used to produce the polymer in Fig. 7.2.

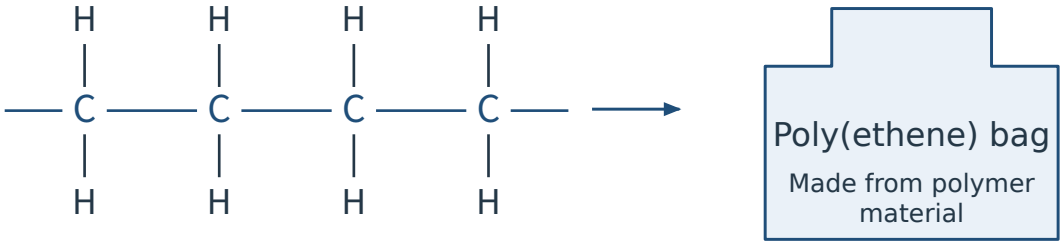
(ii) Name the **type** of condensation polymer in Fig. 7.2.

(iii) Name the **two** homologous series to which the monomers in **(i)** belong.

Type 3: Explain environmental problems from material properties and distinguish recycling routes

Key Ideas **Polymers are large molecules; plastics are materials made from them.** Poly(ethene), for example, is used in plastic bags and other products. One plastic bag contains many polymer molecules, so a short structural segment in a diagram does not represent the entire object. Properties such as durability and resistance to water help the material last during use and also affect what happens after disposal.

Plastics are made from polymers; long-chain structure relates to their uses

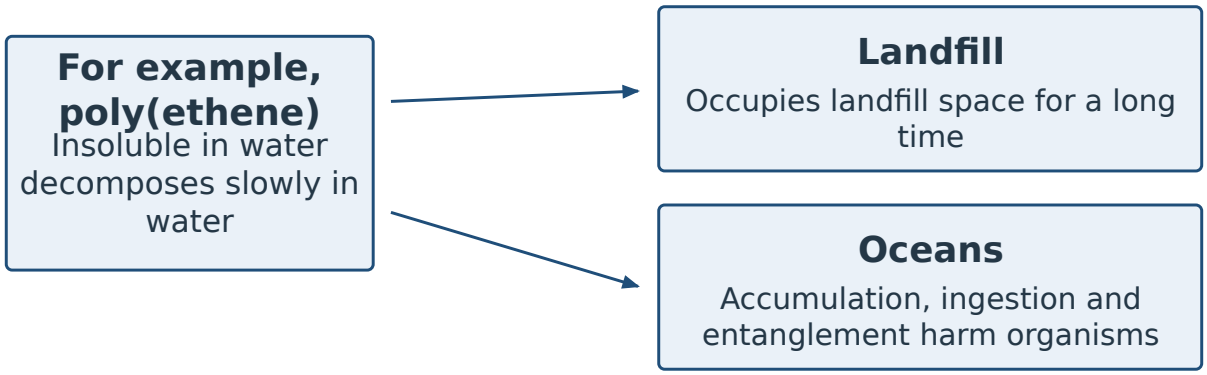


The left shows part of a poly(ethene) molecule; the right represents a plastic product made from many such molecules. Their sizes are not drawn to scale.

In plastics are made from polymers, polymers refers to the large molecules that make up the material. Answer ethene only if asked for the monomer from which poly(ethene) is made. Polymer, plastic product and monomer refer to three different levels; read the question before choosing a name.

Explain landfill and ocean accumulation using slow decomposition. Many common plastics decompose very slowly and occupy landfill space for long periods. Poly(ethene) also does not readily decompose in water, so it accumulates when it reaches the ocean. Animals can be injured by swallowing plastic or becoming entangled in it. Fragmentation into small pieces does not mean complete decomposition of the material.

Slow decomposition gives durability but also allows waste to accumulate



The arrows connect resistance to decomposition with two practical disposal problems. They do not identify the ocean as a suitable waste-disposal site.

To explain why poly(ethene) bags accumulate, write poly(ethene) does not decompose in water. This directly links ineffective removal to continued accumulation. Saying it comes from a hydrocarbon monomer describes

its origin without directly explaining accumulation. Saying it reacts with or dissolves in water does not match poly(ethene)'s properties here.

Burning can release harmful gases, and different plastics have different compositions. Poly(ethene) contains only C and H, so its main complete-combustion products are CO_2 and H_2O. Insufficient oxygen can produce toxic carbon monoxide, CO, and soot. Other plastics or additives may contain additional elements and therefore produce other harmful substances. For an environmental problem caused by burning plastics, explain that toxic gases can form.

Combustion products depend on the material's elements and the oxygen supply

Poly(ethene) contains only C and H	Sufficient oxygen; complete combustion	CO_2 and H_2O
Poly(ethene) contains only C and H	Insufficient oxygen; incomplete combustion	Toxic CO and soot may form
Plastics containing other elements	Depends on polymer and additives	Other harmful gases may also form

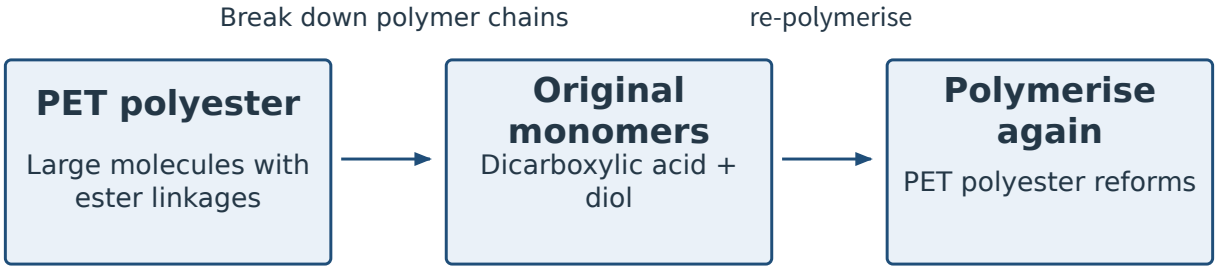
Pure poly(ethene) contains no Cl and cannot produce HCl; burning all plastics is not harmless.

CO is carbon monoxide and HCl is hydrogen chloride. The diagram uses conservation of elements and oxygen conditions; it does not claim identical combustion products for every plastic.

All waste plastics can be burned to give non-toxic products ignores both composition and combustion conditions and is unacceptable. Conversely, do not claim that pure poly(ethene) produces hydrogen chloride, HCl: the reactants contain no Cl, so that would violate conservation of elements. If asked for two disposal challenges, landfill space and ocean accumulation are two distinct answers. Toxic and harmful are not two separate points simply because different words are used.

One PET recycling route returns it to monomers before polymerising it again. PET is a polyester with recurring C(=O)—O ester linkages. Suitable treatment can convert it back to the corresponding dicarboxylic acid and diol, which can then undergo re-polymerisation. A dicarboxylic acid has two COOH groups and a diol has two alcohol OH groups, allowing both to join from two ends into long chains.

Chemical recycling of PET: break it into monomers, then polymerise again



PET is a polyester with C(=O)—O links, not a polyamide.

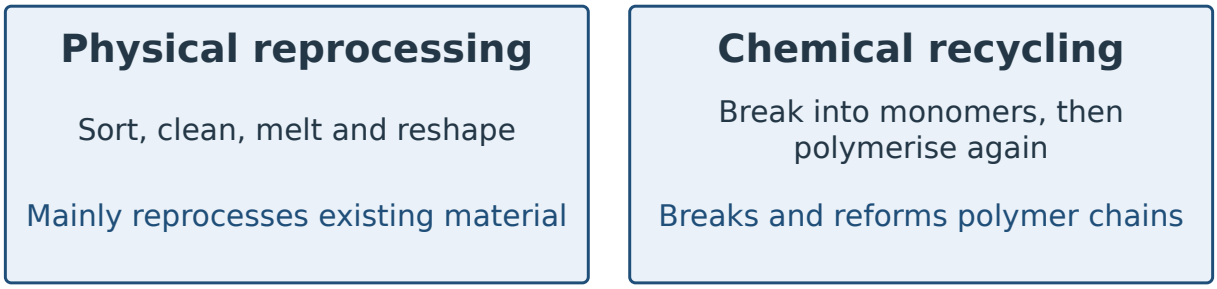
This is a chemical recycling route; it does not mean PET rapidly returns to monomers in the sea.

The arrows represent two chemical changes: breaking down to monomers and polymerising again. Memorising PET's full name is not required by this syllabus.

The statement PET can be broken down into its monomers and re-polymerised contains two actions: returning to monomers, then joining into long chains again. It is correct. PET contains oxygen in ester linkages and is a polyester. Do not confuse it with a polyamide containing C(=O)—NH or with a protein formed from amino acids. The existence of a controlled recycling route does not prove that PET rapidly decomposes on its own in seawater.

Check whether a recycling route actually returns the polymer to monomers. Sorting and cleaning a material, then melting and reshaping it, is one approach to physical reprocessing. Breaking its large molecules into monomers and polymerising them again changes and reforms the polymer chains. Both may be recycling routes, but their actions are not identical merely because both are called recycling.

Physical reprocessing and chemical recycling act at different levels



Monomers and re-polymerised in a PET question identify the route on the right.

The diagram distinguishes two approaches to recycling for interpreting questions. It does not imply that every plastic is suitable for either method.

If a question explicitly mentions monomers and re-polymerised, explain chain breakdown and chain re-formation. Melting alone describes only a physical process and does not account for the chemical recycling specified. If process details are not requested, do not invent temperatures, reagent concentrations or factory conditions.

Poly(ethene) accumulates in oceans because it does not readily _____ in water. Landfilled plastics _____ for long periods, and burning can produce _____. PET is a _____ it can first be converted back into _____ and then _____ again.

Answers: decompose; occupy landfill space; toxic gases; polyester; monomers; polymerised

Recognising the Question Identify how readily the material decomposes, then connect this to specific consequences in landfill, oceans or burning. If PET is described using monomers and re-polymerised, explain return to monomers followed by polymerisation. Keep ester and amide families distinct.

Worked Example 1 (2024·May/June 0620/21 Q37) · Level 2

Which statement about the polymer PET is correct?

A It can be broken down into its monomers and re-polymerised.

B It is an addition polymer.

C It is a polyamide.

D It is made from amino acid monomers.

Solution

PET is a polyester. Suitable chemical treatment can convert it back to its monomers, which can then be polymerised again. A, broken down into its monomers and re-polymerised, is therefore correct. Choose **A**.

B calls PET an addition polymer, which is false: PET forms by condensation of a diol and a dicarboxylic acid. C calls it a polyamide, which is false: it contains C(=O)—O ester linkages, not C(=O)—NH amide linkages. D says it forms from amino acid monomers, also false: amino acids form the natural polyamides called proteins studied here.

Being convertible to monomers by controlled treatment does not mean that PET rapidly decomposes naturally when discarded into the ocean. The question tests its capacity for recycling through chemical conversion.

Worked Example 2 (2025·October/November 0620/31 Q4(h)(ii)) · Level 2

(h) Ethene is polymerised to form the plastic poly(ethene).

(ii) Describe **two** environmental challenges caused by the disposal of plastics such as poly(ethene).

Solution

The question asks for two environmental challenges, so give two different consequences:

First, plastics occupy landfill space for long periods: **Plastics fill landfill sites because they decompose very slowly.** Second, plastics accumulate in oceans and may harm living organisms: **They accumulate in oceans and may harm aquatic organisms.**

These correspond to the official points fills landfill and accumulates in oceans, each an independent problem. Slow decomposition explains their persistence; do not repeat that same idea twice using different words.

The official mark scheme also lists the possible formation of toxic or acidic gases during burning. If using that point, take account of composition and combustion conditions. Poly(ethene) is the example here and contains only C and H, so do not invent Cl-containing products. The landfill and ocean points above already fully meet the request for two challenges.

Practice

1. (2025·May/June 0620/12 Q37) · Level 2

Which statement explains why plastic bags made from poly(ethene) accumulate in oceans?

A Poly(ethene) is a saturated molecule.

B Poly(ethene) is made from hydrocarbon monomers.

C Poly(ethene) reacts with water.

D Poly(ethene) does **not** decompose in water.

2. (2024·October/November 0620/11 Q38) · Level 2

Which statement about the disposal of waste plastics is correct?

A They are put in landfill sites, where they quickly decompose.

B They are burned to produce non-toxic products.

C They accumulate in oceans, where they are harmful to aquatic life.

D They are dissolved in water and pumped into the sea.

Experimental techniques and chemical analysis

12.1 Measuring apparatus and recording data

Type 1: Choose measuring apparatus, read scales accurately and record units and data tables

Key Ideas **Choose apparatus by first identifying what needs to be measured.** Time is commonly measured with a stop-watch in s, temperature with a thermometer in °C, and mass with a balance in g. When observing a changing process, distinguish elapsed time from temperature change.

Choose the instrument and unit after identifying what must be measured

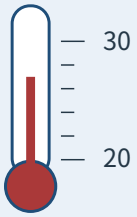
Time: stopwatch

02:13

2 min 13 s=133 s

Check whether the display is minutes:seconds or seconds

Temperature: thermometer



Usually recorded in °C

Mass: balance



Tare before measuring the sample

Convert time from the stopwatch display; the thermometer shows its shape only, with scale omitted.

Each column connects a physical quantity with its apparatus name and unit. The thermometer does not have a fully numbered scale and should not be used to infer a specific temperature.

A stop-watch displaying 02:13 in minutes:seconds means 2 minutes and 13 seconds. Convert the 2 minutes into seconds, then add 13:

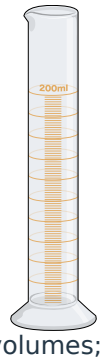
$$2 \times 60 + 13 = 133 \quad \text{s}$$

Do not write 2.13 s. If measuring how long a reaction takes from starting until a colour disappears, the main measuring instrument is a stop-watch. If each trial must also occur at 25°C, a thermometer checks that controlled condition. These are two different tasks.

For liquid volume, distinguish fixed volumes from variable volumes. A volumetric pipette accurately transfers one calibrated fixed volume, such as 25.0 cm³, using a pipette filler. It is not a graduated tube for any chosen volume, such as 26.3 cm³. A burette has fine graduations and a tap, allowing different volumes to be delivered accurately. The volume delivered is the final reading minus the initial reading. A measuring cylinder measures a range of volumes but is generally less precise than a suitable pipette or burette.

Liquid-volume measurement: distinguish fixed volumes from variable volumes

Measuring cylinder



Variable volumes; generally less precise

Not the first choice for accurately measuring 25.00 cm³

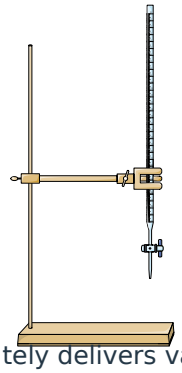
Volumetric pipette



Accurately transfers one fixed volume

Choose the calibrated volume; use a pipette filler

Burette



Accurately delivers variable volumes

Volume delivered = final − initial reading

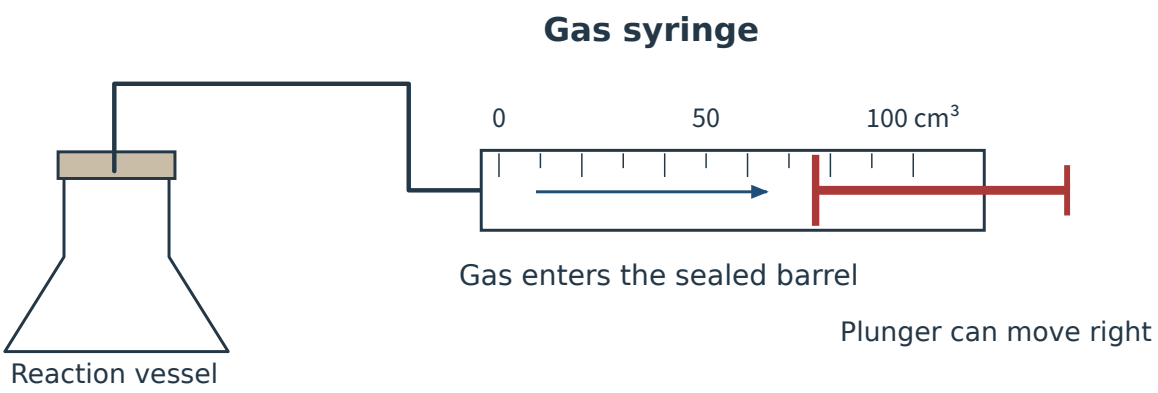
Measuring cylinder: DBCLS, CC BY 4.0; titration apparatus: Servier, CC BY 3.0; both via Bioicons.

The apparatus shapes are for identification; decorative graduations in these small drawings are not data for numerical questions. A single calibration line indicates the pipette's fixed volume.

To measure a fixed 25.0 cm³ accurately, use a 25.0 cm³ volumetric pipette if one is available. To deliver 26.3 cm³ accurately, use a burette. If a multiple-choice question offers only a beaker, measuring cylinder, burette and dropping pipette for measuring 25.00 cm³, choose the burette from those actual options. Beaker graduations usually give rough estimates. A dropping pipette mainly adds liquid drop by drop and does not have the corresponding accurate volume calibration. Do not confuse it with a volumetric pipette because both names contain pipette.

To measure the volume of gas produced, use apparatus that collects the gas completely. The plunger of a gas syringe moves as gas enters, allowing its volume to be read from the scale, commonly in cm³. It differs from a small pipette used to add liquid. Connections between the reaction vessel, delivery tube and syringe must be airtight. The syringe needs a sufficient range and a freely moving plunger; otherwise its reading does not represent all the gas produced.

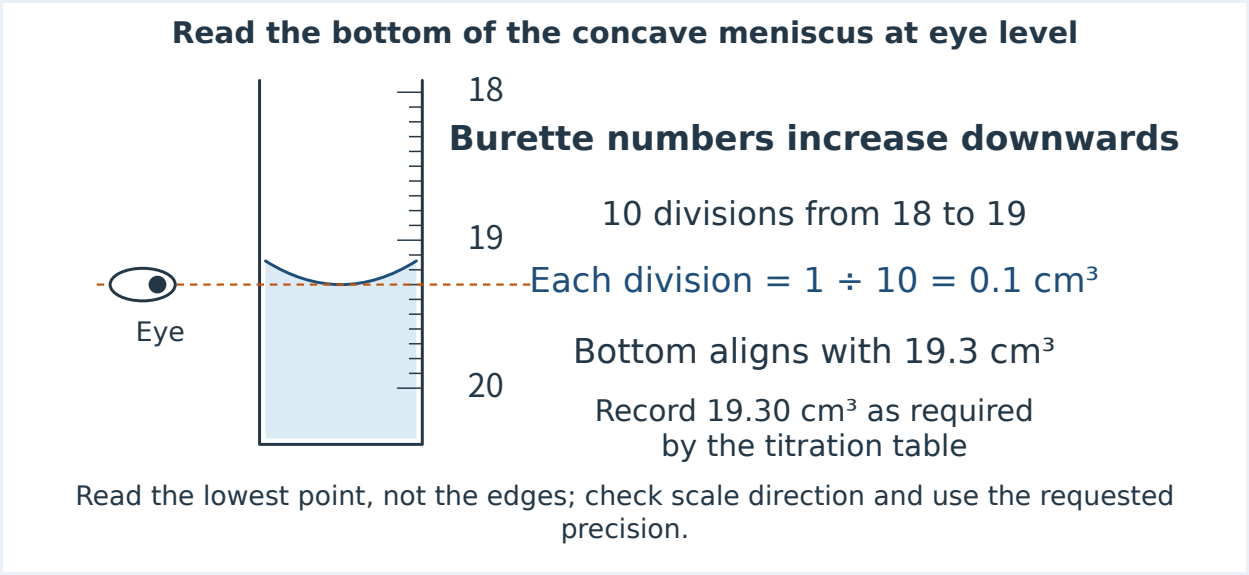
Gas syringe: incoming gas pushes the plunger so volume can be read directly



This shows connections and the direction of movement. The drawn plunger position is not source-question data. Connections must be airtight, the range sufficient and the plunger free to move.

If a reaction is expected to produce about 80 cm³ of gas, a suitable 100 cm³ gas syringe may be appropriate. A 20 cm³ syringe would quickly exceed its range. A leak allows some gas to escape without entering the syringe, giving a volume that is too low. Unlike collection over water, a gas syringe does not require the gas to pass through water. This can be advantageous for water-soluble gases, although the apparatus materials and conditions must still suit the particular gas.

Before reading a scale, find the value of one small division, then read the bottom of the meniscus. Divide the difference between two labelled values by the number of small intervals between them to find the smallest division. Transparent aqueous solutions usually form a concave meniscus in glass volumetric apparatus. Keep your eye level with its lowest point to reduce parallax error. Burette numbers usually increase downwards, unlike the upwards-increasing scale of a measuring cylinder.



The orange dashed line passes through the lowest point of the meniscus at 19.3. The scale follows the labelled values and increases in the direction used on a burette.

There are 10 small intervals between 18 and 19. Together they represent 1 cm³, so each interval is:

$$\frac{19 - 18}{10} = 0.1 \quad \text{cm}^3$$

The bottom of the meniscus is three divisions below 19, giving 19.3 cm³. If the titration table requires two decimal places, record 19.30 cm³. When the level lies between graduations, estimate only to the precision allowed by the question and instrument; for example, a burette with 0.1 cm³ divisions can be read to 0.05 cm³. Do not read the higher edges of the meniscus or reverse the scale direction to obtain 18.7.

Subtract the container when measuring sample mass. Taring resets the balance to zero with the empty container on it, so adding the sample gives its mass directly. Alternatively, record the empty-container mass and the container-plus-sample mass separately and subtract. Tare before adding the sample; do not zero the balance after adding it and report zero as the sample mass.

An empty container has mass: tare it or subtract its mass

Container + sample		Empty container		Sample alone
10.82 g	−	2.16 g	=	8.66 g

Sample mass = 10.82 − 2.16 = 8.66 g

After taring, the balance reads sample mass directly; do not include the container in the sample mass.

All three values belong to one mass measurement. The left gives the total mass, the middle the empty-container mass, and the right the sample mass found by subtraction.

If an empty weighing boat has mass 2.16 g and the boat plus sample has mass 10.82 g, subtract the boat from the total:

$$\text{Sample mass} = 10.82 - 2.16 = 8.66 \quad \text{g}$$

Using 10.82 g directly in n=m/M would include the container in the calculated amount of substance and give a value that is too large. Read the sample mass after taring, or calculate the mass difference before doing the chemistry calculation.

A data table must show what each number represents. Put both the quantity and unit in the heading, such as time/s, temperature/°C or volume/cm³. Use consistent decimal places for readings from the same instrument within a column; different columns may need different precision. Record actual raw readings first, then calculate differences or means or draw graphs. Do not replace measurements with imagined ideal results.

Data tables: put quantity and unit in the heading and use consistent precision within each column

Experiment number	Time / s	Temperature / °C	Gas volume / cm³
1	0	25.0	0.0
2	20	25.0	18.0
3	40	25.0	29.0

Record decimal places justified by the instrument; do not add unsupported digits to look precise.

The table gives a set of measurements with units in the headings. Time is recorded in whole seconds, while temperature and volume follow their respective stated resolutions.

In the second record of the table, 20 means 20 s, 25.0 means 25.0°C and 18.0 means a gas volume of 18.0 cm³. Because units are already in the headings, they need not be repeated in every cell. If the stop-watch is read to whole seconds, do not claim a measurement of 20.00 s merely to match the temperature column. Extra decimal places do not make an experiment more accurate.

A fixed 25.0 cm³ can be measured using the corresponding _____. Accurate variable liquid volumes can be measured with a _____, and gas volumes with a _____. Read a clear concave meniscus at its _____ point, with your eye _____. Sample mass equals _____ minus _____.

Answers: 25.0 cm³ volumetric pipette; burette; gas syringe; lowest; level with it; container-plus-sample mass; empty-container mass

Recognising the Question Identify time, temperature, mass or volume first. For liquid volume, decide whether it is fixed or variable. Check the scale divisions and direction before reading the meniscus at eye level. Calculate sample masses and titres using the appropriate differences.

Worked Example 1 (2025·May/June 0620/11 Q38) · Level 2

A student measures 25.00 cm^3 of dilute hydrochloric acid accurately.

Which piece of apparatus is most suitable?

A beaker

B measuring cylinder

C burette

D dropping pipette

Solution

The task is to measure 25.00 cm³ of dilute hydrochloric acid accurately. Among the four instruments actually supplied, identify one that accurately measures and delivers liquid volume.

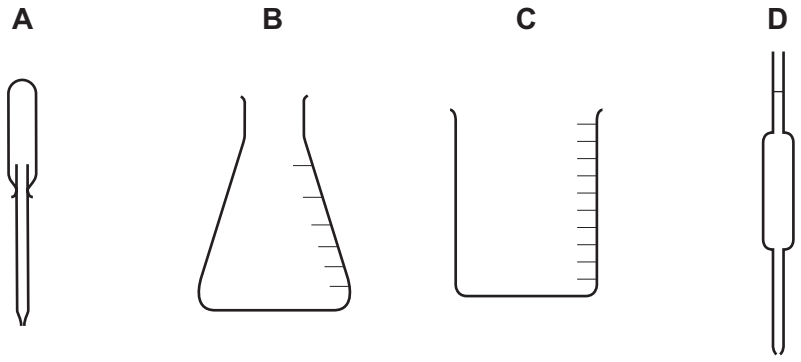
A beaker mainly holds liquids or estimates their volume. A measuring cylinder is generally not the first choice for this level of precision. A burette has fine graduations and a tap, and accurately gives delivered volume as final minus initial reading, so choose **C**. A dropping pipette adds liquid drop by drop; it is not a volumetric pipette with a calibrated fixed capacity.

If a 25.0 cm³ volumetric pipette were offered, it would also be suitable for an accurate fixed volume. It is not an option here, however, so choose from the original four instruments rather than adding your own.

Practice

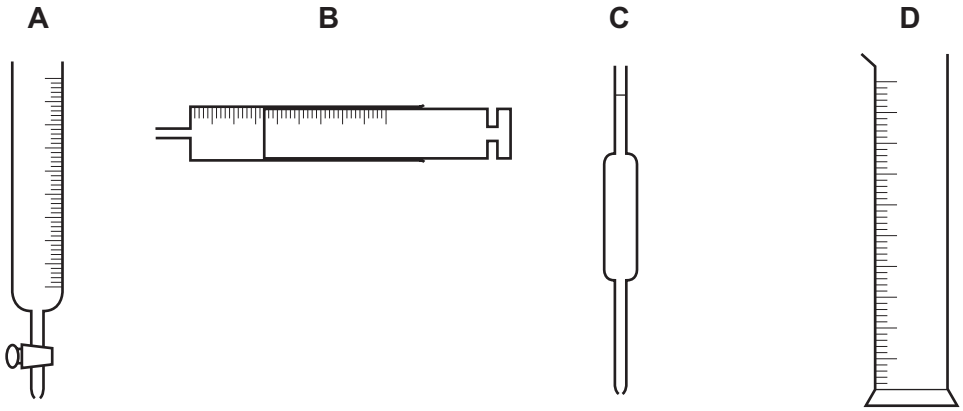
1. (2025·May/June 0620/21 Q38) · Level 2

Which item of apparatus is used to measure 25.0 cm^3 of aqueous sodium hydroxide?



2. (2025·May/June 0620/22 Q38) · Level 2

Which item of apparatus is used to measure exactly 26.3 cm^3 of a liquid?

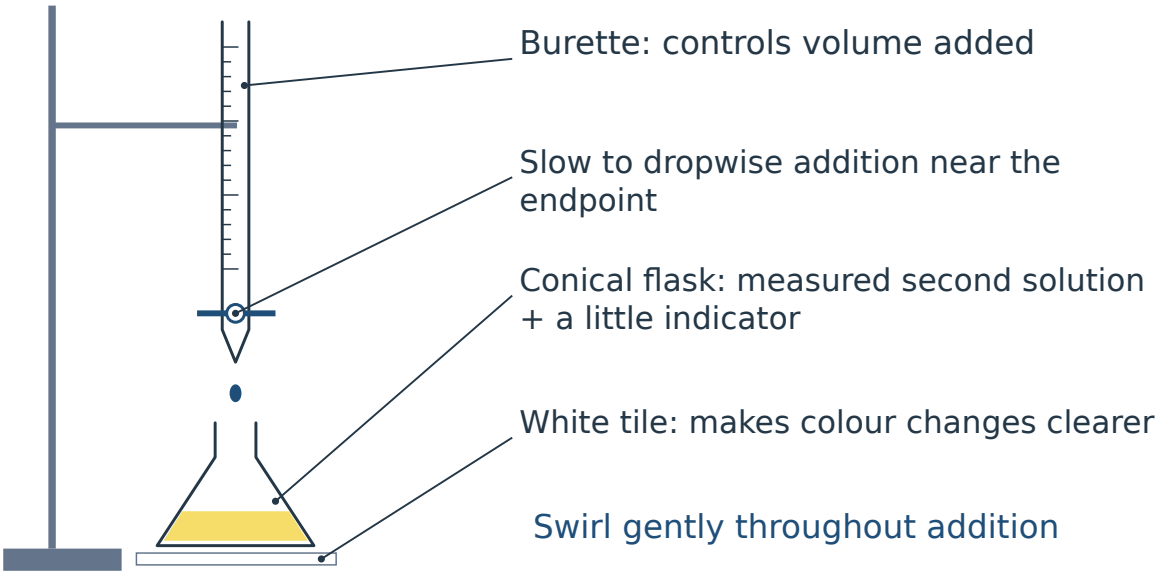


12.2 Titration, meniscus readings and end-points

Type 1: Use burette readings and indicators for quantitative reactions

Key Ideas **A titration accurately measures one solution, then finds the volume of another needed to react with it.** It is not simply mixing two beakers of liquid. Use a volumetric pipette to transfer a fixed sample volume accurately into a conical flask, then add a little suitable indicator. Put the other solution in a burette and add it slowly while gently swirling the flask. Near the end-point, add it drop by drop and stop at the specified permanent colour change.

Titration: add one solution dropwise to a measured volume of another



Burette and conical flask remain open; this apparatus sketch is not a numerical scale-reading diagram.

A volumetric pipette first measures the sample into the conical flask. The other solution is added dropwise from the burette. The indicator shows the end-point, and a white background helps the colour change to be seen.

To titrate 25.0 cm^3 of sodium hydroxide solution with hydrochloric acid, the volumetric pipette measures the 25.0 cm^3 accurately. The conical flask holds the alkali and indicator; its wide base and narrow neck make swirling convenient. The burette controls and measures the acid volume added. A white tile provides a pale background

that makes the end-point colour easier to see. Acid concentrated only where a drop lands has not mixed with the whole alkali sample, so a local colour cannot reliably represent the whole flask. Swirling therefore has a clear purpose. Use a pipette filler with the volumetric pipette; never pipette by mouth.

Rinse each instrument according to its role. Wash the burette and volumetric pipette with distilled water, then rinse each with the solution it will contain or measure. This prevents residual water from diluting an accurately measured solution. Wash the conical flask with distilled water only. A little water remaining in it changes the total volume but not the amount of reactant accurately transferred into it. Do not rinse it with the sample solution, because the residue would add an unmeasured amount of reactant.

Rinsing methods preserve the required concentration and amount of substance

Burette	Rinse with distilled water, then with the solution to be used Prevents residual water diluting the titrant
Volumetric pipette	Rinse with distilled water, then with the solution to be measured Preserves the transferred solution's concentration
Conical flask	Rinse with distilled water; add no extra unmeasured sample solution A little water does not change the measured moles

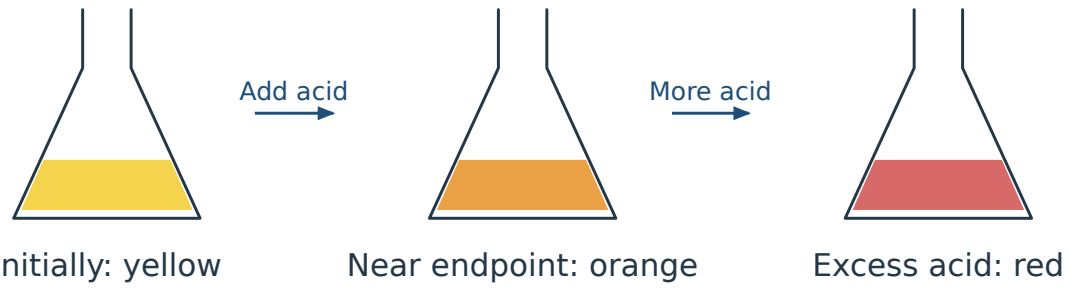
Also fill the burette tip with liquid, remove air bubbles and remove the filling funnel before reading.

The three rows compare the roles of different instruments. A solution rinse must match the solution subsequently contained or measured. Not all glassware should be rinsed with a reactant.

Suppose a volumetric pipette transfers 25.0 cm^3 of alkali into a conical flask already containing 2 cm^3 of distilled water. The alkali is diluted, but its total amount in mol is unchanged, so the required amount and titre of acid should remain the same. By contrast, an air bubble in the burette tip means that some initially delivered liquid fills the tip instead of entering the flask. The recorded volume may therefore be too large. Expel tip bubbles before starting. Remove the filling funnel before taking readings so that residual liquid in it cannot run in and change the reading.

Judge the end-point colour from the indicator, the starting acidity or alkalinity and the direction of addition. Methyl orange is red in acid, yellow in alkali and orange within its transition range. Common questions involving acid added to an alkali or carbonate solution use the change from yellow to permanent orange as the end-point; excess acid would make it redder. This colour sequence does not apply to every titration or indicator. If the original question specifies a mixed indicator and several end-points, record the changes stated there.

Methyl orange: adding acid to alkali changes yellow to orange at the endpoint



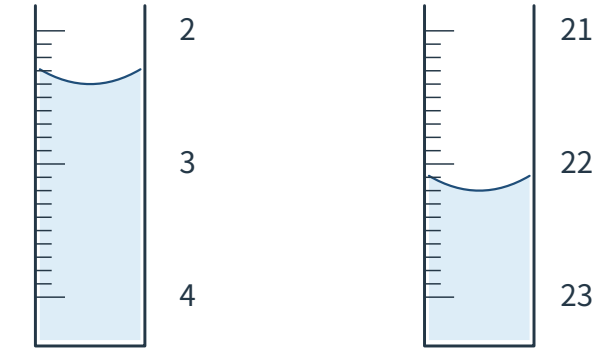
Use the persistent orange endpoint requested here and avoid adding excess acid.
If alkali is added to acid instead, reverse the starting colour and direction of change.

This diagram assumes methyl orange and acid added to an alkaline solution. Not all indicators have these three colours. The colours are illustrative and do not represent exact concentrations or pH values.

Suppose the flask contains potassium carbonate solution with methyl orange and the burette contains dilute sulfuric acid. The flask is initially alkaline and yellow. Add acid slowly while swirling and stop at the required orange colour. Do not write red as the initial colour just because acid is being added: the initial observation concerns the solution already in the flask. Thymolphthalein can be blue on the alkaline side and colourless on the acidic side. Whether to use it and which colour marks the stopping point depend on the experiment and instructions. Do not copy methyl orange's colours.

The titre is the actual volume added: final reading minus initial reading. Determine the smallest scale division, then read with your eye level with the bottom of the meniscus. Burette numbers increase downwards, so delivering liquid normally increases the reading. The initial reading need not be 0; do not alter a genuine reading to 0 just to simplify a calculation. Use the table's units and the precision required by the instrument and question, keeping decimal places consistent for the same type of reading in a column.

A titre is a difference between readings, not the final reading alone



Volume of liquid delivered

= final − initial reading

=22.20−2.40

=19.80 cm^3

Both scales have ten divisions per 1 cm³. The meniscus bottoms are at 2.40 and 22.20 respectively. The coloured background does not change the point that should be read.

The initial reading is 2.40 cm³ and the final reading is 22.20 cm³. Subtract the initial reading from the final reading to find the volume added:

$$22.20 - 2.40 = 19.80 \quad \text{cm}^3$$

22.20 is the final position on the scale, not the volume added since this titration began. The initial 2.40 must be subtracted. The scale has ten intervals per 1 cm³, and the positions of the meniscus bottoms are shown. Here the readings are recorded to two decimal places. If an exam question requires at least one decimal place and its diagram reads directly to 0.1 cm³, complete its table consistently at that specified precision.

Select close repeat results obtained under the same conditions before calculating a mean. A rough titration locates the approximate end-point, so later accurate repeats can be slowed at the right stage. Concordant results must use the same solution concentrations, sample volume, reaction and end-point criterion. If the question requires the rough result or a clear anomaly to be excluded, apply that rule first. Do not average every number just because it appears in the table.

Before averaging, select close, reliable repeats measured under the same conditions

Rough titration	25.00 cm^3	Locates the approximate endpoint; excluded from this average
Accurate repeat 1	24.10 cm^3	Use
Accurate repeat 2	24.15 cm^3	Use
Accurate repeat 3	24.10 cm^3	Use

Here results must differ by no more than 0.10 cm³; the three accurate titres span 0.05 cm³.

Mean = (24.10 + 24.15 + 24.10) ÷ 3 = 24.116... ≈ 24.12 cm³

Accurate titrations are repeated under the same conditions. This example specifies its own agreement criterion; follow the rule in the actual question. Do not include results from different acid concentrations or different end-points in this mean.

For example, suppose three accurate titres must differ by no more than 0.10 cm³. The maximum difference among 24.10, 24.15 and 24.10 is 0.05 cm³, so they qualify. Calculate the mean of these three reliable repeats:

$$\text{Mean titre} = \frac{24.10 + 24.15 + 24.10}{3} = \frac{72.35}{3} \approx 24.12 \quad \text{cm}^3$$

The unrounded result is 24.116… cm³, which becomes 24.12 cm³ to two decimal places. The rough titre of 25.00 is excluded. The 0.10 cm³ criterion is specified for this example; it is not the sole threshold for every examination question. A different acid concentration or a second, different end-point represents different experimental conditions and must not be mixed with these accurate repeats.

