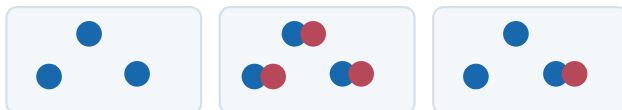


C1 Atomic structure and the periodic table

Chemistry Paper 1 | AQA Combined Science Trilogy | Key ideas to remember

Atoms are conserved

1 Elements, compounds and mixtures



Element

Compound

Mixture

Element: one type of atom.

Compound: different elements chemically bonded in fixed proportions.

Mixture: substances not chemically bonded together; separate by physical methods.

Formula: $\text{H}_2\text{O} = 2 \text{ H atoms} + 1 \text{ O atom}$.

Balanced: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$.

Same number of each atom on both sides.
Change coefficients, **never subscripts**.

Protons identify elements

2 Particles, numbers and isotopes

Particle	Charge	Rel. mass	Location
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus
Electron	-1	Very small	Shells

Atomic number = protons.

Mass number = protons + neutrons.

Neutral atom: electrons = protons.

Na-23: atomic no. 11; **11 p, 12 n, 11 e.**

Isotopes: same protons, different neutrons.

Relative atomic mass: weighted mean.

Chlorine (75% Cl-35, 25% Cl-37):

$(75 \times 35 + 25 \times 37) / 100 = 35.5$.

Groups show patterns

3 Models and electron shells

Dalton: solid spheres.

Thomson: electrons; plum pudding model.

Rutherford: tiny positive nucleus.

Alpha scattering: most pass through → mostly empty space; a few turn sharply → small, dense, positive nucleus.

Bohr: electrons in shells/energy levels.

Chadwick: evidence for neutrons.

First 20 elements: fill inner shells first, using the pattern **2, 8, 8, 2**.

Sodium has 11 electrons: **2, 8, 1**.

Atom radius $\approx 0.1 \text{ nm}$ (10^{-10} m); nucleus much smaller.

4 Separating mixtures

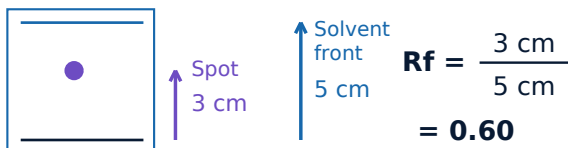
Filtration: insoluble solid from liquid.

Crystallisation: soluble solid from solution.

Simple distillation: collect the solvent; evaporate and condense.

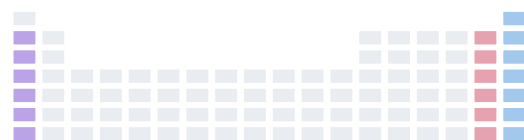
Fractional distillation: liquids with different boiling points.

Chromatography: separates substances by distance travelled.



Both distances start at the **pencil baseline**. Keep the starting solvent level below this line.

5 How the periodic table works



1 Alkali metals

7 Halogens

0 Noble gases

Elements are ordered by **atomic number**.

Period: row; number of occupied shells.

Group: similar properties; same outer electrons (main groups).

Metals: left/centre; conduct, malleable.

Non-metals: right side + hydrogen.

Mendeleev: used mass/properties; left gaps and predicted missing elements.

6 Groups 1, 7 and 0

GROUP 1 - alkali metals

1 outer electron; soft metals.

With water → metal hydroxide + hydrogen.

Down group: more reactive; lower melting points.

GROUP 7 - halogens

7 outer electrons; diatomic, e.g. Cl_2 .

Down group: less reactive; higher melting/boiling points. A more reactive halogen displaces a less reactive one from its halide solution.

GROUP 0 - noble gases

Full outer shell (He has 2); unreactive, single atoms. Boiling points rise down the group.

CHEMISTRY 2 | Structure and bonding

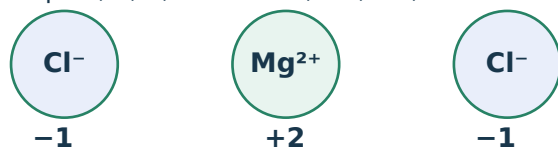
AQA Combined Science Trilogy • Paper 1 • Core knowledge + a labelled Higher Tier reminder

1 Ionic: transfer, then attraction

Metal loses electrons → positive ion.

