

Course Companion

AQA · GCSE · Chemistry

Every topic handout in one volume

exercises.lol

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Atomic structure and the periodic table

AQA · GCSE · Chemistry

Atomic structure and the periodic table: evidence and predictions

A filter can separate sand from salt solution, but cannot separate sodium from chlorine in a salt crystal. The difference is chemical combination. This topic connects the particles in a substance to separation, atomic models and patterns of reactions.

Use a supplied periodic table to identify elements. Count atoms before balancing an equation; count protons before identifying an element. Higher Tier ionic/half-equations and Chemistry-only transition-metal content are labelled below.

Elements, compounds and equations (4.1.1.1)

An **atom** is the smallest part of an **element** that can exist. An element contains atoms with one proton number. It may contain single atoms, such as helium, or molecules, such as O₂. O₂ contains two atoms but only one element.

A **compound** contains two or more elements chemically combined in fixed proportions. H₂O always has two hydrogen atoms for each oxygen atom. Compounds have properties different from their elements and can be separated into elements only by chemical reactions. A **chemical reaction** makes one or more new substances and often has a detectable energy change.

Learn names and correctly capitalised symbols: Co means cobalt, while CO is carbon monoxide. Use the first twenty elements, Groups 1 and 7 and other named elements in this course. The first twenty are:

Atomic number	Name and symbol	Atomic number	Name and symbol
1	hydrogen H	11	sodium Na
2	helium He	12	magnesium Mg
3	lithium Li	13	aluminium Al
4	beryllium Be	14	silicon Si
5	boron B	15	phosphorus P
6	carbon C	16	sulfur S
7	nitrogen N	17	chlorine Cl
8	oxygen O	18	argon Ar
9	fluorine F	19	potassium K

Atomic number	Name and symbol	Atomic number	Name and symbol
10	neon Ne	20	calcium Ca

Formulae give composition: MgCl₂ is magnesium chloride; Na₂O is sodium oxide. A **word equation** names reactants and products: magnesium + oxygen → magnesium oxide. A **symbol equation** uses their formulae.

Worked method – balancing: start with Mg + O₂ → MgO. There are two oxygen atoms on the left and one on the right. Put 2 before MgO, then 2 before Mg: **2Mg + O₂ → 2MgO**. Check Mg: 2 = 2; O: 2 = 2. Change whole-particle coefficients, not the subscripts in a substance’s formula. MgO₂ would be a different formula. Three H₂O molecules contain six H atoms and three O atoms.

Higher Tier – charge as well as atoms: zinc forms Zn²⁺ by losing two electrons: Zn → Zn²⁺ + 2e[–]. Both sides have charge zero overall. A copper ion gains two: Cu²⁺ + 2e[–] → Cu. Adding the half-equations cancels electrons: **Zn + Cu²⁺ → Zn²⁺ + Cu**. Sulfate ions present unchanged in solution are spectator ions and are omitted. Check atoms and total charge after cancellation; do not cancel species that change.

Choosing a separation (4.1.1.2)

A **mixture** contains substances not chemically combined. Each retains its chemical properties. Physical separation makes no new substance; choose a method using the components’ properties and the product you need.

Method	Property used and example
filtration	Insoluble particles cannot pass through filter paper; sand stays as residue while water and dissolved salt form filtrate.
crystallisation	Dissolved solid forms crystals when a concentrated solution cools; recover salt from solution.
simple distillation	A liquid evaporates and then condenses; recover water from salt solution. Non-volatile salt remains in the flask.
fractional distillation	Miscible liquids have different boiling points; repeated evaporation/condensation in a column improves separation of ethanol and water.
chromatography	Dissolved substances move differently between solvent and stationary material; separate colours in ink.

Worked method – pure, dry salt from sand, salt and water: stir to dissolve any remaining salt. Filter; sand remains on the paper and salt solution passes through. Heat the filtrate gently to evaporate some water, then cool to form crystals. Filter the

crystals and dry them between filter papers. Washing the sand with a small amount of water can recover salt solution trapped with it. Filtering alone cannot collect dissolved salt.

Distillation recovers a liquid by evaporation followed by condensation.

Worked choice – recover the water instead: use simple distillation. Heat the salt solution; water vapour leaves the flask, cools in a condenser and collects as liquid water. Evaporation without collecting the vapour loses the desired product. Choose fractional distillation for two miscible liquids; a filter cannot separate them. For chromatography, keep the pencil start line above the solvent so the sample moves up the stationary material instead of dissolving directly into the solvent reservoir.

How evidence changed the atomic model (4.1.1.3)

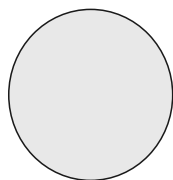
A scientific **model** is a representation that explains evidence. New observations can change or replace it. Early atoms were tiny indivisible spheres. Discovery of the negative electron required internal structure, giving the **plum pudding model**: electrons embedded in a spread-out ball of positive charge.

In alpha scattering, most positive alpha particles passed through a thin metal foil, some were deflected and very few came back. These observations support different conclusions:

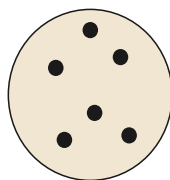
Observation	Nuclear-model conclusion
Most passed straight through	Most of the atom is empty space.
Some were deflected	Positive charge is concentrated and repels the positive alpha particles.
Very few returned towards the source	A tiny, dense central region contains most of the mass.

The **nuclear model** replaced spread-out positive charge with a tiny positive nucleus and surrounding electrons. Bohr then suggested electrons occupy specific distances/energy levels, consistent with observations. Later work identified protons as units of nuclear positive charge. Chadwick’s evidence established neutrons in the nucleus about twenty years after the nuclear model became accepted. Detailed Bohr/Chadwick experimental apparatus is not required.

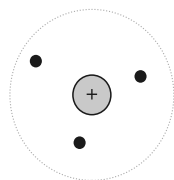
Early model
indivisible sphere



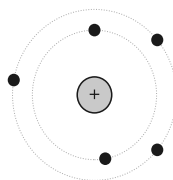
Plum pudding
positive background
with embedded electrons



Nuclear model
mass and positive charge
concentrated in a nucleus



Bohr energy levels
then evidence for
protons and neutrons



Worked explanation: “Some particles bounced back, so the whole atom is solid” contradicts the observation that most passed through. The combined evidence requires a small dense nucleus inside a mostly empty atom. A model must account for the full pattern, not one selected observation.

Counting particles, size and isotopes (4.1.1.4–4.1.1.6)

Particle	Relative mass	Relative charge	Location
proton	1	+1	nucleus
neutron	1	0	nucleus
electron	very small	−1	shells/energy levels

The **atomic number** Z is the proton number and identifies the element. All atoms of that element have the same Z . The **mass number** A is protons + neutrons. A neutral atom has equal proton and electron numbers, so its charges cancel. An ion forms by changing electrons; changing proton number would change the element.

Worked particle count: a chlorine-37 atom has $Z = 17$. Protons = 17; neutrons = $37 - 17 = 20$; electrons = 17. Cl^- has gained one electron, giving 18 electrons while protons/neutrons stay 17/20. For calcium-40, $Z = 20$: Ca^{2+} has 20 protons, 20 neutrons and 18 electrons.

Isotopes are atoms of the same element with different neutron numbers. Chlorine-35 and chlorine-37 both have 17 protons, but 18 and 20 neutrons respectively. They remain chlorine; different mass number does not mean different element.

Typical atom radius is about **0.1 nm** = 1×10^{-10} m; 1 nm = 10^{-9} m. Nucleus radius is less than one ten-thousandth of the atom’s radius, of order 10^{-14} m. Almost all mass is in the nucleus. These are approximate scales, not exact dimensions of every atom.

Worked scale: if an atom’s radius were enlarged to 1 m, a nucleus one ten-thousandth as large would have radius $1/10\,000$ m = 0.0001 m = 0.1 mm. The actual specified nucleus is smaller than that ratio. Compare radius with radius; an atom’s diameter is twice its radius. Atoms of diameter about 2×10^{-10} m would number about 5×10^6

along a 1 mm line, ignoring gaps: $(1 \times 10^{-3})/(2 \times 10^{-10})$.

Relative atomic mass

A_r is an abundance-weighted mean, with no unit. It need not be a whole number. Use the percentage abundances, not an unweighted average:

$$A_r = \frac{\sum(\text{isotope mass} \times \text{percentage abundance})}{100}$$

Worked mean: 75% chlorine-35 and 25% chlorine-37 give $(35 \times 75 + 37 \times 25)/100 = (2625 + 925)/100 = \mathbf{35.5}$. This is closer to 35 because that isotope is more abundant; it lies between 35 and 37. No single atom has “35.5 neutrons” or a mass number of 35.5.

Electron arrangements and periodic position (4.1.1.7–4.1.2.1)

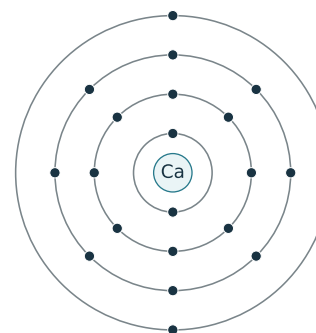
An outer shell electron

is in the highest occupied shell of an atom and helps determine its reactions.

Electrons occupy the lowest available **energy levels** first. For the first twenty neutral atoms, use the GCSE shell model: first shell up to 2, second up to 8, then 8 before the fourth begins. This is the required limited model, not a universal rule for all elements.

Worked arrangement: calcium has $Z = 20$ and therefore 20 electrons. Place 2 in the first shell, 8 in the second, 8 in the third and 2 in the fourth: **2,8,8,2**. A diagram must show those four shells and the same total number of electrons. Chlorine is **2,8,7**, with three occupied shells and seven outer electrons.

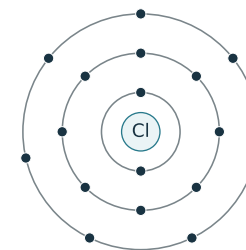
Calcium: 20 electrons



2,8,8,2

Neutral atoms • dots are electrons • shell sizes are schematic, not to scale

Chlorine: 17 electrons



2,8,7

The modern periodic table

orders elements by proton number. A

period

is a row; the number of occupied shells gives the period. A **group** is a column containing elements with related chemical properties. For main Groups 1–7, the group number gives the outer-electron number. Group 0 has a full outer shell:

helium has two electrons, while the other noble gases in this model have eight. This exception prevents treating every group member as having identical electron counts.

Worked prediction: an unknown first-twenty element has 2,8,1. It has 11 electrons/protons, so it is sodium, in Period 3 and Group 1. Its single outer electron helps predict formation of a +1 ion and reactions similar to lithium/potassium. An element with 2,8,7 instead tends to gain one electron in ionic reactions. Position supports predictions; it does not replace reaction evidence.

Developing the periodic table; metals and non-metals (4.1.2.2–4.1.2.3)

Early tables used increasing atomic weight before subatomic particles were known. They were incomplete and strict mass ordering put some elements with unlike neighbours. **Mendeleev** left gaps for undiscovered elements and sometimes changed mass order to keep similar properties together. He predicted properties of missing elements; later discoveries matching those predictions supported his arrangement. Isotope knowledge explained why average atomic-weight order did not always match element identity. Modern ordering uses proton number.

Worked evidence judgement: finding an element with the predicted properties in a gap supports the table more strongly than merely finding any new element. A failed prediction would prompt review of the arrangement. Leaving gaps was a testable proposal, not proof by itself.

Metals, mainly towards the left/bottom, form positive ions in reactions by losing electrons. Non-metals, mainly towards the right/top, do not characteristically form positive ions; they may gain electrons or share them in covalent bonds. Most elements are metals. Metals typically conduct heat/electricity, are shiny and can be bent or shaped; solid non-metals are often brittle and poor conductors. These are characteristic patterns with exceptions, such as conducting graphite, not definitions without exceptions.

Group 0, Group 1 and Group 7 (4.1.2.4–4.1.2.6)

Noble gases are unreactive and do not easily form molecules because their atoms have stable full outer shells. They are monatomic. Their boiling points rise down Group 0 as relative atomic mass increases. Use supplied trend data to predict an intermediate value or whether a later member boils higher; do not claim exact values without data.

Alkali metals in Group 1 have one outer electron and form +1 ions. Lithium, sodium and potassium react with oxygen to form oxides, with chlorine to form chlorides, and with water to form a metal hydroxide and hydrogen. In the school model, $4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O}$; $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$; $2\text{K} + 2\text{H}_2\text{O} \rightarrow 2\text{KOH} + \text{H}_2$. Do not assume every oxygen product under every condition has the same formula; use supplied

formulas when balancing.

With water, lithium fizzes and moves slowly; sodium reacts faster and often melts into a moving ball; potassium reacts still more vigorously and may ignite with a lilac flame. These are observations of supervised demonstrations. Reactivity rises down the group: more occupied shells place the outer electron farther from the nucleus with more shielding, so it is lost more easily despite greater nuclear charge.

Halogens in Group 7 have seven outer electrons and form diatomic molecules: Cl_2 , Br_2 and I_2 . At room temperature chlorine is a pale-green gas, bromine a red-brown liquid and iodine a grey-black solid. Relative molecular mass, melting point and boiling point rise down the group. Reactivity falls: the outer region is farther from the nucleus and more shielded, so attraction for an incoming electron is weaker.

Halogens form ionic halides with metals, such as NaCl , and covalent compounds with non-metals, such as HCl . Do not say all halogen compounds contain ionic bonds.

Displacement occurs when a more reactive halogen replaces a less reactive one in an aqueous halide solution.

Worked displacement: chlorine added to potassium bromide produces bromine, giving an orange/brown solution: $\text{Cl}_2 + 2\text{KBr} \rightarrow 2\text{KCl} + \text{Br}_2$. Chlorine gains electrons from bromide ions. Bromine does not displace chloride ions because it is less reactive. A “no reaction” result can therefore help establish a reactivity order. For Higher Tier: $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$; $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$; net $\text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2$. Atoms and charge are conserved.

Group 0	Group 1	Group 7
Not readily reactive	Reactivity rises ↓	Reactivity falls ↓
Full outer shell Boiling point rises ↓	One outer electron Lost more easily ↓	Seven outer electrons Gain less easily ↓

Down a group, the outer occupied shell is farther from the nucleus.

Transition metals – Chemistry only (4.1.3.1–4.1.3.2)

Use chromium, manganese, iron, cobalt, nickel and copper as examples of **transition metals**. Compared with Group 1 metals, they typically have higher melting points and densities, greater strength/hardness, and lower reactivity with oxygen, water and halogens. Copper does not react with cold water as sodium does. Compare named examples and conditions rather than assuming all metals behave identically.

Many transition elements form ions with different charges and coloured compounds,

and act as **catalysts** . Iron forms Fe^{2+} and Fe^{3+} ; aqueous iron(II) salts can be pale green and iron(III) salts yellow/brown. Aqueous copper(II) salts are often blue. Colour depends on the compound/conditions: not every compound of an element has one fixed colour. Iron is the Haber-process catalyst; nickel catalyses hydrogenation. These examples illustrate typical properties, not properties of every possible transition compound.

Worked formula link: sulfate is SO_4^{2-} . Fe^{2+} gives FeSO_4 ; Fe^{3+} requires two iron ions and three sulfate ions, giving $\text{Fe}_2(\text{SO}_4)_3$. The charges balance to zero in both compounds. Different proportions/iron charges mean different compounds despite the same three element names.

Check yourself

- Choose a separation from component properties and the product required; explain what happens at each stage.
- Link alpha-scattering observations to the nuclear model, rather than memorising a timeline alone.
- Count particles in atoms and ions; calculate a weighted mean; distinguish radius, diameter and scale.
- Represent first-twenty electron arrangements as both numbers and shell diagrams, then predict periodic position.
- Use evidence to explain group trends and displacement; compare named transition-metal properties with Group 1.

2 · Bonding, structure and the properties of matter

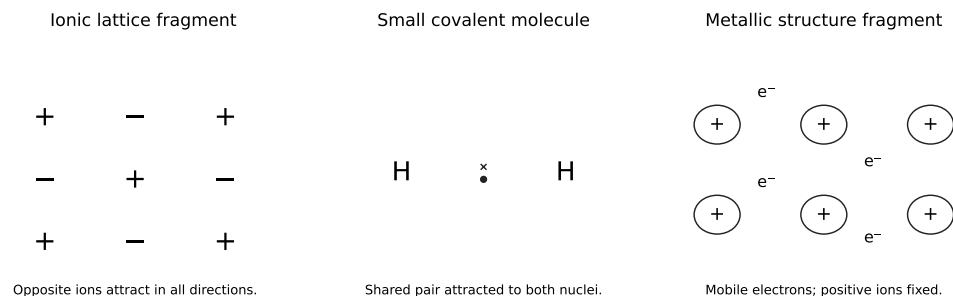
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Bonding, structure and the properties of matter

- Three types of strong chemical bond hold atoms or ions together: ionic (attraction between oppositely charged ions), covalent (shared electron pairs) and metallic (delocalised electrons).
- Structure follows bonding: giant **ionic lattices** , small molecules, polymers, giant covalent structure networks and metal lattices – each predicts melting point and conductivity.
- Carbon shows every trick: **diamond** , **graphite** , **graphene** , fullerenes; nanoparticles (chemistry only) turn **surface area** into properties.

Ionic, covalent and metallic bonding (4.2.1)

Three strong chemical bonds, all explained by **electrostatic forces** :



- Ionic** – the bond is electrostatic attraction between oppositely charged ions. During formation, a metal atom **transfers** outer electrons to a non-metal: metal atoms lose electrons to become **positive ions**; non-metal atoms gain them to become **negative ions**; Groups 1/2 metals and 6/7 non-metals form ions with **noble-gas electronic structures**. Draw the electron transfer with **dot and cross diagrams**; deduce ion charges from group number (Group 1 $\rightarrow +1$, Group 2 $\rightarrow +2$, Group 6 $\rightarrow -2$, Group 7 $\rightarrow -1$).
- Covalent** – non-metal atoms **share pairs of electrons**. Know dot-and-cross diagrams for **H₂, Cl₂, O₂, N₂, HCl, H₂O, NH₃, CH₄**, and line representations for small molecules, polymer repeating units and giant covalent structures.

- **Metallic** – a giant regular structure of positive metal ions with **delocalised outer electrons** free to move through the whole structure; the sharing of these electrons gives strong metallic bonding.

Know the **limitations of models**: dot-and-cross, ball-and-stick, 2D and 3D diagrams all simplify – forces are not drawn, sticks suggest gaps between touching particles, a flat drawing does not fully show 3D arrangement, and a giant structure is represented by only a fragment. Deduce empirical formulae from lattice models; molecular formulae from molecule diagrams.

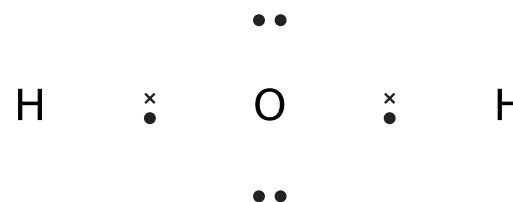
■ Build an electron diagram, then check it

For a small molecule, add the outer electrons from every atom. Choose shared pairs so hydrogen has two electrons around it and the other named atoms have eight. Add the remaining electrons as **lone pairs**. Dots and crosses record where electrons came from; the electrons themselves are identical. A single bond shares two electrons, a double bond four, a triple bond six.