For two end-points in one titration, distinguish cumulative volume from the additional volume between them. If liquid is added continuously without refilling the burette, calculate both cumulative volumes from the same original reading:

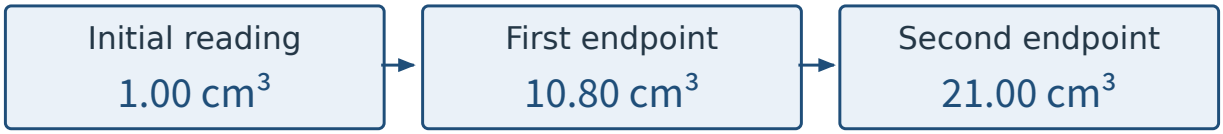
$$\text{Cumulative volume to first end-point} = \text{First end-point reading} - \text{Original reading}$$

$$\text{Total volume to second end-point} = \text{Second end-point reading} - \text{Original reading}$$

Total volume means the cumulative amount added. Only if asked for the volume between the end-points should you calculate the additional middle portion:

$$\text{Additional volume between end-points} = \text{Second end-point reading} - \text{First end-point reading}$$

Two endpoints: total volume always starts from the original reading



To first endpoint: 10.80 − 1.00 = 9.80 cm³

Total to second endpoint: 21.00 − 1.00 = 20.00 cm³

Added between endpoints: 21.00 − 10.80 = 10.20 cm³

Total volume asks for the overall addition; between the endpoints asks only for the extra interval.

The titration is continuous, with no refilling partway through. Each difference has a clearly identified starting and finishing reading.

The original reading is 1.00, the first end-point reading 10.80 and the second end-point reading 21.00, all in cm³.
The total volume added to the first end-point is:

$$10.80 - 1.00 = 9.80 \quad \text{cm}^3$$

The cumulative volume added to the second end-point is:

$$21.00 - 1.00 = 20.00 \quad \text{cm}^3$$

The additional volume between the end-points is:

$$21.00 - 10.80 = 10.20 \quad \text{cm}^3$$

Check by adding the two successive portions:

$$9.80 + 10.20 = 20.00 \quad \text{cm}^3$$

If the table asks for total volume to reach second end-point, entering 10.20 omits the first portion. For separate experimental columns, subtract each experiment's own initial reading, not the starting value from a neighbouring column.

Check reaction coefficients before using titres to compare concentrations. Amount concentration is the amount of dissolved solute divided by solution volume:

$$c = \frac{n}{V}$$

Here c is amount concentration, n is the amount of solute and V is solution volume. If concentration is in mol/dm³, convert V to dm³. The amounts of the two reactants are equal only when they react exactly in a 1:1 ratio; in that case, the solution requiring the larger volume is more dilute. For coefficients in a ratio such as 1:2, convert amounts using the balanced equation first. Volume alone is insufficient.

Check the reaction's mole ratio before comparing concentrations

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Hydrochloric acid	Sodium hydroxide solution
$20.0\ cm^3$	$25.0\ cm^3$
Same moles in a smaller volume: more concentrated	Same moles in a larger volume: more dilute

This equation has a 1:1 ratio, so n(HCl) = n(NaOH) at exact reaction; compare using c = n/V.

If the ratio is not 1:1, use the coefficients first; volume alone is insufficient.

This example specifies a 1:1 stoichiometric ratio and exact reaction. The smaller-volume solution is more concentrated, but this conclusion must follow from the balanced equation.

Suppose 20.0 cm³ of hydrochloric acid exactly neutralises 25.0 cm³ of 0.100 mol/dm³ NaOH solution. First convert the NaOH volume to 0.0250 dm³, then calculate its amount:

$$n(\mathrm{NaOH}) = 0.100 \times 0.0250 = 0.00250 \quad \mathrm{mol}$$

The reaction equation is:

$$\mathrm{HCl + NaOH \rightarrow NaCl + H_2O}$$

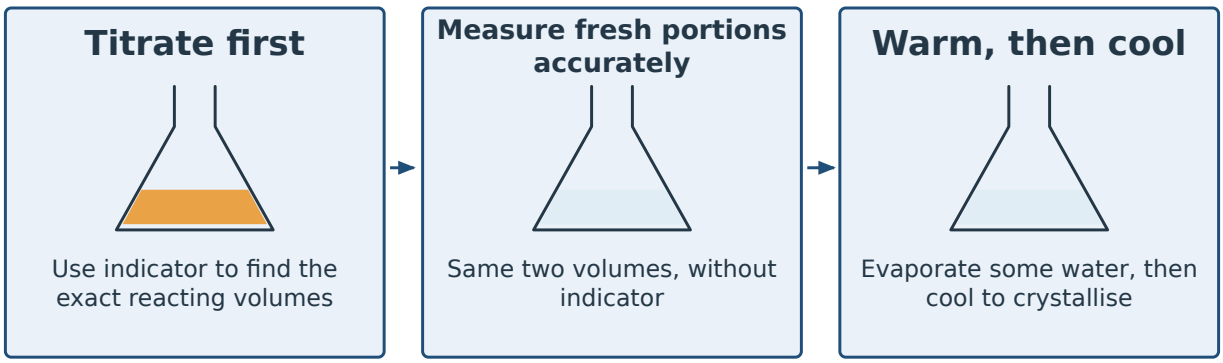
HCl and NaOH react in a 1:1 mole ratio, so the acid also contains 0.00250 mol of HCl. Its volume is 0.0200 dm³, giving:

$$c(\mathrm{HCl}) = \frac{0.00250}{0.0200} = 0.125 \quad \frac{\mathrm{mol}}{\mathrm{dm^3}}$$

0.125 is greater than 0.100, consistent with the same amount of solute occupying a smaller volume at higher concentration. Substituting 20.0 directly into a formula requiring dm³ would give a factor-of-1000 error.

To prepare a pure salt, use a titration to find the volumes, then remake the mixture without indicator. The indicator in the first titration is dissolved and cannot be removed by ordinary filtration as though it were an insoluble solid. React the same accurately measured volumes again, but omit the indicator. Warm the mixture to evaporate some water, then stop at a suitable concentration or crystallisation point and cool to form salt crystals. Finally, as required, filter or remove the crystals, wash them with a little suitable cold solvent and dry them.

For pure salt crystals, do not treat the first indicator-containing mixture as pure



Collect, suitably wash and dry the crystals; use volumes from the accurate titration.

The first mixture contains indicator and establishes the volumes. A fresh second mixture contains no indicator and is concentrated and crystallised. Dissolved indicator is not treated as a solid removable by filtration.

Suppose the first titration shows that 25.0 cm³ of alkali needs exactly 19.80 cm³ of acid. Measure these volumes again using solutions of the same concentrations, then mix them without indicator. The first experiment has already established the exact reacting volumes, so no further colour end-point is needed. Evaporate some water and cool to crystallise. Heat alone does not describe concentration, cooling and obtaining crystals as separate steps. Keep the first mixture used to find volumes distinct from the second mixture used to prepare the pure product.

Before titration, use a _____ to transfer the fixed sample volume accurately. Near the end-point, add the other solution _____ while continuing to _____. Titre=_____–_____. Both cumulative end-point volumes subtract _____. For a mean, use reliable repeat results that are _____. To prepare a pure salt, mix the measured volumes again and _____, then concentrate and _____.

Answers: volumetric pipette; drop by drop; swirl; final reading; initial reading; the original reading for that experiment; close together and obtained under the same conditions; omit the indicator; cool to crystallise

Recognising the Question Identify what the burette and flask contain and which indicator is used. Read the initial and final meniscus bottoms for the same experiment and subtract. With two end-points, distinguish cumulative total volume from the additional volume between them. Results using changed concentrations or indicators are not equivalent repeats for averaging.

Worked Example 1 (2025·May/June 0620/61 Q1) · Level 3

A student prepares some crystals of pure potassium sulfate by reacting aqueous potassium carbonate with dilute sulfuric acid. The equation for the reaction is shown.

$$K_2CO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)$$

The student:

- places 25.0 cm^3 of aqueous potassium carbonate in a conical flask
- adds a few drops of a suitable indicator to the conical flask
- uses a burette to add dilute sulfuric acid to the conical flask until the end-point is reached and the indicator changes colour.

Table 1.1 shows the burette readings the student obtains.

Table 1.1

final burette reading / cm^3	18.8
initial burette reading / cm^3	4.2
volume of dilute sulfuric acid added / cm^3	

(a) Complete Table 1.1 by calculating the volume of dilute sulfuric acid added to the conical flask.

(b) State which solution, aqueous potassium carbonate or dilute sulfuric acid, is the **least** concentrated. Give a reason for your answer.

least concentrated solution

reason

(c) Name a piece of apparatus suitable for measuring the 25.0 cm^3 of aqueous potassium carbonate.

(d) Name a suitable indicator and give the colour change of this indicator at the end-point.

indicator

colour change at end-point to

(e) As the student added the dilute sulfuric acid to the aqueous potassium carbonate they looked for a colour change.

State what the student should do as they add the dilute sulfuric acid to the aqueous potassium carbonate in the conical flask. Do **not** include observations in your answer.

(f) Describe what the student should now do to obtain crystals of pure potassium sulfate.

Solution

First read the experiment and equation. The conical flask contains 25.0 cm³ of potassium carbonate solution and the burette contains dilute sulfuric acid. The given reaction is:

$$K_2CO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)$$

Potassium carbonate and sulfuric acid react in a 1:1 mole ratio. Complete each part from (a) to (f).

(a) The initial burette reading is 4.2 cm³ and the final reading is 18.8 cm³. The acid volume added is:

$$18.8 - 4.2 = 14.6 \quad \text{cm}^3$$

The completed original table is shown below.

Table 1.1

final burette reading / cm^3	18.8
initial burette reading / cm^3	4.2
volume of dilute sulfuric acid added / cm^3	**14.6**

Completed original Table1.1: 18.8−4.2=14.6 cm³. Blue shows added answers; the black table, units and given readings are retained.

(b) The more dilute solution is **potassium carbonate solution**. At exact reaction, the 1:1 coefficients mean equal amounts in mol. Those equal amounts occupy 25.0 cm³ and 14.6 cm³ respectively. Potassium carbonate occupies the larger volume and therefore has the lower concentration. This follows from this reaction ratio and these volumes, not from a rule that an alkali is always more dilute than an acid.

(c) A **volumetric pipette** accurately measures 25.0 cm³. The mark scheme for this question also accepts a burette. Name the instrument, rather than merely saying measure the volume.

(d) Use **methyl orange**, with **yellow to orange** marking the end-point. Potassium carbonate solution is initially alkaline, so the indicator is yellow. Add acid slowly and stop at the required orange colour. Do not use red merely because acid is being added, or replace a specific colour with clear.

(e) **Swirl the conical flask** gently throughout acid addition, mixing each drop with the whole potassium carbonate sample. This part asks for an action, not a separate observation.

(f) Accurately measure fresh portions of 25.0 cm³ of the same potassium carbonate solution and 14.6 cm³ of the same dilute sulfuric acid. Mix them **without indicator**. Warm to evaporate some water, stop heating at a suitable concentration, saturation or crystallisation point, then **cool** to form potassium sulfate crystals. The crystals can then be collected, suitably washed and dried. The official 3 marks focus on remaking the reaction with the established volumes and no indicator, concentrating by evaporation, and cooling to crystallise. The first mixture contains indicator and should not directly supply pure salt crystals. Ordinary filtration cannot remove the dissolved indicator.

Worked Example 2 (2025·October/November 0620/62 Q2(a)) · Level 3

A student investigates the reaction between acidic solution **B** and two different solutions of aqueous sodium hydroxide, solution **C** and solution **D**, using two different indicators.

The student does two experiments.

Experiment 1

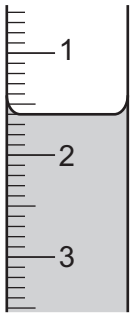
- Rinse a burette with distilled water and then with solution **C**.
- Fill the burette with solution **C**. Run some of solution **C** out of the burette so that the level of the solution is on the burette scale.
- Record the initial burette reading.
- Use a measuring cylinder to pour 25 cm^3 of solution **B** into a conical flask.
- Add five drops of methyl orange indicator **and** five drops of thymolphthalein indicator to the conical flask.
- Stand the conical flask on a white tile.
- Slowly add solution **C** from the burette to the conical flask, while swirling the flask, until the solution changes colour from red to orange. This is the first end-point.
- Record the burette reading at the first end-point.
- Continue to add solution **C** from the burette to the conical flask while swirling the flask. The solution changes colour from orange to yellow.
- Continue to add solution **C**, while swirling the flask, until the solution changes colour from yellow to green. This is the second end-point.
- Record the burette reading at the second end-point.

Experiment 2

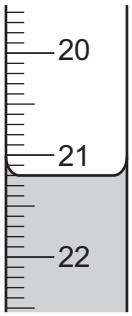
- Empty the conical flask and rinse it with distilled water.
- Repeat Experiment 1 using solution **D** instead of solution **C**.

(a) Use the burette diagrams in Fig. 2.1 and Fig. 2.2 to record the readings for Experiment 1 and Experiment 2 in Table 2.1. Complete Table 2.1.

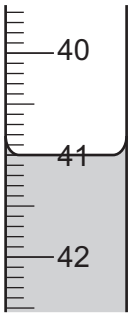
Experiment 1



initial burette reading



burette reading at first end-point



burette reading at second end-point

Fig. 2.1

Experiment 2



initial burette reading



burette reading at first end-point



burette reading at second end-point

Fig. 2.2

Table 2.1

	Experiment 1 using solution **C**	Experiment 2 using solution **D**
burette reading at first end-point / cm^3		
burette reading at second end-point / cm^3		
initial burette reading / cm^3		
volume added from burette to reach first end-point / cm^3		
total volume added from burette to reach second end-point / cm^3		

Solution

Record the two end-points specified in this question separately. The conical flask contains acid B with a mixed indicator, and alkali C or D is added. The original instructions specify a first red-to-orange change, followed on continued addition by orange to yellow, and then the second yellow-to-green change. Follow that stated sequence. Do not import the usual methyl-orange yellow-to-orange sequence for acid added to alkali. The whole of (a) requires six meniscus readings and a table with two columns and five rows.

Experiment 1, solution C. Read the meniscus bottoms: initial **1.6 cm³**, first end-point **21.2 cm³**, second end-point **41.0 cm³**. Subtract the original 1.6 cm³ from both end-point readings:

$$\text{To first end-point} = 21.2 - 1.6 = 19.6 \quad \text{cm}^3$$

$$\text{To second end-point} = 41.0 - 1.6 = 39.4 \quad \text{cm}^3$$

Experiment 2, solution D. The initial reading is **0.4 cm³**, the first end-point **10.2 cm³** and the second **20.1 cm³**. Subtract this experiment's original 0.4 cm³ from both:

$$\text{To first end-point} = 10.2 - 0.4 = 9.8 \quad \text{cm}^3$$

$$\text{To second end-point} = 20.1 - 0.4 = 19.7 \quad \text{cm}^3$$

Complete the original table in its given row order as follows.

Table 2.1

	Experiment 1 using solution **C**	Experiment 2 using solution **D**
burette reading at first end-point / cm^3	**21.2**	**10.2**
burette reading at second end-point / cm^3	**41.0**	**20.1**
initial burette reading / cm^3	**1.6**	**0.4**
volume added from burette to reach first end-point / cm^3	**19.6**	**9.8**
total volume added from burette to reach second end-point / cm^3	**39.4**	**19.7**

Completed original Table2.1. Both end-point volumes subtract the initial reading for the same experiment. All readings and volumes use one decimal place, with cm³ retained from the original table.

The first two rows require the scale readings at the end-points; the third requires the initial reading. Do not fill them in the order in which you inspected the diagrams. The final row explicitly asks for **total volume**, measured cumulatively from that experiment's original reading.

Subtracting the first end-point from the second would instead give:

$$41.0 - 21.2 = 19.8 \quad \text{cm}^3$$

$$20.1 - 10.2 = 9.9 \quad \text{cm}^3$$

These are only the extra volumes between the end-points, not the cumulative volumes required in the total volume row.

All six readings and four calculated volumes are recorded to one decimal place, meeting the requirement for at least one decimal place and consistent recording. The official 5 marks cover the two sets of correct meniscus readings, table placement, four correct volumes and consistent precision. The columns use different solutions, C and D, and the two end-points represent different stages. They are not four equivalent repeat titrations and must not be combined into a mean titre.

Practice

1. (2024·May/June 0620/61 Q1) · Level 2

A student carries out a titration to find the concentration of a sample of dilute hydrochloric acid.

The student:

- adds 25.0 cm^3 of aqueous potassium hydroxide to the apparatus labelled **A** in Fig. 1.1

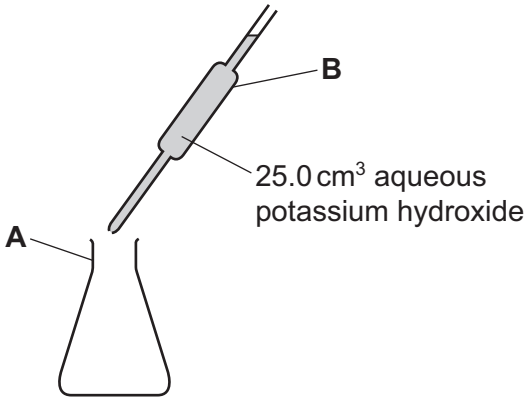


Fig. 1.1

- adds a few drops of a suitable indicator to the apparatus labelled **A**
- uses a burette to add dilute hydrochloric acid to the aqueous potassium hydroxide and indicator mixture in the apparatus labelled **A**.

(a) Name the items of apparatus labelled **A** and **B** in Fig. 1.1.

A

B

(b) The student adds the indicator after the volume of the 25.0 cm^3 of aqueous potassium hydroxide has been measured.

(i) Explain why the student adds an indicator to the aqueous potassium hydroxide.

(ii) Name a suitable indicator.

(c) Describe how the student can determine the volume of dilute hydrochloric acid used in this titration.

(d) The student observes the colour changes that occur as they add dilute hydrochloric acid from the burette.

State one **other** thing the student should do as they add the dilute hydrochloric acid to the aqueous potassium hydroxide.

2. (2024·February/March 0620/62 Q2(a)) · Level 3

A student investigates the reaction between aqueous sodium carbonate and two different solutions of dilute hydrochloric acid, labelled **A** and **B**.

The student does three experiments.

Experiment 1

- Rinse a burette with distilled water and then with dilute hydrochloric acid **A**.
- Rinse a conical flask with distilled water.
- Fill the burette with dilute hydrochloric acid **A**. Run some of the dilute hydrochloric acid out of the burette so that the level of the dilute hydrochloric acid is on the burette scale.
- Record the initial burette reading.
- Use a measuring cylinder to pour 25 cm^3 of aqueous sodium carbonate into the conical flask.
- Add five drops of methyl orange indicator to the conical flask.
- Stand the conical flask on a white tile.
- Slowly add dilute hydrochloric acid **A** from the burette to the conical flask, while swirling the flask, until the solution becomes orange.
- Record the final burette reading.

Experiment 2

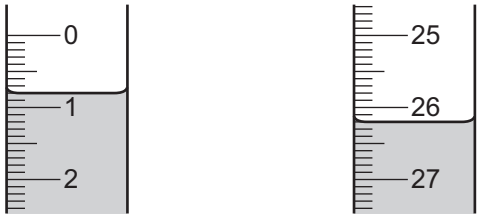
- Refill the burette with dilute hydrochloric acid **A**. Run some of the dilute hydrochloric acid out of the burette so that the level of the dilute hydrochloric acid is on the burette scale.
- Record the initial burette reading.
- Empty the conical flask and rinse it with distilled water.
- Use the measuring cylinder to pour 25 cm^3 of aqueous sodium carbonate into the conical flask.
- Add five drops of thymolphthalein indicator to the conical flask.
- Stand the conical flask on a white tile.
- Slowly add dilute hydrochloric acid **A** from the burette to the conical flask, while swirling the flask, until the solution becomes colourless.
- Record the final burette reading.

Experiment 3

- Repeat Experiment 1, using dilute hydrochloric acid **B** instead of dilute hydrochloric acid **A**.

(a) Use the burette diagrams in Fig. 2.1, Fig. 2.2 and Fig. 2.3 to record the readings for Experiments 1, 2 and 3 in Table 2.1 and complete Table 2.1.

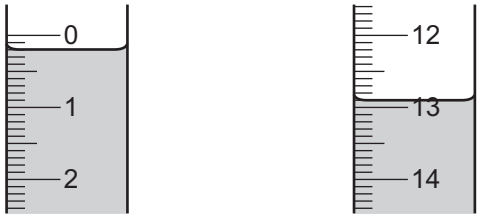
Experiment 1



initial reading final reading

Fig. 2.1

Experiment 2



initial reading final reading

Fig. 2.2

Experiment 3



initial reading final reading

Fig. 2.3

Table 2.1

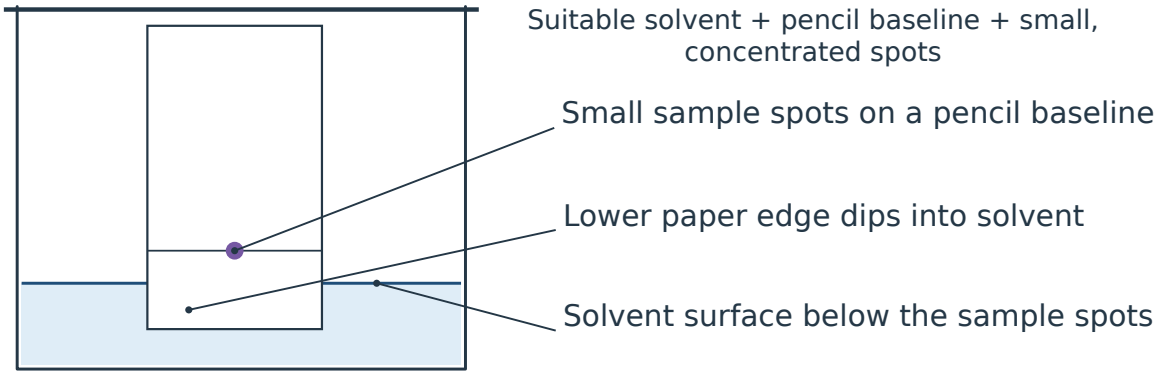
	Experiment 1	Experiment 2	Experiment 3
final burette reading / cm^3			
initial burette reading / cm^3			
volume of dilute hydrochloric acid added / cm^3			

12.3 Paper chromatography

Type 1: Identify mixtures using chromatography apparatus, spots and Rf values

Key Ideas **Let the solvent move along the paper, then see how it separates a mixture.** Paper chromatography separates substances that dissolve in a suitable solvent. Draw a baseline in pencil, place a small sample spot on it and let it dry. Put the lower end of the paper into the solvent, keeping the initial solvent level below the sample spot. Pencil graphite is usually insoluble in the solvent. An ink baseline may introduce additional coloured components that move with the solvent.

At the start, the paper touches solvent but the sample spots stay above the liquid



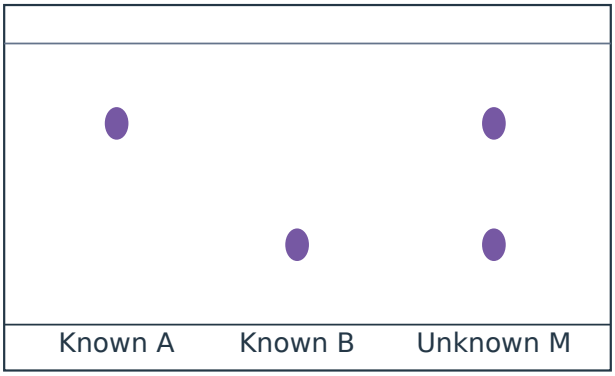
Do not submerge samples in the solvent pool; cover the container to reduce evaporation.

The bottom of the paper enters the solvent while the baseline and sample spot remain above the initial solvent level. This is the starting arrangement, not the final spot positions.

If the solvent covers the sample spot, the sample may dissolve directly into the reservoir instead of separating upwards from a concentrated starting point. If the solvent level is so low that it does not reach the paper, it cannot rise through it. Both conditions matter: the paper must touch the solvent and the sample spot must be above the liquid level. After development, remove the paper and promptly mark the solvent front in pencil before evaporation makes its position disappear.

The number of spots and their positions provide different evidence. With the same paper, solvent and conditions, a pure substance commonly gives one spot, while a mixture that separates gives several. An unknown spot at the same height as a known substance supports the possibility of a shared component. Compare chromatograms under the same development conditions; do not directly compare results obtained with different solvents.

Under the same conditions, compare spot counts and relative positions



A and B each give one spot

M gives two spots, showing a mixture

M's spots match the heights of A and B

Supports A and B being present under these conditions

One spot can contain unresolved components; matching positions support an identity but do not prove it alone.

The three samples develop in the same solvent on the same paper under the same conditions. The two unknown spots align with known A and B respectively.

M has two separate spots, so it is a mixture. Their heights match known A and B, supporting the conclusion that M contains A and B. In a simple question, one spot supports a pure substance, but remember the overlap discussed later: two substances may fail to separate in that solvent. One spot is therefore not absolute proof of purity. Similar colour also does not replace comparing positions and conditions.

Colourless substances can separate too, but their positions must be made visible. Some amino acids in these questions are colourless. Develop the chromatogram in a suitable solvent first, then use a locating agent to make the spots visible. A particular locating-agent name need not be memorised, but explain that its function is to reveal where the separated substances are.

Colourless substances: separate first, then use a locating agent to reveal their positions



A locating agent reveals separated positions; it does not force components with identical Rf values apart.

No coloured spots are visible by ordinary observation on the left. On the right, the same separated positions are made visible. A specific locating-agent name need not be memorised.

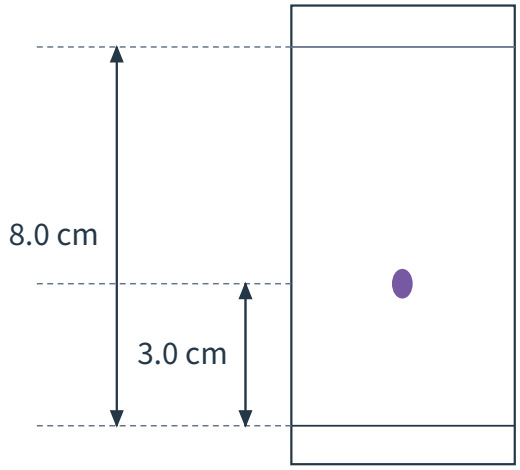
An apparently blank developed chromatogram does not immediately prove that the sample contains no substances. A colourless sample may require a locating agent before observation. Conversely, if the question already shows three clear spots, no locating agent was used does not explain why there are three rather than four. Consider overlapping components instead.

Rf expresses position as a ratio of distances. Measure from the baseline to the centre of the spot, then from the same baseline to the solvent front. The ratio of these distances is Rf:

$$R_f = \frac{\text{Distance moved by substance}}{\text{Distance moved by solvent}}$$

Use the same units for both distances. They cancel on division, so Rf has no unit such as cm or mm. Spots normally lie between the baseline and solvent front, giving Rf values between 0 and 1.

Rf divides two distances measured from the same baseline



Substance distance:
to the spot centre

Solvent distance:
to the solvent front

Rf = substance distance
÷ solvent distance

=3.0÷8.0=0.375

No unit; usually between 0 and 1

The baseline is not the paper bottom, the front is not the paper top, and spots are measured at their centres.

This diagram is drawn accurately to the 3.0 cm:8.0 cm ratio. Both lengths start at the same baseline, and numerator and denominator must use the same length unit.

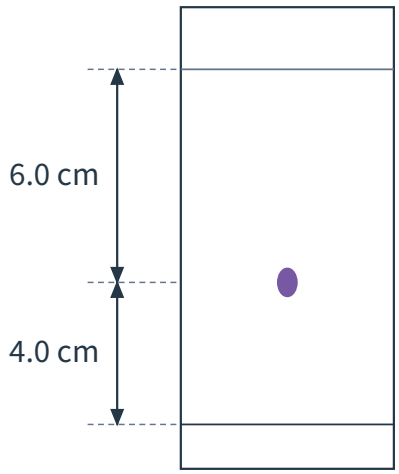
For example, if the substance moves 3.0 cm and the solvent moves 8.0 cm, their matching units allow direct division:

$$R_f = \frac{3.0}{8.0} = 0.375$$

If two decimal places are requested, write 0.38. Do not invert the ratio to 8.0÷3.0 or use the paper's total length as the denominator. In common questions with other conditions held constant, a substance travelling further moves more readily with that solvent. If the solvent is water, the question may allow a comparison of water solubility. Actual movement also depends on interactions with the paper; Rf is not a universal constant determined by solubility alone.

If dimension arrows mark consecutive segments, add them to find the full distance. Some questions give baseline-to-spot and spot-to-solvent-front distances separately. The solvent's total travel is their sum. The first segment already begins at the baseline, so do not subtract the paper-bottom-to-baseline distance again.

For segmented distances, first reconstruct the full baseline-to-front distance



Baseline → spot: 4.0 cm

Spot → front: 6.0 cm

Baseline → front = 4.0 + 6.0

Rf=4.0÷10.0=0.40

Paper margins are not travel distances

The 4.0 cm already starts at the baseline; do not subtract the bottom margin again.

The 4.0 cm and 6.0 cm lengths are consecutive segments, not heights measured from the bottom of the paper. The upper and lower margins are visual references only and are not used in the calculation.

For example, the baseline-to-spot distance is 4.0 cm and the spot-to-front distance is 6.0 cm. The substance travels 4.0 cm, while the solvent travels the whole distance, so add the segments first:

$$\text{Distance moved by solvent} = 4.0 + 6.0 = 10.0 \quad \text{cm}$$

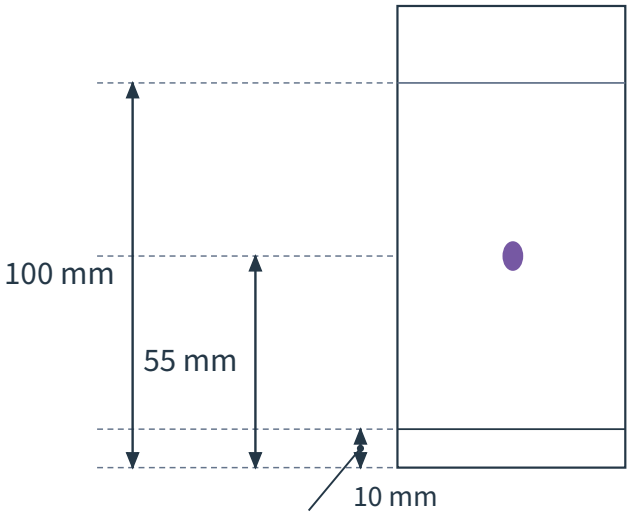
Then calculate the ratio:

$$R_f = \frac{4.0}{10.0} = 0.40$$

Blank paper above or below does not change these distances. Using 4.0÷6.0 mistakes the final segment for the entire solvent travel. Follow each arrow to its starting and finishing points before choosing the numbers to divide.

If heights are measured from the paper's bottom, subtract the baseline height from both. Height gives a position relative to the paper's bottom; travel distance begins at the baseline where the sample started. These coincide only if the baseline is at the bottom. Usually, subtract baseline height from spot height and from solvent-front height, then divide.

For heights measured from the paper bottom, subtract baseline height from both distances



Baseline height: 10 mm

Substance distance
= 55 − 10 = 45 mm

Solvent distance
= 100 − 10 = 90 mm

Rf=45÷90=0.50

Do not use 55 ÷ 100

Identify both ends of each dimension arrow before deciding whether to add or subtract.

All three heights are measured from the same bottom edge. Subtract 10 mm from both the spot and solvent-front heights to obtain their actual travel distances from the baseline.

The baseline is 10 mm high, the spot 55 mm and the solvent front 100 mm. These heights begin at the paper's bottom, but both travel distances must begin at the baseline:

$$\text{Distance moved by substance} = 55 - 10 = 45 \quad \text{mm}$$

$$\text{Distance moved by solvent} = 100 - 10 = 90 \quad \text{mm}$$

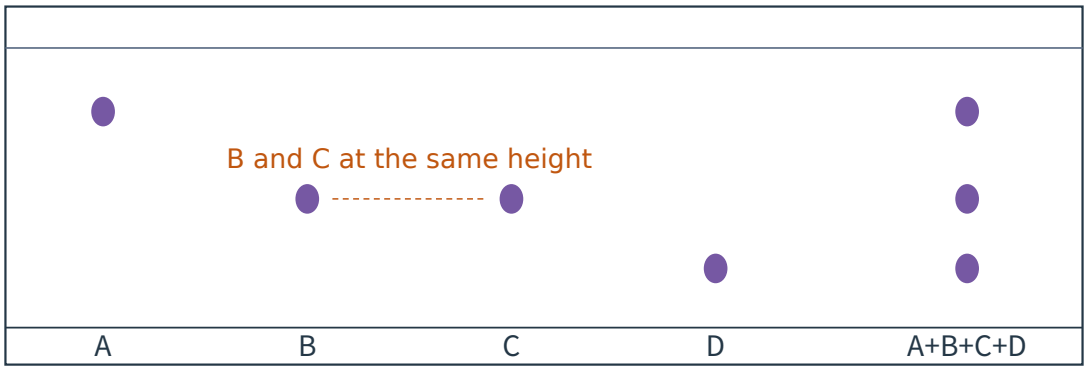
Therefore:

$$R_f = \frac{45}{90} = 0.50$$

Using 55÷100 incorrectly places the common starting point at the paper's bottom. Subtracting 10 from only the numerator leaves the two distances with different starting points and is also wrong. If the original question says not drawn to scale, use the labelled values rather than estimating lengths with a ruler.

Several components may share the same Rf and overlap at one position. Two substances travelling the same distance under the same conditions can form one overlapping spot. Four components therefore do not guarantee four separate spots. If three spots are already visible and other conditions are normal, consider that two substances have not separated.

Four substances can give only three spots when equal Rf values overlap



Under these development conditions, B and C have equal Rf values and cannot give two distinct visible positions.

All five lanes develop under the same conditions. The middle spot in the right-hand lane is an overlap of B and C. Four substances need not produce four separate spots.

If known substances B and C lie at the same height among A, B, C and D, the mixture may show only three spots: A, the shared B-and-C position, and D. A component insoluble in the solvent may remain at the starting point; it does not automatically disappear from the paper. Compare that prediction with whether the original diagram shows a baseline spot. Changing to another suitable solvent or conditions may help separate overlapping substances. A shared spot alone does not prove that they are the same compound.

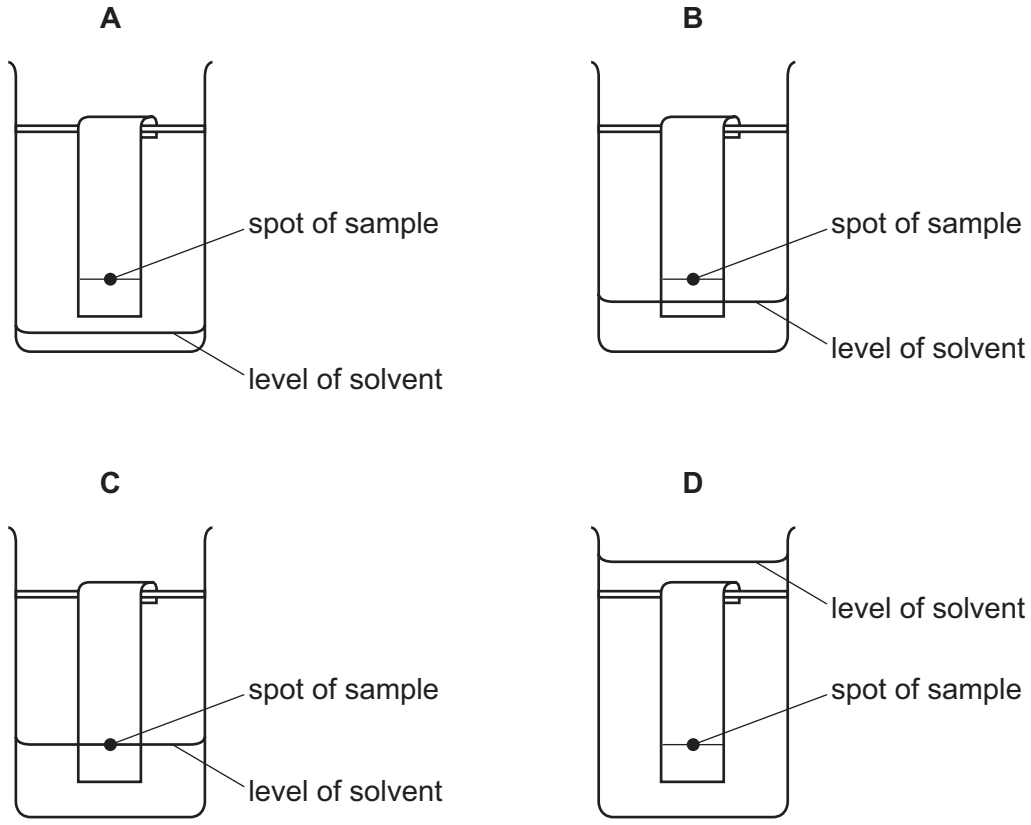
The initial solvent level must be below the _____, while the paper's bottom must _____. A _____ makes colourless spots visible. Both Rf distances begin at the _____ the numerator ends at the _____ and the denominator at the _____. Two substances with the same Rf may _____.

Answers: sample spot/baseline; touch the solvent; locating agent; same baseline; centre of the spot; solvent front; overlap

Recognising the Question Check both that the paper touches the solvent and that the sample is above the liquid level. Interpret spot numbers and positions under the same conditions. For Rf, identify both ends of each dimension arrow: add consecutive segments or subtract the baseline from absolute heights, then divide substance travel by solvent travel.

