Course notes

AQA · GCSE · Chemistry

AQA GCSE · Chemistry: teaching notes

Version: 8462; Version 1.1 (4 October 2019); exams 2018 onwards

These notes cover selected topics of the course; they are not the whole specification.

Assessment and course boundaries

- Foundation and Higher routes must retain Chemistry-only and HT-only statements.
- Paper 1: 4.1–4.5; Paper 2: 4.6–4.10. Each paper is 100 marks, 1 h 45 min, 50%.
- Moles, quantitative titration and equilibrium depth are gated by tier; practical experience is not a substitute written assessment.

Atoms, elements and compounds: read the chemical symbols

Official-unit focus: 4.1 Atomic structure and the periodic table

O is the symbol for oxygen, while O_2 represents an oxygen molecule. Two atoms joined together do not necessarily make a compound: their element identities matter.

An atom is the smallest part of an element that can exist. An element contains one type of atom, defined by proton number. A compound contains two or more elements chemically combined in fixed proportions. Water, H_2O, contains hydrogen and oxygen in a 2:1 atom ratio; oxygen, O_2, is still an element. A chemical reaction forms new substances and often produces a detectable energy change. Separating a compound into its elements requires chemical change.

O_2: one element	H_2O: two elements	$Ca(OH)_2$: 1 Ca, 2 O, 2 H	Co and CO differ

Symbols use an uppercase first letter and, where present, a lowercase second letter: Co is cobalt, whereas CO is carbon monoxide. For the first twenty elements learn H hydrogen, He helium, Li lithium, Be beryllium, B boron, C carbon, N nitrogen, O

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oxygen, F fluorine, Ne neon, Na sodium, Mg magnesium, Al aluminium, Si silicon, P phosphorus, S sulfur, Cl chlorine, Ar argon, K potassium and Ca calcium. Group 1 also includes Rb rubidium, Cs caesium and Fr francium; Group 7 includes F fluorine, Cl chlorine, Br bromine, I iodine and At astatine. Use the supplied periodic table for other specified elements, including Fe iron, Cu copper and Zn zinc. Very radioactive francium and astatine are names to interpret from the table, not school reaction specimens.

Read a formula by identifying each element symbol before applying its subscript. A subscript applies to the symbol immediately before it, or to an entire bracketed group. Name MgO magnesium oxide, NaCl sodium chloride and $Ca(OH)_2$ calcium hydroxide. Compare diagrams of separate atoms, molecules of one element and compound particles; keep particle identity distinct from the number of particles.

■ Checked worked case

Known: one formula unit of $Ca(OH)_2$ contains one Ca, two O and two H, five atoms in all. Three formula units contain 3 calcium, 6 oxygen and 6 hydrogen atoms, fifteen total. The coefficient multiplies the entire formula; it does not change the fixed Ca:O:H ratio of 1:2:2.

■ Common error

A molecule can contain only one element, and some compounds form lattices rather than separate molecules. Heating a mixture to separate it is not automatically a chemical reaction. New substances, rather than an energy change alone, establish chemical change.

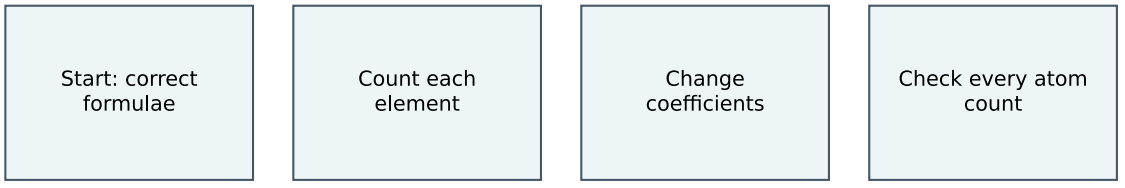
Balance equations by conserving each element

Official-unit focus: 4.1 Atomic structure and the periodic table

Magnesium burns in oxygen to form magnesium oxide. Writing $Mg + O_2 \rightarrow MgO$ names the substances correctly, but it does not yet account for every atom.

A word equation gives reactant and product names: magnesium + oxygen → magnesium oxide. A symbol equation uses correct formulae. Balance it as $2Mg + O_2 \rightarrow 2MgO$: both sides now contain two magnesium atoms and two oxygen atoms. Atoms are rearranged in a reaction, not created or destroyed. The numbers before formulae are coefficients describing relative particle numbers.

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Begin with the actual substance formulae and make an element-count table. Change coefficients to equalize the count for each element. Never change MgO to MgO_2 just to balance oxygen: that would identify a different substance. For hydrogen burning, $2H_2 + O_2 \rightarrow 2H_2O$ conserves four hydrogen atoms and two oxygen atoms. State symbols, when required, distinguish solid (s), liquid (l), gas (g) and aqueous solution (aq).

Use coloured counters to represent atoms and construct the reactants from the stated formulae. Rearrange the same counters into product particles, then record the smallest whole-number coefficients. Confirm each element separately. Models count atoms and bonds schematically; they do not show the detailed collision mechanism or actual particle size.

■ Checked worked case

Known: aluminium reacts with oxygen to form Al_2O_3. Six oxygen atoms are supplied by $3O_2$ and appear in $2Al_2O_3$. Those product units contain four aluminium atoms, so $4Al + 3O_2 \rightarrow 2Al_2O_3$. Final count: Al 4→4 and O 6→6. The coefficients 4:3:2 are a particle ratio, not a ratio of masses in grams.

■ Common error

A balanced equation cannot be obtained by changing a product formula or deleting an element. A coefficient of one is normally omitted. Equal atom counts do not mean equal numbers of molecules, and a balanced equation alone does not establish that a reaction occurs under all conditions.

Higher Tier: conserve charge in ionic and half equations

Higher Tier focus; see section scope notes.

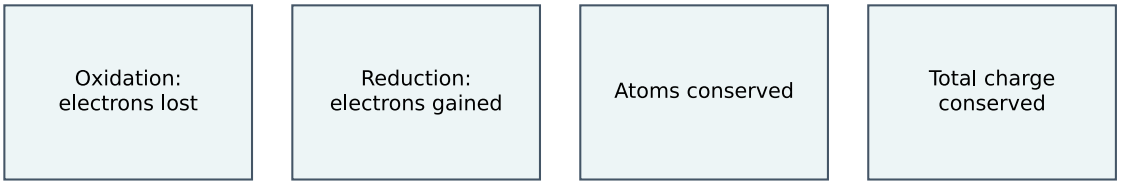
Official-unit focus: 4.1 Atomic structure and the periodic table

An equation can conserve atoms while failing to conserve electrical charge. Ionic reactions need both checks, and a half equation explicitly shows electron transfer.

An ionic equation represents reacting ions and omits spectator ions that remain un-

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changed. For aqueous silver nitrate and sodium chloride, $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ describes the precipitate; Na^+ and NO_3^- remain in solution. A half equation shows electron loss or gain: $Mg \rightarrow Mg^{2+} + 2e^-$ is oxidation, while $Cu^{2+} + 2e^- \rightarrow Cu$ is reduction. Electrons carry negative charge.



In the magnesium half equation, the right-hand total charge is $+2 - 2 = 0$, matching neutral Mg. In copper reduction, $+2 - 2 = 0$ on the left, matching neutral copper. Neutralization can be represented as $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$; the net charge is zero on each side. Electron numbers must match when combining half equations so no free electrons remain in the overall reaction.

First write the relevant ions and products, then balance atoms and charge separately. For a full aqueous equation, expand soluble strong electrolytes into their ions and cancel only identical species present on both sides. Retain solid precipitates, water and other unchanged molecular forms as appropriate. Use supplied ion charges; the equation is not a licence to invent a charge.

■ Checked worked case

Known: aluminium oxidation is $Al \rightarrow Al^{3+} + 3e^-$. Two aluminium atoms therefore release six electrons: $2Al \rightarrow 2Al^{3+} + 6e^-$. Three Cu^{2+} ions accept the same six electrons: $3Cu^{2+} + 6e^- \rightarrow 3Cu$. Combined: $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$. Both sides have total charge +6.

■ Common error

Electrons appear on the product side for oxidation and the reactant side for reduction. Spectator ions are chemically present even though omitted from the net equation. This Higher-only treatment does not make common-tier atom counting, isotope calculations or relative atomic mass Higher-only.

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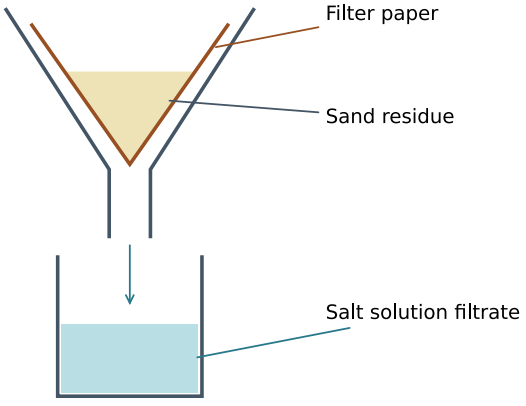
Separate insoluble solids and recover dissolved solutes

Official-unit focus: 4.1 Atomic structure and the periodic table

A filter can remove sand from salt water, but the salt solution passes through. Recovering both solids requires two different physical processes.

A mixture contains elements or compounds that are not chemically combined with one another; each retains its chemical properties. Filtration separates an insoluble solid from a liquid: sand stays as the residue and salt solution passes through as the filtrate. Dissolved ions pass through ordinary filter paper with the solvent. Crystallisation recovers a dissolved solid by concentrating its solution and allowing crystals to form as it cools.

Filtration: insoluble sand retained, dissolved salt passes through



Then concentrate filtrate gently and cool to crystallise the salt

For sand mixed with salt, add water and stir to dissolve the salt, filter the insoluble sand, then gently evaporate some water from the filtrate. Stop concentration before boiling dry and allow the solution to cool; filter and dry the crystals. Washing the sand removes adhering solution, and washing crystals with a little cold suitable solvent removes some surface impurity while limiting dissolution.

Use actual teacher-supervised school separation equipment, eye protection and a risk-assessed heat source. Fold and support the paper in a funnel, collect filtrate in a clean vessel and transfer solution to an evaporating basin. Use a safe end-point such as crystals forming in a cooled test drop. Hot glass can appear cold. Record mass only after the recovered material is dry.

■ Checked worked case

Known: a dry starting mixture contains 6.0 g sand and 10.0 g salt. A separation recovers 5.4 g dry sand and 8.0 g dry salt. Sand recovery = $5.4/6.0\times100 = 90\%$; salt recovery = $8.0/10.0\times100 = 80\%$. Lower recovery can reflect transfer losses or solute left in the remaining solution; it does not show that salt atoms vanished.

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■ Common error

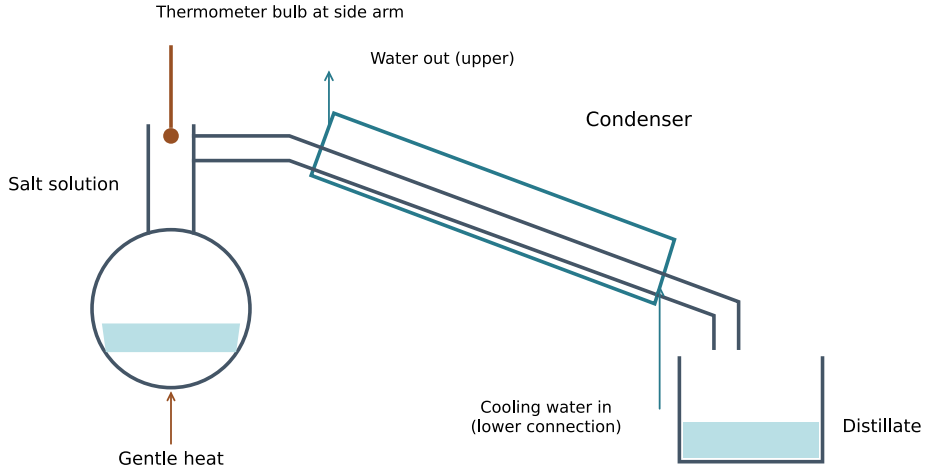
Filtering salt solution does not remove dissolved salt. Crystallisation changes physical arrangement, not the chemical identity of the solute. A wet recovered mass overestimates dry solid mass. Heating every solution to dryness is not a universal purification method because some substances decompose.

Choose distillation or chromatography from the mixture

Official-unit focus: 4.1 Atomic structure and the periodic table

Recovering pure water from salt solution needs the water collected, whereas crystallisation aims to recover salt. The product wanted determines the method.

Simple distillation heats a solution so the volatile solvent vaporises, then cools its vapour in a condenser to collect liquid distillate. Non-volatile salt stays in the flask. Fractional distillation separates miscible liquids with different boiling points using repeated vaporisation and condensation in a fractionating column. The lower-boiling component is enriched in vapour reaching the top; separation is not necessarily perfect.



Simple distillation: open receiving path; salt stays in the flask

Paper chromatography separates soluble components according to their different distributions between the moving solvent and the paper. Draw the baseline in pencil above the solvent level, place a small sample spot and let solvent rise. Components that spend more time in the mobile phase generally travel farther under these conditions. Compare spots with known standards using the same solvent and paper; a single spot alone does not prove purity under every possible method.

For school-supervised distillation, use a risk-assessed apparatus with an open receiving path, never a sealed heated system. Cooling water enters the lower condenser connection and leaves the upper one. Avoid flames with flammable solvents. In chromatography, keep the lid on the vessel as appropriate, mark the solvent front before it dries and measure all distances from the same baseline. Interpretation of Rf is

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developed in Chemical analysis, 4.8.1.3.

■ Checked worked case

Known: a model distillation starts with 60 cm^3 salt solution and collects 42 cm^3 water. Collected volume/starting solution volume = 42/60×100 = 70%. This is a volume-recovery comparison, not evidence of 70% purity or a mass balance. In a supplied chromatogram, a dye travelling 30 mm while the front travels 50 mm has Rf = 30/50 = 0.60.

■ Common error

A condenser liquefies vapour; it does not filter salt. Fractional distillation concerns boiling-point differences, not density ranking. A chromatogram baseline under solvent can let the sample dissolve directly into the reservoir. The method does not create new dye substances.

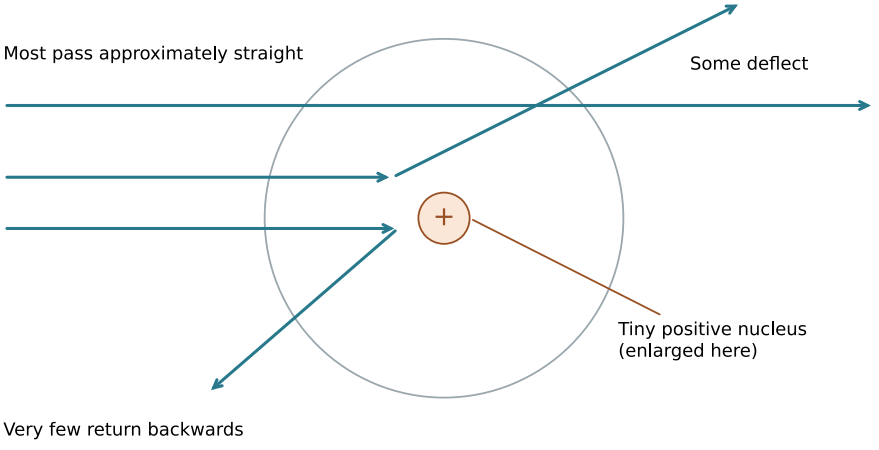
Atomic models change when evidence challenges predictions

Official-unit focus: 4.1 Atomic structure and the periodic table

Most alpha particles passed straight through thin foil, a few changed direction and very few returned backwards. A useful model must account for all three observations.

Before electrons were discovered, atoms were pictured as tiny indivisible spheres. Electron discovery led to the plum pudding model: negative electrons embedded in diffuse positive charge. Alpha-particle scattering challenged this distribution. Most particles passed through, consistent with an atom that is mostly empty space; rare large deflections showed mass and positive charge concentrated in a tiny central nucleus. Positive alpha particles are repelled by a positively charged nucleus.

Schematic paths around one nucleus; not a complete foil or scale model



The nuclear model replaced diffuse positive charge with a dense nucleus and electrons outside it. Bohr added electrons at specific distances or energy levels; his theoretical

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predictions agreed with observations. Later work identified positively charged protons. Chadwick provided evidence for uncharged neutrons in the nucleus about twenty years after the nuclear idea became accepted. A scientific model can develop further without every earlier observation becoming wrong.

Use supplied scattering observations or a computer-free diagram, not a classroom radioactive source. Match each observation to an inference and compare it with the diffuse-charge prediction. Distinguish observed trajectories from inferred structure. A marble analogue illustrates deflection but cannot establish electric charge or reproduce quantum behaviour. Detailed experimental work behind Bohr and Chadwick is outside this specification requirement.

■ Checked worked case

Known: in a fictional sample of 20,000 tracks, 19,800 pass approximately straight, 190 deflect and 10 return backwards. Straight-through percentage = 19,800/20,000×100 = 99%; backwards percentage = 10/20,000×100 = 0.05%. Rare events can be decisive evidence, but these counts are illustrative and are not historical measurements or a direct nucleus-size calculation.

■ Common error

Most particles passing through does not mean atoms have no mass. Backscattering is rare because the dense nucleus occupies little space, not because most nuclei are negatively charged. Shell diagrams are models and do not show electrons following observable miniature planetary tracks.

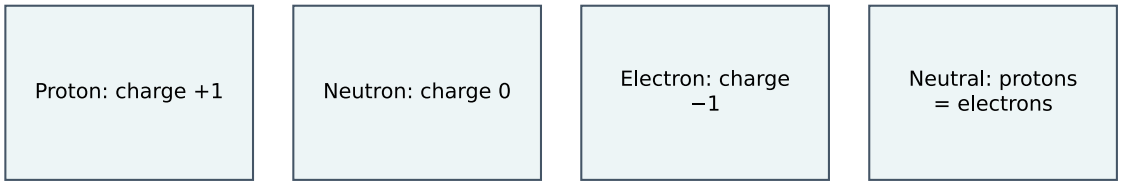
Proton number identifies the element; electrons determine net charge

Official-unit focus: 4.1 Atomic structure and the periodic table

A sodium atom and a sodium ion have different electron counts but still belong to the same element. Changing a neutron count also leaves the element identity unchanged.

A proton has relative electrical charge +1, a neutron 0 and an electron −1. Protons and neutrons are in the nucleus, with electrons outside it in energy levels. Atomic number, Z, is the number of protons. All atoms of a particular element have the same proton number; atoms of different elements have different proton numbers. In a neutral atom, electron number equals proton number, so their charges cancel.

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Net relative charge = proton number − electron number. Losing electrons produces a positive ion; gaining electrons produces a negative ion. Ordinary chemical ion formation changes electrons, not the nucleus. An atom with eleven protons is sodium even when it has ten electrons; its net charge is +1. A model with ten protons is neon, regardless of a similar electron arrangement.

Use counters labelled with charge rather than colours alone. Place protons and neutrons centrally and electrons outside, then add the signed charges. Identify the element from proton number using the periodic table. Explain neutrality as cancellation, rather than saying the atom contains no charged particles. Model distances are deliberately not to scale.

■ Checked worked case

Known: a particle contains 13 protons, 14 neutrons and 10 electrons. Z=13 identifies aluminium. Net relative charge = $13 - 10 = +3$, so the particle is Al^{3+}. The fourteen neutrons contribute no electric charge. If it instead had thirteen electrons, the same nucleus would form a neutral aluminium atom.

■ Common error

Atomic number is not the sum of every particle and cannot be read from electron count for an ion. Neutrality does not mean the proton and electron charges are individually zero. Neutrons affect mass and isotope identity without contributing to the net electric charge.

Isotopes, ions and the scale of the nucleus

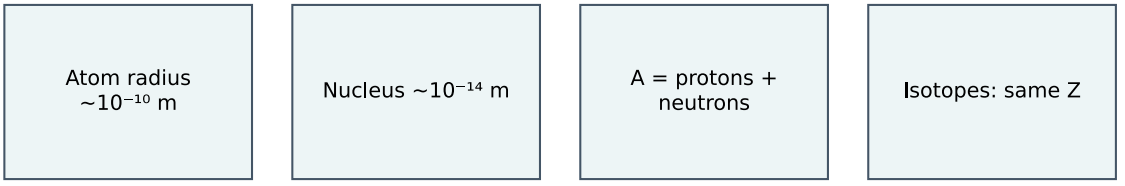
Official-unit focus: 4.1 Atomic structure and the periodic table

Two chlorine atoms can have the same proton number but different masses. Their nuclei contain different neutron numbers, while their element identity remains chlorine.

A typical atom has radius about 0.1 nm, equal to 1×10^{-10} m. Its nucleus has a radius less than one ten-thousandth of the atom's radius, of order 1×10^{-14} m. Almost all atomic mass is in the nucleus. Proton and neutron relative masses are each approxi-

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mately one; an electron's relative mass is very small. Mass number, A, counts protons plus neutrons. Isotopes are atoms of the same element with different neutron numbers.



In nuclide notation the upper-left number is A and the lower-left is Z. Protons=Z; neutrons=A−Z. A neutral atom has Z electrons; a positive ion has lost electrons and a negative ion has gained them. Chlorine-35 and chlorine-37 both have Z=17, with 18 and 20 neutrons respectively. Their mass numbers differ without a change in proton number.

Write a table with separate columns for A, Z, proton, neutron, electron and charge. Check A=protons+neutrons and charge=protons−electrons. For size comparisons convert radii into the same unit before dividing. A nucleus drawn clearly inside a circle is greatly enlarged relative to the atom; label the diagram not to scale and use numbers for a scale calculation.

■ Checked worked case

Known: chlorine-37 as Cl^- has 17 protons, 37−17=20 neutrons and 17+1=18 electrons. In a scale model where atom radius is 1 m and nuclear radius is 1/10,000 of it, the nuclear radius is 0.0001 m = 0.1 mm. This illustrates the order of size, not a precise measured radius for every isotope.

■ Common error

Mass number is an integer particle count, whereas relative atomic mass is an abundance-weighted mean. Ions differ in electrons, isotopes in neutrons. Dividing two radii gives a length ratio; the volume ratio is different. Do not place neutrons outside the nucleus in a shell diagram.

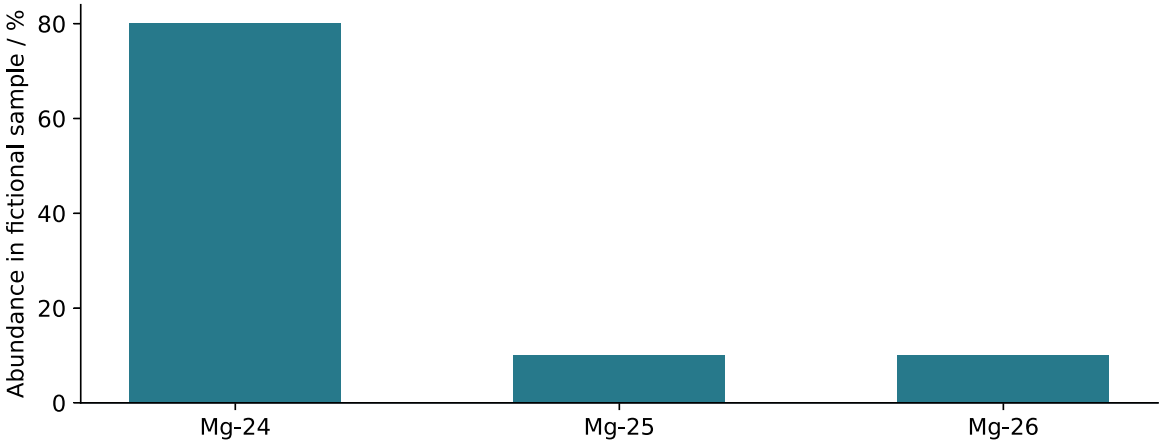
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Relative atomic mass: weight each isotope by its abundance

Official-unit focus: 4.1 Atomic structure and the periodic table

A periodic-table mass need not be a whole number. It describes an average over the natural mixture of isotopes rather than a nucleus containing a fraction of a neutron.

Relative atomic mass, Ar, is an average that accounts for isotope abundance. For GCSE calculations using supplied isotope mass numbers, multiply each mass by its percentage abundance, add the products and divide by 100. More abundant isotopes contribute more to the mean. If frequencies rather than percentages are supplied, divide the weighted sum by total frequency.



For a model chlorine sample containing 75% chlorine-35 and 25% chlorine-37, Ar=(35×75+37×25)/100=35.5. The mean is closer to 35 because that isotope is more abundant. An ordinary unweighted mean would give 36 and incorrectly treat both abundances as equal. Ar is a relative quantity without a gram unit; it is not the mass number of an individual atom.

Check that percentage abundances total 100 before using the percentage formula. A missing second abundance can be found as 100 minus the first only when the sample has exactly two stated isotopes. Estimate where the answer should lie between the isotope masses, then calculate without premature rounding. A bead model with labelled masses illustrates weighted counting without implying all natural samples have the chosen fictional proportions.

■ Checked worked case

Known: a fictional magnesium sample contains 80% mass-24, 10% mass-25 and 10% mass-26. Weighted sum=24×80+25×10+26×10=2,430. Divide by 100 to obtain Ar=24.3. It lies between 24 and 26 and close to the dominant mass-24 isotope. No atom in this model has mass number 24.3.

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■ Common error

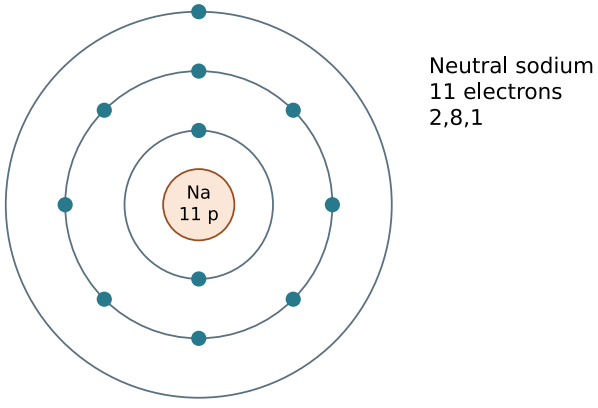
A percentage is divided by 100 once, not twice. Isotope abundance is not the proportion of neutrons inside a nucleus. A changing mean in different samples can reflect a different isotope mixture; it does not show that a single atom acquired a fractional neutron.

Electronic structure of the first twenty elements

Official-unit focus: 4.1 Atomic structure and the periodic table

Sodium has eleven electrons in its neutral atom. Its arrangement 2,8,1 is a distribution across shells, not the three digits of its atomic number.

Electrons occupy the lowest available energy levels, also called shells. For the first twenty elements the GCSE model places two in the first shell, then up to eight in the second, then eight in the third before the fourth begins. Sodium is 2,8,1 and calcium is 2,8,8,2. Draw a central nucleus and concentric shells with the correct electron counts. The drawing represents energy levels; shell radii and electron sizes are not to scale.



Energy-level model; distances and particle sizes not to scale

The first twenty arrangements are H 1; He 2; Li 2,1; Be 2,2; B 2,3; C 2,4; N 2,5; O 2,6; F 2,7; Ne 2,8; Na 2,8,1; Mg 2,8,2; Al 2,8,3; Si 2,8,4; P 2,8,5; S 2,8,6; Cl 2,8,7; Ar 2,8,8; K 2,8,8,1; Ca 2,8,8,2. Verify every sum against the neutral atom's proton number. This simplified sequence should not be extrapolated to every later element or used to claim the third shell can never contain more than eight.

Start with the atomic number from the supplied periodic table. For a neutral atom count that many electrons, fill the inner levels first and write a comma-separated arrangement. Convert it to a diagram by placing the exact number of dots on each ring and label the element. For an ion, adjust electron number before drawing and retain the unchanged nucleus; ion notation and brackets belong with the diagram.

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■ Checked worked case

Known: sulfur has Z=16, so a neutral atom has sixteen electrons. Put two in shell one, eight in shell two and the remaining six in shell three: 2,8,6. An S^{2-} ion has two additional electrons and arrangement 2,8,8. Both retain sixteen protons, while the ion has eighteen electrons and net charge −2.

■ Common error

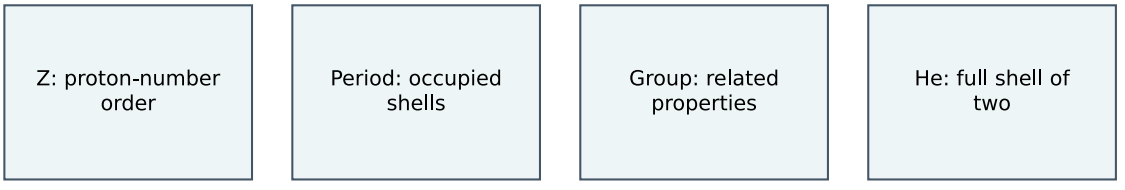
The second number in 2,8,6 is a count, not an electron charge or radius. Filling an outer shell while leaving an available inner level empty is incorrect for these ground-state atoms. Helium is stable with two electrons in its only shell; a full outer shell does not always mean eight.

Periodic positions follow proton number and electron structure

Official-unit focus: 4.1 Atomic structure and the periodic table

Lithium and sodium are separated by several elements in proton-number order but sit in the same column. Their shared outer-electron pattern explains why their reactions resemble one another.

The modern periodic table arranges elements by increasing atomic number. A group is a column containing elements with similar chemical properties; a period is a row. For the main-group first twenty elements, occupied shells give the period, while outer electrons explain group similarities. Magnesium, 2,8,2, is in period 3 and Group 2; chlorine, 2,8,7, is in period 3 and Group 7. AQA uses Group 0 for noble gases, whose outer shell is full.



Group 1 atoms tend to lose one outer electron and form +1 ions; Group 7 atoms can gain one to form −1 ions. Predict similar reaction types within a group, then use the group-specific trend for relative reactivity. Helium has two outer electrons in a full first shell and belongs with Group 0, not Group 2. Hydrogen is an unusual non-metal and should not be described as an alkali metal simply from its printed position.

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Use the supplied periodic table and numerical structures rather than memorizing an outline alone. Find Z, write the arrangement, count occupied shells, identify the outer count and compare the known group. State a prediction and the evidence for it. These main-group rules do not supply a complete electron-configuration treatment of transition metals.

■ Checked worked case

Known: an unknown neutral atom has structure 2,8,8,1. Its electron total is 19, so Z=19 identifies potassium. Four occupied shells give period 4, and one outer electron gives Group 1. It is predicted to form K^+ and to react more vigorously with water than sodium under comparable conditions, following the Group 1 trend.

■ Common error

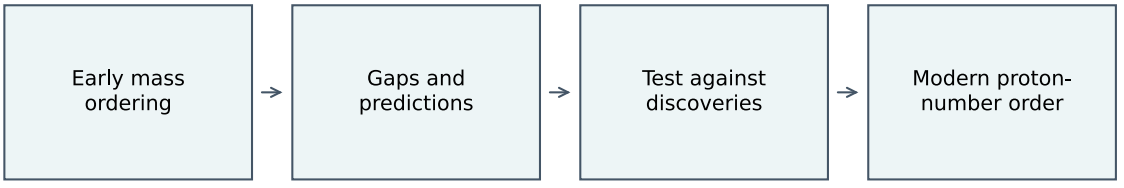
A group is vertical and a period horizontal. Similar properties do not mean identical reaction rates. Outer-shell count is not always the printed modern 1–18 group number, and the simplified first-twenty rule must not be applied indiscriminately to every column.

Mendeleev: gaps and testable predictions

Official-unit focus: 4.1 Atomic structure and the periodic table

An empty position can make a scientific classification stronger when it predicts what an undiscovered element should be like. Mendeleev treated recurring properties as evidence rather than filling every gap.

Early classifications ordered elements by atomic weight before protons, neutrons and electrons were known. Tables were incomplete, and strict mass ordering sometimes put elements into groups with unlike properties. Mendeleev left gaps for undiscovered elements and reversed mass order in some places to retain chemically sensible groups. He used neighbours to predict properties of elements expected in those gaps.



Later discoveries filled gaps with elements whose properties agreed with the predic-

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tions, supporting the classification. Modern order follows atomic number rather than relative atomic mass. Isotopes explain why average mass need not rise strictly with proton number: different isotope masses and abundances influence the mean. The historical improvement was evidence-based prediction, not knowledge of electron shells that had yet to be discovered.

Compare a supplied early table with a modern periodic table. Identify a gap or reversed pair, write the property pattern and decide what discovery would support or challenge the prediction. Use fictional property values for numerical interpolation and label them as such. Agreement with one predicted property is useful evidence but not proof that every aspect of a model is correct.

Checked worked case

Known: in a fictional series a gap lies between densities 4.8 and 6.0 g/cm^3. A simple midpoint estimate is (4.8+6.0)/2 = 5.4 g/cm^3. A discovered value of 5.5 differs by 0.1 g/cm^3. This demonstrates a testable interpolation, not Mendeleev's actual historical data or a universal linear law.

Common error

Mendeleev did not arrange by known proton number. Leaving gaps was a reasoned prediction rather than accidental omission. Isotopes help explain mass-order anomalies; they do not imply the same element occupies several different proton-number positions.

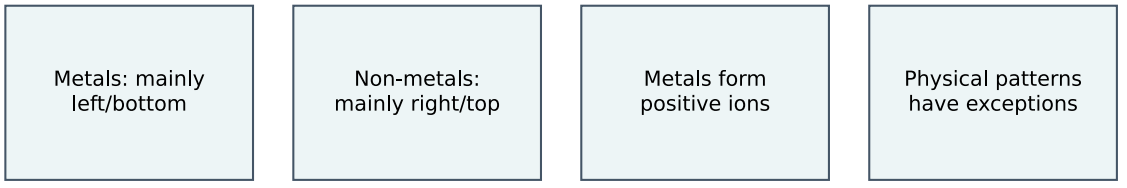
Metals and non-metals: properties, position and ions

Official-unit focus: 4.1 Atomic structure and the periodic table

Sodium and sulfur are both solids in the same period, but their electron arrangements and reactions differ. A single property such as being solid cannot identify a metal.

Most elements are metals, located toward the left and bottom of the periodic table. Non-metals are mainly toward the right and top. Metals characteristically conduct electricity and heat, are often strong and malleable, and react to form positive ions by losing electrons. Non-metals generally conduct poorly and, when solid, are often brittle; many occur as gases. In the specified ionic reactions, non-metals may gain electrons to form negative ions.

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Sodium, 2,8,1, loses its outer electron to form Na^+. Chlorine, 2,8,7, gains one to form Cl^-. Oppositely charged ions can form a compound. Non-metal atoms also share electrons in covalent molecules, rather than always becoming isolated negative ions. Characteristic property patterns have exceptions: graphite conducts electricity and mercury is a liquid metal. Use more than one line of evidence.

Compare teacher-provided property cards, conductivity observations and electronic structures. Choose a classification, explain the electron change in a specified reaction and name a limitation of a physical-property inference. Use pre-approved specimens and supervised low-voltage conductivity apparatus; do not attempt reactions with unknown materials to discover their identity.

■ Checked worked case

Known: four neutral sodium atoms each lose one electron when forming four Na^+ ions. Four electrons are transferred in total; the four nuclei each retain eleven protons. Four chlorine atoms gaining those electrons form four Cl^- ions. The charges balance, but the electron transfer does not turn sodium nuclei into chlorine nuclei.

■ Common error

Losing negative electrons makes an ion positive. Metals do not lose protons during normal chemical reactions. Non-metal does not mean gas, and shiny appearance alone is insufficient evidence. The simplified GCSE ion classification describes the reactions in scope rather than every possible ion in advanced chemistry.

Group 0: stable shells and boiling-point predictions

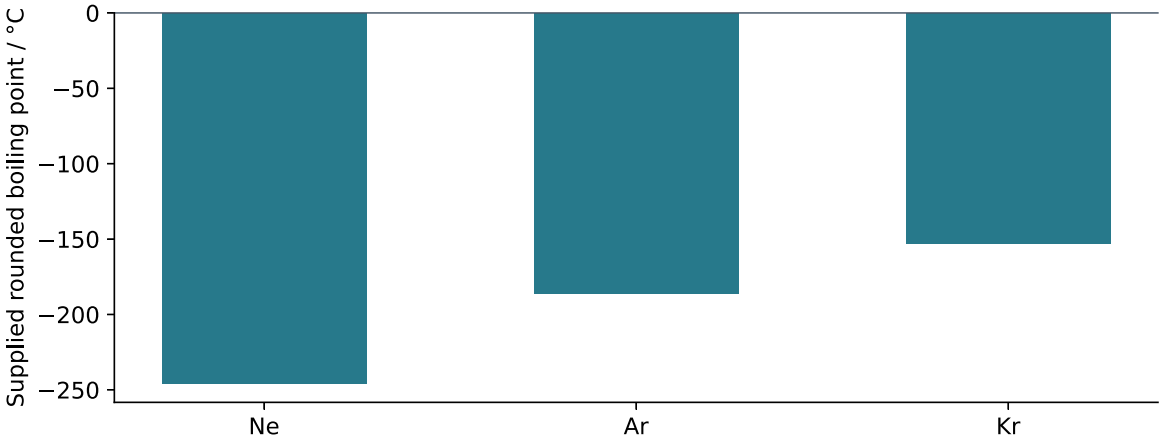
Official-unit focus: 4.1 Atomic structure and the periodic table

Helium has only two electrons but is very unreactive. Stability is about a complete outer shell, not an unconditional target of eight electrons.

Group 0 elements are noble gases. Their atoms have stable outer-electron arrange-

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ments and do not easily form molecules, so they normally occur as separate atoms. Helium has a full first shell of two electrons; the other noble gases in this GCSE treatment have eight in the outer shell. They do not readily gain, lose or share electrons in ordinary reactions. Their boiling points increase with increasing relative atomic mass down the group.



Neon is 2,8 and argon 2,8,8. Compare these stable structures with an alkali metal's one outer electron or a halogen's seven. Unreactivity and boiling point are different properties: a rising boiling point does not imply rising chemical reactivity. Given trend data, predict an ordering; a trend alone rarely determines an exact numerical value for an unmeasured element.

Use supplied structures and a labelled boiling-point table. Keep negative temperature ordering explicit: −186 °C is higher than −246 °C. Plot values when given, then distinguish interpolation from extrapolation. Atomic drawings are not evidence that helium needs another six electrons. More detailed intermolecular-force explanation is developed in Bonding, 4.2.

■ Checked worked case

Known: supplied rounded values are neon −246 °C and argon −186 °C. The increase is $-186 - (-246) = 60$ °C. Argon therefore has the higher boiling point even though both values are below zero. The values illustrate reading the trend and do not imply a 60 °C increase between every consecutive pair.

■ Common error

Noble gases are generally unreactive, not incapable of any reaction under any conditions. Helium belongs in Group 0 with two outer electrons. Boiling separates physical particles and does not remove electrons from atomic shells.

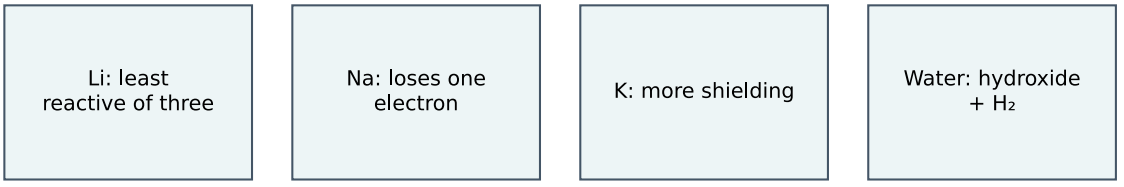
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Group 1: reactions of lithium, sodium and potassium

Official-unit focus: 4.1 Atomic structure and the periodic table

Lithium, sodium and potassium all have one outer electron. Comparable reactions become more vigorous down the group even though proton number also increases.

Group 1 alkali metals have one outer electron and form +1 ions by losing it. With water they produce a metal hydroxide and hydrogen: $2Na + 2H_2O \rightarrow 2NaOH + H_2$. Lithium fizzes and moves on water; sodium also melts into a moving ball; potassium can ignite with a lilac flame in a controlled demonstration. Solutions become alkaline. With chlorine they form white metal chlorides, for example $2Na + Cl_2 \rightarrow 2NaCl$.



With oxygen, these metals tarnish and burn to form oxygen-containing solids; lithium burns crimson, sodium yellow and potassium lilac. The simple oxide model uses $4Li + O_2 \rightarrow 2Li_2O$; sodium and potassium burning products can include peroxide or superoxide, so do not invent a single oxide formula for every demonstration. Reactivity increases Li→Na→K because the outer electron is farther from the nucleus and more shielded by inner shells, so it is lost more easily despite increased nuclear charge.

Interpret teacher demonstrations or supplied observations under school risk assessment. Alkali-metal water reactions need approved very small quantities, protective screens and trained handling; students should not independently scale them up. Compare the same reaction conditions and distinguish qualitative vigour from a measured rate. Oxygen and chlorine demonstrations require their own approved controls.

■ Checked worked case

Known: $2K + 2H_2O \rightarrow 2KOH + H_2$ shows two potassium atoms reacting per hydrogen molecule in the particle model. Six potassium atoms correspond to three H_2 molecules and six KOH formula units. This is a stoichiometric model; it is not instructions for preparing hydrogen or a quantitative mole calculation.

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■ Common error

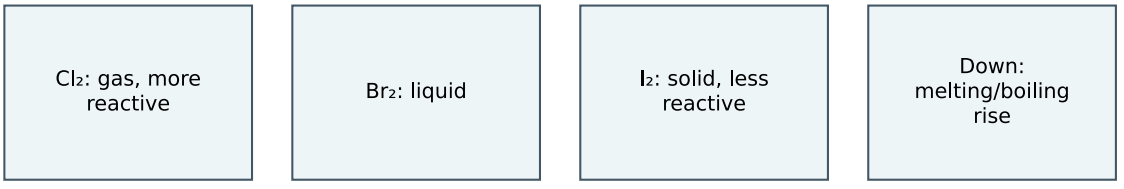
The larger nuclear charge alone does not explain the down-group trend: distance and shielding matter. Group 1 reactions are not all identical in appearance. A purple flame is not evidence that potassium chloride itself is a purple solid, and a water-reaction equation must include hydrogen.

Group 7: molecules, compounds and opposite trends

Official-unit focus: 4.1 Atomic structure and the periodic table

Chlorine is more reactive than iodine, but iodine has higher melting and boiling points. Chemical reactivity and physical state follow different explanations.

Group 7 halogens are non-metals with seven outer electrons. Their elements consist of diatomic molecules such as Cl_2, Br_2 and I_2. Near ordinary room conditions chlorine is a pale green gas, bromine a red-brown liquid and iodine a grey solid forming purple vapour when heated. Down the group, relative molecular mass, melting point and boiling point increase. Reactivity decreases chlorine→bromine→iodine.



A halogen gains an electron to form a -1 halide ion when reacting with a metal, giving an ionic salt such as NaCl. With non-metals it generally forms covalent compounds through shared electron pairs, such as HCl. Down the group the outer shell is farther from the nucleus and more shielded, reducing attraction for an incoming electron despite greater nuclear charge. This is the opposite reactivity trend to Group 1, which loses an electron.

Compare supplied molecular drawings, states and reaction data without handling halogens independently. Chlorine and bromine require teacher-approved containment and ventilation; appearance cards can provide the required evidence safely. Use the same conditions when comparing reactivity. From a trend predict whether an unfamiliar halogen is more or less reactive, and avoid fabricating an exact boiling point.

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■ Checked worked case

Known: using supplied relative atomic masses Cl=35.5 and Br=80, relative molecular masses are Cl_2=2×35.5=71 and Br_2=2×80=160. Bromine has the larger molecular mass, while chlorine is the more reactive element. Neither value has units of grams, and each refers to the diatomic molecule.

■ Common error

A halogen element and its halide ion have different electron counts and properties. Chlorine, Cl_2, is not written as Cl^-. A rising boiling point does not establish a rising rate of electron gain. Solid sodium chloride is white even though chlorine gas is green.

Halogen displacement: predict before observing

Official-unit focus: 4.1 Atomic structure and the periodic table

Chlorine water added to potassium bromide can produce bromine. Reversing the materials does not reverse the reaction: bromine cannot displace the more reactive chlorine.

A more reactive halogen displaces a less reactive halogen from an aqueous solution of its salt. Chlorine displaces bromine and iodine; bromine displaces iodine but not chlorine; iodine displaces neither chlorine nor bromine. For chlorine and potassium bromide, $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$. Potassium remains in the solution while the identity of the free halogen changes.

Added halogen → starting aqueous halide

	Cl^-	Br^-	I^-
Cl_2	No displacement	Br_2 formed	I_2 formed
Br_2	No displacement	No displacement	I_2 formed
I_2	No displacement	No displacement	No displacement

Reactivity: chlorine > bromine > iodine; same-halogen pairs have no net reaction

A displacement table should identify the added halogen, starting halide and expected product before recording colour. Bromine in water is orange/brown and iodine solutions can appear brown; observation must be compared with suitable blanks, because the added halogen also has colour. Same-halogen/halide combinations have no net

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displacement. A missing colour change alone may also reflect low concentration or observation limits.

For AT6 school-supervised work, use approved dilute halogen solutions in small-scale containers under the teacher's risk assessment. Keep solution volumes and concentrations comparable, label each well and include reference colours. Do not generate chlorine or use an unapproved extraction solvent. Record actual observations, then compare with the predicted ranking; this activity is not one of the separately numbered required practicals.

■ Checked worked case

Known: chlorine added to chloride, bromide and iodide solutions gives two predicted displacement reactions out of three tests. Bromine gives one, and iodine zero, under the stated matrix. For $Cl_2 + 2KI \rightarrow 2KCl + I_2$, four KI formula units correspond to two I_2 molecules in the balanced particle model.

■ Common error

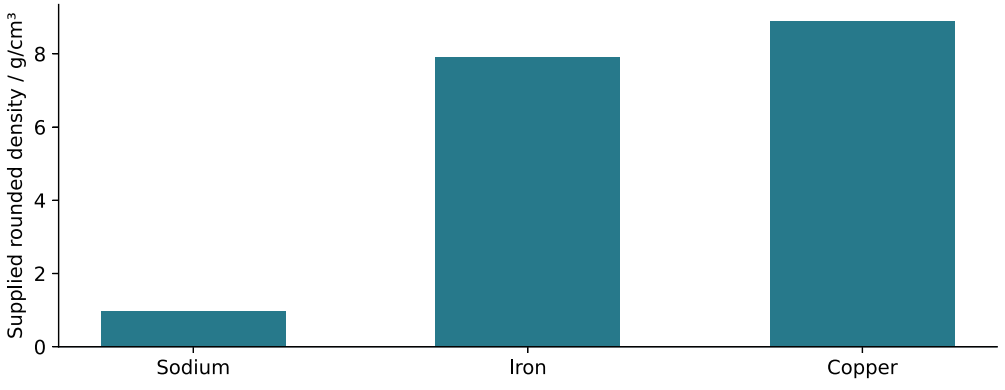
Displacement releases the less reactive halogen, not the more reactive one. A metal-ion spectator does not establish the reactivity order. Do not claim all orange solutions contain bromine without considering the starting reagents and reference observations. The separate HT ionic-equation lesson extends the both-tier prediction.

Chemistry-only: transition metals compared with Group 1

Official-unit focus: 4.1 Atomic structure and the periodic table

A copper wire and a piece of sodium are both metals, yet their behaviour in water and their usefulness in construction differ substantially. Being a metal does not imply Group 1 reactivity.

The specified transition examples are Cr chromium, Mn manganese, Fe iron, Co cobalt, Ni nickel and Cu copper. Compared with Group 1 metals, these typically have higher melting points and densities and are stronger and harder. They are generally less reactive with oxygen, water and halogens. Sodium is soft enough to cut and reacts vigorously with cold water; an iron nail does not show that Group 1 reaction.



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Iron can react with steam under suitable conditions and can rust slowly in oxygen and water; copper does not react with cold water in the Group 1 manner. Many transition metals react with oxygen or halogens when heated, so less reactive does not mean incapable of reaction. Compare the same conditions and specify whether an observation is rapid burning, slow corrosion or no visible change.

Use supplied property data and teacher-approved solid samples to compare density and qualitative hardness. Avoid student handling of alkali metals. Rank measured values rather than extrapolating an exact value for every transition element. For reaction comparisons, teacher demonstrations or source observations must have stated conditions and comparable sample sizes.

■ Checked worked case

Known: supplied rounded densities are sodium 0.97 and iron 7.9 g/cm^3. Equal volumes have mass ratio iron/sodium = 7.9/0.97 ≈ 8.14. A 2.0 cm^3 iron model sample has mass 7.9×2.0=15.8 g. Density does not itself measure reaction rate, hardness or strength.

■ Common error

The listed differences are typical comparisons, not universal laws for every metal. Density and hardness name different properties. Transition metals can react with halogens, and lack of an immediate cold-water reaction does not prove no reaction under different conditions.

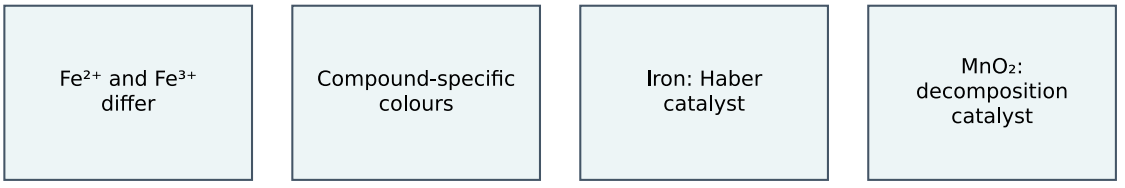
Chemistry-only: variable ions, colours and catalysts

Official-unit focus: 4.1 Atomic structure and the periodic table

Iron forms Fe^{2+} and Fe^{3+} ions in different compounds. A charge must therefore be specified; knowing only that a substance contains iron cannot determine its complete formula.

Many transition metals form ions with different charges, coloured compounds and useful catalysts. Iron(II) and iron(III) show variable charges. Copper(II) aqueous compounds are commonly blue; iron(II) solutions are often pale green and iron(III) solutions yellow/brown. Chromium(III) compounds can be green, manganate(VII) is purple, hydrated cobalt(II) compounds can be pink and nickel(II) solutions commonly green. Colour depends on the particular compound, hydration and conditions.

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Catalysts increase reaction rate without being used up overall. Iron catalyses the Haber process; nickel is used in hydrogenation; manganese dioxide catalyses hydrogen peroxide decomposition. A catalyst changes the route and rate, not the balanced reaction's overall atom count. Variable charges require matching negative-ion charge: $FeCl_2$ contains Fe^{2+} and $FeCl_3$ contains Fe^{3+} with Cl^- in each case.

Use labelled compound photographs, supplied solution observations and given ion charges. Compare an actual teacher-supervised catalysed reaction with an uncatalysed control, maintaining temperature, concentration and volume. Cobalt/nickel compounds and strong oxidizers require approved handling and disposal; reference evidence is appropriate when they are not used in a school experiment. Do not identify an unknown compound solely from colour.

■ Checked worked case

Known: a compound contains Fe^{3+} and O^{2-}. Two Fe^{3+} supply total +6 and three O^{2-} supply −6, giving neutral Fe_2O_3. With Fe^{2+}, one iron ion and one oxide ion balance as FeO. The different iron charges produce different ratios; neither formula is inferred merely from a brown or black appearance.

■ Common error

Many does not mean every transition compound is coloured or every metal has every charge. A catalyst may take part in intermediate steps while being regenerated overall. Its presence does not increase the theoretical amount of product from a fixed limiting reactant or prove an equilibrium composition has changed.

Particles, bonding and bulk properties

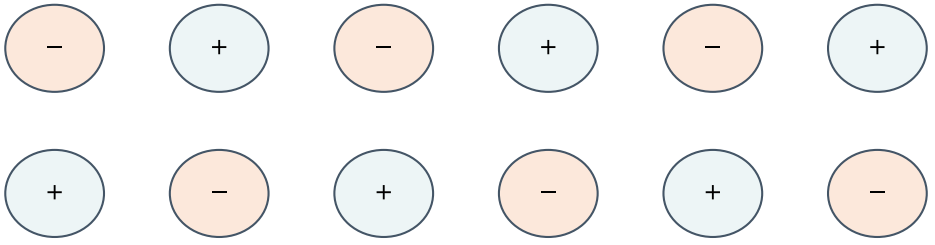
Official-unit focus: 4.2 Bonding, structure and properties

A salt crystal conducts when dissolved but not when solid. The ions exist in both states; their ability to move changes.