Molecule	Total outer electrons	Shared pairs	Lone pairs remaining
H ₂	2	1	0
Cl ₂	14	1	3 on each Cl
O ₂	12	2	2 on each O
N ₂	10	3	1 on each N
HCl	8	1	3 on Cl
H ₂ O	8	2	2 on O
NH ₃	8	3	1 on N
CH ₄	8	4	0

Worked method: water. Oxygen supplies six outer electrons and two hydrogen atoms supply one each: total eight. Two shared O–H pairs use four electrons. The other four form two lone pairs on oxygen. Every hydrogen has two electrons around it; oxygen has eight. The flat electron drawing does not establish the molecule's 3D shape.

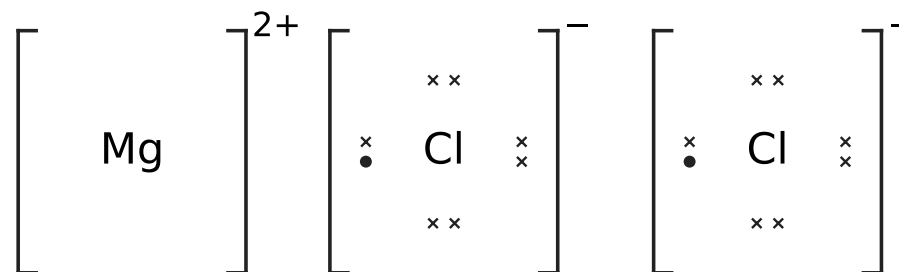
water: dots from O; crosses from H



Flat electron model: positions do not establish the 3D shape.

Worked method: magnesium chloride. Mg starts 2,8,2 and loses two electrons, becoming Mg²⁺ with configuration 2,8. Two chlorine atoms each start 2,8,7 and gain one electron, becoming two Cl⁻ ions with configuration 2,8,8. Draw brackets and charges around every ion. Each chloride has seven original electrons and one gained electron. An outer-electron diagram may leave Mg's former third shell empty; this does not mean the ion contains no electrons.

MgCl₂: one Mg²⁺ and two Cl⁻ ions



Each chloride: seven crosses + one magnesium dot.
Mg²⁺ has configuration 2,8; its old third shell is empty.

The formula follows charge balance: one Mg²⁺ needs two Cl⁻ ions, so MgCl₂. A lattice fragment with 6 positive ions for every 12 negative ions has simplest ratio 1:2. Do not read Mg₆Cl₁₂ as the formula of an individual molecule: the ionic structure is giant. A finite edge of a drawing can also distort apparent counts; use the repeating pattern and charges.

A **line representation** replaces each shared pair with one line. Water is H–O–H, oxygen is O=O and nitrogen N≡N. Lone pairs are omitted in these line formulae. In a polymer, a bracketed repeat unit represents a small section of a very long covalently bonded chain; continuation bonds cross the brackets. In a giant covalent diagram, bonds continue beyond the fragment in a network, rather than stopping at an individual molecule.

Bonding, structure and properties (4.2.2)

States of matter: solid, liquid, gas; melting/freezing at the **melting point**, boiling/condensing at the **boiling point**. Particle theory (small solid spheres) explains the changes; the stronger the forces between particles, the higher the melting/boiling points. (HT) The simple model's limits: no forces shown, all spheres, solid particles. **State symbols:** (s), (l), (g), (aq).

Structure	Bonding/forces	Melting point	Conducts electricity?
ionic lattice	strong electrostatic forces in all directions	high	only when molten or dissolved – ions free to move
small molecules	strong covalent bonds inside , weak intermolecular forces between	low – only weak intermolecular forces are overcome	no – neutral molecules have no mobile ions or delocalised electrons
polymers	strong covalent chains, stronger intermolecular forces	solid at room temperature	no
giant covalent	every atom covalently bonded	very high	generally no; graphite and graphene have mobile delocalised electrons
metals	metallic bonding – delocalised electrons	mostly high	yes – delocalised electrons carry charge (and heat)

Larger molecules → stronger intermolecular forces → higher melting/boiling points. Alloys are harder than pure metals because different-sized atoms **distort the layers**, stopping them sliding.

■ Predict a state and explain a change

At the stated pressure, a pure substance is solid below its melting point, liquid between melting and boiling points, and gas above its boiling point. At a transition temperature, two states can coexist. **Worked example:** X melts at −15 °C and boils at 65 °C. It is solid at −30 °C, liquid at 20 °C and gas at 90 °C. A temperature number alone is not enough without the material's transition data.

In a solid, particles vibrate around fixed positions. During melting they gain enough energy to move past one another but remain close together. During boiling, particles separate much further as attractions are overcome. Cooling reverses these changes and transfers energy to the surroundings. The particles do not swell into larger atoms on heating. An individual atom does not have the melting point, hardness or conductivity of a bulk sample: those properties depend on a collection of particles and their structure.

Use a causal explanation: **structure** → **force or charge carrier** → **property**. For solid NaCl, oppositely charged ions attract strongly in all directions, so melting needs much energy. Its ions are fixed, so it does not conduct. Molten NaCl retains charged ions, but they move and carry charge. For a metal, mobile delocalised electrons carry charge in the solid; for a small molecular liquid, moving neutral molecules alone do not carry an electrical current.

Worked comparison: a polymer chain has strong covalent bonds along it. Relatively strong intermolecular attractions between the large chains make common polymers solid at room temperature. Softening the polymer does not require breaking every covalent bond along every chain. In a giant covalent network such as silica, strong covalent bonds extend throughout the structure, so many bonds must be overcome to melt it. These are different structures even though both contain covalent bonds.

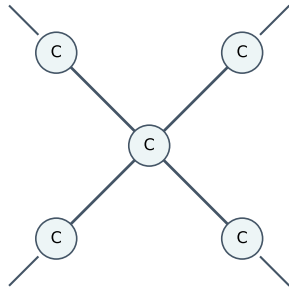
Pure metal layers can slide as the metallic attraction remains. Different-sized atoms in an alloy distort the regular layers, making sliding harder. This explains increased hardness; it does not mean that all pure metals are easily scratched or that an alloy has no movable electrons.

Structure and bonding of carbon (4.2.3)

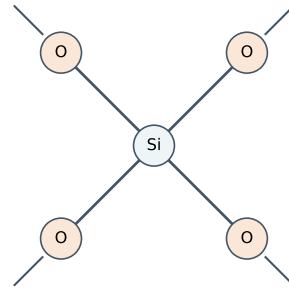
- **Diamond** – a giant covalent structure in which each carbon forms **four** covalent bonds: **very hard, very high melting point, does not conduct**.
- **Graphite** – each carbon forms **three** bonds, layers of hexagonal rings with **no covalent bonds between layers** (soft, slippery – lubricant); **one delocalised electron per atom** → conducts electricity like a metal.
- **Graphene** – a **single layer of graphite**: one atom thick, strong, conducts – electronics and composites.

- **Fullerenes** – hollow molecules of carbon hexagons (plus 5- or 7-membered rings); **Buckminsterfullerene C₆₀** is spherical. **Carbon nanotubes** – cylindrical fullerenes with huge length-to-diameter ratios: nanotechnology, electronics, materials.

Local coordination projections; bonds continue into giant 3D structures



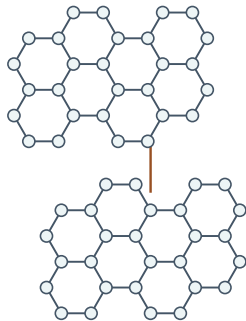
Diamond: four C neighbours



Silica: four O neighbours; each O bridges to another Si

2D projection does not show actual bond angles. Silica oxygens are shared: SiO₂ overall.

Graphite: hexagonal covalent layers, with weaker attraction between layers



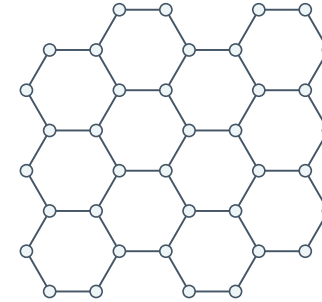
Strong covalent bonds within layers

One delocalised electron per carbon

No covalent bonds between layers

Each interior carbon has three neighbours. Cut edges omit further bonds.

Graphene: one hexagonal carbon layer

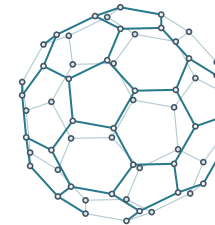


Strong covalent connections

Delocalised electrons

Carbon atoms shown at vertices; cut edges continue. Single-layer schematic.

C₆₀ carbon-bond projection and schematic nanotube shape



C₆₀: spherical cage
60 carbon atoms

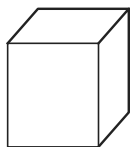


Nanotube: long hollow cylinder
large length/diameter ratio

C₆₀: 60 vertices, 90 bonds; rear edges pale. Nanotube outline omits its carbon-ring bonds.

Bulk and surface properties incl. nanoparticles – chemistry only (4.2.4)

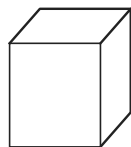
Nanoscience = structures **1–100 nm**; the smallest are of the order of a few hundred atoms. Particle size and atom count are different quantities. Nanoparticles are smaller than **fine particles** (PM_{2.5}, 100–2500 nm), which are smaller than **coarse particles/dust** (PM₁₀ in this specification, 2.5×10^{-6} – 1×10^{-5} m).



side = 100 nm

$$A/V = 0.06 \text{ nm}^{-1}$$

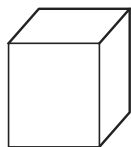
Schematic cube; not drawn to scale.



side = 10 nm

$$A/V = 0.6 \text{ nm}^{-1}$$

Schematic cube; not drawn to scale.



side = 1 nm

$$A/V = 6 \text{ nm}^{-1}$$

Schematic cube; not drawn to scale.

Surface area : volume: as a cube's side **decreases 10×**, SA:V **increases 10×**. Nanoparticles may have different properties from the bulk material because of this high ratio – catalysts (smaller quantities work), medicine delivery, cosmetics, electronics. Concerns: effects inside the body and in the environment are not fully known.

■ Compare equal amounts, not just one particle

For a model cube of side l , surface area $A = 6l^2$ and **volume** $V = l^3$, so $A/V = 6/l$. The quotient has an inverse-length unit. **Worked example:** a 100 nm cube has $A = 60\,000 \text{ nm}^2$, $V = 1\,000\,000 \text{ nm}^3$ and $A/V = 0.06 \text{ nm}^{-1}$. A 10 nm cube has $A = 600 \text{ nm}^2$, $V = 1000 \text{ nm}^3$ and $A/V = 0.6 \text{ nm}^{-1}$: ten times larger. The smaller individual cube has less area, but the same total volume divided into smaller cubes has more total exposed area.

One nanometre is 10^{-9} m. A 20 nm particle is 2.0×10^{-8} m across, about 200 times a typical 0.1 nm atom diameter. It is not one large atom. A 0.2 nm cube would lie below the specified nanoparticle range, so it is not a suitable nanoparticle example.

Evaluate given information. Suppose equal masses of a material remove the same pollutant in 12 minutes as nanoparticles and 40 minutes in bulk form. The tested nano form acts faster; greater exposed area per mass is a plausible explanation. This result alone does not prove that a skin product is safe, cheap or environmentally harmless. Weigh the stated benefit against exposure, possible health/environmental effects and the limits of the evidence. Ask what further measurements would resolve the particular uncertainty. Do not claim that all nanoparticles enter the skin or that all are harmful.

Checklist before you call this topic done

- The three bonds in terms of electrons and electrostatic forces; ion charges from group numbers.
- Dot-and-cross for the eight named molecules; model limitations named.
- The properties table above rebuilt from memory – which forces are overcome on melting.

- Diamond vs graphite vs graphene vs fullerenes, bonding to property; apply this reasoning to an unfamiliar use.
- (Chem) SA:V factor-of-10 rule; the three particle-size bands and nanoparticle uses and concerns.

3 · Quantitative chemistry

AQA · GCSE · Chemistry

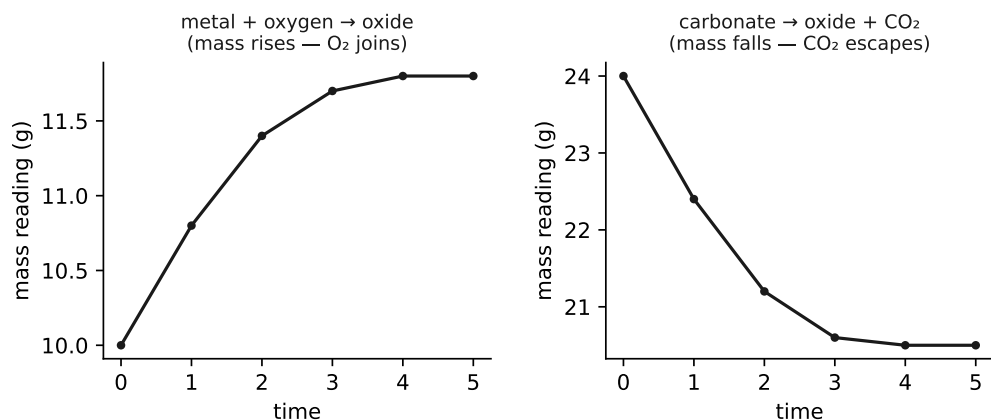
Quantitative chemistry: counting atoms by mass

- Atoms are neither lost nor made: the balanced equation is a mass ledger, and conservation of mass closes every account.
- The mole converts grams into particle counts; moles turn equations into reacting-mass arithmetic.
- Yield, atom economy, molar concentration and gas volumes (chemistry only, with the specified HT demands) finish the quantitative toolkit.

Conservation of mass and equations (4.3.1)

Conservation of mass : no atoms are lost or made, so the mass of products equals the mass of reactants. Symbol equations are **balanced** in atom numbers; know the difference between a **multiplier before a formula** (numbers of units) and a **subscript within it** (atoms in the unit).

Relative formula mass (Mr) = sum of the relative atomic masses in the numbers shown. In a balanced equation, multiply every relative formula mass by its equation coefficient before summing: the totals for reactants and products are equal. **Percentage by mass of an element** = $(Ar \times \text{number of atoms} \div Mr) \times 100\%$.



Mass changes with gases: metal + oxygen → oxide **gains** mass; thermal decomposition of a carbonate **loses** the escaped CO₂ – no law is broken once the gas is counted.

Uncertainty: every measurement carries uncertainty; use the **range of repeated measurements about the mean** to estimate the spread. A mean ± half-range

report is a stated classroom convention, not a rule that automatically includes every source of measurement error.

■ Count atoms before comparing masses

For $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$, the left contains two Mg atoms and two O atoms; so does the right. With Mg = 24 and O = 16, the relative mass total is $2(24) + 32 = 80$ on the left and $2(40) = 80$ on the right. The unweighted calculation $24 + 32 = 40$ is false. Changing the coefficient changes how many formula units react; changing a subscript changes the substance.

Worked formula calculation. For $(\text{NH}_4)_2\text{SO}_4$, count $\text{N}_2\text{H}_8\text{SO}_4$: $M_r = 2(14) + 8(1) + 32 + 4(16) = 132$. Sulfur's mass fraction is $32/132$, so its percentage by mass is $32/132 \times 100 = 24.2\%$. Use the mass of the chosen element throughout the whole formula, including bracket multipliers. Percentages for all elements add to approximately 100%, allowing rounding.

Worked gas accounting. Heating 10.0 g CaCO₃ produces 5.6 g CaO while 4.4 g CO₂ escapes. The balance under the open vessel falls by 4.4 g; the full reaction still accounts for $5.6 + 4.4 = 10.0$ g. During oxidation, oxygen enters from outside the weighed metal sample, so the solid can gain mass. Explain which particles crossed the system boundary. Never seal a heated gas-producing reaction to make a classroom mass demonstration: use an approved apparatus and account for the gas.

Worked uncertainty report. Repeated masses are 2.40, 2.42 and 2.44 g. Mean = $(2.40 + 2.42 + 2.44)/3 = 2.42$ g; range = $2.44 - 2.40 = 0.04$ g. Under the stated half-range convention report 2.42 ± 0.02 g. Plot individual readings on a number line to show their distribution. A close cluster does not prove the balance is correctly calibrated. Retain unusual readings unless there is an evidenced reason to exclude them; report any exclusion and reason.

Moles and reacting masses – HT (4.3.2)

The mole : the mass of one mole of a substance in grams is numerically its **Mr**. One mole contains the **Avogadro constant** of particles – 6.02×10^{23} – the same count of stated particles (atoms, molecules, ions) as a mole of any other substance.

Use the equation and its units directly. $n = m/M$; rearranging gives $m = nM$ and $M = m/n$, where molar mass M has units g/mol and is numerically equal to the relative formula mass. Relative formula mass itself has no unit.

$$\text{moles} = \frac{\text{mass (g)}}{M_r}$$

Equations as mole ratios: $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$ reads “1 mol Mg reacts with 2 mol HCl → 1 mol MgCl₂ + 1 mol H₂”. Given any one mass, calculate all others: mass → moles → ratio → moles → mass.

Balancing from masses: convert each mass to moles, divide by the smallest, clear fractions to whole numbers.

Limiting reactant : the reactant completely used up limits the product; an **excess** of the other ensures completion. Calculate the product from the limiting reactant's moles only.

Concentration : in g/dm³ for all tiers – mass of solute in a given volume; (HT) also **mol/dm³**: moles = concentration × volume(dm³), rearranged as needed.

■ Choose the particle and the equation ratio

Worked mole conversion (HT). For 8.0 g NaOH, $M = 23 + 16 + 1 = 40$ g/mol, so $n = 8.0 \text{ g}/(40 \text{ g/mol}) = 0.20$ mol. Formula-unit count = $0.20 \times 6.02 \times 10^{23} = 1.20 \times 10^{23}$. Each formula unit contains one O and one H, so it contains twice as many O-plus-H atoms as formula units. Always name what is counted: a mole of O₂ molecules contains two moles of oxygen atoms. The Avogadro constant applies to stated ions and electrons as well as atoms and molecules.

Worked reacting mass (HT). Find the CO₂ from complete decomposition of 10.0 g CaCO₃. $M(\text{CaCO}_3) = 100$ g/mol, so $n = 10.0/100 = 0.100$ mol. The equation $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ has a 1:1 ratio, giving 0.100 mol CO₂. Its molar mass is 44 g/mol, so $m = 0.100 \times 44 = 4.40$ g. Reverse the chain if the product mass is supplied: product mass → product moles → reactant moles → reactant mass. Mass ratios are not generally the same as mole ratios.

Worked balancing from masses (HT). Suppose 4.8 g Mg combines with 3.2 g O₂ to form 8.0 g MgO. Amounts are $4.8/24 = 0.20$ mol Mg, $3.2/32 = 0.10$ mol O₂ and $8.0/40 = 0.20$ mol MgO. Divide all by 0.10 to obtain 2:1:2. Write $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ and verify atoms and mass. Use the formula of the actual oxygen reactant, O₂; dividing its mass by 16 would count oxygen atoms instead of oxygen molecules.