Worked Example 1 (2025·May/June 0620/12 Q38) · Level 2

Which diagram shows the correct level of solvent at the start of a chromatography experiment?



Solution

A correct starting arrangement meets two conditions: the lower end of the paper touches the solvent so that it can rise, and the sample spot stays above the solvent level rather than being immersed in the reservoir.

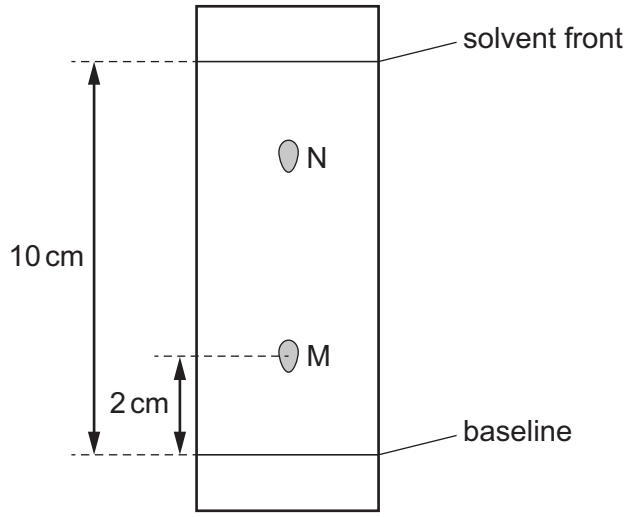
In A, the solvent level is below the paper's lower end, so solvent cannot enter the paper normally. In B, the paper's bottom is immersed while the sample remains above the liquid, satisfying both conditions. In C, the solvent reaches the sample spot; in D, it covers it. Both allow the sample to mix directly with the reservoir.

Choose **B**. A solvent level below the sample alone is not enough to choose A; the paper must actually touch the solvent too.

Worked Example 2 (2023·May/June 0620/22 Q39) · Level 2

Two compounds, M and N, are dissolved in water and separated by chromatography.

The results are shown.



What is the R_f value of M and which compound is most soluble in water?

	R_f value of M	most soluble compound
A	0.2	M
B	0.2	N
C	5.0	M
D	5.0	N

Solution

First calculate M's Rf. M travels 2 cm from the baseline to the spot centre, while the solvent travels 10 cm from the same baseline to the solvent front:

$$R_{f(\mathrm{M})} = \frac{2}{10} = 0.2$$

The matching length units cancel, so Rf has no unit.

Now compare M and N. N travels further. Under this question's shared conditions, with water as solvent, N moves more readily with water, so choose **N** as most soluble in water.

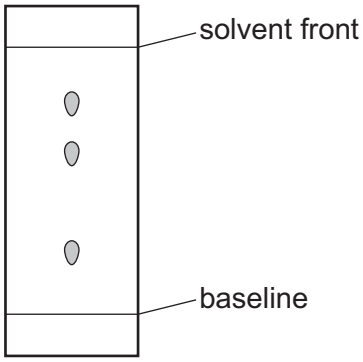
The row containing both 0.2 and N is **B**. Inverting numerator and denominator would incorrectly give 5.0. M's lower Rf also does not mean that it moves more readily with the solvent.

Practice

1. (2025·May/June 0620/21 Q40) · Level 2

A mixture of four different colourless amino acids is analysed by paper chromatography.

The final chromatogram is shown.



Why does the chromatogram only show three spots?

A A locating agent is **not** used.

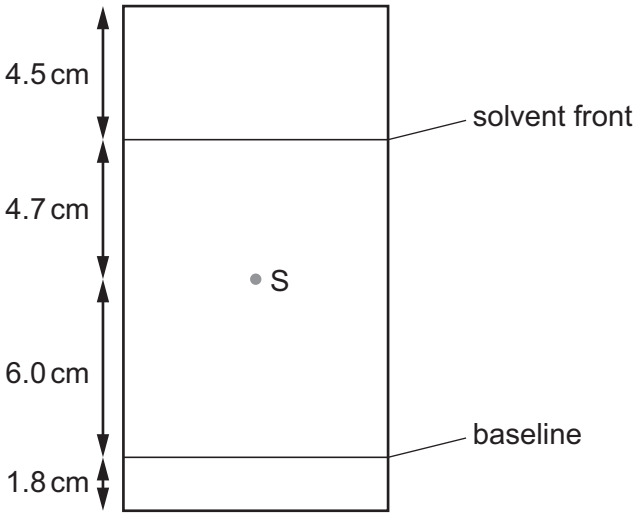
B One of the amino acids is insoluble in the solvent.

C The solvent front is too near the top of the paper.

D Two of the amino acids have the same R_f value.

2. (2023·May/June 0620/23 Q39) · Level 3

The chromatogram obtained from a chromatography experiment on substance S is shown.



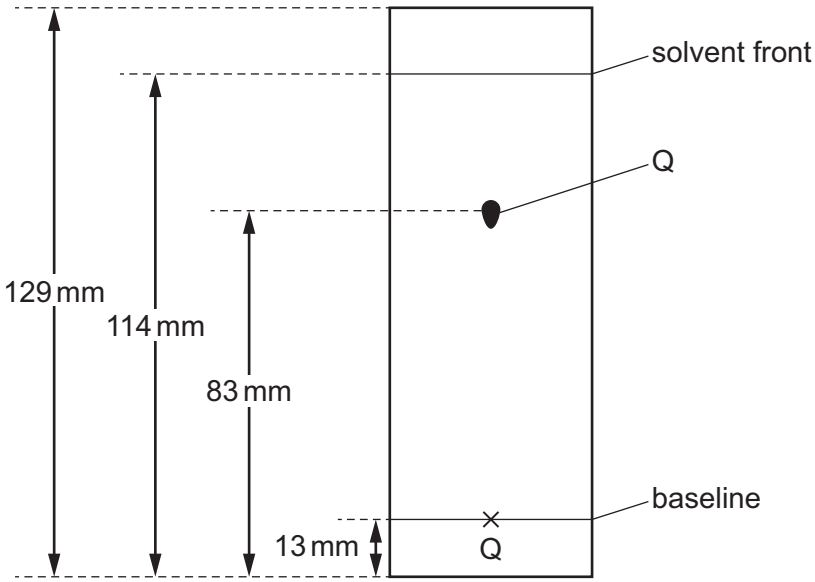
What is the R_f value of S?

A 0.39 **B** 0.46 **C** 0.56 **D** 0.62

3. (2023·October/November 0620/22 Q40) · Level 3

Substance Q is investigated using chromatography.

The chromatogram is shown. The diagram is not drawn to scale.



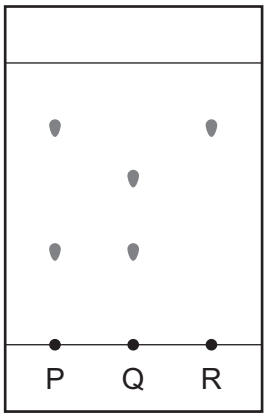
What is the R_f value of Q?

A 0.60 **B** 0.64 **C** 0.69 **D** 0.72

4. (2025·May/June 0620/13 Q39) · Level 3

Three coloured inks, P, Q and R, are tested using paper chromatography.

The chromatogram is shown.



Which statement is correct?

A P is a pure substance.

B P and Q have a colour that could be the same in both inks.

C The top line of the chromatogram is called the baseline.

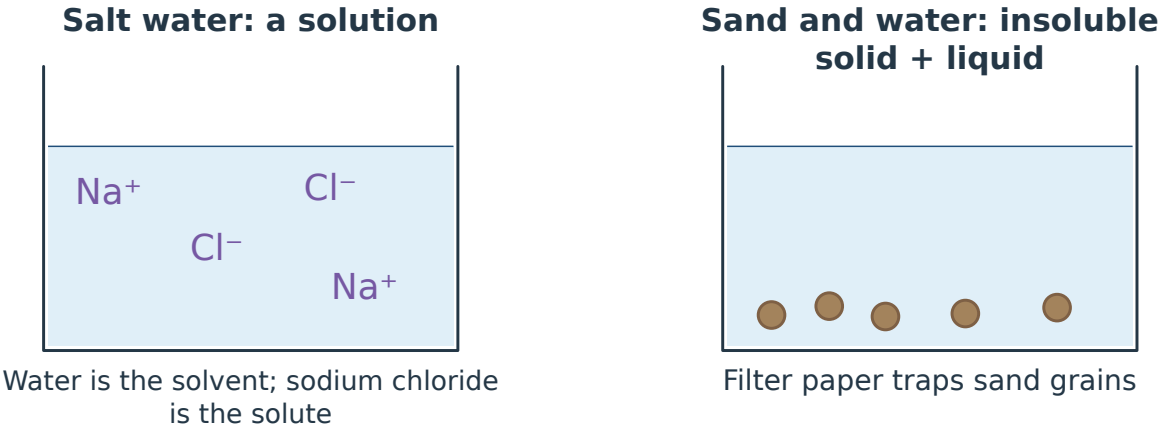
D All the colours in R are also in Q.

12.4 Separation, purification and purity

Type 1: Plan separation and purification using solubility and boiling point

Key Ideas **First decide which substances dissolve and which remain solid.** A solvent dissolves a solute. One or more solutes dissolved in a solvent form a solution. For example, when sodium chloride, NaCl, dissolves in water, H_2O, water is the solvent, sodium chloride is the solute and the whole salt-water mixture is the solution. A transparent mixture can still contain dissolved solute.

Solutes disperse through a solvent; an insoluble solid remains a separate phase



Water is the solvent; sodium chloride is the solute

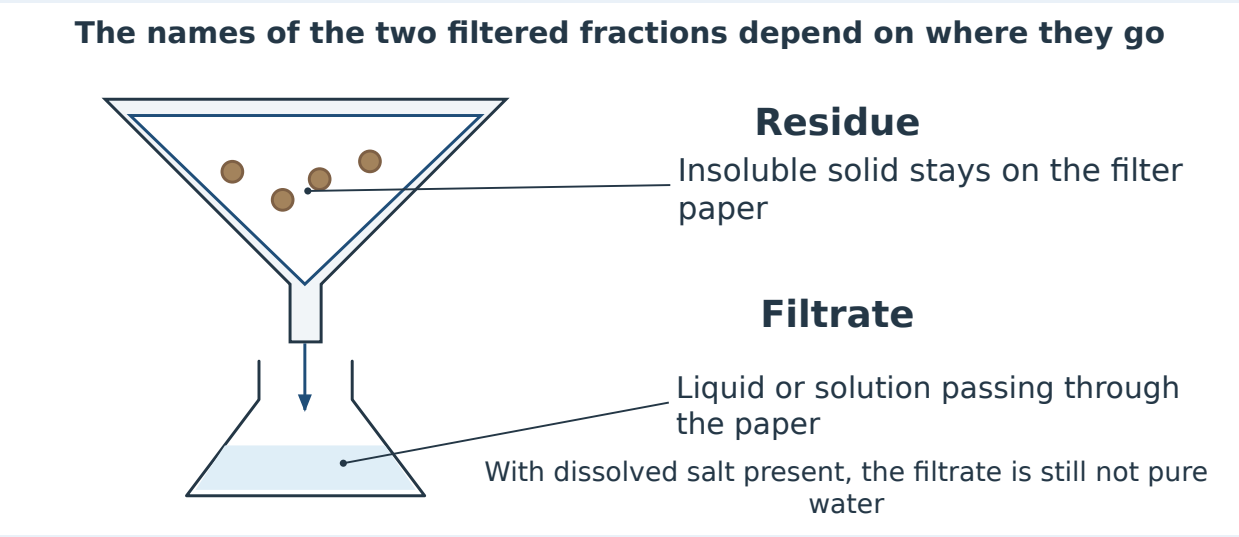
Filter paper traps sand grains

The background represents water; ions on the left illustrate dispersion, not actual size or concentration.

Water is the solvent. Dissolving sodium chloride produces a solution containing Na^+ and Cl^-. Sand does not dissolve and remains as separable solid particles.

Add salt and sand separately to sufficient water and stir. The salt dissolves to form a solution; the sand remains as insoluble particles. On filtration, sand is trapped by the paper while dissolved sodium chloride passes through with the water. Filtering salt solution first to obtain solid salt therefore does not work. Choose a method according to whether each substance is currently dissolved.

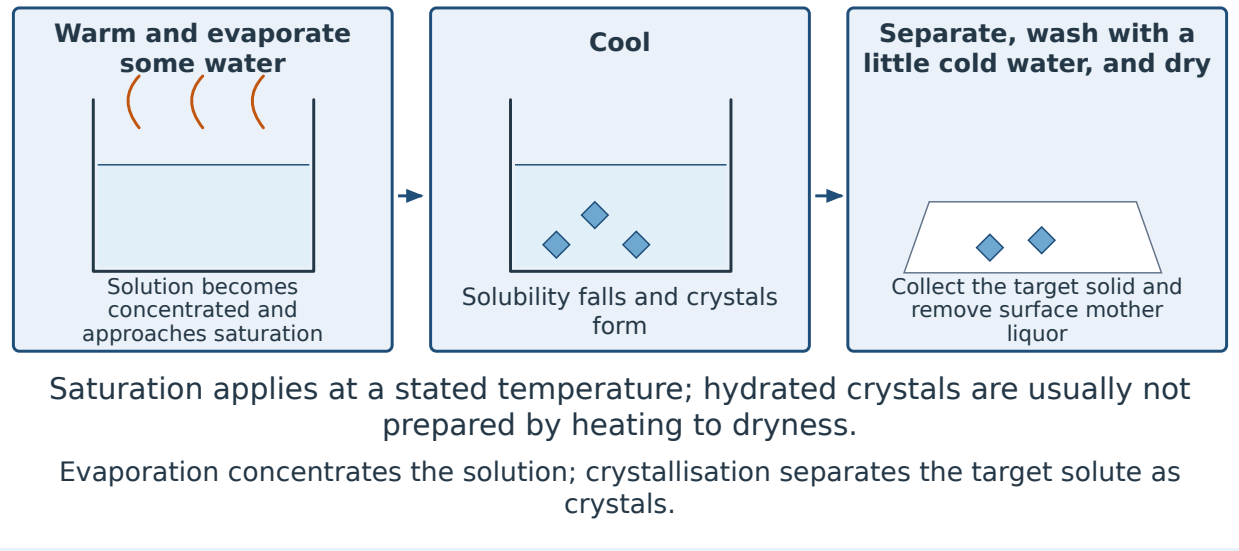
Filtration separates an insoluble solid; name the residue and filtrate by where they go. The solid left on the filter paper is the residue. The liquid or solution passing through is the filtrate. Residue can also describe material left after evaporation, distillation or another process. Its identity depends on the operation, so it does not always mean one particular salt.



Brown particles represent insoluble solid, and blue arrows show the filtrate passing through. Dissolved salt is not drawn as large particles that the paper would trap.

Suppose a reaction leaves excess insoluble oxide and a soluble salt. Filtration retains the excess oxide on the paper and collects the salt solution below. For name the residue, give the oxide's name. For name the filtrate, give the salt's aqueous solution. Solid or liquid alone does not supply the requested substance name, and a salt-containing filtrate is not pure water.

To obtain crystals from a solution, understand saturation and temperature. A saturated solution contains the maximum dissolved concentration of that solute at a specified temperature. In most crystal-preparation examples here, solubility decreases on cooling. Warm the solution and evaporate some water to concentrate it or approach saturation, then cool it to crystallise. Collect the crystals, wash with a little cold water and dry them to remove the mother liquor from their surfaces.

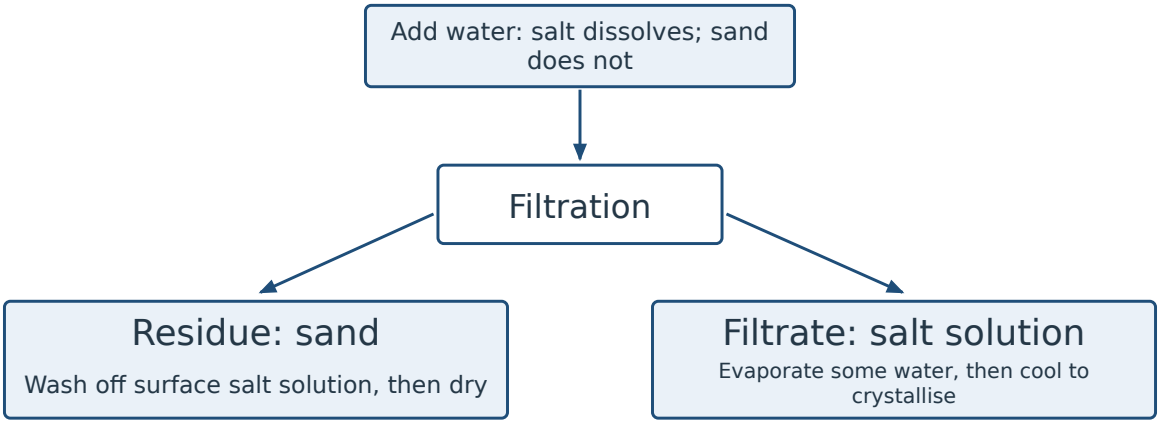


Blue diamonds represent crystals, not their actual sizes. Adapt the procedure to the particular salt. Cooling while water remains in the solution can produce crystals.

If crystals appear when a warm saturated salt solution cools, the amount previously dissolved exceeds what can remain dissolved at the lower temperature. Filtration cannot trap salt that is still dissolved. Strongly heating all the way to dryness may also damage hydrated crystals you intend to prepare. If a multiple-choice question offers evaporation but not crystallisation, recognise evaporation as the concentration step needed to obtain salt from solution and choose accordingly. The later cooling step does not remove evaporation's role.

Choosing a suitable solvent is the first step in separating two solids. It should dissolve one component but leave the other insoluble, without causing an unwanted reaction with a substance to be recovered. Dissolution followed by filtration sends the originally mixed solids to different places, from which they can be recovered separately.

Salt mixed with sand: give the two substances different separation paths



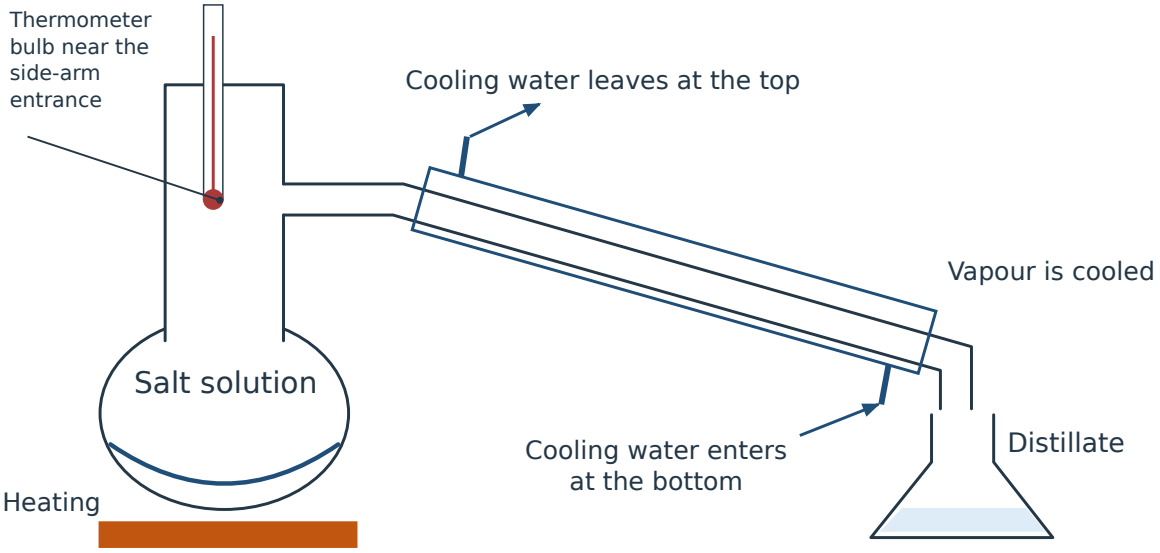
Keep the fraction containing the target; do not discard the required sample after filtration.

Water is chosen because the salt is soluble and sand is insoluble. Each branch retains its substance and completes the corresponding purification.

To recover both salt and sand, add enough water and stir to dissolve the salt. Filter, then wash and dry the sand residue. The filtrate is salt solution; evaporate some water, cool to crystallise, then separate and dry the salt. State which portion is retained at each step. Discarding the filtrate would discard the salt too.

To recover a solvent, vaporise it and then condense it. Simple distillation obtains a solvent from a solution containing a non-volatile solute, such as water from salt water. Heating turns water into vapour, which enters a condenser and cools back into liquid. The collected liquid is the distillate. Salt remains in the original flask and does not distil over in significant amounts with the water vapour.

Simple distillation: vaporise the solvent, then condense and collect it



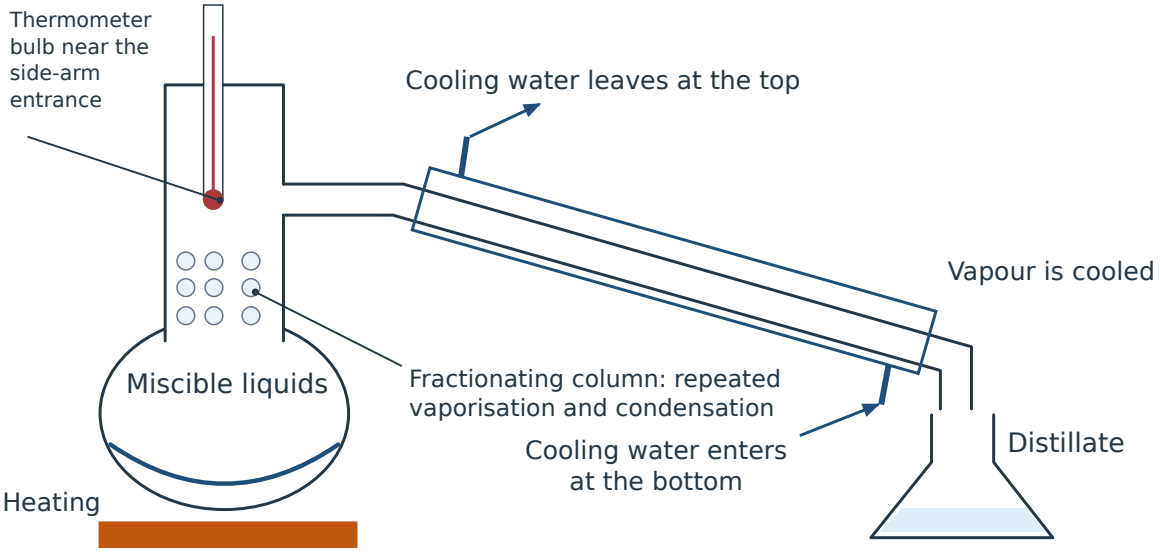
Water vaporises, condenses and is collected; common non-volatile salts remain in the original flask.

The diagram shows apparatus connections. Cooling water enters the condenser at the lower end and leaves at the upper end. The thermometer bulb is near the vapour entrance to the side arm. The receiving end is open to the surroundings; the apparatus is not heated as a sealed system.

If salt is the target, use evaporation and crystallisation. If water is the target, add condensation and collection, so choose distillation. The thermometer bulb near the side-arm entrance measures vapour temperature. Cooling water enters the condenser jacket below and leaves above, helping keep the jacket full. Cooling water flows in the outer jacket while sample vapour passes through the central tube; do not show them mixing.

For miscible liquids, use boiling-point differences and repeated separation. Fractional distillation adds a fractionating column between the flask and condenser. Repeated vaporisation and condensation enrich the rising vapour in the lower-boiling component, which appears first in greater proportion in the distillate. Collecting fractions over appropriate temperature ranges separates liquids with relatively close boiling points more effectively than one simple-distillation stage.

Fractional distillation adds a column to separate liquids of different boiling points more effectively



The lower-boiling component becomes enriched in the first distillate; collect separate temperature ranges.

The diagram shows apparatus connections. Cooling water enters the condenser at the lower end and leaves at the upper end. The thermometer bulb is near the vapour entrance to the side arm. The receiving end is open to the surroundings; the apparatus is not heated as a sealed system.

Ethanol and water are miscible. Filter paper cannot trap one liquid while letting the other through. Their different boiling points allow fractional distillation. The first distillate is richer in the lower-boiling component; do not claim without qualification that one ordinary fractionation gives 100% pure ethanol. If the question asks about a mixture of liquids with different boiling points, use that difference to identify fractional distillation.

Melting and boiling points help identify substances and assess purity. At a given external pressure, a pure substance has a specific melting point and boiling point. An impure sample may melt over a range of temperatures. In questions here involving a non-volatile solute dissolved in a liquid, the usual changes are a lower melting point and a higher boiling point. Compare with pure-substance data at the same pressure, rather than selecting any pair in which both numbers have changed.

Purity from phase-change temperatures: establish the direction, then select matching data

Pure substance: fixed melting and boiling points	Add a soluble, non-volatile impurity at the same pressure
Melting point = 0°C	Melting point below 0°C; broader melting range
Boiling point = 100°C	Boiling point above 100°C

Water with suitable non-volatile solutes illustrates the directions; the size of the change depends on the sample.

External pressure is held constant. The diagram concerns melting-point depression and boiling-point elevation caused by a soluble non-volatile impurity. It does not imply fixed changes for every liquid mixture.

At ordinary atmospheric pressure, pure water melts at 0°C and boils at 100°C. A suitable soluble non-volatile impurity can lower its melting point below 0°C and raise its boiling point above 100°C. Knowing only that an impurity is present does not determine the size of either change. Take care with negative temperatures: −116°C is lower than −114°C. Assess purity using actual measurements, pressure and the supplied reference values. A clear, colourless appearance alone is insufficient.

After filtration, the material on the paper is the _____ and the portion passing through is the _____. A saturated solution must have a specified _____. To recover salt from solution, use _____ to recover water, use _____. Miscible liquids with different boiling points are commonly separated by _____.

Answers: residue; filtrate; temperature; evaporation and crystallisation; simple distillation; fractional distillation

Recognising the Question Identify the target and whether it is soluble or insoluble, solid or a miscible liquid. Track residue and filtrate separately after filtration. Concentrate and crystallise salts, distil solvents, and use boiling-point differences for miscible liquids. In purity questions, compare each phase-change temperature with the reference.

Worked Example 1 (2025·May/June 0620/11 Q39) · Level 1

Excess solid magnesium oxide is added to dilute nitric acid.

Which separation technique is used to remove the excess solid magnesium oxide after the reaction finishes?

A chromatography

B crystallisation

C distillation

D filtration

Solution

After magnesium oxide, MgO, is added to dilute nitric acid, the question explicitly calls it excess solid. Some solid therefore remains when the reaction ends. Use filtration to remove these undissolved particles from the liquid mixture.

Solid MgO remains on the filter paper while the solution containing the product salt passes through, so choose **D**. Crystallisation obtains crystals from solution, distillation vaporises and condenses a liquid for recovery, and chromatography separates dissolved components. None is the direct operation for removing the excess insoluble solid asked about here.

Worked Example 2 (2025·May/June 0620/12 Q39) · Level 1

Which process is used to separate a mixture of liquids with different boiling points?

A dissolving

B crystallisation

C filtration

D fractional distillation

Solution

The two key conditions are mixture of liquids and different boiling points. Use the boiling-point difference to separate them, giving **fractional distillation, D**.

Heating produces vapour, and repeated vaporisation and condensation in the fractionating column enrich it, and then the distillate, in the lower-boiling component. Fractions are collected over suitable temperature ranges. Dissolving only forms a solution, crystallisation obtains crystals, and filtration separates an insoluble solid. These do not use the two liquids' boiling-point difference to separate them.

Worked Example 3 (2025·May/June 0620/22 Q40) · Level 3

Pure ethanol has a melting point of –114 °C and a boiling point of 78 °C.

What are the melting and boiling points of a sample of ethanol with glucose dissolved in it?

	melting point / °C	boiling point / °C
A	–116	77
B	–116	79
C	–112	77
D	–112	79

Solution

Use pure ethanol as the reference: melting point −114°C and boiling point 78°C. Glucose, $C_6H_{12}O_6$, dissolved as a non-volatile solute in ethanol, C_2H_5OH, fits the case of a lower melting point and higher boiling point studied here.

The melting point must be below −114°C, so choose −116°C rather than −112°C. With negative temperatures, −116 is colder and lower. The boiling point must exceed 78°C, so choose 79°C rather than 77°C.

Both columns are correct in **B: −116°C and 79°C**. This selects the option with the correct directions of change. Merely knowing that glucose was added would not independently calculate exact changes of 2°C and 1°C.

Practice

1. (2025·May/June 0620/41 Q3(c)(iii)) · Level 1

This question is about copper and compounds of copper.

(c) Copper(II) sulfate crystals are made by the reaction between copper(II) carbonate and dilute sulfuric acid, using the following steps.

The sulfuric acid has a concentration of 0.100 mol / dm^3.

step 1 Powdered copper(II) carbonate is added to dilute sulfuric acid. The mixture is stirred. A reaction occurs.

step 2 More copper(II) carbonate is added, with stirring, until the reaction stops.

step 3 Unreacted copper(II) carbonate is separated from aqueous copper(II) sulfate by filtration.

step 4 Aqueous copper(II) sulfate is heated until some of the water evaporates.

step 5 The remaining solution is allowed to cool and crystallise.

step 6 The crystals are removed and dried.

(iii) Name the residue in **step 3**.

2. (2025·May/June 0620/43 Q3(c)(ii)) · Level 1

This question is about zinc and compounds of zinc.

(c) Zinc sulfate crystals are made by the reaction between zinc carbonate and dilute sulfuric acid, using the following steps.

step 1 An excess of powdered zinc carbonate is added to dilute sulfuric acid.

step 2 Excess zinc carbonate is separated from aqueous zinc sulfate by filtration.

step 3 Aqueous zinc sulfate is heated until a saturated solution is formed.

step 4 The saturated solution is allowed to cool and crystallise.

step 5 The crystals are removed and dried.

(ii) Name the filtrate in **step 2**.

3. (2025·May/June 0620/21 Q39) · Level 1

Which method is used to obtain copper(II) sulfate crystals from an aqueous solution of copper(II) sulfate?

A chromatography

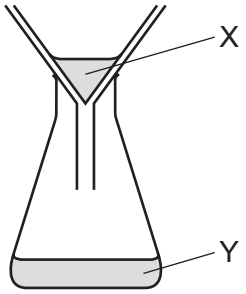
B condensation

C evaporation

D filtration

4. (2025·May/June 0620/22 Q39) · Level 2

A mixture containing an aqueous salt and an insoluble salt is filtered.



Which row describes X and Y?

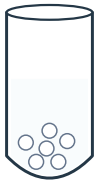
	X	Y
A	solute	pure water
B	solute	filtrate
C	residue	pure water
D	residue	filtrate

12.5 Identifying ions and gases, and flame tests

Type 1: Identify cations using precipitate colours and behaviour in excess reagent

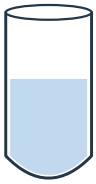
Key Ideas **Distinguish the observation from the identity inferred from it.** When two solutions are mixed and an insoluble solid appears dispersed in the liquid, a precipitate has formed, often abbreviated ppt. A precipitate may remain suspended or slowly settle; it does not need to reach the bottom before it is called a precipitate. White alone is incomplete: write white precipitate. If the liquid becomes uniformly blue without any solid, write blue solution.

State the observation first, then infer the ion from the test conditions



White precipitate

Visible insoluble solid



Blue solution

Uniform colour, no solid particles

White describes only a colour

White precipitate also states that an insoluble solid formed

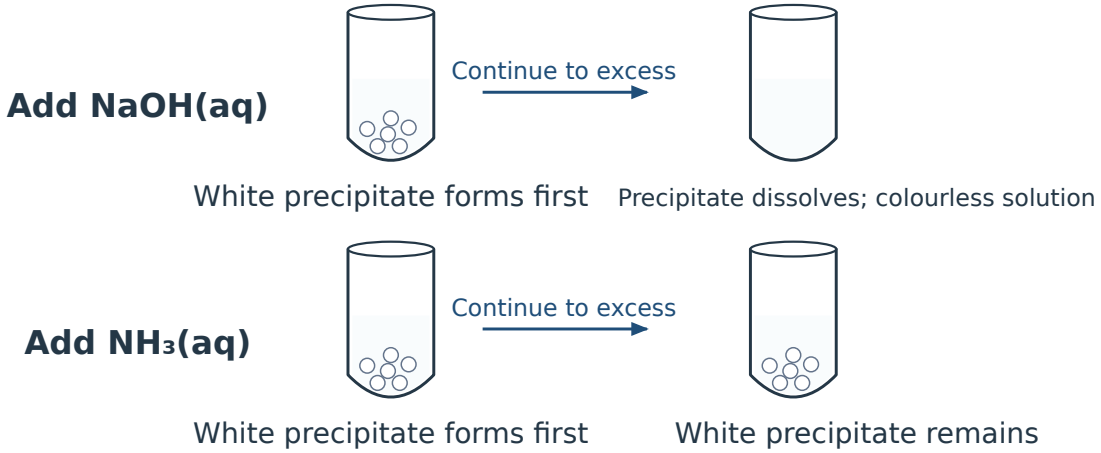
Precipitation ≠ solution colour change

Particles in the test tubes represent precipitate. The uniform liquid on the right represents a coloured solution. A white precipitate alone does not uniquely identify an ion.

If a white solid appears after reagent addition, first record a white precipitate forms. This alone cannot identify Al^{3+}, since Zn^{2+} and Ca^{2+} may also give white precipitates. Narrow the candidates by considering the reagent used and whether further addition dissolves the precipitate.

Dropwise and in excess are two stages in the same tube. Dropwise means adding one drop at a time. In excess means continuing until more reagent is present than the initial reaction requires. NaOH(aq) is aqueous sodium hydroxide, and aqueous ammonia is NH_3(aq); aq indicates an aqueous solution. Both reagents can produce some metal hydroxide precipitates, but their results in excess may differ. To compare the reagents, divide the original solution into two portions and test each separately.

Use two fresh portions of the same aluminium salt: add a little reagent, then excess



Each row uses a fresh sample; do not add ammonia to the tube already treated with NaOH.

Each row starts with a fresh portion of the original solution. Arrows show continued addition of the same reagent to excess within that portion.

For Al^{3+}, NaOH first gives a white precipitate; more NaOH dissolves it to give a colourless solution. A separate portion with ammonia first gives a white precipitate, which remains when more ammonia is added. A complete record includes all four stage results. An initial white precipitate and a final colourless solution are successive observations, not contradictory ones.

Distinguish three white-precipitate candidates by which excess reagent dissolves them. An excess reagent can convert some precipitates into dissolved metal-containing species. The metal remains in the tube; disappearance of the solid does not remove the metal. Both Al^{3+} and Zn^{2+} give colourless solutions in excess NaOH, so this alone does not distinguish them. In a fresh portion with ammonia, the white Zn^{2+} precipitate dissolves in excess, but the Al^{3+} precipitate does not. Ca^{2+} gives a white precipitate insoluble in excess NaOH; with ammonia, it usually gives no precipitate or only a very slight white precipitate.

White precipitates need further comparison using excess reagent

Ion	NaOH: a little → excess	NH_3: a little → excess
Al^{3+}: aluminium ion	White precipitate → dissolves to a colourless solution	White precipitate → insoluble
Zn^{2+}: zinc ion	White precipitate → dissolves to a colourless solution	White precipitate → dissolves to a colourless solution
Ca^{2+}: calcium ion	White precipitate → insoluble	No precipitate, or a very slight white precipitate

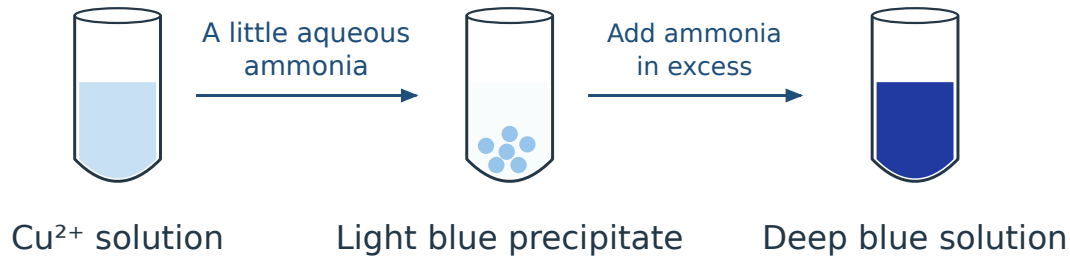
Soluble in excess describes the newly formed precipitate, not the original salt's water solubility.

For Ca^{2+} with aqueous ammonia, retain the range of no precipitate or a very slight white precipitate. Do not replace it with a large white precipitate.

If an unknown gives a white precipitate followed by a colourless solution with NaOH, Al^{3+} and Zn^{2+} remain possible. If a fresh portion also gives a white precipitate that dissolves in excess ammonia, this supports Zn^{2+}. If the white precipitate remains in excess ammonia, this supports Al^{3+}. Use the complete results from both portions, rather than stopping at the first white entry in a table.

For copper(II), distinguish a light blue precipitate from a dark blue solution. A little NaOH or ammonia gives a light blue precipitate with Cu^{2+}. In excess NaOH, it remains insoluble. In excess ammonia, it dissolves and the liquid becomes dark blue. Write the precipitate dissolves to form a dark blue solution, not dark blue precipitate. Copper remains dissolved in the tube; it has not disappeared from the sample.

Copper(II): excess aqueous ammonia gives a deep blue solution



With fresh Cu^{2+} solution and NaOH, a light blue precipitate forms and remains insoluble in excess NaOH.

The final dark blue material is a uniform solution, so no precipitate particles are drawn. It differs from a light blue precipitate remaining in the tube.

Even if ammonia dropwise until in excess is followed by only one answer space, the procedure has two stages. Write the full sequence: a light blue precipitate forms; it dissolves in excess ammonia to give a dark blue solution. This identifies both the initial solid and the final state.