Ionic bonding is electrostatic attraction between oppositely charged ions. A covalent

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bond involves shared electrons. Metallic bonding involves attraction between positive metal ions and delocalized electrons.



Alternating charges in an ionic lattice

To explain a bulk property, name the structure, particles, forces and mobile charge carriers. Simple molecular substances can have strong covalent bonds inside molecules but weak attractions between molecules.

Compare substances using evidence such as melting point, conductivity when solid and molten, and solubility. One property rarely proves a structure; use a pattern of evidence.

■ Checked worked case

Known: an element has atomic number 12 and mass number 24. Protons = 12; neutrons = mass number - atomic number = 24 - 12 = 12. A 2+ ion has electrons = 12 - 2 = 10. Charge changes electron count, not the nucleus.

■ Common error

Melting a simple molecular substance usually overcomes intermolecular attractions; it does not require breaking all covalent bonds within each molecule.

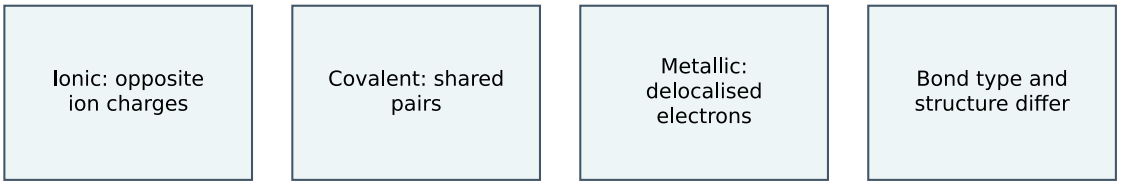
Three strong bonds: identify the attracted particles

Official-unit focus: 4.2 Bonding, structure and properties

Sodium chloride, hydrogen and copper all have strong bonding, but the particles and electron arrangements responsible are different. Name what attracts what before predicting a property.

Ionic bonding is strong electrostatic attraction between oppositely charged ions. It occurs in compounds formed from metals and non-metals after electrons are transferred. Covalent bonding joins atoms sharing pairs of electrons: attraction between the shared electrons and both nuclei holds atoms together. Metallic bonding is attraction between positive metal ions and delocalised electrons throughout a giant metal structure. Metallic bonding occurs in metals and alloys.

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Most non-metal elements and compounds of non-metals have covalent bonding, but their structures can be small molecules, very large molecules or giant networks. An ionic bond is the attraction after ion formation, not the electron transfer itself. A covalent pair is shared, not transferred completely to one nucleus. Metal electrons are free to move through the structure rather than assigned to one pair of atoms.

Sort teacher-provided particle diagrams by charges, shared pairs and delocalised electrons. For each classification state the particles, electrostatic attraction and whether the drawing represents a small molecule or part of a giant structure. Do not classify only from melting point: different structures can produce overlapping property data.

■ Checked worked case

Known: Mg forms Mg^{2+} and oxygen forms O^{2-}. One ion of each gives total charge $+2-2=0$, so MgO is the formula. Its ionic bond is attraction between these charges in a lattice, not a separate covalent pair. By contrast an H_2 molecule has one shared pair and no net charge on the complete molecule.

■ Common error

Strong covalent bonds do not guarantee a high boiling point for a small-molecule substance, because boiling usually separates molecules rather than atoms. An ionic compound is not a collection of independent MgO molecules. Electrons are negative in all three bonding accounts.

Ionic dot-and-cross diagrams: transferred electrons and charges

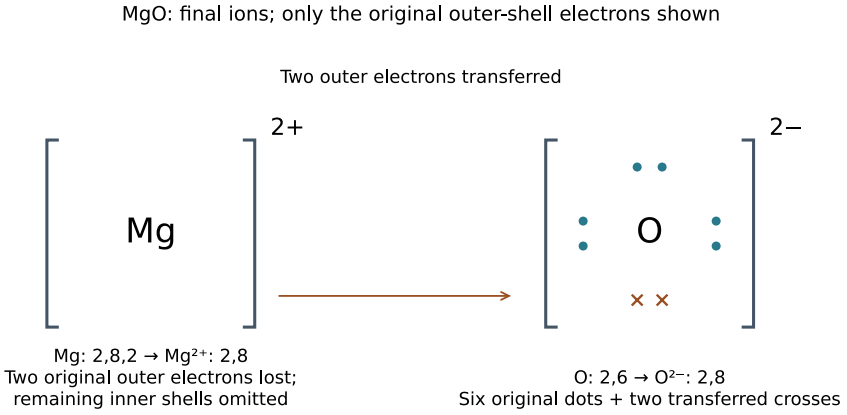
Official-unit focus: 4.2 Bonding, structure and properties

Magnesium loses two electrons while oxygen gains two. A correct diagram shows both the transferred electrons and the final ion charges; a picture of neutral atoms is only the starting point.

Group 1 metals form +1 ions and Group 2 metals +2 ions by losing their outer

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electrons. Group 7 non-metals gain one electron to form −1 ions, and Group 6 non-metals gain two to form −2 ions. In the specified examples these ions have a noble-gas electronic structure. Magnesium changes from 2,8,2 to Mg^{2+} with 2,8; oxygen changes from 2,6 to O^{2-} with 2,8.



Draw final ions in separate square brackets, with charges outside. A dot-and-cross diagram uses different marks for electrons originating from different atoms; transferred electrons keep their original marks. Mg^{2+} has no electrons left in its original third shell, while O^{2-} has six original outer electrons plus two transferred ones. If only outer electrons are shown, state whether the remaining filled shell of the cation is omitted by that convention.

Start from neutral structures and account for every transferred electron. For NaCl transfer one; for MgO transfer two; for $MgCl_2$ transfer one to each of two chlorine atoms; for Na_2O transfer one from each of two sodium atoms to oxygen. Confirm total charge zero and outer-shell completion. Dots and crosses indicate origin only: all electrons have the same physical charge and properties.

■ Checked worked case

Known: $MgCl_2$ contains one Mg^{2+} and two Cl^- ions. Magnesium loses two electrons, one accepted by each chlorine. Charges total $+2+2(-1)=0$. Each chloride has seven original outer electrons and one transferred electron. The simplest ion ratio is 1:2; the diagram must contain both chloride ions.

■ Common error

An ion has the same proton number as its original atom. Do not write Na^{2+} just because two sodium ions appear, or put the transferred electron halfway between final ions as though it were covalently shared. A full shell in these ion examples does not change the element into the noble gas.

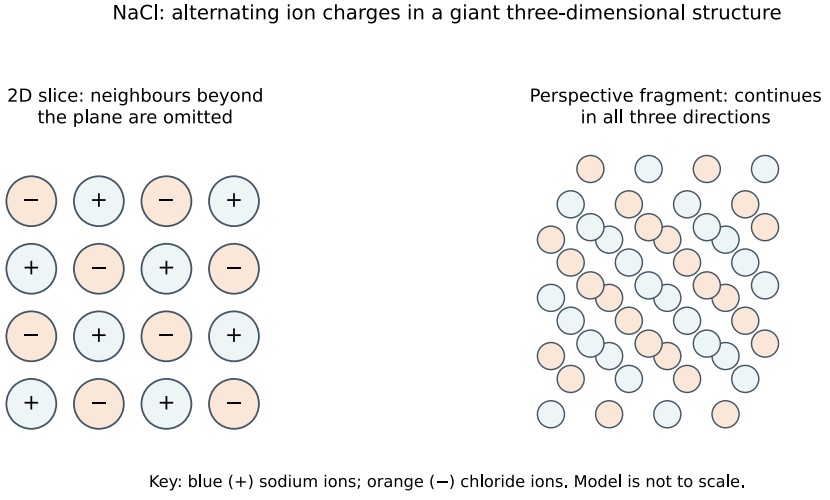
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Giant ionic lattices: formulae and model limits

Official-unit focus: 4.2 Bonding, structure and properties

A sodium chloride crystal contains a repeating arrangement of ions. The formula NaCl records the simplest ratio, not a single isolated pair joined to form a molecule.

An ionic compound is a giant lattice of oppositely charged ions held by strong electrostatic forces acting in all directions. In sodium chloride the Na^+:Cl^- ratio is 1:1. A two-dimensional slice can show alternating charges, while a three-dimensional model shows the repeating arrangement beyond one plane. An interior ion in the sodium chloride structure has six nearest neighbours of opposite charge; an edge drawn on a finite model is a cut through the continuing crystal.



To infer an empirical formula, count ions in a stated representative region and reduce the ratio to the smallest whole numbers. A model with six Mg^{2+} and twelve Cl^- gives $MgCl_2$ after reducing 6:12 to 1:2. Counts in an arbitrary boundary fragment need not themselves give the bulk ratio, so use the representative information or account for shared positions as specified. Knowledge of other named ionic crystal structures is not required.

Compare dot-and-cross, ball-and-stick and space-filling representations. Dot/cross shows electron origin and charge but not the full lattice geometry. Ball/stick makes an arrangement clear but rods are not physical ionic bonds and spaces/radii may be misleading. A 2D slice omits neighbours above and below; a 3D model can hide ions behind others. State the particular useful feature and limitation of each model.

■ Checked worked case

Known: a supplied representative diagram contains four Al^{3+} and six O^{2-} ions. The ratio 4:6 simplifies to 2:3, so its empirical formula is Al_2O_3. Charge check: $2(+3)+3(-2)=0$. This ratio deduction does not require knowing the actual geometry of the aluminium oxide crystal.

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■ Common error

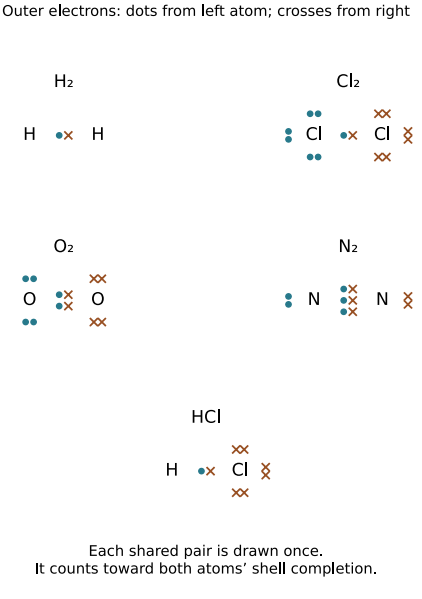
Do not count sticks as electron pairs in an ionic ball-and-stick model. Ions attract several neighbours rather than only the one from which an electron came. A simple lattice slice is a representation of a three-dimensional giant structure, not evidence that the real crystal is a flat sheet.

Shared pairs: hydrogen, chlorine, oxygen and nitrogen

Official-unit focus: 4.2 Bonding, structure and properties

An oxygen molecule contains two shared electron pairs, while nitrogen contains three. Counting shared pairs distinguishes a double from a triple covalent bond.

A covalent bond forms when atoms share a pair of electrons, attracted to both positive nuclei. Hydrogen atoms each contribute one electron to H_2, giving one shared pair and a full first shell of two. Chlorine atoms each have seven outer electrons: in Cl_2 they share one pair and each retains three lone pairs. In HCl, hydrogen and chlorine share one pair; hydrogen has two electrons in its shell and chlorine an octet.



Oxygen has six outer electrons, so O_2 has two shared pairs, a double bond, and two lone pairs on each atom. Nitrogen has five outer electrons, so N_2 has three shared pairs, a triple bond, and one lone pair on each atom. A line diagram writes H–H, Cl–Cl, H–Cl, O=O or N≡N. One line represents one shared pair, not one electron.

Count the available outer electrons first. Place shared pairs between the correct atoms using dots from one and crosses from the other, then place remaining electrons as lone pairs. Check each atom's full outer shell and the total electrons drawn. Compare with

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a line or ball-and-stick model: these make connectivity clear but normally omit lone pairs and electron origin.

■ Checked worked case

Known: O_2 has $2\times6=12$ outer electrons altogether. Four are in its two shared pairs; the remaining eight form four lone pairs, two on each oxygen. Each oxygen counts the four shared electrons plus its own four unshared electrons as an octet. Shared electrons are counted by both atoms for shell completion but drawn only once in the molecule.

■ Common error

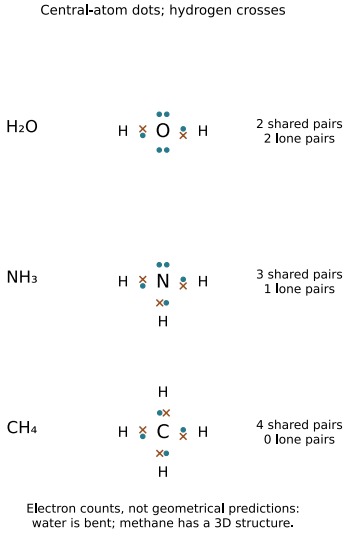
A double bond contains four shared electrons, not two. A chlorine atom's six unshared electrons must not disappear just because a line diagram omits them. These simple electron pictures are models, not a claim that electron positions are fixed dots or that a bond is a literal rod.

Water, ammonia, methane and larger covalent structures

Official-unit focus: 4.2 Bonding, structure and properties

A water molecule has two bonds but also two lone pairs on oxygen. A displayed line formula makes connections easy to see while leaving some electron information out.

In H_2O, oxygen shares one pair with each of two hydrogen atoms and retains two lone pairs. In NH_3, nitrogen shares one pair with each of three hydrogens and retains one lone pair. In CH_4, carbon shares four pairs, one with each hydrogen, and has no lone pairs in the outer-electron diagram. Each hydrogen attains two shell electrons; carbon, nitrogen and oxygen count eight around themselves.



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Covalent substances can consist of small molecules, very large molecules such as polymers, or giant networks such as diamond and silicon dioxide. A polymer has many repeated units linked along each molecule: $[–CH_2–CH_2–]_n$ represents poly(ethene), with bonds passing through brackets and n large. A giant network diagram shows only a small part of continuing covalent connections, not a separate molecule with that exact pictured atom count.

For electron diagrams, use one mark type for the central atom's electrons and another for hydrogen's. Count bonds, lone pairs and all outer electrons. For a molecular formula, count each element in the complete stated molecule. For a polymer, identify the bracketed repeat rather than calling the drawing a small molecule. Ball-and-stick drawings enlarge gaps and do not fix scale; a 2D line formula omits the real 3D arrangement.

■ Checked worked case

Known: methane has $4+4\times1=8$ outer electrons in total, all in four shared pairs. Ammonia also has $5+3\times1=8$: six shared electrons and a two-electron lone pair. A poly(ethene) model containing fifty C_2H_4 repeat units contains 100 carbon and 200 hydrogen atoms in those repeat units; chain-end details are excluded from that simplified count.

■ Common error

A lone pair is not a spare hydrogen or an additional bond. Do not apply molecular formula counting to a clipped giant network without understanding its continuation. The number n is a count of repeat units, not a charge or an electron shell number. Polymer formation is developed further in Organic chemistry.

Metallic bonding: a giant structure with mobile electrons

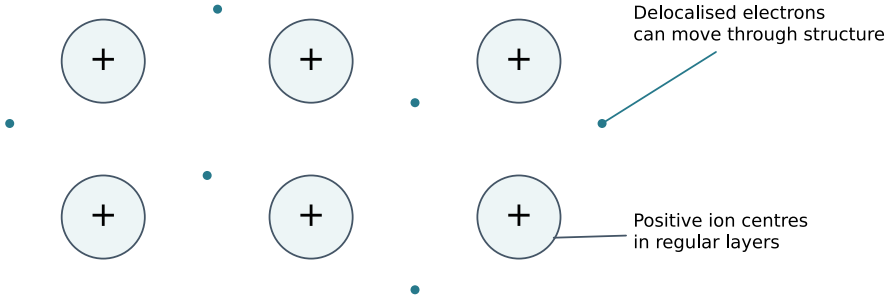
Official-unit focus: 4.2 Bonding, structure and properties

A copper wire conducts as a solid. Unlike a solid ionic crystal, it already contains charge carriers that can move through its giant structure.

A metal has a giant arrangement represented by positive metal ions in a sea of delocalised outer electrons. These electrons are not attached to one atom or one neighbouring pair and can move through the structure. Strong electrostatic attraction between the positive ions and delocalised electrons is metallic bonding. The complete metal is electrically neutral: electron charge balances the positive charge represented by the ions.

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One-electron metal model: positive centres and delocalised electrons



Total positive charge is balanced by electron charge; complete metal is neutral.

The shared electrons extend throughout the structure, so shifting layers can preserve attraction and allow bending. When a potential difference is applied, mobile electrons carry charge; the positive ions remain in their positions in a solid rather than drifting through the wire. The drawing's ions and electron symbols are a model of the bulk bonding, not isolated metal ions mixed with a separate substance.

Compare a metallic diagram with an ionic lattice and a small covalent molecule. Identify whether electrons are local shared pairs or delocalised and whether both positive and negative ions are represented. Use school-approved low-voltage conductivity demonstrations, with no mains circuit and no heating of unknown samples. A schematic can show carriers but not their detailed motions or actual relative sizes.

■ Checked worked case

Known: in a simplified sodium metal model, eight atoms each contribute one outer electron to the delocalised system. The model therefore has eight delocalised electrons balancing eight singly positive ion centres. A twelve-atom magnesium model contributes two each, giving twenty-four electrons. These are electron counts in an explicitly stated model, not a formula for every metal.

■ Common error

The metal does not have a net positive charge simply because the drawing labels positive ion centres. Electrons are delocalised, not lost from the entire piece. Electrical conductivity in a solid metal is different from ion mobility in molten salt. Do not use the sodium model to assign one free electron to every metal atom universally.

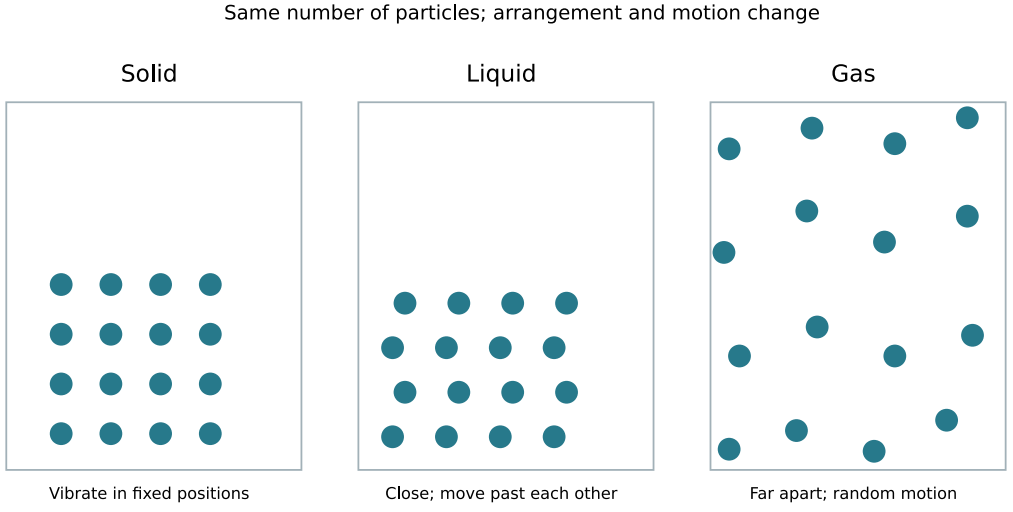
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States and changes: energy overcomes particle attractions

Official-unit focus: 4.2 Bonding, structure and properties

Ice melts while its water molecules remain water molecules. A state change rearranges particles and their motion without necessarily breaking the covalent bonds inside them.

A solid has particles close together vibrating about fixed positions. A liquid has particles still close but able to move past one another. Gas particles are far apart and move freely in random directions. Melting/freezing occur at the melting point and boiling/condensation at the boiling point for the stated pressure. Heating transfers energy: at a state change energy overcomes relevant attractions so arrangement changes; cooling transfers energy away as attractions bring particles closer.



Stronger relevant forces require more energy to overcome and commonly give higher melting and boiling points. Identify the particles and forces before explaining: boiling a molecular liquid overcomes intermolecular attractions, while melting an ionic lattice overcomes strong ion attractions. Bulk properties belong to large collections, not an individual atom that is itself a tiny piece of solid or liquid. At a transition temperature more than one state can coexist.

Given melting and boiling points at one pressure, compare the temperature with both boundaries. Below melting: solid; between: liquid; above boiling: gas. Use teacher-provided observations or an approved water-heating/cooling experiment. Record temperature and state, and distinguish a measured plateau from a schematic curve. Changing pressure or mixtures can alter transition temperatures.

■ Checked worked case

Known: a pure model substance has melting point -12 °C and boiling point 68 °C at the stated pressure. At -20 °C it is solid, at 25 °C liquid and at 80 °C gas. The liquid interval spans $68-(-12)=80$ °C. At exactly -12 °C a melting sample can contain both solid and liquid.

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■ **Common error**

A gas is not always invisible empty space without particles. Particles do not expand into larger atoms when heated. A single atom has no melting point in the bulk-material sense. Temperature need not rise throughout heating when energy is used in a change of state.

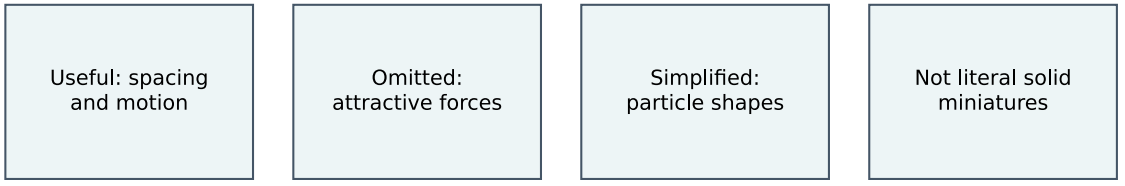
Higher Tier: what the solid-sphere particle model omits

Higher Tier focus; see section scope notes.

Official-unit focus: 4.2 Bonding, structure and properties

A drawing of identical solid balls can show particle spacing, but it cannot by itself explain why one substance boils at a different temperature from another. The missing attractions matter.

The simple particle model represents particles as solid inelastic spheres without forces between them. It usefully compares arrangement, spacing and motion in solids, liquids and gases. Its limitations include absent attractive forces, identical spherical shapes and the impression that particles are filled solid objects. Real particles can be atoms, ions or molecules with different shapes and internal structures; an atom itself is mostly empty space.



Without attractions, the drawing cannot explain energy needed to separate particles or the different melting and boiling temperatures of substances. Treat the balls as symbols for particles, not literal hard miniature pieces of bulk material. A model may succeed at showing close liquid particles while needing an extended force account to explain boiling. Individual particles do not acquire the substance's bulk hardness or melting point.

Compare the same state-change story using a sphere diagram and a labelled force model. For every criticism state what is omitted and the consequence for the explanation. Keep useful features such as particle-number conservation. A model improvement is a justified new representation, not simply adding many details with no explanatory

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purpose.

■ Checked worked case

Known: a supplied model drawing has twelve particles before melting and twelve after. The model conserves particle number while changing arrangement. If a second drawing has eighteen spheres after melting with no particles entering, it incorrectly suggests particle creation: the excess is 18−12=6. This arithmetic tests the diagram, not a physical multiplication of atoms during melting.

■ Common error

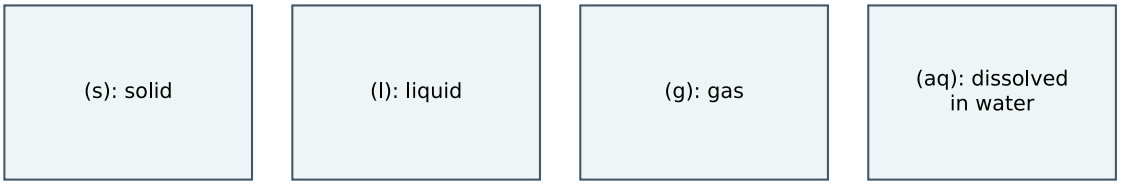
Saying only that the model is unrealistic gives no specific limitation. A force-free model cannot explain different attraction strengths, but this does not make every spacing inference invalid. The specification's explicit evaluation clause is Higher-only; ordinary state predictions remain both-tier.

State symbols: aqueous is a solution, not a pure liquid

Official-unit focus: 4.2 Bonding, structure and properties

NaCl(s) and NaCl(aq) have the same formula but describe different physical forms. The symbol aq means dissolved in water, not a fourth state of pure matter.

Chemical equations use (s) for solid, (l) for liquid, (g) for gas and (aq) for aqueous solution. Pure water is $H_2O(l)$, while dissolved sodium chloride is NaCl(aq). A gas can also dissolve: hydrogen chloride gas is HCl(g), whereas hydrochloric acid solution is HCl(aq). The state depends on conditions and the substance's form, not its formula alone.



For an aqueous precipitation reaction, $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$ distinguishes the new solid from dissolved substances. Magnesium reacting with acid can be written $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$. Correct state labels add evidence about the reaction while the coefficients still conserve atoms.

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Read the stated conditions and observation before labelling: a collected gas, a precipitate, a pure liquid or a solution. Do not infer that every formula containing H is gaseous or that every liquid mixture should be labelled (l). Compare balanced equations with observations from a teacher-approved demonstration and state what each label tells the reader.

■ Checked worked case

Known: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ contains one solid species, two aqueous species and one gas species. There are four species entries even though the HCl coefficient is two. Counting species, atoms and coefficients answers different questions: six Mg atoms correspond to six H_2 molecules in this balanced particle ratio.

■ Common error

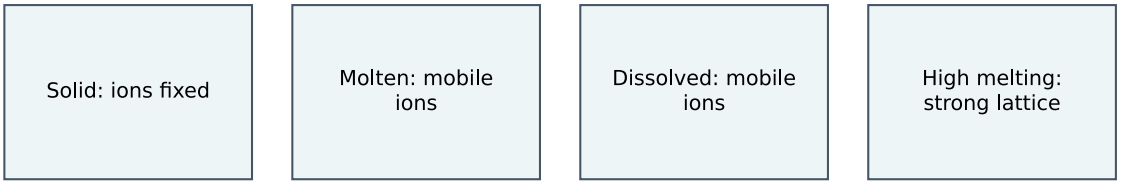
Aqueous does not mean the solute has melted. A precipitate is solid even while surrounded by liquid. State symbols do not change an element symbol or balance a wrong equation. A state must be checked against the supplied temperature, pressure and solvent information.

Ionic properties: strong lattices and mobile ions

Official-unit focus: 4.2 Bonding, structure and properties

A solid salt crystal does not conduct, yet its melt can. The ions have not first appeared on melting; their ability to move has changed.

Ionic compounds have regular giant lattices with strong electrostatic attraction between opposite charges in all directions. Large amounts of energy are needed to overcome many attractions, giving high melting and boiling points. In a solid the ions remain in fixed positions, apart from vibration, and cannot carry charge through the material. When molten, the ions are free to move and conduct. When an ionic compound dissolves in water, mobile ions in the solution can also carry charge.



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The carrier is an ion, not a free electron travelling through the salt. Dissolution separates ions into the solution; it does not simply melt the salt. Not every ionic compound is readily soluble, so explain aqueous conductivity only when a dissolved sample is specified. High melting point supports a strong giant structure, while state-dependent conductivity strengthens the ionic interpretation.

Compare teacher-provided data for solid, molten and dissolved samples. High-temperature molten-salt demonstrations require approved equipment and supervision; students can analyse supplied observations instead. For low-voltage solution tests, keep concentration, electrode spacing and immersion depth comparable and use clean apparatus. A brightness comparison is qualitative unless current is actually measured.

■ Checked worked case

Known: a supplied solution test gives currents 0.18 A and 0.06 A under the same applied voltage and geometry. The current ratio is $0.18/0.06=3$. The first sample carries charge at three times the measured rate in this setup, but the result alone does not determine lattice strength or prove the solutions have equal concentration.

■ Common error

Solid ionic compounds still contain charged ions. Melting does not produce delocalised metallic electrons. Conductivity data must state the sample's form; testing an insoluble powder in water does not establish the properties of a genuinely dissolved solution.

Small molecules: distinguish bonds from intermolecular forces

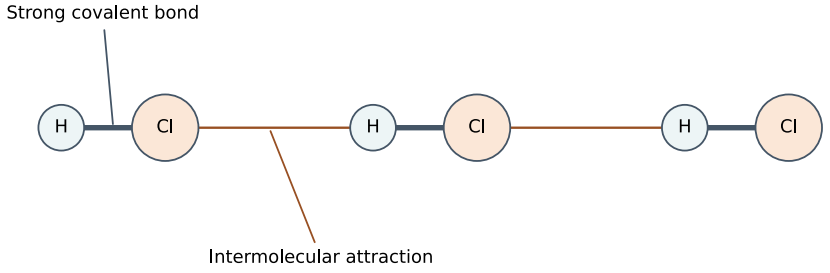
Official-unit focus: 4.2 Bonding, structure and properties

Water can boil without splitting into hydrogen and oxygen. The change separates molecules while the covalent bonds within each molecule remain.

Small molecular substances usually have relatively low melting and boiling points and are often gases or liquids. Strong covalent bonds hold atoms inside each molecule, but weaker intermolecular forces act between molecules. Melting or boiling overcomes these weaker attractions, rather than breaking every covalent bond. Typical neutral molecular substances do not conduct because they have no mobile charged particles.

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Small molecules: strong bonds within; weaker attractions between



Boiling separates molecules; it does not normally break their H–Cl bonds.

For comparable molecules, intermolecular forces generally increase with molecular size, so larger molecules tend to have higher melting and boiling points. Molecular shape and other interactions also matter; a size trend is not an exact universal temperature rule. A compound such as hydrogen chloride can form ions on dissolving, so the pure molecular substance and its reacting aqueous solution must be distinguished.

Use paired molecular drawings that label strong internal bonds and weaker attractions between separate molecules. Identify what is overcome during a specified physical change and what remains intact. Compare supplied boiling-point data for a related series at the same pressure. Do not infer a substance's identity from boiling point alone without composition and other evidence.

■ Checked worked case

Known: supplied rounded boiling points for a fictional related series are −10 °C and 30 °C. The increase is 30−(−10)=40 °C. This is consistent with stronger attractions in the larger member if other relevant features are comparable, but it does not show that an internal covalent bond gained forty units of strength.

■ Common error

Low boiling point does not mean weak covalent bonds. Molecules are usually electrically neutral overall, not devoid of protons and electrons. Dissolving a molecular solute does not always create ions, while a chemical reaction with water can change which particles are present.

Polymers: long molecules and attractions between chains

Official-unit focus: 4.2 Bonding, structure and properties

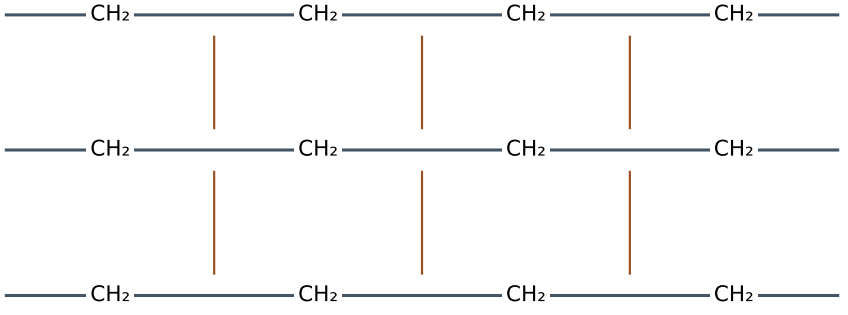
A polymer diagram can show only a few repeat units even though each real molecule contains many. The drawn fragment is not evidence that the material is a giant

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covalent network in all directions.

Polymers have very large molecules made from repeated units. Atoms within each polymer molecule are joined by strong covalent bonds. Attractions between these large molecules are relatively strong compared with those between typical small molecules, so the polymers in this GCSE model are solids at room temperature. A bracketed repeat unit with n large indicates a long molecule continuing beyond the drawn segment.

Separate long covalent chains: shown fragments continue beyond the drawing



Solid lines: covalent backbone. Dotted lines: attractions between chains.

Distinguish bonds along a chain from attractions between separate chains. The size of polymer molecules gives many opportunities for intermolecular attraction. A simplified chain picture helps explain the solid material, but real polymer properties also depend on chain arrangement and other structural features. Do not replace intermolecular attractions with a claim that every neighbouring chain is necessarily joined by covalent bonds.

Compare a small molecular diagram, several long chain fragments and a covalent network. Identify repeated units, chain boundaries and the type of connection shown. Use prepared polymer samples or teacher-provided images, not unapproved heating or burning tests. The $[–CH_2–CH_2–]_n$ model shows poly(ethene); the two bonds through the brackets show continuation of the chain.

■ Checked worked case

Known: two model poly(ethene) chains each contain eighty C_2H_4 repeat units. The pictured repeat-unit content is $2\times80\times2=320$ carbon atoms and $2\times80\times4=640$ hydrogen atoms, with chain ends omitted. That large count does not mean carbon atoms in one chain are covalently connected to every atom in the other chain.

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■ Common error

The label n is not an atom symbol. Polymer molecules can be long while the material contains many separate chains. A drawn dotted attraction between chains is different from a line representing a covalent backbone bond. This lesson's simple solid-polymer model does not claim all polymer materials have identical softening temperatures.

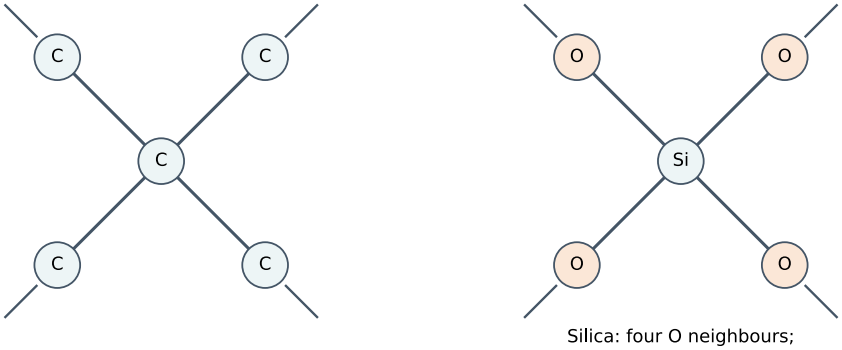
Giant covalent structures: connections continue through the solid

Official-unit focus: 4.2 Bonding, structure and properties

A drawing showing one silicon atom bonded to four oxygens is only a local piece of silica. The oxygen bridges continue into a network, so the pictured fragment is not a separate SiO_4 molecule.

Diamond, graphite and silicon dioxide have giant covalent structures. Strong covalent bonds join atoms through large repeating structures, rather than a substance made of small independent molecules. Many strong covalent bonds must be overcome to melt these structures, requiring much energy and producing very high melting points. They are solids under ordinary conditions.

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: SiO_2 overall.

In silicon dioxide, each silicon is linked to four oxygens and each oxygen bridges two silicons, producing the overall Si:O ratio 1:2. A local drawing may show four surrounding oxygens because each is shared with another silicon elsewhere in the network. Graphite has strongly bonded layers but weaker attractions between layers; its high-temperature behaviour and layer sliding concern different connections.

Trace bonds from one atom to its neighbours and then onward beyond the fragment. Distinguish an uninterrupted network from separate groups with gaps between them. Compare a local coordination drawing and the bulk formula. Models may omit remote

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bonds, distort angles or enlarge spaces; explain what the representation does and does not establish.

■ Checked worked case

Known: a stated representative silica region contains twelve silicon atoms and twenty-four oxygen atoms. The ratio 12:24 reduces to 1:2, giving SiO_2. It is a ratio in a continuing structure, not proof of twelve independent SiO_2 molecules. A local four-oxygen coordination shell cannot be counted without the sharing information to infer the bulk formula.

■ Common error

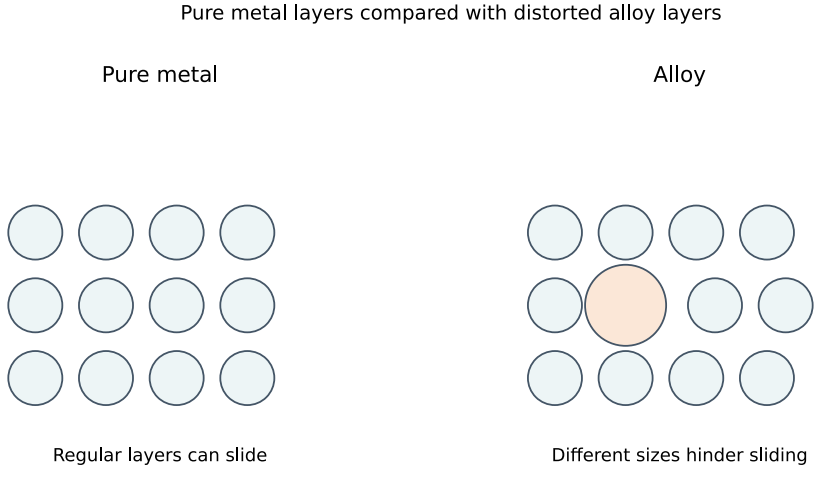
Giant structure does not mean giant individual atoms. Melting a covalent network differs from boiling a small molecular substance because the relevant connections differ. Not all giant covalent substances are electrical insulators: graphite is a key exception with delocalised electrons.

Metals and alloys: distort the layers to resist sliding

Official-unit focus: 4.2 Bonding, structure and properties

A pure metal's regular layers can slide when a force is applied. Introducing different-sized atoms can hinder that sliding, helping explain why an alloy is harder.

Most metals have high melting and boiling points because their giant structures have strong metallic bonding. Atoms in a pure metal are represented in regular layers, which can move past one another while attraction to delocalised electrons persists. This permits bending and shaping. Pure metals can be too soft for particular uses, so they are mixed with other elements to form alloys.



Different-sized atoms distort the regular layers and make sliding more difficult, so many alloys are harder than the corresponding pure metal. Hardness is resistance

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to indentation or scratching, while strength concerns resisting deformation or failure under a load; do not treat the words as identical measurements. An alloy still has metallic bonding and need not have one fixed compound formula.

Compare equal-size atom layers with a model containing some larger atoms. Show the disruption and explain its effect rather than merely saying alloy atoms are bigger. Use supplied hardness data and actual teacher-approved samples with suitable tools. Compare the same test method and conditions; an arbitrary harder sample does not establish how every possible alloy behaves.

■ Checked worked case

Known: supplied model hardness values are 40 units for a pure metal and 70 for an alloy measured by the same method. Increase=70−40=30 units; percentage increase=30/40×100=75%. These fictional values quantify one comparison and are not universal values or a complete explanation of the structural cause.

■ Common error

An alloy is a mixture, not automatically a new compound with fixed proportions. Added atoms hinder sliding without removing all delocalised electrons. A layer model omits real defects and three-dimensional detail, and high melting point is typical rather than universal for metals.

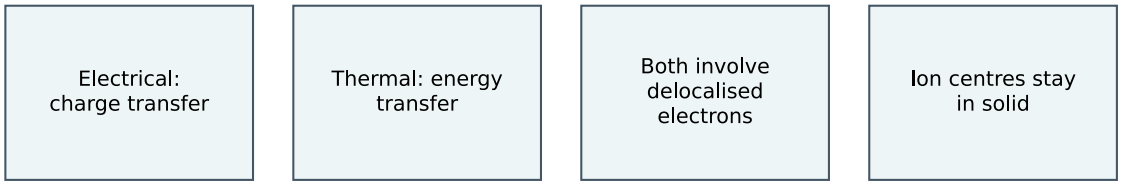
Metal conduction: electrons transfer charge and thermal energy

Official-unit focus: 4.2 Bonding, structure and properties

A metal can conduct without melting. Its delocalised electrons are already mobile in the solid and can transfer charge and energy through the structure.

Metals are good electrical conductors because delocalised electrons can move through the giant structure and carry electrical charge. When a potential difference is applied, the electrons have a net drift that supports current. In a solid metal, the positive ion centres are not the mobile current carriers. Metals are also good thermal conductors because delocalised electrons transfer energy through the material.

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Electrical conduction is transfer of charge; thermal conduction is transfer of energy from a hotter region toward a cooler region. Do not confuse electron drift with the whole wire moving or with electrons being permanently consumed. A metal can conduct thermal energy without being part of an electric circuit. The two explanations share mobile electrons but refer to different observations.

For a supervised low-voltage electrical comparison, keep geometry and temperature controlled and measure voltage and current if interpreting quantitative data. For heat transfer use teacher-approved apparatus and handling, because hot metal may look unchanged. Analyse supplied measurements if direct heating is unsuitable. The simple carrier diagram does not show the detailed speed or microscopic energy distribution.

■ Checked worked case

Known: under a supplied constant-current model, 0.40 C passes a point in 2.0 s. Charge-transfer rate=charge/time=0.40/2.0=0.20 C/s. This is current 0.20 A. It quantifies electrical transfer only; it does not by itself give the amount of thermal energy conducted or the number of electrons in the entire wire.

■ Common error

Metals do not need mobile negative ions to conduct. Electron charge remains negative during heat transfer, and thermal energy is not itself a new particle substance added to the wire. Conductivity can change with temperature and composition, so geometry alone does not fix a universal value.

Diamond: four bonds per carbon in a rigid network

Official-unit focus: 4.2 Bonding, structure and properties

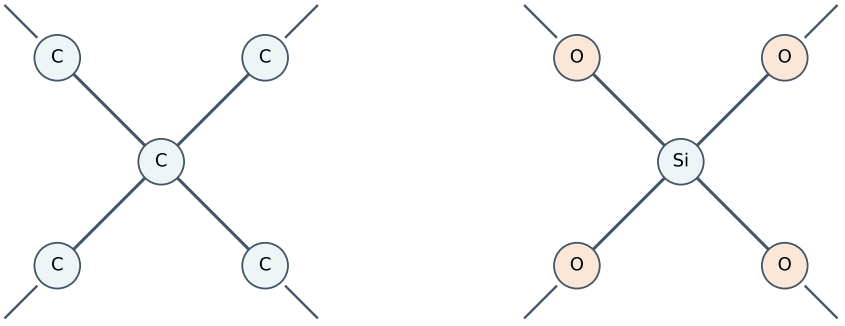
Diamond and graphite both contain carbon, but their bonding arrangements differ. The element name alone does not predict hardness or conductivity.

In diamond each carbon atom makes four covalent bonds with other carbon atoms in a giant three-dimensional structure. Strong connections extend through the solid,

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giving a rigid network and making diamond very hard. Very much energy is needed to overcome many covalent bonds, giving a very high melting point in the specified GCSE account. Its outer electrons are involved in covalent bonds, so it has no delocalised electrons to carry charge and does not conduct electricity.

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: SiO_2 overall.

Explain one property with the appropriate connection: resistance to deformation follows the rigid many-direction network; high melting point follows energy needed to overcome many strong covalent bonds; non-conductivity follows the absence of mobile charged carriers. A drawing with a carbon surrounded by four neighbours shows local coordination, not a complete diamond molecule.

Compare a tetrahedral local model with a graphite-layer model. Count four bonds around an interior diamond carbon and trace continuation beyond the fragment. Mark a two-dimensional picture as a projection of a three-dimensional arrangement. Prepared models or images suffice; there is no need to heat or scratch expensive specimens or infer properties from a gem's colour.

■ Checked worked case

Known: a stated large diamond model contains 120 interior-equivalent carbon sites, each with four bond connections. There are $120 \times 4 = 480$ bond-end counts; each bond has two ends, giving $480/2 = 240$ bonds under this explicitly boundary-free counting convention. Simply multiplying by four double-counts each shared connection.

■ Common error

Diamond is not an ionic lattice of carbon ions. The atoms do not become larger or lose their electrons entirely. An actual finite fragment has boundaries, so the bond-count formula needs the stated interior-equivalent assumption. No claim about current experimental high-pressure melting conditions is required for this GCSE structure explanation.

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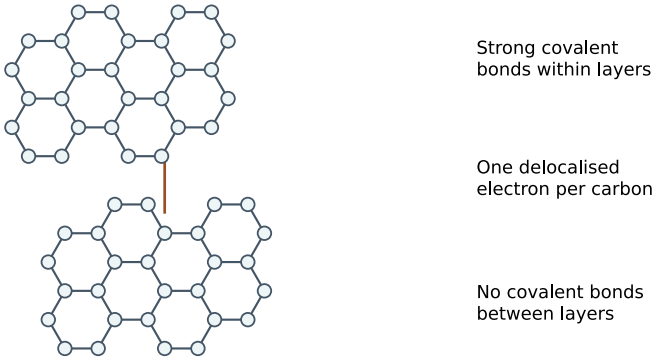
Graphite: strong layers, weak interlayer attractions and mobile electrons

Official-unit focus: 4.2 Bonding, structure and properties

Graphite can be soft enough to leave a pencil mark while still having very strong covalent bonds. Different properties depend on connections within and between its layers.

Each graphite carbon makes three covalent bonds to other carbons in layers of hexagonal rings. Strong bonds inside the layers require much energy to overcome, giving a very high melting point in the GCSE account. There are no covalent bonds between the layers; weaker attractions allow layers to slide past one another, explaining softness and lubricating use. One electron from each carbon is delocalised and can carry charge through the structure.

Graphite: hexagonal covalent layers, with weaker attraction between layers



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

Do not use weak interlayer forces to explain the high melting point: that property involves strong covalent bonding. Conversely, naming strong covalent bonds alone does not explain easy layer sliding. Graphite resembles metals in having delocalised electrons, but its bonding network is layered covalent rather than metallic. Conductivity directions in real graphite are not represented fully by one simple flat drawing.

Inspect a hexagonal-layer model and identify three neighbours for an interior carbon, with bonds continuing at cut edges. Compare layers stacked without covalent connecting lines. A teacher-approved low-voltage test of a prepared graphite rod can provide conductivity evidence. Record contact quality and geometry, and distinguish the rod material from a pencil's mixed graphite/clay core.

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■ Checked worked case

Known: a stated interior graphite-layer model has 90 carbon atoms contributing one delocalised electron each, so it represents 90 such electrons. Its $90 \times 3 = 270$ bond ends correspond to 135 bonds after dividing by two under a boundary-free convention. The electron count and covalent bond count describe different aspects of the same structure.

■ Common error

Graphite does not have four covalent neighbours per carbon like diamond. Weak layer attractions are not absent attraction altogether. A cut edge in a model has fewer drawn neighbours and should not be treated as evidence against the interior three-bond pattern.

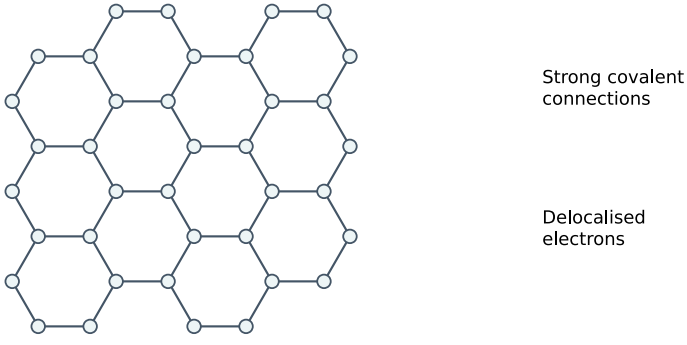
Graphene: one strong conducting carbon layer

Official-unit focus: 4.2 Bonding, structure and properties

A single sheet of graphite's carbon network is graphene. Removing the stacked-layer context changes the material geometry while retaining three covalent neighbours around an interior carbon.

Graphene is a single layer of graphite, a very thin sheet of carbon atoms arranged in hexagonal rings. Each interior carbon is strongly covalently bonded to three others, making the sheet strong. Delocalised electrons allow electrical conduction. Its thinness, strength and conductivity make it useful in electronics and composites, where a material's properties must fit the intended function.

Graphene: one hexagonal carbon layer



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

A composite combines materials so their useful properties can contribute to a designed product. Graphene can reinforce a material, but strength claims need actual data for the particular composite, loading and manufacturing method. Conductivity supports

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electronic applications without proving every graphene-containing product is automatically a good conductor. A drawn lattice edge is a clipped boundary, not the full bulk bonding environment.

Compare a single honeycomb sheet, several graphite layers and a hollow fullerene shape. Identify graphene from the single-layer description and explain one property with bonding or electrons. Use supplied product data rather than asking students to handle loose nanopowders or claiming laboratory-quality graphene has been produced from an ordinary pencil drawing.

■ Checked worked case

Known: a supplied composite test raises supported load from 80 N to 100 N for equal-size samples under one method. Increase=20 N; percentage increase=20/80×100=25%. This illustrates evaluating a particular reinforcement claim, not a universal graphene strength increase or proof of identical performance under a different load direction.

■ Common error

Graphene is not a stack of many layers and is not the same as a hollow C_{60} molecule. A high-strength carbon sheet does not mean every graphene composite has diamond's hardness. Property explanations and commercial performance evidence answer different questions.

Fullerenes and nanotubes: hollow carbon shapes and uses

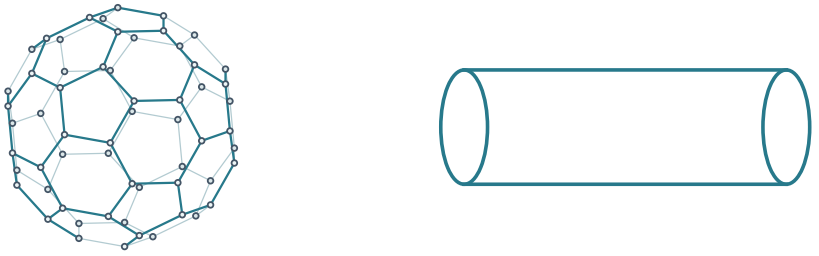
Official-unit focus: 4.2 Bonding, structure and properties

A carbon nanotube has a long hollow cylindrical shape, whereas buckminsterfullerene is approximately spherical. Their shape and dimensions matter when selecting applications.

Fullerenes are molecules of carbon atoms with hollow shapes. Their structures are based on hexagonal carbon rings and may also contain five- or seven-membered rings. Buckminsterfullerene, C_{60}, was the first discovered fullerene and has a spherical cage shape. Carbon nanotubes are cylindrical fullerenes with very high length-to-diameter ratios. Strong carbon bonding and useful electrical/material properties support nanotechnology, electronic and composite applications.

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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.

A hollow molecular cage can be considered for carrying other substances in proposed delivery systems; nanotubes can be used as reinforcing fibres and in electronic materials. A use must be linked to supplied properties and tested for suitability. Do not claim every nanotube has identical conductivity or that a proposed medical use is automatically safe or effective. Graphene is a sheet, graphite stacked layers and a fullerene a hollow shape.

Compare labelled structural descriptions with shape drawings. Distinguish a schematic outline showing geometry from a complete atom-and-bond map: a circle alone does not establish exactly sixty carbons. For nanotube data convert length and diameter into the same unit before finding their ratio. Use contained prepared materials or reference images rather than dispersing loose nanomaterials.

■ Checked worked case

Known: a supplied nanotube is 2.0 micrometres long and 4.0 nm in diameter. Length=2,000 nm, so length/diameter=2,000/4.0=500. The large ratio quantifies its slender shape. Ten separate C_{60} molecules contain $10\times60=600$ carbon atoms, not one giant molecule with an unspecified sixty-unit repeat.

■ Common error

A hollow shape does not make every fullerene an empty container safe for any use. C_{60} is a molecular formula, not a polymer repeat label. Diagrams of cylinders need an explicit schematic label unless actual carbon bonds and topology are correctly shown.

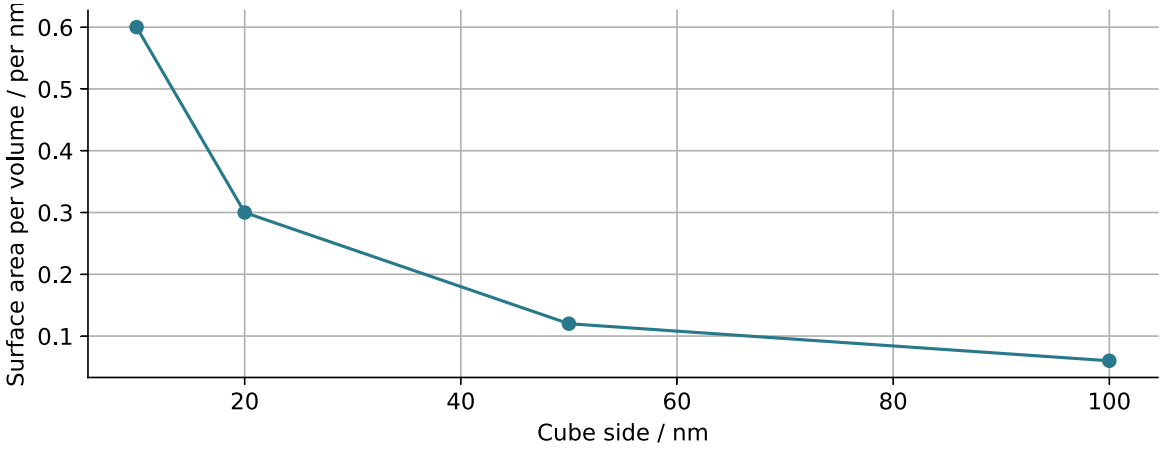
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Chemistry-only: nano sizes and surface-area-to-volume ratio

Official-unit focus: 4.2 Bonding, structure and properties

Breaking a fixed volume into smaller cubes exposes more surface without creating more total material. This helps explain why a small quantity of nanoparticulate material can be effective.