Worked limiting reactant (HT). Mix 4.8 g Mg with 6.4 g O₂. Amounts are 0.20 mol Mg and 0.20 mol O₂. Compare amount divided by coefficient: Mg $0.20/2 = 0.10$; O₂ $0.20/1 = 0.20$. Magnesium runs out first. It makes 0.20 mol MgO, mass = $0.20 \times 40 = 8.0$ g. Oxygen used = $0.20/2 = 0.10$ mol, leaving 0.10 mol = 3.2 g. Check total mass: $4.8 + 6.4 = 8.0 + 3.2 = 11.2$ g. More oxygen cannot increase the product once magnesium is exhausted. Exactly stoichiometric amounts can use up both reactants together; do not force a single limiting label in that case.

Worked mass concentration (both tiers). A 250 cm³ solution contains 3.5 g solute. Convert volume: $250/1000 = 0.250$ dm³. Concentration = $3.5/0.250 = 14$ g/dm³. A 100 cm³ portion contains $m = cV = 14 \times 0.100 = 1.4$ g solute. Volume means the final solution volume, not automatically the volume of water added. At fixed solute mass, doubling volume halves concentration; at fixed volume, doubling solute mass doubles concentration (HT explanation).

Yield and atom economy – chemistry only (4.3.3)

Percentage yield is below 100 % because: the reaction is **reversible**; product is **lost on separation**; reactants react in **unwanted ways**.

$$\% \text{ yield} = \frac{\text{mass actually made}}{\text{maximum theoretical mass}} \times 100$$

Atom economy measures how much of the starting material ends up in the **useful product** – high atom economy matters for sustainability and cost:

$$\text{atom economy} = \frac{\text{coefficient} \times M_r \text{ of desired product}}{\text{sum of coefficient} \times M_r \text{ for all reactants}} \times 100 \%$$

(HT) calculate the theoretical mass from the balanced equation, then the yield.

■ Separate actual yield from equation efficiency

Worked yield. If the theoretical product mass is 8.0 g and the dry recovered product is 6.0 g, yield = $6.0/8.0 \times 100 = 75\%$. A yield above 100% suggests the weighed material is wet, contaminated or affected by measurement/calculation error; it does not show extra atoms were created. When theoretical mass is not supplied, first calculate it from the balanced equation and limiting amount (HT), then compare the same product's actual and theoretical masses.

Worked atom economy. For $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ with CaO desired, atom economy = $56/100 \times 100 = 56\%$. If CO₂ were instead the desired product, it would be 44%. For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, desired-product relative mass is $2 \times 18 = 36$, and total reactant relative mass = $2 \times 2 + 32 = 36$, so atom economy is 100%. Count every coefficient; using only one H₂O would falsely give 50%. These values come from the equation, whereas actual yield comes from the experiment. A 100% atom-economy reaction can still give a low isolated yield.

Choose a route (HT). A fictional route A has 90% atom economy but 40% yield and requires prolonged heating. Route B has 75% atom economy, 95% yield and a useful saleable by-product. A judgement needs more than picking the largest first percentage: compare useful output, energy, speed, equilibrium/reversibility, waste treatment and by-product demand under the supplied conditions. State the tradeoff and missing evidence, rather than declaring one number proves sustainability.

Concentrations in mol/dm³ – chemistry only, HT (4.3.4)

Concentration in **mol/dm³** links moles, mass and volume: moles = $C \times V$; from a titration, knowing the volumes of both solutions and one concentration gives the other – use the balanced equation to relate acid and alkali amounts. Equal volumes do not imply equal amounts, and acid moles equal alkali moles only for a 1:1 equation ratio.

■ Concentration is amount per final solution volume

Use $n = cV$ with V in dm^3 for c in mol/dm^3 . Rearranging gives $c = n/V$ or $V = n/c$. A solution at $0.20 \text{ mol}/\text{dm}^3$ contains 0.20 mol per dm^3 ; it does not contain 0.20 mol in every beaker regardless of volume. To convert molar concentration to mass concentration, multiply by molar mass: $0.20 \text{ mol}/\text{dm}^3 \text{ NaOH} \times 40 \text{ g}/\text{mol} = 8.0 \text{ g}/\text{dm}^3$.

Worked non-1:1 titration. $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$. A 25.0 cm^3 acid sample reacts completely with 20.0 cm^3 of $0.100 \text{ mol}/\text{dm}^3 \text{ NaOH}$. Alkali amount $= 0.100 \times 0.0200 = 0.00200 \text{ mol}$. Acid amount $= 0.00200/2 = 0.00100 \text{ mol}$. Acid concentration $= 0.00100/0.0250 = 0.0400 \text{ mol}/\text{dm}^3$. The volumes are both known, but the equation determines the mole ratio. Check by multiplying acid amount by two to recover the required alkali amount. Titration apparatus and method are developed in Chemical changes; this example interprets completed-reaction data.

Gas volumes – chemistry only, HT (4.3.5)

A given volume of gas contains the same number of moles at the same temperature and pressure – **equal volumes = equal moles**. The volumes of reacting gases (and products) follow the equation's ratio directly, e.g. 2 volumes of hydrogen react with 1 volume of oxygen.

■ Use matching temperature, pressure and gas units

At room temperature and pressure (20°C and one atmosphere), one mole of gas occupies 24 dm^3 in this specification's approximation. Thus $V = n \times 24 \text{ dm}^3$ and $n = V/24$ when V is in dm^3 . Convert cm^3 to dm^3 first. This value does not apply unchanged at any temperature or pressure, and a volume ratio applies to gases measured under the same conditions.

Worked mass to gas volume. For 4.4 g CO_2 at room conditions, $n = 4.4/44 = 0.100 \text{ mol}$, so $V = 0.100 \times 24 = 2.40 \text{ dm}^3 = 2400 \text{ cm}^3$. The reverse is $m = (V/24)M$ with consistent volume units.

Worked gas reaction. $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$. At identical temperature and pressure, $60 \text{ cm}^3 \text{ N}_2$ requires $180 \text{ cm}^3 \text{ H}_2$ and can form $120 \text{ cm}^3 \text{ NH}_3$ if reaction is complete. With only $150 \text{ cm}^3 \text{ H}_2$, hydrogen limits the result: N_2 used $= 150/3 = 50 \text{ cm}^3$, NH_3 formed $= 150 \times 2/3 = 100 \text{ cm}^3$ and N_2 remaining $= 60 - 50 = 10 \text{ cm}^3$. A real reversible reaction may not reach that theoretical conversion. Do not assign the gas volume ratio to a solid or condensed liquid product merely because it has a coefficient in the equation.

Checklist before you call this topic done

- Balance atoms and weighted masses; explain gas mass changes and measurement uncertainty.
- Apply the relevant tier's mole, yield, atom-economy, concentration and gas calculations with units and conditions.

4 • Chemical changes

AQA • GCSE • Chemistry

Chemical changes: choose the reaction from the evidence

A copper wire can stay unchanged in dilute acid while magnesium fizzes. A battery can then force a reaction that would not happen by mixing alone. Use the **reactivity series**, the particles present and the reaction conditions to explain the difference. Sheets 4.1–4.3 contain original practice and authored marking guidance; Higher Tier (HT) demands are labelled.

Reactivity of metals (4.4.1.1–4.4.1.4)

Metal atoms form positive ions. A more reactive metal has a greater tendency to do this. Learn potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper in decreasing reactivity. Carbon and hydrogen are useful reference points; they are not metals.

more reactive ↑	potassium	cold water: hydroxide + hydrogen
	sodium	cold water: hydroxide + hydrogen
	lithium	cold water: hydroxide + hydrogen
	calcium	cold water: hydroxide + hydrogen
	magnesium	cold water: extremely slow
	CARBON	reference for oxide reduction
	zinc	dilute HCl / H_2SO_4 : salt + hydrogen
	iron	dilute HCl / H_2SO_4 : salt + hydrogen
	HYDROGEN	reference for acid / cathode products
	copper	no reaction with cold water or dilute HCl

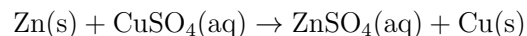
Carbon and hydrogen are non-metals. Compare metals under matching conditions.

At room temperature, potassium, sodium and lithium react with water to form a soluble hydroxide and hydrogen. Sodium usually melts into a ball; lithium floats and fizzes; potassium may burn with a lilac flame. Calcium fizzes and forms calcium hydroxide; the mixture can become cloudy. Magnesium reacts extremely slowly with cold water; zinc, iron and copper show no observable reaction. Compare matching conditions and clean, similar-sized samples. Steam reactions are outside this section.

With dilute hydrochloric or sulfuric acid, magnesium, zinc and iron make a **salt** and hydrogen; copper does not. The more reactive metals above magnesium react

too vigorously for an ordinary classroom acid comparison. Wear eye protection; keep hydrogen away from flames during collection. The specified acid–metal chemistry uses Mg, Zn and Fe with these two acids; do not apply its hydrogen rule to nitric acid.

Displacement occurs when a more reactive metal replaces a less reactive metal from a compound. For zinc in copper(II) sulfate:

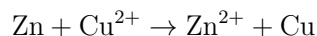


Copper forms and the blue solution fades as copper ions are removed. A result can locate a metal between two known metals; one test alone cannot determine its whole rank.

Metals gain oxygen to make oxides: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$. This is **oxidation**. Loss of oxygen is **reduction**. Gold can occur as the metal itself because it is unreactive. Most metals occur as compounds. Carbon can reduce oxides of metals below it; a metal above carbon needs another method such as electrolysis. Detailed industrial extraction processes are not required here; use any extra process information supplied in a question.

Worked oxygen-transfer example. In $2\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 4\text{Fe} + 3\text{CO}_2$, iron oxide loses oxygen and is reduced. Carbon gains oxygen and is oxidised. The equation conserves four iron atoms, six oxygen atoms and three carbon atoms.

HT: electron transfer. Oxidation is loss of electrons; reduction is gain. For zinc/copper displacement, omit unchanged sulfate ions:



Zinc atoms are oxidised; copper ions, not copper atoms, are reduced. For $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$, Mg loses two electrons and two H^+ ions gain them; chloride is unchanged.

Reactions of acids and making salts (4.4.2)

An acid in water produces H^+ ions. An alkali is a soluble base whose aqueous solution contains OH^- ions; it does not donate electrons to neutralise an acid. **Neutralisation** between acid and base makes salt and water. Carbonates also produce carbon dioxide. A metal reaction instead produces hydrogen.

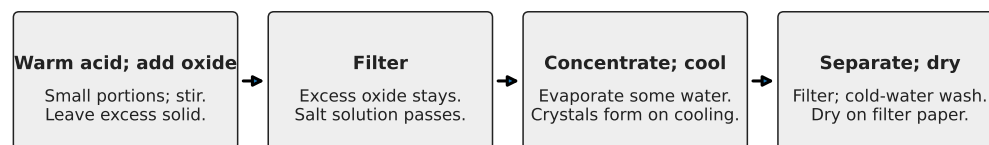
- Hydrochloric acid makes chlorides; nitric acid makes nitrates; sulfuric acid makes sulfates.
- The positive ion comes from the metal, base or carbonate. Balance charges before balancing the equation.
- Example: Ca^{2+} and NO_3^- make $\text{Ca}(\text{NO}_3)_2$, not CaNO_3 .



- $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$ is the ionic neutralisation equation.

Required practical 1: pure dry copper(II) sulfate crystals. Gently warm dilute sulfuric acid, add copper(II) oxide in small portions with stirring until some black solid remains, and filter. Excess insoluble solid shows that the acid has reacted; filtering removes that excess. Heat the blue filtrate gently in a water bath or with an electric heater to concentrate it, without boiling it dry. Leave it to cool and crystallise. Filter the crystals, wash with a little cold distilled water and dry between filter papers. Washing removes solution from the surface; use little cold water to limit crystal loss. Wear eye protection and handle hot apparatus carefully. This written method does not replace supervised practical work.

Pure dry copper(II) sulfate: choose each step by what is soluble

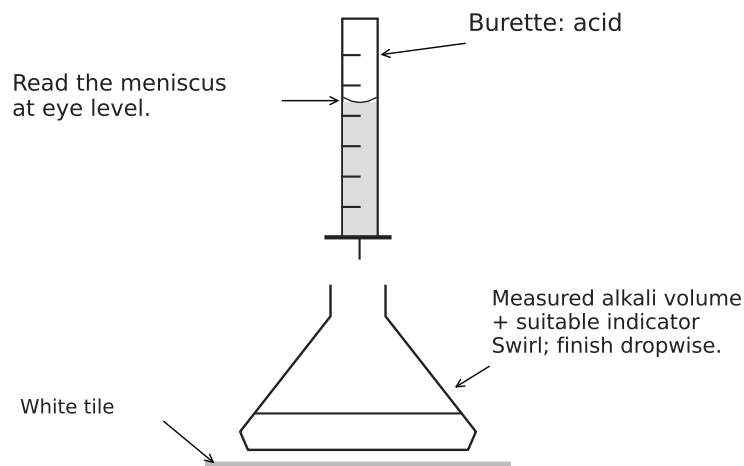


Do not try to filter dissolved acid; do not boil the salt solution dry.

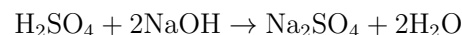
pH scale. At this level, pH runs from 0 to 14: below 7 acidic, 7 neutral, above 7 alkaline. Use universal indicator with its colour chart for an approximate pH, or a suitable pH probe for a reading. Add indicator to a small sample, not the stock bottle. A single narrow-range indicator detects a suitable **titration** endpoint but does not measure the full pH range. Diluting an acid lowers H^+ concentration and brings pH towards 7; it does not make the acid an alkali.

Required practical 2, chemistry only: titration. Rinse a pipette with the alkali, use a pipette filler to transfer a fixed volume into a conical flask, and add a few drops of a suitable indicator. Rinse the burette with the acid; fill it and remove the jet air bubble. Read the bottom of the meniscus at eye level and record initial and final readings. Swirl the flask and add acid slowly, dropwise near the endpoint, over a white tile. The titre is final minus initial reading. Use a rough trial, then repeat accurate runs and average **concordant titres** using the stated tolerance; exclude the rough result. Adding distilled water to the flask washes down splashes without changing the alkali amount. Actual practical work remains required.

Titre = final burette reading – initial burette reading



Worked HT calculation. 20.0 cm³ of 0.100 mol/dm³ NaOH reacts with 25.0 cm³ sulfuric acid. Find the acid concentration and its mass concentration; $M(\text{H}_2\text{SO}_4) = 98$ g/mol.



$$V(\text{NaOH}) = \frac{20.0}{1000} \text{ dm}^3 = 0.0200 \text{ dm}^3$$

$$n(\text{NaOH}) = cV = (0.100 \text{ mol/dm}^3)(0.0200 \text{ dm}^3) = 0.00200 \text{ mol}$$

$$n(\text{H}_2\text{SO}_4) = \frac{n(\text{NaOH})}{2} = \frac{0.00200 \text{ mol}}{2} = 0.00100 \text{ mol}$$

$$c(\text{H}_2\text{SO}_4) = \frac{n}{V} = \frac{0.00100 \text{ mol}}{0.0250 \text{ dm}^3} = 0.0400 \text{ mol/dm}^3$$

$$c_m = cM = (0.0400 \text{ mol/dm}^3)(98 \text{ g/mol}) = 3.92 \text{ g/dm}^3$$

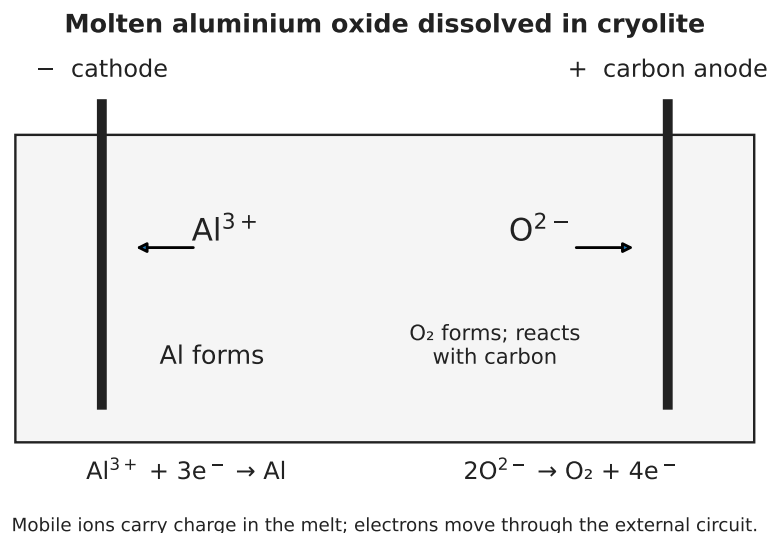
HT: strong and weak acids. A **strong acid** is completely ionised in water; hydrochloric, nitric and sulfuric acids are specified examples. A **weak acid** is only partially ionised; ethanoic, citric and carbonic acids are examples. At a given concentration, a stronger acid has lower pH. Concentrated/dilute describes amount of acid per solution volume, whereas strong/weak describes degree of ionisation. Diluting a strong acid does not turn it into a weak acid. A drop of one whole pH unit means ten times the H⁺ concentration; pH 2 has 100 times the H⁺ concentration at pH 3. Neutrality concerns the balance of H⁺ and OH⁻, not the absence of ions.

Electrolysis (4.4.3.1–4.4.3.5)

An **electrolyte** conducts when its ions can move: an ionic compound may be molten or dissolved in water. Ions cannot carry charge through a solid lattice because they cannot move from place to place. **Electrolysis** uses a current to discharge ions at electrodes and produce elements. Cations move to the negative **cathode** and gain electrons; anions move to the positive **anode** and lose electrons. Ions carry charge in the liquid; electrons carry charge through the external metal circuit.

Molten binary compounds, inert electrodes. Lead bromide gives lead at the cathode and bromine at the anode. Molten zinc chloride gives zinc and chlorine. There is no water to compete in a molten salt.

Aluminium extraction. Aluminium is above carbon. Electrolysis uses much electrical energy for the current and heating. Aluminium oxide dissolves in molten cryolite; the mixture has a lower melting point than pure aluminium oxide, reducing heating demand. Oxygen formed at the positive carbon electrode reacts with it to make carbon dioxide, so this electrode must be replaced. The diagram applies to this molten mixture, not an aqueous aluminium salt.



Aqueous solutions with inert electrodes. Water supplies H⁺ and OH⁻ as additional possibilities. In the specified GCSE rule, a metal below hydrogen deposits at the cathode; if the metal is above hydrogen, hydrogen forms. At the anode, a halide gives its halogen; otherwise oxygen forms. Concentration and electrode material can matter in real systems; use the conditions supplied and do not transfer an inert-electrode prediction to reactive copper electrodes without extra information.

Electrolyte	Cathode	Anode
Aqueous NaCl	hydrogen	chlorine
Aqueous CuCl ₂	copper	chlorine
Aqueous Na ₂ SO ₄	hydrogen	oxygen
Aqueous CuSO ₄	copper	oxygen

Required practical 3. Develop a hypothesis for each solution before testing with inert electrodes. Use labelled separate samples, the same electrode material/spacing, suitable low-voltage DC and matching test time. Observe each electrode separately. Test hydrogen with a lighted splint on a small collected sample (squeaky pop); oxygen relights a glowing splint; chlorine bleaches damp litmus paper. A copper deposit is pink-brown. Wear eye protection, avoid skin contact, and use only small gas quantities with good ventilation and teacher supervision; never smell chlorine. Wash electrodes between solutions to reduce contamination. Compare actual evidence with the hypothesis; absence of visible gas is not proof of no reaction. Written prediction does not replace the practical.