A green precipitate can come from Cr^{3+} or Fe^{2+}. Both ions give green precipitates insoluble in excess ammonia. With NaOH added to a fresh portion and then in excess, the Cr^{3+} precipitate dissolves but the Fe^{2+} precipitate remains. On standing, the upper part of an Fe^{2+} precipitate near the air gradually turns brown as iron(II) is oxidised. Fe^{3+} instead gives a red-brown precipitate from the outset, unlike Fe^{2+}.

Two green-precipitate candidates: distinguish them with excess NaOH

Ion	NaOH: a little → excess	NH_3: a little → excess
Cr^{3+}: chromium(III)	Green precipitate → dissolves	Green precipitate → insoluble
Fe^{2+}: iron(II)	Green precipitate → insoluble	Green precipitate → insoluble
Fe^{3+}: iron(III)	Reddish-brown precipitate → insoluble	Reddish-brown precipitate → insoluble

On standing, the top of the green Fe^{2+} precipitate gradually turns brown where it contacts air.

Fe^{2+} and Fe^{3+} have different charges and different test observations.

Each row gives both tests for one ion. The change on standing applies to the Fe^{2+} precipitate; not every green precipitate turns brown.

If excess ammonia leaves a green precipitate, first list Fe^{2+} and Cr^{3+} as candidates. To distinguish them, add excess NaOH to a fresh portion. Dissolving precipitate supports Cr^{3+}; a remaining green precipitate supports Fe^{2+}. If asked to describe a test and explain the result, include the procedure and what each possible result means.

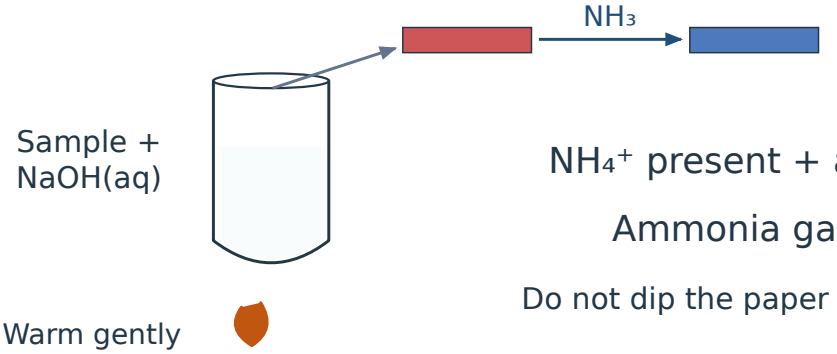
Test for ammonium, NH_4^+, using the ammonia it releases. Ammonium is not a metal ion, so do not apply the coloured-metal-hydroxide-precipitate pattern. Add NaOH and warm the sample. NH_4^+ can release ammonia, NH_3, which turns damp red litmus paper blue at the mouth of the tube. OH^- is the hydroxide ion. It reacts with ammonium to form ammonia and water:

$$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

NH_4^+ and NH_3 differ: the former is an ammonium ion in solution; the latter is an ammonia molecule that can escape as a gas.

Test ammonium by the gas released, not a metal hydroxide precipitate

Damp red litmus paper turns blue



NH_4^+ present + alkali + warming

Ammonia gas, NH_3, forms

Do not dip the paper directly into the alkali

The paper contacts gas at the mouth of the tube, not the NaOH solution itself. Otherwise, a blue colour would not demonstrate the gas's effect.

For ammonium chloride solution, adding NaOH and warming releases gas that turns damp red litmus blue at the tube mouth, supporting NH_4^+. Do not dip the paper directly into NaOH: the alkali itself also turns it blue, so that would not identify the released gas. For observations, state that the paper turns blue. For identify the gas, write ammonia or NH_3.

Some multiple-choice tables combine precipitate and flame tests in one row. Check every column against its actual reagent before accepting the whole row. A soluble sodium or copper salt does not necessarily form

a precipitate with every reagent: substances such as NaOH, Na_2SO_4 and $CuSO_4$ remain dissolved here. Mixing alone does not mean white precipitate. $BaSO_4$ is insoluble and $CaSO_4$ sparingly soluble, so white precipitates may form. Flame colours identify metal ions: Na^+ yellow, Cu^{2+} blue-green, Ba^{2+} light green and Ca^{2+} orange-red.

Check each of the three summary-table columns against its own test

Nitrate solution	Add NaOH	Add dilute H_2SO_4	Flame test
Na^+	No visible change	No visible change	Yellow
Cu^{2+}	Blue precipitate	No visible change	Blue-green
Ba^{2+}	No visible change	White $BaSO_4$ precipitate	Light green
Ca^{2+}	White precipitate	White $CaSO_4$ precipitate	Orange-red

Use a fresh sample each time; a colour alone cannot replace reagent conditions. Sparingly soluble $CaSO_4$ can precipitate.

The three columns describe separate tests on fresh portions, not three reagents added sequentially to one tube. The starting nitrates are all soluble.

If a barium-salt row claims no change with dilute sulfuric acid and a blue-green flame, reject no change because Ba^{2+} and SO_4^{2-} form white $BaSO_4$. Reject blue-green because barium gives light green. A plausible entry in another column cannot make the whole row correct. The later flame-test question type explains the procedure fully.

An unknown sample may contain several ions at once. Several observations need not all be explained by one ion. Interpret each test independently, then combine the evidence into an ion list. Two useful links are that a yellow flame supports Na^+, while dilute nitric acid followed by $AgNO_3$ gives a white precipitate for Cl^- or a yellow one for I^-. A white precipitate dissolved by both excess reagents supports Zn^{2+}.

One sample can contain several ions: use each piece of evidence for its own conclusion

Flame test	Acidify, then add $AgNO_3$	Both reagents added in excess
Yellow supports Na^+	White precipitate supports Cl^-	White precipitate dissolves in excess NaOH and NH_3: supports Zn^{2+}

These results support Na^+, Zn^{2+} and Cl^- together; do not stop after identifying one ion.

The three observations each constrain the sample's composition and should be combined. The next question type explains the white, cream and yellow halide precipitates formed by $AgNO_3$ after acidification.

In the diagram, the yellow flame gives Na^+, the white precipitate with acidified silver nitrate gives Cl^-, and a white precipitate that dissolves in both excess reagents gives Zn^{2+}. Together, write Na^+, Zn^{2+} and Cl^-. If the second

observation changes to a yellow precipitate, change the anion identification, not the yellow sodium flame or zinc's behaviour in excess reagent.

A white precipitate soluble in excess NaOH but insoluble in excess ammonia supports _____. A white precipitate soluble in both excess reagents supports _____. Cu^{2+} forms a _____ in excess ammonia. Gas released by warming NH_4^+ with NaOH makes damp red litmus _____.

Answers: Al^{3+}; Zn^{2+}; dark blue solution; turn blue

Recognising the Question For NaOH or aqueous ammonia tests, identify the cation and give separate observations for a small amount and for excess reagent. For an unknown, combine colour, solubility in excess and the result with the other reagent. To distinguish candidates, specify an additional procedure and its different outcomes, rather than only naming the ions.

Worked Example 1 (2025·May/June 0620/63 Q3(b)) · Level 2

A student tests two solids: solid **L** and solid **M**.

Tests on solid L

Solid **L** is hydrated aluminium chloride.

The student dissolves solid **L** in distilled water to form solution **L**. Solution **L** is divided into four approximately equal portions.

(a) To the first portion of solution **L**, the student adds aqueous sodium hydroxide dropwise and then in excess.

observations when added dropwise

observations in excess

(b) To the second portion of solution **L**, the student adds aqueous ammonia dropwise and then in excess.

observations when added dropwise

observations in excess

Solution

The question states that L is hydrated aluminium chloride, so its aqueous solution contains Al^{3+}. Two separate portions are tested with NaOH and ammonia. The second test does not mean adding another reagent to the first portion.

(a) A little NaOH supplies OH^- and gives a white precipitate with Al^{3+}. Continued addition to excess dissolves it, forming a colourless solution. Write, in order: **a white precipitate forms**; **the precipitate dissolves in excess to form a colourless solution**.

(b) A little ammonia in the fresh portion also gives a white precipitate, but this Al^{3+} precipitate is insoluble in excess ammonia. Write: **a white precipitate forms**; **the white precipitate remains / is insoluble in excess**.

As a final check, compare the end states: a colourless solution in excess NaOH and white solid remaining in excess ammonia. Together they match aluminium ions.

Worked Example 2 (2024·May/June 0620/63 Q3(b)) · Level 3

A student tests two substances: solid **K** and solution **L**.

Tests on solid K

Solid **K** is copper(II) carbonate.

Complete the expected observations.

(a) The student adds about 5 cm^3 of dilute hydrochloric acid to solid **K**.

observations

The solution made in **(a)** is solution **M**. The student divides solution **M** into two portions.

(b) To the first portion of solution **M**, the student adds aqueous ammonia dropwise until it is in excess.

observations

Solution

(a) K is copper(II) carbonate. Carbonates react with dilute acids to release CO_2, so the expected observation on adding dilute hydrochloric acid is **bubbles / fizzing / effervescence**. The reaction is:

$CuCO_3 + 2HCl \rightarrow CuCl_2 + CO_2 + H_2O$

After CO_2 is released, M is a solution containing Cu^{2+}.

(b) Add ammonia to a portion of M, testing Cu^{2+}. A small amount first produces a light blue precipitate. In excess, it dissolves to form a dark blue solution. Write the complete result: **a light blue precipitate forms; it dissolves in excess ammonia to give a dark blue solution**.

Although there is only one answer space, this part covers the sequence from dropwise to in excess. Blue alone omits the precipitate, its dissolution and the final dark blue solution.

Worked Example 3 (2025·October/November 0620/61 Q3(d)) · Level 3

Tests on solid K

Table 3.1 shows the tests and the student's observations for solid **K**.

Table 3.1

tests	observations
test 1 Heat about half of solid **K** in a boiling tube until there is no further change. Hold anhydrous cobalt(II) chloride paper at the mouth of the boiling tube.	condensation forms at the top of the boiling tube the anhydrous cobalt(II) chloride paper changes colour from blue to pink
test 2 The remaining solid **K** is dissolved in distilled water to form solution **K**. Solution **K** is divided into three portions. To the first portion of solution **K**, add aqueous ammonia dropwise and then in excess.	a green precipitate forms the precipitate remains when excess aqueous ammonia is added
test 3 To the second portion of solution **K**, add 1 cm^3 of dilute nitric acid followed by a few drops of aqueous barium nitrate.	a white precipitate forms
test 4 To the third portion of solution **K**, add 3 cm^3 of aqueous sodium carbonate. Test any gas given off.	effervescence is seen and a green precipitate forms the gas turns limewater milky

(d) Two different cations can give the observations in **test 2**.

(i) Identify the **two** possible cations that the observations in **test 2** show could be in solid **K**.

(ii) Describe an additional test that can be carried out on solution **K** to confirm which of the two cations you have identified in **(d)(i)** is in solid **K**.

Explain how the result of this test shows which of these two cations is in solid **K**.

test

explanation

Solution

(d)(i) Use the specified test 2: ammonia gives a green precipitate that remains in excess. Both Fe^{2+} and Cr^{3+} fit, so the two possible cations are **iron(II), Fe^{2+}**, and **chromium(III), Cr^{3+}**.

(ii) Take a fresh portion of solution K, add aqueous NaOH dropwise, then continue to excess. Both candidates first give green precipitates, but their excess-NaOH results differ. Dissolution to a green solution supports Cr^{3+}; a remaining green precipitate supports Fe^{2+}.

A complete answer is: **Add aqueous sodium hydroxide to a fresh portion of solution K, then add it in excess. A green precipitate that dissolves indicates Cr^{3+}; a green precipitate that remains indicates Fe^{2+}.**

This test is useful because the candidates give different outcomes after sharing the same ammonia result. Repeating the ammonia test alone would still not distinguish them.

Practice

1. (2025·May/June 0620/62 Q3(a)) · Level 1

A student tests two solids: solid **J** and solid **K**.

Tests on solid J

Solid **J** is iron(II) iodide.

The student dissolves solid **J** in water to form solution **J**. Solution **J** is divided into four portions.

(a) To the first portion of solution **J**, the student adds aqueous sodium hydroxide dropwise and then in excess.

observations when added dropwise

observations when added in excess

2. (2025·October/November 0620/63 Q3(a)) · Level 1

A student tests two substances: solution **A** and solid **B**.

Tests on solution A

Solution **A** is aqueous iron(III) nitrate.

Record the expected observations.

The student divides solution **A** into three approximately equal portions.

(a) To the first portion of solution **A**, the student adds aqueous ammonia dropwise and then in excess.

observations when added dropwise

observations in excess

3. (2025·February/March 0620/62 Q3(a)) · Level 1

A student tests two substances: solution **A** and solid **B**.

Tests on solution A

Solution **A** is aqueous chromium(III) chloride.

The student divides solution **A** into three approximately equal portions.

Complete the expected observations.

(a) To the first portion of solution **A**, the student adds aqueous sodium hydroxide dropwise until it is in excess.

observations

4. (2025·May/June 0620/11 Q40) · Level 3

Four different aqueous metal nitrates are tested separately with aqueous sodium hydroxide, with dilute sulfuric acid and with a flame test.

Which row shows the correct set of results for the named aqueous metal nitrate?

	aqueous metal nitrate	with aqueous sodium hydroxide	with dilute sulfuric acid	flame test
A	sodium nitrate	no visible change	white precipitate	yellow flame
B	copper(II) nitrate	blue precipitate	no visible change	blue-green flame
C	barium nitrate	white precipitate	no visible change	blue-green flame
D	calcium nitrate	no visible change	white precipitate	yellow flame

5. (2025·October/November 0620/62 Q3(d)) · Level 4

Tests on solid F

Table 3.1 shows the tests and the student's observations for solid **F**.

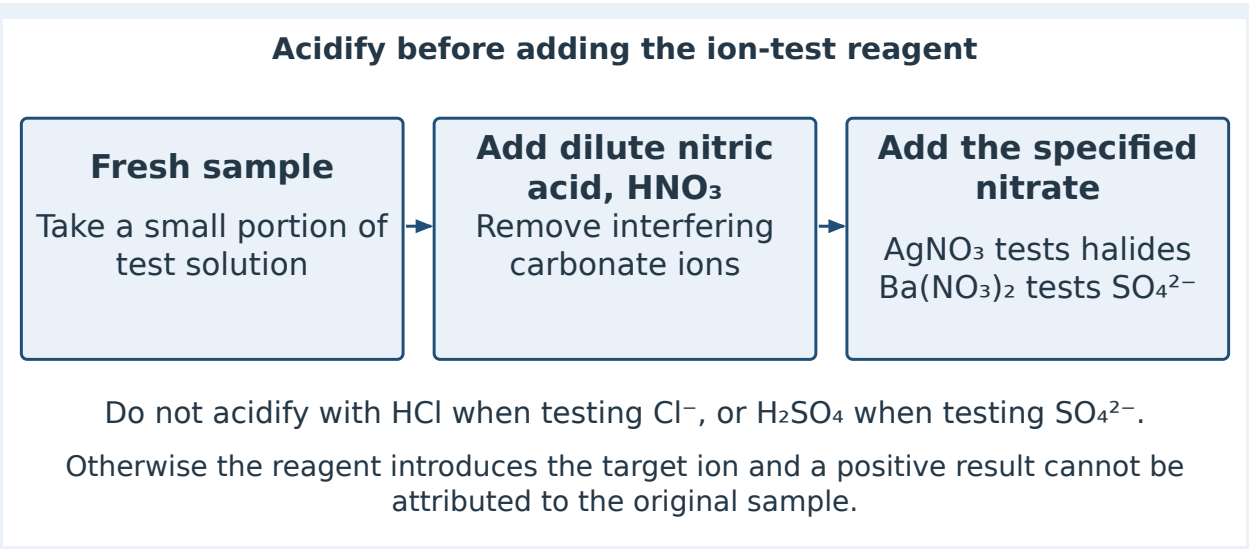
Table 3.1

tests	observations
test 1 Carry out a flame test on solid **F**.	yellow flame
test 2 Dissolve the remaining solid **F** in distilled water to form solution **F**. Divide solution **F** into three portions. To the first portion of solution **F**, add about 1 cm^3 of dilute nitric acid followed by a few drops of aqueous silver nitrate.	yellow precipitate forms
test 3 To the second portion of solution **F**, add aqueous sodium hydroxide dropwise until in excess.	white precipitate forms the precipitate dissolves in excess aqueous sodium hydroxide to form a colourless solution
test 4 To the third portion of solution **F**, add aqueous ammonia dropwise until in excess.	white precipitate forms the precipitate dissolves in excess aqueous ammonia to form a colourless solution

(d) Identify the **three** ions in solid **F**.

Type 2: Identify anions from acidification, reagents and precipitate or gas tests

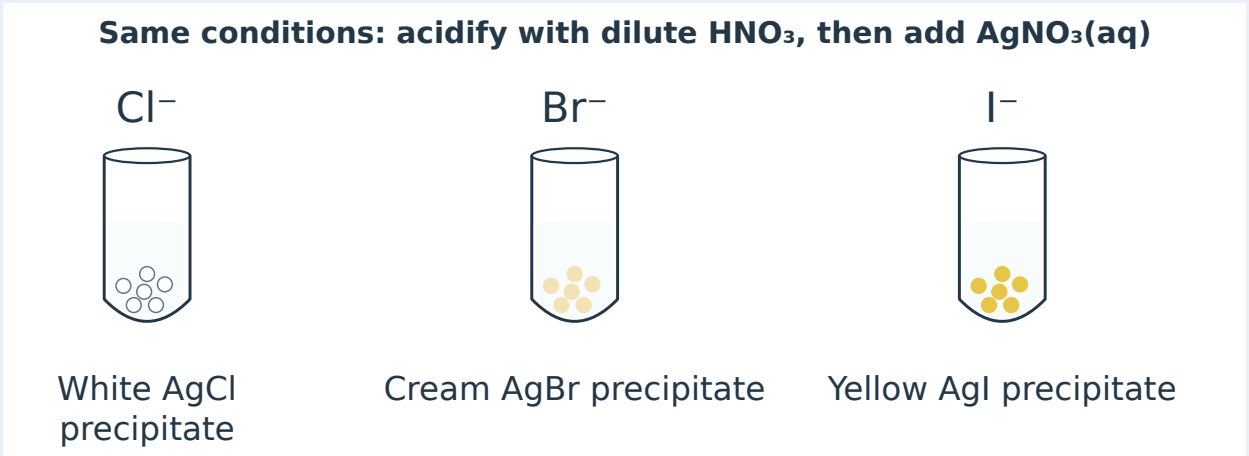
Key Ideas **Check the reagent and procedure before interpreting colour.** Anions are negatively charged ions, such as Cl^- and SO_4^{2-}. For halide or sulfate tests, first add dilute nitric acid to a fresh sample, then the specified reagent. Acid removes carbonate that could form an interfering precipitate. Nitric acid adds neither Cl^- nor SO_4^{2-}. Do not replace it with hydrochloric acid when testing Cl^- or sulfuric acid when testing SO_4^{2-}; otherwise an ion introduced by the reagent could be mistaken for one originally present.



Nitric acid introduces neither Cl^- nor SO_4^{2-}. Removing carbonate interference relies on the previously studied reaction in which carbonate and acid produce a gas.

A sample initially containing no Cl^- acquires Cl^- if hydrochloric acid is added first. Adding $AgNO_3$ can then produce white AgCl, but this does not prove the original sample contained Cl^-. Acidifying with the specified dilute HNO_3 avoids this misleading reagent-derived evidence.

The three halides use the same reagent but give different precipitate colours. Chloride is Cl^-, bromide Br^- and iodide I^-. After acidification, aqueous silver nitrate, $AgNO_3$, produces white AgCl, cream AgBr and yellow AgI respectively. These are solid silver salts, so each complete observation must include precipitate. Changing the metal cation does not make the silver salt of Br^- have the colour of the silver salt of Cl^-.

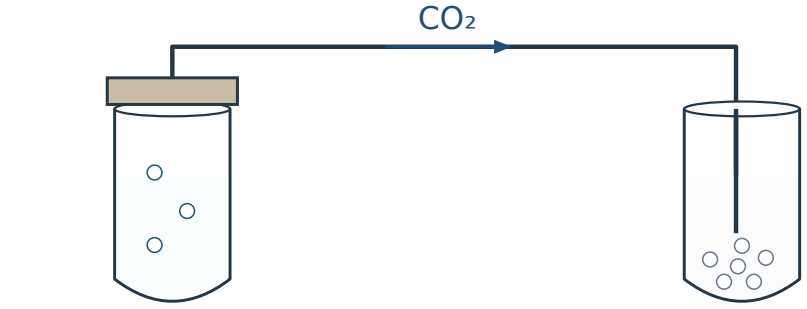


Distinguish cream from pure white and yellow. The precipitate colours are also labelled in words.

NaBr and $CrBr_3$ both contain Br^-. Acidifying separate samples with dilute HNO_3 and adding $AgNO_3$ gives a cream precipitate in both. The sodium or chromium part of the salt name is not enough for this judgement: identify the halide targeted by the test.

For carbonate, react with dilute acid and then identify the gas. Carbonate is CO_3^{2-}. Effervescence after adding dilute acid shows that a gas is formed. Carbonate and acid produce CO_2 and water. Pass the gas through limewater: cloudiness is a positive test for CO_2. Bubbles alone do not uniquely identify CO_2 in every unknown experiment.

Carbonate: bubbles show gas formation; limewater then confirms CO_2



CO_3^{2-} sample + dilute acid

Limewater turns milky

$CO_3^{2-} + 2H^+ \rightarrow CO_2 + H_2O$; bubbling is observed and milky limewater confirms the gas.

A delivery tube carries the gas into limewater, with the gas outlet open. This is a connection diagram, without specified gas volumes or concentrations.

For example, calcium carbonate fizzes with dilute hydrochloric acid:

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

Passing the released gas through limewater produces a milky cloudiness, completing both pieces of evidence for carbonate: acid releases a gas, and that gas is confirmed as CO_2.

Do not give sulfate and sulfite the same test result because their names are similar. Sulfate is SO_4^{2-}. Acidify with dilute nitric acid, then add aqueous $Ba(NO_3)_2$ to produce a white $BaSO_4$ precipitate. Sulfite is SO_3^{2-}. A small amount of acidified aqueous potassium manganate(VII), $KMnO_4$, changes from purple to colourless. This is a redox decolourisation, not formation of a white potassium manganate(VII) precipitate.

SO_4^{2-} and SO_3^{2-} differ by one oxygen atom and require different tests

Sulfate, SO_4^{2-}

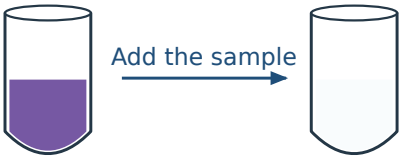
Acidify, then add $Ba(NO_3)_2$



White $BaSO_4$ precipitate

Sulfite, SO_3^{2-}

A little acidified $KMnO_4$ solution



Purple decolourises to colourless

SO_2 gas also decolourises acidified $KMnO_4$; check whether the sample being tested is a gas or a solution.

The white precipitate on the left is barium sulfate. On the right, the purple test reagent loses its colour; the colour change is not shown as a new precipitate.

$(NH_4)_2SO_4$ gives a white precipitate with barium nitrate after acidification. This is evidence for SO_4^{2-}, not NH_4^+. A separate sulfite sample turns acidified $KMnO_4$ from purple to colourless, supporting SO_3^{2-}. Identify the reagent first, then give its matching observation. Do not assign white precipitate to every sulfur-containing anion.

For nitrate, reduce it with aluminium under alkaline conditions, then test for ammonia. Nitrate is NO_3^-. Add NaOH solution and aluminium foil, then warm gently. Any NH_3 produced turns damp red litmus blue at the tube mouth. Aluminium foil is essential to this nitrate test. The NH_4^+ test already studied needs only alkali and warming, without first reducing nitrate. A salt name containing nitrate therefore does not mean that NaOH alone

must release ammonia. When using ammonia as evidence for NO_3^-, also exclude interference from NH_4^+ already present; ammonia from any source is not automatically proof of nitrate.

Nitrate test: retain all three conditions—NaOH, aluminium foil and warming

Possible sample ion	Procedure	Positive result
NH_4^+: ammonium	Add NaOH and warm	NH_3 forms; damp red litmus turns blue
NO_3^-: nitrate	Add NaOH and aluminium foil, then warm	NH_3 forms; damp red litmus turns blue

Nitrate nitrogen is not initially in NH_4^+; aluminium reduces nitrate under alkaline conditions to release ammonia.

Interpret the same positive gas test in relation to the different pretreatments. Warming an ordinary nitrate with NaOH alone does not make it behave as an ammonium salt.

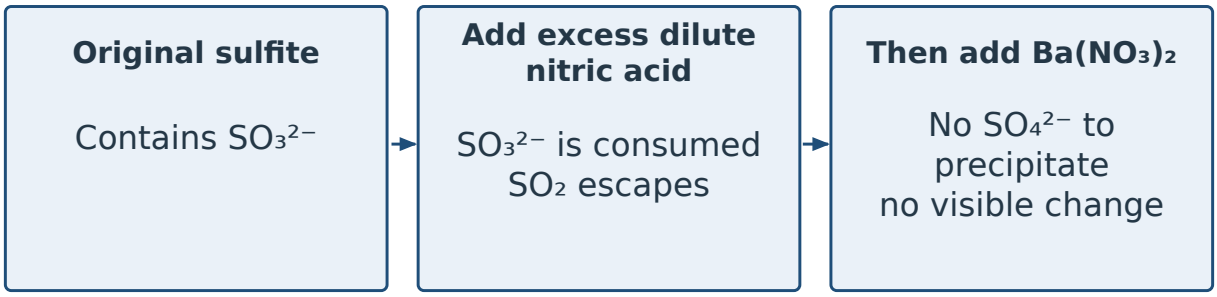
Compare $NaNO_3$ and NH_4Cl. With only NaOH and warming, NH_4Cl releases NH_3; $NaNO_3$ does not undergo this ammonium reaction. Adding aluminium foil as well and warming establishes the specified nitrate-reduction conditions. Keep all four elements in an answer: NaOH, aluminium foil, warming and damp red litmus turning blue.

In a sequence of operations, the sample's composition may change at each step. To the solution produced means that the next reagent is added to the solution remaining after the previous reaction. Sulfite with excess dilute acid can produce escaping SO_2:

$$SO_3^{2-} + 2H^+ \rightarrow SO_2 + H_2O$$

Once it has left the solution, the original SO_3^{2-} cannot still be treated as the ion tested in the next step. Nor can it be changed to SO_4^{2-} without evidence.

The next test examines the solution left after the previous reaction



$$SO_3^{2-} + 2H^+ \rightarrow SO_2 + H_2O$$

Do not count an ion that has already reacted away as still present in the next solution.

This sequence follows the specified initial addition of excess dilute acid. It does not assume that nitric acid converts sulfite to sulfate under the conditions in this question.

Suppose a sulfite has reacted with excess dilute acid and SO_2 has escaped. $Ba(NO_3)_2$ is then added to the remaining liquid. If the question supplies no conditions that form SO_4^{2-}, there is no basis for predicting white $BaSO_4$. Following the given reaction, record no visible change. SO_2 turns acidified $KMnO_4$ from purple to colourless. If the reagent is on white filter paper, the purple disappears and the white paper is revealed.

After acidification and addition of $AgNO_3$, the precipitate colours for Cl^-, Br^- and I^- are _____, _____ and _____ respectively. Test SO_4^{2-} with _____ after acidification. SO_3^{2-} changes acidified $KMnO_4$ from _____ to _____. The NO_3^- test requires _____ as well as NaOH and warming.

Answers: white; cream; yellow; aqueous barium nitrate, $Ba(NO_3)_2$; purple; colourless; aluminium foil

Recognising the Question Check the reagent combination and whether acidification occurs, then identify the anion targeted. Give a complete precipitate colour or positive gas-test result as the observation. For solution produced, track which ions the previous step has consumed before carrying forward the sample's composition.

Worked Example 1 (2025·May/June 0620/62 Q3(c)) · Level 1

A student tests two solids: solid **J** and solid **K**.

Tests on solid J

Solid **J** is iron(II) iodide.

The student dissolves solid **J** in water to form solution **J**. Solution **J** is divided into four portions.

(c) To the third portion of solution **J**, the student adds 1 cm^3 of dilute nitric acid followed by a few drops of aqueous silver nitrate.

observations

Solution

J is iron(II) iodide, so its solution contains I^-. The specified sequence, dilute nitric acid followed by silver nitrate, is the halide test.

I^- and Ag^+ form insoluble yellow AgI, so write **yellow precipitate forms**. Iodide determines this precipitate's colour. The green precipitate from iron(II) with NaOH belongs to a different reagent test and does not apply here.

Worked Example 2 (2025·October/November 0620/63 Q3(b)) · Level 3

A student tests two substances: solution **A** and solid **B**.

Tests on solution A

Solution **A** is aqueous iron(III) nitrate.

Record the expected observations.

The student divides solution **A** into three approximately equal portions.

(b) (i) To the second portion of solution **A**, the student adds excess aqueous sodium hydroxide.

observations

(ii) To the product from **(b)(i)**, the student adds a small piece of aluminium foil. The student gently warms the mixture formed and tests any gas produced.

observations

(iii) Identify the gas produced in **(b)(ii)**.

Solution

(b)(i) A is iron(III) nitrate solution, containing Fe^{3+} and NO_3^-. Excess NaOH gives a **red-brown precipitate** from Fe^{3+}, insoluble in excess NaOH. This is the visible metal-ion test at this stage.

(ii) Adding aluminium foil to the alkaline mixture and warming gently establishes the nitrate-reduction test conditions. Test the gas with damp red litmus paper at the tube mouth. The observation is **the damp red litmus paper turns blue**.

(iii) The gas that turns damp red litmus blue is **ammonia, NH_3**.

The initial red-brown solid and later ammonia release are compatible: the first comes from Fe^{3+}, while the second comes from NO_3^- reacting after aluminium is added. An earlier precipitate does not remove the need to follow the later aluminium-foil and warming steps.

Practice

1. (2024·May/June 0620/61 Q3(b)) · Level 1

A student tests two substances: solution **E** and solid **F**.

Tests on solution E

Solution **E** is aqueous chromium(III) bromide.

Solution **E** is divided into two portions.

Record the expected observations.

(b) To the second portion of solution **E**, the student adds about 1 cm^3 of dilute nitric acid and a few drops of aqueous silver nitrate.

observations

2. (2025·October/November 0620/62 Q3(b)) · Level 1

A student tests two solids: solid **E** and solid **F**.

Tests on solid E

Solid **E** is ammonium sulfate.

Record the expected observations.

The student dissolves solid **E** in distilled water to form solution **E**. The student divides solution **E** into three approximately equal portions.

(b) To the second portion of solution **E**, the student adds about 1 cm^3 of dilute nitric acid followed by a few drops of aqueous barium nitrate.

observations

3. (2025·May/June 0620/12 Q40) · Level 3

Four different colourless solutions are each tested separately with aqueous sodium hydroxide and with acidified aqueous silver nitrate.

Which row shows the results for sodium chloride?

	aqueous sodium hydroxide	acidified aqueous silver nitrate
A	no visible reaction	white precipitate
B	no visible reaction	no visible reaction
C	white precipitate	no visible reaction
D	white precipitate	white precipitate

4. (2024·October/November 0620/61 Q3(c)) · Level 3

A student tests two solids: solid **T** and solid **U**.

Tests on solid T

Solid **T** is lithium sulfite.

Complete the expected observations.

The student dissolves the remaining solid **T** in water to form solution **T**.

The student divides solution **T** into two portions.

(c) (i) To the second portion of solution **T**, the student adds excess dilute nitric acid.
The gas produced turns filter paper soaked in acidified aqueous potassium manganate(VII) white.

Identify the gas produced.

(ii) To the solution produced in **(c)(i)**, the student adds aqueous barium nitrate.

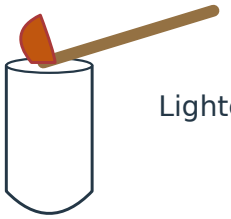
observations

Type 3: Choose gas tests and distinguish procedures from positive observations

Key Ideas **A gas test must pair what you do with what you observe.** Many gases are colourless, so the appearance of a test tube alone cannot identify them. Test hydrogen, H_2, with a lighted splint, which has a flame. A positive result is a pop. Test oxygen, O_2, with a glowing splint, which has a red-hot tip but no flame. Oxygen makes it relight. The splints' starting states and the positive results differ.

H_2 and O_2 both use splints, but their initial states differ

Hydrogen, H_2: lighted splint



Lighted: a flame is present

A squeaky pop is heard

Oxygen, O_2: glowing splint



Glowing: red-hot, without a flame

The splint relights

The left splint is already burning. The right has no flame but still glows red. State both the starting condition and the positive observation.

Put a splint at the tube mouth and a reaction occurs does not say whether the splint began with a flame or what happened. For hydrogen, write place a lighted splint at the mouth of the test-tube; a squeaky pop is heard. For oxygen, write insert a glowing splint; it relights. Do not exchange the results.

Damp litmus distinguishes ammonia turning it blue from chlorine bleaching it. Ammonia, NH_3, turns damp red litmus paper blue. Chlorine, Cl_2, bleaches damp litmus, removing its colour and leaving it white. Red to blue and coloured to white are different observations. The paper must be damp to provide water for the gas's relevant action. Expose it to the gas rather than dipping it directly into an alkaline liquid in the tube.

Damp litmus can turn blue or be bleached: these are different results

Ammonia, NH_3

Expose to the gas



Damp red litmus turns blue

Paper must be damp; do not dip it into the sample liquid

Chlorine, Cl_2

Expose to the gas



Damp litmus is bleached

Paper must be damp; do not dip it into the sample liquid

A white final state means the dye has been bleached. Do not describe bleaching as blue litmus turning red, or describe ammonia as bleaching litmus.

If damp red litmus turns white in a gas, the positive change is bleaching, so choose Cl_2. Do not select NH_3 merely because red litmus is mentioned. If the paper turns blue, use the ammonia test. Although CO_2 can show acidic behaviour, bleaching damp litmus is not its positive test in this section.

Use limewater for carbon dioxide and acidified potassium manganate(VII) for sulfur dioxide. CO_2 turns limewater milky because insoluble calcium carbonate forms. SO_2 turns acidified $KMnO_4$ from purple to colourless. The latter describes the reagent losing colour, not purple SO_2 gas or a white precipitate.

CO_2 and SO_2: choose the correct reagent and describe its change

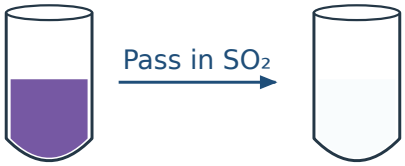
Carbon dioxide, CO_2



Limewater turns milky

Sulfur dioxide, SO_2

Acidified $KMnO_4$ solution



Purple → colourless

State the reagent and positive observation; solution changes or reacts alone is insufficient.

The CO_2 test forms white cloudiness. The SO_2 test removes the purple colour of acidified potassium manganate(VII). The final physical states are different.

A gas that removes the purple colour of acidified $KMnO_4$ gives the positive test for SO_2. If the question also states that a sulfite reacts with an acid to produce it, the reaction source and test result support each other. A gas producing white cloudiness in limewater is identified as CO_2 using the specified test. Name the reagent: solution turns white is not an adequate answer for both.

When tracing a gas back to its source, also use the original reaction conditions. Damp red litmus turning blue first identifies NH_3. Then check whether the starting materials supply NH_4^+ and alkaline conditions. Warming an ammonium salt with an alkali releases ammonia. A suitable basic oxide can also release ammonia from an ammonium salt under the stated heating conditions. For example, CaO reacts with NH_4Cl:

$$CaO + 2NH_4Cl \rightarrow CaCl_2 + 2NH_3 + H_2O$$

Ammonium loses H to become NH_3, while O from the oxide combines with those H atoms to form water. An acidic oxide cannot replace the required basic reactant. Nitrates and ammonium salts are different. If only NaOH and warming are specified, with no aluminium foil, do not automatically apply the nitrate-reduction test.

Work back from a positive gas test to the conditions in the heated mixture

At the tube mouth: damp red litmus turns blue

Gas is NH_3

Sample: ammonium salt + basic substance

Conditions allow ammonia release

$$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

Nitrate is not ammonium; without the stated aluminium reduction conditions, do not treat NO_3^- as NH_4^+.

The left column gives the observed gas evidence; the right gives the chemical conditions needed to explain its source. The columns match evidence and explanation, rather than showing two gases interconverting.

NH_4NO_3 contains NH_4^+, so NaOH and warming alone can release ammonia. $Al(NO_3)_3$ contains NO_3^- but no NH_4^+. If aluminium foil has not been added, nitrate in the name does not establish ammonia release. One substance may contain two testable ions; the actual reagents determine which reaction applies.

Test H_2 with a _____ splint; the positive result is a _____. Test O_2 with a _____ splint; the positive result is that it _____. NH_3 makes damp red litmus _____, while Cl_2 _____ damp litmus. CO_2 makes _____ cloudy, and SO_2 changes acidified $KMnO_4$ from _____ to _____.

Answers: lighted; pop; glowing; relights; turn blue; bleaches it white; limewater; purple; colourless

Recognising the Question Identify the gas from the actual positive splint, litmus or solution test, then check that the given reactants can produce it under the specified conditions. For Describe a test, pair the reagent or apparatus condition with the observation.

Worked Example 1 (2025·February/March 0620/32 Q3(c)) · Level 1

(c) Describe a test for hydrogen.

test

observations

Solution

Describe a positive test that actually identifies hydrogen. Hold a lighted splint at the mouth of a test tube containing a small amount of the gas. Hydrogen gives a squeaky pop.

The two lines can read: test: **hold a lighted splint at the mouth of the test-tube**; observations: **a squeaky pop is heard**.