Non-metal gains → negative ion.

Groups 1, 2, 6, 7 form +1, +2, -2, -1 ions.



Formulae balance charges: NaCl, MgO, MgCl_2 , Na_2O . Use the simplest ratio.

Giant lattice: strong electrostatic attraction between opposite ions in all directions.

High melting point: much energy needed to overcome these attractions.

Conducts molten or dissolved: ions move and carry charge. **Not solid:** ions are fixed. Ion diagrams: full outer shells, brackets, charges.

2 Covalent: shared electron pairs

Covalent bond = shared electron pair attracted to both nuclei. Non-metals share.



Bonds: H 1, O 2, N 3, C 4 in key molecules.

O=O: 2 shared pairs. **N≡N:** 3 shared pairs.

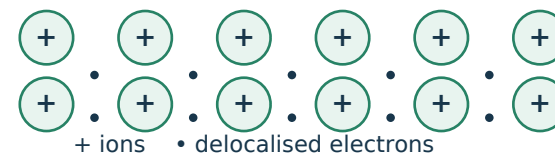
Small molecules: strong bonds inside; weak **intermolecular forces** between molecules.

Low melting/boiling points: little energy overcomes forces **between** molecules. Covalent bonds stay intact.

No conduction: no mobile charged particles. Larger molecules usually have stronger intermolecular forces → higher boiling points.

3 Metals and alloys

Metallic bonding: strong attraction between positive metal ions and **delocalised electrons**.



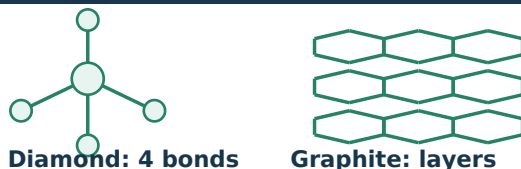
Conduct: delocalised electrons move to carry charge and transfer energy (heat).

Usually high melting points: strong attractions need much energy to overcome.

Pure metal: regular layers can slide, so it can bend and be shaped (**malleable**).

Alloy: metal mixed with other elements. Different-sized atoms distort layers → layers slide less easily → usually **harder**.

4 Giant covalent and carbon sheets



Diamond: 4 bonds per carbon in a rigid 3D network → **hard**; useful for cutting tools. No mobile electrons → **does not conduct**.

Graphite: 3 bonds per carbon in hexagonal sheets. Weak interlayer forces → layers **slide**; soft. One delocalised electron per carbon → **conducts**.

Both: high melting points; many strong covalent bonds need much energy to break.

Silica, SiO_2 : giant covalent; high melting point; does not conduct.

Graphene: one atom-thick sheet; strong and conducting. Used in electronics.

5 Polymers, fullerenes and nanotubes

Polymers: long molecules with strong covalent bonds **along** chains. Relatively strong forces **between** chains → usually solid at room temperature.



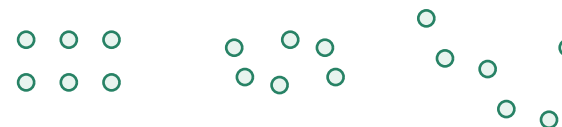
Poly(ethene): n means many repeats

Bonds pass through brackets: chain continues. Each repeat shown contains **2 C and 4 H**.

Fullerenes: hollow carbon molecules. C_{60} : 60 carbons in a spherical cage. Strong covalent bonds inside; weaker forces between separate cages.

Nanotubes: hollow cylinders, long compared to their diameter. Strong covalent bonds → **high tensile strength** → reinforce composites. Suitable nanotubes conduct → electronics.

6 States, symbols and model limits



Solid

Liquid

Gas

Solid: close, fixed positions; vibrate.

Liquid: close, irregular; move past each other.

Gas: far apart; move randomly.

Melting: solid → liquid. **Freezing:** reverse.

Boiling: liquid → gas. **Condensing:** reverse.

Below melting point: **solid**. Between melting/boiling points: **liquid**. Above boiling point: **gas**.

(s) solid; (l) liquid; (g) gas;

(aq) dissolved in water, not just any liquid.

Higher Tier: simple particle models treat particles as solid inelastic spheres and omit forces.