HT: half equation . First balance atoms, then charge using electrons; electrons are on the left for reduction and right for oxidation.



In the oxide equation the reactant is two oxide ions, not peroxide or an oxygen molecule with a charge. Check both oxygen atoms and total -4 charge. In the hydroxide equation four oxygen and four hydrogen atoms occur on each side, with total -4 charge on each side.

Checklist before you call this topic done

- Use observed reactions to order unknown metals; distinguish oxygen-transfer and HT electron-transfer redox.
- Choose reagents and a separation method for a named pure dry salt; deduce its formula from charges.
- Explain pH measurement, dilution, strength and concentration; carry out chemistry titration methods and HT calculations with the balanced ratio.
- Predict molten/aqueous products, explain aluminium extraction, and test an RP3 hypothesis; HT students balance both half-equations.
- Attempt unfamiliar tasks in sheets 4.1–4.3 before viewing the marking guidance. Practical participation and authentic-source integration are separate checks.

5 • Energy changes

AQA • GCSE • Chemistry

Energy changes: follow the energy transfer

A hand warmer heats its surroundings; an instant cold pack cools them. In both, energy is conserved. Follow where energy moves, then decide whether the evidence supports the product's intended use. This topic covers AQA Chemistry 8462 sections 4.5.1.1–4.5.2.2. Bond-energy calculations are **Higher Tier**; cells and fuel cells are **chemistry only**, with fuel-cell half-equations **Higher Tier**.

Exothermic and endothermic changes (4.5.1.1)

An **exothermic** reaction transfers energy from the reacting chemicals to the surroundings. The surroundings warm, and the products have less energy than the reactants by the amount transferred. Combustion, many oxidation reactions and neutralisation are examples. Hand warmers and self-heating cans use exothermic changes.

An **endothermic** reaction takes energy from the surroundings. The surroundings cool; the products have more energy than the reactants. Thermal decomposition and the reaction of citric acid with sodium hydrogencarbonate are examples. Some sports injury packs use endothermic changes. A temperature fall alone does not prove that a new substance formed: physical processes such as dissolving can also change temperature.

Worked temperature example. A reaction mixture changes from 21.0 °C to 15.5 °C without external heating or cooling. Temperature change = final – initial = 15.5 – 21.0 = -5.5 °C. The mixture cooled, supporting an endothermic interpretation: the reacting chemicals took energy from the surroundings. A thermometer measures temperature, not the reaction's energy change directly. No calculation of solution energy or ΔH is required here.

Evaluate a use with evidence. A warmer must reach a suitable temperature for long enough without leaking or overheating. An instant cold pack must cool sufficiently, be available when needed and limit waste. Compare given temperature, duration, storage, cost and reuse information. A chemical change being exothermic does not by itself make an effective hand warmer. A rechargeable battery and a hydrogen fuel cell need a similar purpose-based comparison later.

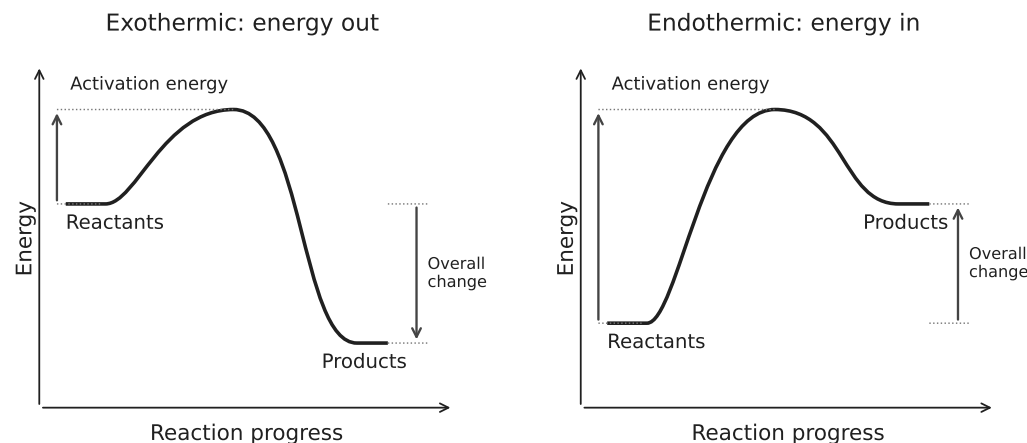
Required practical 4: temperature changes in reacting solutions. Choose one variable, for example the total alkali volume added to a fixed volume of acid. Use teacher-approved solutions, eye protection and an insulated cup supported safely in a beaker. Measure the initial temperature; add a measured alkali volume, replace the lid, stir gently and record the highest temperature. Add further measured portions and record total added volume and temperature. Keep the initial acid volume, concentrations, starting temperatures, cup and reading method comparable. Repeat

the investigation with fresh starting solutions and calculate means for corresponding volumes.

Plot temperature against total added volume. An increase followed by a decrease can locate a maximum; use appropriate best-fit lines on the two regions to estimate their intersection between readings. Do not call the highest measured point an exact end-point. Insulation and a lid reduce heat exchange, but do not remove it. Record actual observations and apparatus during supervised practical work; this written method does not replace that work. Other RP4 comparisons may use acids with metals/carbonates or displacement reactions.

Reaction profiles (4.5.1.2)

A **reaction profile** shows the relative energies of reactants and products as a reaction proceeds. Its horizontal axis is reaction progress, not time. Particles must collide with enough energy for reaction. **Activation energy** is the minimum energy particles need for a reaction to occur.



Draw labelled axes, then separate reactant and product levels. Join them with a smooth curve rising above both levels. Draw the activation-energy arrow from the reactant level to the peak; draw the overall-change arrow between reactant and product levels. Exothermic products are lower; endothermic products are higher. An endothermic reaction still has an activation barrier. A catalyst provides a route with a smaller activation energy; the reactant/product levels and overall change stay the same.

Worked profile reading. On a supplied relative-energy scale, reactants are at 40, the peak at 110 and products at 15. Activation energy is $110 - 40 = 70$ relative units. Overall change is $15 - 40 = -25$ relative units: the reaction is exothermic. Using the peak as the starting level for overall change would confuse the two arrows. These are illustrative profile readings, not a solution calorimetry calculation.

Bond energies – Higher Tier (4.5.1.3)

A **bond energy** gives the energy needed to break a mole of the stated bonds in the supplied model. Breaking reactant bonds requires energy; forming product bonds releases energy. Use the balanced equation to count bonds before calculating. A coefficient multiplies all the bonds in that molecule. Use a supplied double-bond value once for each double bond.

overall energy change = total energy to break bonds – total energy released forming bonds

For an exothermic reaction, forming releases more than breaking requires, so the result is negative. For an endothermic reaction, breaking requires more than forming releases, so the result is positive. Bond formation releases energy even in an endothermic reaction; it simply releases less than was required to break the reactant bonds.

Worked bond-count example. For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, use $\text{H-H} = 436$, $\text{O=O} = 498$ and $\text{O-H} = 464$ kJ/mol. Break two H-H bonds and one O=O bond; form four O-H bonds.

$$\text{broken} = 2(436) + 498 = 1370 \text{ kJ/mol}$$

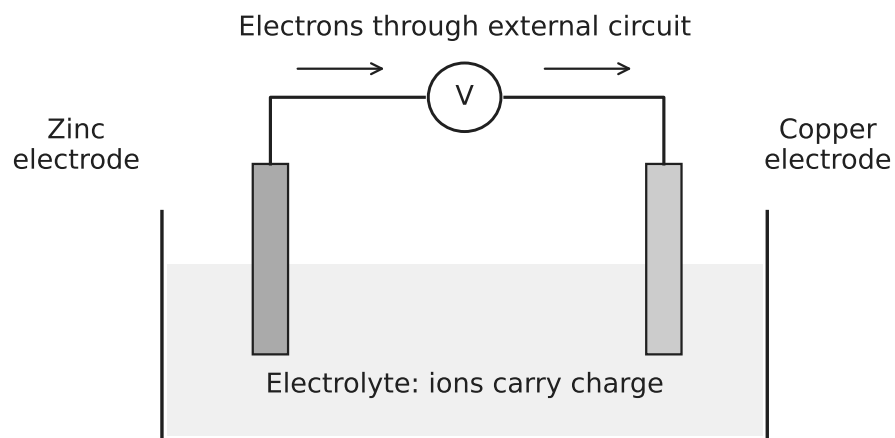
$$\text{formed} = 4(464) = 1856 \text{ kJ/mol}$$

$$\text{overall change} = 1370 - 1856 = -486 \text{ kJ/mol}$$

The reaction is exothermic for the quantities represented by the balanced equation. The supplied bond energies give an approximate model; do not use the thermometer practical to infer this number without another model.

Cells and batteries – chemistry only (4.5.2.1)

A **cell** contains chemicals that react to produce electricity. A simple cell uses two different metal electrodes in an **electrolyte**, connected through an external circuit. Ions carry charge in the electrolyte; electrons move through the wire. Measure the potential difference with a voltmeter. Voltage depends on the electrodes, electrolyte and conditions. Use supplied readings to compare particular cells; a metal's reactivity alone does not fix a universal voltage.



Illustrative Zn/Cu cell; voltage depends on electrolyte and conditions

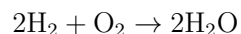
For a controlled comparison, change one electrode while keeping the other electrode, electrolyte and measuring conditions comparable. For example, if a supplied table gives $\text{Zn/Cu} = 1.10 \text{ V}$ and $\text{Fe/Cu} = 0.78 \text{ V}$ in the same electrolyte, Zn/Cu gives the greater voltage under those conditions. Changing the electrolyte is another factor, so it would prevent attributing a difference only to the metal.

A **battery** has two or more cells connected in series to provide a greater voltage when their orientations reinforce one another. Three identical 1.2 V cells connected that way give $3 \times 1.2 = 3.6 \text{ V}$ in the stated ideal model. Do not connect terminals directly together; actual school work needs approved equipment and procedures.

In a **non-rechargeable** cell, the reaction stops when a reactant is used up. Alkaline batteries are the specified example. In a **rechargeable** cell, supplying an external electrical current reverses the chemical reactions. Charging transfers energy into the cell; it does not create energy. Do not attempt to recharge a non-rechargeable cell. Evaluate a device using the given purchase cost, number of uses, charging access, waste, mass and performance. More uses may offset a higher initial cost, but do not automatically prove the lowest environmental impact.

Hydrogen fuel cells – chemistry only (4.5.2.2)

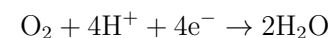
A **fuel cell** receives fuel and oxygen or air from an external supply. Hydrogen is oxidised electrochemically, producing a potential difference; the overall hydrogen–oxygen reaction produces water:



The cell can keep operating while suitable supplies continue. Water is the direct reaction product. Hydrogen production, transport and storage can have other environmental effects, so this equation does not prove zero lifetime emissions. Both

a rechargeable-battery vehicle and a hydrogen-fuel-cell vehicle can have no polluting exhaust at the vehicle. Use given data on refuelling/charging time, range, mass, infrastructure, fuel production and cost to choose for a stated purpose. Do not claim that a fuel cell is always cleaner or cheaper.

Higher Tier: one consistent acidic half-equation model. At the negative electrode, the anode, hydrogen loses electrons and is oxidised. At the positive electrode, the cathode, oxygen gains electrons and is reduced. Electrons flow through the external circuit from the hydrogen electrode to the oxygen electrode.



Each line conserves atoms and charge. Adding them cancels four electrons and four hydrogen ions, leaving the overall equation above. These signs describe a cell supplying electricity; do not copy the electrode signs from externally powered electrolysis. An alkaline model is also possible, but do not mix its hydroxide equations with this acidic pair. Use approved educational equipment or supplied observations; this lesson does not require improvised gas storage.

Check yourself

Classify a change from temperature evidence and explain where energy moves. Evaluate a given warmer/cold-pack design. Plan and interpret the RP4 temperature investigation. Draw both smooth reaction profiles with distinct arrows. For HT, count bonds before subtracting broken minus formed. For chemistry only, interpret cell data, explain series batteries and rechargeability, and compare hydrogen fuel cells with rechargeable cells using a stated evidence boundary. For HT fuel cells, balance both half equations and cancel equal electron numbers.

6 • The rate and extent of chemical change

AQA • GCSE • Chemistry

The rate and extent of chemical change

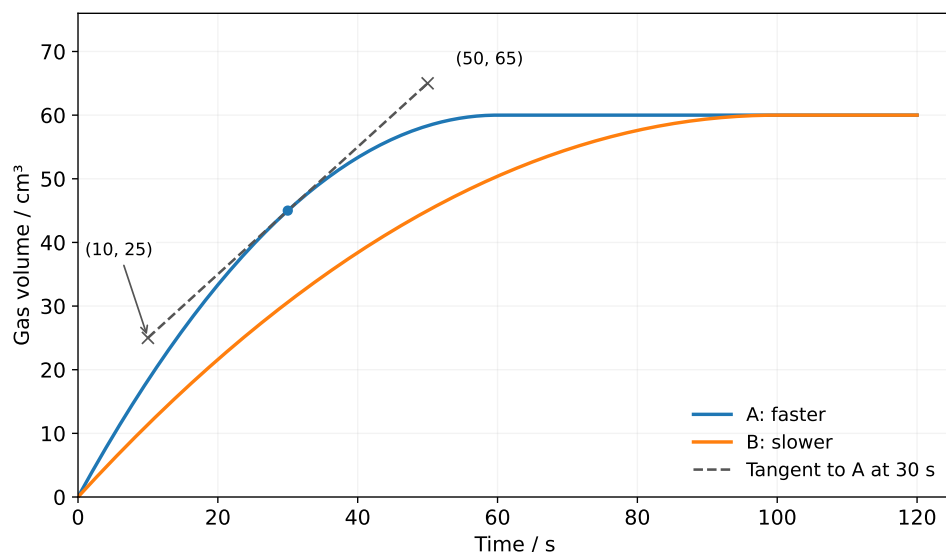
A crushed tablet can react quickly without making more gas than the same tablet left whole. A factory can speed a reaction without increasing its equilibrium yield. Separate **rate of reaction**, how fast change happens, from the amount finally obtained.

This reference covers AQA 8462 sections 4.6.1.1–4.6.2.7. Mean mass/volume rates, drawing/interpreting tangents and equilibrium are both-tier. Numerical tangent gradients, mol/s rates and equilibrium-shift predictions are Higher Tier (HT). Required practical 5 needs both supervised methods as well as written understanding.

Measure a rate (4.6.1.1)

Mean rate is quantity of reactant used divided by time, or quantity of product formed divided by time. These are alternative measurements, not necessarily equal numerical quantities. State the quantity and unit: g/s or cm³/s; HT also uses mol/s. If gas volume changes from 12 to 36 cm³ between 10 and 30 s, use $(36 - 12)/(30 - 10) = 1.2$ cm³/s. Dividing 36 by 30 describes a different interval.

A product–time curve normally rises and flattens as a reactant runs out. A remaining-reactant curve falls and flattens. Compare steepness for rate and plateaux for final quantity. A lower plateau alone does not prove slower chemistry: a different limiting amount or collection loss can change the plateau.



A **tangent** follows the curve's local direction at the chosen time. Draw a straight line touching the curve there; extend it to read two separated points on the tangent. It is not a line joining two nearby points on the curve. Its gradient represents the **instantaneous rate** at that instant; steeper means faster.

HT worked example. The illustrated tangent passes through (10 s, 25 cm³) and (50 s, 65 cm³). Its gradient is $(65 - 25)/(50 - 10) = 1.0$ cm³/s at 30 s. The mean rate from zero to 30 s is instead $45/30 = 1.5$ cm³/s. The tangent can extend above the plateau: its extension is a local construction, not the later gas-volume curve.

HT molar example. Reactant amount falls from 0.018 to 0.006 mol in 40 s. Amount used is 0.012 mol, giving $0.012/40 = 0.00030$ mol/s. Use the positive amount consumed rather than a negative reaction rate from the falling graph.

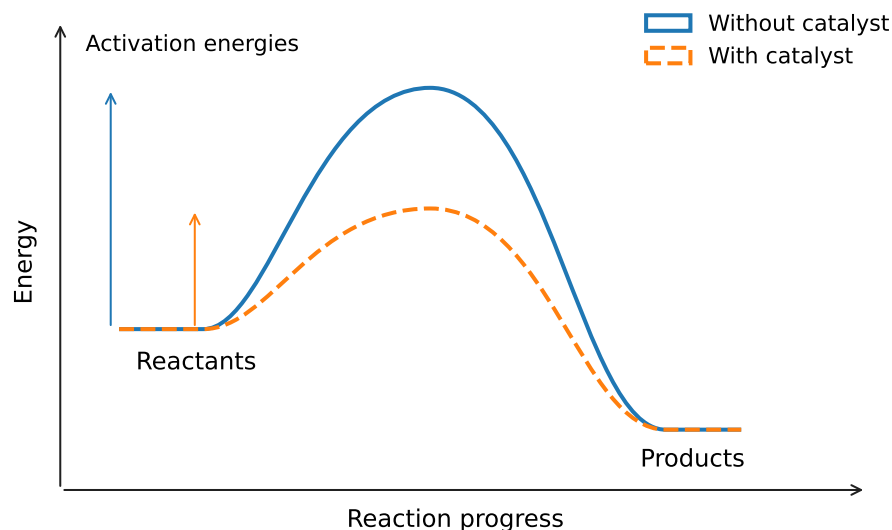
Explain rate changes (4.6.1.2–4.6.1.4)

Collision theory says reacting particles must collide with enough energy.
Activation energy is the minimum energy needed for reaction.

Change	Explanation under comparable conditions
Higher solution concentration	More reactant particles per volume; more frequent collisions.
Higher reacting-gas pressure at fixed temperature	Particles closer together; more frequent collisions.
Smaller pieces of the same solid mass	Larger surface-area-to-volume ratio; more particles exposed for collisions.
Higher temperature	Faster particles collide more frequently; a greater proportion has enough energy.
Add a suitable catalyst	An alternative pathway has lower activation energy.

At fixed temperature, higher concentration does not give each particle more kinetic energy. Breaking a lump into pieces does not increase its mass. The **surface-area-to-volume ratio** compares exposed area with solid volume. Use proportional reasoning only when justified: doubling particles per volume can explain more collisions, but does not prove every measured rate exactly doubles.

A **catalyst** changes rate and is not used up overall. Different reactions need suitable catalysts; **enzymes** are biological catalysts. Identify a catalyst from its effect on rate together with absence from the overall equation, not absence alone. A solvent may also be absent without being a catalyst.



The catalysed curve has a lower maximum relative to the reactants. Both start and finish at the same levels, so overall energy change is unchanged. A catalyst is not an energy source. Faster reaction does not mean more product from the same limiting amount; do not explain catalytic action by saying particles become hotter.