Use lighted, meaning a splint with a flame, rather than glowing, the flameless red-hot splint used for oxygen. A splint alone or pop alone does not fully pair the procedure with its positive result.

Worked Example 2 (2025·October/November 0620/11 Q39) · Level 2

Two different white solids are heated and the gases formed are tested.

The gas from solid 1 turns damp red litmus paper blue.

The gas from solid 2 relights a glowing splint.

Which row identifies the gases formed from each solid?

	solid 1	solid 2
A	ammonia	hydrogen
B	ammonia	oxygen
C	chlorine	hydrogen
D	chlorine	oxygen

Solution

The gas released by solid 1 turns damp red litmus blue, the positive test for NH_3. The first column is therefore **ammonia**.

The gas from solid 2 relights a glowing splint, the positive test for O_2. The second column is therefore **oxygen**. The combination ammonia, oxygen is B.

A calls the second gas hydrogen, but hydrogen gives a pop with a lighted splint. C and D call the first gas chlorine, which bleaches damp litmus. Choose **B**.

Practice

1. (2025·October/November 0620/13 Q40) · Level 1

A piece of damp red litmus paper is put into a test-tube of gas. The litmus paper turns white.

Which gas is in the test-tube?

A ammonia

B carbon dioxide

C chlorine

D oxygen

2. (2025·May/June 0620/13 Q17) · Level 2

Three compounds are listed.

1 aluminium nitrate

2 ammonium chloride

3 ammonium nitrate

All three compounds are heated with aqueous sodium hydroxide.

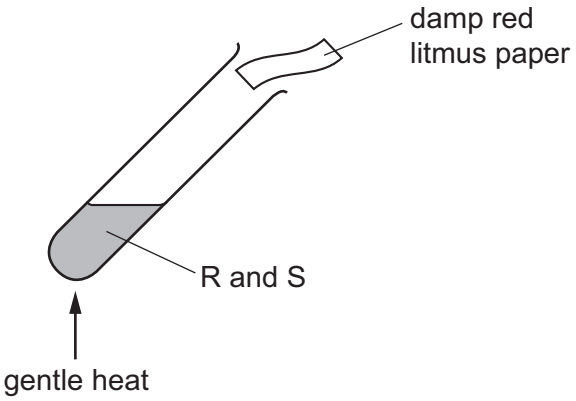
Which compounds produce ammonia?

A 1 and 3 **B** 1 only **C** 2 and 3 **D** 2 only

3. (2025·May/June 0620/11 Q18) · Level 3

A mixture of two substances, R and S, is heated gently.

The damp red litmus paper turns blue.



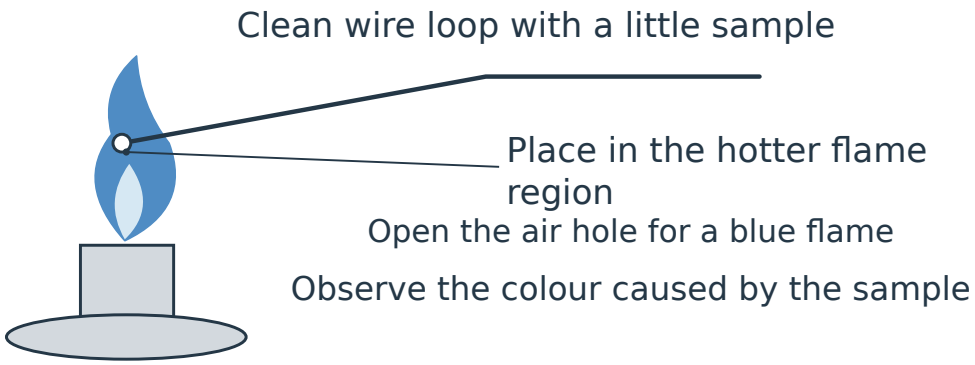
What are R and S?

	R	S
A	a basic oxide	ammonium chloride
B	a basic oxide	sodium nitrate
C	an acidic oxide	ammonium chloride
D	an acidic oxide	sodium nitrate

Type 4: Identify metal ions from flame colours and combine them with other evidence

Key Ideas **A flame test observes how a sample changes a flame's colour.** Introduce a small sample into a suitable Bunsen flame and use the characteristic colour of the metal ion to identify it. Clean the wire loop first and heat it to confirm that no colour remains from a previous sample. Collect a little sample on the loop and place it in the hotter, non-luminous blue flame. A clean splint carrying a little sample may also be used as specified. The yellow safety flame is cooler and has a strong colour of its own, making it unsuitable.

Flame test: place a little sample in a non-luminous blue flame



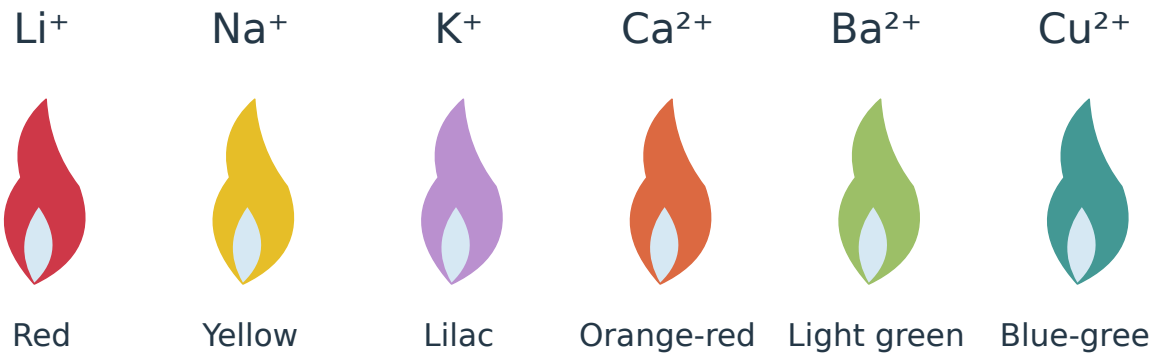
Clean and heat the loop until it gives no contaminating colour; avoid sodium from a previous sample.

The thin line represents wire, with a small sample on the loop at the left. The blue Bunsen flame supplies sufficient heat. Do not mistake the yellow safety flame for a sample's yellow flame-test result.

For Describe how to carry out the flame test, observe the flame colour alone is insufficient. Write: Use a clean wire loop to introduce a small amount of the sample into a non-luminous blue Bunsen flame, and observe the flame colour. This explains both how the sample enters the flame and which flame is needed.

Learn the six metal-ion colours separately, rather than ordering them by intensity. Li^+ gives red, Na^+ yellow, K^+ lilac, Ca^{2+} orange-red, Ba^{2+} light green and Cu^{2+} blue-green. Li^+ and Ca^{2+} both have a red component, but red and orange-red are different standard observations. Distinguish Ba^{2+} light green from Cu^{2+} blue-green as well. Impurities may interfere, and clean apparatus particularly helps reduce yellow colour from residual sodium.

Use the same blue-flame conditions and observe the sample's characteristic colour



Shade can vary with conditions; match each colour name to its ion rather than guessing from similar colour blocks.

The six samples are shown separately; no arrows suggest that one metal's flame colour changes into another's. Colour names beside each flame support reading in monochrome or with differences in colour vision.

If three samples give lilac, light green and blue-green, match them in order to K^+, Ba^{2+} and Cu^{2+}. Establish all three matches before finding the row that agrees completely. Do not guess a whole row from one familiar colour. If asked for metal names, write potassium, barium and copper. If asked for cations, the corresponding charged ions may be given.

A flame colour identifies the cation; identifying the whole salt needs more evidence. NaCl and Na_2CO_3 both contain Na^+ and can give yellow flames. To distinguish them, add dilute acid to fresh portions. Carbonate releases CO_2 bubbles, whereas sodium chloride does not show this gas-producing reaction. Yellow flame supports Na^+, and CO_2 release with acid supports CO_3^{2-}; together they support Na_2CO_3. A shared flame colour is not exclusive to one whole salt.

A yellow flame identifies Na^+; the anion needs a separate test

Salt being compared	Flame test	Add dilute acid to a fresh sample
NaCl	Yellow; contains Na^+	No bubbles
Na_2CO_3	Yellow; contains Na^+	Bubbles; CO_2 turns limewater milky

Equal flame colours do not identify the whole salt; the second result distinguishes the anions.

Both salts contain sodium ions and therefore give the same flame colour. Carbonate releases CO_2 with acid; chloride does not undergo that reaction.

Suppose the only candidates are NaCl, Na_2CO_3, $CaCl_2$ and $CaCO_3$. A yellow flame excludes the two calcium salts. Effervescence with dilute acid then excludes NaCl, leaving Na_2CO_3. If gas confirmation is also requested, pass it through limewater and record cloudiness. If the original question only asks you to choose among the four salts, explain both exclusions.

Use a non-luminous _____ Bunsen flame for flame tests. The colours for Li^+, Na^+ and K^+ are _____, _____ and _____ respectively; those for Ca^{2+}, Ba^{2+} and Cu^{2+} are _____, _____ and _____. A flame colour usually identifies only the salt's _____, rather than the whole salt by itself.

Answers: blue; red; yellow; lilac; orange-red; light green; blue-green; metal cation

Recognising the Question For a procedure, explain how a clean sample enters a blue flame. For observations, match every colour to its metal ion. If the whole salt is requested, combine the flame result with the independent anion test rather than using only one.

Worked Example 1 (2025·February/March 0620/62 Q3(d)) · Level 2

Tests on solid B

Table 3.1 shows the tests and the student's observations for solid **B**.

Table 3.1

tests	observations
test 1 Carry out a flame test on solid **B**.	red coloured flame
test 2 Dissolve the remaining solid **B** in distilled water to form solution **B**. Divide solution **B** into three portions. To the first portion of solution **B**, add a few drops of acidified aqueous potassium manganate(VII).	pale purple solution
test 3 To the second portion of solution **B**, add about 1 cm^3 of aqueous sodium hydroxide.	solution remains colourless
test 4 Warm the product of **test 3** and test any gas given off with damp red litmus paper.	damp red litmus paper remains red
test 5 To the third portion of solution **B**, add about 1 cm^3 of dilute nitric acid followed by a few drops of aqueous barium nitrate.	white precipitate

(d) Describe how to carry out the flame test used in **test 1**.

Solution

This part asks for the procedure for test 1. Repeating red coloured flame, already supplied in the table, does not answer it.

Use a clean wire loop to collect a little solid B or its solution. Adjust the Bunsen burner to a non-luminous blue flame, place the sample-bearing loop in the hotter part and observe the colour. Before changing samples, clean and heat the loop until it gives no residual colour, preventing contamination from being mistaken for B's result.

Write: **Use a clean wire loop to introduce a small amount of solid B into a non-luminous blue Bunsen flame. Observe the flame colour.**

This includes the two core requirements: a blue, non-luminous flame and a wire loop used to introduce the sample. Other tests in the table provide background; this part does not ask you to use them to identify the entire salt.

Worked Example 2 (2025·October/November 0620/61 Q3(a)) · Level 2

A student tests two solids: solid **J** and solid **K**.

Tests on solid J

Solid **J** is lithium chloride.

Record the expected observations.

(a) The student carries out a flame test on solid **J**.

(i) State the colour the Bunsen burner flame becomes during the flame test.

(ii) The student uses a roaring Bunsen burner flame to carry out the flame test.

State why a yellow Bunsen burner flame is **not** suitable for a flame test.

Solution

(a)(i) J is lithium chloride, so Li^+ determines the flame colour: **red**. It is neither sodium's yellow nor calcium's orange-red.

(ii) A yellow Bunsen flame is unsuitable because it is **not hot enough**. Use the hotter roaring blue flame to produce an observable characteristic sample colour. The yellow safety flame's own luminous colour also makes the sample colour harder to identify, but the main reason required here is insufficient heat.

Distinguish the burner's original flame from the colour after the sample is added. Use a blue flame for the test; the lithium sample then gives a red observation.

Practice

1. (2025·May/June 0620/42 Q5(b)(iv)) · Level 1

Alkali metals are reactive elements.

(b) Identify the alkali metal which:

(iv) burns with a lilac flame

2. (2025·October/November 0620/21 Q39) · Level 3

The results of flame tests on three metal salts, X, Y and Z, are shown.

metal salt	flame test colour
X	yellow
Y	lilac
Z	blue-green

Which row identifies the cations present in X, Y and Z?

	X	Y	Z
A	potassium	sodium	calcium
B	potassium	sodium	copper
C	sodium	potassium	calcium
D	sodium	potassium	copper

3. (2025·May/June 0620/13 Q40) · Level 3

The table shows the results of two separate tests on a sample of aqueous T.

test	observation
flame test	yellow flame
add dilute nitric acid	effervescence

What is the identity of T?

A sodium chloride

B calcium chloride

C sodium carbonate

D calcium carbonate

12.6 Integrated practical work: planning, graphing and evaluation

Type 1: Turn an investigation aim into a practical plan with comparable results

Key Ideas **Turn what you want to investigate into a measurable result.** An experimental plan needs more than do the experiment, record results and compare. The independent variable is the factor deliberately changed. The dependent variable is the outcome actually measured. Control variables are conditions kept the same for a fair comparison. Specify the apparatus and quantities first, then when the measurement begins and ends, and finally which result supports which conclusion.

Turn the aim into three questions: what changes, what is measured, and what stays the same?

Change: independent variable	Acid concentration For example, 0.50 and 1.00 mol/dm³
Measure: dependent variable	Time to produce 20 cm³ of gas Use a stopwatch and the same endpoint
Keep fixed: control variables	Magnesium mass and surface area, acid volume, temperature Prevents other factors changing the rate

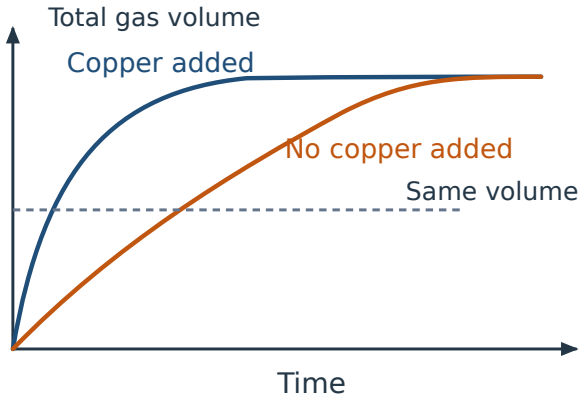
Acid must be sufficient and magnesium must produce more than 20 cm³; otherwise the target may be unreachable.

Time to produce the same gas volume compares reaction rates at different concentrations. The controls serve this specific comparison; simply listing the three variable names is insufficient.

To compare how 0.50 and 1.00 mol/dm³ acid affect the reaction rate with magnesium, use the same magnesium mass and surface area, the same acid volume and the same initial temperature, with acid in excess in both trials. Use a gas syringe and stop-watch to measure the time to collect 20 cm³ of gas. If the times are 40 s and 22 s respectively, the 1.00 mol/dm³ trial has the higher average gas-production rate. First confirm that the magnesium can produce more than 20 cm³; otherwise a trial may never reach the end-point. Do not use longer magnesium ribbon and a higher temperature in one trial and attribute every difference to concentration. Repeat each condition by the same method to check consistency.

To show that a substance is a catalyst, establish both a faster reaction and no net consumption afterwards. Compare gas volume formed in a fixed time or time taken to form a fixed gas volume. Reactant amounts, concentrations, surface areas and temperatures must be comparable, with enough reactant to reach the chosen end-point. Afterwards, recover all of the proposed catalyst, wash, dry and weigh it before comparing its remaining mass with the initial amount added.

To test whether copper is a catalyst, check both rate and recovered mass



Reach the same gas volume under the same conditions

Less time needed with copper

Recover copper after the reaction

Filter → wash → dry → weigh

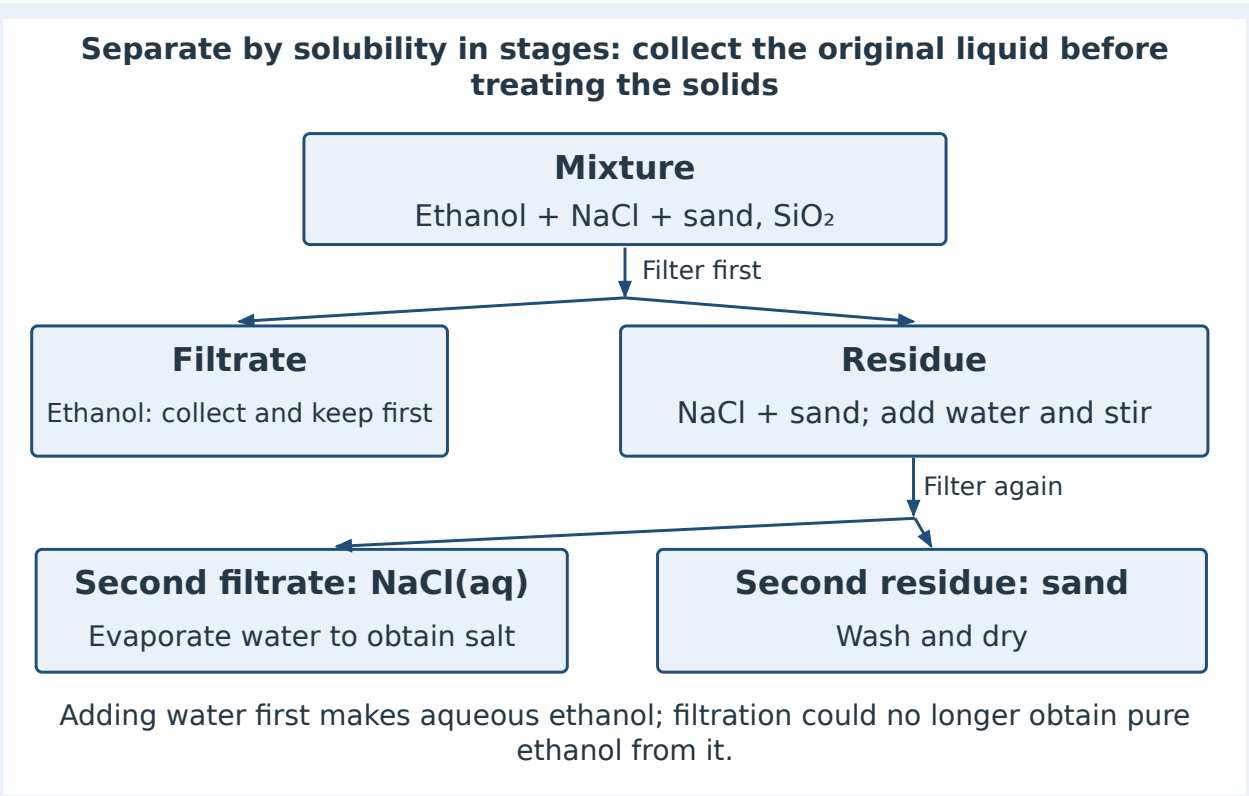
For example, 0.20 g of copper added is still 0.20 g after dry recovery, while the reaction is faster.

Acid is in excess and zinc mass is equal; the trend shows the same final hydrogen amount.

The unscaled curves show trends for the same zinc amount with acid in excess. A catalyst shortens the time to the same gas volume without increasing the final total. Obtain mass evidence only after zinc has fully reacted and copper has been washed and dried.

Suppose the same zinc amount reacts with excess sulfuric acid of the same concentration, and the trial containing 0.20 g copper reaches the specified gas volume sooner. Once zinc has reacted completely, filter off the copper, wash away surface solution and dry it. If its mass remains 0.20 g, both observations support copper acting as a catalyst. More bubbles alone are insufficient: the test substance might itself be an extra gas-producing reactant. An unchanged solid mass alone is also insufficient because a faster reaction has not yet been shown. With equal zinc amounts and excess acid, final hydrogen volumes should match; the catalyst changes how quickly the final volume is reached.

In mixture separation, identify what each filtrate and residue contains at every step. Filtration separates an insoluble solid from a liquid; it does not automatically separate two miscible liquids. Use the supplied solubility table to choose the first solvent, the timing of filtration and the portion finally collected. Explicitly retain each recovered substance. A bare sequence of dissolve, filter, heat can lose track of other required products.

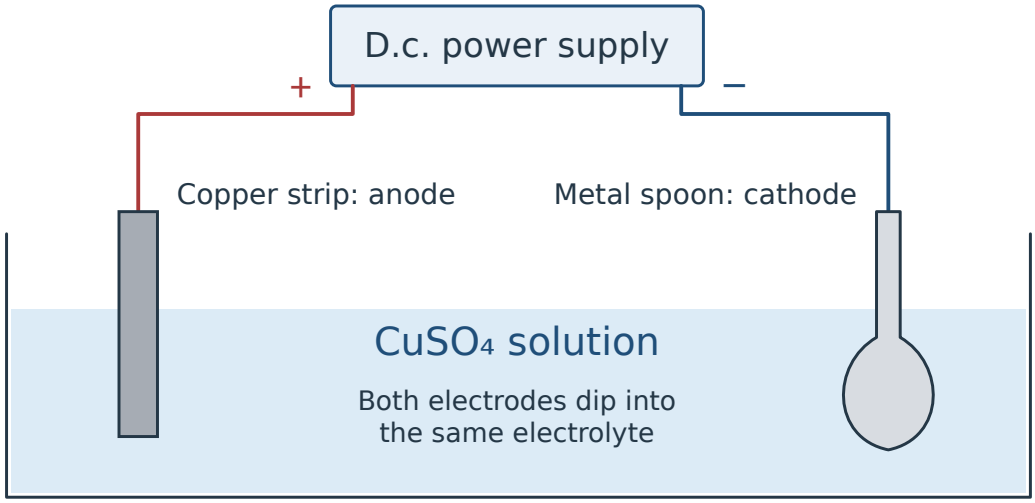


Here NaCl and sand are both taken as insoluble in ethanol, while only NaCl dissolves in the water added later. Arrows track the portions obtained at each separation step.

Consider ethanol mixed with NaCl and sand, with both solids specified as insoluble in ethanol. Filter first and keep the ethanol filtrate. Add water to the residue and stir: NaCl dissolves and sand does not. Filter again, wash and dry the sand, then evaporate water from the second filtrate to obtain salt. Adding water at the start would turn the ethanol into an ethanol–water solution, from which later filtration could not recover pure ethanol. The order follows the substance properties. Before heating, identify which portion is being heated; do not put a mixture still containing flammable ethanol directly over a naked flame without regard to this hazard.

An electroplating plan needs a complete working circuit and a method for finding coating mass. Connect the object to be plated to the negative terminal of a direct-current supply, making it the cathode. The electrolyte contains ions of the desired metal, and that metal can serve as the positive anode. Both electrodes must be immersed and connected separately to the two supply terminals. Do not omit a wire, the power supply or the electrolyte.

Copper plating: object negative, copper strip positive; Cu^{2+} gains electrons on the object



Each supply terminal connects to one electrode; ions carry charge through the electrolyte between them.

The object connects to negative; plating forms on the spoon's immersed surface.

The copper-plating apparatus includes the power supply, two complete wires and two electrodes immersed in electrolyte. The electrodes are not simply short-circuited together by a wire.

To copper-plate a metal spoon, connect a copper strip to the positive terminal and the spoon to the negative terminal, immersing both in $CuSO_4$ solution. Cu^{2+} gains electrons at the spoon and deposits as Cu, while the copper anode replenishes copper ions in solution.

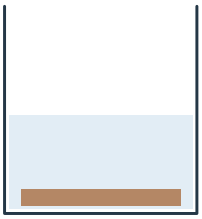
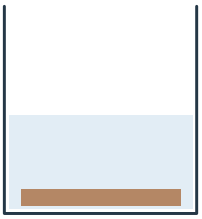
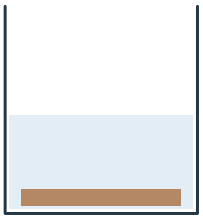
Suppose the clean, dry spoon has mass 40.12 g before plating. After plating, remove it, wash off adhering electrolyte, dry it and obtain 40.38 g. The mass increase is the deposited copper mass:

$$\text{Mass of copper coating} = 40.38 - 40.12 = 0.26 \quad \text{g}$$

Weighing the spoon wet would count residual water and salt as copper. Use clean, dry object masses both before and after plating.

To compare rusting by mass gain, use the same starting mass and exposure time. Rusting incorporates oxygen and other material from the surroundings into the solid; the question states that this increases solid mass. Keep initial dry mass, powder fineness, water volume, temperature, air exposure and duration the same when comparing alloys. Recover all solid at the end, wash and dry it appropriately by the same method, then weigh it. Compare the dry mass gained through rusting, not the mass of a container, water or a sample from which powder has been arbitrarily discarded.

Rusting comparison: equal starting dry masses and exposure conditions; compare dry mass gain

Alloy P	Alloy Q	Alloy R
		
All initially 2.00 g	All initially 2.00 g	All initially 2.00 g
2.06 g after drying	2.02 g after drying	2.00 g after drying

Keep water amount, temperature, air exposure, powder fineness and time equal; recover all solid, dry and weigh.

P gains the most mass, 0.06 g, indicating the fastest rusting under these conditions and over this time.

Use the given masses to compare this set of samples; they do not establish a universal ranking of all alloys. The containers must allow contact with air. Final masses refer to all recovered solids dried in the same way.

Suppose P, Q and R each start at 2.00 g. After the same number of days under the same conditions, all solids are recovered and dried, giving 2.06, 2.02 and 2.00 g. Their mass gains are 0.06, 0.02 and 0.00 g, so P rusted fastest over this common duration. If initial masses were 2 g, 4 g and 6 g, the largest final mass alone would not show the fastest rusting. Different residual water contents could also imitate rusting mass gains. These figures illustrate the comparison; actual alloy rankings must come from actual measurements.

Fastest heating and greatest energy per gram of fuel require different measured outcomes. To find which fuel heats water fastest, compare the time needed for a fixed temperature rise or the temperature rise in a fixed time. Both methods require accurately measured equal water volumes, initial-temperature readings, and controls for container, flame-to-container distance, exposed wick length and other heating conditions. Choose one complete route rather than mixing half of a fixed-time method with half of a fixed-temperature-rise method.

Fastest heating: compare time for a fixed temperature rise, or rise after a fixed time

Method 1: the same 20°C rise

A: 65 s
B: 84 s
C: 107 s. A is fastest. Keep water amount, initial temperature, container and flame position equal

Method 2: the same 60 s heating time

A: 18°C rise
B: 14°C
C: 11°C. A is fastest. Start timing when heating begins

The two data sets represent different methods; choose a method before comparing its readings.

For fastest heating, compare time or ΔT; the question need not become energy per gram of fuel.

The left and right show two alternative experimental designs with independent data sets. Do not combine numbers from the different designs directly.

Route one: three fuels heat equal amounts of water from 20°C to 40°C in 65, 84 and 107 s. Fuel A needs the shortest time for the same 20°C rise and is therefore fastest.

Route two: in a separate experiment, all fuels heat for 60 s and give rises of 18, 14 and 11°C. Fuel A gives the greatest rise in the same time and is fastest.

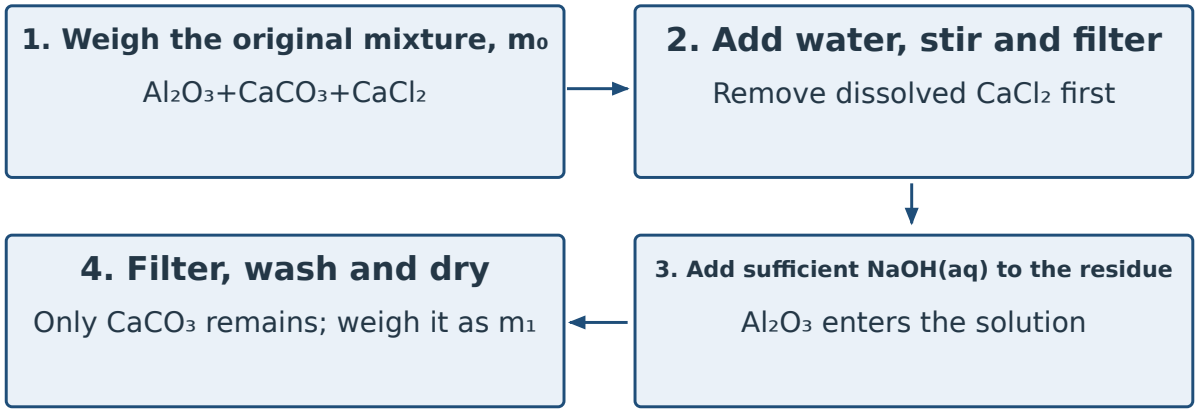
These data sets come from independent experiments. Do not treat 65 s and 18°C as belonging to one record. Measure fuel mass consumed and perform the corresponding energy calculation only if the question separately asks for energy per unit fuel mass.

To find a solid's mass percentage in a mixture, first remove anything else that would be weighed with it. Weigh the original whole mixture as m_0. Use the supplied water solubilities and behaviour with reagents to remove other components in sequence. When only the target remains, wash away soluble impurities, dry it and record its mass as m_1. Use the same mass unit for both:

$$\text{Percentage by mass} = \frac{m_1}{m_0} \times 100\%$$

If a reagent would turn a soluble component into an insoluble solid, remove that component first so it cannot contaminate the final residue.

Mass percentage of a solid in a mixture: separate it cleanly before weighing



If 1.25 g of target is separated from 5.00 g of mixture: mass percentage = 1.25 ÷ 5.00 × 100% = 25.0%.

Do not add NaOH first: $CaCl_2$ would form extra insoluble material and contaminate the $CaCO_3$ being weighed.

The supplied solubilities determine the order: water first, NaOH second. Read steps 1 to 4 from upper left → upper right → lower right → lower left.

For a 5.00 g mixture, first use water to remove $CaCl_2$ and wash the residue. Then use sufficient NaOH solution to transfer Al_2O_3 into solution. Filter, wash and dry the final $CaCO_3$. If its mass is 1.25 g:

$$\text{CaCO}_3\text{ percentage by mass} = \frac{1.25}{5.00} \times 100\% = 25.0\%$$

Do not add NaOH to the original mixture first: the supplied table states that $CaCl_2$ forms another insoluble substance. If this remained in the final residue, the $CaCO_3$ mass would be overestimated. When weighing with filter paper or a weighing dish of known mass, subtract that supporting mass.

Match a safety measure to a specific hazard and action. Risks arise from the acid, alkali, flammable liquid, gas production or heating actually used. State an action that reduces the relevant risk. Be careful does not explain what to do and cannot replace specific measures such as eye protection, avoiding ignition of flammable gas or keeping spare fuel away from flames.

Safety measures must match the actual step: state the hazard and the relevant action

Acid or alkali may splash	Wear eye protection; add small amounts slowly and swirl appropriately
Liquids such as ethanol are flammable	Keep spare fuel away from flames; extinguish and cool before refilling
Zinc and acid produce flammable hydrogen	Use small quantities away from flames; keep the gas path clear

Collect hydrogen in a gas syringe for rate comparisons; do not ignite it or block the outlet.

Each hazard is paired with its operation and control measure. These are specific situations used in this section; be careful does not replace a method.

When comparing hydrogen-production rates from zinc and acid, use small quantities, wear eye protection, keep away from flames and ensure that the route to the gas syringe is unobstructed. These measures reduce acid-splash exposure, prevent hydrogen ignition and avoid gas pressure building in a blocked apparatus. For spirit-burner water heating, keep spare alcohol away from flames and extinguish and cool the burner before refilling. Adapt the measures to the experiment; every possible hazard need not appear in every plan.

For a comparison, define the independent variable as _____, the dependent variable as _____ and controlled conditions as _____. Identifying a catalyst also requires checking after the reaction _____. Before weighing a coating or residue, _____. The denominator for mass percentage is _____. For a fixed temperature rise, the fuel taking the _____ time heats fastest.

Answers: the factor deliberately changed; the quantity actually measured; other factors that could affect the result; that the proposed catalyst has not been consumed overall; wash and dry it; the original total mixture mass; shortest

Recognising the Question Identify whether the goal is comparing rates, recovering pure substances, measuring a coating mass or finding a percentage. Specify apparatus and quantities, start and end points, controls, records and calculations, then judge from actual results. If a circuit diagram is required, include the supply, wires, electrode identities and electrolyte.

Worked Example 1 (2024·February/March 0620/62 Q4) · Level 3

When excess dilute sulfuric acid is added to solid zinc, hydrogen gas and aqueous zinc sulfate are made.

$$Zn(s) + H_2SO_4(aq) \rightarrow H_2(g) + ZnSO_4(aq)$$

Plan an experiment to show that copper is a catalyst for this reaction. Your plan should include how the results of the experiment will show that copper is a catalyst for this reaction.

You are provided with zinc powder, dilute sulfuric acid, copper powder and common laboratory apparatus.

Solution

Aim and criteria. To show that copper catalyses zinc reacting with dilute sulfuric acid, the plan must demonstrate both a faster reaction when copper is added and no net consumption of copper afterwards. The following is a complete practical method.

Compare reaction rates first. Measure a fixed volume of dilute sulfuric acid of fixed concentration, in excess relative to a known mass of zinc powder. Use a conical flask connected to a gas syringe and check that the connections are airtight. Coordinate acid addition with the start of timing. Use a stop-watch to measure the time to collect a specified hydrogen volume, such as 20 cm^3, chosen below the final volume that the zinc can produce and within the syringe range. Record the time.

Repeat with the same zinc mass and particle fineness, the same acid volume and concentration, and an accurately weighed amount of copper powder added this time. Keep initial temperature, vessel, timing start and gas collection the same. Measure the time to collect the same 20 cm^3 of hydrogen. Repeat both sets in the same way to check consistency. A shorter time with copper shows a faster reaction for the same gas volume. Comparing volumes in a fixed time is also valid, but a plan should describe its chosen route completely.

Then check whether copper is consumed. Let the copper-containing reaction continue with sulfuric acid in excess until all zinc has reacted. Filter off the remaining copper, wash away adhering solution with water, dry it and weigh it. Washing removes soluble zinc sulfate and similar residues. Do not weigh a wet residue as though its whole mass were copper.

Judge from both results. If copper shortens the time to the same gas volume and its washed, dried mass remains equal to the mass added, both findings support its role as a catalyst. The original mark scheme awards 6 marks from clear acid and zinc quantities, rate measurement, suitable apparatus, repetition with copper, recovery and weighing, and the two criteria. More bubbles, therefore a catalyst lacks the essential mass evidence. Work on a small scale away from ignition sources and keep the gas outlet clear.

Worked Example 2 (2024·May/June 0620/61 Q4) · Level 3

A **mixture** contains three compounds:

- liquid ethanol
- solid sodium chloride
- solid zinc carbonate.

Table 4.1 gives some information about these three compounds.

Table 4.1

name of compound	solubility in water	solubility in ethanol
ethanol	soluble	
sodium chloride	soluble	insoluble
zinc carbonate	insoluble	insoluble

Describe how to obtain a pure sample of each of the three compounds, ethanol, sodium chloride and zinc carbonate, from the mixture.

You are provided with common laboratory apparatus.

Solution

Use the supplied table to identify where the three substances are. NaCl and $ZnCO_3$ are both insoluble in ethanol. Initially, ethanol is the liquid portion and both solids can be trapped by the first filtration. Do not add water first: it would mix with ethanol and remove the opportunity to recover pure ethanol directly.

First recover ethanol. Filter the original mixture into a clean, dry receiver and retain the filtrate as **ethanol**. Do not add extra ethanol to wash the original mixture into the recovered sample, and do not add water first. This first recovers the liquid already present in the question.

Next recover zinc carbonate. Add distilled water to the first residue and stir until NaCl dissolves fully. $ZnCO_3$ remains insoluble in water. Filter again: the residue is **$ZnCO_3$, zinc carbonate**. Wash it with a little distilled water to remove surface salt solution, then dry it appropriately to obtain pure solid zinc carbonate.

Finally recover sodium chloride. The second filtrate contains NaCl and water. Evaporate the water to obtain solid **NaCl, sodium chloride**. If using concentration, cooling and crystal separation, explicitly finish by obtaining and drying the salt. The liquid heated here is salt water after ethanol has been separated, not the original ethanol-containing mixture.

Check all destinations: first filtrate, ethanol; second residue, washed and dried $ZnCO_3$; second filtrate after water removal, NaCl. The official mark scheme specifically requires correct recovery of ethanol first. Adding water before filtration may separate the solids but fails the goal of obtaining all three pure samples.

Worked Example 3 (2025·February/March 0620/62 Q4) · Level 3

Metal spoons can be electroplated with silver.

Describe how a metal spoon can be electroplated with silver. Include in your answer how you could determine the mass of the silver electroplated onto the metal spoon.

You are provided with solid silver nitrate, a metal spoon, a piece of solid silver, distilled water and common laboratory apparatus.

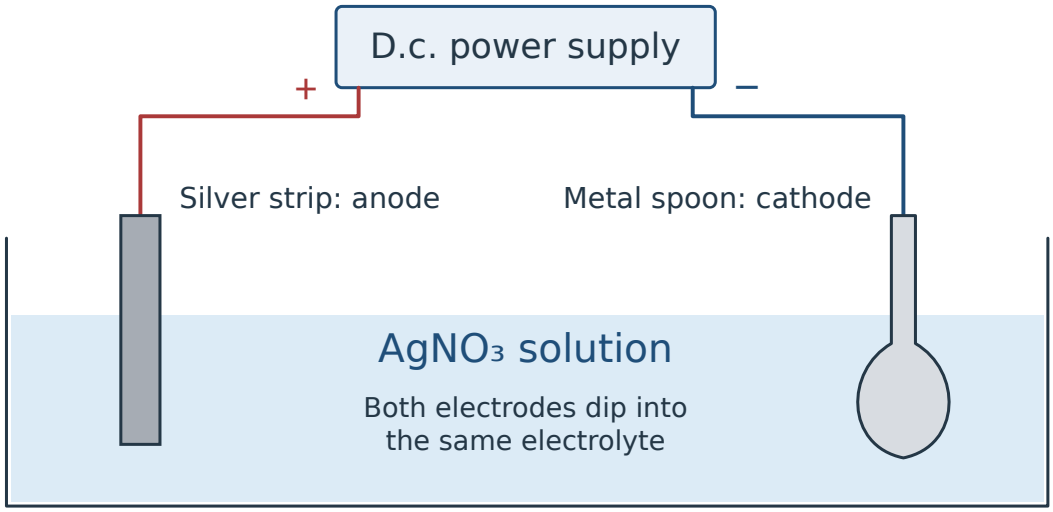
You must include a diagram in your answer.