Nanoscience concerns structures around 1–100 nm. One nanometre is 1×10^{-9} m. The specification uses nanoparticles below fine particles, whose diameters are 100–2,500 nm, and coarse particles at approximately 2,500–10,000 nm. These are its size bands; PM labels in wider air-quality use refer to upper-size cutoffs and should not be silently substituted for the stated classroom intervals. Typical atoms have dimensions of order 0.1 nm, so even a 10 nm particle spans many atomic dimensions.



For a cube side L, area=$6L^2$, volume=L^3 and area/volume=$6/L$. Reducing side length tenfold raises area/volume tenfold. Nanoparticles can have different properties from the same bulk material because a much larger fraction interacts at the surface. More exposed area can improve catalytic effectiveness per mass, but particle shape, aggregation and surface chemistry also affect performance.

Use cubical models as a mathematical illustration and keep length units consistent. Area/volume is a ratio with reciprocal-length units, not simply a percentage. Compare a nano dimension with a stated atomic size by division. A cube model does not imply real nanoparticles are all cubes or all have the same chemical behaviour.

■ Checked worked case

Known: cubes have side 10 nm and 100 nm. The first has area 600 nm^2, volume 1,000 nm^3 and area/volume 0.6 per nm. The second has 60,000 nm^2, 1,000,000 nm^3 and ratio 0.06 per nm. The smaller cube's ratio is ten times larger. A 10 nm length is 10/0.1=100 typical atomic radii under the supplied comparison.

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■ Common error

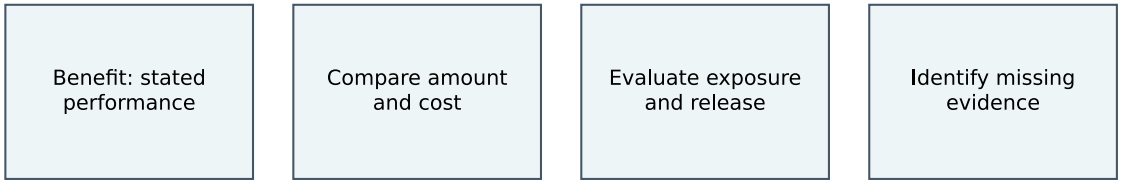
A high surface ratio does not mean a larger total mass. Prefix conversion can change an answer by factors of a thousand. Atom radius versus diameter must be stated in a comparison; the worked ratio explicitly compares length with radius. All risks and properties cannot be inferred from size alone.

Chemistry-only: evaluate a nanoparticle application from evidence

Official-unit focus: 4.2 Bonding, structure and properties

A product may use less active material at nanoscale, yet a performance benefit does not settle questions about exposure or disposal. Evaluate the particular purpose using the information supplied.

Nanoparticles are investigated and used in medicine, electronics, cosmetics and sun creams, deodorants and catalysts. A high surface-area-to-volume ratio can make smaller amounts effective, and size-related properties can provide useful functions. The specification asks students to evaluate an application from given information and recognise possible risks; it does not require memorising a universal list of properties for every nanomaterial.



Consider useful performance, required quantity, cost and durability alongside possible human exposure, environmental release and uncertain long-term effects. Risk depends on the material, particle form, exposure route and amount. A contained particle in a product and an inhalable loose powder need separate exposure evidence. A proposed benefit is not proof of safety, while a possible risk is not proof that harm occurs at every dose.

Use a supplied product evidence table, identifying measured outcomes, comparison conditions and missing information. Weigh advantages and disadvantages for the intended function and make a qualified recommendation supported by evidence. Do not ask students to make a cosmetic, administer a nanomaterial or generate nanopowder.

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Any actual school materials require the approved method and containment.

■ Checked worked case

Known: a fictional catalyst reaches the same measured conversion using 0.30 g nanoscale material instead of 1.20 g bulk material under stated conditions. Mass reduction=(1.20−0.30)/1.20×100=75%. That comparison supports lower required quantity in this trial; it says nothing by itself about inhalation risk, lifespan, cost or environmental release.

■ Common error

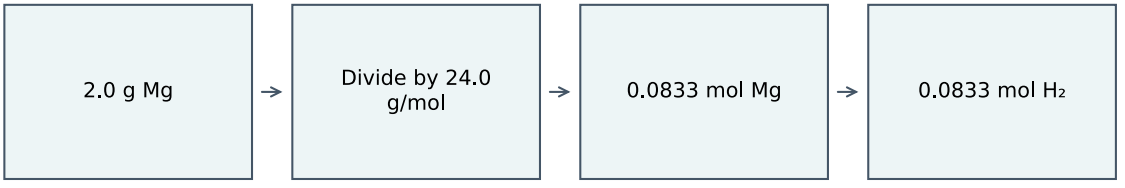
Do not infer clinical effectiveness from a classroom surface calculation. Evaluate the provided data rather than claiming nano always means better or more dangerous. A conclusion can identify a promising function while requiring stronger exposure or disposal evidence before wider use.

Amounts, equations and limiting reagents

Official-unit focus: 4.3 Quantitative chemistry

The smallest mass of reactant is not necessarily the limiting reagent. The balanced equation compares particle amounts, not grams directly.

The mole measures amount of substance. Use molar mass to convert mass into amount. Balanced equation coefficients give mole ratios; they do not give equal masses.



Calculate the amount available for each reactant and divide by its coefficient. The smaller ratio limits the reaction. Use that reactant to calculate the maximum product before comparing actual yield.

Write the balanced equation first, include units in molar masses, then convert each given mass or solution volume into amount. Convert cubic centimetres to cubic decimetres before using concentration in moles per cubic decimetre.

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■ Checked worked case

Known: 2.0 g of Mg reacts with excess acid. Use amount = mass/molar mass. With molar mass Mg = 24.0 g per mole, amount Mg = 2.0/24.0 = 0.0833 mol. In Mg + 2HCl → MgCl2 + H2, amount H2 = amount Mg = 0.0833 mol.

■ Common error

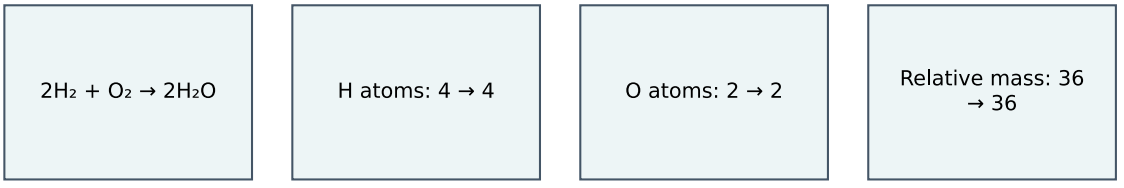
Excess acid means acid does not limit the stated calculation. A coefficient of 2 before HCl does not double the hydrogen amount.

Conservation of mass: count the complete system

Official-unit focus: 4.3 Quantitative chemistry

A sealed reaction vessel weighs the same before and after a chemical reaction even when a gas appears. The atoms have been rearranged inside the weighed system.

The law of conservation of mass states that atoms are neither made nor lost in a chemical reaction. The total mass of products equals the total mass of reactants when every substance is included. A balanced symbol equation has the same number of atoms of each element on both sides. It does not require the same number of separate molecules on each side.



For $2H_2 + O_2 \rightarrow 2H_2O$, the coefficient 2 multiplies an entire formula. The subscript 2 belongs to the preceding element, so one water molecule contains two H atoms and one O atom. Both sides contain four H atoms and two O atoms. Using supplied relative atomic masses H=1 and O=16 gives reactant totals 2×2 + 32 =36 and product total 2×18=36. The coefficients 2:1:2 are not gram ratios.

Make an atom-count table before doing a mass calculation. Identify the physical boundary of the weighed system: vessel, liquid, solid and any contained gas. For an actual supervised demonstration, use an approved arrangement that accommodates any pressure change; do not improvise a heated sealed vessel. Compare readings only after correcting for anything added or removed from the balance.

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■ Checked worked case

Known: an enclosed model reaction starts with 4.0 g A and 7.0 g B, and has no other reactants. The combined products weigh 11.0 g. If one product weighs 8.5 g, the other weighs 11.0−8.5=2.5 g. This mass accounting uses the complete stated system and does not identify either substance chemically.

■ Common error

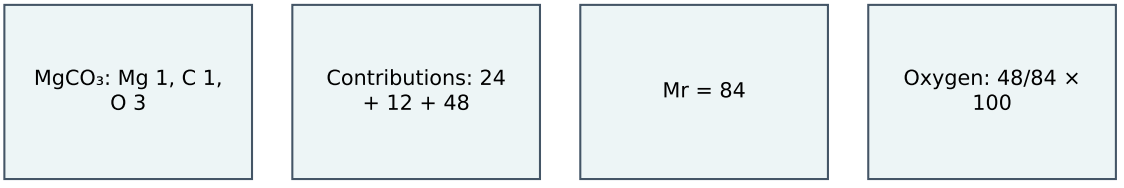
Conservation does not mean that every individual substance keeps its own mass. Reactants are converted into products. Altering a subscript to balance an equation changes the substance, whereas changing a coefficient changes the number of particles. Include material crossing the system boundary before calling a result a failure of conservation.

Relative formula mass and the mass percentage of an element

Official-unit focus: 4.3 Quantitative chemistry

Calcium hydroxide contains more than one oxygen and hydrogen atom. Reading its brackets correctly is essential before calculating either formula mass or calcium percentage.

Relative formula mass, Mr, is the sum of the relative atomic masses of all atoms shown in a chemical formula. Relative atomic and formula masses are comparative numbers without gram units. For $Ca(OH)_2$, supplied Ar values Ca=40, O=16 and H=1 give Mr=40+2×(16+1)=74. The subscript outside the brackets multiplies both O and H.



An element's percentage by mass is its total relative mass contribution divided by Mr, multiplied by 100. Calcium contributes 40 of the 74 units in $Ca(OH)_2$, giving 40/74×100=54.1% to three significant figures. Oxygen contributes 32 and hydrogen 2. These percentages describe the compound's fixed mass proportions, not the proportion of different atoms counted equally.

List element, atom count, Ar and mass contribution in four columns. Add contri-

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butions for Mr, then select the requested element's contribution for the numerator. In a balanced equation include coefficients: $2Mg + O_2 \rightarrow 2MgO$ gives $2\times24+32=80$ on the left and $2\times40=80$ on the right. Use the periodic-table values supplied with a particular question rather than guessing more precise data.

■ Checked worked case

Known: $MgCO_3$ has supplied Ar values Mg=24, C=12 and O=16. Mr=24+12+3×16=84. Oxygen contributes 48, so oxygen percentage=48/84×100=57.1%. The carbon percentage is 12/84×100=14.3%; it is not one fifth just because one of five atoms is carbon.

■ Common error

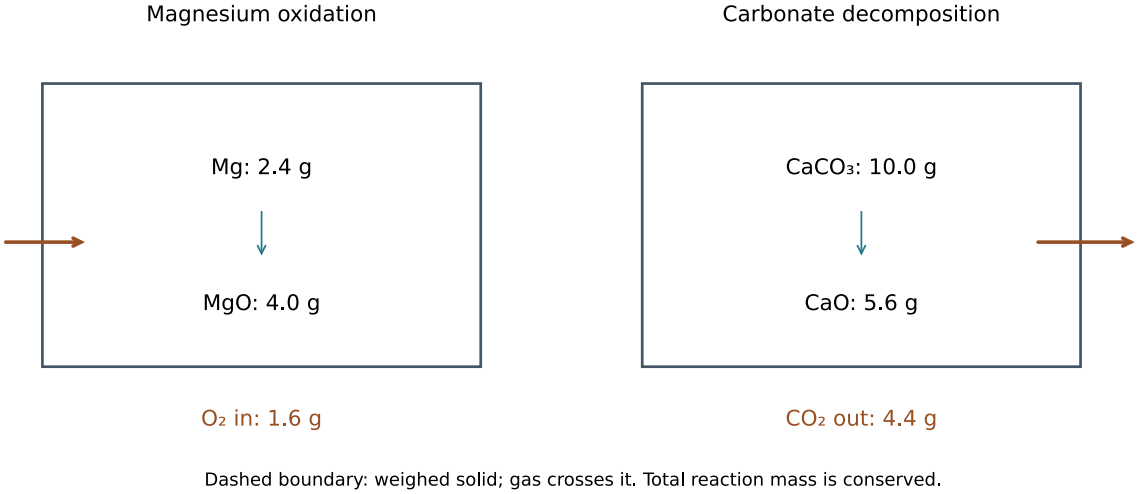
Do not multiply the whole formula by an internal subscript. A coefficient belongs to an equation amount, not the Mr of one formula unit. Percentages should total about 100, allowing rounding. A formula mass is not automatically a measured sample mass, and percentage by mass differs from percentage of atoms.

An open vessel: explain mass entering or escaping as gas

Official-unit focus: 4.3 Quantitative chemistry

A heated metal can gain mass, while a heated carbonate can leave a lighter solid. Neither observation contradicts conservation when gas transfer is counted.

In an open system, gas may enter from or escape to the surrounding air. Magnesium reacts with oxygen as $2Mg + O_2 \rightarrow 2MgO$. The oxide contains the original magnesium plus oxygen from outside, so it weighs more than the original metal. Calcium carbonate decomposes as $CaCO_3 \rightarrow CaO + CO_2$. The remaining solid is lighter because carbon dioxide has escaped, not because atoms have vanished.



Dashed boundary: weighed solid; gas crosses it. Total reaction mass is conserved.

Identify which material is on the balance at each reading. In the metal example,

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oxygen joins the weighed solid; in the carbonate example, gas leaves the weighed apparatus. In particle terms the same atoms are redistributed among substances, with some crossing the selected system boundary. A sealed-system total includes all products; an open-residue measurement does not.

For a teacher-approved school investigation, record the clean dry vessel mass, sample mass and final cooled vessel-plus-solid mass. Use appropriate heat protection, ventilation and an approved method. Loss of solid through spitting or incomplete reaction can change the reading too. Do not seal heated gas-producing apparatus merely to demonstrate conservation, and do not assume every measured difference is the intended gas transfer.

■ Checked worked case

Known: 10.0 g of a model carbonate leaves 5.6 g solid oxide after complete decomposition with no other losses. Escaped gas mass=10.0−5.6=4.4 g. The 5.6 g oxide plus 4.4 g gas equals the original 10.0 g. For 2.4 g Mg forming 4.0 g MgO, oxygen gained=1.6 g.

■ Common error

A gas has mass even if it is invisible. Mass gain does not imply extra magnesium atoms were created. Cooling before weighing reduces temperature-related weighing problems. If apparatus mass changes or solid is spilled, the simple gas-only explanation needs revision rather than being forced to fit the result.

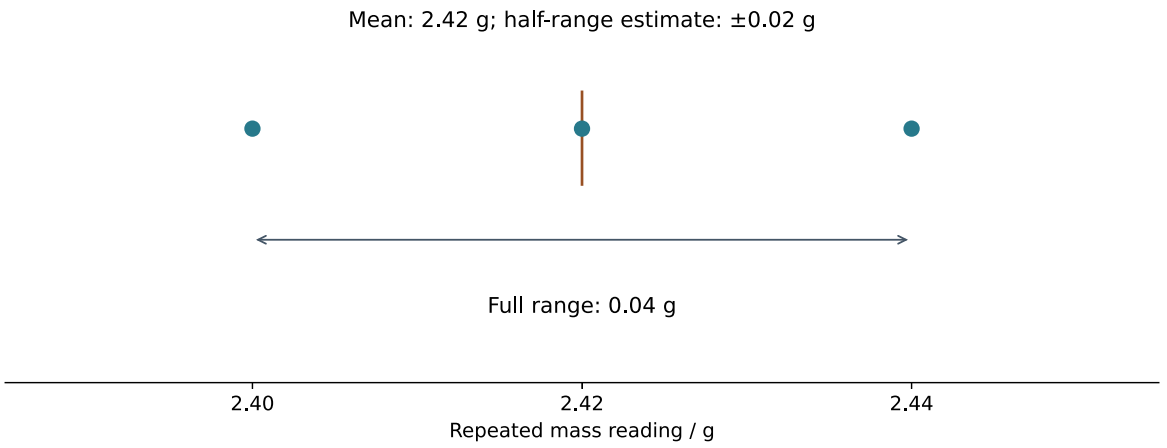
Repeated measurements: mean, spread and estimated uncertainty

Official-unit focus: 4.3 Quantitative chemistry

Repeated mass measurements can differ slightly even when the same method is used. Reporting only the mean hides evidence about that spread.

Every measurement has some uncertainty. Repeat readings help show the distribution of results. Calculate the mean by adding the readings and dividing by their number; calculate the range as highest minus lowest. A common school estimate is half the range, reported as mean plus or minus that estimate when the distribution supports it. State that convention rather than claiming it is a universal statistical confidence interval.

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For 2.40, 2.42 and 2.44 g, the mean is 2.42 g and the full range is 0.04 g. Half-range is 0.02 g, so the stated estimate is 2.42±0.02 g. Here the mean lies midway between the extremes. For an asymmetric distribution the extreme deviations from the mean differ; display the actual results as well as the chosen uncertainty estimate.

Plot a dot for each repeat on a labelled mass axis; coincident results can stack vertically. Keep units and sensible decimal places. Investigate an unusual value before excluding it, record any exclusion reason and preserve the original readings. More repeats reveal random spread but cannot by themselves remove a systematic bias, such as an uncorrected balance zero.

■ Checked worked case

Known: 5.1, 5.2, 5.2 and 5.3 cm give mean 5.2 cm and range 0.2 cm. Under the stated half-range convention, estimated uncertainty is ±0.1 cm. This says something about the observed repeat spread. It does not prove the true length is inside that interval or that the ruler's resolution equals 0.2 cm.

■ Common error

Uncertainty is not the same as a mistake. Zero range from a coarse instrument does not establish perfect measurement. Do not halve an instrument resolution and label it repeat range. Compare like quantities and express uncertainty in the same unit as the measured value.

Higher Tier: moles connect mass to stated particles

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

One mole of carbon atoms and one mole of carbon dioxide molecules contain the same number of stated particles, although their masses and atom contents differ.

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Chemical amount is measured in moles, symbol mol. One mole contains 6.02×10^{23} of the stated particles, the GCSE value of the Avogadro constant. For a substance with relative formula mass Mr, one mole has mass numerically Mr grams; its molar mass has unit g per mol. Calculate amount n=mass in grams divided by molar mass. For 11 g CO_2 with Mr=44, n=11/44=0.25 mol.



State what is being counted. A mole of CO_2 molecules contains one mole of carbon atoms and two moles of oxygen atoms, three moles of atoms altogether. A mole of NaCl formula units corresponds to one mole of Na^+ ions and one mole of Cl^- ions in the ionic lattice. Calling either sample simply a mole of particles without specifying which particles can make an answer ambiguous.

Write mass, molar mass and amount in separate columns. Convert kilograms or milligrams into grams before dividing. Rearrange n=m/M to m=nM or M=m/n as needed. Calculate particle number as $n\times6.02\times10^{23}$ and retain standard form; divide particle number by the same constant to find amount. Use sensible significant figures and keep relative mass distinct from molar-mass units.

■ Checked worked case

Known: 0.50 mol water has mass $0.50\times18=9.0$ g and contains 3.01×10^{23} water molecules. Each molecule has two H atoms, so the hydrogen-atom count is 6.02×10^{23}. The three atoms per molecule do not change the number of water molecules in the sample.

■ Common error

A mole is a counting amount, not a fixed gram mass for every substance. Do not use Mr as if it already contained a gram unit. The Avogadro value refers to stated particles, and multiplying by the number of atoms in each molecule is a separate step.

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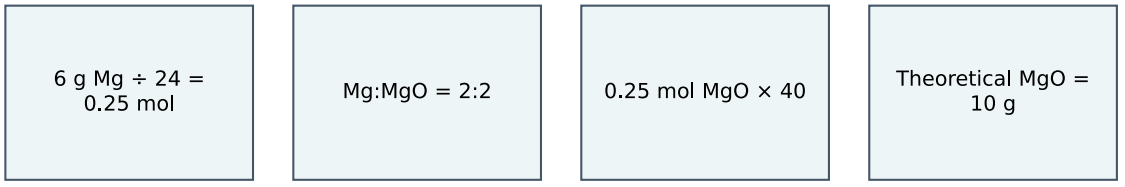
Higher Tier: use the equation ratio between two mole amounts

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

A balanced equation links amounts of different substances. It cannot be used by multiplying the starting mass directly by the product coefficient.

Use three steps: convert the known mass to moles, apply the balanced coefficient ratio, then convert the target amount to mass. For $Mg + 2HCl \rightarrow MgCl_2 + H_2$, one mole Mg requires two moles HCl and forms one mole each of $MgCl_2$ and H_2. The ratio concerns amounts, while each substance has its own molar mass.



For $2Mg + O_2 \rightarrow 2MgO$, supplied molar masses Mg=24, O_2=32 and MgO=40 g per mol give mass proportions 48:32:80. These differ from coefficients 2:1:2. A calculation from one reactant assumes sufficient other reactant and complete reaction unless the question supplies a limiting amount. If the known quantity is product mass, work backwards through the same ratio.

Balance the equation first and underline the known and requested substances. Write a mole row underneath the coefficients, labelling every substance. Avoid rounding a small mole amount early. State the sufficient-reactant and complete-conversion assumptions. Numerical theoretical masses do not include recovery loss, side reactions or reversible-equilibrium effects.

■ Checked worked case

Known: 6.0 g Mg gives 6.0/24=0.25 mol Mg. Its coefficient equals MgO's coefficient, so theoretical MgO amount=0.25 mol and mass=0.25×40=10.0 g. Oxygen required is 0.125 mol, mass 0.125×32=4.0 g. The calculated total 10.0 g equals 6.0+4.0 g.

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■ **Common error**

Coefficients do not replace formula masses. A two-to-two coefficient ratio simplifies to one-to-one. A larger product mass can include atoms from another reactant rather than contradict conservation. The mass calculation is a theoretical prediction under stated assumptions, not proof that an experiment will recover every gram.

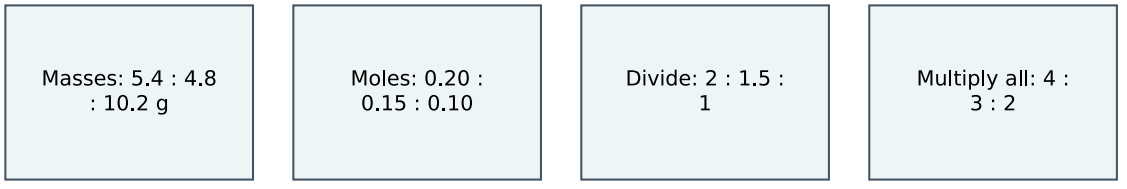
Higher Tier: derive balancing numbers from reacting masses

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

A reaction's measured masses may be in a ratio such as 2.4:1.6:4.0, while its equation coefficients are 2:1:2. Converting masses to amounts reveals the correct ratio.

For known substance formulae, divide each reacting mass by its molar mass to obtain amounts. Divide all amounts by the smallest one, then multiply the entire ratio if needed to obtain small whole numbers. Use those numbers as coefficients before the supplied formulae. Never change the formulae themselves just to make the ratio look simpler.



For Mg, O_2 and MgO with molar masses 24, 32 and 40 g per mol, supplied masses 2.4, 1.6 and 4.0 g give 0.10, 0.050 and 0.10 mol. Dividing by 0.050 gives 2:1:2, hence 2Mg + O_2 → 2MgO. The oxygen formula is O_2, so its molar mass is 32, not the atomic value 16.

Organise substance, mass, molar mass, amount and simplified ratio in a table. When a ratio includes 1.5, multiply every term by two instead of rounding that term to one or two. Check the resulting atom counts and conserved total mass. Reacting masses must exclude unreacted excess or unrelated apparatus; otherwise they cannot directly establish the equation ratio.

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■ Checked worked case

Known: masses 5.4 g Al, 4.8 g O_2 and 10.2 g Al_2O_3 give amounts 0.20, 0.15 and 0.10 mol using molar masses 27, 32 and 102. Ratio=2:1.5:1, so multiply all terms by two to obtain 4:3:2. The equation $4Al + 3O_2 \rightarrow 2Al_2O_3$ conserves four Al and six O atoms.

■ Common error

Whole-number coefficients are an equation ratio, not a demand to round every experimental amount independently. A product may contain a bracketed formula whose full molar mass must be calculated. Rearranging n=m/M gives M=m/n, but a derived M still needs the correct substance identity to interpret it.

Higher Tier: find which reactant limits the product

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

Having more grams of a reactant does not necessarily make it the excess reactant. The equation ratio and molar masses decide which supply runs out first.

The limiting reactant is completely used up under the complete-reaction model and sets the maximum product amount. An excess reactant remains because more was supplied than the balanced equation needs. Convert both reactant masses to amounts before comparing. For $A + 2B \rightarrow C$, one mole A requires two moles B; compare n(A) with n(B)/2, not the raw mole amounts alone.



For $Mg + 2HCl \rightarrow MgCl_2 + H_2$, 0.20 mol Mg would require 0.40 mol HCl. If only 0.30 mol HCl is supplied, acid limits and only 0.15 mol Mg can react. Maximum H_2 amount is 0.15 mol, and 0.05 mol Mg remains. This assumes reaction proceeds as written and does not describe incomplete conversion caused by slow rate or equilibrium.

Write available amount, required coefficient and available amount divided by coefficient for each reactant. The smaller scaled supply determines the maximum reaction extent.

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Multiply that extent by the target product coefficient. To find excess remaining, subtract the amount consumed from the amount supplied. Avoid declaring a limiting reactant solely from a visual guess or whichever mass number is smaller.

■ Checked worked case

Known: 4.8 g Mg is 4.8/24=0.20 mol. With 0.30 mol HCl, magnesium consumed=0.15 mol, mass=3.6 g. Remaining Mg=4.8−3.6=1.2 g. Hydrogen forms at 0.15 mol, mass=0.15×2=0.30 g using supplied molar mass H_2=2 g per mol.

■ Common error

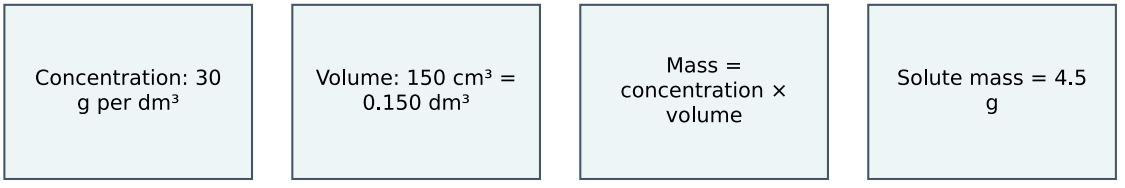
Excess means present beyond the required reacting amount, not an impurity. Adding more excess reactant alone does not raise the theoretical product after the limiting reactant is exhausted. Recovered yield can be lower than this maximum for other reasons; do not confuse limiting amount with percent recovery.

Solution concentration by mass: convert volume before calculating

Official-unit focus: 4.3 Quantitative chemistry

A bottle labelled 8 grams per cubic decimetre contains 8 g solute in each 1 dm^3 of solution. A quarter of that solution volume contains a quarter of the solute mass.

Mass concentration states the mass of dissolved solute in a given volume of solution. With concentration c in grams per dm^3, mass m in grams and solution volume V in dm^3, m=cV. One dm^3 equals 1,000 cm^3. Thus 250 cm^3 equals 0.250 dm^3, not 250 dm^3. Use final solution volume rather than assuming the solvent volume is identical.



For a uniformly mixed solution, the amount of solute taken is proportional to the sample volume. A 0.250 dm^3 sample of an 8.0 grams per dm^3 solution contains 2.0 g solute. The unit calculation cancels dm^3 and leaves grams. This shared-tier numerical task is distinct from the embedded Higher-only explanation of how changing mass or solution volume changes concentration.

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Write the concentration unit beside the value and convert volume before multiplying. Mark whether a number refers to solute, solvent or solution. For an actual teacher-approved preparation, weigh the selected solute, dissolve it and make the solution up to the required final volume using appropriate school apparatus. Check complete dissolution and safe handling; the written calculation does not replace actual supervised technique.

■ Checked worked case

Known: 30 grams per dm^3 solution is sampled at 150 cm^3. Volume=150/1,000=0.150 dm^3, so solute mass=30×0.150=4.5 g. A 300 cm^3 portion contains 9.0 g under the same uniform-concentration condition. The sample's total solution mass is not given by this solute calculation.

■ Common error

Do not multiply concentration by a cm^3 number when the concentration uses dm^3. Dissolved solute is still present even if invisible. A concentration is not the total mass of the bottle, and equal solution volumes need not contain equal solute masses when concentrations differ.

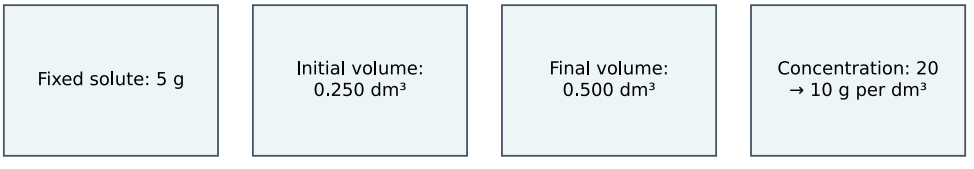
Higher Tier: explain how mass and volume change concentration

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

Adding water to a solution can lower concentration while leaving the dissolved solute mass unchanged. Concentration describes a ratio, not whether solute has disappeared.

Mass concentration c=m/V compares solute mass with final solution volume. At fixed volume, doubling dissolved mass doubles concentration. At fixed dissolved mass, doubling final solution volume halves concentration. If both change, calculate the new ratio: multiplying mass and volume by the same factor leaves concentration unchanged. This explanatory relationship is an embedded Higher-only requirement in 4.3.2.5.



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During simple dilution with pure solvent and no losses or reaction, the solute mass stays constant. Therefore $c_1V_1=c_2V_2$ when both concentrations use the same mass units and both volumes the same volume units. The final volume includes the original sample and added solvent; it is not the volume of water added alone. Dissolution and dilution should not be confused with a chemical transformation of solute.

Identify what is held constant before stating a direction of change. Rearrange c=m/V to V=m/c or m=cV. For a planned school dilution, use appropriate measured transfer and make up to the required final volume under teacher supervision. Concentrated acids require the school's specific safe procedure; the generic arithmetic is not an instruction to mix hazardous liquids.

■ Checked worked case

Known: 5.0 g solute in 0.250 dm^3 solution gives 20 grams per dm^3. Diluting to final volume 0.500 dm^3 gives 10 grams per dm^3. Alternatively, keeping final volume 0.250 dm^3 and increasing dissolved solute to 10.0 g gives 40 grams per dm^3. The two operations change different parts of the ratio.

■ Common error

Do not say concentration always rises when volume rises: the held quantity matters. The equation assumes a uniform solution and that the stated mass is dissolved. Evaporation may also remove volatile solute, so constant solute mass must be a stated condition rather than automatically inferred.

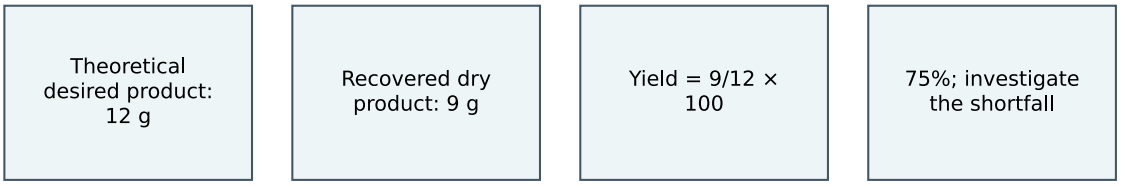
Chemistry-only: percentage yield compares recovered and possible product

Official-unit focus: 4.3 Quantitative chemistry

A calculation predicts 10 g product, but only 8 g dry product is recovered. The missing recovered mass does not establish that atoms were destroyed.

Actual yield is the amount of desired product obtained. Theoretical yield is the maximum predicted by the reaction calculation under its stated conditions. Percentage yield equals actual product mass divided by theoretical product mass, multiplied by 100. When the theoretical mass is supplied, this calculation is both-tier Chemistry-only content; deriving that theoretical mass from reactants and an equation is separately Higher-only.

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Recovery may be below the maximum because a reversible reaction does not go to completion, product is lost during separation, or reactants undergo side reactions. Some atoms can remain in reactants or enter unwanted products; some desired product can remain dissolved or on apparatus. Conservation concerns all atoms and substances, while yield concerns the recovered desired product.

Record a dry product mass rather than including retained solvent or unrelated material. Write actual and theoretical values in the same unit before dividing. Inspect a proposed explanation against the method: transfer loss, incomplete crystallisation or reversible conversion are different causes and need different evidence. School-supervised separations remain actual tasks; this written evaluation supports them rather than replacing them.

■ Checked worked case

Known: a supplied theoretical product mass is 12.0 g and dry recovered mass is 9.0 g. Percentage yield=9.0/12.0×100=75%. The shortfall is 3.0 g desired product relative to the maximum, not necessarily 3.0 g spilled solid. The cause cannot be assigned uniquely from these two mass values alone.

■ Common error

Divide actual by theoretical, not the reverse. A yield above 100% prompts a check for wet/impure product, incorrect mass data or assumptions; it does not prove extra atoms were created. A catalyst can speed reaching a result without changing the stoichiometric maximum from fixed limiting reactants.

Higher Tier, Chemistry-only: derive the theoretical yield first

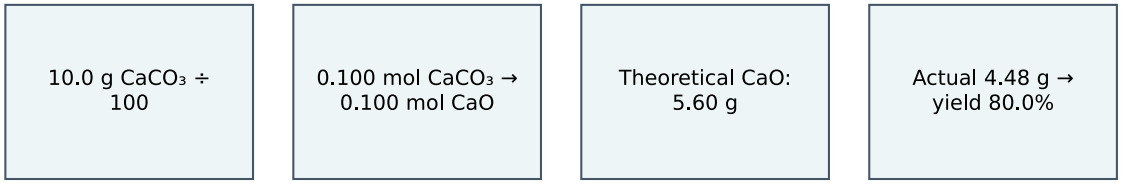
Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

If a question gives only a reactant mass and recovered product, there is one extra step before finding yield: determine the theoretical product mass from the equation.

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Convert the stated limiting-reactant mass into moles, use the balanced coefficient ratio and multiply by product molar mass. The result is theoretical yield under complete conversion with sufficient other reactants. Then use actual/theoretical×100 for percentage yield. This mass derivation is Higher-only, while dividing two supplied product masses is common-tier Chemistry-only work.



For $CaCO_3 \rightarrow CaO + CO_2$, the amount ratio $CaCO_3$:CaO is 1:1. Supplied molar masses are 100 and 56 g per mol. A 10.0 g $CaCO_3$ sample contains 0.100 mol and can give 0.100 mol CaO, mass 5.60 g. The other 4.40 g is theoretical carbon dioxide, so the product-residue mass is lower than the original carbonate mass even at complete conversion.

Label theoretical and actual masses clearly. State sample purity and which reactant limits if relevant. Use unrounded intermediate values when multiplying a ratio. Check the total product masses against reactant mass for a decomposition with no added reactant. For an actual school experiment, follow approved heating and handling procedures; calculation alone does not establish purity or completion.

■ Checked worked case

Known: 10.0 g pure $CaCO_3$ theoretically gives 5.60 g CaO. If 4.48 g dry CaO is recovered, yield=4.48/5.60×100=80.0%. Dividing 4.48 by 10.0 instead would compare different substances and incorrectly give 44.8%. Recovery must be compared with the possible mass of the same desired product.

■ Common error

A smaller solid residue is not itself low yield when gas is an expected product. Impure starting material changes the available reactant mass. Do not use the actual product mass to calculate the theoretical amount and then claim an independent yield result.

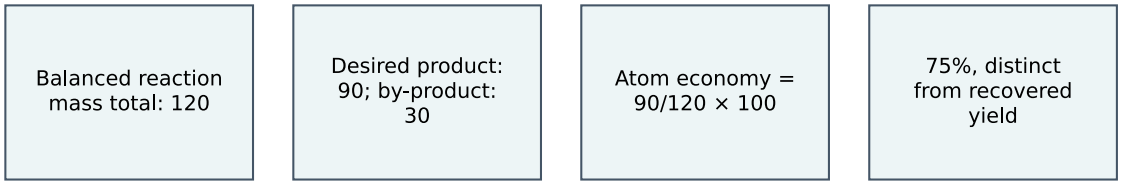
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Chemistry-only: atom economy follows the balanced reaction

Official-unit focus: 4.3 Quantitative chemistry

A reaction can recover all the product it theoretically makes while directing many starting atoms into an unwanted by-product. Yield and atom economy measure different things.

Atom economy measures the share of starting-material mass that appears in the desired product according to the balanced equation. Percentage atom economy equals the relative formula mass contribution of the desired product divided by the sum of all reactant relative mass contributions, multiplied by 100. Include balancing coefficients in both contributions where needed.



For $CaCO_3 \rightarrow CaO + CO_2$ with relative masses 100, 56 and 44, choosing CaO as desired product gives $56/100\times100=56\%$. Choosing CO_2 as the desired product for a different purpose gives 44%. Conservation still accounts for all 100 mass units. A high atom economy can reduce waste and raw-material cost, but it does not alone establish safety, energy demand or commercial suitability.

Balance the equation, identify the desired product explicitly, and write every coefficient-weighted relative mass. For $2H_2 + O_2 \rightarrow 2H_2O$, desired water contribution is $2\times18=36$ and reactant total is $2\times2+32=36$, so atom economy is 100%. This uses relative mass bookkeeping rather than an experimental product weighing. Both-tier calculations are distinct from Higher-only comparison of complete reaction pathways.

■ Checked worked case

Known: a supplied balanced reaction uses total reactant relative-mass contribution 120 and produces desired-product contribution 90 plus by-product contribution 30. Atom economy=$90/120\times100=75\%$. An experiment could still recover only half its theoretical desired product, giving 50% yield while the balanced reaction retains 75% atom economy.

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■ Common error

Do not use recovered product mass in the atom-economy formula. Do not omit a reactant just because it is a gas. High yield cannot remove a stoichiometric by-product. If a by-product has a use, that may improve the pathway's practical value, but keep the stated desired-product definition explicit.

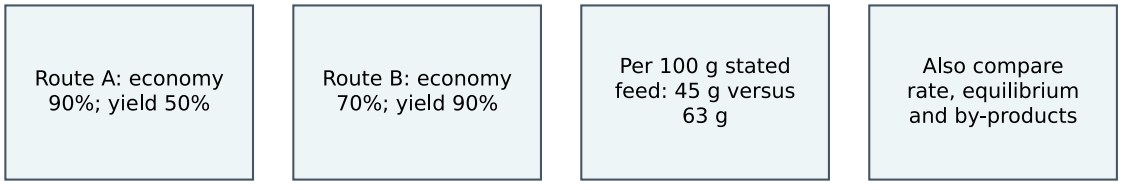
Higher Tier, Chemistry-only: choose a reaction pathway from several criteria

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

A high atom-economy pathway may be slow or give a low recovered yield. A justified industrial choice compares the relevant evidence and states the decision priority.

Use supplied data about atom economy, yield, reaction rate, equilibrium position and usefulness of by-products. Atom economy describes stoichiometric allocation of starting material; yield describes desired product obtained against its theoretical maximum. Rate affects output over time, while equilibrium may limit conversion. Useful saleable by-products may reduce waste or improve economics, but their market and processing requirements need evidence.



Imagine route A with atom economy 90%, yield 50% and a slow rate, and route B with atom economy 70%, yield 90% and a faster rate under supplied comparable conditions. Neither percentage alone decides every use. If equal starting mass could theoretically enter product according to these fractions, recovered desired mass per 100 g starting material is 45 g for A and 63 g for B under this simplified comparison.

State what the data hold constant: feedstock mass, purity, conditions, reaction time or cost. Rank criteria against the question's objective and acknowledge one disadvantage of the chosen route. Do not equate fast reaction with low energy cost unless the conditions support that inference. A favourable equilibrium position is different from a catalyst speeding the approach to it.

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■ Checked worked case

Known: a supplied route directs 80% of 200 g reactant mass into theoretical desired product, giving 160 g. At 75% recovery, actual desired mass is 120 g. This product-of-fractions illustration assumes the atom-economy fraction applies to the stated total feed and no reactant excess is included; other industrial datasets may require a different basis.

■ Common error

Higher atom economy alone does not prove an entire process is sustainable. Toxicity, energy, separation and resource availability can matter when supplied. Keep yield and atom economy denominators separate, and avoid inventing cost or environmental facts absent from the evidence.

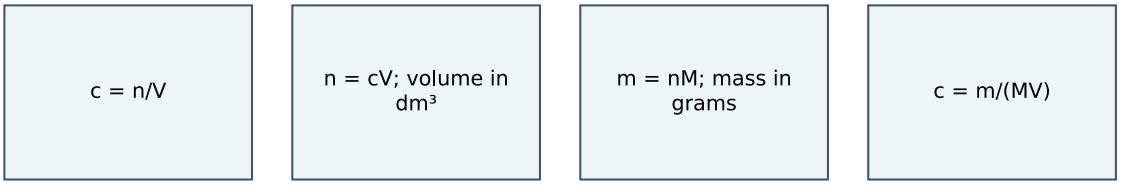
Higher Tier, Chemistry-only: concentration in moles per solution volume

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

A concentration of 0.20 moles per cubic decimetre tells the amount of dissolved solute, not its gram mass. The molar mass is needed to connect amount and mass.

Molar concentration c=n/V uses solute amount n in mol and final solution volume V in dm^3. Rearrange to n=cV or V=n/c. To find solute mass use m=nM, where M is molar mass in grams per mol. Combining gives c=m/(MV). The same gram mass of two different solutes can have different mole amounts because their molar masses differ.



Convert cm^3 to dm^3 by dividing by 1,000. A 250 cm^3 portion of 0.20 moles per dm^3 solution contains $0.20 \times 0.250 = 0.050$ mol solute. If solute molar mass is 40 g per mol, that amount weighs 2.0 g. At fixed amount, greater final solution volume means lower molar concentration. At fixed final volume, more solute moles means higher

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concentration.

Write units beside concentration and volume before multiplying. Identify which chemical species concentration describes. A bottle's molar concentration normally refers to the stated solute formula amounts, not automatically the total amount of every ion after dissociation. Actual solution preparation requires a measured final volume, complete dissolution and the teacher-approved apparatus and procedure.

■ Checked worked case

Known: 4.0 g NaOH with supplied molar mass 40 g per mol is 0.100 mol. In final volume 0.500 dm^3 its concentration is 0.100/0.500=0.200 moles per dm^3. A 100 cm^3 aliquot contains 0.0200 mol and 0.800 g NaOH. This is solute content rather than total solution mass.

■ Common error

Do not substitute grams directly into c=n/V. Do not use the volume of solvent added when a final solution volume is specified. Molar concentration and mass concentration are related through molar mass, but their numerical values and units are different.

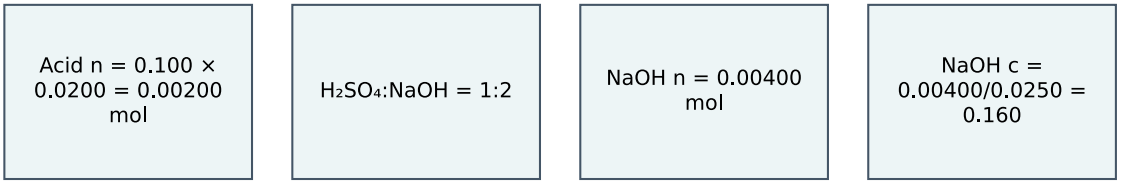
Higher Tier, Chemistry-only: titration concentrations follow the equation ratio

Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

Equal volumes of acid and alkali do not necessarily react completely with each other. The concentrations and balanced coefficients both determine the required quantities.

For reacting solutions, calculate known amount with n=cV, use the balanced coefficient ratio and divide the unknown amount by its volume. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$ means one mole acid reacts with two moles alkali. A 1:1 shortcut is valid for HCl and NaOH, but not for this sulfuric-acid equation.



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A pipette supplies the measured sample volume and a burette delivers the titre. The titre is final minus initial burette reading, not the final reading alone. Use consistent close results under the school's stated concordance rule and exclude a rough trial from the accurate mean. A suitable indicator end-point approximates the required neutralisation condition; it is not evidence that every solution must be pH 7.

Perform titration only as an actual supervised school task using approved dilute solutions, eye protection and the specified pipette filler. Rinse and use the apparatus as instructed, remove the filling funnel before readings and add dropwise near the end-point. Record readings at appropriate precision. These calculation exercises do not count as completing Required Practical 2, whose real method belongs in chemical changes.

■ Checked worked case

Known: 25.0 cm^3 NaOH requires 20.0 cm^3 of 0.100 mol per dm^3 H_2SO_4. Acid amount=0.100×0.0200=0.00200 mol. Alkali amount=2×0.00200=0.00400 mol. NaOH concentration=0.00400/0.0250=0.160 mol per dm^3. The ratio factor comes from the equation, not from the two measured volumes.

■ Common error

Use the correct unknown volume in the final division. Mean titres must be selected from appropriate trials; averaging a rough result can spoil an accurate calculation. Keep practical skill evidence separate from written calculations and do not invent measurements as if they were collected by students.

Higher Tier, Chemistry-only: gas volumes at room temperature and pressure

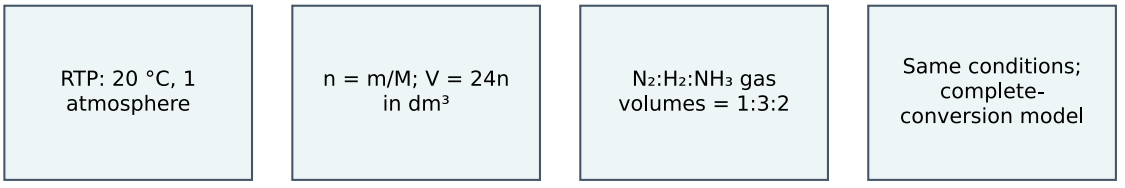
Higher Tier focus; see section scope notes.

Official-unit focus: 4.3 Quantitative chemistry

One mole of hydrogen and one mole of carbon dioxide occupy the same gas volume at the same temperature and pressure. Their different masses do not change that equal-amount volume rule.

At the specification's room temperature and pressure, 20 °C and 1 atmosphere, one mole of gas occupies 24 dm^3. Calculate V=24n in dm^3, or n=V/24. For a gas mass, first calculate n=m/M, then volume. Conversely, a gas volume gives amount and then mass using its molar mass. This fixed molar-volume value is tied to the stated conditions.

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Equal gas amounts occupy equal volumes at identical conditions, so balanced coefficients give gaseous volume ratios. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ gives ratio 1:3:2 for gas volumes measured at the same temperature and pressure. This ratio does not apply to a liquid or solid volume. A predicted product volume assumes sufficient reactants and the stated complete-conversion model; real equilibrium can reduce conversion.

Convert cm^3 to dm^3 consistently and label every volume's conditions. Do not use 22.4 dm^3 per mol, a value associated with different reference conditions, in this RTP task. Use teacher-provided gas data or approved low-risk gas collection under supervision; do not propose making ammonia or hydrogen in an improvised apparatus. Water-vapour and collection losses can affect real measurements.

■ Checked worked case

Known: 4.4 g CO_2 with molar mass 44 g per mol gives 0.100 mol, hence 2.40 dm^3 at RTP. In the ammonia equation, 6.0 dm^3 H_2 can theoretically react with 2.0 dm^3 N_2 to form 4.0 dm^3 NH_3 at identical conditions. There are four reactant gas volumes for two product volumes; total gas volume need not be conserved even though atoms and mass are.

■ Common error

The number 24 is not a universal gas volume at every temperature and pressure. Coefficients concern gas volume only under matched conditions. Equal gas volumes can have different masses, and a reduction in total gas volume does not mean atoms disappeared.

Redox and electrolysis

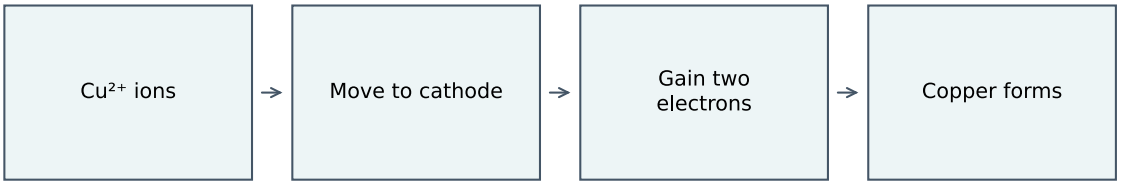
Official-unit focus: 4.4 Chemical changes

An aqueous salt solution can produce different electrode products from the molten salt. Water introduces competing species into the system.

Oxidation is loss of electrons and reduction is gain of electrons. In electrolysis, cations move toward the cathode and anions toward the anode. Reduction occurs at the

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cathode.



Predict products using the specified electrolyte and electrode material. In an aqueous solution, hydrogen or oxygen may form because water-related species compete. Molten salts contain only the ions of the salt.

Use a low-voltage direct-current supply, approved electrodes, and the school risk assessment. Collect gases only by an approved method. Keep chlorine demonstrations teacher-controlled; do not ask students to generate hazardous gases independently.

■ Checked worked case

Known: a copper ion gains two electrons. Half-equation: $Cu^{2+} + 2e^- \rightarrow Cu$. One mole of Cu^{2+} requires two moles of electrons. For 0.050 mol of copper, electron amount $= 2 \times 0.050 = 0.100$ mol.

■ Common error

Electrode signs depend on the cell type. In an electrolytic cell the cathode is negative; reduction remains the defining process at a cathode in every cell.

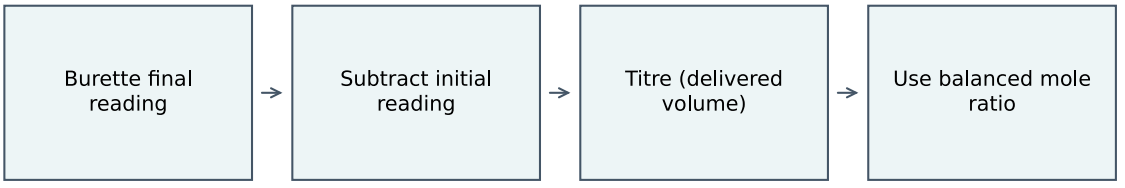
Titration and a defensible concentration

Official-unit focus: 4.4 Chemical changes

A burette reading is not the delivered volume. The titre is the difference between final and initial readings, and both readings have uncertainty.

A titration measures the amount of one solution needed to react with a known amount of another. The equation gives the mole ratio. An indicator endpoint approximates the equivalence point when a suitable indicator is used.

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Calculate the known amount first, apply the stoichiometric ratio, then divide by the unknown solution volume in cubic decimetres. Use concordant titres as required by the school method and report the accepted values.

Rinse the burette with its solution and the pipette with the solution it transfers. Rinse the flask with distilled water. Add titrant slowly near the endpoint, swirl, and read the meniscus at eye level. Use a white tile and appropriate eye protection.

■ Checked worked case

Known: 25.0 cubic centimetres of acid reacts 1:1 with 20.0 cubic centimetres of 0.100 mol per cubic decimetre alkali. $n = cV = 0.100 \times 0.0200 = 0.00200$ mol. Acid amount is 0.00200 mol. $c = n/V = 0.00200/0.0250 = 0.0800$ mol per cubic decimetre.

■ Common error

Adding distilled water to the flask changes volume but not the transferred amount of analyte. Do not average a rough titre with carefully measured concordant values.

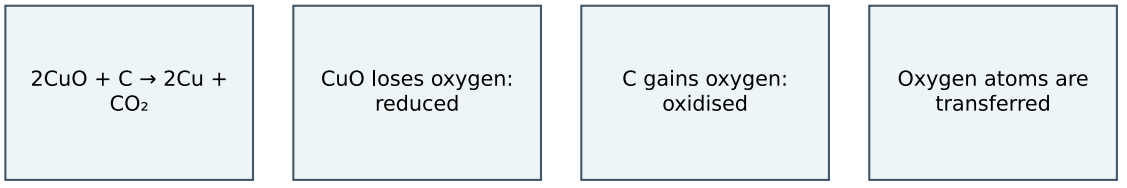
Metal oxides: oxidation gains oxygen, reduction loses it

Official-unit focus: 4.4 Chemical changes

A strip of magnesium can form a white solid after reacting with oxygen. The new substance contains oxygen as well as the starting metal.