Required practical 5: two concentration methods (4.6.1.2)

Develop a hypothesis before measuring: higher acid concentration should give a steeper gas-volume curve with matched magnesium; higher thiosulfate concentration should reach the same cloudiness threshold sooner. Both approaches are required. Follow the approved school procedure; written answers do not certify participation.

Gas-volume method. Adopt the source-reviewed handbook activity: equal cleaned 3 cm magnesium ribbons react with 50 cm³ teacher-prepared dilute hydrochloric acid at two concentrations. Collect hydrogen through a delivery tube into an inverted water-filled measuring cylinder; a gas syringe is an allowed alternative. Keep ribbon mass/length, oxide removal, acid volume, starting temperature and apparatus the same. Add magnesium, fit the bung promptly and start timing consistently. Record volume, for example every 10 s, until constant. Repeat and plot both curves on the same axes.

Check for leaks and an unobstructed gas path; never seal a gas-producing reaction without an outlet. Use eye protection, a stable collector and no ignition source near hydrogen. A delayed bung loses early gas. Water collection needs suitable gas solubility and consistent readings; a syringe can stick or leak. Equal final volumes require equal limiting magnesium and enough acid in both trials.

Colour/turbidity method. Turbidity is cloudiness from suspended particles. The reviewed activity uses 40 g/dm³ sodium thiosulfate stock. Mix stock

and water to a constant 50 cm³, then add the same 10 cm³ dilute hydrochloric acid. Stock/water volumes 10/40, 20/30, 30/20, 40/10 and 50/0 cm³ give pre-acid concentrations 8, 16, 24, 32 and 40 g/dm³. These labels describe the dilution before adding acid, not the final 60 cm³ mixture.

Place the same flask on a black cross, add acid, swirl consistently and time until the cross disappears. Sulfur makes the mixture cloudy. Keep temperature, total volume, flask/viewing depth, acid amount, lighting and endpoint judgement consistent. Repeat each dilution and compare mean times. At a comparable threshold, 1/mean time is a relative-rate proxy: 60 and 30 s give 0.0167 and 0.0333 per second, so the second proxy is twice as large. This is not a measured mass rate or proof that the whole reaction has finished.

Sulfur dioxide is released and can aggravate breathing difficulties. Use ventilation, eye protection and teacher-selected controls; do not lean over the flask or inhale fumes. Record raw repeats and justify any anomaly exclusion. Averaging cannot remove systematic viewing differences. Reducing stock volume without replacing water changes viewing depth as well as concentration.

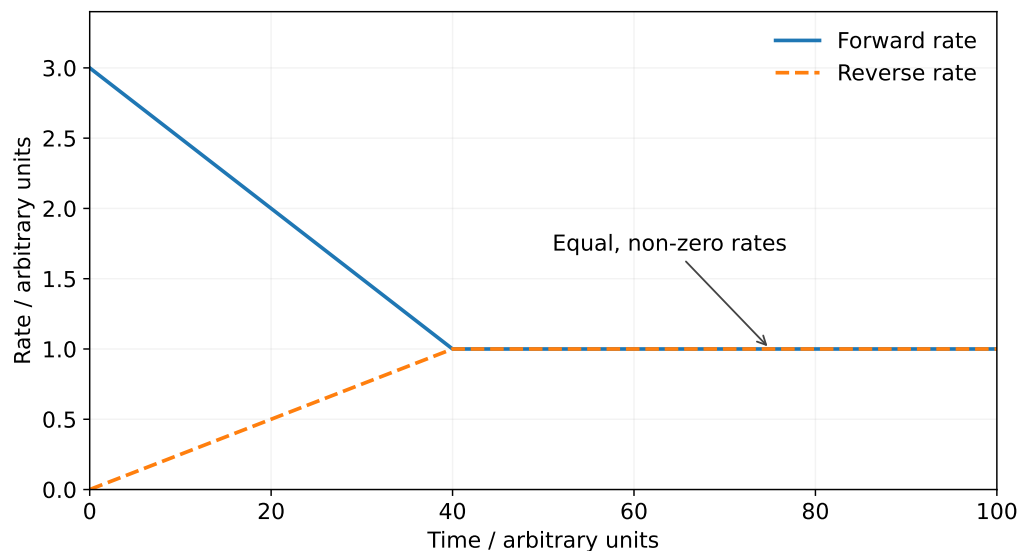
Reverse a reaction and its energy transfer (4.6.2.1–4.6.2.2)

In a **reversible reaction**, products can form the original reactants. Use $\rightleftharpoons$; changing conditions can change the direction favoured. Heating ammonium chloride forms ammonia and hydrogen chloride; cooling allows the gases to form ammonium chloride again. Preserve the substances in both directions.

Hydrated copper sulfate is blue. Heating removes water to form white anhydrous copper sulfate; adding water reverses this change. Dehydration is endothermic and hydration exothermic. For the same amount, energy transfers have equal magnitude and opposite directions. If dehydration takes in 72 kJ, reverse hydration releases 72 kJ for the corresponding amount; it does not release 144 kJ.

Dynamic equilibrium (4.6.2.3)

A **closed system** prevents reactants/products escaping. **Dynamic equilibrium** means forward and reverse reactions continue at equal rates. Concentrations remain constant but need not be equal. A sealed vessel with constant 0.60 mol reactant and 0.20 mol product can be at equilibrium if the rates are equal.



Closing a vessel allows equilibrium but does not establish it instantly. Starting with reactants, forward rate can fall while reverse rate rises as product forms. At equilibrium particles still react both ways; a flat concentration trace does not mean reaction stopped. An escaping product prevents this closed-system description from applying.

Higher Tier: predict a changed equilibrium (4.6.2.4–4.6.2.7)

Equilibrium position describes relative reactant/product amounts. **Le Chatelier principle** predicts a response opposing a condition change until a new equilibrium is reached. Use qualitative predictions here; no equilibrium-constant calculation is required.

Concentration. Adding a reactant immediately raises its concentration. Subsequent net forward reaction uses some added reactant and forms more product. Removing a reactant favours reverse; adding product favours reverse; removing product favours forward. Concentrations change until equilibrium returns. The disturbed concentration need not return exactly to its original value; rates are unequal during adjustment.

Temperature. Heating favours the endothermic direction; cooling favours exothermic. With an exothermic forward reaction, heating reduces equilibrium product amount and cooling increases it. With an endothermic forward reaction, heating increases product amount and cooling reduces it. Read the given forward energy change. Faster reaction after heating does not decide which equilibrium mixture contains more product.

Gas pressure. Compression at fixed temperature favours fewer gas molecules; expansion favours more. Count gaseous coefficients, not chemical names or solids. In $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$, four gaseous molecules become two, so compression favours ammonia. In $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$, both sides have two, giving no predicted po-

sition shift. This rule concerns changing volume, not adding a reactant and changing composition.

A suitable catalyst speeds both directions and reaches equilibrium sooner, without changing final position at the same conditions. In an industrial exothermic process, low temperature may favour yield but make production slow. A moderate temperature can balance rate, yield, energy use and cost. Pressure choices also need equipment/safety costs. Use supplied data: maximum yield alone is not an economic judgement.

Check yourself

- Choose quantity and interval; draw/interpret a tangent. HT calculates tangent and molar rates.
- Explain all five factors; distinguish faster reaction from final amount.
- Plan both RP5 methods with hypotheses, controls, repeats, observation limits and hazards.
- Explain reversibility, opposite energy transfers and equal continuing rates in a closed system.
- HT separates disturbance from adjustment, counts gas coefficients and predicts heating/cooling for either energy direction.

7 · Organic chemistry

AQA · GCSE · Chemistry

Organic chemistry: choose the reaction

A refinery can separate crude oil, break large molecules into smaller ones, or use small molecules to make polymers. These are three different processes. Distillation changes which molecules are collected together; cracking changes molecules; polymerisation joins monomers into chains. Choose the process from the evidence, not from the word “heat”.

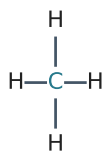
This reference covers AQA 8462 sections 4.7.1–4.7.3. The named organic families and addition polymers are Separate Chemistry content on both tiers. **Higher Tier (HT)** is marked for weak-acid explanations, condensation polymers and amino-acid condensation. Recognising natural-polymer monomers is both-tier content.

Crude oil and alkane structures (4.7.1.1)

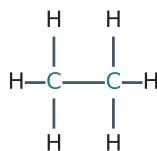
Crude oil is a finite resource found in rocks. It formed from ancient biomass, mainly plankton buried in mud. Formation is too slow to replace present consumption on a human timescale. It is a mixture of many compounds, most of them **hydrocarbons**: compounds containing hydrogen and carbon only. Ethanol contains oxygen, so it is not a hydrocarbon.

Most crude-oil hydrocarbons are **alkanes**. Their open-chain saturated series has general formula C_nH_{2n+2} . Carbon makes four bonds and hydrogen one. A molecular formula counts atoms; a displayed formula shows every atom and bond. The four required names are methane, ethane, propane and butane.

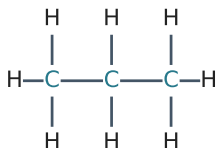
Methane · CH_4



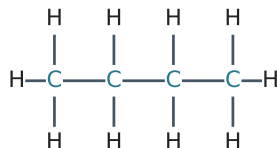
Ethane · C_2H_6



Propane · C_3H_8



Butane · C_4H_{10}

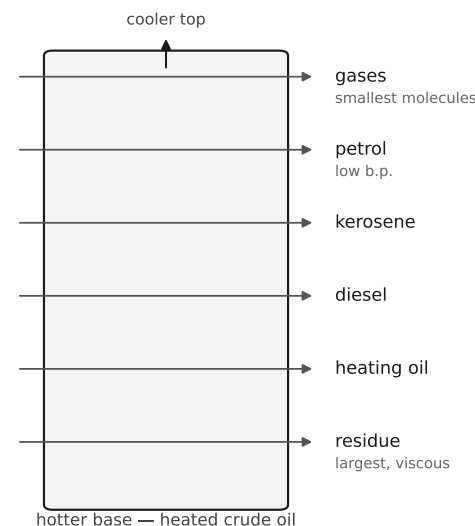


Worked check. A supplied open-chain alkane has six carbon atoms. Its hydrogen count is $H = 2n + 2 = 2 \times 6 + 2 = 14$, giving C_6H_{14} . This does not give a formula for

the whole crude-oil mixture. Names beyond the first four are not required recall here.

Fractions and fuel choices (4.7.1.2–4.7.1.3)

Fractional distillation separates the mixture physically. Heat crude oil so that many components vaporise. Vapours rise through a column that is hotter at the bottom and cooler at the top. Higher-boiling components condense lower down; lower-boiling components rise further before condensing. The lowest-boiling gases can leave at the top. Each **fraction** contains several hydrocarbons with similar boiling points and carbon numbers; it is not one pure compound. Covalent carbon chains are not broken by boiling.



For comparable hydrocarbons, increasing molecule size increases boiling point and **viscosity**, but decreases **flammability**. Viscosity means resistance to flow. A viscous fuel may flow poorly through a narrow supply tube; easy ignition matters for starting a burner. Do not substitute a density trend or claim one property determines every use.

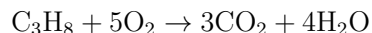
Fractions supply petrol, diesel oil, kerosene, heavy fuel oil and liquefied petroleum gases. They also provide **feedstock**: starting materials for solvents, lubricants, polymers and detergents. Fuels are burned for energy; feedstocks are converted into other useful chemicals.

Worked choice. Fictional fractions P and Q have comparable flow times of 8 s and 24 s. P also ignites more readily. P better meets a stated need for easy flow and ignition. The time ratio $r = 24/8 = 3$ describes this test; it does not prove Q's viscosity is exactly three times P's without a calibrated relation.

Combustion and cracking (4.7.1.3–4.7.1.4)

Complete combustion oxidises carbon to carbon dioxide and hydrogen to water, releasing energy. Start with the fuel formula and balance carbon, then hydrogen, then oxygen. Do not change the formula subscripts to balance an equation.

Worked combustion. Propane contains three C atoms and eight H atoms, so one molecule forms three carbon dioxide and four water molecules. These products contain ten O atoms, requiring five oxygen molecules:



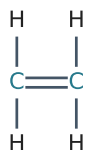
Cracking breaks large hydrocarbons into smaller, more useful molecules. Catalytic cracking passes vaporised hydrocarbons over a hot catalyst. Steam cracking mixes hydrocarbon vapour with steam at a very high temperature. Recall these general conditions, not one compulsory numerical temperature. Products include alkanes and alkenes. Smaller fuels are in high demand; alkenes are chemical feedstock for polymers and other products.

Worked cracking. In $\text{C}_{10}\text{H}_{22} \rightarrow \text{C}_8\text{H}_{18} + \text{X}$, subtract atoms: $C(X) = 10 - 8 = 2$ and $H(X) = 22 - 18 = 4$. Thus $\text{X} = \text{C}_2\text{H}_4$, ethene. Both C and H balance. Real cracking can give a mixture; do not claim every large molecule always splits into precisely these two products.

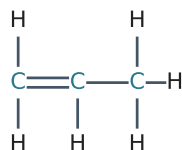
Alkene structures and addition (4.7.2.1–4.7.2.2)

An **alkene** has a carbon–carbon double bond, $\text{C}=\text{C}$. The double bond is its **functional group**, which gives characteristic reactions. Open-chain alkenes with one $\text{C}=\text{C}$ have formula C_nH_{2n} . They are **unsaturated**: two fewer H atoms than the corresponding alkane. Formula alone does not locate the double bond. Butene and pentene below show one terminal-double-bond arrangement.

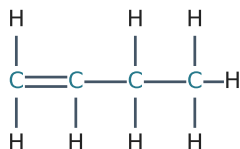
Ethene · C_2H_4



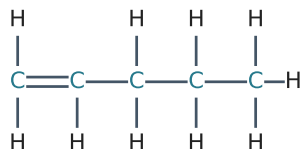
Propene · C_3H_6



Butene · C_4H_8



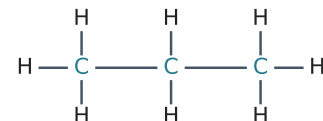
Pentene · C_5H_{10}



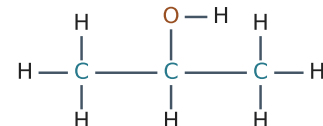
In **addition**, $\text{C}=\text{C}$ becomes $\text{C}-\text{C}$ and atoms from the other reactant attach to the two former double-bond carbons. Keep the carbon skeleton and existing hydrogens. Check carbon valency four after drawing.

- Hydrogen: use heated nickel catalyst; one H joins each carbon, making an alkane.
- Water: use steam, high temperature and pressure, with a phosphoric-acid catalyst; H and OH add across $\text{C}=\text{C}$, making an alcohol. For a longer alkene, show the stated OH position rather than assuming the molecular formula fixes it.
- Halogens: one atom from Cl_2 , Br_2 or I_2 joins each carbon. Chlorine/bromine addition is described at room temperature; do not promise identical speed or completion for iodine. No hydrogen is removed in the addition model.

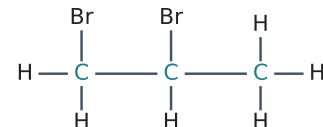
Propene + $\text{H}_2 \rightarrow$ propane · heated nickel



Propene + steam $\rightarrow$ shown OH-position product · acid catalyst



Propene + $\text{Br}_2 \rightarrow$ dibromo product · no H removed



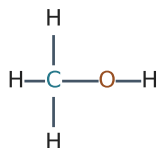
Bromine-water test. In the controlled alkane/alkene comparison, an alkene changes orange bromine water to colourless. The alkane control stays orange under these conditions. This supports unsaturation; it does not uniquely identify ethene. Use approved microscale demonstrations, not independent handling of elemental halogens.

Alkenes tend to burn in air with smoky flames because combustion is incomplete. With sufficient oxygen their complete combustion still forms carbon dioxide and water. For ethene, $\text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$. Hydrogen addition is a different reaction from burning in oxygen.

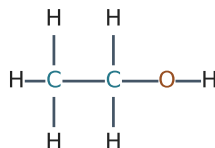
Alcohols and fermentation (4.7.2.3)

Alcohols contain $-\text{OH}$. Learn methanol, ethanol, propanol and butanol. The longer examples shown are the primary straight-chain forms; another OH position can change oxidation behaviour.

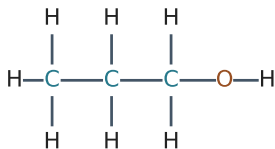
Methanol



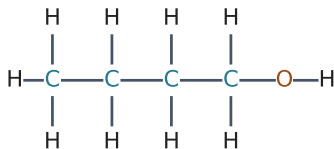
Ethanol



Propanol · primary form shown

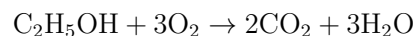


Butanol · primary form shown



Describe four changes for these taught alcohols: sodium gives fizzing as hydrogen forms; combustion releases energy; adding to water gives an aqueous mixture, with mixing behaviour depending on the alcohol and amounts; an oxidising agent converts the taught primary form to its corresponding carboxylic acid. Longer alcohols mix less readily with water than the shorter members. Do not claim unlimited solubility for all four.

Alcohols are fuels, solvents and chemical feedstocks. Their ability to burn supports fuel use, while dissolving suitable substances supports solvent use. Only balanced combustion equations are required here for alcohol reactions. For ethanol, count its own oxygen atom as well as supplied oxygen:



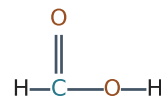
Fermentation makes aqueous ethanol from sugar solution using yeast under warm, oxygen-free conditions. Yeast enzymes convert sugar to ethanol and carbon dioxide. Cold conditions slow enzyme action; excessive heat damages the enzymes/yeast. The product is an ethanol solution, not pure ethanol. Industrial ethene hydration uses steam and a catalyst instead of yeast and sugar.

For a supervised comparison, keep sugar concentration, yeast amount, temperature and collection time matched; include a yeast-free control and record raw gas observations. An approved airlock lets carbon dioxide escape while restricting incoming air. Never tightly seal a gas-producing vessel. Gas alone does not measure ethanol purity or prove every bubble came from fermentation. Written planning does not certify the practical activity.

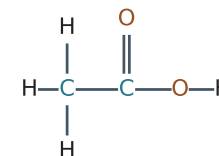
Carboxylic acids and esters (4.7.2.4)

Carboxylic acids contain -COOH , with C=O and O-H attached to the same carbon. That carbon counts in the chain. The first four are methanoic, ethanoic, propanoic and butanoic acids.