Solution

Prepare the electrolyte and initial mass measurement. Dissolve the supplied solid **silver nitrate, $AgNO_3$**, in distilled water to make the electrolyte. Clean, wash and dry the metal spoon, then weigh it and record its initial mass, m_1.

Connect the complete direct-current circuit shown below. The silver strip is connected to the positive terminal as the anode. The metal spoon is connected to the negative terminal as the cathode. Immerse both in aqueous silver nitrate, with a separate wire from each electrode to the supply. Keep the electrodes apart, then switch on to electroplate.

m25/62/4 complete answer diagram: silver to positive, metal spoon to negative



Each supply terminal connects to one electrode; ions carry charge through the electrolyte between them.

The object connects to negative; plating forms on the spoon's immersed surface.

A complete silver-electroplating arrangement uses the supplied silver, metal spoon, silver nitrate and distilled water. Dissolve solid $AgNO_3$ first. Connect the spoon to the negative terminal and the silver strip to the positive terminal.

At the spoon, silver ions gain electrons and deposit as a coating:

$$Ag^+ + e^- \rightarrow Ag$$

The silver anode can lose electrons and replenish silver ions in solution:

$$Ag \rightarrow Ag^+ + e^-$$

Correct electrode identities ensure that silver deposits on the intended spoon.

Then determine the coating mass. After plating for a period, disconnect the supply and remove the spoon. Wash off adhering electrolyte with distilled water, dry it and measure its final mass, m_2. Subtract the initial mass, m_1, to obtain the silver coating mass:

$$\text{Mass of silver coating} = m_2 - m_1$$

For example, if the initial mass is 30.20 g and the final mass is 30.46 g:

$$30.46 - 30.20 = 0.26 \quad \text{g}$$

The original question supplies no actual measured masses. Explain how to measure both, then give the subtraction formula.

The question explicitly requests a diagram. Spoon connected to the negative terminal alone does not replace drawing the supply and complete circuit. Include washing and drying after plating; otherwise adhering solution would be incorrectly counted as silver. The official 6 marks are drawn from the two weighings, electrolyte preparation, complete circuit, negative spoon, silver electrode, washing and drying, and mass subtraction.

Practice

1. (2024·May/June 0620/63 Q4) · Level 3

Iron rusts slowly when exposed to water and air. There is an increase in mass when iron rusts as the insoluble solid hydrated iron(III) oxide forms.

Cast iron, mild steel and stainless steel are all alloys that contain iron.

Plan an experiment to find which of these three alloys rusts most quickly when exposed to water and air. Include in your answer how the results of the experiment will show which alloy rusts most quickly.

You are provided with powdered cast iron, powdered mild steel, powdered stainless steel and common laboratory apparatus.

2. (2025·October/November 0620/62 Q4) · Level 3

Liquid alcohols can be burned as fuels in spirit burners. The energy given out can be used to heat water in a boiling tube as shown in Fig. 4.1.

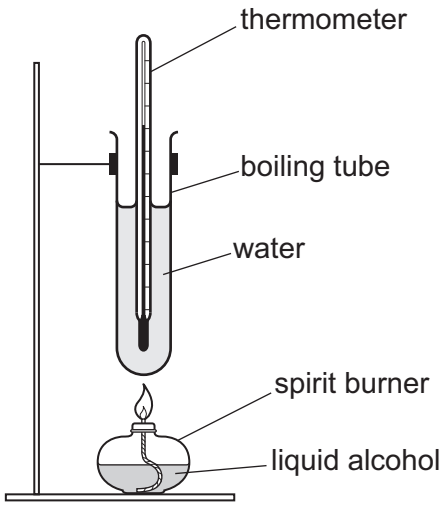


Fig. 4.1

Describe how to investigate which of three liquid alcohols, ethanol, propanol and butanol, heats the water in the boiling tube the fastest.

You are provided with three spirit burners, samples of the three alcohols, water and common laboratory apparatus.

3. (2024·October/November 0620/63 Q4) · Level 4

A mixture contains three solid compounds:

- aluminium oxide
- calcium carbonate
- calcium chloride.

Table 4.1 gives some information about the three compounds in the mixture.

Table 4.1

name of compound	solubility in water	effect of adding aqueous sodium hydroxide
aluminium oxide	insoluble	reacts to form a soluble compound
calcium carbonate	insoluble	no effect
calcium chloride	soluble	reacts to form an insoluble compound

Plan an experiment to find the percentage by mass of calcium carbonate in the mixture. Your plan must include how you will calculate the percentage by mass of calcium carbonate in the mixture.

You are provided with a sample of the mixture, distilled water, aqueous sodium hydroxide and common laboratory apparatus.

Type 2: Plot best-fit lines and curves, then use interpolation, extrapolation, gradients and intercepts

Key Ideas A results table brings repeated measurements together; a graph makes the relationship between them visible. **Plotting** means placing each actual data pair at its correct coordinates and using a suitable trend line to answer the question. It is not just drawing an attractive line by eye. First identify the quantities on both axes and the quantity ultimately requested.

1. Complete the instrument readings and table calculations first. For a thermometer, find two numbered marks and count the small intervals between them. Divide the difference between the numbers by the number of intervals to find the °C per division. Read the top of the liquid column, not the thermometer's height on the page. If the liquid ends between marks, estimate as required; for example, halfway between 1°C divisions gives 0.5°C. If one decimal place is required, write whole temperatures as 27.0°C and keep calculated temperature changes to one decimal place too.

Burette values usually **increase downwards** as liquid is delivered. Read the lowest point of a concave meniscus at eye level. Its scale reading is not directly the volume still in the tube. The actual volume added in one titration is the **titre**, calculated as final reading minus initial reading.
For example, with an initial reading of 0.6 cm³ and a final reading of 18.4 cm³, the delivered volume is:

$$18.4 - 0.6 = 17.8 \quad \text{cm}^3$$

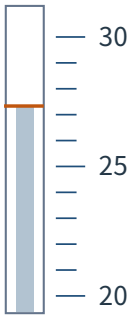
For a stop-watch with minute and second readings, read both. Since 1 min equals 60 s, convert 1 min 13 s to matching units before adding:

$$60 + 13 = 73 \quad \text{s}$$

Do not record 13 s or 1.13 s. If asked for the nearest whole number of seconds, record an integer. Use the actual scale drawn on each instrument. Existing table rows do not justify guessing the remaining readings as an arithmetic sequence.

Identify scale direction and the smallest division before reading

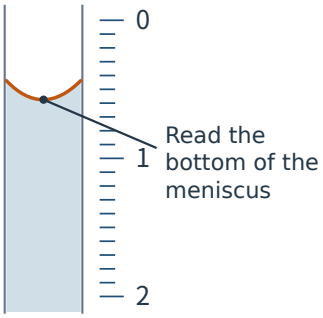
Thermometer: increases upwards



27.3°C lies between 27 and 28

Read the diagram and record the requested precision

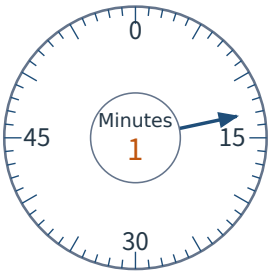
Burette: increases downwards



Reading: 0.6 cm³; not the volume remaining

Volume used = final − initial reading

Stopwatch: minutes first, then seconds



1min 13s=60+13=73s

Central number gives the minutes

Read each instrument using its own shown scale. In an answer, still check the actual graduations in the question diagram.

Subtract the correct starting value to find a change. First establish whether the relevant temperature is the temperature at a specified time or the experiment's maximum temperature. Then subtract that trial's initial temperature:

$$\text{Temperature change} = \text{Temperature at that time or maximum} - \text{Initial temperature}$$

If initial temperatures differ between trials, use the initial temperature in each row. If one experiment is measured continuously over time, use that experiment's temperature at time=0. Subtracting the preceding time's reading would calculate a different quantity.

Calculate the requested change before plotting it on the axes

Experiment A

Initial 20.0°C → later 26.0°C

Temperature change = 26.0 − 20.0

=6.0°C

Experiment B

Initial 23.0°C → later 27.0°C

Temperature change = 27.0 − 23.0

=4.0°C

B ends hotter, but A has the larger temperature change; these are different quantities.

The two experiments start at different temperatures. Each temperature change subtracts that experiment's own initial value.

Calculate the two changes separately. A rises from 20.0°C to 26.0°C:

$$\Delta T_A = 26.0 - 20.0 = 6.0 \quad ^\circ\text{C}$$

B rises from 23.0°C to 27.0°C:

$$\Delta T_B = 27.0 - 23.0 = 4.0 \quad ^\circ\text{C}$$

ΔT_A and ΔT_B denote the temperature changes in A and B. B finishes hotter, but A has the larger change. If the vertical axis is temperature change, plot 6.0 and 4.0, not 26.0 and 27.0. At time=0, each experiment's change is 0.0 even though their actual starting temperatures differ.

Some table entries come from the method rather than an instrument. Repeating the other steps while changing only water volume means that the specified solid mass stays unchanged. If a particular solution and water must total 50 cm³, subtract that row's solution volume from 50 cm³:

$$\text{Water volume} = 50 - \text{Volume of that solution}$$

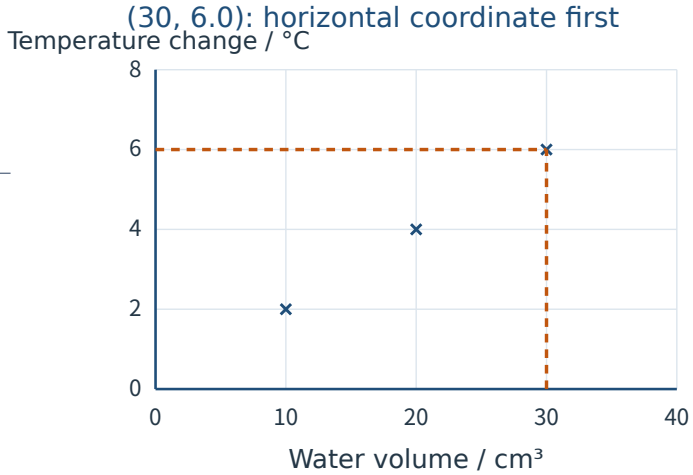
All volumes in this expression are in cm³. If a further 5 cm³ of acid is specified separately, do not automatically subtract it from this 50 cm³. First identify which components the stated total includes.

2. Choose scales and locate coordinates. The deliberately changed independent variable usually goes on the horizontal axis. The measured or required calculated dependent variable goes on the vertical axis. Follow any printed axis labels and horizontal scale; do not swap the axes. For example, with water volume horizontal and temperature change vertical, a row containing 30 cm³ and 6.0°C gives (30,6.0). Coordinates list the horizontal value first.

Each table row gives one point; every scale division must represent the same interval

Data table

Water volume / cm³	Temperature change / °C
10	2.0
20	4.0
30	6.0



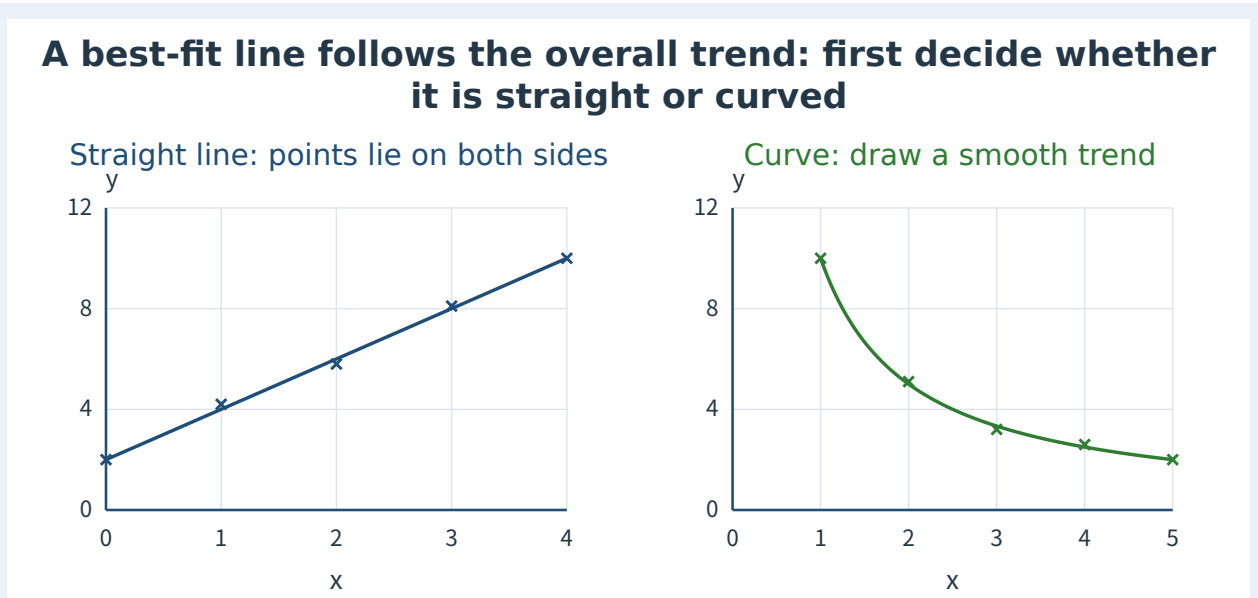
The x-value 30 and y-value 6 come from the same row; temperature change is not the maximum temperature.

The data are paired as horizontal and vertical coordinates on linear scales. Dashed lines show how to locate a point; they are not additional experimental curves.

A **linear scale** uses equal spacing for equal numerical differences along an axis. If each large square represents 2°C, the marks 0, 2, 4, 6 and 8 must be evenly spaced. Do not suddenly change the last interval to 10°C per square just to fit a point. The two axes may use different units and values per square. Find the largest value before choosing the vertical range: all points must fit and use the grid well. The formal marking requirement for this question type is that the plotted data occupy at least half the vertical-axis height. When 0 and the grid are supplied, add clear, evenly spaced scale labels beside the original grid.

Plot small crosses or clear dots. Locate the horizontal value first, then find the height from the vertical value. Pair each row once. A horizontal axis starting at 10 or 20 is valid if that is the printed scale; do not treat its left edge as 0. After plotting, check against the table row by row, especially decimals, changes calculated from different initial temperatures and any supplied origin.

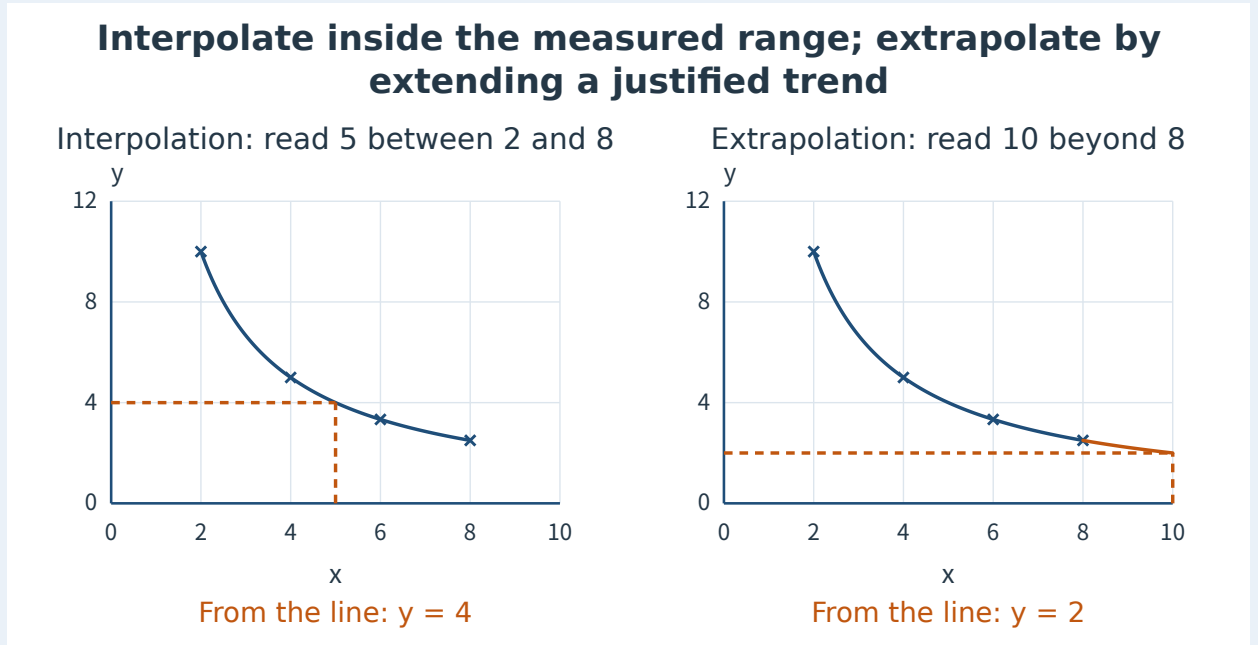
3. Best fit does not mean joining successive points with straight segments. Experimental readings may contain small errors. A **line of best fit** represents the overall relationship, staying close to most points with deviations reasonably distributed on either side. If a straight line is explicitly requested, draw one with a ruler. If a curve is requested or the data clearly bend, draw a smooth curve instead of a jagged sequence joining neighbouring points. The line need not pass through every point, but each plotted point must remain accurate.



Fit the two data sets separately. A line may reasonably miss individual points, while the points themselves remain accurately located from the original data.

If one point lies well away from the rest, first recheck the instrument reading, calculation and plotting. Do not silently change a value or delete a point to obtain a desired trend. Do not force every graph through the origin. Use the origin as a constraint only when the question requires a start at (0,0) or the data and physical meaning support it. A curve flattening means the change is slowing; it does not mean the vertical-axis value has fallen to 0.

4. Interpolate or extrapolate from the fitted line. For a given horizontal coordinate, follow a vertical line to the **fitted line**, then move horizontally to the vertical axis to read the result. These vertical and horizontal guides show clearly how you worked out your answer. They must meet the actual fitted line used, rather than stopping in nearby empty space.



The same curve is used to compare interpolation and extrapolation. The solid orange section is the extrapolation; orange dashed lines guide readings to the axes.

Estimating within the measured horizontal range is **interpolation**. With measurements at 2, 4, 6 and 8, estimate y at x=5 by meeting the curve vertically and reading across to y=4. Since the relationship is curved, simply averaging two neighbouring y values is not always valid.

Estimating beyond the measured range is **extrapolation**. To find y at x=10 when data stop at 8, first extend the existing smooth trend to 10, then read y=2. Do not abruptly change direction from the final point or straighten the curve without evidence. A printed x=10 tick does not mean a measurement was made there. Extrapolation depends on the **data range**, not the grid boundary. It assumes the trend continues and is generally less reliable further from the measured range.

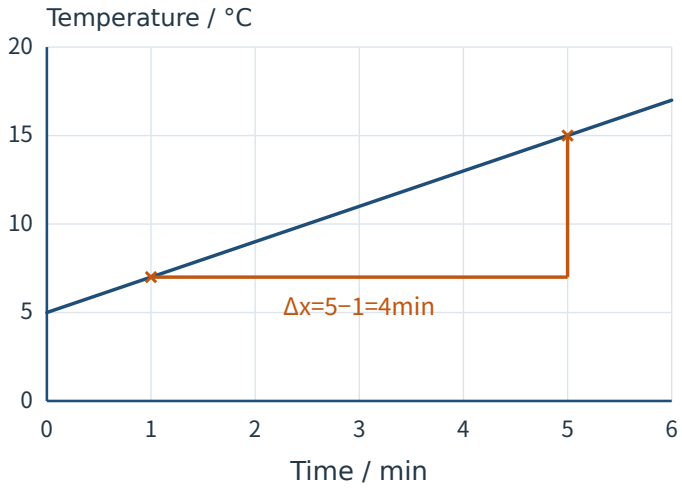
Keep the unit and make the answer consistent with your own line. Reasonable fitting and reading differences are allowed. If the official marking says correct reading from their graph, it does not require everyone to memorise an identical last digit. Each completed graph below states the value read from **the line shown**, displaying both the method and the result.

5. Gradient measures how quickly one quantity changes with another. Move rightwards along a straight line and calculate the change in vertical coordinate and the change in horizontal coordinate. Divide them to obtain the gradient. Δ means change: Δy is final minus initial vertical coordinate, and Δx is final minus initial horizontal coordinate:

$$\text{Gradient} = \frac{\text{Change in vertical coordinate}}{\text{Change in horizontal coordinate}} = \frac{\Delta y}{\Delta x}$$

Choose two widely separated, clearly readable points on the fitted line and draw a large right-angled triangle. These points need not be original experimental data points. Read each side using the corresponding axis scale. Dividing physical centimetre lengths on the paper does not directly give an experimental gradient with the correct units.

Straight-line gradient: use two well-separated points on the line



Choose (1, 7) and (5, 15)

Δy=15−7=8°C

Gradient = Δy ÷ Δx

=8 ÷ 4=2°C/min

Not 15 ÷ 5

The straight line does not pass through the origin. Calculate its gradient from differences between two coordinates; its unit is the vertical-axis unit divided by the horizontal-axis unit.

The points shown are (1 min,7°C) and (5 min,15°C). Calculate both coordinate changes first:

$$\Delta y = 15 - 7 = 8 \quad ^\circ\text{C}$$

$$\Delta x = 5 - 1 = 4 \quad \text{min}$$

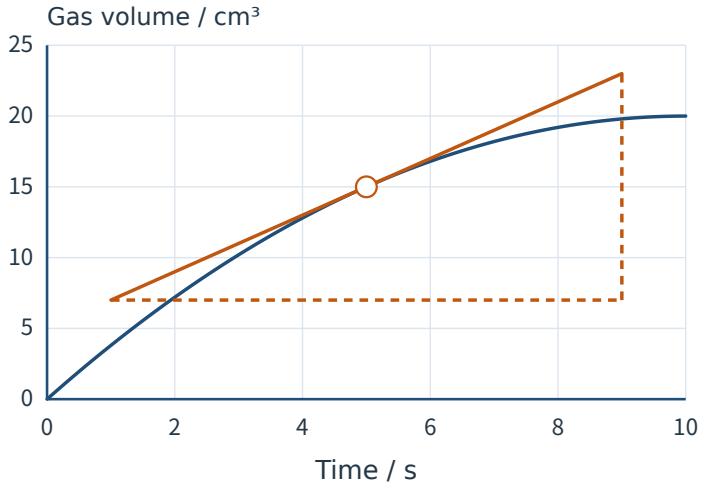
Then divide the vertical change by the horizontal change:

$$\text{Gradient} = \frac{8}{4} = 2 \quad ^\circ\frac{\text{C}}{\text{min}}$$

The line does not pass through the origin, so 15÷5 ignores its vertical intercept and is not its gradient. A line falling to the right has negative Δy and positive Δx, giving a negative gradient. If asked for the magnitude of the rate of decrease, explain the sign and magnitude as the wording requires.

A curve has different gradients at different points. To find the gradient at a specified time, draw a **tangent** at that point, matching the curve's local direction. Choose two widely separated points on the tangent and calculate Δy÷Δx in the same way. Joining two distant points on the curve gives a secant and the average change over that interval, not the instantaneous gradient at an intermediate point.

Gradient at one point on a curve: first draw a tangent there



Required time: 5 s

Choose two points on the orange tangent

(1, 7) and (9, 23)

Gradient = (23 − 7) ÷ (9 − 1)

$=2\text{cm}^3/\text{s}$

The tangent gradient gives the rate at that instant; use two tangent points, not arbitrary curve points.

The orange straight line has the same local direction as the blue curve at (5,15), so it is the tangent there.

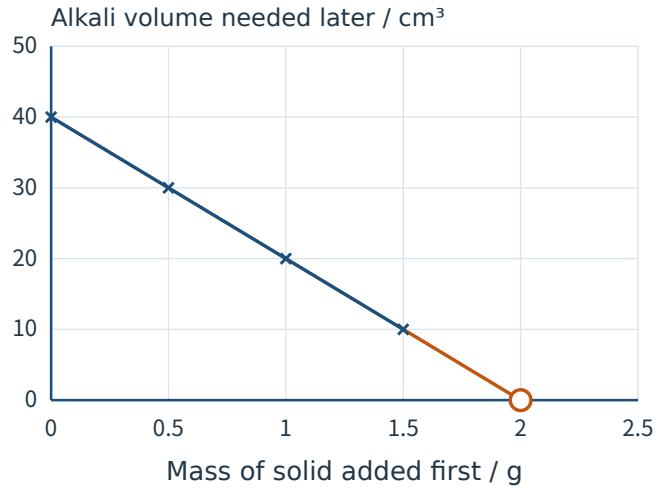
To find the gradient at 5 s, use the readable points (1 s,7 cm³) and (9 s,23 cm³) on the orange tangent. Divide their volume change by their time change:

$$\text{Gradient} = \frac{23-7}{9-1} = 2 \quad \frac{\text{cm}^3}{\text{s}}$$

This means that the gas volume is increasing by about 2 cm³ per second at that moment. Use this example to learn numerical tangent calculations. The four original examination questions in this question type practise instrument tables, best fit, interpolation and extrapolation; they do not contain a numerical tangent part.

6. An intercept is where a line meets an axis. At the horizontal-axis intercept, y=0; at the vertical-axis intercept, x=0. Read the axis label before interpreting that zero. Suppose a solid first consumes some acid, then an alkali titrates the acid remaining. If the vertical axis gives the later alkali volume, y=0 means no acid remains to need alkali. Its horizontal coordinate then gives the solid mass that just consumes all the acid.

For an intercept, identify the axis and explain what the zero means



x-intercept: y = 0

No alkali needed afterwards

The first solid has used up all the acid

From the graph: x = 2.0 g

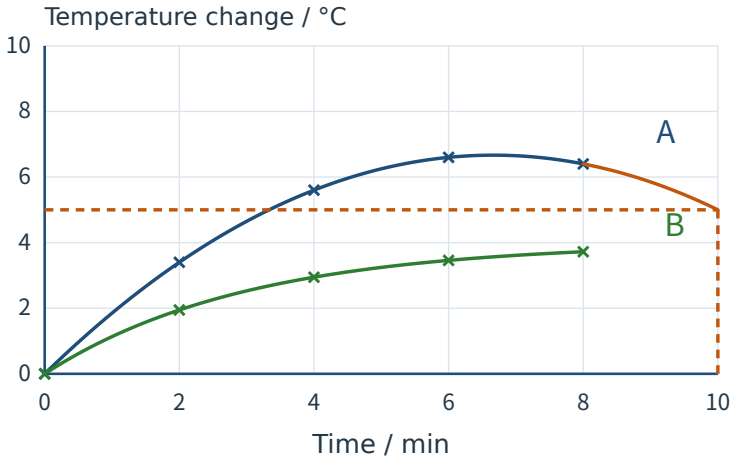
The y-intercept 40 is not the required solid mass

The graph compares solid mass with the later alkali volume. Blue represents measured data and orange extends to the horizontal axis. It does not continue into meaningless negative reagent volumes.

Solid masses of 0, 0.5, 1.0 and 1.5 g correspond to later alkali volumes of 40, 30, 20 and 10 cm³. Extend the straight line to y=0 and read 2.0 g on the horizontal axis. The 40 cm³ vertical intercept answers how much alkali is needed with no solid added, not the requested solid mass. Negative volumes on an extension beyond the horizontal axis do not represent physically measurable amounts.

For two data sets on one graph, pair, fit and label each separately. Do not join a point from A to a point from B or draw two unlabelled curves without experiment 1/2 identities. If both temperature-change curves must start at (0,0), subtract each experiment's own initial temperature and use their common change origin. Different actual initial temperatures do not prevent both changes from beginning at zero.

Label both curves and extrapolate only the one requested



Vertical axis: change from each initial temperature

Both therefore start at (0, 0)

A may rise and then fall

Extrapolate A; do not read B instead

A flat curve need not mean a zero reading

identify the curve, then use guide lines.

Each curve represents its own experiment's temperature change. Only A is extended in orange; dashed lines read from the specified time to A's extension.

Finally, check the requested actions: every reading and calculation is complete; scales are consistent and use the grid well; all points are correctly placed; the fit is the requested straight line or curve; reading guides meet the specified curve; the target lies inside or outside the measured range as identified; and units and curve labels are correct. Each check corresponds to a specific measurement or graph-reading action.

Recognising the Question Identify whether the vertical axis shows raw readings, changes or volume differences. Then distinguish plot, best fit, interpolate, extrapolate and intercept. Complete the table and graph before giving a reading that can be checked against them.

Worked Example 1 (2025·May/June 0620/62 Q2(c)) · Level 3

A student investigates the temperature change when anhydrous lithium chloride dissolves in water.

The student does five experiments.

Experiment 1

- Use a 50 cm^3 measuring cylinder to pour 40 cm^3 of distilled water into a 100 cm^3 beaker.
- Use a thermometer to measure the initial temperature of the water.
- Add a 2.0 g sample of anhydrous lithium chloride to the water in the beaker.
- Continually stir the mixture in the beaker using the thermometer.
- Measure the highest temperature reached by the mixture in the beaker.
- Empty the beaker and rinse the beaker with distilled water.

Experiment 2

- Repeat Experiment 1 using 30 cm^3 of distilled water instead of 40 cm^3.

Experiment 3

- Repeat Experiment 1 using 25 cm^3 of distilled water instead of 40 cm^3.

Experiment 4

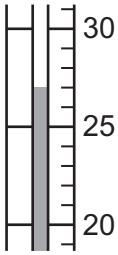
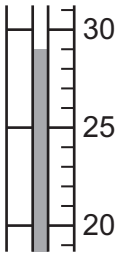
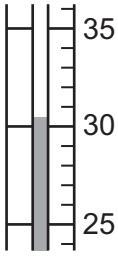
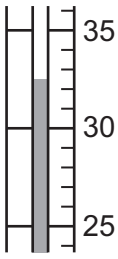
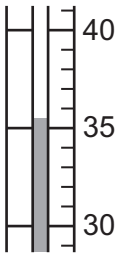
- Repeat Experiment 1 using 20 cm^3 of distilled water instead of 40 cm^3.

Experiment 5

- Repeat Experiment 1 using 15 cm^3 of distilled water instead of 40 cm^3.

(a) Use the information in the description of the experiments and the thermometer diagrams to complete Table 2.1.

Table 2.1

experiment	mass of anhydrous lithium chloride / g	volume of distilled water / cm^3	initial temperature / °C	thermometer diagram for highest temperature reached / °C	highest temperature reached / °C	temperature change / °C
1	2.0	40	22.5			
2		30	22.5			
3		25	22.5			
4		20	22.0			
5		15	22.0			

(b) Complete a suitable scale on the *y*-axis and plot your results from Experiments 1 to 5 on Fig. 2.1.

Draw a line of best fit.

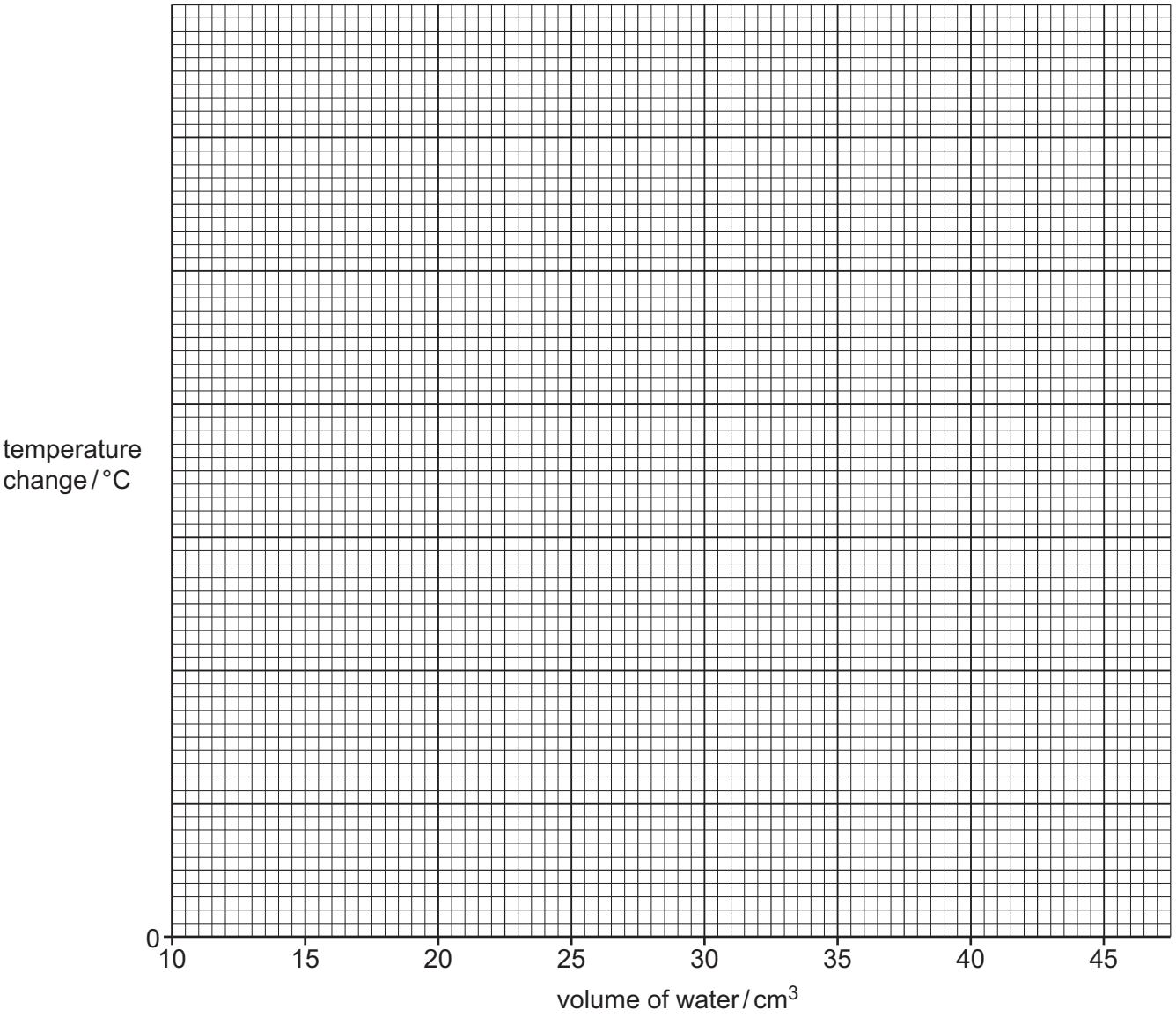


Fig. 2.1

(c) Extrapolate the line on your graph in Fig. 2.1 to deduce the temperature change when Experiment 1 is repeated with 45 cm^3 of water instead of 40 cm^3 of water.

Show clearly on Fig. 2.1 how you worked out your answer.

Solution

Identify all the parts included here. This unit contains (a), reading five thermometers and completing the table; (b), choosing the vertical scale, plotting and fitting; and (c), extrapolating to 45 cm^3 of water. Anhydrous lithium chloride is LiCl without water of crystallisation in the solid. Anhydrous does not mean that no water is added in the experiment. All five trials use 2.0 g of solid as specified, changing only water volume.

(a) Complete the table. Each small thermometer division is 1°C; a liquid column halfway between marks is read to 0.5°C. The readings in order are 27.0, 29.0, 30.5, 32.5 and 35.5°C.
For row 1, subtract that row's initial temperature of 22.5°C:

$$27.0 - 22.5 = 4.5 \quad ^\circ\text{C}$$

Row 4 must use its own initial temperature of 22.0°C:

$$32.5 - 22.0 = 10.5 \quad ^\circ\mathrm{C}$$

Experiment	LiCl mass/g	Water volume/ cm³	Initial temperature/°C	Maximum temperature/°C	Temperature change/°C
1	2.0	40	22.5	27.0	4.5
2	2.0	30	22.5	29.0	6.5
3	2.0	25	22.5	30.5	8.0
4	2.0	20	22.0	32.5	10.5
5	2.0	15	22.0	35.5	13.5

Enter 2.0 g in all five mass rows. Record every temperature and change to one decimal place, including 27.0 and 8.0. Do not plot the raw temperature 30.5 on a temperature-change axis.

(b) Plot and fit. The printed horizontal axis is water volume and starts at 10 cm³. Here, the vertical scale runs evenly from 0 to 14°C, with 2°C per 10 small squares, or 0.2°C per small square. The maximum, 13.5°C, nearly reaches the top, so the data use more than half the vertical height and every point fits. Plot (40,4.5), (30,6.5), (25,8.0), (20,10.5) and (15,13.5) accurately, then draw a smooth decreasing best-fit curve that gradually flattens as water volume increases.