CHEMISTRY 3 | Quantitative chemistry

AQA Combined Science Trilogy • Core knowledge first • Purple panels are Higher Tier

1 Conservation and balancing

No atoms are created or destroyed.

Total reactant mass = total product mass.

Include **all gases**. In a closed system, matter cannot enter or leave.

Missing mass: 12 g carbon + 32 g oxygen = **44 g carbon dioxide**. Add or subtract.



H: 4 each side. O: 2 each side.

Change **front coefficients**, never subscripts.
A coefficient multiplies the **whole formula**.

3H₂O: 6 H and 3 O atoms.

Check every element; use the simplest ratio.

2 Formula mass and percentages

M_r = sum of A_r values for every atom.

Relative formula mass has **no unit**.

Use the A_r values given in the question.

H₂O: (2 × 1) + 16 = **18**.

Ca(OH)₂: 40 + 2 × (16 + 1) = **74**.

The bracket 2 multiplies both O and H.

% = element contribution / M_r × 100

Count **all atoms** of the chosen element.

Use masses, not the fraction of atoms.

Ca: 40

Total M_r = 100

Calcium in CaCO₃: 40/100 × 100 = **40%**.

Oxygen in H₂O: 16/18 × 100 = **88.9%**.

Keep extra digits; round the final answer.

5 Higher Tier: moles and masses

1 mol = 6.02 × 10²³ stated particles.

Molar mass (g/mol) is numerically equal to M_r.

n = mass / molar mass

Mass = n × molar mass.

9 g H₂O: 9/18 = **0.50 mol**.

Molecules = 0.50 × 6.02 × 10²³

= **3.01 × 10²³ water molecules.**

Mass → moles → ratio → mass



From 12 g Mg with excess oxygen:

1. Mg = 12/24 = **0.50 mol**.

2. Mg:MgO = 1:1 → **0.50 mol MgO**.

3. MgO mass = 0.50 × 40 = **20 g**.

Coefficients give **mole ratios, not gram ratios**.

Convert mg or kg to **g** before using molar mass.

3 Gases and measurement uncertainty

Solid gains mass: gas enters and reacts.



2.4 g Mg → 4.0 g MgO: **1.6 g O₂ gained**.

Solid loses mass: a product gas escapes.



10.0 g → 5.6 g solid: **4.4 g CO₂ released**.

Mass is conserved when the gas is included.

Readings: 4.8, 5.0, 5.2 g

Mean = (4.8 + 5.0 + 5.2)/3 = **5.0 g**.

Range = 5.2 – 4.8 = **0.4 g**.

If using **half the range**, uncertainty = ±0.2 g.

Report **5.0 ± 0.2 g**. Follow the stated method.

Precise: readings close together.

Accurate: close to the true value.

Repeats do not remove a systematic error.

4 Concentration in g/dm³

Concentration = mass / volume

c = m/V · solute mass in g; volume in dm³.

Use the **final solution volume**.

1000 cm³ = 1 dm³

cm³ → dm³: divide by 1000.

250 cm³ = 0.250 dm³; 500 cm³ = 0.500 dm³.

Example: 10 g in 500 cm³ of solution.

c = 10/0.500 = **20 g/dm³**.

Mass = concentration × volume

For 250 cm³ at 20 g/dm³:

m = 20 × 0.250 = **5.0 g**.

Volume = mass / concentration.

HT: fixed volume, double mass → double c.

Fixed mass, double volume → half c.

6 Higher Tier: ratios and limits

Find coefficients from masses:

Convert every mass to moles; divide by the smallest amount; scale to whole numbers.

1:0.5:1 → 2:1:2. Check the atom counts.

1:1.5:1 → 2:3:2; do not round 1.5 to 2.

Limiting reactant: used up first, limiting product.

Compare **available moles / coefficient**.

The smaller result limits the reaction.



12 g Mg: 12/24 = 0.50 mol; ÷2 = **0.25**.

4 g O₂: 4/32 = 0.125 mol; ÷1 = **0.125**.

Oxygen limits. MgO = 2 × 0.125 = 0.25 mol.

Maximum MgO mass = 0.25 × 40 = **10 g**.

Mg used = 0.25 × 24 = 6 g; **6 g Mg remains**.