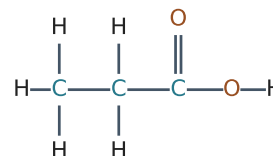
Methanoic acid



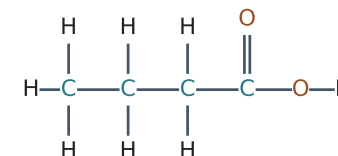
Ethanoic acid



Propanoic acid



Butanoic acid

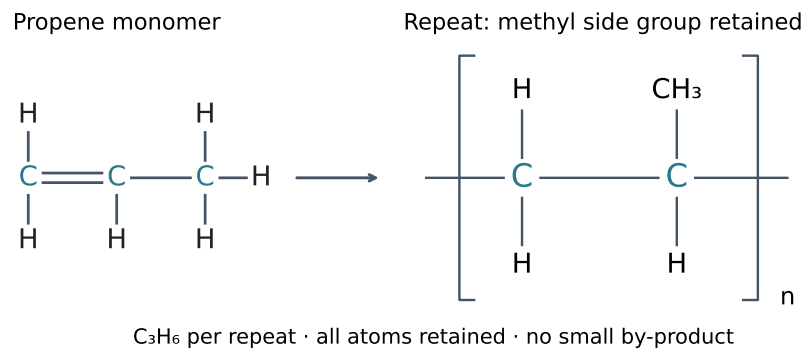


They dissolve in water to form acidic solutions. With carbonates they form a salt, water and carbon dioxide: fizzing is an observation, and carbon dioxide turns limewater milky. With an alcohol they form an **ester** and water. Ethanoic acid + ethanol gives **ethyl ethanoate** + water. Recognise the ester link C(=O)-O-C ; do not call its entire group an alcohol OH. No other ester name or balanced acid-reaction equation is required recall here. Teacher-approved ester preparation uses selected reagents, controlled warming and safe waste handling.

HT: weak means partially ionised. A weak carboxylic acid forms only some hydrogen ions in water. At the same stated molar acid concentration, a strong monoprotic acid produces more H^+ and has a lower pH. Weak does not mean dilute: concentration and fraction ionised are different ideas. Compare like conditions before using pH as evidence of acid strength.

Addition polymers and side groups (4.7.3.1)

Many alkene **monomers** join by **addition polymerisation**. Only the two former C=C carbons become neighbouring backbone positions; attached side groups stay attached. Every monomer atom remains in the **repeating unit**, and no small-molecule by-product forms.

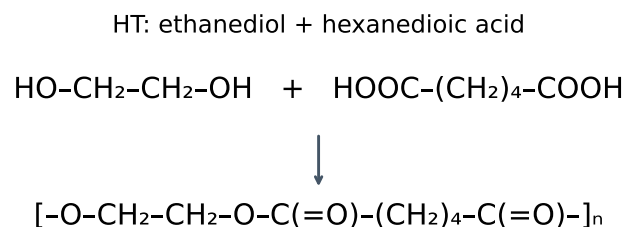


Worked method. Trace the two $\text{C}=\text{C}$ carbons in $\text{CH}_2=\text{CH}-\text{CH}_3$. Replace $\text{C}=\text{C}$ by $\text{C}-\text{C}$; copy the hydrogens and CH_3 side group. Add single continuation bonds through both brackets and write n outside: the interior is $-\text{CH}_2-\text{CH}(\text{CH}_3)-$. Each backbone carbon still has four bonds. Reverse the method to find the monomer: remove continuation bonds and restore $\text{C}=\text{C}$ between those two carbons. Do not put every carbon of the molecule into a three-carbon backbone.

For twelve shown propene-derived units, total carbon count $N_{\text{C}} = 12 \times 3 = 36$: 24 backbone and 12 side-group carbons. Unspecified chain end groups are not included in this repeat-interior count.

HT: condensation and amino acids (4.7.3.2–4.7.3.3)

Condensation polymerisation joins monomers with two functional groups while losing a small molecule, usually water in these examples. One functional group at each end lets the chain extend. A diol with two $-\text{OH}$ groups and a diacid with two $-\text{COOH}$ groups form a **polyester** with repeated ester links.

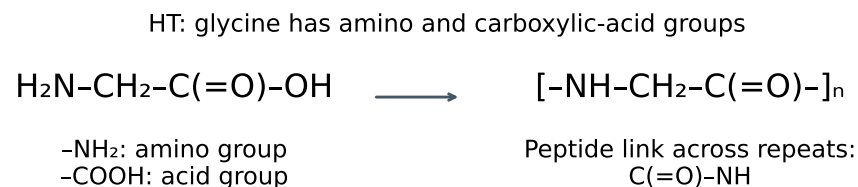


Ester links: $\text{C}(=\text{O})-\text{O}$ · each actual link releases one H_2O

Ideal repeat interior; finite-chain end groups omitted

The diol contributes $-\text{O}-\text{CH}_2-\text{CH}_2-\text{O}-$; the diacid contributes $-\text{C}(=\text{O})-(\text{CH}_2)_4-\text{C}(=\text{O})-$. Each actual link releases one water molecule. A finite open chain made from three diols and three diacids joins six molecules through five links, losing five waters. Ideal repeating-unit shorthand omits end groups; do not apply its water count blindly to a finite chain or a ring.

An **amino acid** carries an amino group and a carboxylic-acid group. Glycine is $\text{H}_2\text{N}-\text{CH}_2-\text{COOH}$. Condensation forms a **polypeptide**; different amino acids can occur within one chain to make proteins.



Each actual joining link releases one H_2O

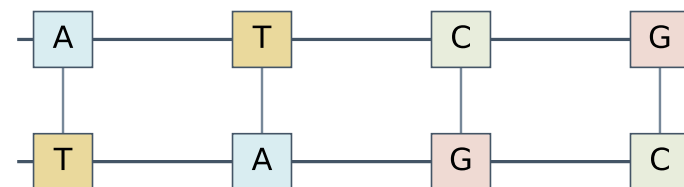
Repeat interior; finite-chain terminal groups omitted

Worked finite-chain check. Four glycine molecules in one open chain make three joining links and three water molecules. If supplied relative masses are glycine 75 and water 18, the chain's relative mass is $M_r = 4 \times 75 - 3 \times 18 = 246$. This calculation uses the supplied model; no additional mechanism is required.

Natural polymers and final checks (4.7.3.4)

Most DNA molecules have two polymer chains made from four types of **nucleotide**, arranged as a **double helix**. DNA encodes genetic instructions. A nucleotide is the whole monomer, not just its base label. Proteins have amino-acid monomers; starch and cellulose both have glucose monomers. Sharing a monomer does not make two polymers identical: their arrangements and roles can differ.

DNA: two chains · four nucleotide types · flattened before twisting



Each base-labelled block represents a whole nucleotide; not to molecular scale

Check yourself: explain physical separation versus cracking; balance a new combustion equation; draw a displayed alkene addition product without losing H; distinguish fermentation from steam hydration; recover a monomer from a repeat; and count actual links in a supplied finite condensation chain. Mark HT explanations separately from both-tier recognition.

8 · Chemical analysis

AQA · GCSE · Chemistry

Chemical analysis: choosing evidence

A bottle may say “pure milk”, yet milk contains several substances. A green flame may suggest copper, yet a second colour can be hidden in a mixture. **Chemical analysis** combines a suitable test, a recorded observation and a justified conclusion.

An observation is what you see; a conclusion is what that observation supports.

This topic covers specification 4.8. Both tiers study purity, formulations, chromatography and the four gas tests. All ion tests and instrumental methods here are **Separate Chemistry, both tiers**, including flame emission spectroscopy. Required practicals 6 and 7 need actual supervised practical work and recorded results; paper examples prepare you but do not certify practical completion.

Purity and designed mixtures (4.8.1.1–4.8.1.2)

A **pure substance** is a single element or compound, not mixed with another substance. Pure oxygen is an element; pure water is a compound, despite containing two chemically combined elements. Milk with nothing added is still a mixture. “Natural” and “chemically pure” answer different questions.

Pure elements and compounds melt and boil at specific temperatures under stated conditions. Compare a measured melting interval with a reference using calibrated apparatus and comparable conditions. A mixture often melts over a wider interval, and an impurity often lowers a solid’s melting point. Do not claim that every mixture has the same temperature pattern or that every impurity changes a boiling point in the same direction; pressure also affects boiling temperature.

Worked evidence. A reference melts at 80.0 °C. Sample A melts from 79.8 to 80.2 °C; B from 72.0 to 78.0 °C using the same apparatus. A’s narrow interval near the reference supports greater purity. B’s lower, broader interval supports impurity. A’s matching interval alone does not uniquely prove identity; another substance could have similar data.

A **formulation** is a mixture designed as a useful product. Each component has a purpose and its amount is carefully measured to achieve the required properties. Paint can contain pigment for colour and a binder to hold it on a surface. Fuels, cleaning agents, paints, medicines, alloys, fertilisers and foods can be formulations. Use supplied product information: there is no need to memorise proprietary ingredients. A mixture produced accidentally is not automatically a formulation.

Chromatography and a complete Rf method (4.8.1.3)

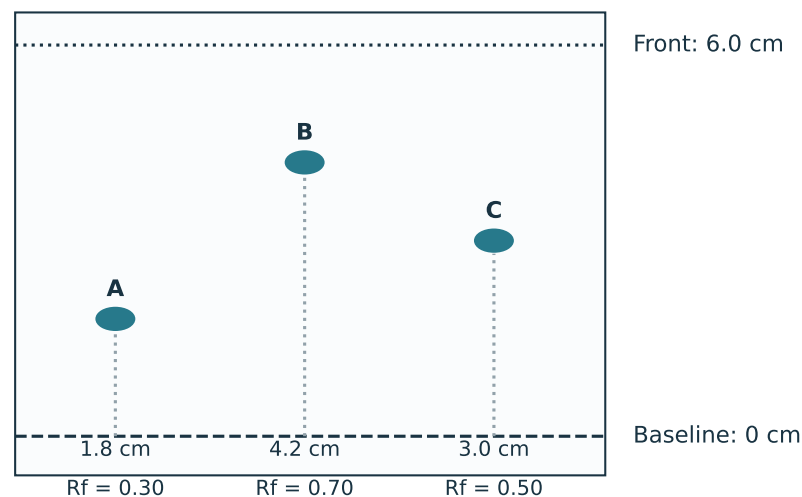
Chromatography separates a mixture because its substances distribute differently between a **stationary phase** and a **mobile phase**. In this GCSE paper model, paper is stationary and solvent moves through it. Substances carried more readily by the solvent travel farther; substances held more strongly by the stationary phase travel less. Travel distance is not decided by particle mass alone.

For required practical 6, use approved coloured substances and solvent under school supervision. Draw a pencil **baseline** above the solvent level; ink could dissolve and create extra spots. Place small concentrated sample/reference spots on the line and let them dry. Stand the paper in the solvent with the spots above its surface. Let solvent rise; remove the paper before the front reaches the top and mark the **solvent front** immediately, before evaporation hides it. Record the chromatogram and actual distances.

$$R_f = \frac{\text{distance from baseline to centre of spot}}{\text{distance from baseline to solvent front}}$$

Both distances use the same baseline and unit. **Rf value** has no unit. Measure to the centre of a spot, not its edge or from the bottom of the paper. A spot cannot travel beyond the solvent front in this model, so a value above 1 flags a measurement or division error.

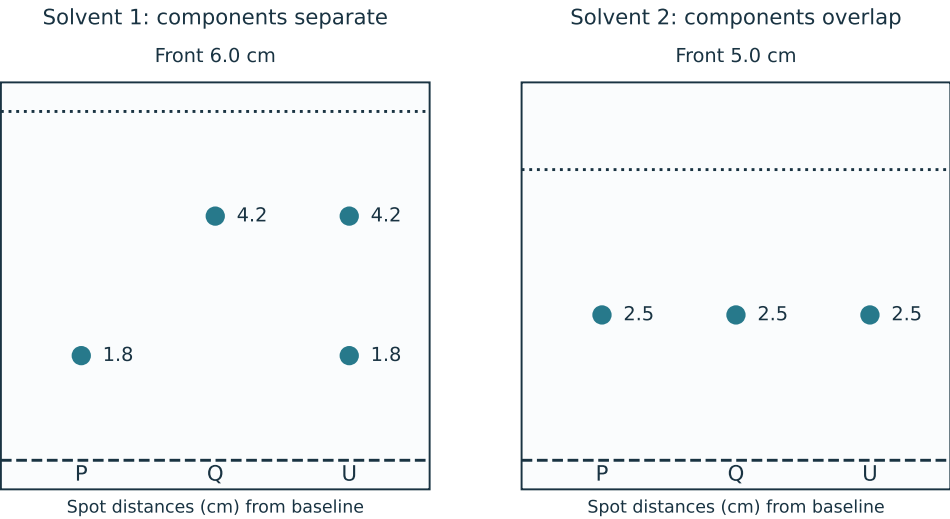
Distances measured from baseline to spot centre



Worked calculation. A spot travels 3.6 cm and the solvent front travels 6.0 cm:

$$R_f = \frac{3.6 \text{ cm}}{6.0 \text{ cm}} = 0.60$$

A matching spot/Rf with a known reference under the **same conditions** supports identification. Changing solvent can change Rf and whether substances separate. Two spots show at least two detected components. A pure compound gives one spot in all suitable solvents, but the reverse inference is limited: one observed spot can contain two substances that travel together, or miss an undetected component. Use another solvent or another test when the evidence is insufficient.



Transfer. In solvent 1, P has Rf 0.30, Q 0.70, and unknown U has both spots. In solvent 2, P and Q each have Rf 0.50, and U shows one spot at 0.50. Solvent 1 supports P and Q in U; solvent 2's single spot does not overturn that evidence or prove purity. These are fictional reference substances used to teach interpretation.

Four gas tests: pair the action and observation (4.8.2.1–4.8.2.4)

Gas	Test	Positive observation
hydrogen	Hold a burning splint at the open end of a test tube of gas.	A pop sound, often called a squeaky pop.
oxygen	Insert a glowing splint into the gas.	The glowing splint relights.
carbon dioxide	Bubble through or shake with limewater, aqueous calcium hydroxide.	Limewater turns milky/cloudy.
chlorine	Expose damp litmus paper to the gas in the approved setup.	Litmus is bleached white.

For oxygen, a splint that is already burning cannot demonstrate **relighting**. Let the flame go out while retaining a glow, then use the specified test. Do not claim this observation proves oxygen uniquely among every possible gas; it identifies oxygen

among the four gases in this syllabus context.

Carbon dioxide forms a white solid in limewater. The required positive observation is cloudiness; clearing on prolonged addition is not needed to identify it in this test. Chlorine gas and chloride ions need different tests. Litmus turning red alone is not the specified chlorine result: bleaching white is the key observation.

Worked decision. An unknown relights a glowing splint. Conclude oxygen from the positive test, rather than guessing from the reaction that produced it. A failed pop test alone does not identify which of the other gases is present. Use separate samples and the appropriate positive test. Actual gas tests use small approved samples and teacher-controlled apparatus; chlorine is handled only in the school's approved contained setup, never generated as an unsupervised exercise.

Metal ions: flames, precipitates and excess reagent (4.8.3.1–4.8.3.2)

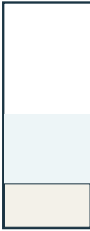
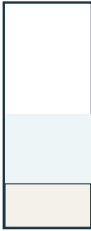
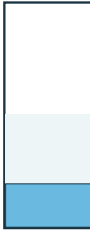
For a **flame test** , use the school's approved clean wire/reference procedure. Compare unknown and known samples, cleaning between tests to prevent contamination. Record the observed colour before naming the ion. A mixture can mask a colour, so an absent visible colour does not establish an absent ion.

Metal ion	Flame colour
lithium	crimson
sodium	yellow
potassium	lilac
calcium	orange-red
copper	green

Add sodium hydroxide solution to a fresh portion to look for an insoluble hydroxide **precipitate** . Then add excess sodium hydroxide when the white-precipitate distinction is needed.

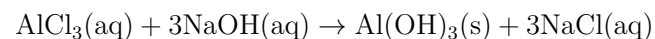
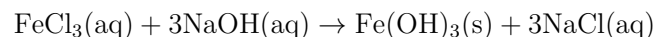
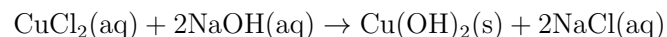
Cation	Initial hydroxide observation	In excess sodium hydroxide
Al ³⁺	white precipitate	dissolves
Ca ²⁺	white precipitate	remains
Mg ²⁺	white precipitate	remains
Cu ²⁺	blue precipitate	remains
Fe ²⁺	green precipitate	remains
Fe ³⁺	brown precipitate	remains

NaOH: initial precipitate colour; label states behaviour in excess

Al^{3+}	Ca^{2+}	Mg^{2+}	Cu^{2+}	Fe^{2+}	Fe^{3+}
					
Dissolves	Remains	Remains	Remains	Remains	Remains

Worked identification. A white precipitate alone leaves Al^{3+} , Ca^{2+} and Mg^{2+} possible. If it dissolves in excess NaOH, identify aluminium among these listed ions. If it remains, calcium and magnesium are not distinguished by this test alone; an orange-red flame supports calcium. Do not confuse copper's green flame with its blue hydroxide precipitate, or iron(II)'s green precipitate with iron(III)'s brown one.

Write balanced **full symbol equations** for the specified insoluble hydroxides. Balance the hydroxide groups and salt products; brackets show that each metal can need more than one OH group:



For CaCl_2 , MgCl_2 or FeCl_2 , the corresponding $\text{M}(\text{OH})_2$ needs 2NaOH and forms 2NaCl . These formation equations are both-tier knowledge here. An equation for the aluminium precipitate dissolving to sodium aluminate is not required. Ionic/half equations retain their separate Higher Tier scope elsewhere.

Anions: choose the right acid and fresh sample (4.8.3.3–4.8.3.5)

Ion	Reagents and procedure	Positive evidence
carbonate	Add dilute acid; test the gas with limewater.	Effervescence and gas turning limewater milky support carbonate.
chloride	Dilute nitric acid, then silver nitrate solution.	White silver chloride precipitate.
bromide	Dilute nitric acid, then silver nitrate solution.	Cream silver bromide precipitate.
iodide	Dilute nitric acid, then silver nitrate solution.	Yellow silver iodide precipitate.
sulfate	Dilute hydrochloric acid, then barium chloride solution.	White barium sulfate precipitate.

Use **fresh sample portions** for different tests, so a previous reagent does not add an ion that creates a false result. In the halide test, use nitric acid rather than hydrochloric acid, which would introduce chloride. Do not swap the acid/reagent pairs between the halide and sulfate tests. “A white precipitate” without its reagent is insufficient to distinguish these several possibilities.

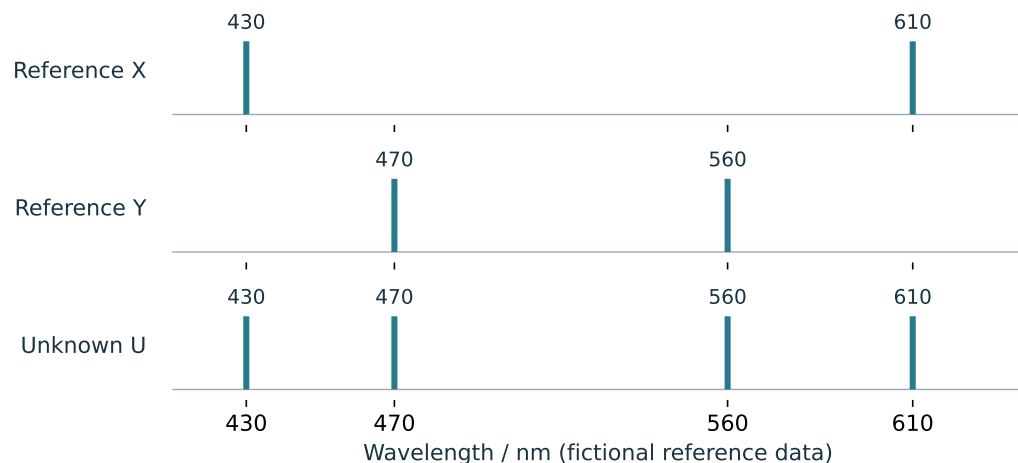
Worked combined result. Fresh portions of a soluble unknown give a yellow flame and a white precipitate with dilute nitric acid followed by silver nitrate. These observations support sodium ions and chloride ions respectively; sodium chloride is consistent with both if the task specifies one pure salt. The halide test does not identify sodium, and the flame test does not identify chloride. Record uncertainty if the sample may be a mixture.