Metals react with oxygen to form metal oxides. In the oxygen definition, oxidation is gain of oxygen and reduction is loss of oxygen. Magnesium is oxidised in $2Mg + O_2 \rightarrow 2MgO$. A metal oxide can be reduced if oxygen is removed from it. These definitions identify changes in substances, not merely a change in colour or measured temperature.

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For $2CuO + C \rightarrow 2Cu + CO_2$, copper oxide loses oxygen and is reduced, while carbon gains oxygen and is oxidised. The same oxygen atoms leave the oxide and appear in carbon dioxide. The two changes occur together in this reaction. Oxygen is an element; oxide is a compound containing oxygen with another element, so the names are not interchangeable.

Use atom models or supplied before/after formulae to trace oxygen. Balance each element and identify the substance receiving or losing oxygen. Actual burning-metal or heated-oxide work needs the teacher-approved method, heat protection and ventilation. A bright magnesium flame must not be viewed directly; a written example is not an instruction to ignite a sample independently.

■ Checked worked case

Known: a stated metal sample begins at 3.6 g and forms 6.0 g oxide with complete reaction and no losses. Oxygen incorporated=6.0−3.6=2.4 g. The metal's gain of oxygen identifies oxidation. The oxide mass includes material from the surrounding oxygen, so this is compatible with mass conservation.

■ Common error

Reduction does not mean every mass decreases, and oxidation does not mean every colour change. The common-tier oxygen account is distinct from the wider electron definition taught in the separate HT lesson. Some redox changes involve no oxygen and therefore need that other definition.

The reactivity series: compare reactions at room temperature

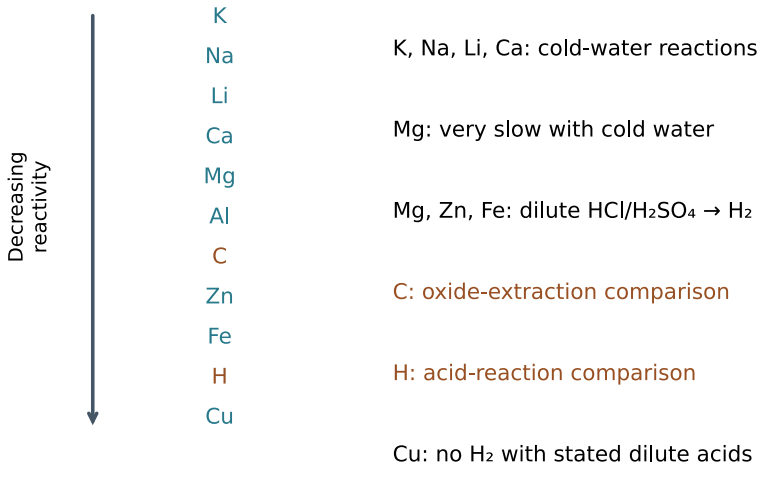
Official-unit focus: 4.4 Chemical changes

Copper does not release hydrogen from dilute hydrochloric acid, while magnesium reacts readily. The contrast helps locate them relative to hydrogen in the reactivity series.

For the specified metals, descending reactivity is potassium, sodium, lithium, calcium,

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magnesium, zinc, iron, copper. Include carbon above zinc and hydrogen below iron when using extraction or acid rules; aluminium is commonly shown above carbon in the fuller series. More reactive metals more readily form positive ions. In this section comparisons are at room temperature, not reactions with steam.



Potassium, sodium and lithium react vigorously with cold water to produce hydroxides and hydrogen, with potassium generally most vigorous. Calcium also reacts, producing hydrogen and calcium hydroxide. Magnesium reacts very slowly with cold water, while zinc, iron and copper show no appreciable reaction under this school comparison. With dilute hydrochloric or sulfuric acid, Mg, Zn and Fe produce hydrogen, with Mg more vigorous and Fe slower; Cu does not. Very reactive alkali metals are not student acid-test samples.

Compare suitable teacher-approved metal samples under controlled acid concentration, volume, temperature and exposed surface conditions. Observe bubbles and metal consumption, but do not treat raw bubble counts from different-sized samples as a pure reactivity measure. Use supplied evidence or a teacher demonstration for highly reactive metals. Steam reactions and detailed unusual conditions are outside this stated recall scope.

■ Checked worked case

Known: supplied observations show metal A reacts with cold water; B does not appreciably react with water but releases hydrogen in dilute acid; C does neither. Under comparable conditions, A>B>C is supported. A separate acid test collects 36 cm^3 gas in 60 s, giving mean collection rate 0.60 cm^3 per second; this is method-specific evidence, not a universal rate constant.

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■ Common error

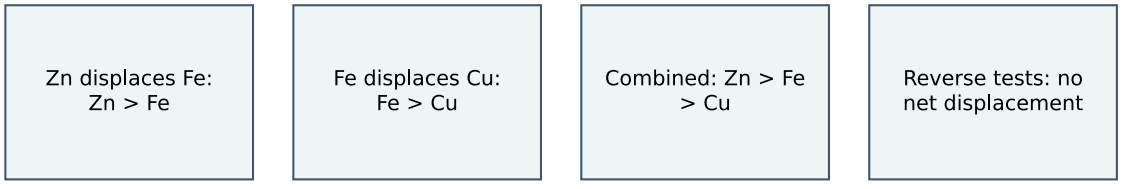
No visible reaction does not mean a substance can never react under any condition. A protective layer or unequal surface area can affect observations. Hydrogen and carbon are comparison positions, not metals. Do not mix the source's room-temperature scope with magnesium-steam reactions.

Displacement evidence: build a consistent metal order

Official-unit focus: 4.4 Chemical changes

Iron in copper(II) sulfate solution can acquire a copper coating. The more reactive metal takes the place of the less reactive metal in the compound.

A more reactive metal displaces a less reactive metal from a compound. Fe + $CuSO_4$ → $FeSO_4$ + Cu shows iron replacing copper, so the evidence supports Fe>Cu. Copper placed in iron(II) sulfate does not displace iron under the same ordinary comparison. Displacement is a chemical change producing different substances, not simply dissolving a solid.



For supplied results, translate each successful displacement into a comparison arrow: A displaces B implies A>B. Combine comparisons into an order and check for contradictions. If A displaces B and B displaces C, A>B>C is consistent. Lack of visible reaction alone needs care because an unsuitable method, protective layer or short observation time can hide a reaction.

Use approved dilute salt solutions, clean matched metal samples, fixed volumes and comparable observation times. Record colour, deposit and sample changes without assigning identity from colour alone. A control with no added metal helps distinguish background change. Dispose of metal-containing solutions through the school's specified collection route; do not test highly reactive metals in water-based salt solutions independently.

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■ Checked worked case

Known: supplied results show zinc displaces iron and copper, and iron displaces copper. The consistent order is Zn>Fe>Cu. If 8.0 g of a supplied recoverable-metal model gives 6.4 g recovered deposit, recovery is 6.4/8.0×100=80%. That recovery calculation does not determine the reactivity order; the displacement comparisons do.

■ Common error

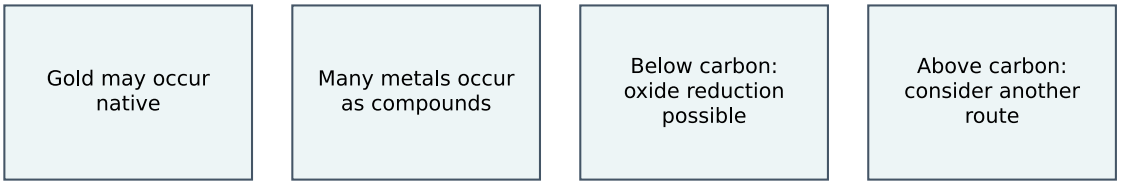
A metal cannot normally displace itself to produce a net change. Do not call the solution deposit the newly formed more reactive metal. Half and ionic equations belong to the separate HT treatment; the common-tier displacement conclusion can be explained using compounds and the series.

Extracting metals: carbon can reduce suitable metal oxides

Official-unit focus: 4.4 Chemical changes

Gold can occur as the metal itself, but many useful metals occur in compounds. Extraction must separate the metal chemically from the other elements.

Very unreactive metals such as gold can be found native, as the metal. Most metals are found as compounds. Metals less reactive than carbon can be extracted from their oxides by reduction using carbon. The oxide loses oxygen and produces metal; carbon gains oxygen. Metals above carbon require another route in this GCSE account, including electrolysis where suitable.



For $2CuO + C \rightarrow 2Cu + CO_2$, copper oxide is reduced and carbon oxidised in oxygen terms. Carbon cannot be assumed to reduce every metal oxide under useful extraction conditions. Aluminium is more reactive than carbon and its extraction uses electrolysis, not simply this copper-oxide route. Detailed blast-furnace process recall is not required by this section's stated limits.

When given an extraction description, locate the feed compound, reducing substance,

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desired metal and by-products. Evaluate supplied energy, cost, purity or waste data against a stated purpose rather than inventing industrial details. Any actual heated-oxide school work uses the approved small-scale method and technician-selected substances; industrial extraction is a reference case, not a school construction task.

■ Checked worked case

Known: a supplied ore model contains 4% recoverable metal by mass. A 250 kg batch therefore contains $0.04 \times 250 = 10$ kg recoverable metal before processing losses. If 8 kg is obtained, recovery is 80%. This ore-content calculation supports a supplied process comparison but does not imply that every carbon reduction has the same yield or ore composition.

■ Common error

Native means occurring as the elemental metal, not merely being dug from a mine. Oxygen loss identifies oxide reduction; lower ore mass alone does not. A process may be chemically possible while poor economically or environmentally under supplied conditions.

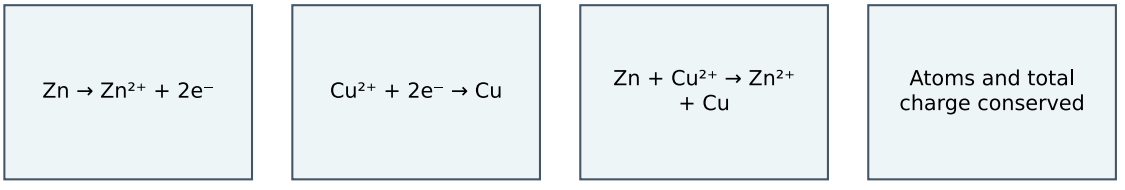
Higher Tier: displacement as electron loss and gain

Higher Tier focus; see section scope notes.

Official-unit focus: 4.4 Chemical changes

Zinc reacting with copper(II) ions transfers electrons even though the net ionic equation contains no oxygen. The electron definition extends redox beyond oxygen reactions.

Oxidation is loss of electrons; reduction is gain of electrons. $Zn \rightarrow Zn^{2+} + 2e^-$ is oxidation and $Cu^{2+} + 2e^- \rightarrow Cu$ is reduction. Combining gives $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$. Sulfate ions in a zinc–copper sulfate experiment remain spectator ions and do not appear in this net ionic equation.



Check both atom counts and total charge. The combined left charge is +2 and the

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right charge is +2. Zinc atoms become positive ions by losing electrons; copper ions become neutral metal atoms by accepting them. Identify the actual species reduced, Cu^{2+}, rather than saying the already-neutral copper product gains further electrons.

Write the two half equations, multiply them if needed to make equal electron numbers, then cancel electrons. Cancel only unchanged spectator ions in a full ionic equation. Use the reactivity series to predict whether a proposed displacement is supported; a charge-balanced equation alone does not guarantee that the reaction will occur.

■ Checked worked case

Known: $2Al \rightarrow 2Al^{3+} + 6e^-$ and $3Cu^{2+} + 6e^- \rightarrow 3Cu$ combine as $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$. Two aluminium atoms supply six electrons and three copper ions accept them. Both sides have charge +6. These particle coefficients are not gram masses.

■ Common error

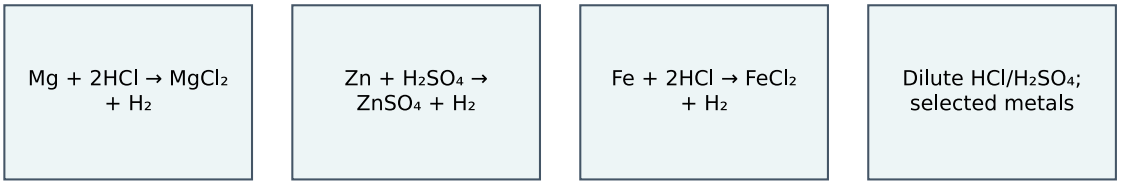
Oxidation and reduction occur together in the overall electron transfer. Positive-ion formation by a metal is electron loss, not loss of protons. Electrons cancel from the full reaction but remain useful in each half equation. This entire source section is Higher-only.

Acids with metals: choose the salt and hydrogen products

Official-unit focus: 4.4 Chemical changes

Magnesium in dilute hydrochloric acid produces bubbles and a dissolved salt. Hydrogen is the gas, while chloride from the acid contributes to the salt identity.

Some metals react with acids to make a salt and hydrogen. In this source recall scope use magnesium, zinc and iron with dilute hydrochloric or sulfuric acid. Hydrochloric acid gives chlorides and sulfuric acid gives sulfates. $Mg + 2HCl \rightarrow MgCl_2 + H_2$ and $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ are balanced examples. Iron with dilute acid forms iron(II) salts in the stated school account.



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The metal must be suitable relative to hydrogen. Copper does not normally release hydrogen from these dilute acids. Do not extend this acid–metal pattern to nitric acid or concentrated acid reactions, which are outside the stated scope and can produce different outcomes. With a metal oxide, the usual products are salt and water instead of hydrogen.

Use teacher-approved dilute acid and selected small metal samples with eye protection and no nearby ignition source. Record bubbles and metal disappearance. A permitted small collected hydrogen sample gives a squeaky pop with the teacher-approved lit-splint test; do not apply a flame to the generating vessel or a large gas accumulation. Temperature, surface area and acid concentration affect the observed rate.

■ Checked worked case

Known: $Fe + 2HCl \rightarrow FeCl_2 + H_2$ balances one Fe, two Cl and two H atoms on each side. Four represented iron atoms require eight HCl molecules and form four H_2 molecules. This counting exercise uses equation particle ratios without requiring the Higher-only electron explanation of this reaction.

■ Common error

Not every acid–solid reaction releases hydrogen. Carbonates release carbon dioxide, while many bases produce water. Salt formulae need correct ion charges. Common-tier product prediction remains separate from the embedded Higher-only electron redox interpretation.

Higher Tier: hydrogen ions are reduced in acid–metal reactions

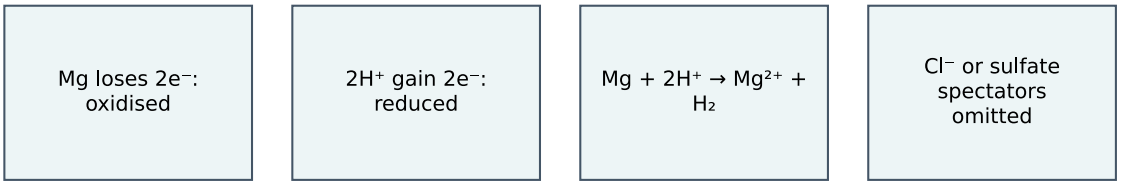
Higher Tier focus; see section scope notes.

Official-unit focus: 4.4 Chemical changes

The hydrogen in bubbles from magnesium and acid has come from aqueous hydrogen ions. Those ions gain electrons supplied by the magnesium.

For $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$, magnesium loses two electrons and is oxidised. The hydrogen ions gain electrons and are reduced. The half equations are $Mg \rightarrow Mg^{2+} + 2e^-$ and $2H^+ + 2e^- \rightarrow H_2$. Combining cancels the transferred electrons. Chloride or sulfate ions remain in solution as appropriate spectators.

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Charge is +2 on both sides of the net equation. Two H^+ ions are needed to form one H_2 molecule; gaining one electron each makes their total gain two electrons. The metal product is an ion, while the hydrogen product is a neutral diatomic molecule. Electron gain must be assigned to hydrogen ions, not to chloride ions just because hydrochloric acid was used.

Separate the reactant species before identifying redox. In a supplied equation locate any metal atom becoming a positive ion and hydrogen ions becoming gas. Balance each half equation by atoms and charge. Do not generalise the simple net equation to every acid or every metal; the source specifies Mg, Zn and Fe with dilute hydrochloric/sulfuric acids.

■ Checked worked case

Known: $Fe \rightarrow Fe^{2+} + 2e^-$ combines with $2H^+ + 2e^- \rightarrow H_2$. Five represented Fe atoms release ten electrons, which reduce ten H^+ ions to five H_2 molecules. The electron count explains the two-to-one H^+:H_2 ratio without changing the number of iron nuclei.

■ Common error

Hydrogen gas formation is reduction in this account, even though no oxygen has been removed. Oxidation does not mean gaining positive charge by adding protons. This embedded HT explanation does not make common-tier salt names and product equations HT-only.

Neutralisation: select products and balance salt charges

Official-unit focus: 4.4 Chemical changes

Copper oxide with sulfuric acid gives copper sulfate and water. Changing the acid to nitric acid changes the salt's negative ion, even though copper stays the positive ion.

Acid plus a metal oxide or hydroxide gives salt and water. Acid plus a metal carbonate gives salt, water and carbon dioxide. An alkali is a soluble base; some bases are

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insoluble. Hydrochloric acid supplies chloride salts, nitric acid nitrate salts and sulfuric acid sulfate salts. The base, alkali or carbonate supplies the positive ion.



Use charges to make a neutral formula: Cu^{2+} and $SO_4{}^{2-}$ give $CuSO_4$; Mg^{2+} and $NO_3{}^-$ give $Mg(NO_3)_2$. $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$ is balanced. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ has two chloride ions per calcium ion and one carbon dioxide molecule per carbonate formula unit.

Identify the acid first, then the positive ion and required ratio. Use brackets when more than one polyatomic ion is needed, keeping the nitrate or sulfate group intact. Confirm atom counts in the full equation. Observe carbon dioxide with the approved limewater test where appropriate, rather than treating all bubbles as hydrogen.

■ Checked worked case

Known: aluminium ions Al^{3+} combine with sulfate $SO_4{}^{2-}$. Two aluminium ions supply +6 and three sulfates supply −6, giving $Al_2(SO_4)_3$. The formula contains two Al, three S and twelve O atoms. This formula exercise uses supplied ions and does not require conducting a reaction with aluminium oxide in class.

■ Common error

Water is a product of oxide/hydroxide neutralisation, while carbonate neutralisation also produces carbon dioxide. A base need not dissolve in water to react with acid. A salt's name depends on both ions, and the polyatomic-ion subscript inside its own formula must not be altered to balance charge.

Required Practical 1: prepare pure, dry soluble salt crystals

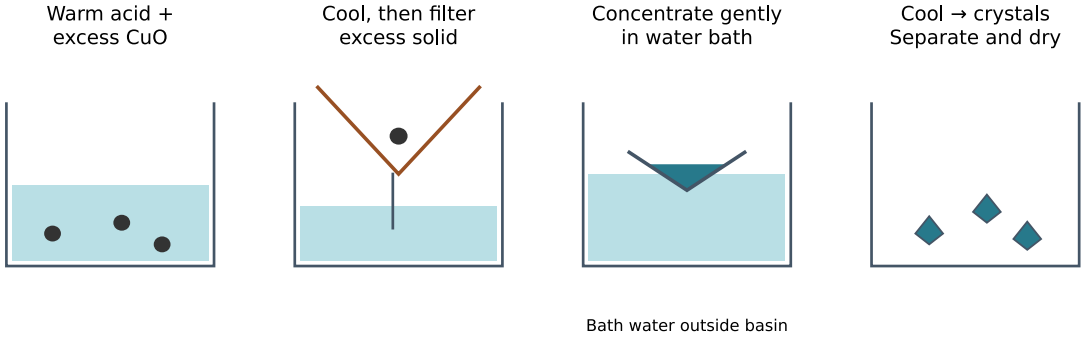
Official-unit focus: 4.4 Chemical changes

Copper oxide is an insoluble black solid; copper sulfate solution is blue. Adding excess oxide lets the remaining acid react, then filtration removes the unused solid.

Required Practical 1 prepares a pure, dry soluble salt from an insoluble oxide or carbonate. For copper sulfate use dilute sulfuric acid and copper(II) oxide: CuO +

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$H_2SO_4 \rightarrow CuSO_4 + H_2O$. Warm the dilute acid gently under the approved method, then add oxide in small portions with stirring until a little remains unreacted. The excess indicates that acid is used up under the method's assumptions.



$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$; not to scale; approved school apparatus required

Allow cooling and filter to remove excess insoluble oxide. The filtrate contains dissolved copper sulfate, so it passes through filter paper. Gently concentrate it in an evaporating basin using a water bath or electric heater, then allow cooling to form crystals. Separate the crystals and dry gently as directed. Evaporating completely to dryness is not the crystal-preparation endpoint.

Carry out the school-supervised task with the specified dilute reagents, eye protection and risk-assessed heat equipment. Record the actual sequence, observations and recovered dry sample. The handbook's copper sulfate example uses warmed acid, then a water bath for evaporation. Avoid overheating the solution; hot apparatus can look cold. The teacher/technician determines appropriate crystal handling and disposal.

■ Checked worked case

Known: a supplied preparation records empty collection vessel 18.40 g and vessel plus dry crystals 21.65 g. Crystal mass=21.65−18.40=3.25 g. A wet sample would include retained solution and give an inflated recovered mass. This measured mass is not a theoretical yield unless starting amounts and product form are also specified.

■ Common error

Filtration removes excess solid, not the dissolved salt. Heating concentrates the solution; cooling promotes crystals rather than a new neutralisation reaction. A written sequence and fabricated observations do not establish that RP1 was completed: retain the student's actual supervised work and records.

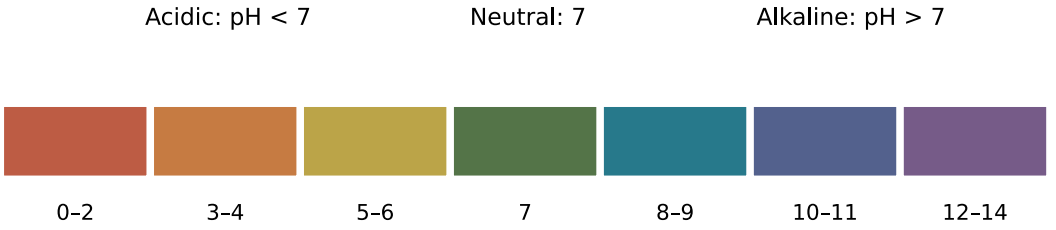
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The pH scale and hydrogen–hydroxide neutralisation

Official-unit focus: 4.4 Chemical changes

Universal indicator gives a range of colours as an acid is added to an alkali. A pH probe can record a numerical change, but it still needs an appropriate measurement method.

In this GCSE aqueous account, acids produce hydrogen ions H^+ and alkalis contain hydroxide ions OH^-. On the usual 0–14 school scale, pH below 7 is acidic, pH 7 neutral and pH above 7 alkaline. Universal or wide-range indicator gives approximate pH by matching its colour with the supplied chart; a suitable calibrated probe provides a numerical reading.



$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Indicative colours only: use the chart for the actual universal indicator

Neutralisation between acid and alkali forms water: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. The positive and negative charges cancel, and atoms are conserved. Excess acid after mixing leaves the solution acidic; excess alkali leaves it alkaline. Neutralisation does not automatically mean that equal volumes of every pair of solutions will produce pH 7.

Use clean samples and small consistent indicator quantities or a correctly calibrated, rinsed pH probe. Match colours under suitable lighting and read the chart for the indicator actually used. During a supervised strong-acid/strong-alkali comparison record the measured pH against added volume. Indicator colour estimates and probe readings have different uncertainties; do not invent precision from a broad colour band.

■ Checked worked case

Known: supplied solutions P, Q and R have pH 3, 7 and 11. P is acidic, Q neutral and R alkaline under this school model. If ten represented H^+ ions react with seven represented OH^- ions, seven water molecules form and three H^+ ions remain. This particle illustration explains excess acid without requiring a Higher-only logarithmic concentration calculation.

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■ Common error

A neutral solution contains ions; neutral does not mean ion-free. Alkali is not the same word as any insoluble base. A pH difference is not a simple gram-mass difference, and the tenfold hydrogen-concentration interpretation is in the separate HT strong/weak-acid lesson.

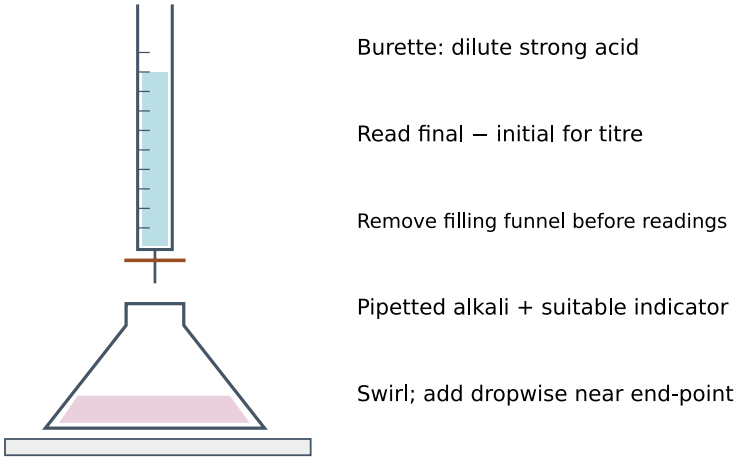
Chemistry-only Required Practical 2: measure reacting volumes

Official-unit focus: 4.4 Chemical changes

A burette lets you add acid in small measured steps to a known alkali volume. A suitable indicator shows when to stop; it does not calculate the concentration by itself.

Required Practical 2 determines reacting acid and alkali volumes. A volumetric pipette and filler transfer a known alkali volume into a conical flask. A burette holds the chosen dilute strong acid: hydrochloric, nitric or sulfuric acid in the source scope. Use a suitable indicator, a white tile and swirling. Add dropwise near the end-point and record the delivered volume.

Clamped vertical burette; white tile; apparatus schematic



Titre equals final burette reading minus initial reading. Read the appropriate meniscus at eye level, with the burette vertical and the filling funnel removed. A rough trial locates the endpoint; subsequent careful trials supply an accurate mean under the school's stated concordance rule. Phenolphthalein is pink in alkali and becomes permanently colourless at the handbook's acid-into-alkali endpoint. Universal indicator is unsuitable for a sharply defined accurate endpoint.

Perform the actual school-supervised task with approved dilute solutions and eye protection. Rinse apparatus as instructed to avoid unwanted dilution, fill the burette tip and use the pipette's specified draining method. Never pipette by mouth. Retain

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initial/final readings, trials, selected mean and exclusion reason. Foundation students measure reacting volumes; concentration determination is the separate Higher extension.

■ Checked worked case

Known: an accurate trial begins at 1.20 cm^3 and ends at 19.55 cm^3, so titre=18.35 cm^3. A second accepted titre is 18.25 cm^3; mean=(18.35+18.25)/2=18.30 cm^3. The school supplies its acceptable concordance tolerance; this example does not impose a qualification-wide 0.10 cm^3 rule.

■ Common error

Do not average the rough trial automatically or read only the final burette value. Adding water to the flask does not change the amount of pipetted solute if nothing is lost. Written methods are preparation/review; they do not replace the student's actual titration performance.

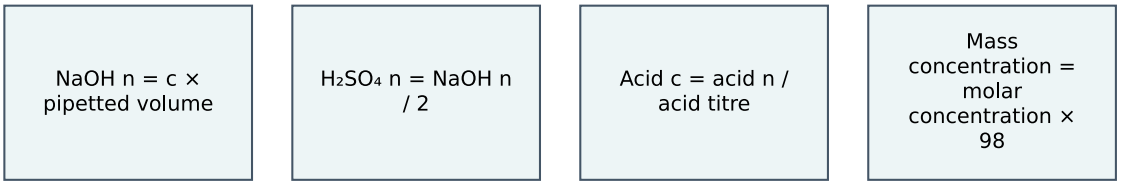
Higher Tier, Chemistry-only: determine acid concentration from RP2 data

Higher Tier focus; see section scope notes.

Official-unit focus: 4.4 Chemical changes

A known sodium hydroxide solution can determine an unknown sulfuric-acid concentration. The acid amount is half the alkali amount because the equation requires two NaOH per H_2SO_4.

For $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$, n(acid)=n(NaOH)/2. First calculate alkali amount as concentration times pipetted volume in dm^3. Divide by two, then divide acid amount by mean acid titre in dm^3. Convert acid molar concentration to mass concentration by multiplying by its molar mass. For supplied H=1, S=32 and O=16, H_2SO_4 has molar mass 98 g per mol.



The handbook PDF page 35 correctly states the equation and the 2:1 ratio, but prints

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a reversed half-mole statement afterwards. The correct inference is acid moles = half alkali moles. The balanced equation gives the correct amount relationship. The common-tier RP2 task still concerns accurate reacting volumes, while this numerical determination is Higher-only.

Use the measured accurate mean from the student's actual supervised task, including the stated trial selection. Record known solution concentration and volume, convert cm^3 to dm^3, and label each chemical's amount. If the question supplies different acid identity or coefficient ratio, change the calculation accordingly. A calculated concentration is not a direct burette reading.

■ Checked worked case

Known: 25.0 cm^3 NaOH at 0.100 mol per dm^3 contains 0.00250 mol. Acid amount=0.00125 mol. With acid mean titre 12.50 cm^3, acid concentration=0.00125/0.01250=0.100 mol per dm^3. Mass concentration=0.100×98=9.80 grams per dm^3. Both units describe the same solution using different quantities.

■ Common error

Do not multiply the alkali amount by two for this acid. Do not use centimetre-cubed volumes without conversion when concentration uses dm^3. The result assumes the stated pure-solution reaction, correct endpoint and sufficiently accurate measurements.

Higher Tier: acid strength differs from acid concentration

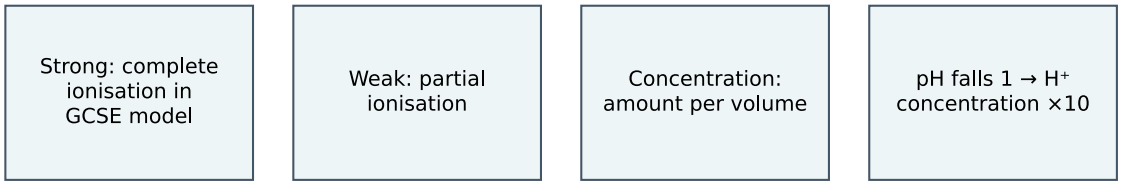
Higher Tier focus; see section scope notes.

Official-unit focus: 4.4 Chemical changes

A dilute strong acid and a concentrated weak acid describe two different comparisons. Strong is about ionisation; concentrated is about how much acid is present per solution volume.

A strong acid is completely ionised in aqueous solution in this GCSE account. Hydrochloric, nitric and sulfuric acids are the named strong examples. A weak acid is only partially ionised; ethanoic, citric and carbonic acids are named examples. Dilute and concentrated describe acid amount per volume, whereas weak and strong describe degree of ionisation. One term cannot replace the other.

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At a given concentration under comparable conditions, the stronger acid has lower pH because it produces greater hydrogen-ion concentration. Each decrease of one pH unit corresponds to a tenfold increase in H^+ concentration. A decrease of two units means a hundredfold increase. Use whole-number pH comparisons in this source scope; do not treat pH as a linear concentration scale.

Compare labelled solutions of known identity and concentration using supplied data or teacher-approved dilute samples and a calibrated pH probe. Separate a strength comparison at fixed concentration from a concentration comparison using the same acid. The GCSE strong-acid wording is a model; do not silently turn it into advanced assumptions about every dissociation step of concentrated sulfuric acid.

■ Checked worked case

Known: solution A has pH 2 and solution B pH 5 under the stated conditions. A has 10^3=1,000 times the H^+ concentration of B. The pH values alone do not identify the acids or their total acid concentrations. A low pH can reflect both ionisation behaviour and the amount present.

■ Common error

Weak does not mean harmless and strong does not mean concentrated. Dilution does not change an acid's named strong/weak classification. A neutral pH 7 solution is not ion-free; it has balanced acid–alkali character in the school account. Avoid predicting pH from a label alone without concentration and conditions.

Electrolysis: mobile ions move to oppositely charged electrodes

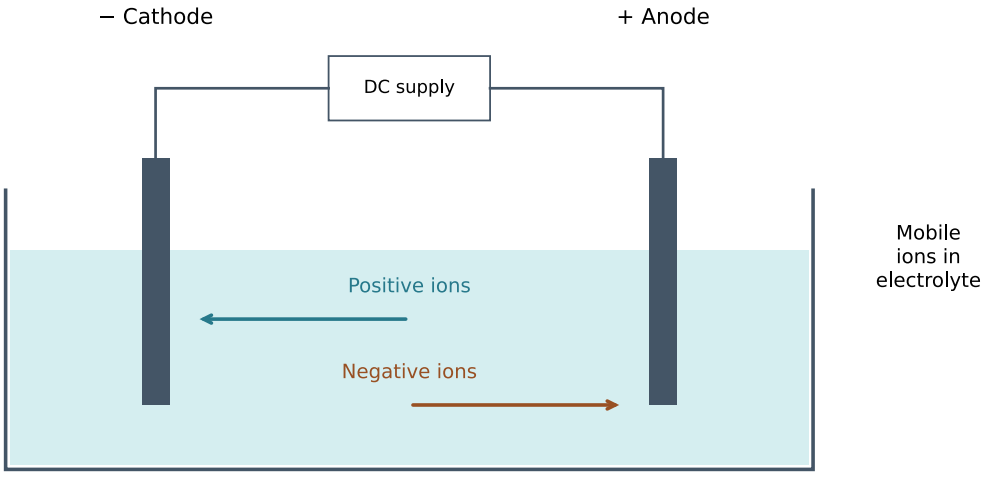
Official-unit focus: 4.4 Chemical changes

An ionic solid does not conduct through fixed ions, but its melt can. Connecting a direct-current supply to that liquid can cause different products at its two electrodes.

An electrolyte is a liquid or solution containing mobile ions that can conduct electricity. Melting an ionic compound or dissolving a suitable ionic compound frees ions to

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move. In electrolysis, the negative electrode is the cathode and attracts positive ions; the positive electrode is the anode and attracts negative ions. Discharged ions form products at the electrode surfaces.



Powered electrolysis: electrodes kept apart; electrons travel in external circuit

Charge travels through the electrolyte by moving ions and through the external circuit by electrons. Positive ions are not positive electrons, and solid salt ions do not become mobile simply because a wire touches the solid. Electrolysis uses electrical energy to drive chemical change. The polarity names here apply to an externally powered electrolytic cell, not a blanket rule for every electrochemical cell.

Draw a labelled direct-current supply, two separated electrodes and the electrolyte. Mark supply polarity, electrode identity and arrows for each ion type. Actual school work uses approved low-voltage equipment and solutions, with electrodes held apart to prevent short circuits. Record products and observations at the named electrode instead of merely saying bubbles occurred somewhere.

■ Checked worked case

Known: a labelled model has electrode A connected to the negative supply terminal and B to the positive. Cu^{2+} moves towards A, while Cl^{-} moves towards B. A is the cathode and B the anode. In a particle-count illustration, five Cu^{2+} and ten Cl^{-} have balanced total charge $+10-10=0$; equal ion counts are not required for neutrality.

■ Common error

The electrode signs attract ions but do not imply that every positive ion forms its metal in an aqueous solution; water-derived species can compete. That product-selection rule is separate. Higher-only half equations are also separate from this common-tier movement and apparatus model.

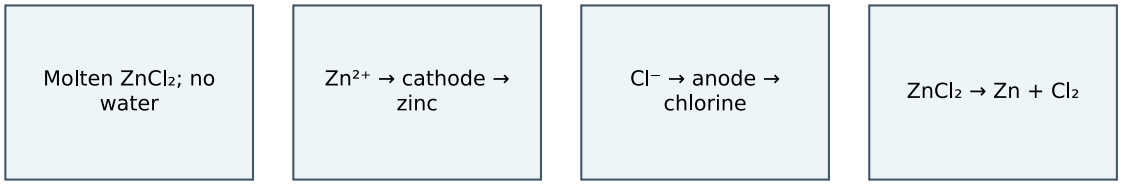
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Molten binary compounds: metal at the cathode, non-metal at the anode

Official-unit focus: 4.4 Chemical changes

Molten sodium chloride and sodium chloride dissolved in water do not give the same cathode product. The molten compound has no competing water-derived hydrogen ions.

A binary ionic compound contains two elements. When its melt is electrolysed with inert electrodes, its metal forms at the negative cathode and its non-metal at the positive anode. Molten lead bromide gives lead and bromine; molten sodium chloride gives sodium and chlorine. Ions must be free to move, so the solid form is not the conducting electrolyte in this model.



Identify the positive metal ion and negative non-metal ion from the formula. The cathode product is the metal element, while a halogen anode product is diatomic, such as Br_2 or Cl_2. Do not write bromide as the elemental product. The formula's ratio is determined by charge, and the product equation must preserve every atom.

Use diagrams, acquired observations or a teacher-approved demonstration rather than prescribing independent molten lead bromide work. The specification names anhydrous zinc chloride as a safer alternative, but it still needs an approved risk assessment and heated apparatus. Hot corrosive melts and toxic gas products cannot be treated like an ordinary room-temperature salt solution.

■ Checked worked case

Known: $ZnCl_2$ contains one Zn^{2+} for two Cl^- ions. Its overall decomposition is $ZnCl_2 \rightarrow Zn + Cl_2$. Six represented formula units can give six zinc atoms and six chlorine molecules, containing twelve chlorine atoms. This equation is common-tier product accounting; the electron half equations are taught separately at Higher Tier.

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■ Common error

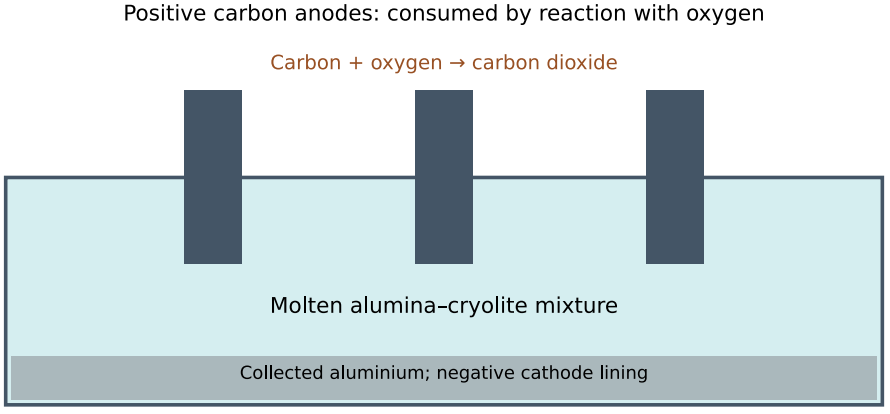
Molten is not the same as aqueous. Water changes the available species and product rules. Inert means the electrode is not intended to take part in the stated reaction, not that every material survives all conditions. Products need both the element name and the correct electrode.

Aluminium extraction: a molten mixture and a consumed carbon anode

Official-unit focus: 4.4 Chemical changes

Aluminium oxide has a very high melting point. Industrial electrolysis uses a molten oxide–cryolite mixture instead of simply melting pure oxide at its own high melting temperature.

Electrolysis extracts metals from molten compounds when they are too reactive for carbon reduction or react with carbon. It requires energy both to maintain the molten electrolyte and to provide electrical current. Aluminium manufacture uses aluminium oxide dissolved in molten cryolite. The mixture melts at a lower temperature than pure aluminium oxide, reducing the heating demand.



Cryolite lowers operating temperature; heating and current still need energy

Aluminium forms at the negative cathode. Oxygen associated with oxide discharge at the positive carbon anode reacts with carbon, forming carbon dioxide in the GCSE account. The carbon anode is consumed and needs continual replacement. It is therefore not an inert electrode like the one assumed in a simple product-prediction exercise.

Interpret a labelled industrial cell using supply polarity, electrolyte composition, metal collection and anode material. Use supplied electricity, temperature or electrode-replacement data to evaluate cost or resource demand. This is an industrial reference case, not a school task to build a hot extraction cell. Avoid adding unrequired plant details that obscure the two required explanations.

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■ Checked worked case

Known: a supplied operating comparison states heating-energy demand of 900 units for the pure-oxide scenario and 600 for the mixture scenario on the same basis. Reduction=300 units, percentage reduction=300/900×100=33.3%. This fictional dataset illustrates a stated comparison; it is not a published industrial efficiency claim or total-life-cycle energy estimate.

■ Common error

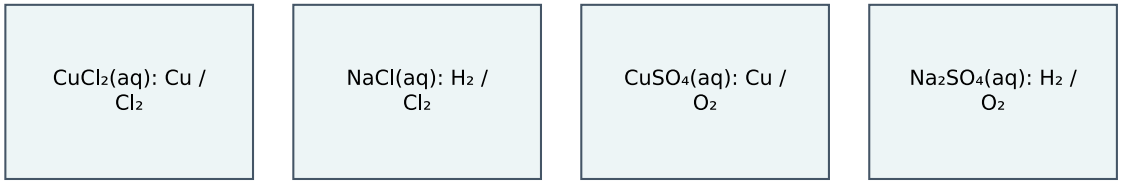
Cryolite does not provide the aluminium metal being extracted from the alumina feed in the simplified account. Lower heating demand does not remove the electrical-current requirement. Carbon consumption follows its chemical reaction at the anode, not simply mechanical wear. Do not predict oxygen collection unchanged at an anode that reacts with it.

Aqueous electrolysis: water-derived species change the products

Official-unit focus: 4.4 Chemical changes

Aqueous sodium chloride gives hydrogen at the cathode rather than sodium. The solution contains water-derived species as well as the dissolved salt ions.

For the specification's inert-electrode school model, hydrogen forms at the cathode if the dissolved metal is more reactive than hydrogen; a less reactive metal such as copper can form instead. At the anode, oxygen forms unless halide ions are present, when the halogen is predicted. The GCSE account describes water supplying hydrogen and hydroxide ions that can be discharged.



Aqueous copper(II) chloride therefore gives copper at the cathode and chlorine at the anode. Sodium chloride solution gives hydrogen and chlorine under the approved handbook conditions. Copper(II) sulfate gives copper and oxygen, while sodium sulfate gives hydrogen and oxygen. State that these are the specified single-solute, inert-

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electrode predictions; concentration and electrode material can affect actual competing reactions outside this simple rule.

Identify the dissolved positive and negative ions, then include water-derived ions in the explanation. Apply cathode and anode rules separately. Use only school-approved dilute solutions and a low-voltage supply, with fixed electrode spacing and appropriate ventilation. Chlorine-containing observations must follow the technician-approved small-scale method; do not smell gases or generate large quantities.

■ Checked worked case

Known: four supplied solutions are $CuCl_2$, NaCl, $CuSO_4$ and Na_2SO_4 in water with inert electrodes. Two have copper cathode products and two hydrogen; two have chlorine anode predictions and two oxygen. This classification counts predictions from stated conditions, not proof of identical product rates or recovery across the four solutions.

■ Common error

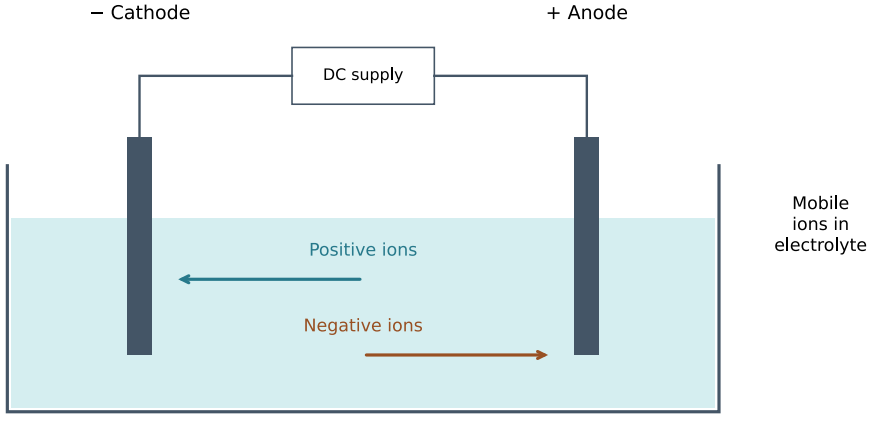
Dissolved chloride gives chlorine, not solid chloride ions, at the anode in the stated model. Aqueous sulfate does not produce sulfur by simply stripping its name from the formula. Do not apply the molten binary rule to water-containing solutions.

Required Practical 3: test an electrolysis-product hypothesis

Official-unit focus: 4.4 Chemical changes

Comparing copper chloride with sodium chloride lets a student test a prediction about cathode products while keeping chloride as the common anion.

Required Practical 3 is an actual investigation of aqueous electrolysis using inert electrodes and a developed hypothesis. A suitable prediction is that copper chloride will give a copper deposit at the cathode, while sodium chloride will give hydrogen under the approved conditions; both will show chlorine at the anode. State the reason using metal reactivity and the specified aqueous product rules before observing.



Powered electrolysis: electrodes kept apart; electrons travel in external circuit

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The handbook uses carbon rods, low-voltage direct current and approved $CuCl_2$/NaCl solutions. A brown/red deposit supports copper at the negative electrode; gas at the positive electrode bleaches damp blue litmus under the approved test. A bubble observation alone does not identify a gas. Label each observation with electrode sign and keep the products separate.

The teacher/technician approves concentrations, small quantities, voltage, duration, ventilation and disposal. Keep the carbon rods apart and fixed to avoid a short circuit. Use the prescribed contained small-scale gas test and never inhale chlorine. Record matched solution volumes, electrode spacing, supply setting and elapsed time. Retain actual results, labelled apparatus and a conclusion comparing them with the hypothesis.

■ Checked worked case

Known: a supplied model comparison records gas collection 12 cm^3 after 4 min and 18 cm^3 after 6 min under stated equal conditions. Both mean collection rates are 3 cm^3 per minute. Equal rates do not establish gas identity; the electrode location, predicted reaction and appropriate product test provide that evidence.

■ Common error

Do not label invented measurements as a completed practical. Inert-electrode predictions need review if the electrode material reacts. Changes in concentration, spacing or supply can alter the comparison. The handbook notes chlorine production and short controlled operation; follow the current school risk assessment rather than treating a written lesson as full practical permission.

Higher Tier: electrode half equations conserve atoms and charge

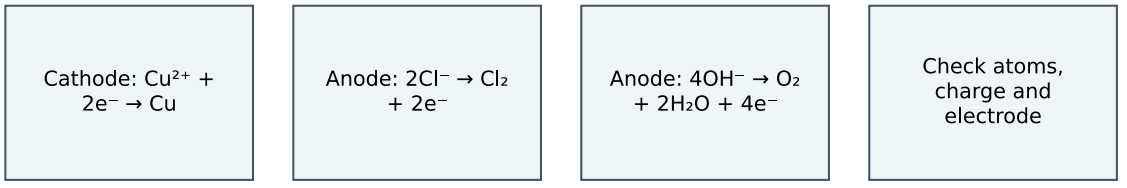
Higher Tier focus; see section scope notes.

Official-unit focus: 4.4 Chemical changes

At a cathode, copper ions become copper by accepting electrons. At an anode, chloride ions become chlorine by releasing electrons. The two electrode changes need separate equations.

Cathode reduction examples are $Cu^{2+} + 2e^- \rightarrow Cu$, $Al^{3+} + 3e^- \rightarrow Al$ and $2H^+ + 2e^- \rightarrow H_2$. Anode oxidation examples are $2Cl^- \rightarrow Cl_2 + 2e^-$ and $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$. Electrons appear on the reactant side when gained and the product side when lost. The OH^- equation conserves both atoms and total charge.

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For $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$, oxygen count is four on each side and hydrogen count four on each side. Charge is -4 on both sides. For $2Cl^- \rightarrow Cl_2 + 2e^-$, chlorine is diatomic and the electron total balances the two negative charges. A correct formula is needed before balancing; changing Cl_2 to Cl would misidentify the product.

Write the species and intended products first, then balance atoms, then charge with electrons. Check electrode sign and oxidation/reduction identity. Multiply a half equation only when matching overall electron transfer or the question's required amount. Distinguish the basic metal/halide molten examples from water-derived aqueous examples and the industrial reacting carbon anode.

■ Checked worked case

Known: three represented Cu^{2+} ions require six electrons to form three Cu atoms. Four represented Al^{3+} ions require twelve. Six represented Cl^- ions form three Cl_2 molecules and release six electrons. This count supports the HT equation model; it does not require calculating electrical charge, current efficiency or Faraday's constant outside the GCSE objectives.

■ Common error

Electrons do not disappear from charge accounting just because they are not atoms. Anode oxidation is electron loss in this powered electrolysis model. Positive-ion discharge does not mean gaining positive electrons. The entire 4.4.3.5 section and the half-equation requirement across 4.4.3 are Higher-only.

Energy transfers: explain warming and cooling

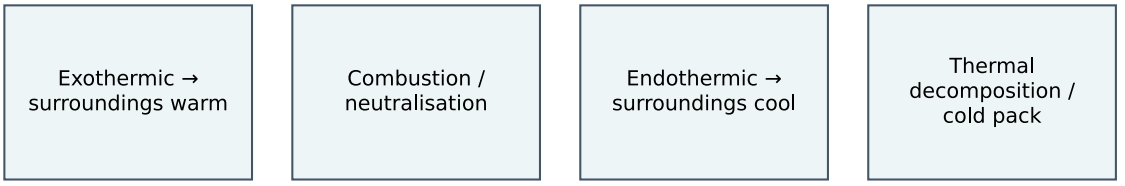
Official-unit focus: 4.5 Energy changes

A hand warmer releases energy to your hands. An instant cold pack takes energy from its surroundings. The direction of energy transfer explains their different uses.

Energy is conserved in chemical reactions. An exothermic reaction transfers energy to the surroundings, which become warmer; products have less energy than reactants by

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the amount transferred. Combustion, many oxidation reactions and neutralisation are exothermic examples. An endothermic reaction takes energy from the surroundings, which cool; products have more energy than reactants. Thermal decomposition and citric acid reacting with sodium hydrogencarbonate are endothermic examples.



Distinguish the reacting chemicals from their surroundings. A thermometer in a reacting solution measures the temperature of that solution, which receives or supplies energy during the chemical change. A temperature rise supports an exothermic interpretation under the stated conditions; a fall supports endothermic behaviour. Heating a vessel externally can obscure this evidence. Not every process involving cooling is a chemical reaction: new substances must also be formed.

Evaluate a supplied warmer or cold-pack design against its purpose: suitable temperature, duration, risk of leakage, storage and single-use waste. Use information supplied for the particular product. Many self-heating cans exploit exothermic changes; some sports injury packs use endothermic changes. These examples explain energy transfers and do not recommend applying an untested chemical mixture to skin.

■ Checked worked case

Known: a supervised reaction starts at 21.0 °C and reaches 28.5 °C. The temperature rise is 28.5−21.0=7.5 °C, supporting an exothermic interpretation with no external heater. A second starts at 22.0 °C and falls to 17.0 °C: the change is −5.0 °C and its cooling magnitude is 5.0 °C. Energy has been transferred, not destroyed.

■ Common error

AQA 4.5.1.1 requires measuring temperature changes, not calculating transferred energy using heat capacity or calculating molar enthalpy. Those quantities are not substitutes for the required temperature evidence. A warmer that becomes hotter is not necessarily more suitable if it can burn skin or runs out too quickly.

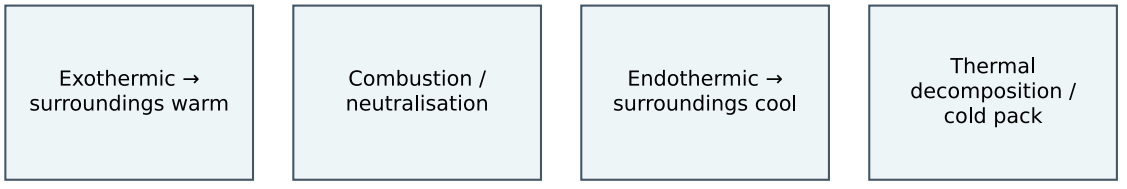
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Energy transfers: explain warming and cooling

Official-unit focus: 4.5 Energy changes

A hand warmer releases energy to your hands. An instant cold pack takes energy from its surroundings. The direction of energy transfer explains their different uses.