Adding more of the excess reactant alone cannot make more product.

CHEMISTRY 4 | Chemical changes

AQA Combined Science Trilogy • Core knowledge first • Purple panels are Higher Tier

1 Reactivity and extraction

More reactive → less reactive

K	Na	Li	Ca	Mg	Al	C
Zn	Fe	H	Cu	Ag	Au	

C and H are non-metal reference points.

More reactive metals form positive ions more readily; they displace less reactive metals.



Below carbon: reduce the oxide with carbon.

Above carbon: e.g. aluminium, use electrolysis.

Gold: may occur as the metal itself.

Oxidation: gain O. Reduction: lose O.



CuO is reduced; carbon is oxidised.

2 Acids, salts and pH

Acid + suitable metal → salt + hydrogen

Acid + base / alkali → salt + water

Acid + carbonate → salt + water + CO₂

HCl → **chlorides**; H₂SO₄ → **sulfates**;
HNO₃ → **nitrates** (reactions with bases).

First salt name comes from the metal ion.

Neutral salt formula: balance ion charges.



Acid < 7

Neutral 7

Alkali > 7



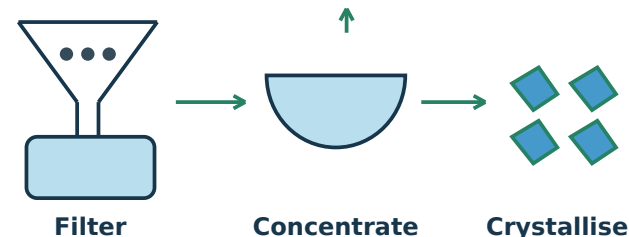
Universal indicator estimates pH; a probe gives a numerical reading. An **alkali is a soluble base**.

CO₂ test: limewater turns milky.

3 Required practical: soluble salt

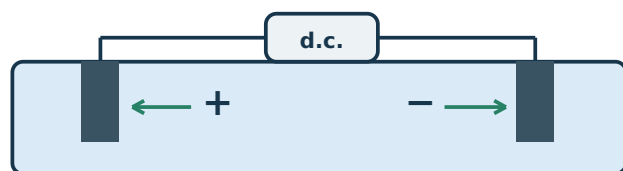


Make pure, dry copper sulfate crystals.



1. Warm dilute acid gently; stop heating.
2. Add **excess** oxide to use up all the acid.
3. Filter off excess oxide; keep the solution.
4. Gently evaporate **some** water, then cool.
5. Filter crystals; rinse with a little cold distilled water. Blot dry. Wear eye protection.

4 Electrolysis: predict the products



Cathode (-)

Anode (+)

Molten or aqueous: mobile ions carry charge.

Solid: ions fixed, so does not conduct.

Molten binary: metal at cathode;
non-metal at anode.

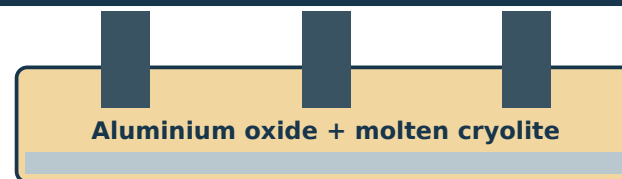
Aqueous, inert electrodes:

Cathode: metal above H → hydrogen;
metal below H → the metal.

Anode: halide → halogen; otherwise oxygen.

CuSO₄(aq): Cu at (-); O₂ at (+).

5 Aluminium and practical evidence



Al collects at the negative cathode

Cryolite lowers the melting temperature, so less heating energy is needed.

Oxygen reacts with the carbon anodes → **CO₂**.
Carbon is used up: **replace the anodes**.

Practical: graphite electrodes, low-voltage d.c.;
test each product. Control conditions; repeat.
Eye protection; ventilation for chlorine.

H₂: lit splint → squeaky pop.

O₂: glowing splint relights.

Cl₂: bleaches damp litmus white.

6 Higher Tier: redox and acids

OIL RIG: Oxidation Is Loss of electrons;
Reduction Is Gain of electrons.

Cathode = reduction; anode = oxidation.

Displacement: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$



Balance **atoms and total charge**.

Strong acid: fully ionised in water.