Required practical 7 involves actual school-supervised flame and ion tests on known/unknown ionic compounds. Use approved small quantities, clean apparatus and fresh portions; record reagent, observation and inference separately. Class paper results and the tables here do not replace doing and documenting the practical.

Instrumental results and reference spectra (4.8.3.6–4.8.3.7)

An **instrumental method** can identify substances accurately, detect small amounts (**sensitivity**), and give rapid results. Accurate means close to the accepted value; sensitive means detecting small quantities. They describe different advantages. Instruments still need suitable references, calibration and controlled measurements.

Spectroscopy analyses light. In **flame emission spectroscopy**, put the sample into a flame and pass its emitted light through a spectroscope. The output is a **line spectrum**. Compare line positions with supplied reference spectra to identify metal ions. A mixture can contain lines from more than one metal. Use supplied intensity/concentration calibration to measure concentration; brightness without a calibrated comparison does not supply an absolute concentration.



Worked spectrum. Reference X has lines at 430 and 610 nm; Y has lines at 470 and 560 nm. The unknown has all four lines, supporting X and Y within these references. These fictional positions are given data, not wavelengths to memorise. A missing faint line can reflect the detection limit; do not assert that every absent line proves zero concentration.

Worked calibration. Under a stated linear calibration, 2.0 mg/dm³ gives intensity 20 and 6.0 mg/dm³ gives intensity 60. Intensity 35 lies 15/40 of the way between them:

$$c = 2.0 \text{ mg/dm}^3 + \frac{35 - 20}{60 - 20} \times (6.0 - 2.0) \text{ mg/dm}^3 = 3.5 \text{ mg/dm}^3$$

Stay inside the supplied calibrated range. This calculation uses an explicitly given model; do not assume every instrument is linear at every concentration. Choose an instrument when the task needs detection of a small amount, several metals or a measured concentration; choose a simple chemical test when an appropriate qualitative observation is sufficient.

9 · Chemistry of the atmosphere

AQA · GCSE · Chemistry

Chemistry of the atmosphere: explain the evidence

An engine can produce several invisible gases. A thermometer alone cannot tell us which gas caused a change. This topic links observations to chemical explanations: how air changed, how radiation transfers energy, and how the products of burning depend on the fuel and conditions.

Use a three-part answer: **source or action** → **chemical process** → **consequence**. Distinguish a measured result from an interpretation. The numbered sections below cover AQA 8462 4.9.1.1–4.9.3.2, both tiers. Earlier topics supply balanced equations, percentages, combustion and the carbon cycle; practical performance is separate.

Today's atmosphere: fractions and samples (4.9.1.1)

For about 200 million years, the main proportions have been much as today: roughly **four-fifths nitrogen** and **one-fifth oxygen**. Small amounts include noble gases, carbon dioxide and water vapour. The more precise familiar dry-air approximations are about 78% nitrogen and 21% oxygen; these do not imply that the remaining 1% is all carbon dioxide. Water-vapour content varies, so specify dry air if using a dry-air composition.

Worked sample. Estimate the nitrogen volume in 250 cm³ of air using the syllabus approximation. Assume all gas volumes refer to the same temperature and pressure.

$$V(\text{N}_2) = f(\text{N}_2)V(\text{air}) = \frac{4}{5} \times 250 \text{ cm}^3 = 200 \text{ cm}^3$$

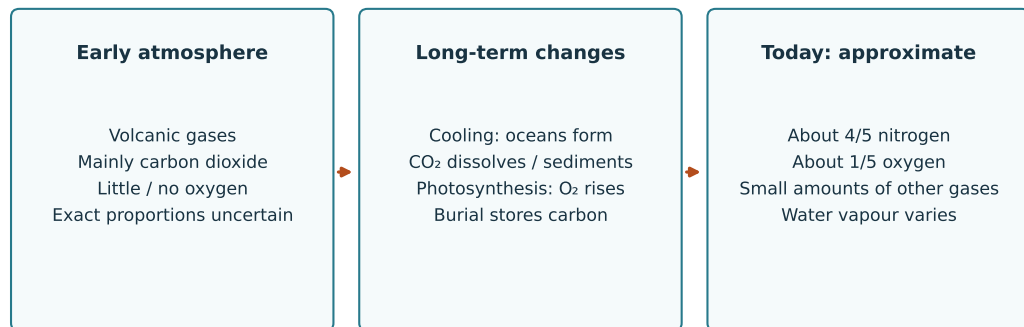
The matching oxygen estimate is $\frac{1}{5} \times 250 = 50 \text{ cm}^3$. These rough fractions deliberately ignore the small components; do not combine 80% nitrogen with a separate exact 21% oxygen as though they were one consistent dataset.

An early atmosphere: a theory with limited evidence (4.9.1.2)

Earth formed about 4.6 billion years ago. Direct atmospheric measurements from that time do not exist; geological material has also changed over this long time. Evidence for the earliest atmosphere is limited, so theories develop when evidence improves. Ancient ice records later atmospheres, not a sample of the air from Earth's formation.

One theory is that intense **volcanic activity** during the first billion years released the gases forming the early atmosphere. It was mainly carbon dioxide, with little or no oxygen, possibly like the present atmospheres of Mars and Venus. Volcanoes also released nitrogen, water vapour and small amounts of methane and ammonia. A comparison with another planet supports a possible explanation; it is not proof that

Earth's gas percentages were identical.



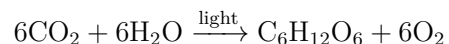
Qualitative sequence — boxes are not measured gas fractions or a time scale

Condensation changes water vapour into liquid. As Earth cooled, water vapour **condensed** into liquid water, forming oceans. Carbon dioxide dissolved in the water; carbonates precipitated and formed sediments, removing carbon dioxide from the atmosphere. Nitrogen gradually built up. Do not say cooling changed water into carbon dioxide or that volcanoes supplied the later oxygen rise.

Evidence judgement. A modern volcano releases carbon dioxide, nitrogen and steam. This makes a volcanic origin plausible, but it does not measure the exact composition of the atmosphere billions of years ago. State both the supporting observation and the limitation.

How oxygen increased (4.9.1.3)

Algae first produced oxygen by **photosynthesis** about 2.7 billion years ago. Plants evolved later. Over the next billion years, oxygen gradually increased to a level that enabled animals to evolve. This was a long change, not an instant jump to today's percentage.



Carbon dioxide + water → glucose + oxygen, using light energy. Count six carbon, twelve hydrogen and eighteen oxygen atoms on each side. Photosynthesis removes carbon dioxide and releases oxygen; respiration has the opposite overall gas effect. A plant does both, so consider the balance of processes when discussing net change.

How carbon dioxide decreased: rocks and fuels (4.9.1.4)

Several processes removed carbon dioxide. Dissolving in oceans and forming carbonate sediments began before plants. Later, algae and plants used carbon dioxide to make biomass by photosynthesis. Burial kept some carbon out of the atmosphere for very

long periods.

- **Limestone** forms from accumulated calcium carbonate sediments, including shells and skeletons of marine organisms, which become rock. Carbon in this carbonate is stored outside the atmosphere.
- **Coal** forms from buried plant material, compressed and changed over millions of years under conditions that limit complete decay.
- **Crude oil** and **natural gas** form mainly from the buried remains of small marine organisms, changed by heat and pressure over millions of years.

Coal, crude oil and natural gas are **fossil fuels**. They do not form from carbon dioxide simply freezing or dissolving. Burning these stores returns carbon mainly as carbon dioxide. Rock storage is long-term, not literally permanent: geological processes can return carbon too.

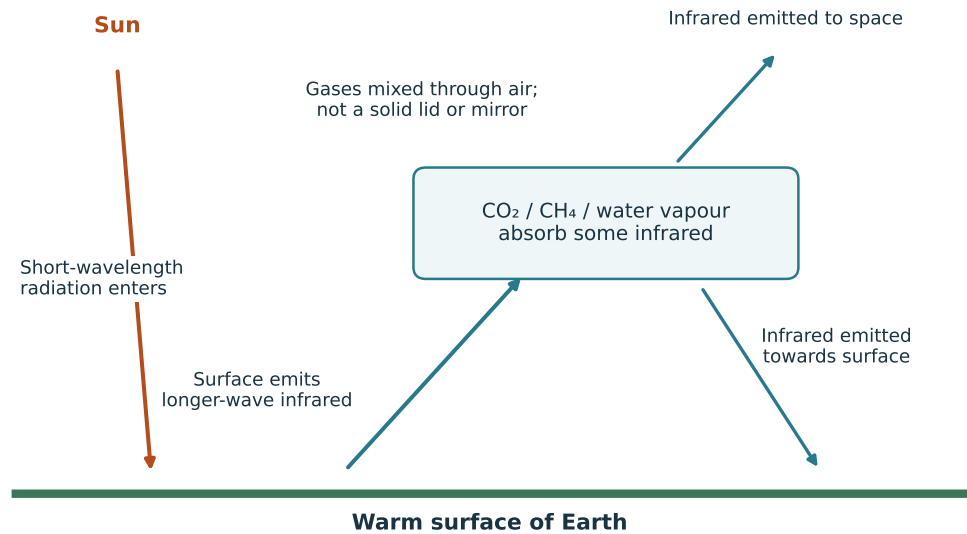
Worked explanation. Why can both limestone formation and coal formation lower atmospheric carbon dioxide? Limestone stores carbon in carbonate rock after removal from water/air. Plants first remove carbon dioxide by photosynthesis; burial of their remains then stores that carbon in coal rather than returning all of it by decay. Name the route as well as the final store.

Greenhouse gases and energy transfer (4.9.2.1)

A **greenhouse gas** absorbs and emits infrared radiation. Water vapour, carbon dioxide and methane are greenhouse gases. The natural **greenhouse effect** keeps Earth's surface warm enough for life; the question about human influence concerns an increase in this warming.

Short-wavelength radiation from the Sun passes through the atmosphere and warms the surface. The warm surface emits longer-wavelength **infrared radiation**.

Greenhouse gases absorb some of this outgoing infrared and emit infrared in all directions, including towards the surface. Some radiation escapes to space. Increased greenhouse-gas amounts can reduce the initial rate of energy loss for the same surface temperature, so the surface warms towards a new energy balance.



The gases are mixed through the air, not a solid lid. Absorption and emission are not mirror reflection. Do not describe the ozone hole as the cause, or say that greenhouse gases stop all radiation leaving Earth.

Human activities and judging climate evidence (4.9.2.2)

Recall at least two activities for **each** gas. Burning fossil fuels increases carbon dioxide. Deforestation reduces carbon dioxide removal by photosynthesis and may release stored carbon through burning or decay. Methane increases through cattle farming and decay of organic waste in landfill; flooded rice fields can also release methane.

When reading a climate report, check the measured variable, units, dates, geographical coverage and whether the graph shows a trend or just one unusual year. A relationship between two graphs is evidence to examine, not by itself a complete causal explanation. Mechanism measurements and several independent datasets strengthen an explanation.

Peer review means other specialists examine a study's methods, evidence and reasoning before publication. It can reveal weaknesses; it does not guarantee every conclusion is correct. Communicating results clearly helps people judge the evidence. A headline based on one selected period or an unexplained graph may be biased. Complex climate systems require simplified models, which have assumptions and uncertainties; uncertainty about the exact size of an effect is not evidence that no effect exists.

Original data example. A fictional classroom dataset gives a global mean temperature anomaly of 0.30°C in period A and 0.70°C in period B. This is a difference from a stated reference average, not Earth's absolute temperature.

$$\Delta T = T_B - T_A = 0.70 - 0.30 = 0.40 \text{ }^{\circ}\text{C}$$

This supports warming between these two period means. It does not establish a temperature rise in every place or every year, and does not alone identify the cause. Compare uncertainties and other evidence before a broader conclusion.

Climate effects: scale, risk and uncertainty (4.9.2.3)

Global climate change involves long-term patterns, not today's weather. Four potential effects are rising sea levels, changing rainfall, changes in the frequency or severity of extreme weather, and changes in habitats/species distributions.

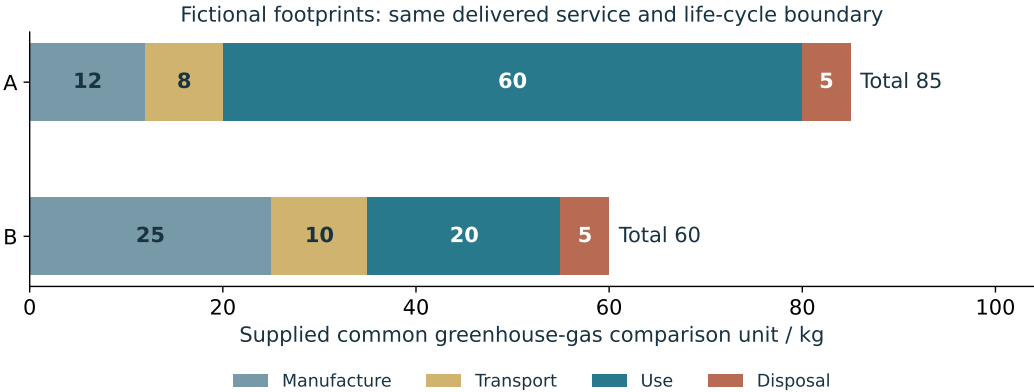
Link each effect to a consequence. Rising seas can flood low coastal land; changing rainfall can cause water shortages or damage crops; severe events can harm people and infrastructure; shifting habitats can threaten organisms unable to move or adapt. Warming can raise sea level through thermal expansion of seawater and melting land ice. Melting floating sea ice does not have the same direct sea-level effect as adding meltwater from land.

Discuss **scale** (how many people/which area and over what time), **risk** (likelihood and seriousness), and environmental implications rather than listing disasters. Exact future regional effects depend on emissions, model assumptions and complex feedbacks. Give a conditional prediction supported by the supplied evidence, not a guarantee that every location will become hotter and drier.

Carbon footprints: totals, choices and limits (4.9.2.4)

A **carbon footprint** is the total carbon dioxide and other greenhouse gases emitted over the full life cycle of a product, service or event. Include raw materials, manufacture, transport, use and disposal where they apply. A footprint is not just carbon dioxide from using an object and is not the mass of carbon atoms in it. When a problem supplies all gases in a common comparison unit, use that unit consistently; conversion factors are not required unless given.

Compare the same service, such as 100 journeys of the same distance, over the same lifetime and **system boundary**. A product with greater manufacturing emissions can still have lower total emissions after repeated use.



Worked comparison, fictional supplied data. Both options deliver the same service. A has 12 kg manufacture + 8 kg transport + 60 kg use + 5 kg disposal; B has 25 + 10 + 20 + 5 kg in the same supplied greenhouse-gas comparison unit.

$$F_A = 12 + 8 + 60 + 5 = 85 \text{ kg}, \quad F_B = 25 + 10 + 20 + 5 = 60 \text{ kg}$$

$$\text{reduction} = \frac{F_A - F_B}{F_A} \times 100\% = \frac{25}{85} \times 100\% \approx 29.4\%$$

B has the lower total despite higher manufacturing emissions. This conclusion depends on the supplied use stage and equal service; it does not prove B has less water demand or less waste.

Reduce carbon dioxide through efficient equipment, insulation, reduced fossil-fuel use, suitable lower-emission energy sources or protecting forests. Reduce methane through preventing organic waste, capturing landfill gas and changing suitable farming practices. Explain limitations: initial cost, available infrastructure, energy supply, land needs, competing food demands or public acceptance. Captured methane still releases carbon dioxide if burned, but prevents direct methane release and can supply useful energy. No single action guarantees zero footprint.

Predict combustion products from fuel and conditions (4.9.3.1)

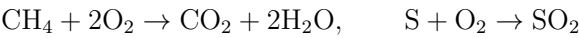
Most fuels contain carbon and/or hydrogen and may contain sulfur. First inspect the fuel’s elements; then inspect oxygen supply and temperature. Complete combustion of a **pure hydrocarbon** in sufficient oxygen produces carbon dioxide and water. This statement does not mean every fuel, under every condition, produces only these two gases.

With limited oxygen, **incomplete combustion** can produce carbon monoxide and carbon particles (soot), as well as some carbon dioxide and water.

Particulates include solid particles and unburned hydrocarbons released to air. Burning sulfur impurities forms sulfur dioxide. At high temperatures, nitrogen

and oxygen in air can react to form oxides of nitrogen, even when the fuel contains no nitrogen.

Fuel/condition supplied	Supported prediction and reason
Pure hydrogen, sufficient oxygen	Water; no carbon in the fuel to form carbon dioxide or soot
Pure methane, sufficient oxygen	Carbon dioxide and water; complete hydrocarbon combustion
Methane, limited oxygen	Carbon monoxide and/or soot possible; not enough oxygen for full oxidation
Carbon fuel containing sulfur	Carbon oxides and sulfur dioxide; sulfur is oxidised
High-temperature engine using sulfur-free fuel	Nitrogen oxides still possible from nitrogen/oxygen in air



Worked prediction. Removing sulfur from an engine fuel reduces sulfur dioxide; it does not automatically remove nitrogen oxides because their nitrogen can come from air. More oxygen may reduce incomplete-combustion products, but does not remove carbon dioxide from complete combustion.

Pollutant properties, effects and appropriate action (4.9.3.2)

Pollutant	Important property/effect	Action linked to its source
Carbon monoxide, CO	Toxic, colourless and odourless; reduces blood’s oxygen transport	Improve combustion/ventilation and use appropriate detection; smell cannot identify it
Sulfur dioxide, SO ₂	Respiratory problems and acid rain	Remove sulfur or treat exhaust gases with a suitable alkali
Oxides of nitrogen, NO _x	Respiratory problems and acid rain	Reduce formation or use suitable exhaust treatment
Particulates	Health/respiratory problems and global dimming	Improve combustion and use suitable particle filters

Acid rain damages ecosystems and reacts with some buildings and stone.
Global dimming is reduced sunlight reaching the surface because of par-

tibles; it is not the same process as greenhouse-gas absorption of outgoing infrared. A filter that catches particles does not necessarily remove dissolved/gaseous carbon monoxide. Distinguish CO, the poisonous incomplete-combustion gas, from CO₂, a greenhouse gas.

When an exhaust treatment question gives a reagent, use its stated reaction. Calcium oxide can neutralise sulfur dioxide to calcium sulfite; further oxidation can produce sulfate. Do not silently write calcium sulfate as the product of CaO + SO₂ alone: an extra oxygen source is required. A suitable catalytic converter can use carbon monoxide to reduce nitrogen monoxide to nitrogen while forming carbon dioxide; it does not simply split every nitrogen oxide into nitrogen and oxygen without other chemistry.