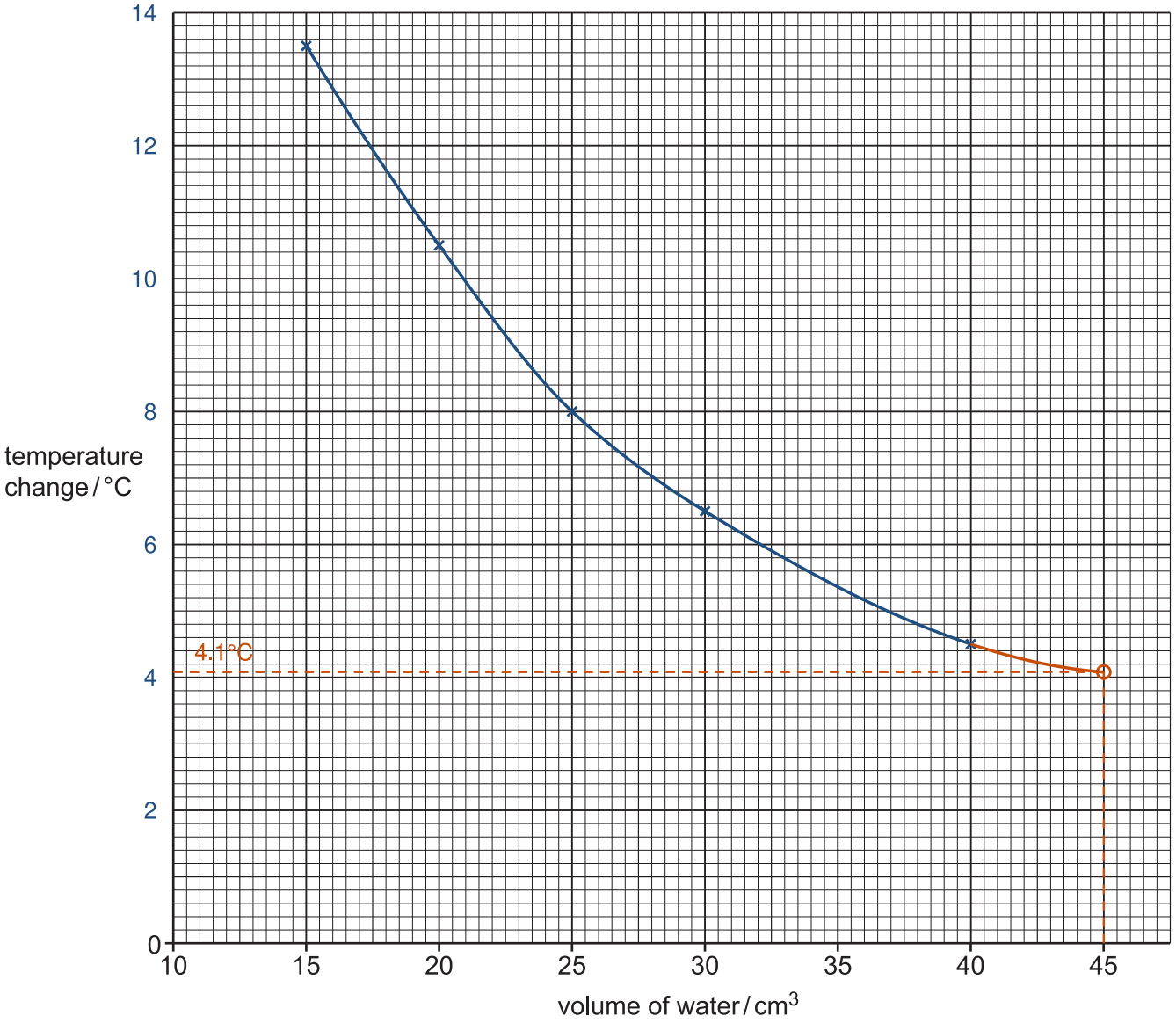


Fig. 2.1

The original grid is completed with a vertical scale, five actual data points and a reasonable smooth fit. The orange extension reaches 45 cm³, and dashed guides show the reading from this curve.

(c) Extrapolate beyond the data. The largest measured water volume is 40 cm³, so 45 cm³ lies outside the range. Extend the curve from (b) smoothly to 45, then go vertically from 45 on the horizontal axis to the extension and horizontally back to the vertical axis. The curve shown gives **about 4.1°C**. Report a temperature change with °C; do not add the initial temperature to turn it into a maximum temperature. The official criteria are a smooth extension, a correct reading from the student's own extension and the unit. 4.1°C is the value for this graph, not a universally required final digit for every reasonable fit.

Worked Example 2 (2025·October/November 0620/61 Q2(c)) · Level 3

A student investigates the reaction between dilute hydrochloric acid and aqueous sodium hydroxide.

The student does five experiments.

Experiment 1

- Fill a burette with aqueous sodium hydroxide.
- Run some of the aqueous sodium hydroxide out of the burette so that the level of the aqueous sodium hydroxide is on the burette scale.
- Record the initial burette reading.
- Use a volumetric pipette to add 25.0 cm^3 of dilute hydrochloric acid to a conical flask.
- Stand the conical flask on a white tile.
- Add five drops of thymolphthalein indicator to the conical flask.
- Slowly add aqueous sodium hydroxide from the burette to the conical flask, while swirling the flask, until the solution just changes colour. This is the end-point.
- Record the final burette reading.

Experiment 2

- Empty the conical flask and rinse it with distilled water.
- Refill the burette with aqueous sodium hydroxide.
- Run some of the aqueous sodium hydroxide out of the burette so that the level of the aqueous sodium hydroxide is on the burette scale.
- Record the initial burette reading.
- Use the volumetric pipette to add 25.0 cm^3 of dilute hydrochloric acid to the conical flask.
- Add 0.25 g of sodium hydrogencarbonate to the conical flask and swirl the flask.
- Stand the conical flask on the white tile.
- Add five drops of thymolphthalein indicator to the conical flask.
- Slowly add aqueous sodium hydroxide from the burette to the conical flask, while swirling the flask, until the solution just changes colour. This is the end-point.
- Record the final burette reading.

Experiment 3

- Repeat Experiment 2 using 0.50 g of sodium hydrogencarbonate instead of 0.25 g.

Experiment 4

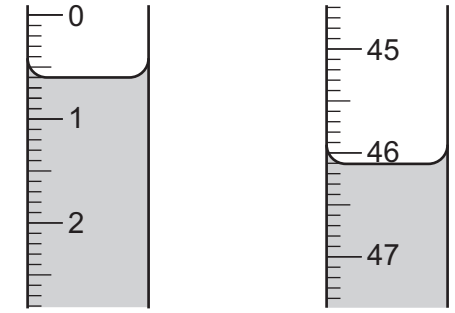
- Repeat Experiment 2 using 1.00 g of sodium hydrogencarbonate instead of 0.25 g.

Experiment 5

- Repeat Experiment 2 using 1.50 g of sodium hydrogencarbonate instead of 0.25 g.

(a) Use the information in the descriptions of Experiments 1 to 5 and the burette diagrams in Fig. 2.1 and Fig. 2.2 to complete Table 2.1.

Experiment 1

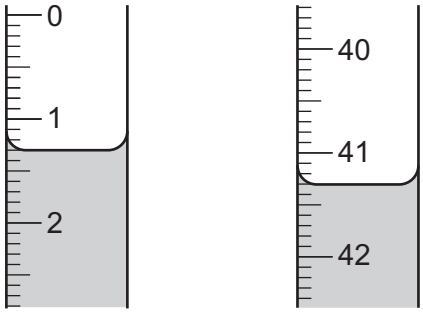


initial burette reading

final burette reading

Fig. 2.1

Experiment 2



initial burette reading

final burette reading

Fig. 2.2

Table 2.1

	Experiment 1	Experiment 2	Experiment 3	Experiment 4	Experiment 5
volume of dilute hydrochloric acid / cm^3					
mass of sodium hydrogencarbonate / g	0.00	0.25	0.50	1.00	1.50
final burette reading / cm^3			36.0	24.1	13.5
initial burette reading / cm^3			1.4	0.3	0.5
volume of aqueous sodium hydroxide added to reach the end-point / cm^3					13.0

(b) Complete a suitable scale on the *y*-axis and plot the results from Experiments 1 to 5 in Table 2.1 on Fig. 2.3. Draw a straight line of best fit.

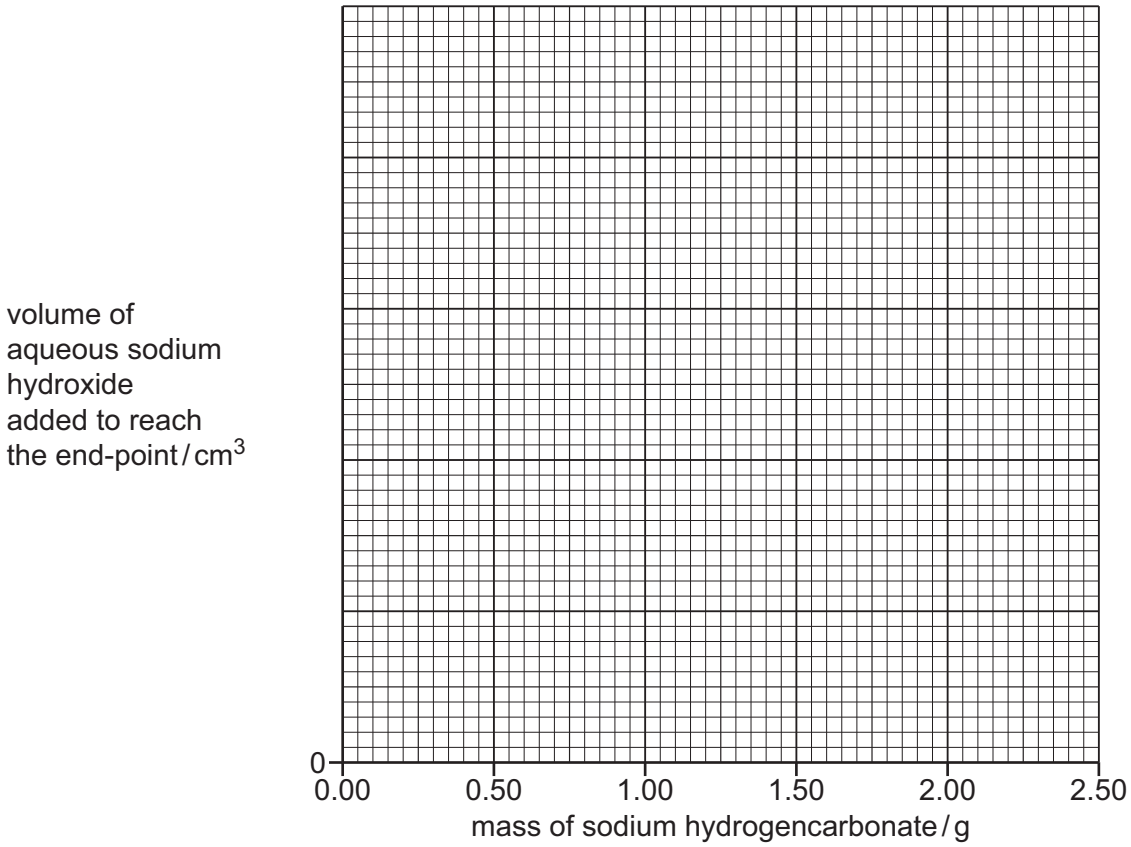


Fig. 2.3

(c) The sodium hydrogencarbonate added in Experiments 2 to 5 neutralises some of the dilute hydrochloric acid in the conical flask. The sodium hydroxide used in the titration neutralises the remaining dilute hydrochloric acid.

Extrapolate the line on your graph in Fig. 2.3 and deduce the mass of sodium hydrogencarbonate needed to neutralise **all** of the dilute hydrochloric acid in the conical flask.

Show clearly **on Fig. 2.3** how you worked out your answer.

Solution

Extract initial and final burette readings separately from the experimental conditions. This unit includes (a), the complete table; (b), a straight best-fit line; and (c), extrapolation to exact complete neutralisation of the acid. Every trial uses 25.0 cm³ of the same hydrochloric acid, changing only the mass of sodium hydrogencarbonate, $NaHCO_3$, added first. Thymolphthalein signals the titration end-point by colour change. Calculate the actual NaOH solution delivered from the burette readings.

(a) Read each burette and subtract initial from final. In trial 1, the initial reading is 0.6 cm³ and the final reading 46.1 cm³:

$46.1 - 0.6 = 45.5 \quad \text{cm}^3$

In trial 2, the initial reading is 1.3 cm³ and the final reading 41.3 cm³:

$41.3 - 1.3 = 40.0 \quad \text{cm}^3$

Burette readings increase downwards, so do not reverse the two columns. For each remaining trial, subtract its own initial reading from its final reading:

$36.0 - 1.4 = 34.6 \quad \text{cm}^3$

$24.1 - 0.3 = 23.8 \quad \text{cm}^3$

$13.5 - 0.5 = 13.0 \quad \text{cm}^3$

Experiment	HCl volume/cm³	$NaHCO_3$ mass/g	Initial reading/cm³	Final reading/cm³	NaOH volume/cm³
1	25.0	0.00	0.6	46.1	45.5
2	25.0	0.25	1.3	41.3	40.0
3	25.0	0.50	1.4	36.0	34.6
4	25.0	1.00	0.3	24.1	23.8
5	25.0	1.50	0.5	13.5	13.0

Enter 25.0 for all five hydrochloric-acid volumes. Record instrument volume readings and calculated titres to one decimal place. Retain the solid masses exactly as supplied, such as 0.00 and 0.25, without rounding them unnecessarily.

(b) Draw the required straight line. Use solid masses 0.00, 0.25, 0.50, 1.00 and 1.50 g horizontally and the **actual delivered volumes** 45.5, 40.0, 34.6, 23.8 and 13.0 cm³ vertically. Here the vertical scale runs evenly from 0 to 50 cm³, with 10 cm³ per 10 small squares. Plot all five points accurately and use a ruler for one straight line close to them. Do not join them as separate segments or force the fit through the origin.

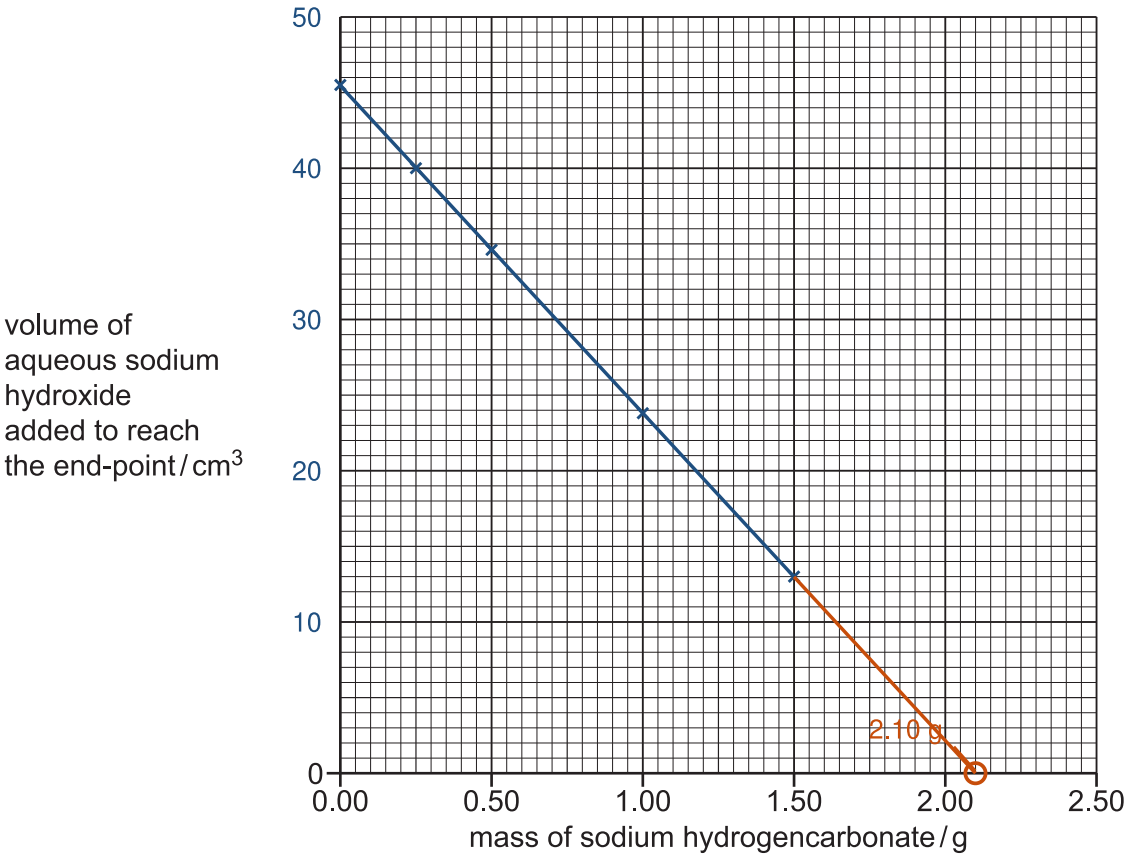


Fig. 2.3

The five points on the original grid come from the complete titration table. The orange straight-line extension meets the horizontal axis at about 2.10 g in this graph.

(c) Explain why the horizontal intercept is required. $NaHCO_3$ added first neutralises some acid; the later NaOH neutralises only what remains. If $NaHCO_3$ just neutralises all the hydrochloric acid, no NaOH is needed, so the vertical-axis volume is 0.0 cm³. Extend the straight line rightwards to y=0 and read its horizontal coordinate. The line shown gives **about 2.10 g**.

Do not read the vertical intercept of 45.5 cm³ or use the largest measured mass, 1.50 g. The intercept lies beyond the measured range, so show the straight-line extension. The official marks follow the student's reasonable best-

fit line and its horizontal-intercept reading, with g as the unit. This graph's 2.10 g does not impose identical final digits on all reasonable plots.

Practice

1. (2025·May/June 0620/63 Q2(d)(ii)) · Level 3

A student investigates how the rate of reaction between dilute hydrochloric acid and aqueous sodium thiosulfate changes as the concentration of the aqueous sodium thiosulfate decreases. During the reaction, the solution slowly becomes cloudy. As the solution becomes cloudy, it becomes more difficult to see through the solution.

The student does five experiments.

Experiment 1

- Use a 50 cm^3 measuring cylinder to pour 50 cm^3 of aqueous sodium thiosulfate into a 100 cm^3 beaker.
- Use a 10 cm^3 measuring cylinder to pour 5 cm^3 of dilute hydrochloric acid into the beaker containing the aqueous sodium thiosulfate.
- Immediately start a stop-watch and stir the contents of the beaker.
- Stand the beaker on a printed sheet and look down from above the beaker as shown in Fig. 2.1.

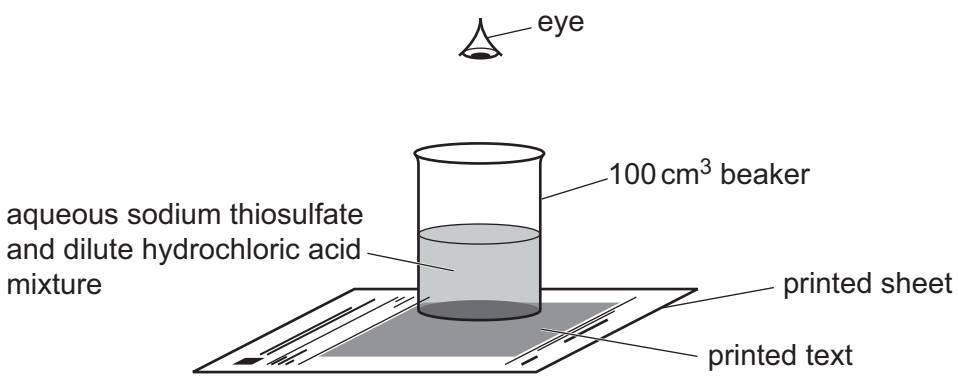


Fig. 2.1

- When the text on the printed sheet is **not** visible, stop the stop-watch and record the time to the nearest whole number of seconds.
- Empty the contents of the beaker and rinse the beaker with distilled water.

Experiment 2

- Repeat Experiment 1 using 40 cm^3 of aqueous sodium thiosulfate instead of 50 cm^3. Use the 50 cm^3 measuring cylinder to add 10 cm^3 of distilled water to the beaker **before** adding the dilute hydrochloric acid.

Experiment 3

- Repeat Experiment 2 using 30 cm^3 of aqueous sodium thiosulfate and 20 cm^3 of distilled water.

Experiment 4

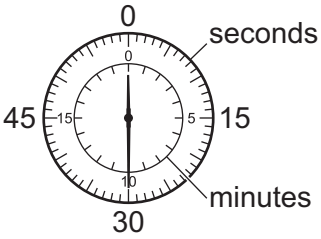
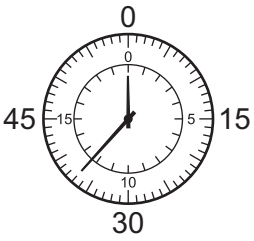
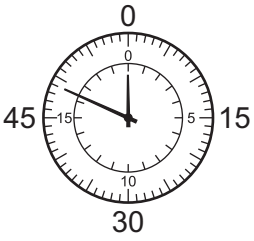
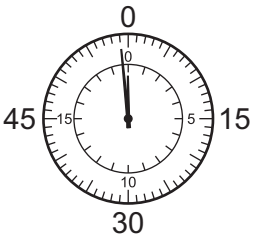
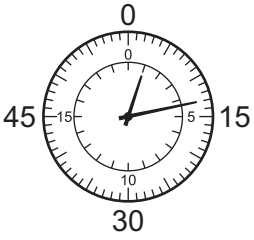
- Repeat Experiment 2 using 25 cm^3 of aqueous sodium thiosulfate and 25 cm^3 of distilled water.

Experiment 5

- Repeat Experiment 2 using 20 cm^3 of aqueous sodium thiosulfate and 30 cm^3 of distilled water.

(a) Use the information in the description of the experiments and the stop-watch diagrams to complete Table 2.1.

Table 2.1

experiment	volume of aqueous sodium thiosulfate / cm^3	volume of distilled water / cm^3	volume of dilute hydrochloric acid / cm^3	stop-watch diagram	time taken for text to **not** be visible / s
1	50	0	5		
2					
3					
4					
5					

(b) Write a suitable scale on the *y*-axis and plot the results from Experiments 1 to 5 on Fig. 2.2.

Draw a curve of best fit.

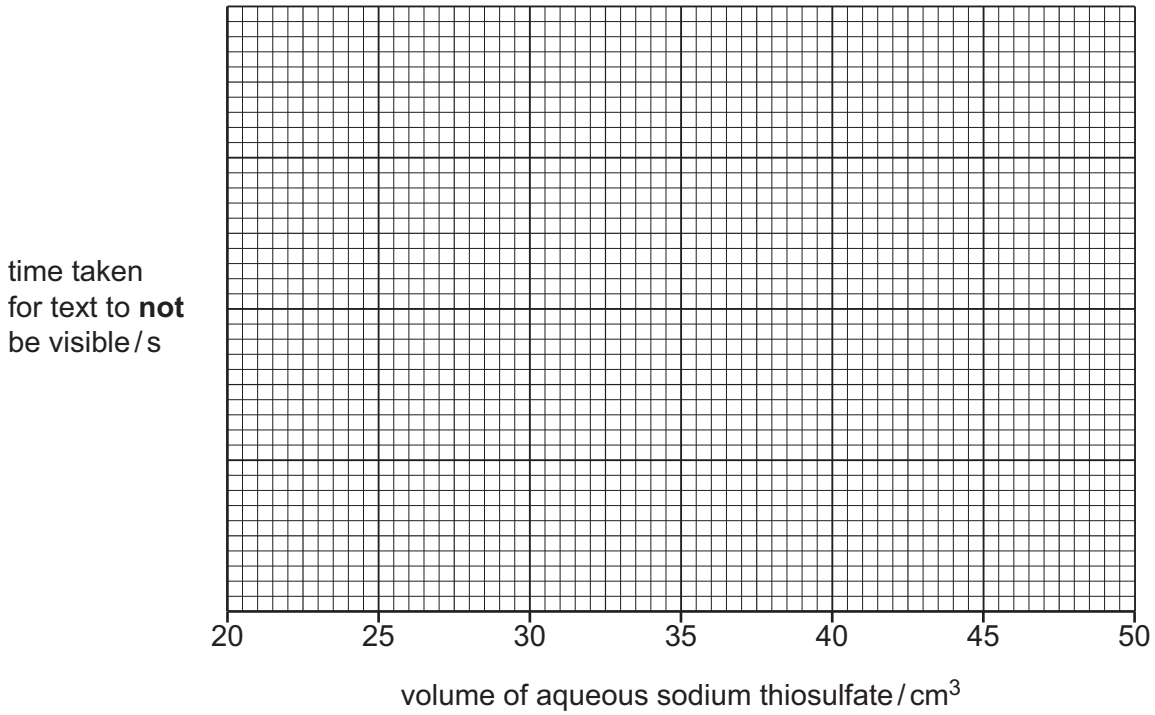


Fig. 2.2

(d) In each experiment the total volume of aqueous sodium thiosulfate and distilled water is constant.

(i) Calculate the volume of distilled water needed when the volume of aqueous sodium thiosulfate is 37 cm^3.

(ii) From your graph in Fig. 2.2, deduce the time for the text to **not** be visible when the volume of aqueous sodium thiosulfate is 37 cm^3.

Show clearly on Fig. 2.2 how you worked out your answer.

2. (2025·February/March 0620/62 Q2(c)) · Level 3

A student investigates the temperature change when aqueous copper(II) sulfate reacts separately with two different metals, zinc and iron.

The student does two experiments.

Experiment 1

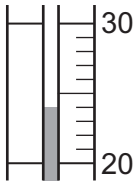
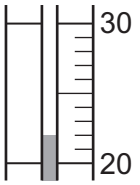
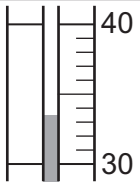
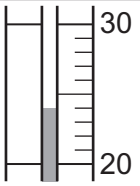
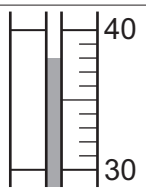
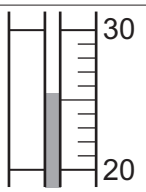
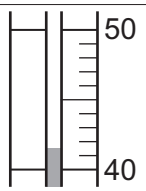
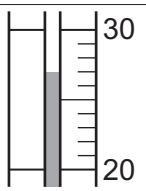
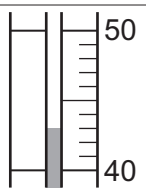
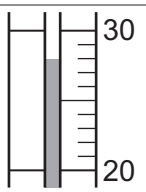
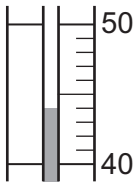
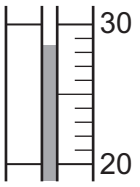
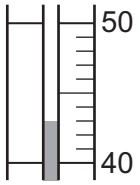
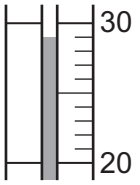
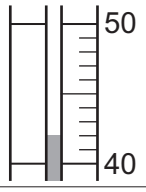
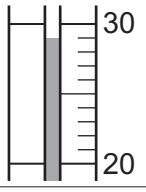
- Use a 50 cm^3 measuring cylinder to pour 30 cm^3 of aqueous copper(II) sulfate into a 100 cm^3 beaker.
- Use a thermometer to measure the temperature of the contents of the beaker. This is the temperature at time = 0 s.
- Add 3 g of zinc powder to the beaker. At the same time start a stop-watch.
- Using the thermometer, continually stir the mixture in the beaker.
- Measure the temperature of the mixture every 30 seconds for 210 seconds.
- Empty and rinse the beaker with distilled water.

Experiment 2

- Repeat Experiment 1, using 3 g of iron powder instead of 3 g of zinc powder.

(a) Complete Table 2.1 by using the thermometer diagrams and calculating the temperature changes from the temperature at 0 s.

Table 2.1

	Experiment 1			Experiment 2		
time in s	thermometer diagram	temperature in °C	temperature change since time = 0 s in °C	thermometer diagram	temperature in °C	temperature change since time = 0 s in °C
0		24.0	0.0		22.0	0.0
30						
60						
90						
120						
150						
180						
210						

(b) Complete a suitable scale on the *y*-axis, and plot your results for Experiment 1 and Experiment 2 on Fig. 2.1. The point at (0, 0) has been plotted for you.

Draw **two** curves of best fit. Both curves must start at (0, 0).

Label both curves.

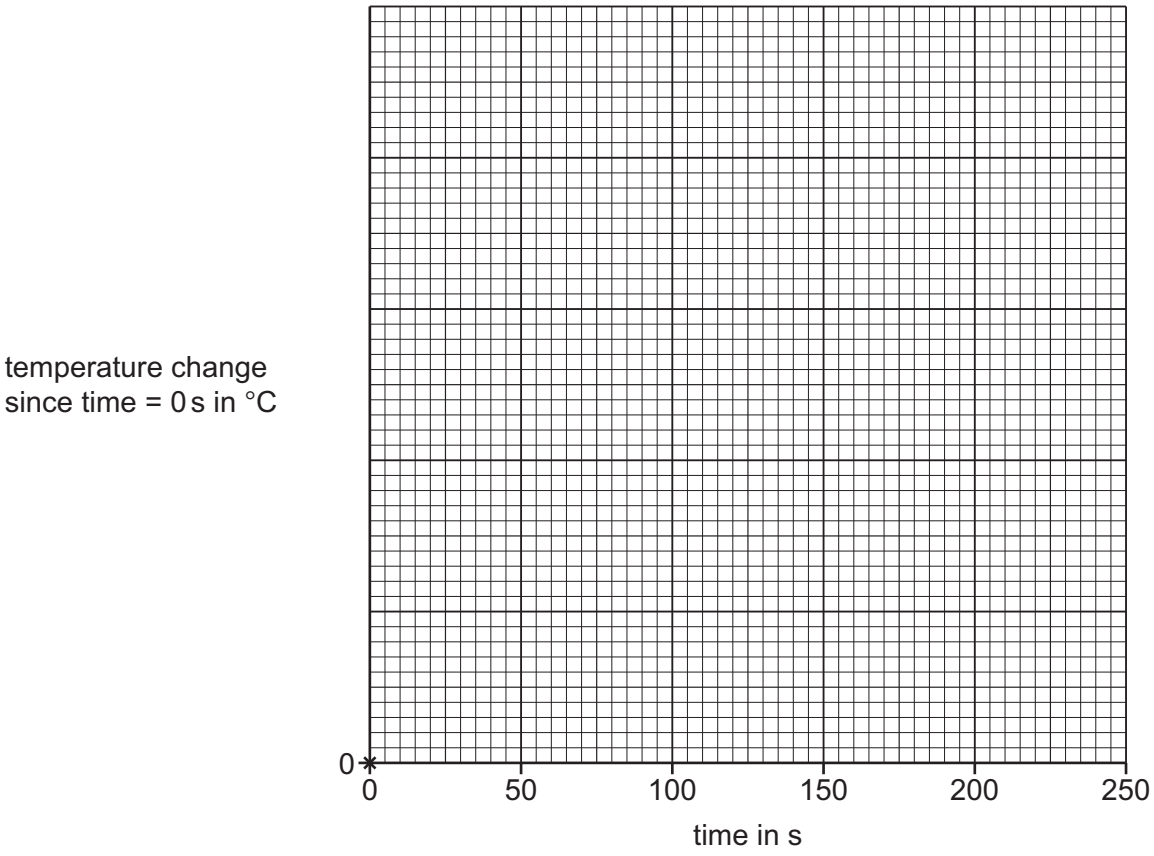


Fig. 2.1

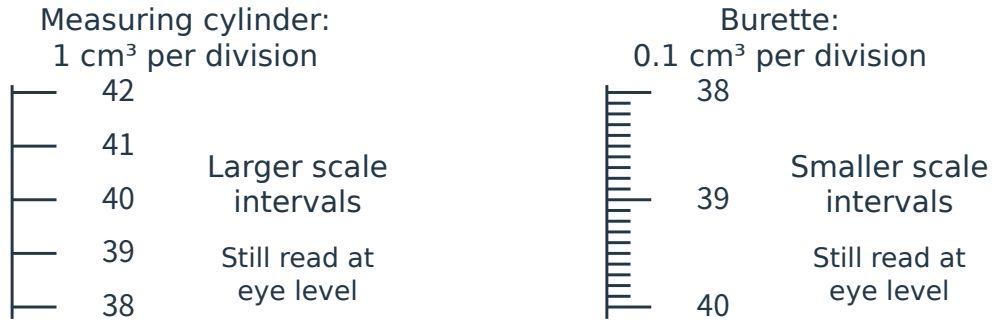
(c) Extrapolate the curve for Experiment 1 on your graph in Fig. 2.1 to deduce the temperature change since time = 0 s in Experiment 1 after 240 seconds.

Show clearly **on Fig. 2.1** how you worked out your answer.

Type 3: Explain how experimental changes affect measured results

Key Ideas **Explain which source of error an improvement changes.** Accuracy describes how close a result is to the true value being measured. Close repeat results show consistency, but do not guarantee freedom from a shared bias. For a specified volume, an instrument with a suitable range, finer graduations and proper working condition can improve measurement and delivery. Correct eye level and rinsing are still needed. Changing apparatus without a reason is not a complete answer.

Fine graduations improve precision, but the instrument must also suit the volume measured



Use the question diagram and instrument specification for actual graduations and tolerances; these are not universal specifications.

Improvement: choose a more suitable precise instrument → finer, more reliable volume readings → more accurate quantities.

The two scales represent instruments with different graduations. No meniscus is shown, so these are not numerical reading questions. The actual choice still depends on the required volume and instrument range.

Suppose 40 cm³ of liquid is required. The given measuring cylinder has 1 cm³ divisions, while the given burette has 0.1 cm³ divisions and enough capacity. The burette's finer scale gives a more detailed measurement of delivered volume and helps measure the required amount accurately. Write use a burette, because it measures the volume more accurately than the given measuring cylinder. This compares the two specified instruments, not every model of measuring cylinder. A burette still requires final minus initial reading, not just its final scale position.

Use repeats to check consistency, then calculate a mean where appropriate. Repeats allow results under the same conditions to be compared and can reveal occasional timing, observation or recording problems. An anomalous result lies noticeably apart from the others. Check the record and procedure and repeat it, rather than deleting it merely to obtain a preferred answer. Repetition alone cannot remove a leak, heat loss or the same dilution error that acts in the same direction every time.

Repeats reveal inconsistency; only comparable results should be averaged

Repeat 1	Repeat 2	Repeat 3	Repeat 4
34s	35s	35s	56s

56 s is clearly different → check records and procedure → repeat under the same conditions

If 56 s is confirmed as a timing error, the other mean = (34 + 35 + 35) ÷ 3 ≈ 34.7 s

Repeated gas leaks, heat loss or identical dilution can bias every result in the same direction.

Do not reject a result merely because it is unexpected; give the evidence, checks and conditions for exclusion.

Four times are given, with 56 s assumed to have been confirmed as a timing error. Repeats check consistency; calculate a mean only from justified valid data.

For four times to the same end-point, 34, 35, 35 and 56 s, 56 s clearly differs from the other three and should first be checked and repeated. Only after confirming it as a timing error should the other three valid values be averaged:

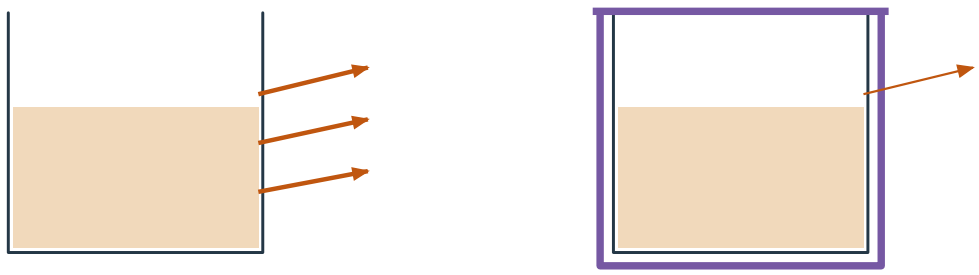
$$\text{Mean time} = \frac{34 + 35 + 35}{3} \approx 34.7 \quad \text{s}$$

The unrounded value is 34.666… s. Without a confirmed problem, do not discard 56 s without justification.

In another experiment, a similar gas volume might leak out every time. Three results could agree closely but still give a mean that is too low. Repair the leak; simply doing more trials does not solve it.

Reducing heat loss improves measurements of an exothermic temperature rise. Heat flows from the warmer reaction mixture to cooler surroundings, so the thermometer records a smaller rise than the reaction could produce with better insulation. A polystyrene cup insulates better than an ordinary glass beaker, and a suitable lid also reduces heat transfer. Connect the explanation explicitly: better insulation, less heat lost and a more reliable measured temperature rise.

For an exothermic temperature rise, insulation reduces heat transfer to the surroundings



Ordinary glass beaker | Polystyrene cup + suitable lid

Heat flows from the warmer mixture to cooler surroundings; insulation reduces losses without eliminating them.

For the same exothermic reaction, an ordinary cup might give $\Delta T = 8°C$ and improved insulation $\Delta T = 10°C$.

Arrows show heat transfer from the warmer mixture to cooler surroundings. Arrow numbers and widths indicate a trend, not a quantitative energy balance. The example temperature rises are not inevitable values for a particular real experiment.

Suppose equal reactant quantities at the same initial temperature give an 8°C rise in an ordinary cup and a 10°C rise after insulation is improved. Changing the cup has not made the reactants release extra energy. Less heat escapes before the maximum is read, so more of the reaction's heat appears as a temperature rise in the mixture. Actual values depend on measurement; 8 and 10 here explain the cause. More accurate alone does not identify reduced heat loss and may be incomplete in a two-mark explanation.

Whether a water bath helps depends on the role of temperature in the experiment. When studying concentration's effect on rate, temperature may be controlled, and a thermostatically controlled water bath can keep trials at the same temperature. When measuring the temperature change caused by dissolving or reacting, the rise itself is the result of interest. An external constant-temperature bath exchanges heat with the mixture and tends to hold it at the set temperature, changing the rise being measured.

Whether a water bath is suitable depends on what the experiment measures

Aim: compare rates at a fixed temperature

A constant-temperature bath controls temperature and prevents it affecting the rate comparison
for example, all concentrations react at 25°C

Aim: measure the reaction's own temperature rise

A constant-temperature bath absorbs or supplies heat, pulling the sample towards its set temperature and reducing or changing the rise being measured

State the experimental aim, then explain the effect on the result; a water bath is not automatically an improvement.

The right-hand diagram uses an external constant-temperature water bath. This differs from controlling temperature while investigating reaction rate.