Weak acid: partly ionised.

Concentration: amount of acid per volume.

Strong can be dilute; weak can be concentrated.

pH 4 → pH 3 → pH 2

Each pH step down: **H⁺ concentration ×10**.
pH 2 has **100×** the H⁺ concentration of pH 4.

CHEMISTRY 5 | Energy changes

AQA Combined Science Trilogy • Core knowledge first • Purple panels are Higher Tier

1 Energy in or energy out?



Exothermic: transfers energy **to** the surroundings. Products have **less** chemical energy than reactants.

Endothermic: takes energy **from** the surroundings. Products have **more** chemical energy than reactants.

Energy is conserved: transferred, not created or destroyed. EXO = exits; ENDO = enters.

2 Examples and temperature change

Exothermic: combustion, many oxidation reactions, neutralisation.

Uses: hand warmers; self-heating cans.

Endothermic: thermal decomposition; citric acid + sodium hydrogencarbonate.

Uses: some instant cooling / injury packs.

$$20\text{ }^{\circ}\text{C} \rightarrow 31\text{ }^{\circ}\text{C}$$

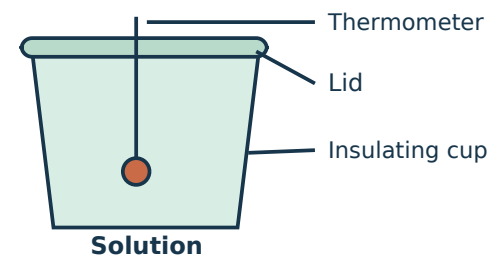
$$\text{Rise} = 31 - 20 = 11\text{ }^{\circ}\text{C}$$

Change = highest or lowest – start.

A rise is evidence for an exothermic reaction; a fall is evidence for an endothermic reaction.

Compare packs using the data: temperature, duration, cost per use and waste. Hottest is not always best.

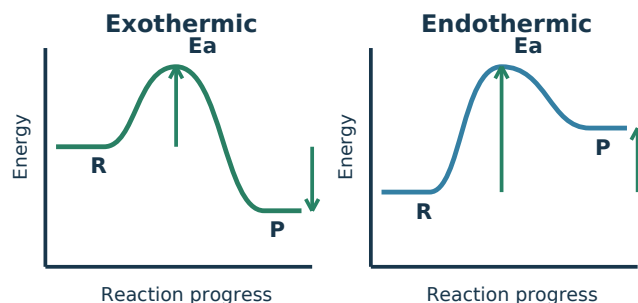
3 Required practical 10



1. Wear eye protection; support cup in beaker.
2. Measure fixed solution volumes. Bring both to the same start temperature; record it.
3. Mix, fit lid, stir; read temperature often.
4. Record the maximum (or minimum); subtract the start temperature.

Change: acid concentration. **Control:** volumes, alkali concentration, start temperature, apparatus. **Repeat:** calculate means.

4 Reaction profiles



Exothermic: P lower. Endothermic: P higher.

Activation energy (E_a): minimum energy needed for reacting collisions. Arrow: **reactants up to the peak**.

Overall change: between R and P levels, not from the peak. The x-axis is **not time**.

5 Higher Tier: bonds and energy

BREAK bonds → energy IN

FORM bonds → energy OUT

Overall energy change = breaking – forming.

Negative → **exothermic**.

Positive → **endothermic**. Use **kJ/mol**.

Exothermic: forming releases more than breaking needs. **Endothermic:** breaking needs more than forming releases.

Count all bonds, including coefficients.

$2\text{H}_2\text{O}$ has **4 O-H** bonds.

2O_2 has **2 O=O** double bonds. Use the supplied double-bond energy once for each O=O.

6 Higher Tier: worked calculation



Bond energies supplied in kJ/mol:

H-H = 436; Cl-Cl = 243; H-Cl = 431.



BREAK: 1 H-H and 1 Cl-Cl
 $436 + 243 = 679\text{ kJ/mol}$.

FORM: 2 H-Cl
 $2 \times 431 = 862\text{ kJ/mol}$.

Overall: $679 - 862 = -183\text{ kJ/mol}$.
Negative → **exothermic**.

For 1 mol of reaction as written, making 2 mol HCl.