Energy is conserved in chemical reactions. An exothermic reaction transfers energy to the surroundings, which become warmer; products have less energy than reactants by the amount transferred. Combustion, many oxidation reactions and neutralisation are exothermic examples. An endothermic reaction takes energy from the surroundings, which cool; products have more energy than reactants. Thermal decomposition and citric acid reacting with sodium hydrogencarbonate are endothermic examples.



Distinguish the reacting chemicals from their surroundings. A thermometer in a reacting solution measures the temperature of that solution, which receives or supplies energy during the chemical change. A temperature rise supports an exothermic interpretation under the stated conditions; a fall supports endothermic behaviour. Heating a vessel externally can obscure this evidence. Not every process involving cooling is a chemical reaction: new substances must also be formed.

Evaluate a supplied warmer or cold-pack design against its purpose: suitable temperature, duration, risk of leakage, storage and single-use waste. Use information supplied for the particular product. Many self-heating cans exploit exothermic changes; some sports injury packs use endothermic changes. These examples explain energy transfers and do not recommend applying an untested chemical mixture to skin.

■ Checked worked case

Known: a supervised reaction starts at 21.0 °C and reaches 28.5 °C. The temperature rise is 28.5−21.0=7.5 °C, supporting an exothermic interpretation with no external heater. A second starts at 22.0 °C and falls to 17.0 °C: the change is −5.0 °C and its cooling magnitude is 5.0 °C. Energy has been transferred, not destroyed.

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■ Common error

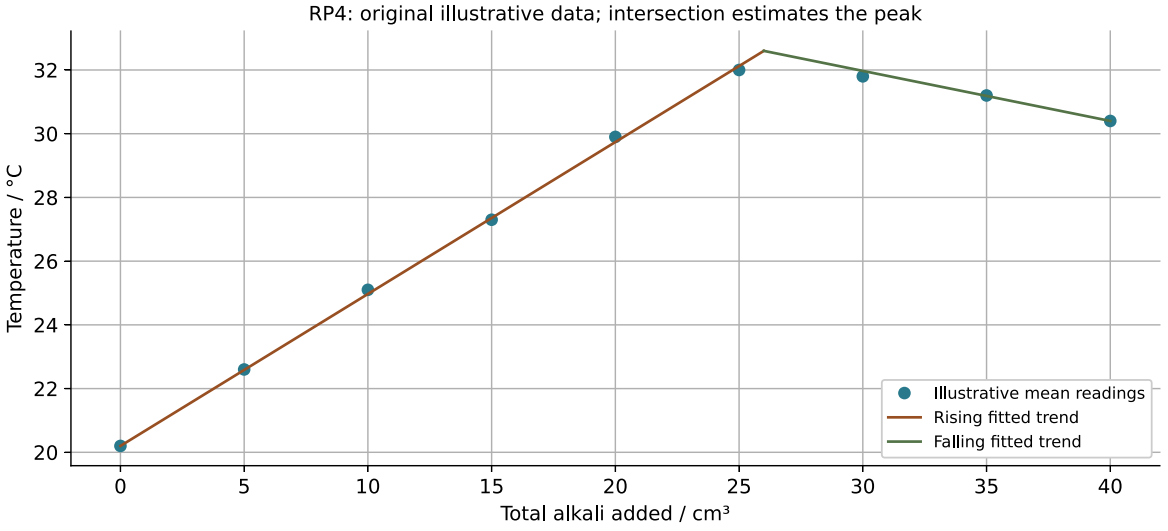
AQA 4.5.1.1 requires measuring temperature changes, not calculating transferred energy using heat capacity or calculating molar enthalpy. Those quantities are not substitutes for the required temperature evidence. A warmer that becomes hotter is not necessarily more suitable if it can burn skin or runs out too quickly.

Required practical 4: investigate temperature changes in solutions

Official-unit focus: 4.5 Energy changes

Adding alkali to acid first raises the temperature, but later additions can make the mixture cooler. A falling part of this graph does not automatically mean neutralisation became endothermic.

Required practical 4 investigates variables affecting temperature change in reacting solutions: suitable acid–metal, acid–carbonate, neutralisation or displacement reactions. In the handbook neutralisation activity, a polystyrene cup is supported in a beaker and fitted with a lid and thermometer. Students start with 30 cm^3 dilute hydrochloric acid, record its temperature, then add teacher-prepared sodium hydroxide in successive 5 cm^3 portions up to 40 cm^3.



After each addition, replace the lid, stir gently and record the highest temperature. Repeat the whole investigation with fresh starting solutions and calculate means at each added volume. Plot mean temperature against total alkali volume added. Draw appropriate best-fit lines through the rising and falling regions and estimate their intersection. It estimates the peak between discrete readings; do not join every noisy point and call the highest measured point exact.

Use the school-approved risk assessment, eye protection and technician-prepared solutions. Sodium hydroxide presents a particular eye hazard. Keep initial acid volume,

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solution concentrations, starting temperatures, cup, stirring and reading procedure comparable. Record actual observations and an apparatus sketch in the laboratory; this lesson prepares and interprets that work, rather than certifying performance from a written answer. A lid and insulation reduce heat exchange but do not eliminate it.

■ Checked worked case

Known: an original illustrative pair of runs at one volume gives 30.4 °C and 30.8 °C, so mean temperature=(30.4+30.8)/2=30.6 °C. Starting mean temperature is 20.2 °C, giving a rise of 10.4 °C. After acid is exhausted, further cooler alkali adds no comparable neutralisation warming, while mixing and energy transfer to surroundings can lower temperature. This does not establish an endothermic neutralisation.

■ Common error

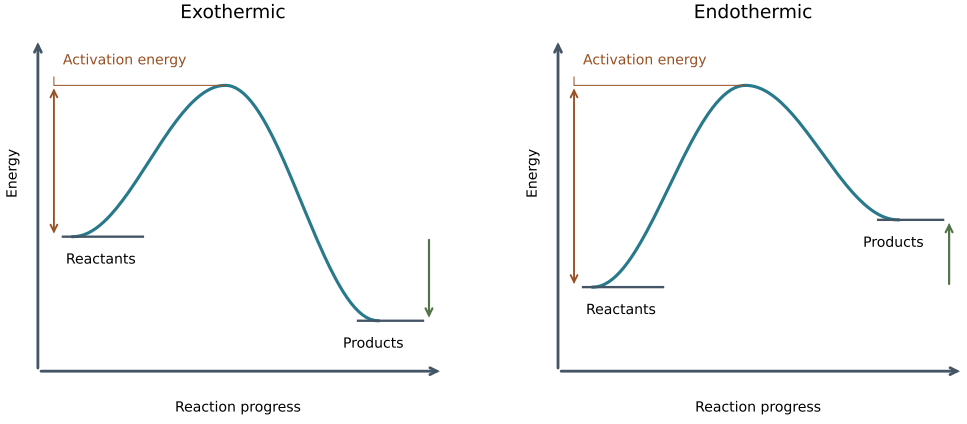
Do not prepare concentrated stock chemicals independently, heat this mixture to force a peak, or remove inconvenient repeats without recording a reason. A greater added volume changes both reactant amount and total liquid volume, so the graph is not a simple measure of reaction energy per mole. No heat-capacity calculation is required here. Bond-breaking explanations belong to the separate Higher lesson.

Reaction profiles: distinguish the barrier from the overall change

Official-unit focus: 4.5 Energy changes

An exothermic reaction may still need a spark to start. Releasing energy overall does not remove the initial energy barrier that reacting particles must overcome.

Reactions occur when particles collide with sufficient energy. Activation energy is the minimum energy particles must have for the reaction to occur. A reaction profile plots energy vertically against reaction progress horizontally. The reactant level leads up to a curved peak, then down to the product level. The activation-energy arrow runs from the reactant level to the peak; it is not measured from the graph's arbitrary zero.



Green arrows: overall change; both profiles have an activation barrier

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In an exothermic profile products lie below reactants, showing an overall energy decrease of the reacting chemicals and transfer to surroundings. In an endothermic profile products lie above reactants, showing energy taken from surroundings. Both profiles can have an activation barrier. The vertical difference between reactants and products is the overall change, distinct from the larger climb to the peak. Horizontal distance is not elapsed time or a measured reaction rate.

Draw axes first, mark clearly different reactant and product levels, then join them with a smooth curve rising above both. Label reactants, products and activation energy. Add a separate vertical arrow for the overall energy change. Compare the two diagrams using level differences, not the apparent length of the curve. Use relative energy units if numbers are supplied; they illustrate a model rather than requiring calorimetry.

■ Checked worked case

Known: a supplied relative-energy model has reactants at 30, peak at 85 and products at 10 units. Activation energy is 85−30=55 units; the reacting chemicals decrease by 30−10=20 units, so it is exothermic. If products instead lie at 50 with the same reactants and peak, the increase is 20 units and the reaction is endothermic. The initial barrier remains 55 units.

■ Common error

A peak is not an intermediate product that must be collected. Do not label reaction progress as time, equate activation energy to overall energy change or assume exothermic reactions always occur rapidly at room temperature. The source requires qualitative profiles; numerical relative levels are a way to practise reading their geometry.

Higher Tier: count bonds before calculating the energy change

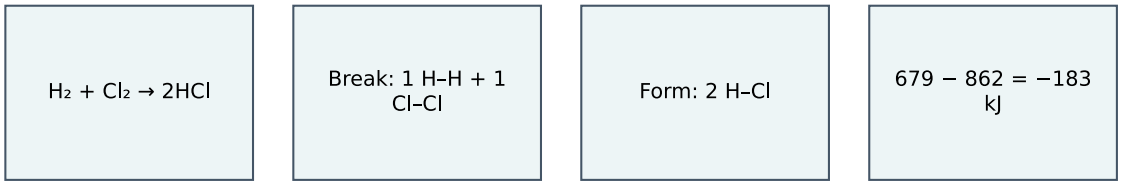
Higher Tier focus; see section scope notes.

Official-unit focus: 4.5 Energy changes

Breaking a bond needs energy. An exothermic reaction releases energy overall because formation of new bonds releases more than was needed to break the old ones.

Energy must be supplied to break bonds in reactants. Energy is released when bonds form in products. Using supplied bond energies, overall energy change=total energy needed for bonds broken−total energy released for bonds formed. A negative answer means bond formation releases more than breaking requires, so the reaction is exothermic. A positive answer means the reverse and the reaction is endothermic.

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Use the balanced equation and actual bond types. A coefficient multiplies all bonds in that molecule. For $H_2 + Cl_2 \rightarrow 2HCl$, break one H–H and one Cl–Cl and form two H–Cl bonds. For $2H_2 + O_2 \rightarrow 2H_2O$, break two H–H bonds and one O=O bond, and form four O–H bonds. An O=O double bond uses its supplied double-bond value once; do not double a single-bond value.

Make separate broken and formed lists with counts, multiply each supplied value by its count, total each list, then subtract in the stated order. These values give an approximate model for the reaction quantities represented by the equation. Check units and sign and interpret the result in words. Do not infer this calculated value from a thermometer reading without a different model; AQA's solution practical does not require that calculation.

■ Checked worked case

Known: supplied bond energies are H–H 436, Cl–Cl 243 and H–Cl 431 kJ per mole of bonds. For $H_2 + Cl_2 \rightarrow 2HCl$, breaking requires 436+243=679 kJ; forming releases 2×431=862 kJ. Overall change=679−862=−183 kJ for the equation amounts: exothermic. The released magnitude is 183 kJ, but the signed change is negative. The balanced equation is essential to the count.

■ Common error

Do not call breaking bonds exothermic, subtract formed from broken in the reverse order, or count atoms as if each were a bond. A negative energy change is not negative activation energy. This entire source section is Higher-only, while drawing the profiles remains common-tier content.

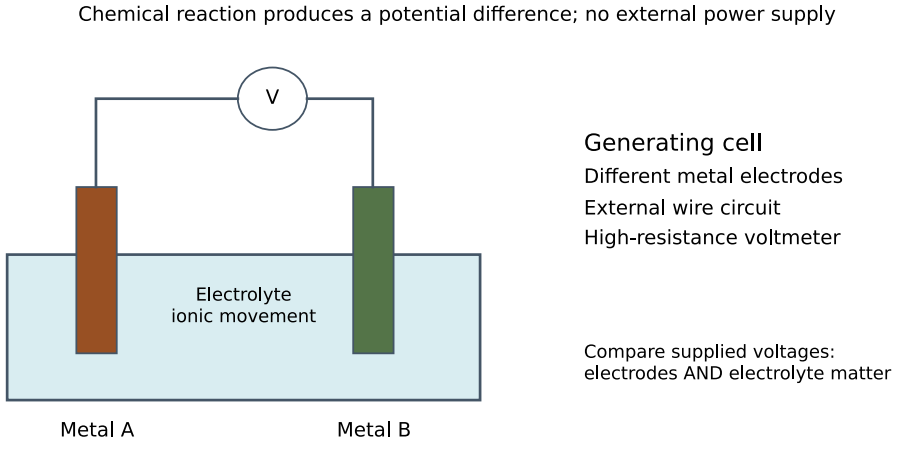
Separate Chemistry: cells produce a potential difference

Official-unit focus: 4.5 Energy changes

Two different metal electrodes touching an electrolyte can produce a potential difference. This chemical cell supplies electricity, unlike electrolysis driven by an external power supply.

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Chemical reactions in cells produce electricity. A simple cell consists of two different metals in contact with an electrolyte and connected through an external circuit. The electrolyte permits ionic movement, while electrons move through the metal wires. The voltage produced depends on electrode materials and the electrolyte. A battery contains two or more cells connected in series to give a greater voltage when their polarities are aligned.



Use supplied readings or reactivity information to compare cells under specified conditions. Do not promise a universal voltage from a metal pair alone: electrolyte choice and conditions also matter. Connecting identical cells in series adds their potential differences in the ideal stated model; reversed orientation subtracts one cell's contribution. Voltage measures potential difference, not the amount of chemical reactant remaining or the rate of all reactions.

For actual school work use approved metal strips and electrolytes, a high-resistance voltmeter and consistent exposed areas, separation, temperature and cleaning. Keep electrodes apart so they do not directly short together. Read the sign and magnitude with the stated meter connections. Do not attach a bench power supply to this generating-cell comparison or attempt to charge an unsuitable cell. Detailed commercial cell chemistry is not required by this section.

■ Checked worked case

Known: an original supplied table gives zinc–copper cell readings of 0.90 V in electrolyte A and 0.65 V in electrolyte B under stated conditions. This supports an electrolyte effect, not a claim about every zinc–copper cell. Three identical 0.90 V cells in aligned series give an ideal total of 2.70 V. Two aligned and one reversed would give 0.90 V in that simplified model.

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■ Common error

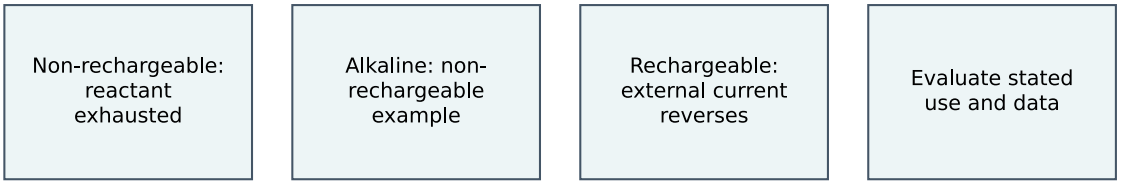
The electrolyte does not carry free electrons in the same way as the wire. A battery is not just any one cell in this specification's terminology. An externally powered electrolytic cell has a different purpose and electrode context. Do not transfer an electrode sign rule between the two apparatus without identifying how it is operating.

Separate Chemistry: choose cells using lifespan and charging evidence

Official-unit focus: 4.5 Energy changes

A torch cell eventually stops supplying electricity because the chemical reactants are used up. A rechargeable cell is designed so an external current can reverse the reaction.

Non-rechargeable cells stop producing electricity when one of the reactants has been used up. Alkaline cells are non-rechargeable examples in this specification. Rechargeable cells and batteries can be recharged because an external electrical current reverses the chemical reaction. Charging transfers energy into the device; it does not create unlimited energy or permanently restore every cell to its original condition.



Choose a cell for its intended use. Compare required voltage, available operating time, mass, initial cost, replacement frequency, charging access and disposal information supplied. Repeated use can make a rechargeable option cheaper over time despite greater initial cost. A rarely used emergency device might place more weight on storage performance and immediate availability. These conclusions depend on stated data, not a universal rule that one cell type is always best.

Read the manufacturer's stated type and use only the approved charger and school procedure for an actual recharge demonstration. Do not charge alkaline cells or construct an improvised charger. A written comparison can use fictional costs safely: state whether electricity, charger purchase and end-of-life disposal are included. Keep environmental and financial criteria distinct so a low purchase price is not automatically called lower waste.

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■ Checked worked case

Known: a supplied non-rechargeable option costs £2 per complete use. A rechargeable option costs £12 initially and £0.10 per use to charge, including the stated charger. At ten uses costs are £20 versus £12+10×£0.10=£13, a £7 saving for the rechargeable option. At one use its £12.10 cost exceeds £2. This comparison assumes both meet the same performance need and remain usable.

■ Common error

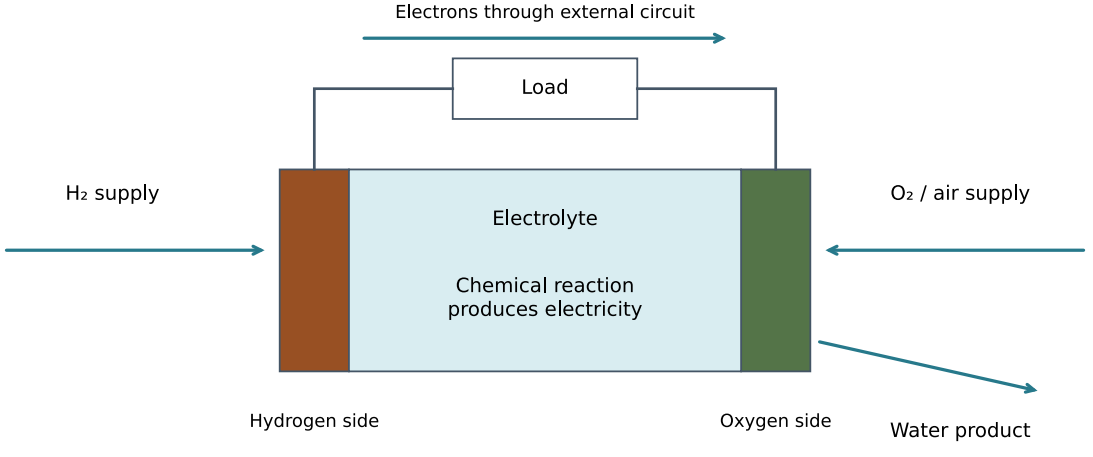
Do not call recharging the replacement of consumed chemicals with no energy input. Non-rechargeable does not mean unable to produce electricity repeatedly before exhaustion. A rechargeable lifetime is finite; a cost projection beyond supplied cycle life is unsupported. This is both-tier separate Chemistry, without advanced electrode mechanisms.

Separate Chemistry: evaluate a continuously supplied fuel cell

Official-unit focus: 4.5 Energy changes

A hydrogen fuel cell continues operating while hydrogen and oxygen are supplied from outside. Its local water product does not by itself show how the hydrogen was produced.

Fuel cells receive fuel such as hydrogen and oxygen or air from an external source. The fuel is oxidised electrochemically, producing a potential difference. In a hydrogen fuel cell the overall reaction is $2H_2 + O_2 \rightarrow 2H_2O$. This transfers chemical energy electrically to a circuit rather than requiring the fuel to burn in a flame. External fuel supply distinguishes it from a rechargeable battery containing its stored reactants.



$2H_2 + O_2 \rightarrow 2H_2O$; continuous external fuel supply

Compare both devices against a purpose. Supplied information may include mass, operating time, refuelling or charging time, infrastructure, fuel storage and cost. Wa-

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ter is the direct product of the hydrogen–oxygen reaction; hydrogen production and transport may have other environmental effects. A conclusion about total emissions needs data for those stages. Do not infer zero lifecycle carbon emissions merely from the absence of carbon in the local equation.

Use a teacher-approved educational fuel-cell kit or supplied observations; never improvise compressed hydrogen storage or a hydrogen–oxygen gas mixture. Record how fuel is supplied, where the electrical circuit connects and what product forms. A comparison task must use the given boundary: device-only, fuel production or a wider lifecycle. Detailed commercial engineering is not required, and the electron half equations are in the separate Higher lesson.

■ Checked worked case

Known: a supplied model device runs 4 hours per fuel cartridge; a rechargeable device runs 3 hours per full charge. Twelve uninterrupted hours require three cartridges or four battery charge-equivalents, ignoring the stated changeover time. This is an operating-time comparison, not proof of lower cost or environmental impact. In the balanced chemical equation, six represented H_2 molecules need three O_2 and form six H_2O molecules.

■ Common error

A fuel cell does not manufacture its own limitless hydrogen, and oxygen is a reactant, not its waste product. A rechargeable battery requires charging energy while the fuel cell needs a continuing external fuel supply. Neither device is automatically the best under every combination of cost, storage and environmental conditions.

Higher Tier: combine hydrogen fuel-cell half equations

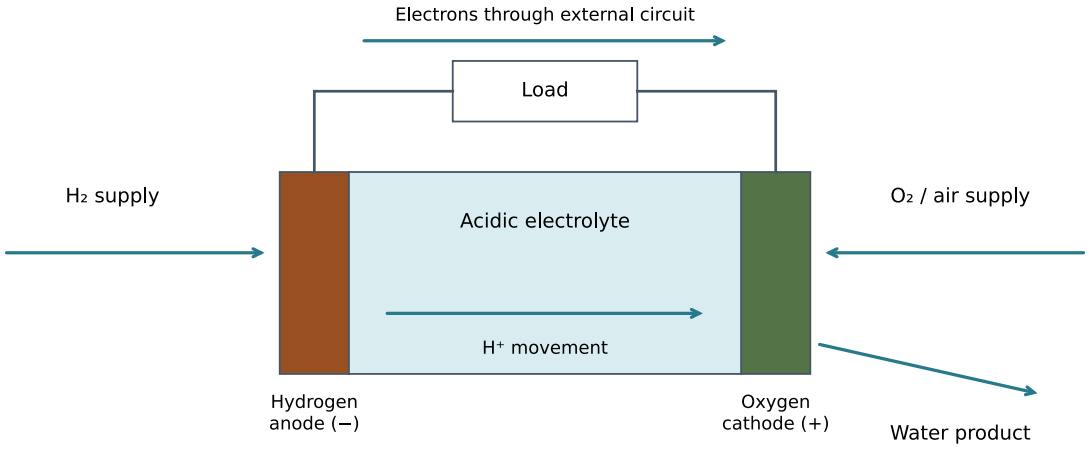
Higher Tier focus; see section scope notes.

Official-unit focus: 4.5 Energy changes

A hydrogen fuel cell separates oxidation and reduction at two electrodes. The same number of electrons must be released at one electrode and accepted at the other.

For the explicitly acidic-electrolyte model, hydrogen oxidation is $2H_2 \rightarrow 4H^+ + 4e^-$. Oxygen reduction is $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$. Adding the half equations cancels four hydrogen ions and four electrons, leaving $2H_2 + O_2 \rightarrow 2H_2O$. Oxidation means electron loss and reduction electron gain. These half equations are the embedded Higher-only part of the source section.

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$2H_2 + O_2 \rightarrow 2H_2O$; generating cell, explicitly acidic model

Check atoms and charges separately. The hydrogen half equation has four hydrogen atoms on each side and zero net charge because +4 from H^+ balances −4 from electrons. The oxygen half equation has two oxygen atoms and four hydrogen atoms on each side, again zero net charge. In this generating cell electrons leave the hydrogen anode through the external circuit and reach the oxygen cathode. The anode is negative and the cathode positive while the cell supplies electricity.

Begin with one consistent electrolyte model. Balance hydrogen with H^+, oxygen with water and charge with electrons, then scale to equal electron numbers and cancel. Do not mix an acidic half equation with an alkaline one containing OH^-. The electrode names identify oxidation at the anode and reduction at the cathode; signs must be interpreted for a generating fuel cell, rather than copied from an externally powered electrolytic cell.

■ Checked worked case

Known: starting from $H_2 \rightarrow 2H^+ + 2e^-$, multiply by two to match the four electrons consumed by $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$. Two represented H_2 molecules supply four electrons and one O_2 molecule accepts four. If three O_2 molecules react in the same scaled model, twelve electrons pass through the external circuit and six H_2O molecules form. Electrons cancel from the overall equation.

■ Common error

Do not put electrons on the reactant side for hydrogen oxidation, allow them to remain in the overall reaction, or say they pass through the electrolyte like they do through a metal wire. Common-tier students still need the overall reaction and fuel-cell evaluation, but these half equations and charge reasoning are Higher-only.

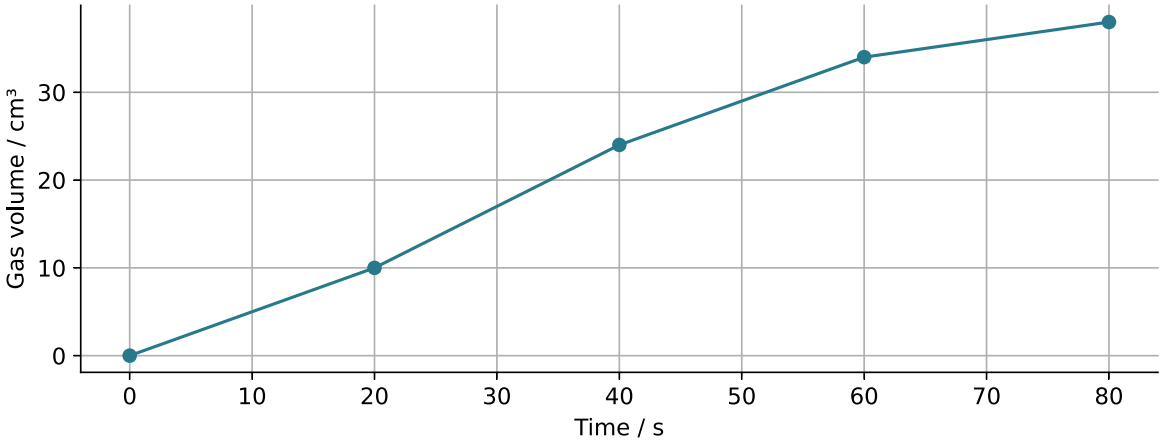
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Rates, catalysts and reliable endpoints

Official-unit focus: 4.6 Rate and extent of chemical change

A faster reaction finishes sooner, but it need not make more product. Rate and final yield answer different questions.

Reaction rate describes reactant used or product formed per time. Higher temperature increases the fraction of collisions with enough energy. A catalyst provides an alternative pathway with lower activation energy.



A product-time graph has a steeper gradient where rate is larger. A tangent estimates instantaneous rate; a secant gives average rate over an interval. The final plateau reflects the total collected product under the stated conditions.

For gas production, check apparatus for leaks, start timing consistently and record volume at regular intervals. Keep concentration, reactant amount and surface area controlled when changing temperature.

■ Checked worked case

Known: gas volume increases from 10 to 34 cubic centimetres between 20 and 60 s. Average rate = change in volume/change in time. Rate = (34 - 10)/(60 - 20) = 0.60 cubic centimetres per second. This is not necessarily the instantaneous rate at 40 s.

■ Common error

A catalyst does not change the final equilibrium composition at fixed conditions. A mass-loss method cannot detect all reactions, and losing gas through a leak biases a collection experiment.

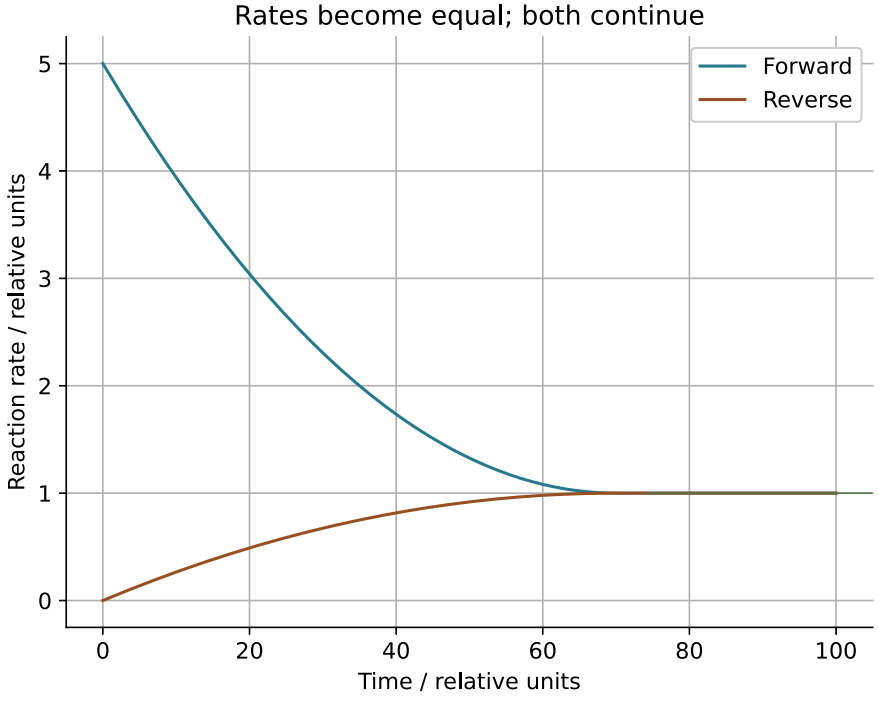
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Dynamic equilibrium: equal rates, continuing reactions

Official-unit focus: 4.6 Rate and extent of chemical change

A closed reaction vessel can show constant composition while particles still react in both directions. Equality of rates keeps the overall amounts unchanged.

When a reversible reaction occurs in apparatus preventing escape of reactants and products, equilibrium is reached when forward and reverse reactions occur at exactly the same rate. The state is dynamic: both directions continue. At fixed conditions the amounts of reactants and products stay constant, but they need not be equal. A closed system prevents loss of matter; it does not necessarily prevent energy transfer to maintain a set temperature.



Beginning with mostly reactants, the forward rate may be high while the reverse rate is initially low. As products accumulate, reverse reaction becomes possible more often until the rates match under the stated model. A reaction-rate graph should show convergence to the same non-zero rate. A composition graph can level off at different reactant and product amounts. Identify which quantity is on the axis before claiming two curves must meet.

Use an original particle-count model or supplied data with fixed conditions and no matter lost. Count forward and reverse events over equal time windows. Their difference gives net change for a one-to-one A $\rightleftharpoons$ B model. Do not build a sealed heated chemical apparatus to demonstrate the idea; safe school demonstration arrangements require approved containment and temperature control. A count model illustrates balance without proving a particular real chemical rate law.

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■ Checked worked case

Known: in a closed A ⇌ B model, 18 forward and 18 reverse events occur in one second. Net B change=18−18=0 even though 36 reaction events occur. If 70 A and 30 B particles remain on average, these unequal amounts are compatible with equilibrium. At an earlier interval with 20 forward and 12 reverse events, net B increases by eight and the model is not yet at equilibrium.

■ Common error

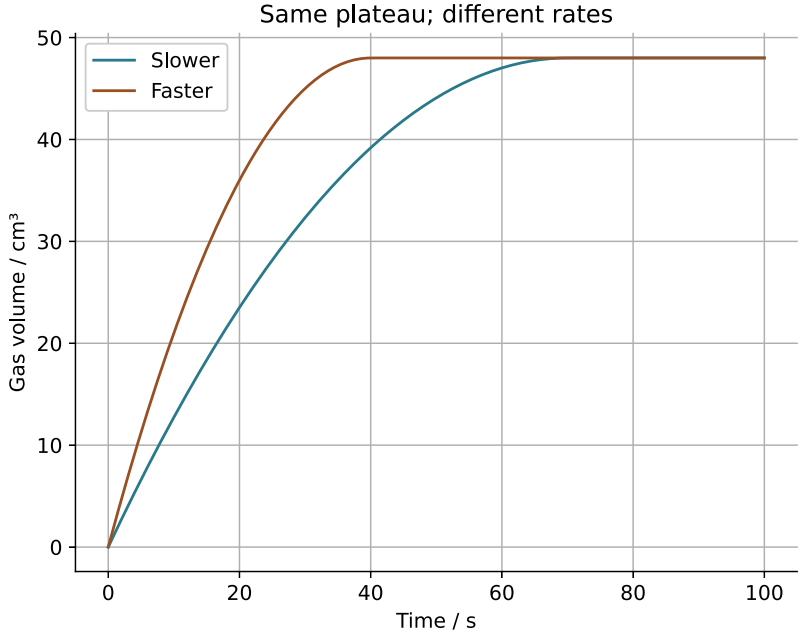
Constant amounts do not mean particles are motionless or reaction has stopped. A steady open-flow system can also have constant measured amounts, so constancy alone does not establish closed-system equilibrium. Equal mass, concentration and molecule count are not the required equality: rates are. Condition-change predictions are separately Higher-only.

Mean rates: use the change over the chosen interval

Official-unit focus: 4.6 Rate and extent of chemical change

Two reactions can produce the same final gas volume but reach it at different times. The final amount and the speed of formation answer different questions.

Mean rate is quantity of reactant used divided by time taken, or quantity of product formed divided by time taken. Common-tier quantities include mass in grams and gas volume in cm^3, giving g per second or cm^3 per second. If a graph already shows an accumulated quantity, use the change between the two stated times, not the final reading alone. A product curve normally rises and becomes less steep as reactants are used up.



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Plot time on the horizontal axis and measured quantity on the vertical axis with labelled units and sensible scales. A steeper product-time curve means faster formation; a horizontal plateau means no further measured product forms. A tangent touches the local curve direction at the chosen point and its steepness indicates rate there. Drawing and interpreting tangents is both-tier; calculating their numerical gradients is in the Higher lesson. A straight chord between distant points gives an interval mean, not the local rate.

Collect readings at regular intervals with a consistent start. Check whether the graph is product formed, reactant used up or reactant remaining. Remaining reactant falls; its negative slope represents a positive rate of consumption when expressed as the amount used. Compare curves at the same time and under stated conditions. Do not assume a plateau proves every reactant is exhausted: a limiting reactant, incomplete collection or another stopped process needs consideration.

■ Checked worked case

Known: gas readings are 12 cm^3 at 20 s and 36 cm^3 at 60 s. Quantity formed during that interval=36−12=24 cm^3; time=60−20=40 s. Mean rate=24/40=0.60 cm^3 per second. A separate mass-loss example records 0.80 g escaping gas over 40 s, giving 0.020 g per second if that gas loss is the measured reaction quantity.

■ Common error

Do not divide 36 by 40 for the interval calculation or call a final volume a rate. A tangent line is a graphical model of the local direction, not a new experimental trace. Mol/s calculations are Higher-only. Curves with the same plateau can have different rates, and a larger plateau alone does not prove a faster initial reaction.

Higher Tier: calculate tangent gradients and molar rates

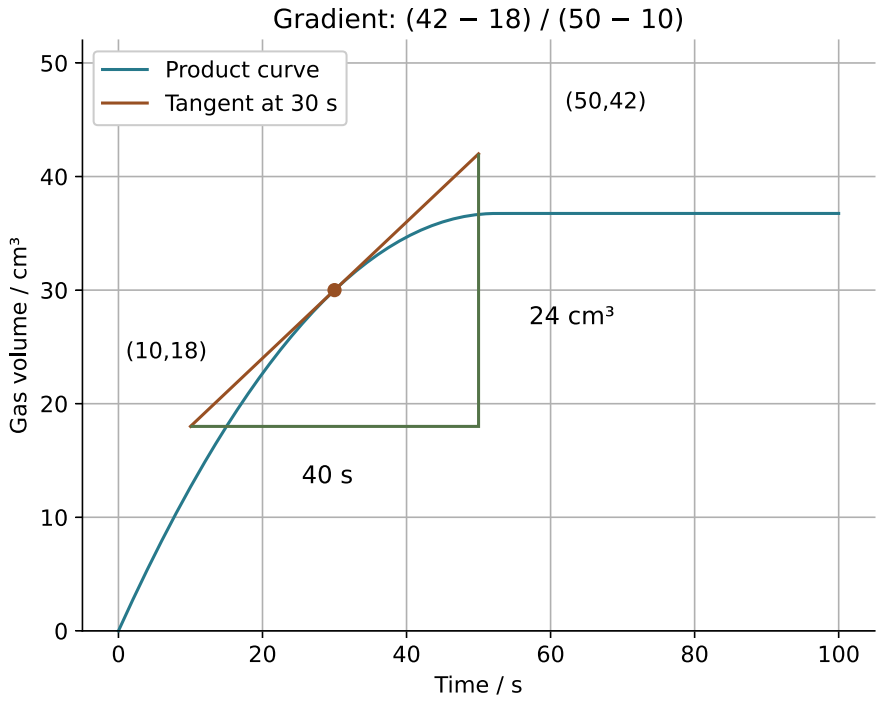
Higher Tier focus; see section scope notes.

Official-unit focus: 4.6 Rate and extent of chemical change

A curve can flatten during a reaction. Its mean rate over a long interval can therefore differ from its rate at one particular instant.

An instantaneous rate is estimated by the gradient of a tangent to the quantity–time curve at the selected time. Gradient=vertical change/horizontal change using two well-separated points on the tangent itself. For a product-volume graph this gives cm^3 per second. Higher-tier quantity can also be in moles, giving mol per second. The units come from the actual axes, not from a remembered numerical example.

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First mark the requested time, draw a tangent following the smooth curve's local direction, then select two readable points on that line. They need not be measured data points or lie on the curve elsewhere. A larger triangle reduces the relative effect of reading uncertainty. Tangent estimates can differ slightly between reasonable drawings. A chord linking two curve points measures an interval mean and must not be presented as the tangent gradient.

Show each coordinate pair, differences, division and final unit. For reactant remaining, the tangent gradient is negative; report its positive magnitude when asked for rate of reactant consumption. Convert time or amount units before dividing: milliseconds and seconds are not interchangeable, nor are millimoles and moles. This numerical tangent calculation and the molar-rate requirement are the embedded Higher clauses, while common-tier students still draw and interpret tangent steepness.

■ Checked worked case

Known: points read from a supplied tangent are (10 s,18 cm^3) and (50 s,42 cm^3). Gradient=(42−18)/(50−10)=24/40=0.60 cm^3 per second at the tangent's stated contact time. In a separate amount measurement, 0.012 mol product forms in 30 s; mean molar rate=0.012/30=0.00040 mol per second. This second number is a mean, not automatically an instantaneous rate.

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■ **Common error**

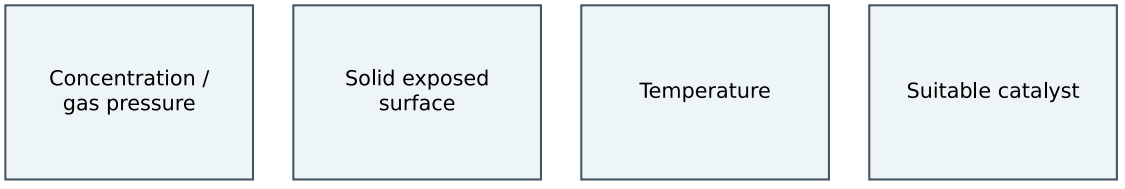
Do not take points from the original curve when a tangent calculation is requested. A negative gradient of reactant remaining does not mean particles react backwards. Do not convert cm^3 to moles without supplied conditions or an appropriate relation. The tangent contact time must be distinguished from the endpoints used for the gradient triangle.

Five rate factors: change the conditions, keep comparisons fair

Official-unit focus: 4.6 Rate and extent of chemical change

Powdered calcium carbonate reacts with acid faster than equal-mass large pieces. More exposed surface can speed the reaction without increasing the carbonate amount.

Reaction rates are affected by concentration of reactants in solution, pressure of reacting gases, surface area of solid reactants, temperature and catalysts. Increasing concentration or gas pressure generally increases collision frequency under the stated comparison. Increasing exposed solid surface gives more sites for collision. Raising temperature increases collision frequency and the fraction with sufficient energy. A suitable catalyst supplies a lower-activation-energy pathway.



State exactly what changes and what is controlled. For equal masses of the same carbonate with acid in excess, smaller pieces can reach the same final carbon dioxide amount sooner. If increasing acid concentration also changes the limiting reactant amount, both speed and final amount can change; this is a different comparison. Higher gas pressure is relevant to reacting gases, not automatically a general explanation for every liquid reaction.

Choose one independent variable and measure a defensible rate response: initial curve steepness, quantity formed in a fixed time, or time to a defined comparable endpoint. Keep other factors controlled, including temperature, total liquid volume, solid mass and exposed surface where appropriate. Repeat measurements and retain observations. A catalyst must be suitable for that reaction; a biological enzyme is not a universal

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catalyst for unrelated chemical changes.

■ Checked worked case

Known: an original comparison reaches a fixed 30 cm^3 gas endpoint in 40 s for larger pieces and 20 s for smaller pieces. Endpoint mean rates are 30/40=0.75 and 30/20=1.50 cm^3 per second, a twofold ratio. This does not prove that the instantaneous rate was exactly doubled at every time or that the smaller pieces yield twice as much gas overall.

■ Common error

Do not confuse concentration with total volume or pressure with temperature. Cutting a solid raises exposed area without changing the identity of its atoms. A hotter reaction can still have the same limiting-reactant yield. This source section is both-tier; detailed equilibrium shifts are separate Higher content.

Required practical 5: test concentration with gas-volume curves

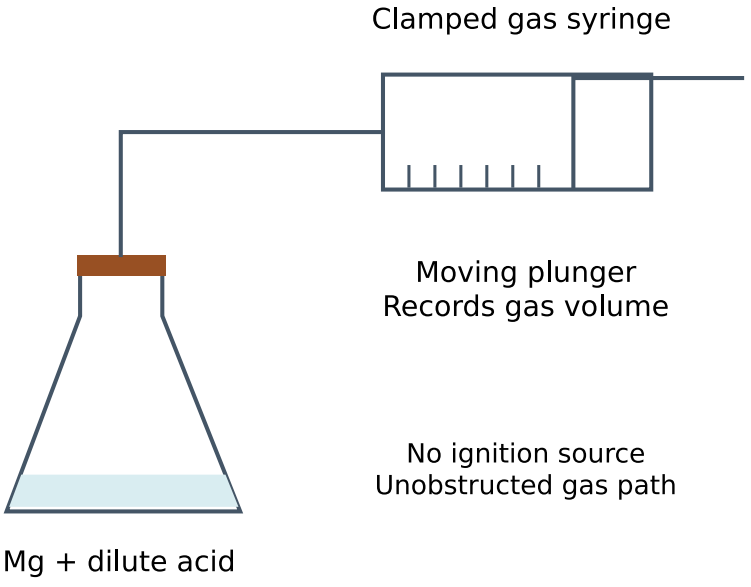
Official-unit focus: 4.6 Rate and extent of chemical change

More concentrated acid can react faster with magnesium. A gas-volume curve tests that prediction, but a delayed bung or a leak can lose the early hydrogen.

RP5 requires a concentration investigation using both gas-volume measurement and colour or turbidity change. In the gas activity, equal cleaned 3 cm magnesium ribbon lengths react with 50 cm^3 of teacher-prepared dilute hydrochloric acid of two concentrations. Hydrogen is collected through a delivery tube into an inverted water-filled measuring cylinder in a trough. The handbook permits a gas syringe as an alternative. Develop a hypothesis before measuring.

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RP5: permitted syringe alternative



Fit the collector and check for unobstructed movement and an open gas path into the collector. Add magnesium, replace the bung promptly and start timing consistently. Record gas volume at suitable intervals, for example every 10 s, until it becomes constant. Plot both volume–time curves on the same axes and compare steepness at comparable times. Identical final volumes require equal limiting magnesium and enough acid; real collection losses or sample differences can change recorded plateaux.

Perform the approved actual school experiment with eye protection, a stable clamped collector and no ignition sources near hydrogen. Use matched ribbon mass/length and consistent oxide removal, acid volume, initial temperature and apparatus. Never seal gas-producing apparatus with no outlet. Repeat trials where feasible and compare methods. A syringe avoids some water-collection problems but can stick or leak; a water collector needs consistent reading conditions and gas not excessively soluble in water.

■ Checked worked case

Known: original illustrative readings give 18 cm^3 after 20 s in the lower-concentration trial and 30 cm^3 after 20 s in the higher one, both starting at zero. Endpoint mean rates are 0.90 and 1.50 cm^3 per second. The latter is faster over that interval; the ratio is $1.50/0.90 \approx 1.67$. These fictional values are not the handbook's technician data or a rule that rate exactly tracks concentration in every reaction.

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■ Common error

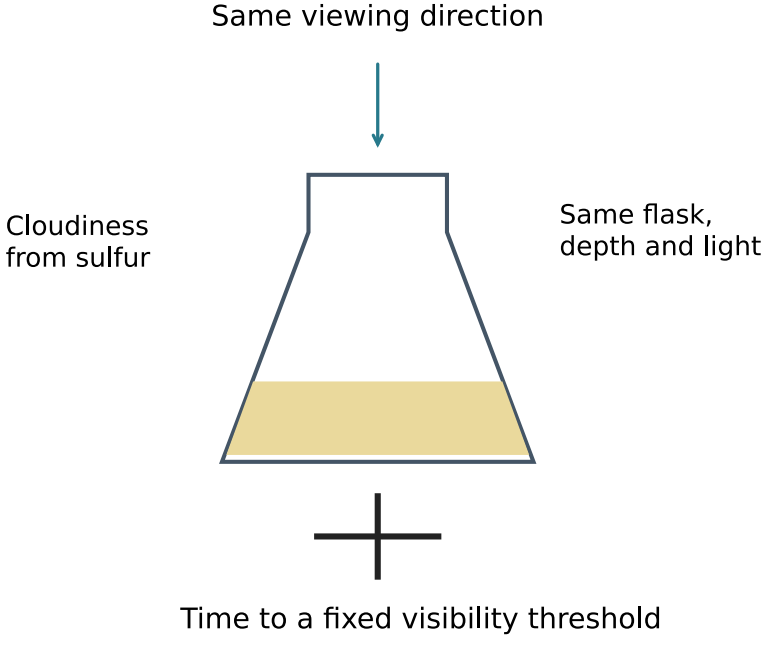
A written graph answer cannot certify practical participation. Do not ignore gas lost before fitting the bung, interpret a smaller collected plateau automatically as slower chemistry, or use unequal magnesium pieces as a concentration-only test. The molar concentration labels identify provided solutions; calculating a mol/s rate remains Higher-only.

Required practical 5: compare a defined disappearing-cross endpoint

Official-unit focus: 4.6 Rate and extent of chemical change

Sulfur particles make an acid–thiosulfate mixture cloudy until a black cross is no longer visible. The endpoint is an observation threshold, not the instant all reactants have been used up.

The RP5 turbidity activity dilutes 40 g/dm^3 sodium thiosulfate stock with water to a constant 50 cm^3 before adding the same 10 cm^3 dilute hydrochloric acid. Use stock/water volumes 10/40, 20/30, 30/20, 40/10 and 50/0 cm^3, giving pre-acid concentrations 8,16,24,32 and 40 g/dm^3. Put the same flask on a printed black cross, add acid, swirl consistently and time until the cross is no longer visible.



The concentration labels refer to the 50 cm^3 pre-acid dilution; adding acid further dilutes the mixture consistently. Maintain total volumes, flask, viewing depth, lighting, temperature and endpoint judgement. Repeat and calculate mean endpoint time. For a comparable fixed turbidity threshold, 1/mean time is a relative rate proxy: a shorter

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time means a larger proxy. It does not measure a known mass of sulfur per second or the exact completion time of the entire reaction.

Carry out only the school-approved method with ventilation and eye protection. Sulfur dioxide is released and can aggravate breathing difficulties; do not lean over the vessel or deliberately inhale fumes. The teacher selects controls and an appropriate alternative observation arrangement for affected students. Record raw times and reasons for any justified anomalous-result exclusion; do not discard values merely because they spoil a trend. Compare repeats within a method before claiming reproducibility across groups.

■ Checked worked case

Known: original repeat endpoint times are 48,50 and 52 s. Mean time=(48+50+52)/3=50 s; relative proxy=1/50=0.020 per second. A second concentration gives mean 25 s and proxy 0.040 per second, twice as large under the same observation threshold. This does not mean sulfur production was exactly constant within either trial.

■ Common error

A disappearing cross is not a direct gas-volume reading, and one method alone does not fulfil both RP5 approaches. Do not dilute stock by reducing solution volume without replacing water, because depth and total volume would also change. A subjective endpoint needs consistent judgement; repeat averaging cannot correct a systematic viewing difference.

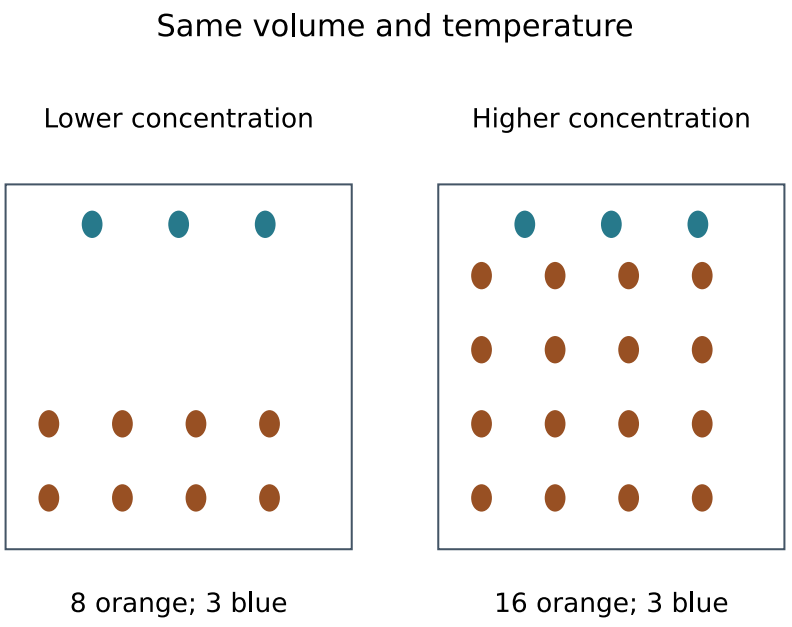
Collision theory: concentration and reacting-gas pressure

Official-unit focus: 4.6 Rate and extent of chemical change

Putting more reacting particles into the same volume makes encounters more frequent. It does not automatically give each particle more energy.

Collision theory requires reacting particles to collide with each other and have sufficient energy. Activation energy is the minimum particle energy needed for reaction. Increasing reactant concentration in a solution places more reacting particles in a given volume, so collisions between reactants happen more often. Compressing reacting gases at the same temperature raises pressure and particle concentration, increasing collision frequency. More frequent successful collisions generally increase rate.

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Keep frequency distinct from collision energy. At unchanged temperature, raising concentration alone does not mean particles move faster or each has more energy. In a simple stated model with one collision partner held fixed, doubling the other partner's particle concentration can double encounter opportunities. It is not a universal measured rate law for every mechanism or for simultaneous changes to both reactants. Different reactions require suitable evidence for exact proportional relationships.

Compare same-volume particle drawings using equal particle sizes and the same stated temperature. Count only the labelled reacting species, not solvent particles as if they were all reactants. For gas compression specify that temperature is controlled; actual compression can otherwise change temperature too. School experiments use approved dilute solutions and supplied gas data rather than student-built pressure vessels. Translate a rate claim into collision frequency and sufficient-energy reasoning.

■ Checked worked case

Known: a supplied same-volume model increases reacting particles from 20 to 40 while the other partner and temperature stay fixed. Particle concentration doubles, consistent with twice as many encounter opportunities in the stipulated model. Compressing a fixed gas sample from 200 to 100 cm^3 at the same temperature similarly doubles particles per unit volume. These ratios do not provide an exact collision count from a still picture.