To compare rates at two concentrations at 25°C, first placing both in a 25°C water bath can support a fair comparison. If the task instead measures how much a salt raises water temperature from 25°C on dissolving, keeping the cup in a bath maintained at 25°C transfers added heat to the bath and suppresses or alters the

measured change. State the original aim before explaining an apparatus change's effect. Constant temperature is not automatically an advantage.

For rinsing errors, track concentration and amount in mol separately. When changing the acid in a burette, wash with water to remove the previous acid, then rinse with the new acid to replace residual water. Omitting the second step lets water dilute the new acid. Each cm³ then contains fewer moles, so reaching the same end-point requires a larger acid volume. Rinsing the conical flask with the carbonate solution after water would leave extra unmeasured carbonate and also increase the acid requirement. A little distilled water in the flask does not increase sample moles.

Trace rinsing errors through concentration or mole changes to the required titre

Burette rinsed only with water; titrant does not replace residual water
Acid is diluted: fewer moles per unit volume
→ Larger volume needed for the same endpoint

Conical flask then rinsed with the carbonate sample solution
Extra unmeasured carbonate remains in the flask
→ More acid needed; titre too large

Only a little distilled water remains in the flask
Total measured carbonate moles unchanged
→ Acid moles for exact reaction unchanged

A water rinse first removes the previous acid; rinsing afterwards with the new acid prevents its dilution.

The three situations are independent comparisons, each retaining its original measured sample and end-point. Explain volume bias through chemical quantities, rather than merely calling the apparatus dirty.

Suppose a sample originally needs 20.0 cm³ of 0.100 mol/dm³ acid to reach the specified end-point. Since 20.0 cm³ is 0.0200 dm³, the acid amount required is:

$$n = 0.100 \times 0.0200 = 0.00200 \quad \text{mol}$$

If residual water dilutes the titrant to 0.0800 mol/dm³, the volume needed to supply the same 0.00200 mol is:

$$V = \frac{0.00200}{0.0800} = 0.0250 \quad \text{dm}^3$$

$$0.0250 \quad \text{dm}^3 = 25.0 \quad \text{cm}^3$$

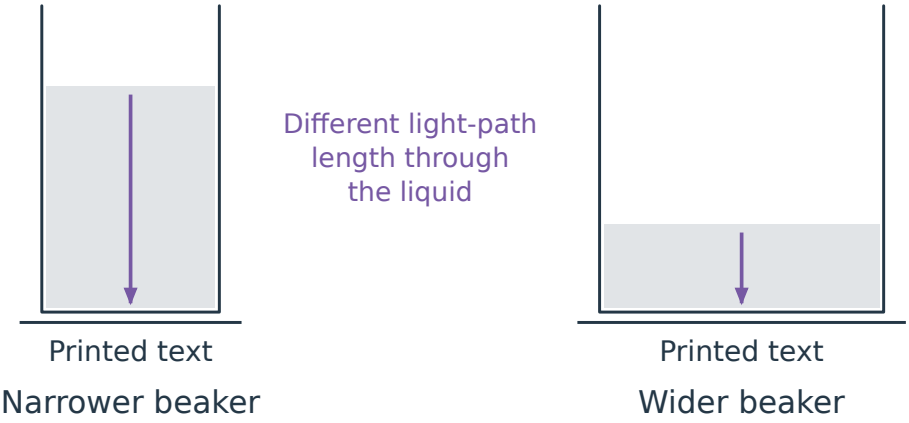
25.0 cm³ exceeds the original 20.0 cm³, so the titre increases.

A different error carries extra sample into the flask, increasing the acid moles required in the first place. It can also increase the volume, but for a different reason. Explain that reason rather than calling both errors contamination and therefore inaccurate.

The disappearing-text end-point also depends on liquid depth. Viewed from above, light from the text below the vessel passes through the full liquid layer. The same total volume in a wider beaker forms a shallower layer and shorter light path. The time when the text disappears therefore changes even if acid and sodium thiosulfate concentrations, temperature and chemical reaction rate remain unchanged. Keep beaker shape, width, liquid volume and viewing method the same throughout a comparison.

Disappearing-text experiments need the same vessel geometry: equal volume is shallower in a wider beaker

Same solution volume, concentration and reaction conditions



The text disappears at a different time because the observed endpoint changes; this alone does not show a different reaction rate.

Cross-sections of cylindrical vessels show that the same volume forms a shallower layer in the wider vessel. Grey identifies the liquid only. Arrows represent optical path length through it, not a chemical reaction direction.

The same 55 cm³ of reaction mixture is shallower in a wide beaker than in a narrow one. From this geometry, with other conditions unchanged and observation from above, the shallower layer generally needs more cloudiness to hide the same text, so the observed time may be longer. This direction is an inference from the specified geometry. If an examination question asks only why the beaker should not change, state that changing liquid depth changes the measured time for the text to disappear. Do not invent a number of seconds or equate a changed observational end-point directly with a slower reaction.

Repeats mainly check _____ of results and do not automatically remove shared _____. An insulated cup reduces _____. A constant-temperature water bath may alter the _____ that the experiment is meant to measure. Diluting the acid usually makes the titre at the same end-point _____. Changing beaker size changes liquid _____ and therefore the _____ for the text to disappear.

Answers: consistency; systematic bias; heat loss to the surroundings; temperature change; increase; depth; time

Recognising the Question Connect each change to the original aim: which reading the apparatus affects, which heat loss insulation reduces, whether residual water changes concentration or moles, and which optical path beaker width changes. Repeats check consistency. An error acting in the same direction each time requires its cause to be corrected, not just more trials.

Worked Example 1 (2025·May/June 0620/62 Q2(g)) · Level 3

A student investigates the temperature change when anhydrous lithium chloride dissolves in water.

The student does five experiments.

Experiment 1

- Use a 50 cm^3 measuring cylinder to pour 40 cm^3 of distilled water into a 100 cm^3 beaker.
- Use a thermometer to measure the initial temperature of the water.
- Add a 2.0 g sample of anhydrous lithium chloride to the water in the beaker.
- Continually stir the mixture in the beaker using the thermometer.
- Measure the highest temperature reached by the mixture in the beaker.
- Empty the beaker and rinse the beaker with distilled water.

(f) Explain why the results obtained would be more accurate if the beaker used in each experiment was replaced by a polystyrene cup.

(g) **(i)** Explain why using a burette instead of a measuring cylinder is an improvement.

(ii) Explain why standing the beaker in a water-bath is **not** an improvement.

Solution

Establish the original aim first. Anhydrous LiCl dissolves in water, and the student records the initial water temperature and maximum mixture temperature to investigate the temperature change on dissolving. Parts (f), (g)(i) and (g)(ii) evaluate three different apparatus changes; answer each separately.

(f) Replace the glass beaker with a **polystyrene cup**. Polystyrene is a better insulator and **reduces heat transfer to the surroundings**. Less heat is lost before the maximum temperature is measured, so the recorded rise better represents the change caused by dissolving. For the 2 marks, include both better insulation and reduced heat loss.

(g)(i) A **burette** instead of the supplied measuring cylinder **measures the specified water volume more accurately**. The improvement is in volume measurement, not directly in thermometer sensitivity. The official 1 mark accepts the comparison of greater accuracy than the original measuring cylinder.

(g)(ii) A constant-temperature **water bath** is unsuitable for this aim because it tends to hold the beaker at the set temperature. Heat released or absorbed on dissolving is exchanged with the bath, altering or suppressing the temperature change of interest. The original mark scheme's key point is that the bath keeps the temperature constant. A bath can be useful when temperature is a control variable, but here temperature change is what must be measured.

Worked Example 2 (2024·February/March 0620/62 Q2(d)) · Level 3

A student investigates the reaction between aqueous sodium carbonate and two different solutions of dilute hydrochloric acid, labelled **A** and **B**.

The student does three experiments.

Experiment 1

- Rinse a burette with distilled water and then with dilute hydrochloric acid **A**.
- Rinse a conical flask with distilled water.
- Fill the burette with dilute hydrochloric acid **A**. Run some of the dilute hydrochloric acid out of the burette so that the level of the dilute hydrochloric acid is on the burette scale.
- Record the initial burette reading.
- Use a measuring cylinder to pour 25 cm^3 of aqueous sodium carbonate into the conical flask.
- Add five drops of methyl orange indicator to the conical flask.
- Stand the conical flask on a white tile.
- Slowly add dilute hydrochloric acid **A** from the burette to the conical flask, while swirling the flask, until the solution becomes orange.
- Record the final burette reading.

Experiment 2

- Refill the burette with dilute hydrochloric acid **A**. Run some of the dilute hydrochloric acid out of the burette so that the level of the dilute hydrochloric acid is on the burette scale.
- Record the initial burette reading.
- Empty the conical flask and rinse it with distilled water.
- Use the measuring cylinder to pour 25 cm^3 of aqueous sodium carbonate into the conical flask.
- Add five drops of thymolphthalein indicator to the conical flask.
- Stand the conical flask on a white tile.
- Slowly add dilute hydrochloric acid **A** from the burette to the conical flask, while swirling the flask, until the solution becomes colourless.
- Record the final burette reading.

Experiment 3

- Repeat Experiment 1, using dilute hydrochloric acid **B** instead of dilute hydrochloric acid **A**.

(d) At the start of Experiment 3 the burette is rinsed with distilled water and then with dilute hydrochloric acid **B**.

(i) Identify the substance removed from the burette when it is rinsed with distilled water at the start of Experiment 3.

(ii) Describe how the result of the titration would change if the burette was **not** rinsed with dilute hydrochloric acid **B** after it had been rinsed with water.

(iii) Explain why the conical flask is **not** rinsed with aqueous sodium carbonate after it is rinsed with water.

Solution

Identify which acid is now being introduced. The first two titrations use hydrochloric acid A; experiment 3 changes to hydrochloric acid B. The whole of (d) asks why the burette is first washed with water and then rinsed with B, and why the conical flask is not additionally rinsed with sodium carbonate.

(i) Washing the burette with distilled water before experiment 3 removes **dilute hydrochloric acid A remaining from the previous experiment**. It is not removing B, which has not yet been introduced.

(ii) Without a subsequent rinse with B, water remaining in the burette **dilutes** the newly added acid. The same sodium carbonate sample still requires the same acid moles at the same end-point, but the diluted acid contains fewer moles per cm^3. Therefore a **larger acid volume is needed and the titre increases**. State the direction of the volume change; inaccurate alone does not answer how the result changes.

(iii) Do not rinse the conical flask with sodium carbonate solution because retained solution would **increase the amount of sodium carbonate in the flask** beyond the accurately measured 25 cm^3 sample. The extra carbonate moles affect the acid volume needed for neutralisation. Residual distilled water merely dilutes the measured sample without adding carbonate moles, so the two rinsing liquids have different effects. The official point is the extra sodium carbonate solution introduced. A complete explanation connects that addition to the loss of the specified sample amount.

Practice

1. (2025·May/June 0620/61 Q2(e)) · Level 1

A student investigates how the rate of reaction of magnesium ribbon with dilute acid changes as the concentration of the acid is changed. The student uses five solutions of the same acid, **A**, **B**, **C**, **D**, and **E**. Each solution has a different concentration. The acid is in excess in all experiments.

The student does five experiments.

(e) Explain why repeating each experiment is an improvement.

2. (2025·February/March 0620/62 Q2(g)) · Level 2

A student investigates the temperature change when aqueous copper(II) sulfate reacts separately with two different metals, zinc and iron.

The student does two experiments.

Experiment 1

- Use a 50 cm^3 measuring cylinder to pour 30 cm^3 of aqueous copper(II) sulfate into a 100 cm^3 beaker.
- Use a thermometer to measure the temperature of the contents of the beaker. This is the temperature at time = 0 s.
- Add 3 g of zinc powder to the beaker. At the same time start a stop-watch.
- Using the thermometer, continually stir the mixture in the beaker.
- Measure the temperature of the mixture every 30 seconds for 210 seconds.
- Empty and rinse the beaker with distilled water.

Experiment 2

- Repeat Experiment 1, using 3 g of iron powder instead of 3 g of zinc powder.

(g) Describe **two** changes to the **apparatus** that will improve the results of this investigation. For each change, explain why it will improve the results.

change 1

explanation 1

change 2

explanation 2

3. (2025·May/June 0620/63 Q2(f)) · Level 3

A student investigates how the rate of reaction between dilute hydrochloric acid and aqueous sodium thiosulfate changes as the concentration of the aqueous sodium thiosulfate decreases. During the reaction, the solution slowly becomes cloudy. As the solution becomes cloudy, it becomes more difficult to see through the solution.

The student does five experiments.

Experiment 1

- Use a 50 cm^3 measuring cylinder to pour 50 cm^3 of aqueous sodium thiosulfate into a 100 cm^3 beaker.
- Use a 10 cm^3 measuring cylinder to pour 5 cm^3 of dilute hydrochloric acid into the beaker containing the aqueous sodium thiosulfate.
- Immediately start a stop-watch and stir the contents of the beaker.
- Stand the beaker on a printed sheet and look down from above the beaker as shown in Fig. 2.1.

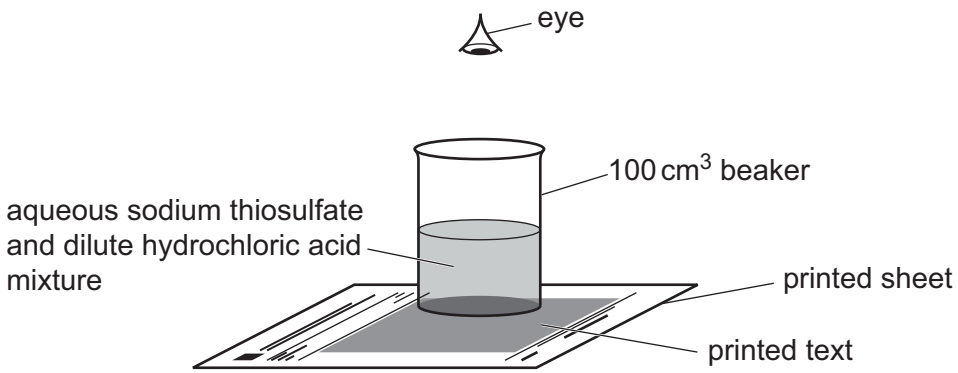


Fig. 2.1

- When the text on the printed sheet is **not** visible, stop the stop-watch and record the time to the nearest whole number of seconds.
- Empty the contents of the beaker and rinse the beaker with distilled water.

(f) Explain why it is important **not** to change the size of the beaker to a larger beaker during the investigation.

Appendix A: Names and Formulae

This appendix groups elements and elemental substances, ions, compounds, structures and materials. Entries within each table are alphabetical. Find the name used in the question, then check its formula on the right. Retain any conditions or qualifications attached to the name.

Subscripts count atoms; superscripts give charge. A subscript outside brackets multiplies the whole group inside. For example, $Al_2(SO_4)_3$ contains 2 Al, 3 S and 12 O atoms. Capitalisation matters: Co is cobalt, while CO is carbon monoxide.

1 Element Names and Symbols

An element symbol can identify the element or, when counting particles, one atom. H, N, O, F and Cl are not the formulae of their usual elemental molecules; see the next table for those molecules.

English name	Formula / structure notes
aluminium	Al
argon	Ar
barium	Ba
beryllium	Be
boron	B
caesium	Cs
calcium	Ca
carbon	C
chlorine	Cl
chromium	Cr
cobalt	Co
copper	Cu
erbium	Er
europium	Eu
fluorine	F
francium	Fr
gold	Au
helium	He
hydrogen	H
iron	Fe
krypton	Kr
lead	Pb
lithium	Li
magnesium	Mg
neon	Ne
nickel	Ni
nitrogen	N
oxygen	O
phosphorus	P
platinum	Pt
potassium	K
rubidium	Rb
silicon	Si

English name	Formula / structure notes
silver	Ag
sodium	Na
sulfur	S
tennessine	Ts
tin	Sn
xenon	Xe
zinc	Zn
zirconium	Zr

2 Elemental Molecules and Structures

The subscript 2 in H_2 means two hydrogen atoms per molecule. Diamond and graphite are both written C, but have different atomic arrangements and properties.

English name	Formula / structure notes
bromine	Br_2
chlorine	Cl_2
diamond	C Like graphite, this is elemental carbon; the atomic arrangement differs.
fluorine	F_2
graphite	C
hydrogen	H_2
iodine	I_2
nitrogen	N_2
oxygen	O_2 Dissolved oxygen is still O_2; do not confuse it with oxide ions, O^{2-}.

3 Cations

A superscript + means the whole ion is positively charged. The 2 in Cu^{2+} gives its charge, not two copper atoms; the subscript 4 in NH_4^+ counts four hydrogen atoms.

English name	Formula / structure notes
aluminium ion	Al^{3+}
ammonium ion	NH_4^+
barium ion	Ba^{2+}
calcium ion	Ca^{2+}
chromium(III) ion	Cr^{3+}
copper(II) ion	Cu^{2+}
hydrogen ion	H^+
iron(II) ion	Fe^{2+}
iron(III) ion	Fe^{3+}
lithium ion	Li^+
potassium ion	K^+
silver ion	Ag^+
sodium ion	Na^+
zinc ion	Zn^{2+}

4 Anions

A superscript – means the whole ion is negatively charged. SO_4^{2-} is one sulfate ion containing one sulfur and four oxygen atoms, with two units of negative charge overall.

English name	Formula / structure notes
bromide ion	Br^-
carbonate ion	CO_3^{2-}
chloride ion	Cl^-
ethanoate ion	CH_3COO^-
hydroxide ion	OH^-
iodide ion	I^-
manganate(VII) ion	MnO_4^-
nitrate ion	NO_3^-
oxide ion	O^{2-}
phosphate ion	PO_4^{3-}
sulfate ion	SO_4^{2-}
sulfide ion	S^{2-}
sulfite ion	SO_3^{2-}

5 Inorganic Compounds

A molecular formula counts atoms in one molecule. An ionic formula gives the composition ratio of ions: NaCl does not represent a separate sodium chloride molecule. Roman numerals such as (II) and (III) distinguish compounds of an element in different oxidation states.

English name	Formula / structure notes
aluminium nitrate	$Al(NO_3)_3$
aluminium oxide	Al_2O_3
aluminium sulfate	$Al_2(SO_4)_3$
ammonia	NH_3 See the aqueous-solutions table for aqueous ammonia.
ammonium chloride	NH_4Cl
ammonium dichromate	$(NH_4)_2Cr_2O_7$
ammonium nitrate	NH_4NO_3
ammonium phosphate	$(NH_4)_3PO_4$
ammonium sulfate	$(NH_4)_2SO_4$
barium chloride	$BaCl_2$
barium hydroxide	$Ba(OH)_2$
barium nitrate	$Ba(NO_3)_2$
barium oxide	BaO
barium sulfate	$BaSO_4$
calcium carbonate	$CaCO_3$
calcium chloride	$CaCl_2$
calcium hydroxide	$Ca(OH)_2$
calcium nitrate	$Ca(NO_3)_2$
calcium oxide	CaO
calcium phosphate	$Ca_3(PO_4)_2$
calcium silicate	$CaSiO_3$
calcium sulfate	$CaSO_4$

English name	Formula / structure notes
carbon dioxide	CO_2
carbon monoxide	CO
chlorine(I) oxide	Cl_2O
chromium(III) bromide	$CrBr_3$
chromium(III) chloride	$CrCl_3$
chromium(III) oxide	Cr_2O_3
cobalt(II) chloride	$CoCl_2$ Composition formula of the salt.
compound in chromite	$FeCr_2O_4$ Composition of the named compound; the whole ore is not a pure substance.
compound in iron pyrite	FeS_2 Composition of the named compound.
copper(II) carbonate	$CuCO_3$
copper(II) chloride	$CuCl_2$
copper(II) hydroxide	$Cu(OH)_2$
copper(II) nitrate	$Cu(NO_3)_2$
copper(II) oxide	CuO
copper(II) sulfate	$CuSO_4$ Salt composition; do not add water of crystallisation when no hydration number is given.
cryolite	Na_3AlF_6
hydrogen chloride	HCl; HCl(g) HCl(g) is hydrogen chloride gas; see the aqueous-solutions table for hydrochloric acid.
hydrogen peroxide	H_2O_2
hydrogen sulfide	H_2S
iodine monobromide	IBr
iodine monochloride	ICl
iron(II) chloride	$FeCl_2$
iron(II) hydroxide	$Fe(OH)_2$
iron(II) iodide	FeI_2
iron(II) sulfate	$FeSO_4$
iron(III) chloride	$FeCl_3$
iron(III) hydroxide	$Fe(OH)_3$
iron(III) nitrate	$Fe(NO_3)_3$
iron(III) oxide	Fe_2O_3 Hydrated rust is not pure anhydrous Fe_2O_3.
lead(II) bromide	$PbBr_2$
lead(II) chloride	$PbCl_2$
lead(II) nitrate	$Pb(NO_3)_2$
lead(II) sulfate	$PbSO_4$
lithium chloride	LiCl Composition formula of the salt.
lithium hydroxide	LiOH
lithium oxide	Li_2O
lithium sulfite	Li_2SO_3
magnesium carbonate	$MgCO_3$
magnesium chloride	$MgCl_2$
magnesium nitrate	$Mg(NO_3)_2$
magnesium oxide	MgO

English name	Formula / structure notes
magnesium sulfate	$MgSO_4$ If a question gives only the salt name, do not infer an unspecified hydration number.
manganese(IV) oxide	MnO_2
nickel(II) nitrate	$Ni(NO_3)_2$
nitric acid	HNO_3
nitrogen dioxide	NO_2
nitrogen monoxide	NO
phosphoric acid	H_3PO_4
phosphorus(V) oxide	P_4O_{10}
potassium bromide	KBr
potassium carbonate	K_2CO_3
potassium chloride	KCl
potassium dichromate(VI)	$K_2Cr_2O_7$
potassium hydroxide	KOH
potassium iodide	KI
potassium manganate(VII)	$KMnO_4$
potassium nitrate	KNO_3
potassium phosphate	K_3PO_4
potassium sulfate	K_2SO_4
potassium sulfide	K_2S
silicon(IV) oxide / silicon dioxide	SiO_2
silver bromide	$AgBr$
silver chloride	$AgCl$
silver iodide	AgI
silver nitrate	$AgNO_3$
sodium bromide	$NaBr$
sodium carbonate	Na_2CO_3
sodium chloride	$NaCl$ Main component of table salt.
sodium hydrogencarbonate	$NaHCO_3$
sodium hydroxide	$NaOH$
sodium nitrate	$NaNO_3$
sodium oxide	Na_2O
sodium sulfate	Na_2SO_4
sodium sulfide	Na_2S
sodium sulfite	Na_2SO_3
sodium thiosulfate	$Na_2S_2O_3$
sulfur dioxide	SO_2
sulfur trioxide	SO_3
sulfuric acid	H_2SO_4
tin(IV) oxide	SnO_2
vanadium(V) oxide	V_2O_5
water / steam	H_2O; $H_2O(g)$ (g) means gaseous water.
zinc carbonate	$ZnCO_3$
zinc hydroxide	$Zn(OH)_2$
zinc oxide	ZnO

English name	Formula / structure notes
zinc sulfate	$ZnSO_4$ Do not add a fixed hydration number to crystals when none is stated.
zinc sulfide	ZnS

6 Aqueous Solutions, Anhydrous Substances and Hydrates

(aq) means dissolved in water; water of crystallisation follows the centred dot. $CuSO_4(aq)$ and $CuSO_4 \cdot 5H_2O$ are not interchangeable. If no hydration number is given, retain an unspecified x or state only the salt composition; do not invent a number.

English name	Formula / structure notes
anhydrous cobalt(II) chloride	$CoCl_2$ No water of crystallisation.
anhydrous copper(II) sulfate	$CuSO_4$ No water of crystallisation; distinguish this from the pentahydrate below.
anhydrous lithium chloride	LiCl The question explicitly states anhydrous.
aqueous ammonia	$NH_3(aq)$ Ammonia dissolved in water; distinguish it from ammonia gas.
aqueous cobalt(II) chloride	$CoCl_2(aq)$
aqueous copper(II) sulfate	$CuSO_4(aq)$ (aq) means dissolved in water, not a number of waters in a crystal.
aqueous nitric acid	$HNO_3(aq)$
aqueous sulfuric acid	$H_2SO_4(aq)$
hydrated aluminium chloride	$AlCl_3 \cdot xH_2O$ The hydration number is unspecified; do not assume x = 6.
hydrated cobalt(II) chloride	$CoCl_2 \cdot 6H_2O$ This formula explicitly contains six waters of crystallisation.
hydrated copper(II) sulfate	$CuSO_4 \cdot 5H_2O$ This formula explicitly contains five waters of crystallisation.
hydrated iron(III) oxide (rust)	$Fe_2O_3 \cdot xH_2O$ x varies; no single hydration number is specified.
hydrochloric acid	HCl(aq) Hydrogen chloride dissolved in water.
limewater / calcium hydroxide solution	$Ca(OH)_2(aq)$
oleum	$H_2S_2O_7$ (representation used in this question) Read the formula as used in the Contact-process question; not every oleum sample is this one pure compound.

7 Organic Names and Formulae

One molecular formula may allow different atom connections. Structural formulae such as CH_3CH_2OH give more connectivity information. If the name or question diagram leaves the structure unspecified, retain that limitation. Entries are alphabetical for lookup from the question.

English name	Formula / structure notes
2-aminopropanoic acid / alanine	$H_2NCH(CH_3)COOH$
but-1-ene	$CH_2=CHCH_2CH_3$
but-2-ene	$CH_3CH=CHCH_3$
butan-1-ol	$CH_3CH_2CH_2CH_2OH$
butan-2-ol	$CH_3CH(OH)CH_2CH_3$
butane	C_4H_{10}

English name	Formula / structure notes
butane-1,4-diamine	$H_2NCH_2CH_2CH_2CH_2NH_2$
butane-1,4-diol	$HOCH_2CH_2CH_2CH_2OH$
butanedioic acid	$HOOCCH_2CH_2COOH$
butanoic acid	$CH_3CH_2CH_2COOH$
butyl butanoate	$CH_3CH_2CH_2COOCH_2CH_2CH_2CH_3$
butyl ethanoate	$CH_3COOCH_2CH_2CH_2CH_3$
calcium ethanoate	$(CH_3COO)_2Ca$
chloroethane	C_2H_5Cl
chloromethane	CH_3Cl
copper(II) ethanoate	$(CH_3COO)_2Cu$
decane	$C_{10}H_{22}$
1,2-dibromoethane	CH_2BrCH_2Br
ethane	C_2H_6
ethane-1,2-diamine	$H_2NCH_2CH_2NH_2$
ethane-1,2-diol	$HOCH_2CH_2OH$
ethanedioic acid	$HOOCCOOH$
ethanoic acid	CH_3COOH
ethanol	C_2H_5OH; CH_3CH_2OH Both representations describe the same substance.
ethene	C_2H_4; $CH_2=CH_2$ Left: molecular formula; right: the carbon–carbon double bond is shown.
ethyl butanoate	$CH_3CH_2CH_2COOCH_2CH_3$
ethyl ethanoate	$CH_3COOCH_2CH_3$
ethyl methanoate	$HCOOCH_2CH_3$
ethyl propanoate	$CH_3CH_2COOCH_2CH_3$
glucose	$C_6H_{12}O_6$
lactic acid	$CH_3CH(OH)COOH$
magnesium ethanoate	$(CH_3COO)_2Mg$
methane	CH_4
methanoic acid	$HCOOH$
methanol	CH_3OH
methyl ethanoate	CH_3COOCH_3
methyl methanoate	$HCOOCH_3$
methyl propanoate	$CH_3CH_2COOCH_3$
2-methylpropane	$CH_3CH(CH_3)CH_3$ Branched C_4H_{10} structure; distinguish it from straight-chain butane.
octane	C_8H_{18}
pentane	C_5H_{12}
prop-2-enoic acid	$CH_2=CHCOOH$
propan-1-ol	$CH_3CH_2CH_2OH$
propan-2-ol	$CH_3CH(OH)CH_3$
propane	C_3H_8
propane-1,3-diol	$HOCH_2CH_2CH_2OH$
propanedioic acid	$HOOCCH_2COOH$
propanoic acid	CH_3CH_2COOH
propene	C_3H_6; $CH_2=CHCH_3$
propyl butanoate	$CH_3CH_2CH_2COOCH_2CH_2CH_3$
sodium ethanoate	CH_3COONa

8 Groups, General Formulae and Structural Scope

Some entries describe part of a structure or a family; others give the composition of a structure in a particular question. Short lines are bonds, R is a variable side group, and [carbon skeleton] stands for an unspecified carbon framework. These are not all unique small-molecule formulae.

English name	Formula / structure notes
amide linkage	—C(=O)—NH— A linkage within a chain, not the molecular formula of a whole polymer.
amino acid	H_2N—CH(R)—COOH NH_2 and COOH attach to the same carbon here; R is a variable side group, not an element.
amino group	—NH_2
an eight-carbon alkene	C_8H_{16} The compound is specified as an alkene; this molecular formula does not locate C=C.
butanol	C_4H_9OH The OH position and full carbon skeleton are unspecified, so no unique structure is implied.
carboxyl group	—COOH
compound B, May/June 2025 Paper 42, Question 6	$C_5H_6O_4$ Molecular formula of the structure in that question only; B is not a general substance name.
diamine	H_2N—[carbon skeleton]—NH_2
dicarboxylic acid	HOOC—[carbon skeleton]—COOH This structure has two carboxyl groups; the carbon skeleton may vary.
diol	HO—[carbon skeleton]—OH The bracketed words represent the remaining carbon skeleton, not element symbols.
ester linkage	—C(=O)—O— =O denotes a double bond between this carbon and oxygen.
ethyl group	—CH_2CH_3 The short line on the left is a bond to the rest of the structure.
five-carbon saturated monohydric alcohol	$C_5H_{12}O$ Atom counts apply within the stated alcohol family; different structures remain possible.
four structures, May/June 2025 Paper 12, Question 10	C_4H_8; C_3H_6O; C_3H_6O; C_3H_6O Listed left to right as in the question; the last three have equal atom counts but different connections.
hydrocarbons	Contain only the elements C and H A class of compounds, with no single shared molecular formula.
methyl group	—CH_3
oxides of nitrogen	NO_x Collective notation for nitrogen oxides, not one substance with a fixed x.
PET (a polyester)	Contains —C(=O)—O— linkages Only the characteristic linkage is shown, not a complete PET repeat unit.
poly(ethene)	[—CH_2—CH_2—]$_n$ Brackets contain the repeat unit; n is the number of repetitions.
poly(propene)	[—CH_2—CH(CH_3)—]$_n$ Not the molecular formula of one fixed small molecule.
polyamide, including nylon and proteins	Repeated amide linkages in the chain Characteristic linkage: —C(=O)—NH—; the rest of the chain can differ.
polyester, including PET	Repeated ester linkages in the chain Characteristic linkage: —C(=O)—O—; there is no single formula for all polyesters.
propanol	C_3H_7OH Neither this name nor this formula specifies the OH position; do not assume propan-1-ol.

9 Materials, Mixtures and Indicators

Air, alloys, ores and many indicators have no single small-molecule formula. The two indicator formulae listed are for reference; prepared indicator solutions also contain solvent and other components.

English name	Formula / structure notes
activated carbon	Mainly C C identifies the carbon component, not a guarantee that every material sample is impurity-free.
air	Gas mixture; no single molecular formula
bauxite	Ore; no single molecular formula The whole ore is not equivalent to pure Al_2O_3.
brass	Alloy; no single molecular formula An alloy of copper and zinc.
chlorophyll	Class name; no single formula listed
coke / coal / petroleum	Materials or mixtures; no single molecular formula Coke is mainly carbon; coal and petroleum contain many components.
enzyme / catalase	No single small-molecule formula Enzymes form a class; catalase names a particular enzyme. One simple formula cannot describe all enzymes.
gasoline / petrol	Mixture; no single molecular formula
ink	Mixture; no single molecular formula
litmus	No single molecular formula listed Contains several components; its use depends on acid–base colour changes.
methyl orange (sodium salt)	$C_{14}H_{14}N_3NaO_3S$ Composition of the sodium salt; a prepared indicator solution is not a pure substance.
stainless steel	Alloy; no single molecular formula
steel	Alloy; no single molecular formula
tap water / river water	Mixture; no single molecular formula Water molecules are H_2O; actual samples may also contain dissolved or suspended substances.
thymolphthalein	$C_{28}H_{30}O_4$ For name and formula reference; acid–base testing still uses its colour change.
universal indicator	Mixture; no single molecular formula

Appendix B: Examination Reference Tables

These are the qualitative-analysis notes supplied with the original examination. Check the reagent, heating, excess addition and observed colour together. Chapter 12 explains the tests and gives worked examples.

Notes for use in qualitative analysis

Tests for anions

anion	test	test result
carbonate, CO_3^{2-}	add dilute acid, then test for carbon dioxide gas	effervescence, carbon dioxide produced
chloride, Cl^- [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
bromide, Br^- [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	cream ppt.
iodide, I^- [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
nitrate, NO_3^- [in solution]	add aqueous sodium hydroxide, then aluminium foil; warm carefully	ammonia produced
sulfate, SO_4^{2-} [in solution]	acidify with dilute nitric acid, then add aqueous barium nitrate	white ppt.
sulfite, SO_3^{2-}	add a small volume of acidified aqueous potassium manganate(VII)	the acidified aqueous potassium manganate(VII) changes colour from purple to colourless

Tests for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium, Al^{3+}	white ppt., soluble in excess, giving a colourless solution	white ppt., insoluble in excess
ammonium, NH_4^+	ammonia produced on warming	–
calcium, Ca^{2+}	white ppt., insoluble in excess	no ppt. or very slight white ppt.
chromium(III), Cr^{3+}	green ppt., soluble in excess	green ppt., insoluble in excess
copper(II), Cu^{2+}	light blue ppt., insoluble in excess	light blue ppt., soluble in excess, giving a dark blue solution
iron(II), Fe^{2+}	green ppt., insoluble in excess, ppt. turns brown near surface on standing	green ppt., insoluble in excess, ppt. turns brown near surface on standing
iron(III), Fe^{3+}	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc, Zn^{2+}	white ppt., soluble in excess, giving a colourless solution	white ppt., soluble in excess, giving a colourless solution

Tests for gases

gas	test and test result
ammonia, NH_3	turns damp red litmus paper blue
carbon dioxide, CO_2	turns limewater milky
chlorine, Cl_2	bleaches damp litmus paper
hydrogen, H_2	‘pops’ with a lighted splint
oxygen, O_2	relights a glowing splint
sulfur dioxide, SO_2	turns acidified aqueous potassium manganate(VII) from purple to colourless

Flame tests for metal ions

metal ion	flame colour
lithium, Li^+	red
sodium, Na^+	yellow
potassium, K^+	lilac
calcium, Ca^{2+}	orange-red
barium, Ba^{2+}	light green
copper(II), Cu^{2+}	blue-green

Source: Cambridge IGCSE Chemistry 0620, May/June 2025, Paper 61, Notes for use in qualitative analysis.

B.2 Periodic Table

The next page keeps the examination periodic table in its original orientation; turn the book sideways to read it. Use relative atomic masses explicitly supplied in a question first; otherwise refer to this table. It also states that 1 mol of gas occupies 24 dm³ at room temperature and pressure.

Source: Cambridge IGCSE Chemistry 0620, May/June 2025, Paper 11.

The Periodic Table of Elements

Group																	
I	II											III	IV	V	VI	VII	VIII
							1 H hydrogen 1										2 He helium 4
3 Li lithium 7	4 Be beryllium 9											5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
11 Na sodium 23	12 Mg magnesium 24											13 Al aluminium 27	14 Si silicon 28	15 P phosphorus 31	16 S sulfur 32	17 Cl chlorine 35.5	18 Ar argon 40
19 K potassium 39	20 Ca calcium 40	21 Sc scandium 45	22 Ti titanium 48	23 V vanadium 51	24 Cr chromium 52	25 Mn manganese 55	26 Fe iron 56	27 Co cobalt 59	28 Ni nickel 59	29 Cu copper 64	30 Zn zinc 65	31 Ga gallium 70	32 Ge germanium 73	33 As arsenic 75	34 Se selenium 79	35 Br bromine 80	36 Kr krypton 84
37 Rb rubidium 85	38 Sr strontium 88	39 Y yttrium 89	40 Zr zirconium 91	41 Nb niobium 93	42 Mo molybdenum 96	43 Tc technetium –	44 Ru ruthenium 101	45 Rh rhodium 103	46 Pd palladium 106	47 Ag silver 108	48 Cd cadmium 112	49 In indium 115	50 Sn tin 119	51 Sb antimony 122	52 Te tellurium 128	53 I iodine 127	54 Xe xenon 131
55 Cs caesium 133	56 Ba barium 137	57–71 lanthanoids	72 Hf hafnium 178	73 Ta tantalum 181	74 W tungsten 184	75 Re rhenium 186	76 Os osmium 190	77 Ir iridium 192	78 Pt platinum 195	79 Au gold 197	80 Hg mercury 201	81 Tl thallium 204	82 Pb lead 207	83 Bi bismuth 209	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –	113 Nh nihonium –	114 Fl flerovium –	115 Mc moscovium –	116 Lv livermorium –	117 Ts tennessine –	118 Og oganesson –

Key

atomic number
atomic symbol
name
relative atomic mass

lanthanoids	57 La lanthanum 139	58 Ce cerium 140	59 Pr praseodymium 141	60 Nd neodymium 144	61 Pm promethium –	62 Sm samarium 150	63 Eu europium 152	64 Gd gadolinium 157	65 Tb terbium 159	66 Dy dysprosium 163	67 Ho holmium 165	68 Er erbium 167	69 Tm thulium 169	70 Yb ytterbium 173	71 Lu lutetium 175
actinoids	89 Ac actinium –	90 Th thorium 232	91 Pa protactinium 231	92 U uranium 238	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure (r.t.p.).

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Measuring cylinder: DBCLS, via Bioicons, licensed under CC BY 4.0. Asset: graded-measuring-cylinder. English labels and explanations of use are added.

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