Independent transfer check

A town replaces sulfur-containing fuel with sulfur-free fuel and fits particle filters. Its report claims “all climate and breathing problems are solved”. Explain the two changes’ likely benefits, then reject the overclaim using the remaining carbon dioxide, possible carbon monoxide and high-temperature nitrogen oxides. State what measurements would be needed before judging the size of any improvement. Use the independent questions in sheets 9.1–9.3 to check the full set of methods.

10 • Using resources

AQA • GCSE • Chemistry

Choose a resource route by the job it must do

A town needs safe water, a bridge that resists corrosion and fertiliser for farms. The cheapest raw material is not automatically the best choice: treatment, lifetime, energy and waste all matter. Compare the same useful service, state the conditions and explain the chemistry behind a decision.

This reference covers 4.10.1.1–4.10.4.2 for Separate Chemistry. Alternative biological metal extraction and equilibrium/rate evaluation of the Haber process are **Higher tier (HT)**. Written RP8 questions assess knowledge; actual supervised analysis and purification require a practical record.

Resources, replacement and sustainable development (4.10.1.1)

Earth’s resources provide warmth, shelter, food and transport. Agriculture supplements natural supplies of food, timber and fibres. Natural wool/cotton can be supplemented or replaced by synthetic fibres such as polyester; a synthetic product is not automatically more sustainable.

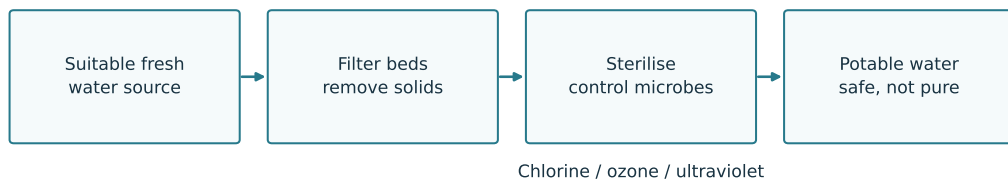
A **finite resource** has a limited supply that is not replaced on the time scale of use, such as a metal ore or fossil fuel. A **renewable resource** can be replenished on a useful time scale, such as sustainably managed timber. Renewable does not mean unlimited: cutting trees faster than they grow still reduces the stock.

Sustainable development meets current needs without compromising future generations’ ability to meet theirs. Chemistry can improve crop production, material lifetime and industrial efficiency. Evaluate an improvement using evidence, including its costs and wastes.

Worked scale comparison. Fictional resource stocks are 20 tonnes and 20 000 tonnes. The ratio is $20\,000/20 = 1000 = 10^3$: three orders of magnitude. Check common units first. A large stock alone does not establish sustainability; extraction rate and replenishment also matter.

Potable water: choose the treatment for the source (4.10.1.2)

Potable water is safe to drink, with sufficiently low dissolved salts and harmful microbes. It is not chemically pure water because it contains dissolved substances. Clear-looking water may still contain salts, microbes or dissolved chemicals.



Salt water → distillation or reverse osmosis → salt removal, with energy input

In the UK, rain supplies fresh water collecting in the ground, lakes and rivers. Choose an appropriate source, pass water through filter beds to remove suspended solids, then **sterilise** using chlorine, ozone or ultraviolet light. Filtration does not remove all dissolved salts; sterilising controls microbes rather than removing salts.

If fresh supplies are limited, **desalination** removes salt from sea/salt water. Distillation boils water and condenses its vapour, leaving most dissolved salts behind. **Reverse osmosis** uses pressure to drive water through a suitable membrane while retaining salts. Both need substantial energy: heating/condensing or producing pressure. Select a route using the local source, energy supply and required quality, rather than treating ordinary filtration as desalination.

RP8 knowledge. Compare water samples using pH and dissolved-solids measurements, then investigate distillation under teacher supervision. For residue, measure a known sample volume, weigh a dry container, evaporate water safely, cool and reweigh; repeat heating/cooling to constant mass where appropriate. Residue mass is final minus empty-container mass. Distillation needs collection and cooling of vapour, not just boiling a sample dry. Never drink laboratory samples or assume one measurement certifies safety.

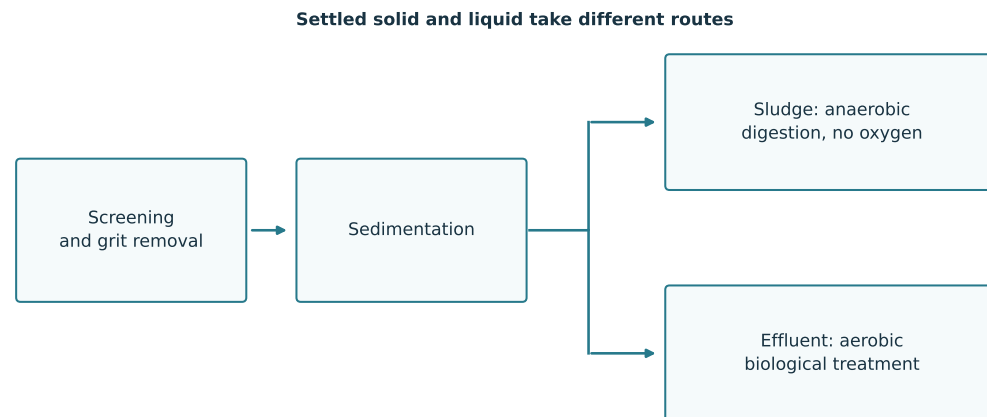
Worked residue. A 50.0 cm^3 sample leaves 0.120 g solid. Convert $50.0 \text{ cm}^3 = 0.0500 \text{ dm}^3$, so concentration = $0.120/0.0500 = 2.40 \text{ g/dm}^3$. This measures residue under the stated method, not microbial safety.

Waste water follows two treatment branches (4.10.1.3)

Sewage and agricultural waste water need removal of organic matter and harmful microbes. Industrial waste can also contain harmful chemicals requiring suitable additional treatment. Treatment before environmental release does not automatically make the water potable.

Screening removes large objects and grit removal takes out heavy grit. **Sedimentation** separates settled **sludge** from liquid **effluent**. These then take different routes: **anaerobic digestion** treats sludge without oxygen; **aerobic biological treatment** uses oxygen to treat organic matter in

the effluent. Do not send both products through one imagined serial chain.



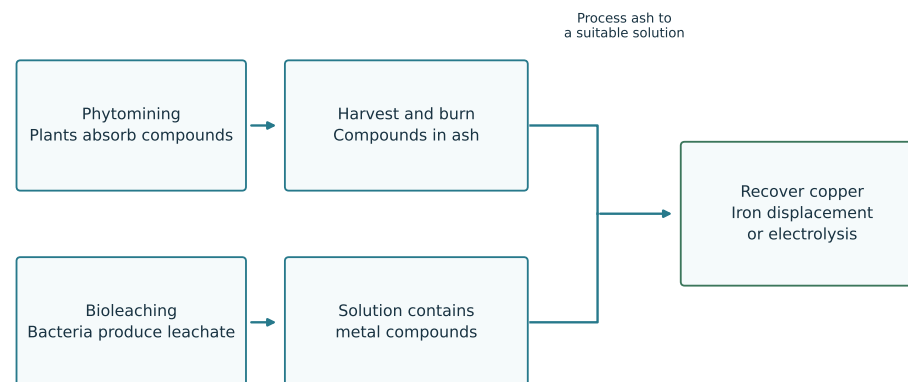
Industrial chemicals may need additional treatment

For comparable local conditions, suitable ground water usually needs less treatment than waste water; salt water needs energy-intensive salt removal. This is a reasoned comparison, not an absolute ranking for every contaminated source. Check dissolved substances, organic matter, microbes and available processes.

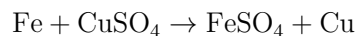
Recover copper from low-grade sources (4.10.1.4, HT only)

Copper ores are finite and high-grade supplies become scarcer. **Phytomining** uses plants to absorb metal compounds. Harvest and burn the plants; their ash contains those compounds. Further processing is needed: the plant has not produced a piece of pure copper.

Bioleaching uses bacteria to produce a **leachate** solution containing metal compounds. Copper can be recovered from a suitable solution by displacement with scrap iron or by electrolysis. Both routes need a recovery stage after obtaining the metal-containing material.



HT: obtaining a metal compound is not the same as obtaining the pure metal



Iron is more reactive than copper and displaces it. Biological routes can avoid digging, moving and disposing of very large amounts of rock, but compare processing time, yield, energy, land use and waste using supplied information. “Biological” does not mean free, instant or impact-free.

Life cycle assessment: compare an equal service (4.10.2.1)

A **life cycle assessment** (LCA) considers extraction/processing of raw materials; manufacture/packaging; use/operation; and disposal at the end of useful life. Include transport and distribution at each stage. Compare the same service and lifetime, such as carrying the same shopping load the same number of times.

Water, resources, energy and some wastes can be quantified from suitable data. Measurement and assumptions still have uncertainty. Combining different pollutant effects into one score requires value judgements; an LCA is not purely objective. A disposal-only comparison may be useful for that stage but cannot establish the best whole-life product.

Worked equal-service decision. Fictional paper-bag figures for a shopping task total $6 + 4 + 2 = 12$ energy units. A reusable plastic bag needs 30 units to make and 1 unit per task to clean. After n tasks, paper needs $12n$ and plastic $30 + n$. Equality gives $30 = 11n$, so $n = 2.73$; the reusable option uses less energy from the third complete task under these assumptions. This energy result does not also prove lower water use, litter or every pollutant effect. Record each category and the real reuse rate.

Reduce, reuse and recycle with a material reason (4.10.2.2)

Reducing material use can lower extraction, processing and waste. Reusing a whole glass bottle after collection and washing avoids remaking it; recycling crushes and remelts glass into another product. Metals can be melted and recast or reformed. Some scrap steel can be added to iron from a blast furnace, reducing the amount extracted from ore.

Separation requirements depend on the required final properties. Mixed materials and contamination may require more sorting; collection, transport, washing and remelting still use resources. Recyclable describes a possible route, not proof that collection and processing actually happen. Compare those burdens with avoided mining/quarrying and processing using supplied data.

Transfer. If recycling a stated batch uses 8 MJ but new production uses 50 MJ for the same material and quality, the supplied saving is $50 - 8 = 42$ MJ, or $42/50 \times 100 = 84\%$. Check whether collection was included before using that number in a claim.

Corrosion: barrier or sacrificial protection? (4.10.3.1)

Corrosion destroys a material through chemical reactions with its environment. Iron rusts only when both water and oxygen are available. In a comparison, ordinary damp air allows rusting; dry air lacks water; boiled water covered with oil limits oxygen entry. Keep the iron samples and exposure conditions comparable. One damp nail without controls cannot show that both substances are necessary.



Keep nail material, exposure time and temperature comparable. Schematic, not to scale.

Paint, grease and suitable electroplating form barriers. A scratch can expose iron to oxygen and water. Zinc galvanising adds **sacrificial protection**: zinc is more reactive than iron and oxidises preferentially, so nearby iron can remain protected at a scratch while suitable zinc contact remains. A less reactive coating cannot give that same protection. Aluminium’s thin adherent oxide layer protects its surface; it is not evidence that aluminium never reacts.

Match alloys to properties and uses (4.10.3.2)

Most everyday metals are alloys. Different atom sizes can disrupt regular metal layers and make sliding harder; do not claim every **alloy** has identical properties.

Alloy	Composition or feature	Suitable use and reason
Bronze	Copper and tin	Statues: useful strength and durability
Brass	Copper and zinc	Fittings/instruments: useful working and corrosion properties
Gold jewellery	Gold with silver/copper/zinc	Harder wearing than pure soft gold
High-carbon steel	Iron with a higher carbon content	Hard tools; strength but brittleness limits some uses
Low-carbon steel	Iron with a lower carbon content	Shaped car-body panels; softer and easier to form
Stainless steel	Iron with chromium and nickel	Cutlery: hard and corrosion resistant
Aluminium alloys	Low density	Aircraft structures where low mass matters

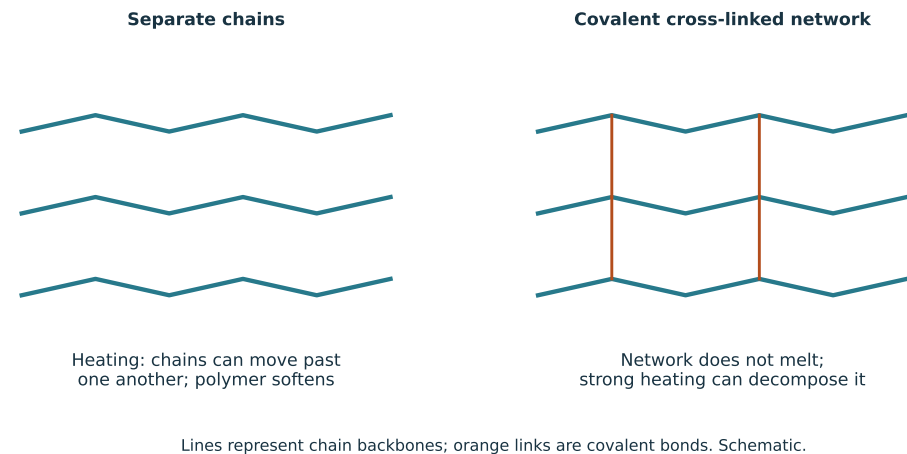
Gold purity is measured in carats: 24 **carat** is pure gold, and 18 carat is $18/24 = 75\%$ gold by mass. A 12 g ring then contains $12 \times 18/24 = 9$ g gold and 3 g other metals. For an unfamiliar alloy, use its supplied composition/property data rather than guessing from its name.

Ceramics, polymers and composites (4.10.3.3)

Soda-lime glass is made by heating sand, sodium carbonate and limestone. Borosilicate glass uses sand and boron trioxide and melts at a higher temperature. Clay ceramics, such as bricks and pottery, are made by shaping wet clay and heating it in a furnace. Hardness, brittleness and heat resistance affect the job each can do.

Low-density and high-density poly(ethene) both come from ethene. Different production conditions lead to different chain structures: branching makes close packing harder in LD material; straighter HD chains pack more closely. The same monomer does not force the same density or flexibility.

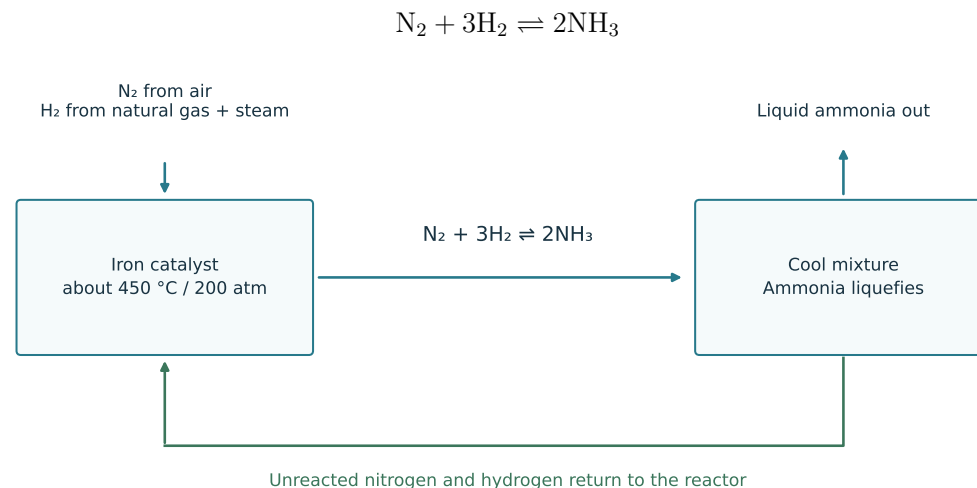
A **thermosoftening polymer** has separate chains without the cross-linked network considered here. Heating lets chains move past one another as inter-molecular attractions are overcome, so it softens/melts and can be remoulded. Melting does not require breaking the chain's covalent backbone. A **thermosetting polymer** has covalent cross-links joining chains into a network; it does not melt on heating, and sufficiently strong heating can decompose it. Remelting is therefore not a suitable recycling route for that network.



A **composite** has a **matrix** /binder surrounding and binding a **reinforcement** of fibres or fragments. Examples include glass fibres in resin, carbon fibres in polymer resin, and concrete with aggregate in a cement-based binder. “Carbon fibre” alone does not identify both components. Select materials by the required combination of strength, density, flexibility, temperature resistance and cost, using quantitative data when supplied.

Haber manufacture and the HT compromise (4.10.4.1)

The **Haber process** manufactures ammonia for nitrogen fertilisers. Nitrogen comes from air; hydrogen can be obtained from natural gas reacted with steam. Purified gases pass over iron at about 450 °C and about 200 atmospheres.



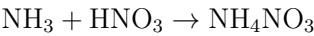
The reaction is reversible. Cooling the mixture liquefies ammonia for removal; unreacted nitrogen and hydrogen return to the reactor. Recycling improves use of feedstock, but does not prove zero waste or zero energy demand. These process steps are required at both tiers.

HT reasoning. The forward reaction is exothermic, so lower temperature favours equilibrium ammonia yield but slows reaction. Higher temperature raises rate but reduces equilibrium yield. Higher pressure favours ammonia because four reactant gas molecules become two product molecules, but strong equipment and compression cost money and energy. The industrial conditions balance rate, yield and cost; they are not the conditions for maximum possible equilibrium yield. Iron speeds both directions and approach to equilibrium without changing its position.

Worked feed check. With 2 mol nitrogen and 6 mol hydrogen, the theoretical maximum is 4 mol ammonia if all reactants converted. If only 1 mol nitrogen reacts on a pass, 3 mol hydrogen reacts and 2 mol ammonia forms; 1 mol nitrogen and 3 mol hydrogen remain for recycling. The per-pass conversion is not the same as overall conversion after repeated recycling.

NPK fertilisers: choose salts and compare manufacture (4.10.4.2)

An **NPK fertiliser** is a formulation of salts supplying nitrogen, phosphorus and potassium for plant growth. Ammonia can make ammonium salts and nitric acid. For example, ammonia neutralises nitric acid:



Potassium chloride, potassium sulfate and phosphate rock are obtained by mining. Phosphate rock is not directly suitable as fertiliser; treatment makes usable soluble salts. Recall these products:

Acid treatment of phosphate rock	Products to recognise
Nitric acid	Calcium nitrate and phosphoric acid
Sulfuric acid	Single superphosphate: calcium phosphate fertiliser with calcium sulfate
Phosphoric acid	Triple superphosphate: calcium phosphate fertiliser without the sulfate component

The phosphate fertiliser in single/triple superphosphate is calcium dihydrogen phosphate. Do not label every acid treatment simply “superphosphate”, and do not confuse phosphoric acid with a potassium source.

For the same salt, compare laboratory and industrial routes using supplied information. A careful laboratory neutralisation/titration followed by evaporation and crystallisa-

tion can produce a pure solid: a titration is not inherently an impure product. Industry can integrate ammonia/acid manufacture, control continuous large-scale output and recover heat, while requiring equipment, energy and emissions management. Explain the actual trade-off; “industry is always better” is not an evaluation.

Independent transfer check

A school wants drinking water, a reusable container and fertiliser. Explain why filtration alone cannot desalinate seawater, why a reusable product needs an equal-service LCA, and why fertiliser manufacture still consumes resources when Haber gases are recycled. Then compare scratched paint with scratched zinc on iron. State the missing data or condition for each decision rather than claiming one option solves every problem.