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■ **Common error**

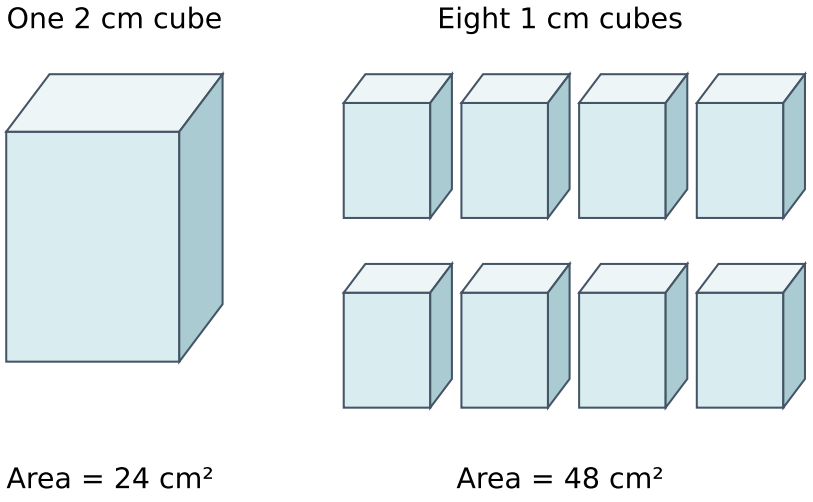
Do not say greater concentration lowers activation energy or that any collision necessarily reacts. Pressure reasoning concerns gas reactants. Solvent volume changes, temperature drift and changed chemical species can invalidate a simple model comparison. The model explains a direction of change; an exact numerical rate ratio needs its stated assumptions or measured evidence.

Temperature and solid size: explain two different rate effects

Official-unit focus: 4.6 Rate and extent of chemical change

Warming particles and crushing a solid can both speed reactions, but they act differently. Temperature changes particle energies; crushing exposes more contact surface.

At higher temperature particles move faster and collide more frequently. Collisions are more energetic, so a greater fraction has at least the activation energy and can react. Temperature does not lower the activation energy of the unchanged pathway. For a reacting solid, only accessible surface contacts the other reactant. Smaller pieces at the same total solid volume have a greater surface-area-to-volume ratio, exposing more sites and increasing collision frequency.



A cube model makes the size comparison explicit. A cube side L has area $6L^2$ and volume L^3. Cutting it into eight half-side cubes preserves total volume but doubles total area. This model assumes newly exposed faces are accessible and not clumped or coated. Real powders need appropriate containment; aggregate formation can reduce accessible area. Do not claim every ten-degree temperature rise always doubles rate:

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exact effects depend on reaction and conditions.

Investigate one factor at a time with teacher-approved dilute reactants and moderate controlled temperatures. Use a water bath where appropriate and measure the reacting mixture's temperature, not only the bath label. Compare equal solid masses with the same composition and control acid concentration and volume. Fine powders can create exposure hazards; use the approved particle sizes rather than grinding unknown materials independently. Explain the measured trend through the relevant collision mechanism.

■ Checked worked case

Known: an original model cube of side 2 cm has volume 8 cm^3 and area 24 cm^2. Eight cubes of side 1 cm have total volume 8 cm^3 and total area $8\times6=48$ cm^2, twice the original area. For this accessible-face model the increased contact area supports faster reaction with the same solid amount; it does not alone predict an exact measured doubling of rate.

■ Common error

Do not say hot particles become physically larger, or that smaller pieces have intrinsically lower activation energy. Surface area and total amount are different. Raising temperature changes the energy distribution; cutting changes geometry. Both can increase successful collision frequency, but explanations should not swap their causes.

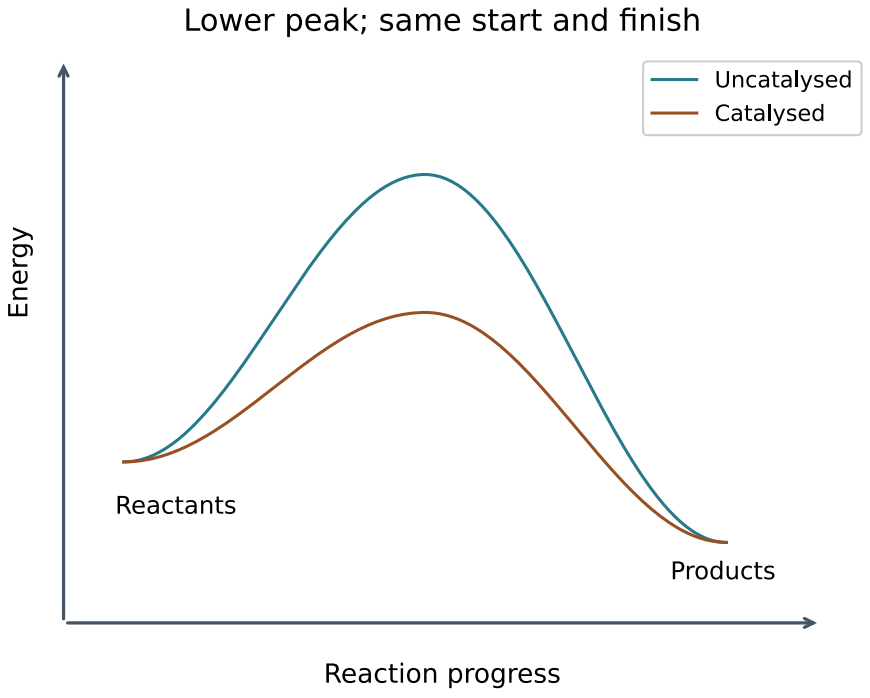
Catalysts: a lower barrier, the same overall energy change

Official-unit focus: 4.6 Rate and extent of chemical change

A catalyst lets a reaction proceed faster without appearing as a consumed reactant in the overall equation. It changes the route, not the identities of the starting and finishing substances.

Catalysts increase reaction rate by providing a different reaction pathway with a lower activation energy. They are not used up overall, although they may participate in steps and be regenerated. Different reactions need different catalysts; enzymes act as catalysts in biological systems. On a reaction profile, the catalysed curve has a lower peak but the same reactant and product levels as the uncatalysed curve. The overall energy change is therefore unchanged.

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At the same temperature more colliding particles can meet the smaller activation requirement. This does not mean the catalyst adds energy to every particle or increases temperature by definition. Evidence for catalytic behaviour combines rate increase and regeneration/no net consumption. An unchanged small sample alone does not prove catalysis unless the reaction-rate effect is shown. The overall equation omits a regenerated catalyst from its stoichiometric reactant and product lists.

Compare teacher-approved catalysed and uncatalysed trials with matching reactant quantities, concentrations, temperature and collection apparatus. Record time-dependent product measurements rather than only final volume. Use the specified catalysts in the course where required; do not memorise unrelated industrial names. Actual hydrogen peroxide catalyst demonstrations need the school's selected dilute solutions and protection because rapid gas evolution and warming can occur.

■ Checked worked case

Known: a supplied profile has reactants at 20, products at 5 and uncatalysed peak at 80 relative units. A catalysed peak is 50. Both paths decrease overall by 15 units; the barriers are visibly different. A fixed endpoint reached in 60 s without catalyst and 20 s with it gives a threefold ratio of mean rates to that endpoint, assuming the same measured amount.

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■ Common error

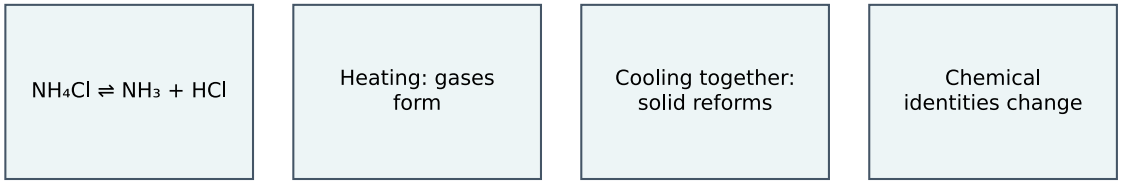
Do not lower the product energy when drawing the catalysed path or describe the catalyst as consumed in the overall reaction. A catalyst does not increase the fixed limiting-reactant amount of product. At equilibrium it speeds both directions without changing the final composition under fixed conditions; this statement does not require equilibrium-constant calculations.

Reversible reactions: the products can reform the reactants

Official-unit focus: 4.6 Rate and extent of chemical change

Heating ammonium chloride can produce ammonia and hydrogen chloride. Cooling the gases together can reform the original ammonium chloride: both directions are chemical reactions.

In a reversible reaction, products can react to produce the original reactants. Write A + B ⇌ C + D with two opposed half arrows. Conditions can favour a direction. The specified example is ammonium chloride ⇌ ammonia + hydrogen chloride: heating favours decomposition, while cooling the gases together reforms ammonium chloride. Each direction forms substances with different chemical identities from its starting materials.



The symbol does not mean that all substances must be present in equal quantities or that the reaction continually switches completely from one side to the other. A reversible reaction may reach equilibrium when substances are contained under suitable fixed conditions, but reversibility alone is not proof that an open apparatus has reached equilibrium. Melting/freezing is a reversible physical change; it does not by itself demonstrate this chemical product-to-reactant relationship.

Use supplied observations or a teacher demonstration under the approved risk assessment. Ammonia and hydrogen chloride gases are irritant; students must not generate or smell them independently. Record where gas forms and where the white solid reappears, distinguish observation from chemical explanation and retain the condition labels. Do not improvise a sealed heated vessel. Write a balanced equation with

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suitable state symbols only for the stated conditions.

■ Checked worked case

Known: $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$ conserves one N, four H and one Cl atom on each side. One represented ammonium chloride formula unit gives one ammonia and one hydrogen chloride molecule. Six decomposed formula units give six of each gas, twelve gas molecules total in this particle model. Cooling permits those pairs to reform six formula units.

■ Common error

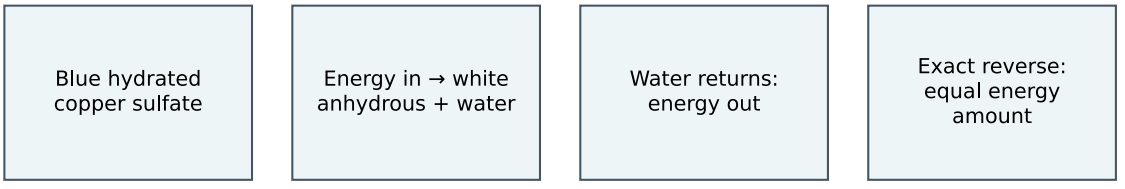
The gases do not need to have the same mass merely because their molecule counts match. A reversible arrow does not give equal rates at every moment. White solid forming on cooling must be interpreted with the identified gases; colour alone cannot establish the reaction. Written interpretation does not replace safe supervised observations.

Reverse reactions: energy transfers reverse direction

Official-unit focus: 4.6 Rate and extent of chemical change

Removing water from blue hydrated copper sulfate requires energy. Returning water to the white anhydrous material reverses the chemical change and releases energy.

If a reversible chemical reaction is exothermic in one direction, it is endothermic in the opposite direction. The same energy amount is transferred for the exact reversed change under matching conditions and quantities. In the specified example, hydrated copper sulfate (blue) changes to anhydrous copper sulfate (white) and water by an endothermic change; rehydration is exothermic. Labels should identify direction as well as whether energy enters or leaves.



Products with greater energy in the forward endothermic reaction become the higher-energy starting materials for the reverse exothermic reaction. Reversing the equation swaps the levels. The equal transferred amounts do not mean equal activation energies

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for the two directions: the climb to the same peak starts from different levels. That distinction can be read qualitatively without introducing an advanced thermodynamic calculation.

Use teacher-approved demonstration samples or supplied before/after evidence. Copper salts need controlled handling and waste collection; adding water to hot material or heating unknown compounds is inappropriate. Record colour alongside the identified substances and temperature evidence. Water loss from a hydrate is a specific chemical change, not a rule that every blue substance becomes white on heating or that every colour change is reversible.

■ Checked worked case

Known: in a fictional level model hydrated starting materials lie at 10 relative units and the dehydrated products at 30. The forward increase is 20 units; the exact reverse decreases by 20. In an illustrative event count, eight forward changes transfer eight equal energy packets into the chemical system; eight reverse changes transfer eight packets out. This is conservation reasoning, not solution calorimetry or a measured molar enthalpy.

■ Common error

Compare the same reaction amount and conditions before claiming equal transferred amounts. A temperature rise in one wet sample and a smaller temperature fall in a different heated sample are not directly comparable energy measurements. Do not confuse total energy change with activation barrier or use the example to calculate an unrequired heat-capacity quantity.

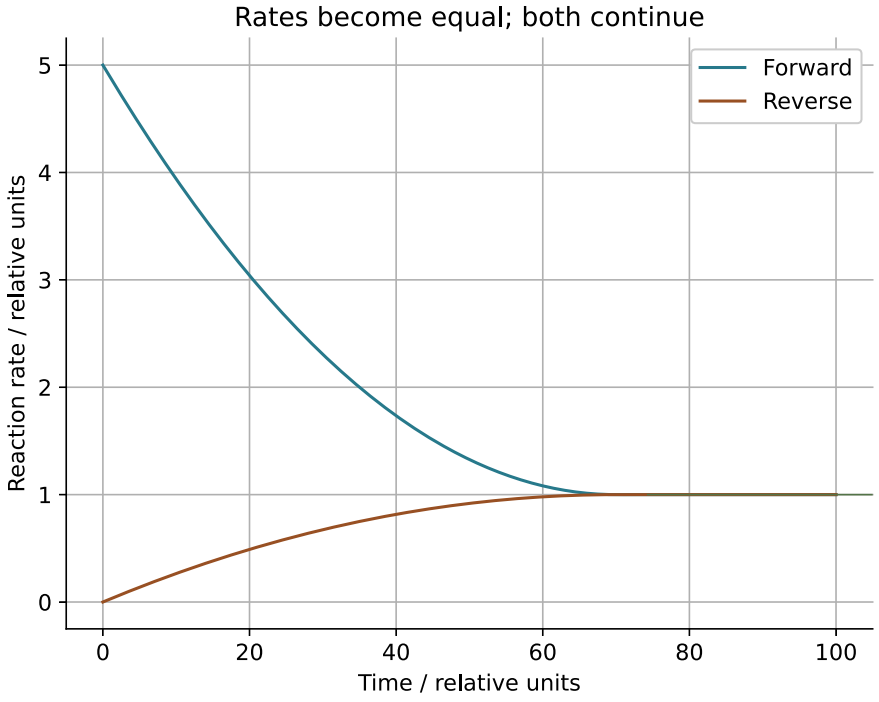
Dynamic equilibrium: equal rates, continuing reactions

Official-unit focus: 4.6 Rate and extent of chemical change

A closed reaction vessel can show constant composition while particles still react in both directions. Equality of rates keeps the overall amounts unchanged.

When a reversible reaction occurs in apparatus preventing escape of reactants and products, equilibrium is reached when forward and reverse reactions occur at exactly the same rate. The state is dynamic: both directions continue. At fixed conditions the amounts of reactants and products stay constant, but they need not be equal. A closed system prevents loss of matter; it does not necessarily prevent energy transfer to maintain a set temperature.

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Beginning with mostly reactants, the forward rate may be high while the reverse rate is initially low. As products accumulate, reverse reaction becomes possible more often until the rates match under the stated model. A reaction-rate graph should show convergence to the same non-zero rate. A composition graph can level off at different reactant and product amounts. Identify which quantity is on the axis before claiming two curves must meet.

Use an original particle-count model or supplied data with fixed conditions and no matter lost. Count forward and reverse events over equal time windows. Their difference gives net change for a one-to-one A $\rightleftharpoons$ B model. Do not build a sealed heated chemical apparatus to demonstrate the idea; safe school demonstration arrangements require approved containment and temperature control. A count model illustrates balance without proving a particular real chemical rate law.

■ Checked worked case

Known: in a closed A $\rightleftharpoons$ B model, 18 forward and 18 reverse events occur in one second. Net B change=18−18=0 even though 36 reaction events occur. If 70 A and 30 B particles remain on average, these unequal amounts are compatible with equilibrium. At an earlier interval with 20 forward and 12 reverse events, net B increases by eight and the model is not yet at equilibrium.

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■ Common error

Constant amounts do not mean particles are motionless or reaction has stopped. A steady open-flow system can also have constant measured amounts, so constancy alone does not establish closed-system equilibrium. Equal mass, concentration and molecule count are not the required equality: rates are. Condition-change predictions are separately Higher-only.

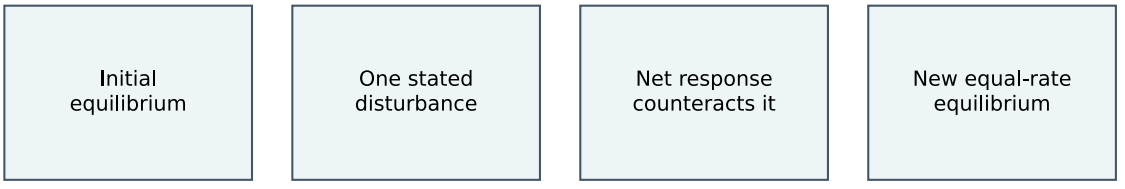
Higher Tier: predict the response to an equilibrium disturbance

Higher Tier focus; see section scope notes.

Official-unit focus: 4.6 Rate and extent of chemical change

An equilibrium mixture responds when a condition changes. Its response opposes the disturbance, but does not necessarily restore the original amounts exactly.

Equilibrium amounts depend on reaction conditions. Le Chatelier's Principle states that when a system at equilibrium is disturbed, it responds to counteract the change. After a change of concentration, temperature or reacting-gas pressure, the system adjusts until forward and reverse rates are equal again. Predict qualitatively which direction is favoured using the equation and the changed condition. This entire source section is Higher-only.



Begin with the system already at equilibrium, identify one imposed change and select the appropriate response rule. Increased reactant concentration favours its consumption; removing product favours forming more. Warming favours the endothermic direction. Compression of a gaseous equilibrium favours the side with fewer gas molecules. These are composition responses, not the same as saying every faster reaction makes a larger final yield.

Use supplied equations, forward energy direction, states and before/after data. Mark a shift arrow as net adjustment toward a new equilibrium, not cessation of the other reaction. Keep other conditions specified. A suitable catalyst increases both rates and

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reduces time to equilibrium without shifting the final equilibrium composition at fixed conditions. No equilibrium-constant expression or numerical constant calculation is required by this GCSE section.

■ Checked worked case

Known: a supplied one-to-one mixture initially contains 40 A and 60 B particles at equilibrium. After an imposed change and re-equilibration it contains 30 A and 70 B. Net B increase is ten, so the observed response favoured products. Total remains 100 in this stated closed conversion model. These data show a direction of shift but alone do not identify which condition was changed.

■ Common error

Do not describe counteracting as entirely cancelling the original disturbance. A catalyst does not count as a concentration addition of reactant. A shift toward products does not mean all reactants are consumed, and equal rates at the new equilibrium do not require the old mixture ratio. Different perturbations can produce the same direction of observed composition change.

Higher Tier: add a reactant or remove a product

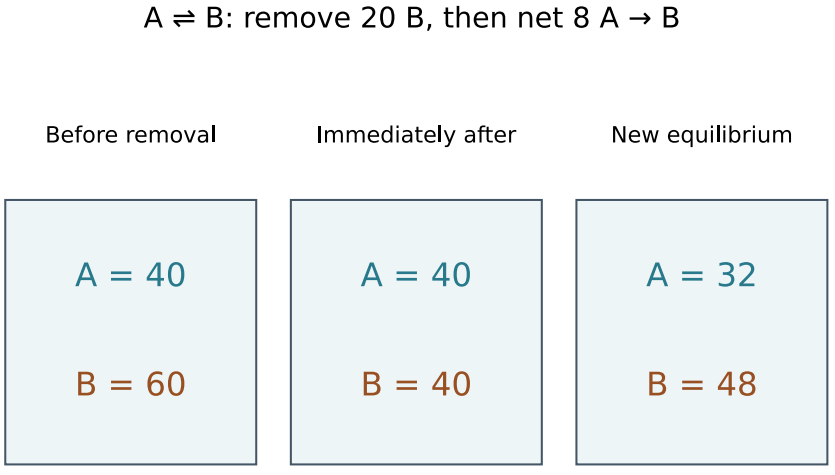
Higher Tier focus; see section scope notes.

Official-unit focus: 4.6 Rate and extent of chemical change

Removing a product can draw a reversible system toward forming more of it. The response partly replaces what was removed; it does not imply product concentration never fell.

Changing a reactant or product concentration disturbs an equilibrium. Increasing a reactant concentration favours a net forward reaction, consuming some added reactant and forming more products until equilibrium returns. Decreasing a product concentration also favours forward reaction. Increasing a product favours the reverse reaction, while removing a reactant favours reverse formation of that reactant. Use the actual balanced equation to identify each side.

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Separate the immediate imposed change from the response. If product is removed, its amount first falls; a subsequent net forward change can raise it from that lowered value without necessarily restoring the original level. At the new equilibrium both directions again occur at equal rates. A plot may show a sudden jump in the changed species followed by slower adjustment; the other species need not have the same immediate jump.

Use original fictional particle or concentration data with a stated fixed temperature and appropriate mixture model. Identify the species directly changed, then predict which direction uses it or replaces it. Real solution mixing can change volume as well as amount, so concentration comparisons need that information. The simple one-to-one particle model is an explanation aid; it is not a numerical equilibrium-constant method.

■ Checked worked case

Known: A ⇌ B starts with 40 A and 60 B particles. Removing 20 B immediately leaves 40 A and 40 B. The subsequent response converts a net eight A to B, producing 32 A and 48 B at the supplied new equilibrium. B rose by eight after removal but remains twelve below its original 60. Matter removal was the imposed disturbance; subsequent conversion conserves the remaining total of 80.

■ Common error

Do not claim the product-removal operation itself immediately increases product concentration. A forward shift need not mean the final product amount exceeds its pre-removal value. Adding an unrelated substance is not automatically the same as adding a reactant. This whole section is Higher-only and requires qualitative interpretation, not memorised K calculations.

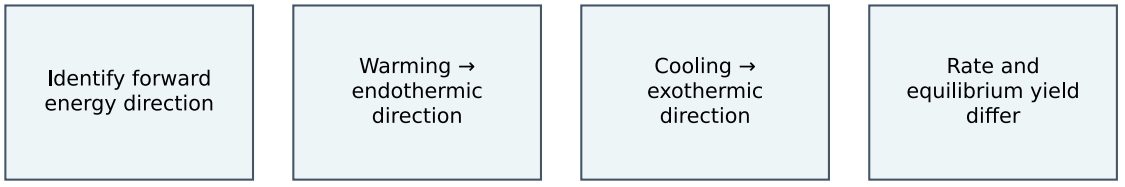
Higher Tier: warming favours the endothermic direction

Higher Tier focus; see section scope notes.

Official-unit focus: 4.6 Rate and extent of chemical change

Heating can make reactions faster yet reduce the equilibrium amount of a desired product. Rate and final equilibrium composition need separate explanations.

Increasing temperature favours the endothermic direction, which takes in energy and counteracts warming. Decreasing temperature favours the exothermic direction, which releases energy. If the written forward reaction is endothermic, warming increases the relative product amount and cooling decreases it. If forward is exothermic, warming decreases relative products and cooling increases them. Always state which direction the given energy label describes.



Both forward and reverse reactions can become faster on warming, yet their relative rate balance changes, causing a net shift before a new equilibrium is reached. For an exothermic forward industrial reaction, a low temperature may favour yield while slowing production; an appropriate temperature choice can be a compromise. A catalyst can help rate without turning that exothermic forward direction into an endothermic one or increasing its equilibrium yield at fixed conditions.

Read a supplied equation and mark forward endothermic/exothermic before writing a shift conclusion. Use data collected after equilibrium at each stated temperature, with other relevant conditions specified. A temporary product increase during heating is not necessarily the final equilibrium composition. This lesson predicts direction and

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evaluates supplied evidence; no numerical equilibrium constant or thermodynamic derivation is required.

■ Checked worked case

Known: in a fictional exothermic A $\rightleftharpoons$ B mixture at equilibrium, the supplied product fraction is 70% at a lower temperature and 55% at a higher one. Warming favours the reverse endothermic direction, consistent with the 15 percentage-point decrease. In a stated total of 200 represented particles, product counts are 140 and 110; the decrease is 30. These numbers are illustrative, not actual chemical yield data.

■ Common error

Do not write heating always increases equilibrium yield. The term exothermic must attach to a direction, not to the reversible pair without qualification. Percentage-point change is not the same as percentage change relative to the original value. A faster approach to equilibrium does not prove a more product-rich equilibrium.

Higher Tier: count gaseous coefficients before predicting a shift

Higher Tier focus; see section scope notes.

Official-unit focus: 4.6 Rate and extent of chemical change

Compressing an equilibrium gas mixture favours the side with fewer gaseous particles. Count coefficients from the balanced equation before naming the favoured product.

For gaseous equilibria, increasing pressure by compression at fixed temperature shifts toward the side with fewer gas molecules in the balanced equation. Decreasing pressure by expansion favours the side with more gas molecules. Count gaseous coefficients, not every atom inside each formula. For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, the left side has four gas molecules to the right side's two, so compression favours ammonia.

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$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$



Reactants

4 gas molecules

Products

2 gas molecules

Compression at fixed T → fewer gas molecules

Expansion → more gas molecules; count coefficients

If the gas-molecule counts are equal on both sides, changing pressure by compression gives no pressure-driven composition shift in this model. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ has two on each side. If solids or liquids occur in a supplied equation, do not count their coefficients as gas molecules. Compression can also affect rates by increasing gas particle concentration; a rate effect alone is not proof of an equilibrium shift.

Write state symbols and a gas-count total under each side. Specify compression or expansion of the reacting mixture at constant temperature: merely adding an unrelated gas does not justify blindly applying the same rule without a volume/pressure description. Evaluate a supplied industrial choice using yield, rate, equipment cost and operating constraints. Do not ask students to construct high-pressure reaction vessels.

■ Checked worked case

Known: in $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, gas coefficients total three on the left and two on the right. Compression favours the product side; expansion favours reactants. One thousand represented stoichiometric forward sets convert 3,000 gas molecules into 2,000, a decrease of 1,000 in this counting model. The equation count predicts shift direction, not a specific percentage conversion.

■ Common error

Do not count oxygen atoms to decide the gas-count rule. A higher pressure does not always favour products; it favours the smaller gaseous count. Equal counts mean no pressure-driven shift under the stipulated compression model, not that rate can never change. This entire source section is Higher-only.

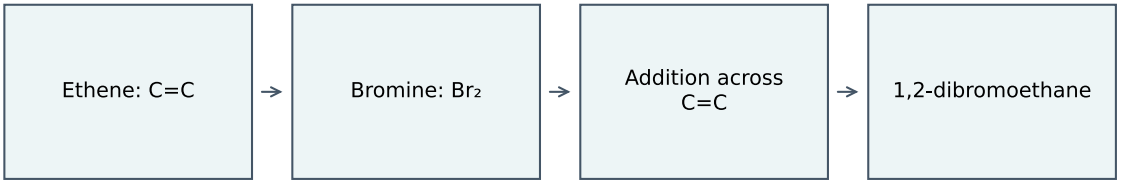
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Organic structures and reaction pathways

Official-unit focus: 4.7 Organic chemistry

Two compounds can have the same molecular formula but different structures. Their functional groups help predict which reactions they undergo.

A homologous series shares a functional group and general formula. Structural isomers share a molecular formula but differ in atom connections. Alkenes contain a carbon-carbon double bond.



Distinguish addition, substitution, oxidation and polymerization by tracing bonds before and after reaction. Conditions and reagents belong to the reaction arrow; they are not interchangeable labels.

Draw displayed formulae with correct carbon and hydrogen valencies. For the specified alkene additions, keep the carbon skeleton and track the atoms added across C=C. Compare the supplied structure before and after reaction. GCSE 8462 does not require an advanced multistep synthesis route in this focus.

■ Checked worked case

Known: ethene adds bromine across its double bond. The two-carbon skeleton stays intact and each carbon gains one bromine atom, giving 1,2-dibromoethane. One mole of ethene reacts with one mole of bromine in this addition reaction.

■ Common error

Bromine decolourization provides evidence of unsaturation in an appropriate test. It is not proof that an unknown sample is specifically ethene.

Crude oil: a finite mixture, not one compound

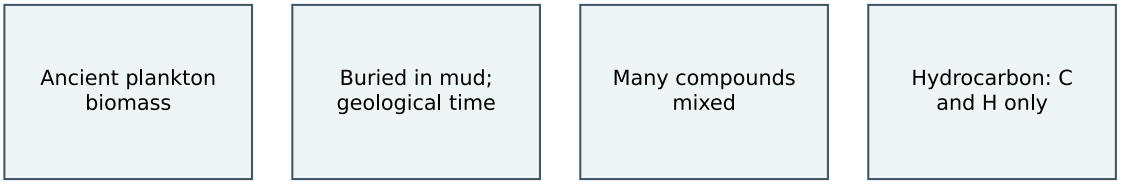
Official-unit focus: 4.7 Organic chemistry

A refinery receives crude oil containing many compounds. It cannot assign the whole liquid one molecular formula, because a mixture has no single fixed chemical compo-

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sition.

Crude oil is a finite resource found in rocks. It formed from ancient biomass, mainly plankton buried in mud, over geological time. Its rate of formation is far too slow to replace present consumption on a human timescale. Crude oil contains a very large number of compounds; most are hydrocarbons. A hydrocarbon molecule contains only hydrogen and carbon atoms. Most hydrocarbons in crude oil belong to the alkane family.



A mixture contains different substances together without a fixed bonding ratio between those substances. Each individual compound still has its own definite formula. C_3H_8 is a hydrocarbon; C_2H_5OH contains oxygen and therefore is not a hydrocarbon, even though it contains carbon and hydrogen. Carbon dioxide is not a hydrocarbon either. Finite describes a limited resource, not a prediction that every deposit will disappear on one exact date.

Read every element symbol before classifying a formula. Then distinguish statements about an individual molecule from statements about the whole mixture. Use supplied refinery composition data to discuss which components are present and their proportions. A classroom model can use labelled molecular cards; it does not require obtaining or heating crude oil. Keep origin, composition and use as separate parts of an explanation.

■ Checked worked case

Known: an original sample model contains 20 C_3H_8 molecules, 10 C_2H_6 molecules and 5 oxygen-containing molecules. There are 35 molecules represented, of which 30 are hydrocarbons. Hydrocarbon number percentage = $30/35\times100 = 85.7\%$ to one decimal place. This is a count percentage, not a mass percentage or a measured composition of all crude oils.

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■ Common error

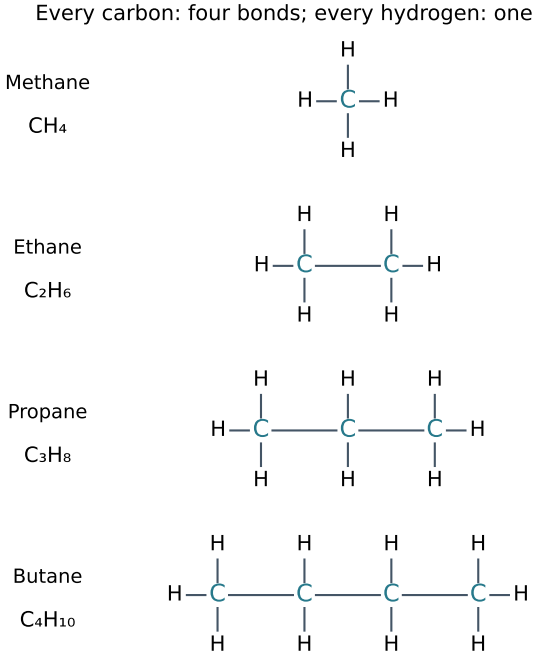
Do not call crude oil pure, renewable on a human timescale, or a single alkane. Biomass origin does not make every biomass-derived fuel a fossil fuel. A substance containing carbon is not automatically a hydrocarbon: all of its elements must satisfy the definition. Models and percentages describe only the supplied sample.

Alkanes: connect the formula to the displayed bonds

Official-unit focus: 4.7 Organic chemistry

Ethane is written C_2H_6, but its displayed formula also shows where the six hydrogen atoms attach. Both representations describe the same molecule at different levels of detail.

A homologous series is a family of compounds with the same functional group and general formula and similar chemical reactions. The general formula of the alkane homologous series is CnH2n+2. For the first four members, methane is CH_4, ethane C_2H_6, propane C_3H_8 and butane C_4H_{10}. Successive members differ by CH_2. A displayed formula shows each atom and bond; carbon makes four bonds and hydrogen one. In these alkanes all carbon–carbon bonds are single. Their molecules are saturated.



For a straight chain, each end carbon has three hydrogen atoms, while an internal carbon has two. Methane has no carbon neighbour and so has four hydrogens. Count the bond lines at each carbon to check a displayed structure rather than adding hydrogen randomly. The general formula identifies the expected molecular composition of this

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series; molecular composition alone need not establish a unique arrangement of atoms.

Construct the first four molecules using a model kit or the original displayed diagrams. Count carbons, total hydrogens and bonds at each atom. Translate between displayed and molecular formulae, then test a supplied formula against H=2n+2. A question may supply a longer-chain formula without requiring its specific name. Keep the required names limited to the four in the section.

■ Checked worked case

Known: for n=3, expected hydrogen number is H=2n+2=2×3+2=8, giving propane C_3H_8. Its end carbons contribute 3+3 hydrogen atoms and its middle carbon contributes 2, also giving 8. For a supplied C_6 formula, the alkane series predicts 14 hydrogens; its specific name is not needed for this calculation.

■ Common error

Do not confuse the small subscript with the coefficient in an equation. C_2H_4 does not satisfy the alkane formula. A carbon with five ordinary covalent bonds is not a valid displayed alkane here. These flat drawings represent bonding, not actual straight planar molecules, bond angles or relative atomic sizes.

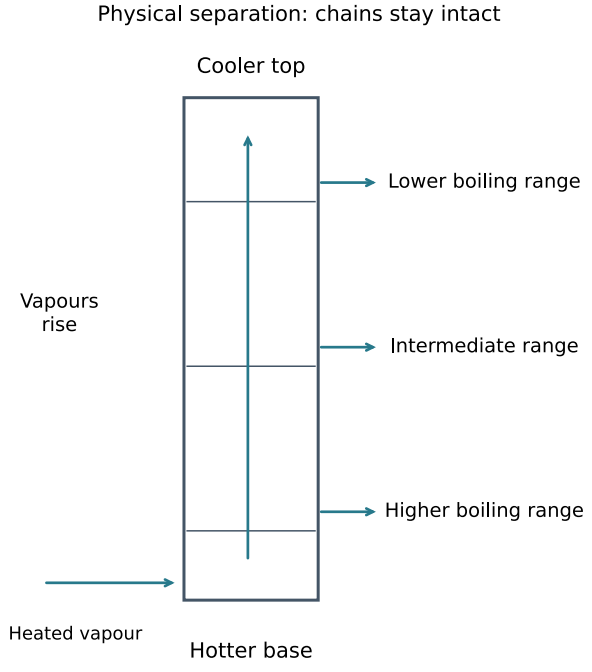
Fractional distillation: separate by evaporation and condensation

Official-unit focus: 4.7 Organic chemistry

A distillation column can collect a fuel fraction containing many compounds. Separating a useful boiling range does not make every molecule in that collected liquid identical.

Crude oil is heated so many components evaporate. Vapour enters a fractionating column that is hotter at the bottom and cooler at the top. Components condense in regions cool enough for them to become liquid. Higher-boiling hydrocarbons condense lower down; lower-boiling ones travel farther up before condensing. Each collected fraction contains molecules with similar carbon numbers and a range of boiling points, rather than one pure substance.

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Evaporation and condensation are physical changes: molecules are separated without breaking their carbon chains. A sufficiently volatile fraction can remain gaseous at the top; heavy material can remain near the base. Fuels include petrol, diesel oil, kerosene, heavy fuel oil and liquefied petroleum gases. Fractions also provide feedstock for the petrochemical industry, producing materials such as solvents, lubricants, polymers and detergents. Carbon atoms can bond to each other in families of related structures; this helps explain the wide variety of natural and synthetic carbon compounds.

Trace a labelled vapour through the temperature-gradient diagram and explain where it condenses using its supplied boiling range. Compare a fraction's intended fuel use with a feedstock use. Do not replace an explanation with a memorised order of names alone; names of other specific fractions are not required. School demonstrations must use an approved substitute apparatus and risk assessment, not a sealed heated crude-oil container.

■ Checked worked case

Known: an illustrative column region is at 180 °C. A component with a supplied boiling point of 250 °C can condense as vapour cools into this region; one with a boiling point of 100 °C can remain gaseous and rise farther. The difference between those supplied boiling points is $\Delta T = 250 - 100 = 150$ °C. These fictional values demonstrate the separation principle, not official fraction boundaries.

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■ Common error

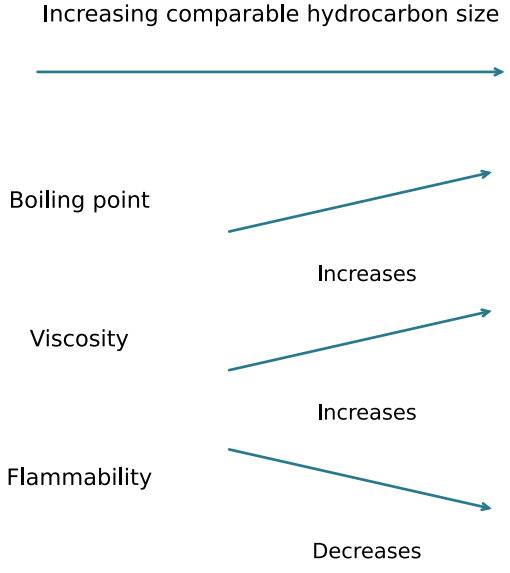
Do not say large molecules rise farther because they are heavier or that the column creates new alkanes. Position depends on boiling behaviour within a temperature gradient. A collected fraction is not necessarily pure, and industrial temperatures vary. Cracking is a later chemical process, not another name for fractional distillation.

Hydrocarbon size: boiling point, viscosity and flammability

Official-unit focus: 4.7 Organic chemistry

A fuel must flow through its supply system and ignite under its intended conditions. A bigger hydrocarbon molecule is not automatically a more suitable fuel for every device.

As hydrocarbon molecular size increases, boiling point generally increases, viscosity increases and flammability decreases. Viscosity describes resistance to flow: a more viscous liquid flows less readily. Flammability describes how readily a material catches fire under comparable conditions. Smaller hydrocarbons are generally more volatile and ignite more readily; larger ones need higher temperatures to boil and are harder to ignite.



Boiling changes intermolecular separation, not chain bonds

Keep the three comparisons separate. Boiling point concerns liquid becoming gas, viscosity concerns flow, and flammability concerns ignition and burning. Stronger intermolecular attractions in larger comparable hydrocarbon molecules help explain higher boiling points. Boiling does not normally break the covalent bonds within the

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molecules. The section limits recalled property trends to these three; density or a precise energy-per-molecule trend is not required here.

Use supplied property data to select a fuel and justify it against a stated purpose. A ranking is meaningful only if temperature and test conditions are comparable. A teacher-supervised investigation can compare flow times or supplied ignition observations using approved small samples and controlled equipment. Students must not ignite unknown fuels independently, and a flame comparison alone cannot establish every property.

■ Checked worked case

Known: in a fictional equal-volume flow test at the same temperature, a smaller-molecule liquid takes 8 s to flow through an opening and a larger-molecule liquid takes 24 s. Flow-time ratio=24/8=3. This supports greater resistance to flow for the second liquid under that method. It does not prove its viscosity is exactly three times as large without a validated relation.

■ Common error

Do not confuse high viscosity with low boiling point or say every fuel's ignition behaviour is controlled by molecular size alone. A longer flow time is evidence tied to the apparatus and conditions. These broad trends apply to comparable hydrocarbons; they are not universal predictions for every carbon compound or mixed commercial fuel.

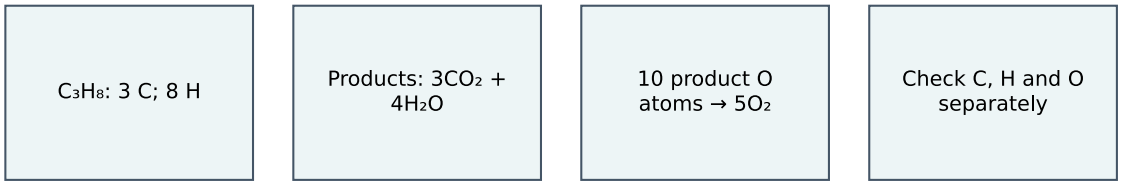
Complete combustion: conserve atoms when hydrocarbons burn

Official-unit focus: 4.7 Organic chemistry

The water formed when propane burns was not stored as liquid water inside the fuel. Its hydrogen atoms combine with oxygen during a reaction that also forms carbon dioxide.

Combustion of hydrocarbon fuels releases energy. During complete combustion with sufficient oxygen, carbon is oxidised to carbon dioxide and hydrogen to water. Begin with hydrocarbon + oxygen → carbon dioxide + water. Use the given hydrocarbon formula to count carbon and hydrogen atoms, then balance oxygen. This chemical reaction makes new substances, unlike the physical separation in the fractionating column.

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One carbon atom requires one carbon dioxide molecule, while each pair of hydrogen atoms forms one water molecule. Oxygen atoms on the product side come from both products, so count both before finding O_2. If an intermediate oxygen coefficient is a half number, multiply every coefficient by two for smallest whole-number coefficients. Never alter the given hydrocarbon's subscripts to force balance.

Write correct formulae first, balance carbon and hydrogen, then oxygen, and check all three element totals. State the oxygen condition in the explanation. Supplied combustion observations can support energy release; they do not alone prove every carbon atom became carbon dioxide. Incomplete combustion products and atmospheric effects belong to their other course sections. Use teacher-approved demonstrations or written evidence rather than independent fuel-burning trials.

■ Checked worked case

Known: propane is C_3H_8. Three carbon atoms give $3CO_2$ and eight hydrogen atoms give $4H_2O$. Product oxygen count=$3\times2+4=10$, requiring $5O_2$. The balanced equation is $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$. Check C:3→3, H:8→8 and O:10→10. For ethane, $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$ avoids a fractional coefficient.

■ Common error

Equal atom totals do not imply equal numbers of molecules or equal masses of carbon dioxide and water. A smoky flame is not proof of complete combustion. This task needs the supplied formula and atom conservation, not memorised combustion equations for every named fuel or an unrequired enthalpy calculation.

Cracking: chemical change makes useful smaller molecules

Official-unit focus: 4.7 Organic chemistry

Fractional distillation can separate a long-chain fraction, but separation cannot change a molecule into a shorter one. Cracking supplies a chemical route to smaller molecules for fuels and manufacturing.

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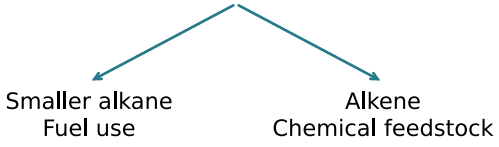
Cracking breaks larger hydrocarbons into smaller molecules. In catalytic cracking, vaporised hydrocarbon passes over a hot catalyst. In steam cracking, hydrocarbon vapour is mixed with steam and heated to a high temperature. General conditions are required, rather than one universal numerical temperature. Products include alkanes and alkenes. Alkenes are more reactive and react with bromine water, changing it from orange to colourless.

Supplied cracking example: count both elements

$$C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$$

Carbon: 10 = 8 + 2

Hydrogen: 22 = 18 + 4



Alkene + bromine water: orange → colourless

High demand for small-molecule fuels makes some cracking products useful as fuels. Alkenes can be used to make polymers and other chemicals. A cracking equation is an atom-conservation example; different products can form depending on conditions. A formula supplied in the equation must be preserved. The full named alkene structures and addition reactions are separate Chemistry-only 4.7.2 requirements, not evidence that this introductory lesson completes them.

Count carbon and hydrogen on each side of the supplied equation, solve for a missing formula or coefficient, and check both elements. Contrast cracking with distillation: cracking forms new substances, while distillation separates existing ones. Bromine-water observations need approved teacher-controlled microscale equipment and ventilation; students must not improvise a heated cracking apparatus or expose themselves to bromine. Written interpretation prepares practical reasoning without certifying laboratory work.

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■ Checked worked case

Known: a supplied model reaction is $C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$. Carbon count is 10=8+2 and hydrogen count 22=18+4. The smaller alkane can serve as fuel and the alkene as chemical feedstock. Four represented starting molecules in this particular equation form four molecules of each product, eight product molecules in all. This is not a claim that all cracking always doubles molecule count.

■ Common error

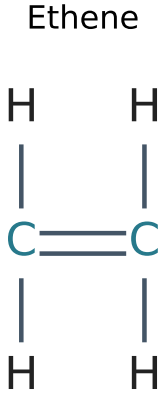
Do not put oxygen into this supplied cracking equation or call it combustion. Steam in steam cracking does not require inventing water as a product in every simplified equation. Bromine water becoming colourless supports an alkene interpretation in the specified hydrocarbon comparison; the observation is not universal identification of any unknown substance.

Chemistry-only: recognise the first four alkenes

Official-unit focus: 4.7 Organic chemistry

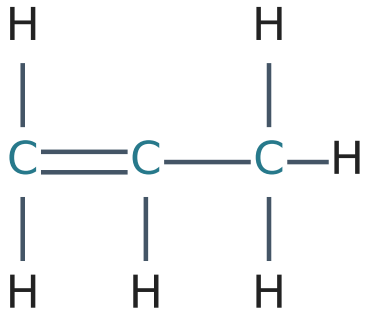
Propane and propene both have three carbon atoms, but propene has two fewer hydrogen atoms. Its carbon–carbon double bond accounts for the difference in the displayed structure.

Alkenes are hydrocarbons containing a carbon–carbon double bond. The functional group C=C helps determine their characteristic reactions. The homologous series with one double bond in an open chain has general formula CnH2n. Its first four members are ethene C_2H_4, propene C_3H_6, butene C_4H_8 and pentene C_5H_{10}. They are unsaturated: each has two fewer hydrogens than the alkane with the same carbon number.

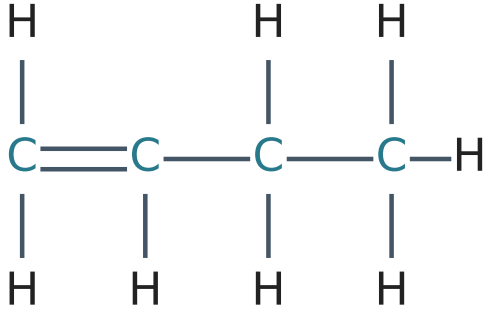


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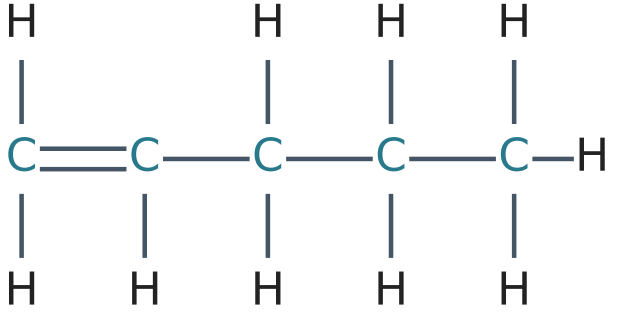
Propene



Butene: shown terminal C=C



Pentene: shown terminal C=C



A displayed formula shows all atoms and all bond lines. Count a double bond as two of carbon's four bonds. In the terminal-double-bond structures shown, the first carbon

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has two hydrogens and the next has one, except ethene where each has two. A longer chain finishes with CH_3. Butene and pentene can have different atom arrangements; the diagrams show one stated straight-chain arrangement, rather than claiming their molecular formula uniquely fixes a structure.

Build the four illustrated molecules using a model kit, then translate each into a displayed drawing and molecular formula. Check hydrogen valency one and carbon valency four. Explain what a flat drawing reveals about connectivity and what it omits about the real three-dimensional molecule. Recognise a supplied longer member from its double bond and formula without requiring a specific name beyond the four stated members.

■ Checked worked case

Known: propane has C_3H_8 while propene has C_3H_6. Hydrogen difference=8−6=2 atoms. For a supplied open-chain member with six carbon atoms and one C=C, H=2n=2×6=12. Its corresponding alkane would have H=2n+2=14. These relations count composition; they do not locate the double bond by themselves.

■ Common error

Do not draw an extra hydrogen on a carbon already making four bonds, or identify every CnH2n formula as an alkene without checking structure: other families can share a molecular formula. Unsaturated does not mean the sample contains no hydrogen. This is separate Chemistry content on both tiers, not a Higher-only extension.

Chemistry-only: hydrogen addition and alkene combustion

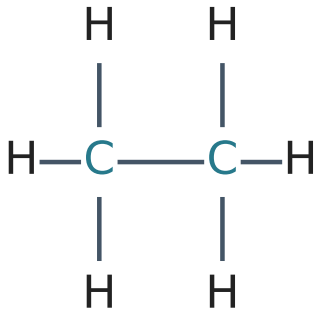
Official-unit focus: 4.7 Organic chemistry

Propene becomes propane when hydrogen adds across C=C. Its three-carbon skeleton stays intact; this reaction adds hydrogen rather than breaking the chain into shorter fragments.

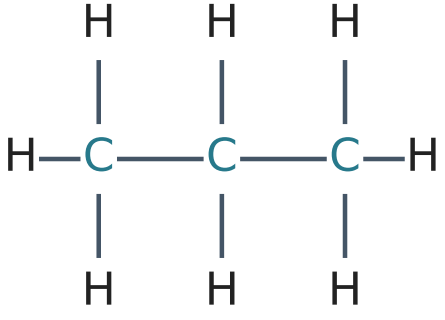
Alkenes react with hydrogen in the presence of a heated nickel catalyst to form alkanes. A representative teaching condition is about 150 °C with nickel; it describes the reaction condition, not a student procedure. One H_2 adds across one C=C: its two hydrogen atoms attach to the two formerly double-bonded carbons and the carbon–carbon bond becomes single. This is hydrogenation, an addition reaction.

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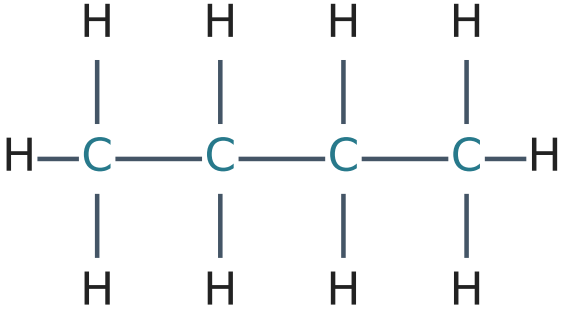
Ethene + hydrogen: displayed product



Propene + hydrogen: displayed product

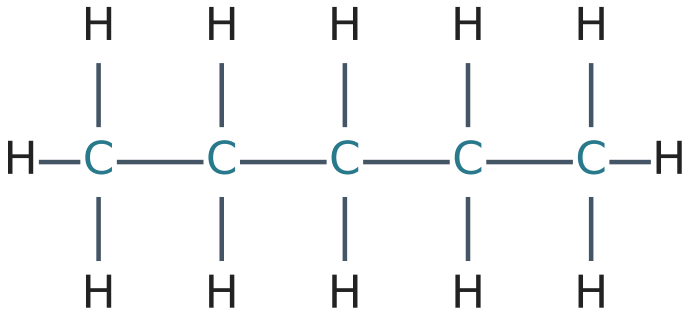


Butene + hydrogen: displayed product



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Pentene + hydrogen: displayed product



An addition reaction joins atoms to the molecule across the multiple bond without removing the existing carbon skeleton. Check the new bond counts: each carbon still makes four bonds. Separately, alkenes burn in oxygen like other hydrocarbons. With sufficient oxygen, complete combustion gives carbon dioxide and water. In air they tend to burn with smoky flames because incomplete combustion produces carbon particles; smoke does not establish complete combustion.

Draw the given alkene first, replace C=C by C–C and add one H to each affected carbon. Retain every other atom. Use the individual displayed product panels for all four named starting members. For combustion, balance carbon, hydrogen and then oxygen. Analyse approved demonstration evidence rather than conducting a pressurised hydrogen/catalyst experiment: combustible gases, heated catalysts and oxygen must remain under trained control.

■ Checked worked case

Known: $C_3H_6 + H_2 \rightarrow C_3H_8$ conserves C=3 and H=8 on each side. In a represented particle batch, six propene molecules consume six H_2 molecules and form six propane molecules if addition is complete. Complete ethene combustion is $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$, with two C, four H and six O atoms on each side.

■ Common error

Do not add two hydrogen atoms to only one carbon or leave the double bond unchanged after addition. A catalyst is not a stoichiometric reactant. The hydrogenation example does not prescribe universal operating conditions for every industrial alkene. A smoky flame indicates incomplete combustion in the stated context, not that water can never form.

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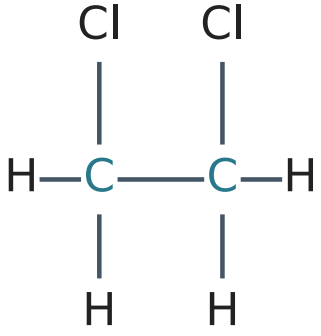
Chemistry-only: chlorine, bromine and iodine addition

Official-unit focus: 4.7 Organic chemistry

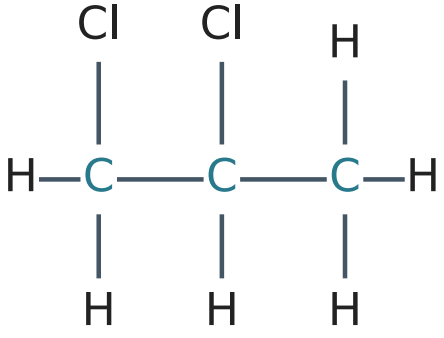
Bromine water loses its orange colour when it reacts with an alkene. A displayed product explains where the two bromine atoms went instead of saying that bromine simply disappeared.

Alkenes react with halogens by addition across C=C. A halogen molecule X_2 supplies one atom to each formerly double-bonded carbon, and C=C becomes a single bond. Here X may be Cl, Br or I. Chlorine and bromine addition are readily described at room temperature without a required UV lamp or nickel catalyst; iodine addition can be slower and less favourable, so the drawing represents the stated addition product rather than promising identical test speed for every halogen.

Ethene + Cl_2: displayed product

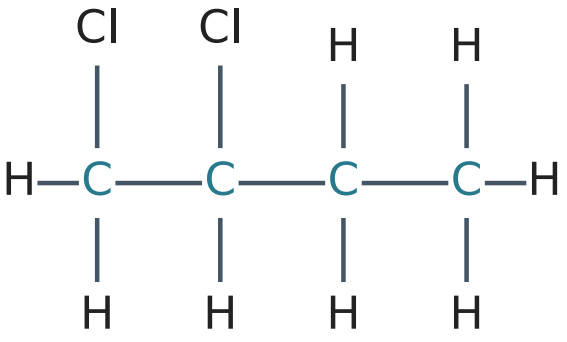


Propene + Cl_2: displayed product

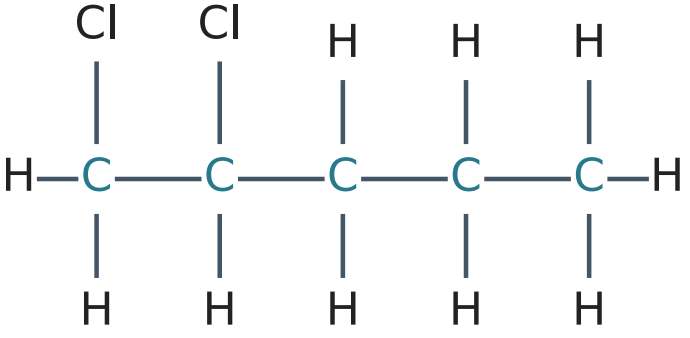


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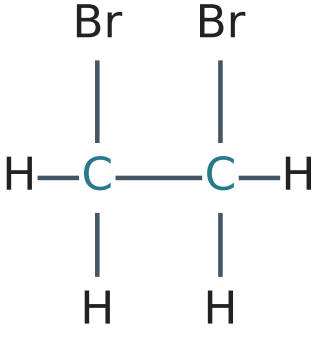
Butene + Cl_2: displayed product



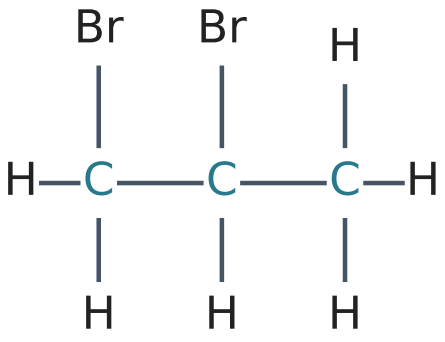
Pentene + Cl_2: displayed product



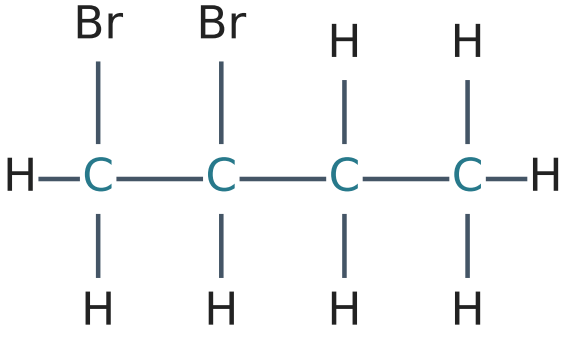
Ethene + Br_2: displayed product



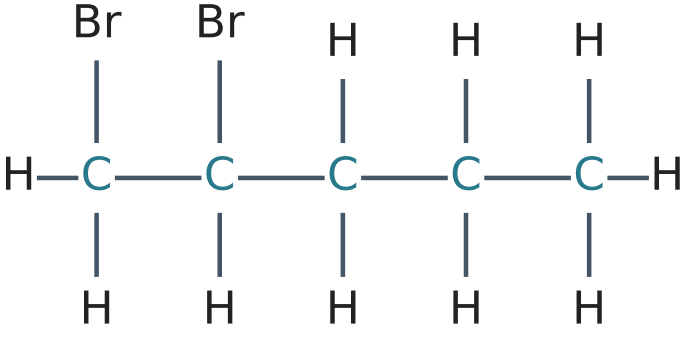
Propene + Br_2: displayed product



Butene + Br_2: displayed product

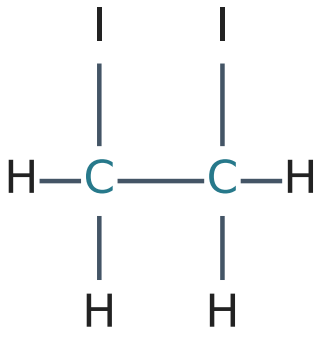


Pentene + Br_2: displayed product

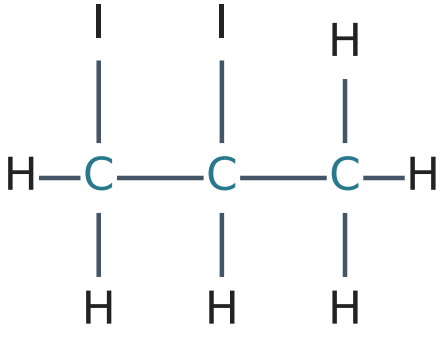


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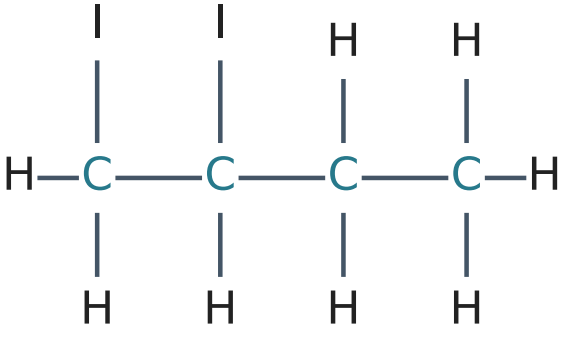
Ethene + I_2: displayed product



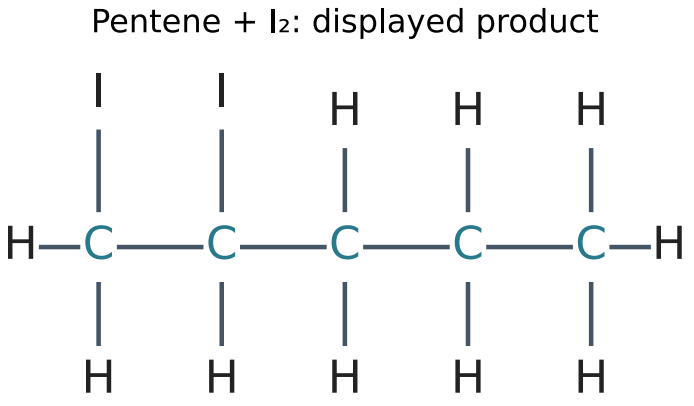
Propene + I_2: displayed product



Butene + I_2: displayed product



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For ethene plus bromine, $CH_2{=}CH_2$ becomes $BrCH_2{-}CH_2Br$. Each carbon retains its two hydrogen atoms and gains one bromine, giving four bond contributions. Bromine water changes orange to colourless in the specified reaction; an alkane control does not give the same rapid addition under these conditions. A carbon skeleton and hydrogen count are retained while two halogen atoms are added. The displayed panels distinguish chlorine, bromine and iodine products for all four named starting alkenes.

Draw the original molecule, mark the two double-bonded carbons, change their link to single and attach one halogen to each. Check every carbon and every hydrogen rather than replacing an H with Br. Record actual colour against a blank/control in teacher-approved microscale demonstrations. Halogens need approved containment, eye protection and ventilation; students must not generate chlorine, handle elemental bromine or improvise iodine reaction conditions independently.

■ Checked worked case

Known: $C_4H_8 + Br_2 \rightarrow C_4H_8Br_2$ accounts for four C, eight H and two Br atoms. Three molecules undergoing one addition each gain six bromine atoms in total. With Cl_2 the corresponding product has two chlorine atoms per molecule; with I_2 it has two iodine atoms. Those supplied product formulae do not describe polymerisation or substitution.

■ Common error

Do not remove a hydrogen in an addition drawing, or claim UV light is the required catalyst for the alkene test. That confuses addition with alkane substitution. Decolourisation is evidence for unsaturation in the controlled comparison, not unique proof of ethene or universal identification of any unknown liquid. A colourless endpoint alone cannot demonstrate that a sample is safe.

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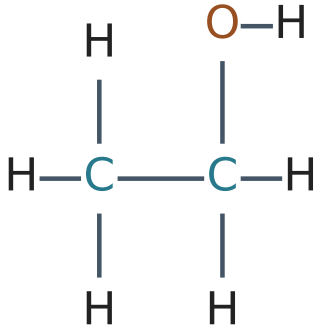
Chemistry-only: steam addition forms an alcohol

Official-unit focus: 4.7 Organic chemistry

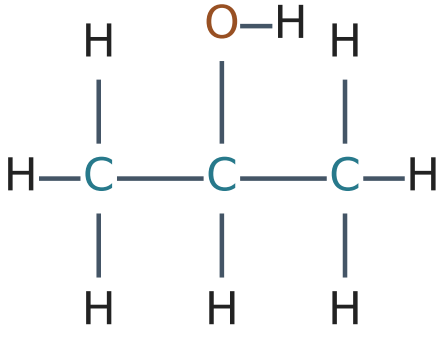
Ethene plus steam can form ethanol. The oxygen atom comes from the water added across C=C; it was not already hidden in the hydrocarbon.

Hydration is addition of water to an alkene. Industrial ethene hydration uses steam at high temperature and pressure with a phosphoric-acid catalyst; about 300 °C and 60 atmospheres are representative conditions. One H_2O supplies H to one formerly double-bonded carbon and OH to the other, while C=C becomes single. Ethene forms ethanol: $C_2H_4 + H_2O \rightarrow C_2H_5OH$. The carbon skeleton is retained.

Ethene + water: stated OH-position product

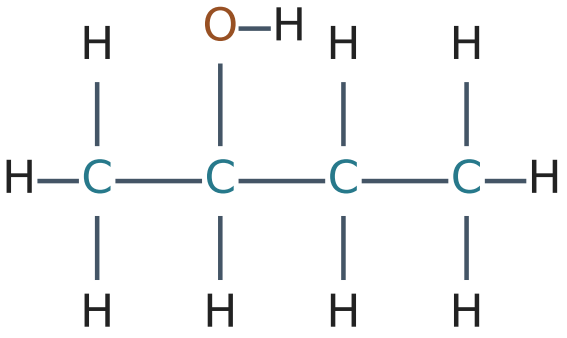


Propene + water: stated OH-position product

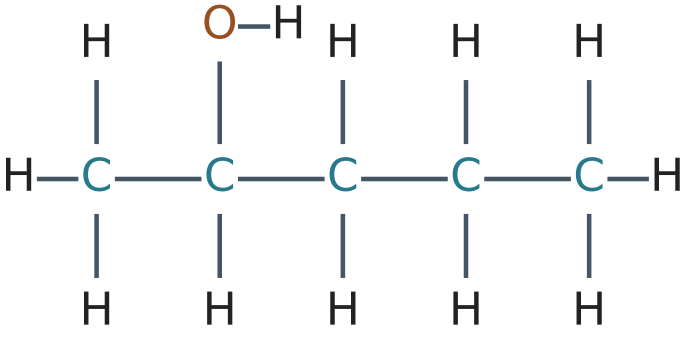


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Butene + water: stated OH-position product



Pentene + water: stated OH-position product



The product is an alcohol, not a hydrocarbon, because it now contains oxygen. Count the H inside OH as well as those bonded directly to carbon. Ethene's two carbons are equivalent in this simple addition. Unsymmetrical starting alkenes can give different OH positions: the longer terminal-chain panels show the usual secondary-alcohol arrangement with OH on the second carbon. These drawings identify stated products without adding a required advanced mechanism or demanding positional product names beyond the course scope.

Draw all bonds including C–O and O–H, and check carbon four, oxygen two and hydrogen one. Compare the product with the input molecule and steam rather than adding OH alone and losing the extra H. The paired first-four displays show before/after connectivity. These industrial conditions are reference information, not a school pressurised synthesis task. Actual school work can use models and supplied process evidence under supervision.

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■ Checked worked case

Known: ethene has four hydrogen atoms. Adding water contributes two, so ethanol contains six H in total: five attached to carbon and one to oxygen. Its total atom count is 2 C + 6 H + 1 O = 9. A displayed product from C_3H_6 plus water has C_3H_8O, eleven atoms in total. The formula alone does not specify the OH position.

■ Common error

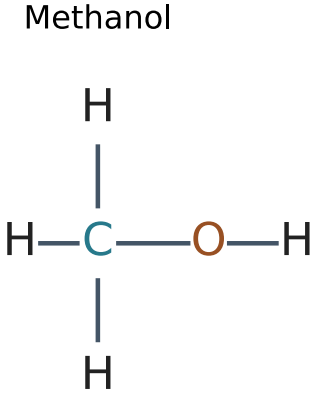
Do not use cold liquid water alone as the specified ethene hydration condition or substitute nickel for phosphoric acid. Hydrogen addition gives an alkane; steam addition gives an alcohol. Industrial reaction need not convert every molecule in one pass. Hydration is distinct from the yeast fermentation route to an aqueous ethanol solution.

Chemistry-only: alcohol structures, names and uses

Official-unit focus: 4.7 Organic chemistry

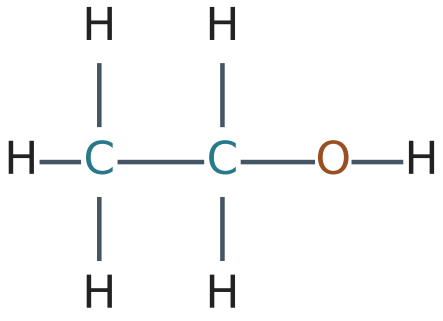
Ethanol can dissolve substances in a formulation and can also serve as a fuel. Its useful roles follow its structure and properties, rather than making it interchangeable with water.

Alcohols contain the –OH functional group attached to a carbon framework. The first four named members here are methanol CH_3OH, ethanol CH_3CH_2OH, propanol $CH_3CH_2CH_2OH$ and butanol $CH_3CH_2CH_2CH_2OH$. These displays use the straight-chain primary arrangements for the last two names. Oxygen makes two bonds, one to carbon and one to H. The total hydrogen count includes the H in OH.

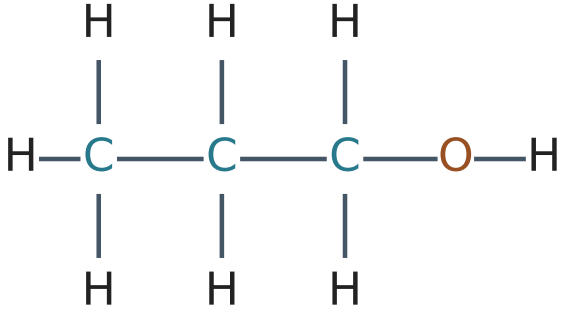


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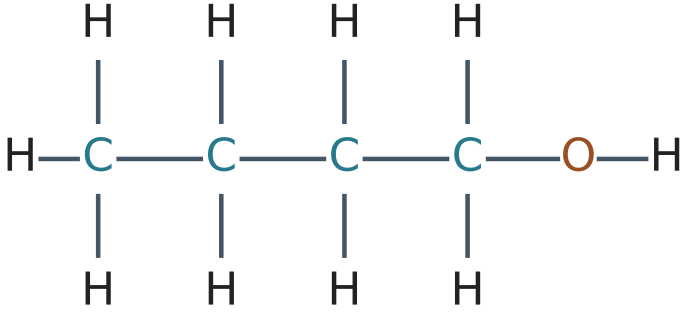
Ethanol



Propanol: shown primary form



Butanol: shown primary form



The first-four alcohols are used as fuels, solvents and starting materials for making other chemicals. Methanol is an industrial fuel/feedstock and solvent; ethanol is used

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as a fuel and solvent in suitable formulations; propanol and butanol serve as solvents and chemical feedstock. The selected use must be tied to actual product information, concentration and purpose. None of these industrial uses implies a sample is suitable to drink or apply to skin.

Compare molecular, condensed and displayed formulae. Locate OH, count the carbon chain and identify the given member. Distinguish a single alcohol molecule from an aqueous solution containing many water molecules as well. Use teacher-provided labelled structures or approved product evidence; do not identify a liquid by taste or smell. Specific names beyond the four stated members are not required for the section.

■ Checked worked case

Known: CH_3CH_2OH contains two C, six H and one O atom, nine altogether. Four represented ethanol molecules contain $4\times9=36$ atoms, including four oxygen atoms. Methanol has CH_4O when all atoms are collected into its molecular formula, while CH_3OH makes the OH group clearer. Neither way of writing it changes its identity.

■ Common error

Do not label every molecule with oxygen an alcohol: an OH within COOH has a different functional-group context. The alcohol OH is covalently bonded, not a free hydroxide ion. The displayed primary propanol/butanol examples do not claim every possible atom arrangement with the same molecular formula has the same reactions.

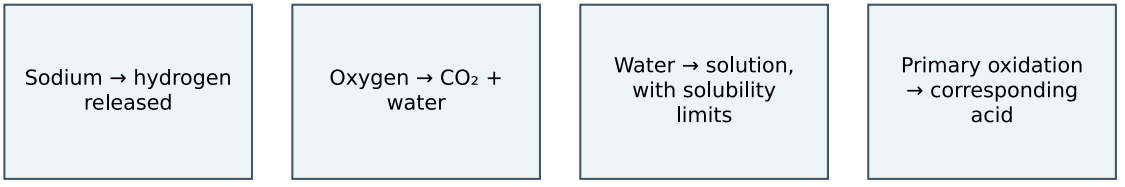
Chemistry-only: four observations for the first alcohols

Official-unit focus: 4.7 Organic chemistry

Ethanol burns to carbon dioxide and water, yet its reaction with sodium releases hydrogen. Naming the reagent prevents these two observations being confused.

The first-four primary alcohols shown react with sodium to release hydrogen and form a sodium-containing organic product. They burn in sufficient oxygen to form carbon dioxide and water, releasing energy. They dissolve in water to form solutions, though solubility decreases along these examples and butanol dissolves to a more limited extent. With a suitable oxidising agent under appropriate heating, the shown primary alcohols can be oxidised to the corresponding carboxylic acids.

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For the displayed primary examples, methanol can give methanoic acid, ethanol ethanoic acid, propanol propanoic acid and butanol butanoic acid. The OH carbon's position matters; a different alcohol arrangement cannot automatically be assigned the same oxidation product. In an approved test, acidified potassium dichromate(VI) can change orange to green as it oxidises the primary alcohol. The GCSE requirement asks for observations/descriptions, not aldehyde mechanisms or balanced equations for these non-combustion reactions.

Record reagent, conditions and observation in separate columns. Describe solution formation with water as mixing/dissolving, not proof that water has made a new alcohol. Any sodium or oxidising-agent demonstration requires the school's risk assessment, trained handling, eye protection and approved small quantities; chromium(VI) reagents need controlled waste. Students should analyse reference observations if the school does not approve the demonstration. Do not add sodium to an aqueous alcohol mixture to infer an alcohol-only reaction.

■ Checked worked case

Known: complete ethanol combustion is $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$. Carbon count is 2, hydrogen 6 and oxygen 7 on each side; the fuel already contributes one O atom. Two ethanol molecules consume six O_2 molecules in this represented complete-combustion equation. Balancing methanol gives $2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$.

■ Common error

Do not forget the oxygen already in the alcohol when balancing combustion, or call the hydrogen from sodium reaction carbon dioxide. Methanol and other alcohols can be toxic and flammable; miscibility does not establish safety. A written response or a supplied colour change does not certify completion of an actual school practical.

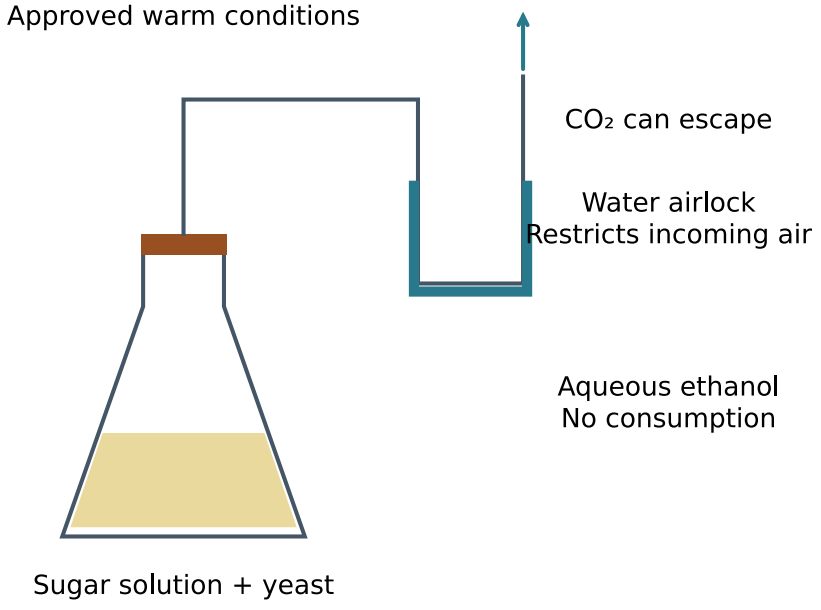
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Chemistry-only: yeast fermentation produces aqueous ethanol

Official-unit focus: 4.7 Organic chemistry

Bubbles from a yeast/sugar mixture are evidence of gas production. They do not directly measure ethanol concentration or show that every molecule of sugar has fermented.

Yeast enzymes convert sugars in aqueous solution into ethanol and carbon dioxide in anaerobic conditions. A warm temperature, commonly around 30–40 °C in teaching examples, supports fermentation; low temperature slows activity and excessive heat damages the enzymes or kills yeast. Oxygen is excluded from the fermentation mixture while carbon dioxide can escape through an approved arrangement. The result is an aqueous solution of ethanol, not pure ethanol.



Teacher-approved apparatus; never tightly seal a gas-producing vessel

Maintain the school-selected temperature and sugar concentration, use the prepared yeast and record actual gas or mass observations over time. An airlock can let gas leave without freely admitting outside air; it is not a licence to tightly seal a gas-producing bottle. The rate can change as substrate runs low or ethanol inhibits yeast. Gas bubbling is qualitative unless quantity is measured, and a silent vessel is not by itself proof that no ethanol is present.

For an actual supervised activity, compare matched labelled mixtures with a yeast-free control and the same volumes, sugar concentration and temperature. Retain raw observations and repeat comparisons. Analyse differences cautiously because gas leakage, dissolved carbon dioxide and contamination affect measurements. Any later separation of ethanol requires approved equipment and no flame near flammable vapour. Products of school fermentation must not be consumed.

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■ Checked worked case

Known: three illustrative gas readings over the same interval are 18, 21 and 21 cm^3. Mean volume=(18+21+21)/3=20 cm^3. A yeast-free control gives 2 cm^3, so the difference is 18 cm^3 under this measurement method. This does not directly determine ethanol volume without an appropriate relation and complete gas accounting. The route differs from adding steam to ethene with an acid catalyst.

■ Common error

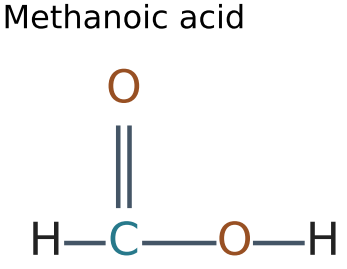
Do not replace yeast with nickel, use the industrial hydration temperature for living yeast or describe fermentation as requiring an oxygen supply. A mass loss can include escaping gas but needs its own controls. No balanced fermentation equation is required by this alcohol section, and the task is preparation for authentic supervised work rather than a replacement written examination.

Chemistry-only: recognise acids and explain carbonate observations

Official-unit focus: 4.7 Organic chemistry

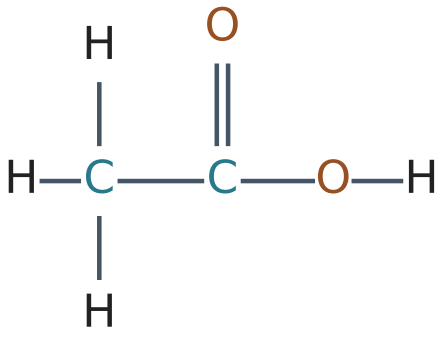
Ethanoic acid contains both C=O and O–H within one COOH group. Treating its OH as an ordinary alcohol group would predict the wrong family and reactions.

Carboxylic acids contain the –COOH functional group. The first four named members are methanoic acid HCOOH, ethanoic acid CH_3COOH, propanoic acid CH_3CH_2COOH and butanoic acid $CH_3CH_2CH_2COOH$. The carbon in COOH counts as part of the total carbon number. It has a double bond to one oxygen and a single bond to an O–H group. The individual displayed panels show every atom and bond.

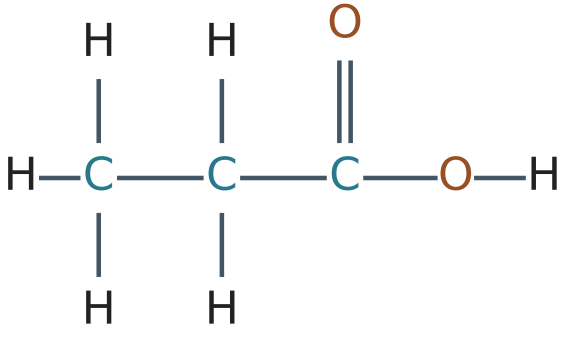


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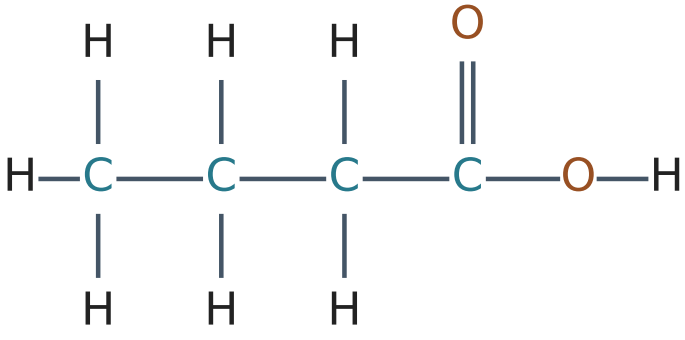
Ethanoic acid



Propanoic acid



Butanoic acid



These first-four acids dissolve in water to make acidic solutions. A suitable indicator or measured pH demonstrates acidity under the stated conditions. They react with

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carbonates to form a salt, carbon dioxide and water. For example, ethanoic acid with sodium carbonate forms sodium ethanoate plus the gas and water. Carbon dioxide can turn limewater milky in an approved product test. Reaction with alcohols forms an ester and water, developed in the separate ester lesson.

Compare a supplied acid structure with the named member by counting all carbons, including COOH. For an approved supervised carbonate comparison, use teacher-selected dilute acids, comparable volumes and carbonate quantities, and record effervescence or gas measurements. Keep a path for gas to escape or reach the approved collector. Indicator colour identifies acidity, not the precise acid; different concentration or temperature can affect reaction rate.

■ Checked worked case

Known: CH_3CH_2COOH has three C, six H and two O atoms, eleven altogether. The acid name is propanoic acid: the two-carbon CH_3CH_2 fragment is joined to a third carbon in COOH. Four represented molecules contain $4\times11=44$ atoms and eight oxygen atoms. Counting only CH_3CH_2 would give the wrong family member.

■ Common error

Do not omit the carboxyl carbon or require a balanced carbonate equation where the specification asks only for a description. Common-tier acidity and reaction observations remain separate from the Higher explanation of partial ionisation. A weak-acid description does not establish low concentration or safe handling, and acidity alone does not prove that an unknown is ethanoic acid.

Chemistry-only: an acid and an alcohol form an ester

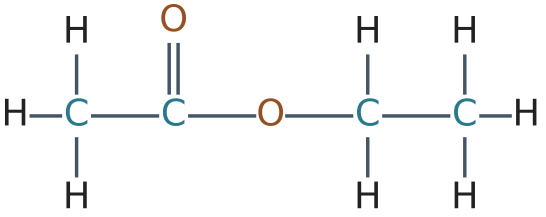
Official-unit focus: 4.7 Organic chemistry

Warming ethanol with ethanoic acid under an approved acid-catalysed method can make ethyl ethanoate. The name keeps information about both starting organic compounds.

A carboxylic acid reacts with an alcohol to form an ester and water. Ethanoic acid plus ethanol forms ethyl ethanoate plus water. A suitable acid catalyst and warming support the laboratory reaction; school demonstrations use a controlled water bath and approved small quantities because the reagents and product can be flammable and concentrated catalyst is hazardous. The course requires the ester name ethyl ethanoate, not a memorised list of other ester names.

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Ethyl ethanoate: acid and alcohol carbon frameworks retained



Ethanoic acid + ethanol → ethyl ethanoate + water

$C_4H_8O_2$; ester link C(=O)–O–C

The illustrated product contains a C(=O)–O–C connection. The acid-derived carbon framework and alcohol-derived carbon framework both remain. It is different from simply mixing an acid with water, from sodium carbonate reaction and from alkene addition. Formation need not be complete in one trial. A characteristic smell may be supplied as reference evidence but is not a unique identification test or a reason to inhale concentrated vapour.

Use the word equation and the displayed model to trace carbon and oxygen atoms. In an approved teacher-led demonstration, record the starting reagents, warming/catalyst conditions and evidence without treating a fruity description alone as conclusive. A comparison may use supplied spectra or labelled products if no practical is approved. Do not improvise a flame, sealed heated vessel, concentrated acid addition or consumption of the product. Actual school evidence remains separate from written practice.

■ Checked worked case

Known: ethanol has two carbon atoms and ethanoic acid has two, so ethyl ethanoate has four. Its molecular formula $C_4H_8O_2$ has fourteen atoms: 4+8+2=14. The remaining atoms form water in the model. Two represented ester molecules contain 28 atoms. The word equation is sufficient for the acid-reaction requirement; balanced symbol-equation recall is not imposed.

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■ Common error

Do not confuse the acid catalyst with a substance consumed in the overall word equation, or name the product ethanol because an alcohol was used. An ester is not the same functional group as its parent acid. The water product does not prove every condensation process or polymer structure has been taught; those are separate 4.7.3 requirements.

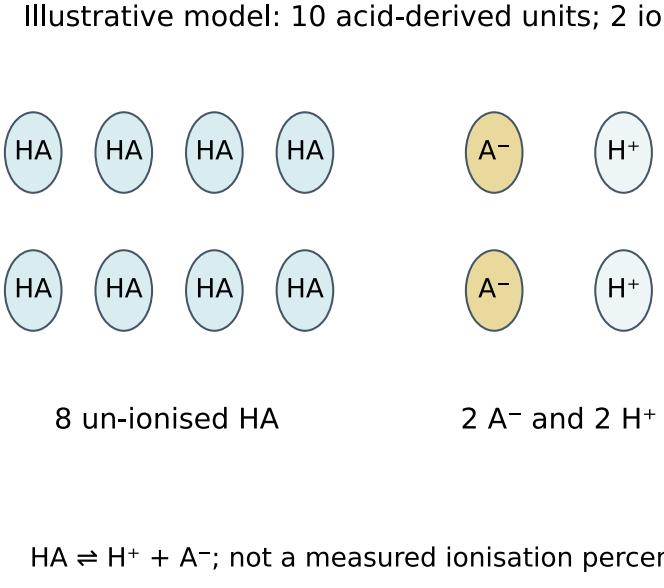
Higher Tier: carboxylic acids are only partially ionised

Higher Tier focus; see section scope notes.

Official-unit focus: 4.7 Organic chemistry

An ethanoic acid solution can be acidic while most acid particles remain un-ionised. Weak describes the extent of ionisation, not a guarantee of a dilute or harmless solution.

Carboxylic acids are weak acids: in aqueous solution they are only partially ionised. Some acid molecules form hydrogen ions and the corresponding negative carboxylate ions, while others remain un-ionised. At the same stated concentration, a weak acid generally has fewer hydrogen ions and a higher pH than a strong acid in this GCSE comparison. The reversible model reflects a balance between ionisation and recombination.



Concentration describes the amount of acid in a volume; strength describes its extent of ionisation. These are independent labels. Comparing a dilute strong acid with a concentrated weak acid does not establish the simple same-concentration pH ordering.

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For whole-number pH differences, a decrease of one pH unit means ten times the hydrogen-ion concentration. No equilibrium constant, pKa calculation or advanced weak-acid approximation is required here.

Use teacher-provided concentrations and pH data or an approved supervised pH measurement. Keep temperature and measurement calibration comparable. State what is measured and what explanation depends on equal concentration. A particle model should conserve total acid-derived units and electrical charge: each ionised unit supplies one positive H^+ and one negative carboxylate in the illustrated monoprotic example. Particle counts are a model, not a measured degree of ionisation for all samples.

■ Checked worked case

Known: solution A has pH 2 and B pH 4. The difference is 2, so hydrogen-ion concentration ratio A:B=10^2=100:1. If a fictional closed model contains 100 acid-derived units with 5 ionised, it has 95 un-ionised acid molecules, 5 carboxylate ions and 5 H^+ ions. That count illustrates partial ionisation without asserting an actual ethanoic acid percentage.

■ Common error

Do not say weak acids cannot react with carbonates or have no hydrogen ions. A concentrated weak acid may still have substantial hazards. The pH ratio concerns free H^+, not automatically the ratio of total acid concentrations. This whole explanatory clause is Higher-only; recognising acids and describing their water/carbonate/alcohol reactions remain both-tier.

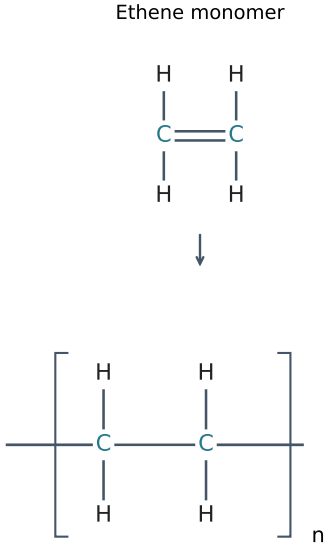
Chemistry-only: an alkene becomes an addition polymer

Official-unit focus: 4.7 Organic chemistry

Ethene is a small molecule with C=C. Poly(ethene) contains a long chain with single carbon–carbon bonds: joining changes the bonding arrangement while retaining the atoms of the ethene units.

In addition polymerisation many small monomer molecules join to form a very large polymer molecule. For ethene, C=C becomes a single bond within each two-carbon unit, and each carbon also bonds along the new chain. Poly(ethene) has repeating unit $–CH_2–CH_2–$. No other molecule is formed in this addition reaction, so the repeating unit contains the same two carbon and four hydrogen atoms as one ethene monomer.

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Repeating interior; end groups not shown
Single bonds cross brackets; all monomer atoms retained

Draw the two-carbon backbone horizontally, keep both hydrogens on each carbon and show a single bond extending out of each side of the brackets. Put n outside the brackets to represent many repeats. A continuation bond joins the neighbouring unit; it is not an extra hydrogen or a separate monomer still containing C=C. Carbon has four bonds when the continuation bonds are counted. The bracketed unit represents the repeating interior of the chain, not its complete end-group structure.

Use a molecular kit or atom-labelled cards to build several ethene molecules. Replace each C=C by a single link, then join their carbon backbones without removing any atoms. Translate that model into a short chain segment and a bracketed repeating unit. Reverse the reasoning for a supplied unit: identify the adjacent backbone carbons that came from the double bond, restore C=C and remove the continuation bonds when drawing the monomer.

■ Checked worked case

Known: a model chain interior contains 25 ethene-derived repeating units. Each contributes two carbon atoms, so the interior contains $25\times2=50$ carbon atoms and $25\times4=100$ hydrogen atoms. These counts concern the stated repeating units; they do not specify the actual initiator-derived end groups or claim every polymer molecule has exactly that chain length.

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■ Common error

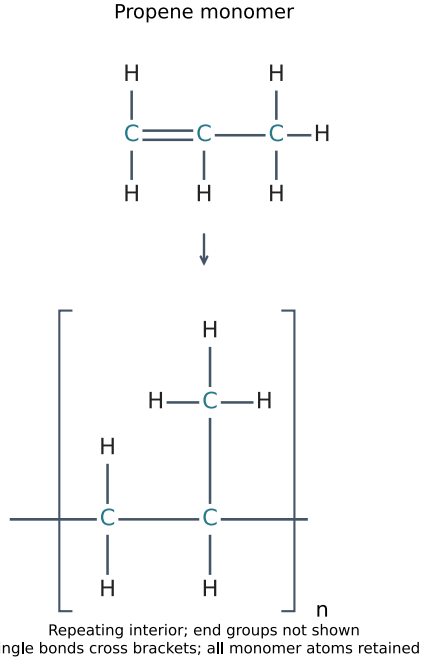
The polymer does not retain a double bond in every repeating unit, and addition polymerisation does not release water. Do not draw disconnected brackets around a complete alkane molecule: the chain must continue through single bonds. This separate Chemistry objective applies to both tiers. Detailed industrial mechanisms and named initiators are not required by this section.

Chemistry-only: keep side groups when drawing a repeating unit

Official-unit focus: 4.7 Organic chemistry

Propene has three carbon atoms, but only two belong to C=C. In poly(propene), those two form the repeating backbone while the third stays as a CH_3 side group.

Propene can be written $CH_2{=}CH{-}CH_3$. Addition polymerisation gives the repeating unit $-CH_2-CH(CH_3)-$. The double-bond carbons become neighbouring backbone carbons with single continuation bonds. The methyl side group remains attached to the same carbon; it is not cut off or automatically inserted as a third backbone position. Each repeating unit retains C_3H_6, the same atoms as the monomer, because no small-molecule by-product is formed.



Trace the two C=C carbons first and label their attached atoms or groups before changing any bond. In the repeat, the CH_2 carbon bonds to two hydrogens and two backbone neighbours. The CH carbon bonds to one hydrogen, the CH_3 group and two backbone neighbours. Both have four bonds. A supplied $CH_2{=}CHCl$ example

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similarly gives –CH_2–CHCl–: chlorine stays attached. This supplied structure is for transferring the model, not an extra list of commercial names to memorise.

Draw a given alkene in a fixed orientation, circle the double-bond carbons and copy their side groups into the repeating unit. Replace the double bond with a single one, add continuation bonds and brackets, then check every atom and valency. For a reversed task, remove the continuation bonds and restore C=C between the corresponding backbone carbons. Equivalent reversed drawings describe the same connectivity; placement on the page is not a different polymer identity.

■ Checked worked case

Known: twelve propene-derived repeating units contain 12×3=36 carbon atoms in total. Twenty-four are backbone carbon atoms, while twelve are in methyl side groups. They also contain 12×6=72 hydrogen atoms in the stated units. Distinguishing backbone from total carbon prevents an incorrect claim that the three-carbon monomer must produce a three-carbon linear repeating backbone.

■ Common error

Do not change CH_3 into CH_2 while leaving it attached as a side group, add water as an addition product or keep the original C=C. A polymer drawing does not establish biodegradability, recycling availability or mechanical strength by itself. Those claims require separate property and life-cycle evidence. The monomer-to-repeat representation applies on both tiers.

Higher Tier: two-ended monomers form a polyester

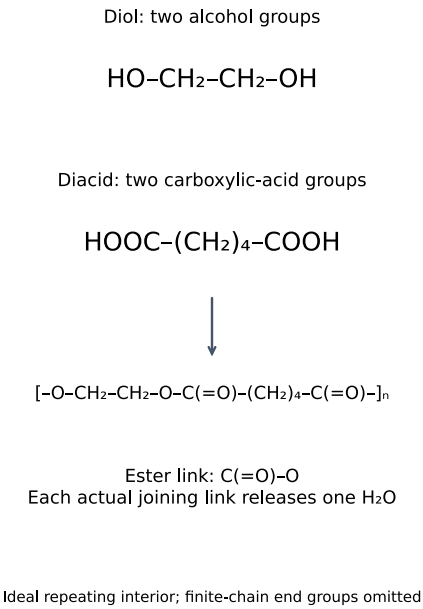
Higher Tier focus; see section scope notes.

Official-unit focus: 4.7 Organic chemistry

An alcohol molecule with two OH groups can join at both ends. Together with a molecule bearing two COOH groups, it can form a chain rather than stopping after a single ester link.

Condensation polymerisation joins monomers with two functional groups, usually losing small molecules such as water. The specification example pairs ethanediol HO–CH_2–CH_2–OH with hexanedioic acid HOOC–$(CH_2)_4$–COOH. Each monomer has two of the same functional group. Alcohol and carboxylic-acid ends react to form ester links C(=O)–O, leaving functional groups available to extend the chain. The resulting polyester contains residues of both monomer types.

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Read the ideal repeating interior as –O–CH_2–CH_2–O–C(=O)–$(CH_2)_4$–C(=O)–. The CH_2–CH_2 segment came from the diol, and the four-CH_2 segment and carbonyl carbons came from the diacid. At each joining event, H from an alcohol OH and OH from an acid COOH make water. The remaining oxygen links the carbonyl carbon to the diol segment. Unlike addition polymerisation, the repeat therefore does not retain every atom of the original monomer pair.

Build a paper or model-kit chain from two-ended diol and diacid cards. Join only compatible ends, mark each ester link and place one water card beside each actual joining event. A single OH/COOH pair cannot extend indefinitely at both ends. Compare the repeat with both starting molecules and explain which functional groups enabled growth. This modelling task is not permission to perform an unapproved polymer synthesis in the laboratory.

■ Checked worked case

Known: a finite, unbranched chain is built from three diol and three diacid molecules, with no ring or extra cross-links. Six molecules require five joining links, so five H_2O molecules are formed. The schematic long-chain specification equation uses the ideal repeating-unit convention; exact finite-chain water counts depend on remaining end groups. Counting actual links avoids treating that convention as a measured end-group formula.

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■ Common error

A polyester contains many ester links; the name does not mean every polymer is an ester. Condensation can form other link types, and the lost small molecule need not always be water. Retain the carbonyl double bond in C(=O)–O. This whole objective is Higher Tier Chemistry-only content; it must not make common-tier addition polymerisation Higher-only.

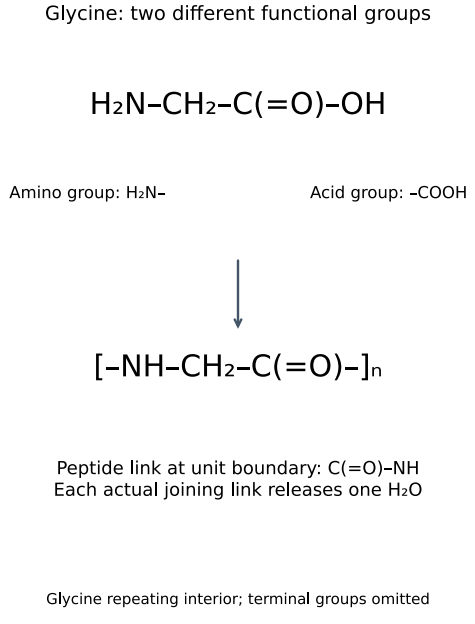
Higher Tier: amino acids join to form polypeptides

Higher Tier focus; see section scope notes.

Official-unit focus: 4.7 Organic chemistry

Glycine $H_2N–CH_2–COOH$ has two different functional groups in one molecule. Its amino end and carboxylic-acid end let it join with other amino acids while forming water.

An amino acid has an amino group and a carboxylic-acid group. In the stated glycine example, the amino group is $H_2N–$ and the acid group is –COOH, with CH_2 between them. Condensation links the acid-derived carbonyl carbon of one molecule to the amino nitrogen of another. The C(=O)–NH connection is a peptide link. Repeated joining forms a polypeptide, represented for glycine by the repeating interior $–NH–CH_2–C(=O)–$.



Each joining event loses OH from one carboxylic-acid group and H from one amino group as H_2O. A chain can extend because functional groups remain at its ends. Different amino acids can combine in the same chain to make proteins; their order

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and folding affect the resulting protein. Glycine-only repetition is a simple model, not a claim that every natural protein contains one repeating amino-acid type. Detailed zwitterion, stereochemistry and protein-structure classifications are not required here.

Use labelled amino-acid cards with distinguishable amino and acid ends. Form a short open chain, mark each peptide link, retain the carbonyl double bonds and account for each water molecule. Then build a chain using several different amino-acid labels to represent sequence variation. Keep chemistry evidence separate from a dietary claim: identifying a peptide bond does not establish the biological role, nutritional value or safety of an unknown sample.

■ Checked worked case

Known: four glycine molecules form a finite open chain with three links. They release three water molecules. Using supplied relative masses glycine=75 and water=18, the chain relative mass is $4\times75-3\times18=246$. This calculation includes its remaining terminal groups, unlike multiplying an ideal interior repeat mass alone. It illustrates conservation of matter; the smaller mass is in water, not destroyed atoms.

■ Common error

A peptide link C(=O)–NH differs from an ester link C(=O)–O. Do not remove the carbonyl oxygen or lose an entire amino group. The one-water-per-link finite-chain calculation excludes cyclic or branched arrangements. Amino-acid condensation chemistry is HT only, while recognising amino acids as protein monomers in 4.7.3.4 applies on both tiers.

Chemistry-only: DNA is made from nucleotide monomers

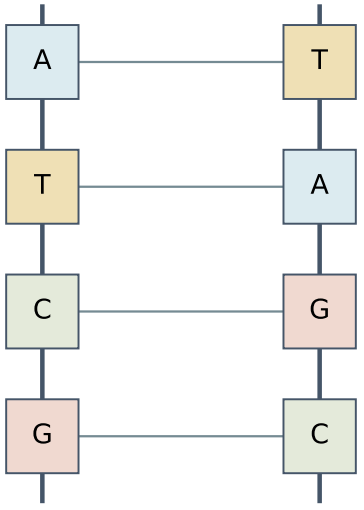
Official-unit focus: 4.7 Organic chemistry

A DNA model has repeated building blocks along two chains. The order of its four nucleotide types can encode information; it is not a chain made from amino acids or glucose.

DNA, deoxyribonucleic acid, is a very large molecule essential for life. It encodes genetic instructions for the development and functioning of organisms and viruses. Most DNA molecules consist of two polymer chains arranged as a double helix. Each chain is built from nucleotide monomers. There are four nucleotide types, commonly identified by the bases A, T, C and G; a base alone is only part of the whole nucleotide.

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Two-chain ladder model before twisting



Each block = one nucleotide, labelled by its base
A↔T and C↔G; shapes and distances not to scale

The chains provide a repeated structural framework while the sequence of nucleotide types varies. A ladder model can distinguish the two continuous chains from paired positions across them. In the familiar double-stranded DNA model, A pairs with T and C with G. Twisting that ladder models the double helix. Colour-coded blocks are a representation: they are not actual nucleotide shapes, accurate atom arrangements or a diagram of separate bases joined without a backbone.

Build an actual classroom model with four labelled nucleotide-card types. Arrange a short sequence along one chain and the complementary sequence along the other, then twist or compare with a supplied three-dimensional model. State what the model shows about repeated units and variable sequence, and what it omits about dimensions and chemical bonds. No human DNA collection or biological sampling is necessary for this representation task.

■ Checked worked case

Known: a short double-stranded model has 18 paired positions. Each position contains one nucleotide from each chain, so there are $18\times2=36$ nucleotides. If one chain contains six A, four T, five C and three G, its partner contains six T, four A, five G and three C. Counting the whole model therefore gives equal A/T and equal C/G totals, without claiming every DNA sequence has equal quantities of all four types.

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■ Common error

A nucleotide is not an amino acid, and four nucleotide types does not mean DNA is only four nucleotides long. A double helix has two polymer chains, not one chain split into four. Recognising DNA and its monomers is Chemistry-only content on both tiers. Additional molecular detail is an explanatory model, not a new requirement to reproduce advanced nucleotide chemistry.

Chemistry-only: match natural polymers to their monomers

Official-unit focus: 4.7 Organic chemistry

Starch and cellulose both use glucose building blocks, but one stores carbohydrate and the other supports plant cell walls. The identity of the monomer alone does not determine the complete polymer structure.

Natural polymers include proteins, starch, cellulose and DNA. Proteins are made from amino acids; starch and cellulose are made from glucose monomers; DNA chains are made from nucleotides. Natural means produced in biological systems, not a promise that a substance is safe, edible or rapidly biodegradable. Each family can contain very large molecules assembled from many smaller building blocks.



Proteins can use different amino acids in the same chain, with sequence and folding contributing to their functions. Starch and cellulose use glucose but differ in how their building blocks are joined and arranged, producing different properties. Starch stores carbohydrate in plants; cellulose gives strength to plant cell walls. The required Chemistry comparison is monomer identity and polymer recognition, not memorising advanced alpha/beta linkage names or complete structural formulae of carbohydrates.

Sort labelled structure or identity cards into monomer/polymer pairs, then justify each match aloud using the building-block evidence. Place starch and cellulose in separate polymer groups even though their glucose cards match. If using a teacher-approved starch observation with iodine, record the actual sample and observed colour; a result for starch does not identify every glucose-containing polymer. Keep practical

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observations, structural inferences and biological-use statements in separate columns.

Checked worked case

Known: a simplified chain diagram labels 15 glucose-derived units in a cellulose segment and 12 in a starch segment. The total is 15+12=27 glucose-derived units. This counts the supplied model units, not free glucose dissolved in either sample. If the protein model beside them has 20 amino-acid residues, those residues must not be added to a glucose-unit count just because all three diagrams represent polymers.

Common error

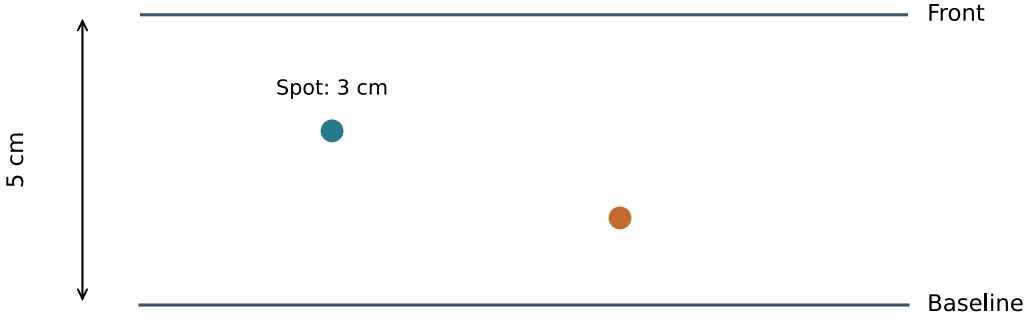
Do not say cellulose has different monomers merely because its properties differ from starch, or that all proteins contain only glycine. A polymer's monomer type and arrangement are separate evidence. Common-tier recognition of amino-acid protein monomers does not require the HT condensation mechanism, and common-tier DNA recognition does not require an invented written practical examination.

Chemical tests and analytical confidence

Official-unit focus: 4.8 Chemical analysis

A coloured flame is useful evidence, but a contaminated wire can give a misleading result. Analytical conclusions depend on the method and controls.

Chemical analysis identifies substances or measures amount. Chromatography separates components because they distribute differently between stationary and mobile phases. A pure substance has characteristic physical properties under stated conditions.



Rf is distance travelled by a component divided by distance travelled by the solvent front, both measured from the baseline. Compare under the same conditions; an Rf value alone does not establish identity across different solvents.

Use pencil for the baseline, keep spots above solvent level, mark the solvent front promptly, and run known references alongside unknowns. For ion tests, use clean

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equipment and separate aliquots to avoid carrying reagents into later tests.

■ Checked worked case

Known: spot travels 3.0 cm; solvent front travels 5.0 cm. Rf = spot distance/front distance. Rf = 3.0/5.0 = 0.60. Rf has no unit. Two matching Rf values support identification only when other evidence and conditions agree.

■ Common error

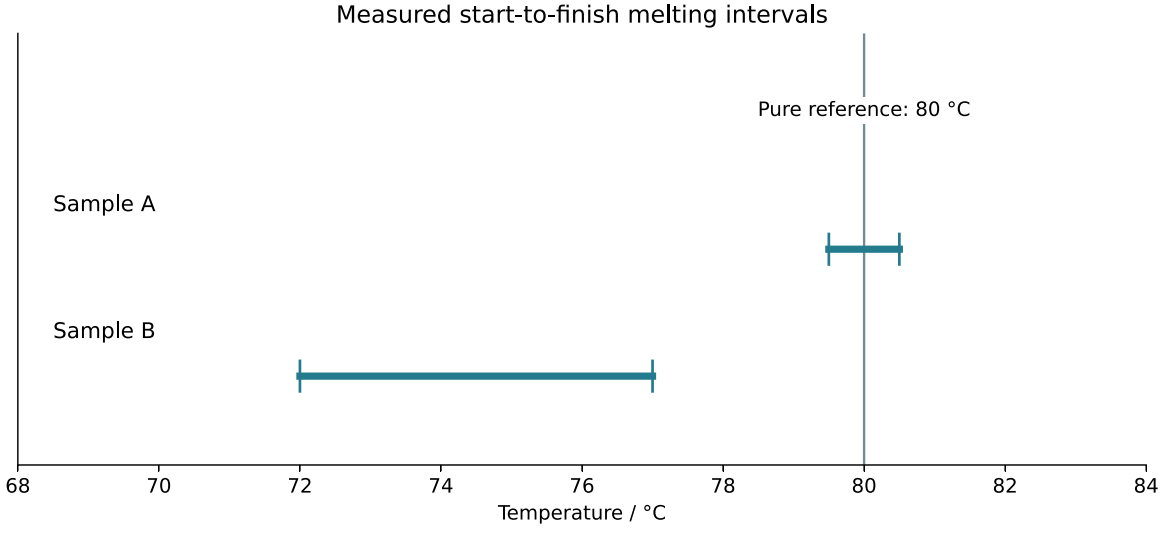
A single spot can conceal substances that co-elute. Ink on the baseline can dissolve and create extra spots.

Chemical purity: use melting and boiling evidence

Official-unit focus: 4.8 Chemical analysis

A bottle labelled pure milk still contains water, proteins, fats and other substances. Chemically pure and nothing added describe different things.

In chemistry a pure substance is one element or one compound, not mixed with another substance. Pure oxygen and pure water satisfy that definition despite water containing two elements chemically combined. Milk is a mixture even if it has no additives. Pure elements and compounds melt and boil at specific temperatures for the stated conditions. A mixture commonly melts over a range; an impurity often lowers and broadens the melting interval relative to a pure solid.



Compare the measured transition with a reference at the same pressure and with calibrated apparatus. A sharp matching melting point supports purity and identity, while a broadened or shifted interval can indicate an impurity. Boiling temperature depends on pressure and the mixture; do not claim every impurity must shift every boiling point in the same direction. A single matching temperature is supporting evidence rather than unique proof of identity, because different substances can have

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similar values.

Read the temperature at which melting begins and the temperature at which the last solid disappears. Record an interval rather than selecting its midpoint to hide a broad transition. Compare repeat measurements and instrument resolution. Any actual school melting-point or boiling demonstration uses approved samples and controlled heating; never heat an unknown flammable material or close a heated vessel to obtain a more convenient reading.

■ Checked worked case

Known: pure reference X melts at 80 °C under the stated conditions. Sample A melts from 79.5 to 80.5 °C; sample B from 72 to 77 °C. Interval widths are 1.0 and 5 °C. B has a broader, lower transition and is more consistent with impure X. The measurement resolution and possibility of another material still need consideration.

■ Common error

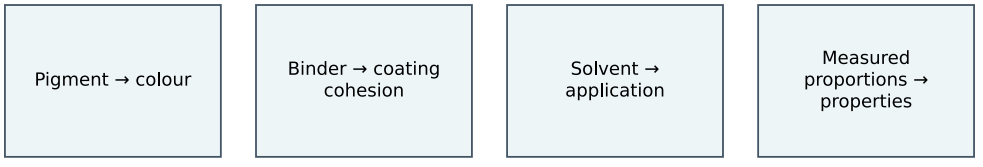
Pure does not mean natural, harmless, colourless or a single type of atom. A compound can be pure. Do not identify an impurity solely by the size of the melting-point change or ignore pressure when comparing boiling data. This section applies on both tiers.

Formulations: each measured component has a purpose

Official-unit focus: 4.8 Chemical analysis

Paint is designed to give colour, spread over a surface and dry into a useful coating. Its components have different purposes; a random mixture would not reliably do the same job.

A formulation is a mixture designed as a useful product. Each chemical has a particular purpose, and carefully measured quantities give the required properties. In a supplied paint example, pigment provides colour, a binder holds the coating together and solvent helps application before evaporating. Formulations also include fuels, cleaning agents, medicines, alloys, fertilisers and foods. Their proportions may be adjusted without changing them into a single compound.



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Use the supplied information to connect component, quantity and intended property. Too much solvent may make a paint spread easily but reduce coating thickness, while the right binder/pigment balance affects coverage and adhesion. These are product-specific trade-offs rather than a universal recipe. A formulation's measured composition is different from the fixed chemical proportions inside a compound. Proprietary component names are not required knowledge.

Compare fictional product data using the same amount of painted surface or another stated service. Calculate component quantities for a chosen batch, then check that the total agrees with the target mass. Students can analyse labelled prepared mixtures or a teacher-approved low-risk model; the task is not permission to compound medicines, fuels or cleaning chemicals independently. A useful product claim requires measured performance, not just a plausible ingredients list.

■ Checked worked case

Known: a fictional 200 g coating uses 30% pigment, 45% binder and 25% solvent by mass. Component masses are $0.30\times200=60$ g, $0.45\times200=90$ g and $0.25\times200=50$ g. Their total is 200 g. Doubling the batch doubles each mass; it does not change any percentage. The numerical formulation is original teaching data, not a commercial product recipe.

■ Common error

Formulation does not mean chemically pure, and not every accidental mixture is designed for a useful purpose. A label saying natural gives no evidence about the proportions or safety of its components. Do not infer an exact formulation from colour alone. Both-tier students identify and reason from supplied composition and purpose.

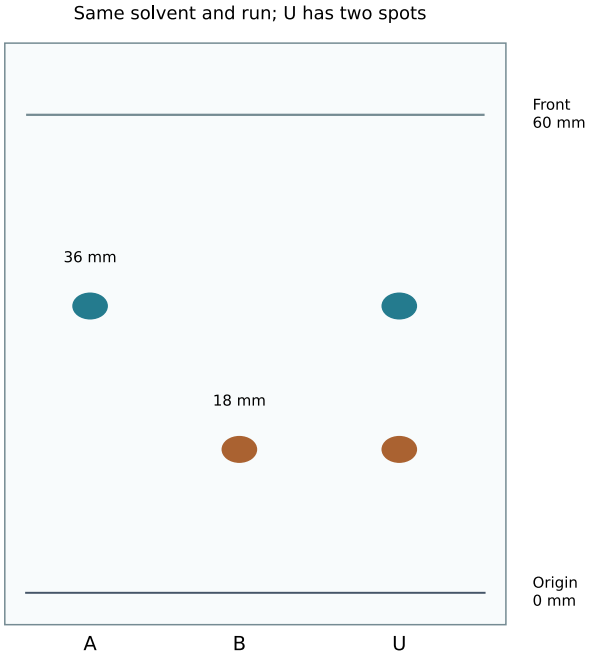
Chromatography: phases, spot positions and Rf

Official-unit focus: 4.8 Chemical analysis

Two dyes begin at the same pencil line but finish at different heights. They were carried by the same solvent; their different distributions between phases cause the separation.

Chromatography separates a mixture and provides evidence for identifying components. The mobile phase moves through the stationary phase. In paper chromatography the paper supports the stationary phase while solvent moves up the sheet. Components distribute differently between the phases: a component carried more in the mobile phase moves farther under the stated conditions. The solvent front marks the farthest point reached by the solvent, not the highest dye spot.

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Rf is distance moved by the substance divided by distance moved by the solvent. Measure both from the origin line, using the centre of the spot for the substance. The ratio has no unit and ordinarily lies between zero and one in this model. Compare unknown and reference spots using the same paper, solvent and conditions. A different solvent can change the separation and Rf; it is not a universal identifying number for one substance.

Read each lane separately and count distinct spots. Multiple spots show multiple detectable components. A single spot is consistent with a pure compound, but a mixture may have components that overlap or are not detected in that solvent. Use another suitable solvent or independent evidence to strengthen a purity judgement. Plot measured spot centres clearly and state when broad or streaked spots make the centre uncertain; do not choose a convenient edge.

■ Checked worked case

Known: solvent front is 60 mm from the origin and an unknown spot centre is 36 mm from it. Rf=36/60=0.60. A reference at 36 mm in the same run supports the same detectable component. Another unknown spot at 18 mm has Rf=18/60=0.30, so the unknown lane contains at least two separated components. Matching one spot does not identify the whole mixture.

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■ Common error

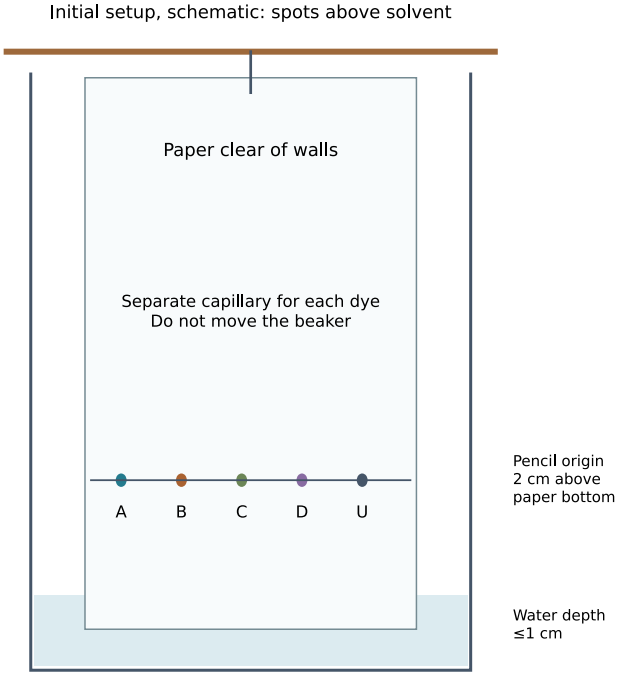
Do not measure from the paper edge, divide solvent distance by spot distance or attach millimetres to Rf. Higher travel is not evidence that a molecule weighs more or that the paper is moving. This both-tier objective includes calculation and appropriately rounded answers; an unreviewed advanced chromatographic technique is not required.

Required practical 6: separate and identify actual food dyes

Official-unit focus: 4.8 Chemical analysis

An unknown food colouring can contain two known dyes and another dye absent from the reference set. A good result must allow an unmatched spot to remain unmatched.

RP6 investigates separation and comparison of coloured substances and requires Rf calculation. The handbook uses four known food colourings A–D, an unknown U, chromatography paper, water and separate capillary tubes. Its teacher notes suggest U contain two known dyes and a third not in A–D. Known single dyes are preferable references; a commercial colouring may itself be a mixture. Written predictions do not certify that the practical has been performed.



Draw a horizontal pencil origin 2 cm above the lower paper edge. Mark positions away from the edges, place small concentrated spots using separate clean tubes and label in pencil. The handbook aims for 2–3 mm spots and no more than 1 cm water depth, keeping the origin above the water while the paper bottom dips into it. Support the paper without touching the beaker walls and avoid moving the vessel during the run.

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Stop before the solvent reaches the paper top, remove the paper and mark the solvent front immediately in pencil. Let the paper dry before measuring from the origin to each spot centre and front. Preserve the actual chromatogram, measurements, solvent identity and calculated Rf. Compare U with references in the same run; record an additional spot as unidentified rather than inventing a name. The school method and risk assessment govern capillary handling and any drying equipment.

■ Checked worked case

Known: in an original model run the front travels 50 mm. A, B and an unmatched U spot travel 10, 30 and 40 mm respectively. Their Rf values are 0.20, 0.60 and 0.80. If U also has spots at 10 and 30 mm, it supports A and B plus another detectable component. These model readings are practice data; the practical record must contain the student's real observations.

■ Common error

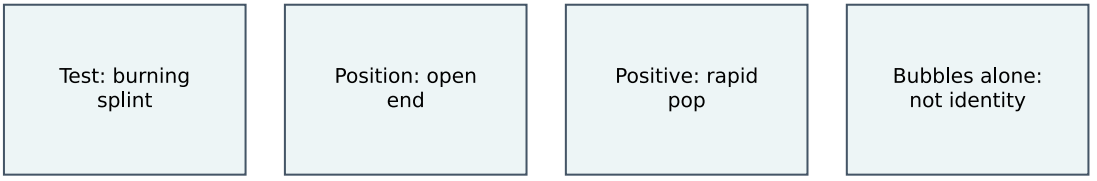
Ink can move and interfere, so the baseline and labels use pencil. An origin submerged in solvent lets samples dissolve into the reservoir. Large spots can overlap; touching walls can distort the front. No spot matching A–D does not mean the unknown contains no dye. RP6 applies on both tiers and assesses actual separation skills alongside written interpretation.

Gas identification: hydrogen gives a squeaky pop

Official-unit focus: 4.8 Chemical analysis

Bubbles from a metal reacting with acid suggest a gas is being made. They do not alone identify hydrogen; the specified test supplies separate evidence.

The hydrogen test uses a burning splint held at the open end of a test tube containing the gas. Hydrogen burns rapidly with a pop sound. State both the test and observation: burning splint at the open end, followed by a pop. A glowing splint relighting belongs to oxygen, while limewater becoming cloudy belongs to carbon dioxide. Merely saying use a splint leaves its condition and the identifying result unspecified.



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A reaction prediction and a gas test answer different questions. A known reactive metal with dilute acid can produce hydrogen, but bubbles alone could be another gas in a different system. A positive pop supports hydrogen under the approved test conditions. A weak or absent observation can reflect a small or diluted sample as well as absence of hydrogen, so an inconclusive trial should not be rewritten as a definite negative.

An actual school gas test requires teacher-selected small quantities, approved collection, an open testing path and controlled ignition. Keep any bulk gas-generating vessel and stored flammable material away from the test flame. Never tightly seal an active gas reaction or scale up an explosive demonstration. Record the actual splint condition, sampling method and observed result, with a labelled apparatus sketch; written identification is preparation rather than laboratory completion.

■ Checked worked case

Known: an original observation record contains six separately approved gas-test trials. Five give a clear pop and one is inconclusive. Clear positive proportion=$5/6\times100\approx83.3\%$. This is a proportion of recorded outcomes, not the hydrogen purity or a recommended number of ignition trials. An inconclusive trial remains explicitly uncertain and cannot be counted as proof of another gas.

■ Common error

Do not insert a glowing splint and call any relighting a hydrogen test. Colourless appearance does not distinguish hydrogen from other colourless gases. The identifying pop is rapid combustion, not evidence that a metal itself popped. This both-tier lesson teaches the specified test and its interpretation; it does not require an invented written practical qualification component.

Gas identification: oxygen relights a glowing splint

Official-unit focus: 4.8 Chemical analysis

A splint whose flame has gone out can still glow. If it relights inside the sampled gas, the specified observation supports oxygen.

The oxygen test uses a glowing splint inserted into a test tube of gas. The splint relights in oxygen. The splint is glowing before it is inserted, rather than already carrying a flame; this makes relighting an observable change. Write glowing splint and relights together in an identification answer. A burning splint giving a pop is the hydrogen test and is not interchangeable with this method.

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Oxygen supports combustion but is not itself described as a flammable fuel in this test. Increased burning is evidence about combustion conditions; the school identification relies on the characteristic relighting observation. Colourless gas or a known source reaction alone is weaker evidence than a correctly performed test. A sample mixed with air or contaminated during collection may give less clear observations, so retain uncertain results rather than inventing an identification.

For actual supervised testing, the teacher chooses the source, collection method and small sample quantity. Prepare the splint with a clearly glowing tip under the approved method, then place it into the collected gas as instructed. Use eye protection and keep ignition arrangements separate from incompatible flammable substances. Record the starting splint state and whether it relights; do not describe the flame's colour as the required identifying observation.

■ Checked worked case

Known: a fictional comparison table records clear relighting in four of five suitable sample observations. The clear-positive proportion is $4/5\times100=80\%$. This illustrates recording and interpretation, not oxygen concentration. Counting outcomes cannot convert the test into a quantitative purity measurement without an independently calibrated method and appropriate evidence.

■ Common error

A splint already burning cannot demonstrate relighting from a glow. Do not write oxygen burns with a pop, or assume every gas that permits some burning is pure oxygen. The qualitative test gives identification support under suitable conditions and has observation limits. Both tiers need the specified test; performing it in school remains distinct from answering a written gas question.

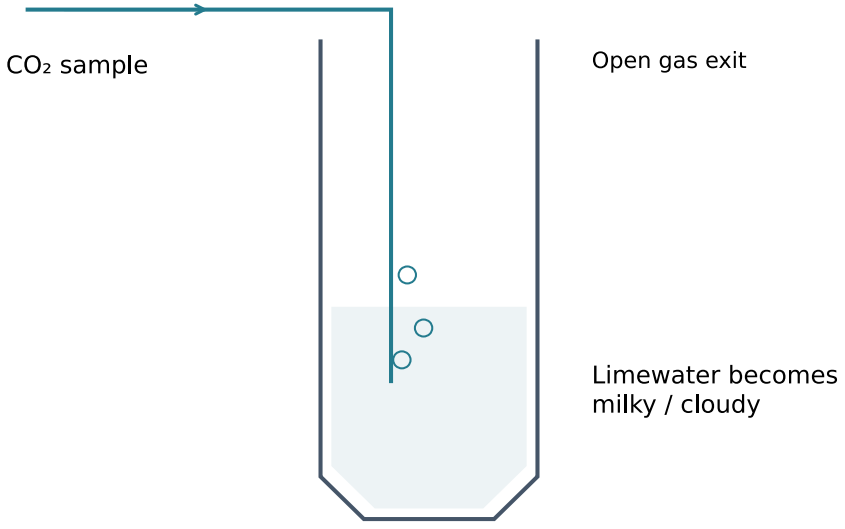
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Gas identification: carbon dioxide clouds limewater

Official-unit focus: 4.8 Chemical analysis

A carbonate reacting with acid releases gas. Bubbling a small sample through limewater gives a cloudy appearance, adding identification evidence to the reaction prediction.

Limewater is an aqueous solution of calcium hydroxide. Carbon dioxide shaken with limewater or bubbled through it turns the solution milky or cloudy. The white insoluble calcium carbonate produced makes it cloudy: $CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$. A precipitate is a newly formed insoluble solid dispersed in the liquid, not simply an unreacted starting powder or a colourless gas bubble.



Approved small-scale test; no sealed gas reaction

State the reagent and the positive result together. A flame going out is not a unique carbon-dioxide test because several gases fail to support combustion. Under continued excess carbon dioxide the cloudiness can eventually disappear by further reaction; that optional observation does not invalidate the initial specified milky result. Do not substitute that extension for the basic test or make students memorise an unrequired second equation.

Use teacher-approved dilute reagents and an open gas-delivery path that prevents pressure buildup. For a carbonate demonstration, add the selected dilute acid only under the school method, then direct a small gas sample into fresh limewater. Avoid suckback and unapproved gas generation. Record the initial clarity, reagent identity and actual change, comparing a suitable blank when provided. Distinguish the measured observation from an inferred gas identity.

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■ Checked worked case

Known: an original observation table contains 12 samples tested with the same fresh limewater method; nine become cloudy and three remain clear. Cloudy-result fraction=9/12=0.75, or 75%. This does not measure CO_2 percentage in any individual sample. A clear result may be inconclusive if gas transfer failed; a carefully documented method matters more than counting positives alone.

■ Common error

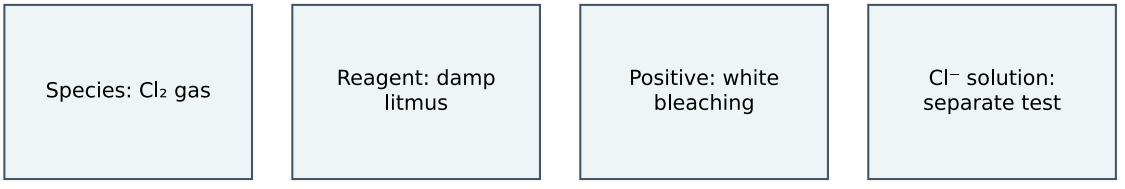
Limewater is not just any calcium-containing liquid, and cloudiness is not a green flame. Do not report a carbonate ion as a gas: the carbonate reaction produces carbon dioxide, whose test provides the observation. The source equation conserves Ca, C, O and H, but a balanced equation does not replace the actual school evidence record.

Gas identification: chlorine bleaches damp litmus

Official-unit focus: 4.8 Chemical analysis

Chloride ions in sodium chloride do not behave like chlorine gas. An identification answer must specify the tested species rather than relying on the shared word chlorine.

The specified chlorine test uses damp litmus paper placed into the gas. Chlorine bleaches the paper so it turns white. The paper must be damp; dry litmus is not the stated method. Blue litmus may first turn red before bleaching, but the identifying final observation is bleaching white. Describe both the damp paper and the observed bleaching, rather than reporting an acid colour change alone.



Chlorine is Cl_2, a molecular element, while chloride is Cl^-, a negative ion in compounds and solutions. Their tests are different: gas bleaching identifies chlorine under suitable conditions, whereas acidified silver nitrate precipitates chloride ions. Pale green gas appearance can support the description but does not justify smelling a sample. Not every substance containing chlorine atoms releases chlorine gas or bleaches paper in the same way.

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Chlorine is harmful to breathe. Actual testing uses only teacher-approved small quantities, suitable ventilation/containment and controlled handling under the current school risk assessment. Never sniff the gas, generate it from improvised cleaning products or independently repeat an electrolysis setup to obtain it. Students can interpret source observations if direct exposure is unsuitable; the teacher records what actual practical experience and demonstration evidence were provided.

■ Checked worked case

Known: a supplied qualitative record has five damp papers becoming white and three tests reported inconclusive. There are eight records total, and clear-bleaching proportion=$5/8\times100=62.5\%$. That proportion describes the observation dataset, not the chlorine concentration or the method's universal accuracy. An inconclusive result remains distinct from an observed non-bleaching result.

■ Common error

Do not confuse bleaching with the simple red/blue indicator response, write dry litmus as the required reagent or use a smell test. Chloride precipitation is a separate solution-ion test. A written explanation does not authorise chlorine handling or certify a supervised task was performed. The gas-test objective applies on both tiers.

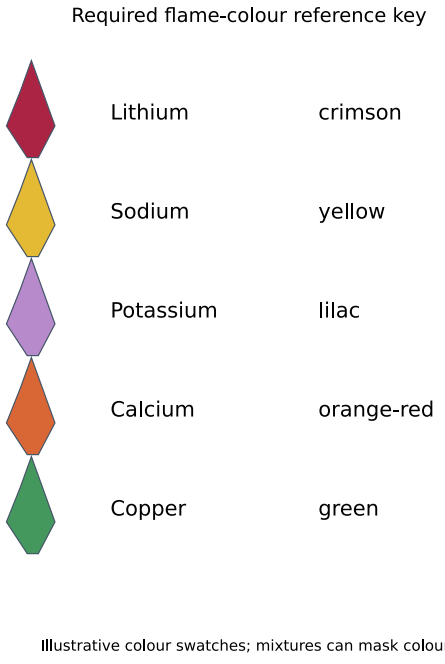
Chemistry-only: identify five metal ions by flame colour

Official-unit focus: 4.8 Chemical analysis

A yellow flame can dominate a mixed sample even when another ion is present. A clean wire and reference colours matter when interpreting a flame test.

Flame tests identify some positive metal ions, or cations, in compounds. The required colours are lithium crimson, sodium yellow, potassium lilac, calcium orange-red and copper green. State the ion and observation together. These are flame colours, not necessarily the colour of the original solid or solution. The source does not require memorising other metal-ion flame colours.

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Illustrative colour swatches; mixtures can mask colours

A sample is introduced into a suitable flame using the school-approved wire or soaked-splint method. Compare its emission with known references under the same conditions. A mixture may have masked colours; intense sodium emission can obscure another response. Yellow therefore supports sodium in the tested conditions but does not prove no other ions are present. Contamination of the wire, sample or apparatus can produce a misleading result.

For actual RP7 school work, use labelled known compounds, approved small amounts and a blue Bunsen flame under supervision. Clean the nichrome wire by the approved method between samples; the handbook suggests fine emery paper and explicitly excludes concentrated hydrochloric-acid watch glasses for GCSE students. Record actual observed colour before naming the cation. Prepared reference photographs or a teacher demonstration can support interpretation but do not silently certify individual practical performance.

■ Checked worked case

Known: an original record contains five unknown samples with flames crimson, yellow, lilac, orange-red and green. Using the required reference table identifies Li, Na, K, Ca and Cu respectively. If three of ten properly recorded samples are lilac, their sample proportion is $3/10\times100=30\%$; it is not potassium concentration or the proportion of potassium atoms in one sample.

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■ Common error

Do not identify chloride ions from a sodium-yellow flame or infer a salt's complete formula from its cation alone. A negative or masked result may be inconclusive. A flame colour is qualitative evidence; concentration measurement requires a calibrated instrument and appropriate data. Separate Chemistry flame-test content applies to both tiers.

Required practical 7: identify both ions in an actual unknown salt

Official-unit focus: 4.8 Chemical analysis

A lilac flame tells you about the positive ion, while acidified barium chloride tells you about the negative ion. A single ionic compound needs both identities before its formula can be justified.

RP7 requires chemical tests on unknown single ionic compounds covering the specified flame, hydroxide, carbonate, halide and sulfate evidence. The handbook's basic activities establish known flame and anion results, then repeat them on the unknown. Teachers also ensure the 4.8.3.2 hydroxide comparisons are covered in the school programme. A written table of predicted colours is useful preparation but is not an actual practical record.

Original model evidence: potassium sulfate

Test	Observed unknown	Inference
Flame	Lilac	K^+ supported
Acid → limewater	No positive result	No carbonate evidence
HCl + $BaCl_2$	White precipitate	SO_4^{2-} supported
HNO_3 + $AgNO_3$	No precipitate	No halide evidence

Fresh portion per test; real RP7 record uses actual observations

Establish reference results with labelled known cation compounds and carbonate/sulfate/chloride/bromide/iodide salts. Use fresh unknown portions for different tests. Keep flame-wire cleaning, dilute-acid identities and precipitating

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reagents exact. Record carbonate gas and limewater, sulfate with dilute HCl/barium chloride, and halides with dilute nitric acid/silver nitrate. Compare subtle silver-halide colours side by side. The handbook suggests potassium sulfate as a convenient unknown; the school can choose another suitable single compound.

Use the approved station arrangement, separate pipettes, current school risk assessment and supervised small quantities. The teacher demonstrates pipette gas transfer to limewater and controls Bunsen use, caustic reagents, soluble barium salts and silver waste. Keep raw observations, test conditions, any no-change/inconclusive result and justified inference. Do not erase contradictory results; check contamination, repeat under approved conditions and retain the explanation. No proprietary reagent-making or concentrated-acid cleaning is assigned to students.

■ Checked worked case

Known: an original unknown gives a lilac flame, no carbonate response, white precipitate in the acidified barium-chloride test and no halide result. Within the stated references this supports K^+ and $SO_4{}^{2-}$. Two +1 potassium ions balance one −2 sulfate ion, giving K_2SO_4. Five formula units contain ten potassium ions and five sulfate ions. Those quantities explain the inferred formula, not observations supposedly collected by the student.

■ Common error

Do not reuse an HCl-treated portion for the silver-nitrate halide test, or infer the anion from a flame. A matrix can leave ambiguity or conflicting evidence; report that honestly instead of inventing a unique salt. The practical uses actual supervised tasks and both-ion reasoning, not an invented written performance exam or an automatically certified completion checkbox.

Chemistry-only: colour and excess alkali identify metal hydroxides

Official-unit focus: 4.8 Chemical analysis

Three different metal-ion solutions can initially make white precipitates with sodium hydroxide. Adding excess reagent supplies another observation: the aluminium precipitate dissolves.

Adding sodium hydroxide solution forms insoluble hydroxides with the specified cations. Aluminium, calcium and magnesium give white precipitates, but only the aluminium hydroxide precipitate dissolves in excess sodium hydroxide in this reference set. Copper(II) gives blue, iron(II) green and iron(III) brown precipitates. A precipitate is an insoluble solid formed by the reaction, not a solution that merely changes colour.

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Cation	NaOH added	Excess NaOH
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

Record initial precipitation separately from what happens on adding excess reagent. A white precipitate that dissolves in excess supports aluminium; one remaining supports calcium or magnesium within the supplied set. Use additional suitable evidence, such as calcium's orange-red flame, when a single observation leaves alternatives. The Roman numerals in iron(II) and iron(III) distinguish ion charges and their different hydroxides. Colour should be compared with known results in similar conditions.

Carry out only the approved supervised small-scale comparison using labelled solutions, separate clean portions and selected dilute sodium hydroxide. Add the reagent as instructed, observe the precipitate, then add excess and record whether it dissolves. Sodium hydroxide is corrosive at sufficient concentration; school controls govern eye protection, handling and waste. Students need not write an equation for producing sodium aluminate: the specification requires the observation but excludes that equation.

■ Checked worked case

Known: supplied unknown A forms white precipitate then clears in excess alkali; B forms white and remains; C forms brown. A supports Al^{3+}, B leaves Ca^{2+}/Mg^{2+} unresolved, and C supports Fe^{3+}. If four of eight white-precipitate samples dissolve in excess, their count proportion is $4/8\times100=50\%$. It does not mean an individual sample is half aluminium hydroxide.

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■ Common error

Do not treat all white precipitates as aluminium, mix up iron hydroxide colours or call an original blue solution a blue precipitate without observing a solid. Excess reagent and initial addition are different test stages. The classification applies on both tiers; the separate equation lesson teaches the required balanced formation equations.

Chemistry-only: balance insoluble-hydroxide formation equations

Official-unit focus: 4.8 Chemical analysis

Iron(II) chloride and iron(III) chloride need different quantities of sodium hydroxide in their balanced formation equations. The metal-ion charge affects the hydroxide formula.

For a supplied soluble metal chloride and sodium hydroxide, metal hydroxide precipitates while sodium chloride remains dissolved. Magnesium and calcium form $Mg(OH)_2$ and $Ca(OH)_2$. Copper(II) and iron(II) form $Cu(OH)_2$ and $Fe(OH)_2$. Aluminium and iron(III) form $Al(OH)_3$ and $Fe(OH)_3$. OH inside brackets is a hydroxide group; its subscript multiplies both oxygen and hydrogen. Do not change the compound formula while balancing.

Balanced formation equations, one stated metal at a time

$$MgCl_2 + 2NaOH \rightarrow Mg(OH)_2 + 2NaCl$$

$$CaCl_2 + 2NaOH \rightarrow Ca(OH)_2 + 2NaCl$$

$$CuCl_2 + 2NaOH \rightarrow Cu(OH)_2 + 2NaCl$$

$$FeCl_2 + 2NaOH \rightarrow Fe(OH)_2 + 2NaCl$$

$$AlCl_3 + 3NaOH \rightarrow Al(OH)_3 + 3NaCl$$

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

Reactant solutions (aq); metal hydroxide precipitate (s)
NaCl remains aqueous; excess-Al reaction not required

The divalent examples have form $MCl_2(aq) + 2NaOH(aq) \rightarrow M(OH)_2(s) + 2NaCl(aq)$, with M the stated Mg, Ca, Cu or Fe(II). For Al and Fe(III), $MCl_3(aq) + 3NaOH(aq) \rightarrow M(OH)_3(s) + 3NaCl(aq)$. These are full symbol equations, suitable for the both-

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tier requirement. Separate HT ionic-equation work may simplify them, but it is not needed to recognise the common-tier formation equations here.

Write actual names and formulae first, then count metal, chlorine, sodium, oxygen and hydrogen on both sides. Keep the precipitate solid and the stated starting solutions aqueous. Use the visible precipitate observation to justify (s), rather than assuming every product is solid. Compare the six formation equations with the identification table. The aluminium excess-alkali dissolution observation is separate; the specification does not require a sodium-aluminate production equation.

■ Checked worked case

Known: $FeCl_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaCl$ conserves Fe 1, Cl 3, Na 3, O 3 and H 3 on each side. Two stated $FeCl_3$ formula units correspond to six NaOH units and two $Fe(OH)_3$ units in this particle ratio. For $FeCl_2$ the analogous hydroxide coefficient is two, not three, because the compound formula is different.

■ Common error

A coefficient before NaOH counts complete formula units; a subscript in $Fe(OH)_3$ gives composition. Do not write $FeOH_3$ or remove sodium/chloride from a full symbol equation without explicitly changing to an ionic representation. These particle counts do not require a mole calculation or imply that any unknown chloride should be handled without the approved school method.

Chemistry-only: identify carbonate through its carbon dioxide

Official-unit focus: 4.8 Chemical analysis

An unknown salt fizzes when dilute acid is added. The bubbles alone are not enough: testing the gas with limewater supplies the required carbon-dioxide evidence.

Carbonates react with dilute acids to release carbon dioxide. The gas turns fresh limewater milky or cloudy, supporting carbonate ions in the tested salt within the stated reference set. Name both stages: add dilute acid to the sample, then transfer or bubble the evolved gas through limewater. A carbonate ion remains part of the starting compound; it is carbon dioxide gas that reaches the limewater.

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Stage	Observation	Inference
Dilute acid	Effervescence	Gas produced
Gas → limewater	Milky / cloudy	CO_2 supported
Combine evidence	Two-stage result	Carbonate supported

For sodium carbonate and hydrochloric acid, $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$ illustrates the source of the gas. The CO_2 test uses aqueous calcium hydroxide and forms insoluble calcium carbonate. Effervescence means gas bubbles, not a unique gas identity. A flame being extinguished is also non-specific: several gases do not support combustion, so it is weaker than the specified limewater result.

RP7's acquired student method uses a small gas transfer with a teat pipette into limewater. The teacher demonstrates this and several transfers may be needed before cloudiness is clear. Use the school-selected dilute acid and small quantities with eye protection, no mouth pipetting and no sealed gas-producing vessel. Record actual effervescence, transfer conditions and limewater change; a failed transfer should be documented as an observation limit, not renamed as another anion.

■ Checked worked case

Known: in the balanced sodium-carbonate particle model, five Na_2CO_3 formula units correspond to five CO_2 molecules and ten HCl units. The carbon count is five on each side. In a supplied reference matrix, carbonate fizzes and clouds limewater while sulfate and the three halides show no change under the same test. This supports identification from reagent-specific evidence rather than salt appearance alone.

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■ Common error

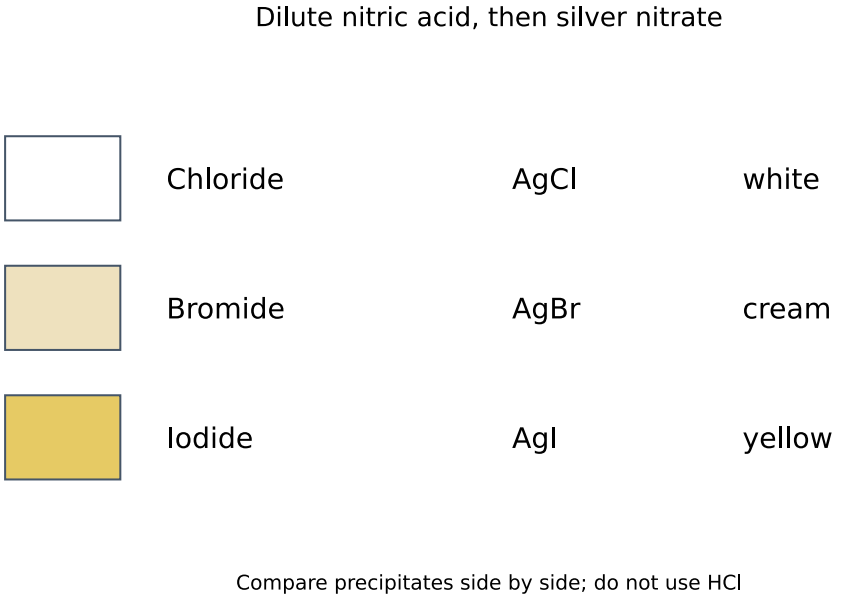
The handbook's optional alternative mentions a lit splint going out; preserve its source text, but do not treat that as unique CO_2 identification. The student method and specification explicitly use limewater. Do not confuse carbonate with carbon dioxide or infer the cation from the acid test. Both-tier Chemistry-only students need the correct two-stage interpretation.

Chemistry-only: acidified silver nitrate distinguishes three halides

Official-unit focus: 4.8 Chemical analysis

Three pale precipitates look similar until they are compared side by side. The acid and precipitating reagent must be correct before a colour can identify the halide.

Halide ions in solution form precipitates with silver nitrate in the presence of dilute nitric acid. Chloride gives white silver chloride, bromide cream silver bromide and iodide yellow silver iodide. State the reagent sequence and observation together: acidify with dilute nitric acid, add silver nitrate, then record the precipitate colour. The solution-ion test differs from damp-litmus bleaching by chlorine gas.



Nitric acid does not introduce a halide ion that would contaminate this test. Using hydrochloric acid instead introduces chloride and can produce a misleading white precipitate with silver nitrate. Acidification also removes carbonate interference in the standard comparison. The school method governs quantities and handling. Silver-halide colours are subtle, so compare known results side by side under the same light

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rather than relying on a tiny colour difference from memory.

Use a fresh clean portion of each solution, selected dilute nitric acid and approved silver nitrate under school supervision. Keep pipettes separate and record observations before assigning identities. Silver salts and acids require the current risk assessment, eye protection and controlled waste collection. Do not mix every reagent into the same unknown tube; previous tests may introduce ions that affect the next test. Additional ammonia-solubility chemistry is not required by this section.

■ Checked worked case

Known: $NaBr + AgNO_3 \rightarrow AgBr + NaNO_3$ illustrates a one-to-one particle ratio for the precipitate. The supplied observation is cream, identifying bromide within the reference set. Four NaBr formula units reacting completely with the stated four $AgNO_3$ units give four AgBr units. These counts explain conservation; they do not turn colour into a measured concentration.

■ Common error

Do not say chloride gives a yellow precipitate, replace nitric acid with hydrochloric acid or confuse precipitation colour with a flame colour. No visible result may reflect concentration or contamination limits and should be recorded honestly. A single anion test cannot identify the unknown's cation. This separate Chemistry section applies on both tiers.

Chemistry-only: acidified barium chloride detects sulfate

Official-unit focus: 4.8 Chemical analysis

A white solid alone does not tell you sulfate is present. You need to know that the sample was tested with the specified acidified barium chloride reagent.

Sulfate ions in solution produce a white precipitate with barium chloride in the presence of dilute hydrochloric acid. The solid is barium sulfate. Record acidification, barium chloride addition and the white precipitate together. The same word white appears in chloride's silver-nitrate test and several hydroxide tests, so colour must be interpreted with reagent identity and conditions.

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Step	Purpose / result
Fresh solution	Avoid previous-test contamination
Dilute HCl	Remove carbonate interference
Barium chloride	White $BaSO_4$ precipitate

Do not acidify with sulfuric acid: it introduces sulfate

Dilute hydrochloric acid helps remove carbonate interference before barium chloride is added. Barium salts can otherwise precipitate with carbonate as well. Do not use sulfuric acid to acidify the sample: it introduces sulfate and could produce a false positive. This requirement differs from the halide test, where nitric acid avoids introducing chloride. The acid choice is therefore part of the chemical evidence, not a interchangeable label.

For actual RP7, take a fresh portion of the solution, add the teacher-selected dilute hydrochloric acid, then the approved barium chloride solution. Use clean pipettes and compare with known reference results. Soluble barium salts require careful handling and controlled waste under the school risk assessment. Record the observation before naming sulfate; avoid contaminating the stock or treating the white material as safe merely because insoluble barium sulfate is the intended product.

■ Checked worked case

Known: $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$ illustrates one sulfate unit forming one barium-sulfate unit. Three Na_2SO_4 formula units give three $BaSO_4$ units and six NaCl units in the supplied particle model. The sulphur count remains three on each side. A matching white result with these reagents supports sulfate within the stated reference set; it does not give an anion concentration.

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■ Common error

Do not identify sulfate from a white precipitate produced by silver nitrate, or use sulfuric acid before barium chloride. A carbonate gas test and an acidified sulfate precipitation test use different observations. Negative or uncertain results remain qualified by sampling and detection limits. Both tiers learn the specific reagent/result pair; no extra advanced barium chemistry is imposed.

Chemistry-only: accuracy, sensitivity and speed of instruments

Official-unit focus: 4.8 Chemical analysis

An instrument can detect a trace too small to give an obvious visible precipitate. That advantage is sensitivity, while closeness to a reference concentration concerns accuracy.

Instrumental methods detect and identify elements and compounds using measured signals. The specification's advantages are accuracy, sensitivity and rapid measurement. Accuracy means agreement with the accepted or true value in the stated comparison; sensitivity here means detecting small amounts or concentrations; rapid means obtaining results quickly. A test's ability to give a repeatable colour is useful, but does not automatically quantify a trace concentration.

Original data: three distinct advantages

Feature	Instrument A	Comparison
Reading (reference 10)	9.8	8.9
Detection threshold	0.1 unit	2 units
Measurement time	5 s	60 s

Compare an instrument and a chemical test against the same task. Detecting a small metal-ion concentration is different from identifying a high-concentration salt by flame colour. Instrumental outputs can be numerical and easier to compare objectively than a subtle cream/white judgement. Instruments still need appropriate references, calibration, clean samples and correct operation. A fast result can be wrong if contami-

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nation or an unsuitable reference invalidates the method.

Use a supplied table of reference concentrations, measured values, minimum detectable concentrations and run times. Link each advantage to the relevant column rather than repeating three adjectives with no evidence. Consider school access, preparation and cost as practical limitations when information is provided. This section asks for advantages over the stated chemical tests; it does not require operating an unacquired advanced instrument or memorising unrelated spectroscopic mechanisms.

■ Checked worked case

Known: the supplied reference is 10.0 concentration units. Instrument A reads 9.8 and B reads 8.9, with absolute errors 0.2 and 1.1 units. A is closer in this comparison. If A detects 0.1 unit and the visible test detects 2 units, A has the lower detection threshold. If measurement times are 5 and 60 s, the stated time reduction is 60−5=55 s. These original data illustrate distinct advantages, not specifications of a commercial instrument.

■ Common error

Do not call a lower detection threshold greater accuracy without comparison to a known value. Repeated readings can be precise but systematically wrong. A negative result below a detection threshold does not prove zero analyte. Instrumental advantages do not erase the need for supervised practical work or justify inventing a qualification performance exam.

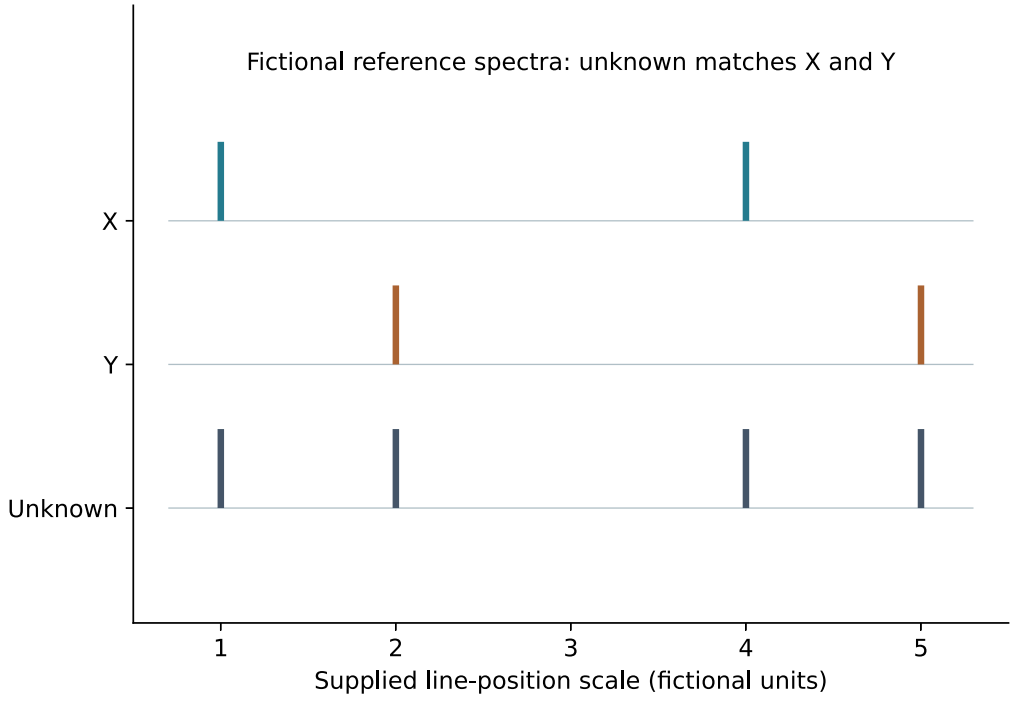
Chemistry-only: match emission lines and use calibration data

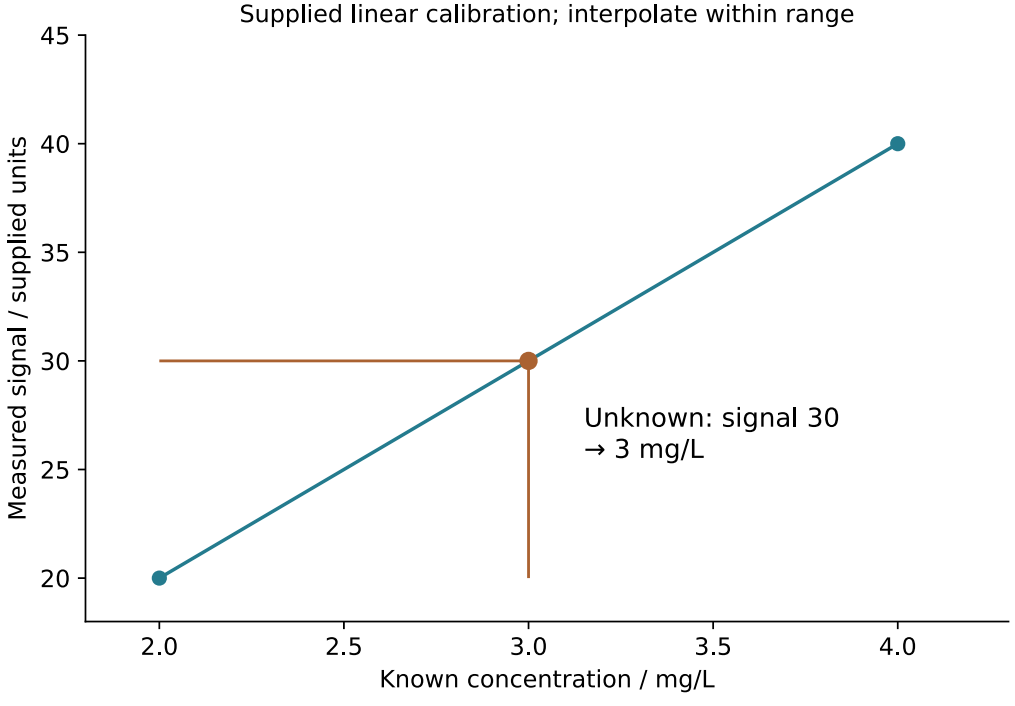
Official-unit focus: 4.8 Chemical analysis

A mixed flame can hide a visible colour, but an instrument separates the emitted light into lines. Comparing line positions with references can reveal more than one metal ion.

In flame emission spectroscopy, the sample is put into a flame and emitted light passes through a spectroscope. Its line-spectrum output is compared with a reference set to identify metal ions in solution. The positions of several matching lines support an identity. An unknown may contain lines from more than one reference. The method can also measure concentration using calibrated signal data for a selected ion under controlled conditions.

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Separate line position from signal strength. Position supports identity; intensity at a suitable line, compared with known concentrations, supports quantity. Use the supplied wavelength scale or reference table, not a memorised list of invented atomic wavelengths. The original figure's reference X/Y positions are explicitly fictional teaching data. Check several lines and overlapping signals; one shared line alone can leave

alternatives. No electron-transition or advanced instrument mechanism is required here.

For a supplied calibration, plot the measured reference signal against known concentration and read an unknown within the calibrated range. Use the same sample preparation, instrument conditions and background correction described by the data. Avoid extrapolation beyond the references without evidence. A teacher-approved handheld spectroscope observation can show line structure; it does not by itself supply quantitative concentration calibration or replace the RP7 chemical-test experience.

■ Checked worked case

Known: fictional reference X has lines at positions 1 and 4, while Y has lines at 2 and 5 on the supplied position scale. The unknown shows all four, supporting X and Y. Separately, a provided calibration for one selected ion gives signals 20 and 40 at concentrations 2 and 4 mg/L, with linear zero-background relation in that interval. Signal 30 therefore gives 3 mg/L. These are supplied-model results, not measured atomic wavelengths or a universal intensity law.

■ Common error

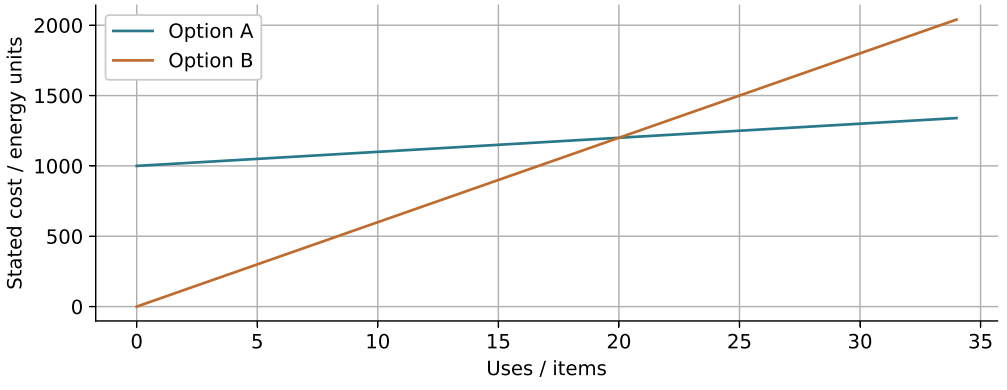
Do not infer concentration from the horizontal line position, or identify a spectrum by total brightness alone. A matching line pattern supports identity only within the supplied references and resolution. Record uncertain or unexplained lines. The current section limits interpretation to flame emission with appropriate reference data; it does not demand an advanced mass-spectrum or infrared-spectrum course.

Atmosphere, resources and life-cycle decisions

Official-unit focus: 4.9 Chemistry of the atmosphere; 4.10 Using resources

A reusable container can require more energy to manufacture than a single-use one. Its impact depends on how often it is used and what happens at disposal.

A life-cycle assessment considers raw materials, manufacture, transport, use and end-of-life processes. Greenhouse gases absorb and emit infrared radiation; pollution and climate effects are related but distinct questions.



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Define the functional unit before comparing products. The same delivered service, such as carrying one litre of water a hundred times, is fairer than comparing one object with another regardless of lifetime.

List system boundaries, energy sources and assumptions. Compare water demand, emissions and waste separately before making a judgement. Explain whose priorities affect the decision and where the data are uncertain.

■ Checked worked case

Known: a reusable item requires 1,000 energy units to make and 10 per use; a disposable item requires 60 per use. Equality occurs when 1,000 + 10n = 60n. Rearranging gives 1,000 = 50n, so n = 20 uses under this simplified model.

■ Common error

A break-even model is sensitive to cleaning, transport and disposal assumptions. Recyclable does not guarantee that an item is actually collected and recycled.

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