

Edexcel GCSE Chemistry — Chemistry

Topic 3

Chemical changes

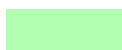
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Free edition

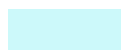
Revision notes for one topic, free to download.

These are the summary notes only. The full edition of this booklet also carries the worked examples, the required practicals and the formula sheets, with every topic in one place.

Colour Codes



Definitions



Diagrams



Notes and warnings



Sub-topic headings

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Define Alkali

- a soluble base; in solution, a source of hydroxide ions, OH^- .

Define Indicator

- a substance that changes colour with acidity or alkalinity.

Define Titration

- adding one solution from a burette to a measured volume of another until the end point, to find the volumes that react exactly.

Define Precipitate

- an insoluble solid formed when two solutions are mixed.

Define Electrolysis

- the process in which electrical energy, from a d.c. supply, decomposes electrolytes.

Define Inert electrode

- graphite or platinum; it does not react during electrolysis.

HIGHER TIER

Define HT ONLY

- loss of electrons; **reduction**

Define Acid

- in solution, a source of hydrogen ions, H^+ .

Define Base

- any substance that reacts with an acid to form a salt and water only.

Define Salt

- the compound formed when an acid's hydrogen is replaced by a metal or ammonium ion.

Define Neutralisation

- a reaction between an acid and a base.

Define Electrolyte

- an ionic compound, molten or dissolved in water, decomposed by an electric current.

Define Cation / anion

- a positively / negatively charged ion, migrating to the negative **cathode** / positive **anode**.

Oxidation

HIGHER TIER

Define HT ONLY

- gain of electrons.

3.1 Acids (3.1–3.21)

Acids, alkalis and the pH scale (3.1–3.5)

- Acids in solution are sources of hydrogen ions (H^+); alkalis in solution are sources of hydroxide ions (OH^-) (3.1):
 $\text{HCl(aq)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$; $\text{NaOH(aq)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$.
- A neutral solution has a pH of 7; acidic solutions have lower pH values and alkaline solutions higher pH values (3.2), on a scale of 0–14.

Indicators (3.3). Learn all three:

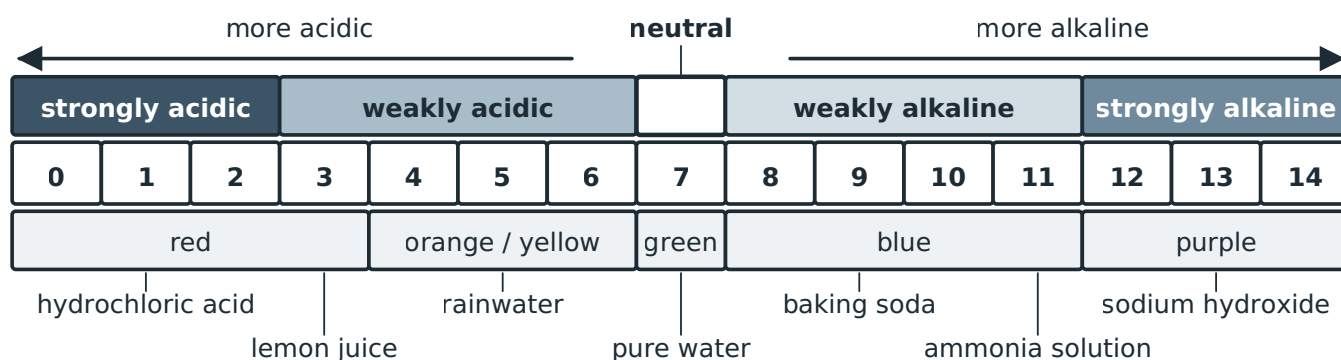
Indicator	In acid	In neutral	In alkali
Litmus	Red	Purple	Blue
Methyl orange	Red	Orange	Yellow
Phenolphthalein	Colourless	Colourless	Pink

HIGHER TIER

Universal indicator and a pH meter give the actual pH; the three above only show which side of neutral you are on — which is what makes them good for titration end points.

HT ONLY Recall that the higher the concentration of hydrogen ions in an acidic solution, the lower the pH; and the higher the concentration of hydroxide ions in an alkaline solution, the higher the pH (3.4). **HT ONLY** Recall that as hydrogen ion concentration increases by a factor of 10, the pH decreases by 1 (3.5).

The pH scale and the colours of universal indicator

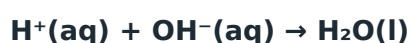


Universal indicator (a wide range indicator) gives these colours. A pH probe is more accurate.

Acids and alkalis

Acids produce H^+ ions in aqueous solution.
Alkalis produce OH^- ions in solution.
pH 7 is neutral, below 7 acidic, above 7 alkaline.

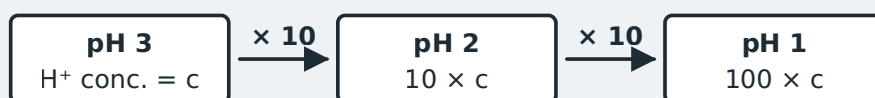
Neutralisation



Hydrogen ions react with hydroxide ions to make water. acid + alkali $\rightarrow$ salt + water

HT only — pH and hydrogen ion concentration

As the pH decreases by one unit, the concentration of hydrogen ions increases by a factor of 10.



A strong acid is fully ionised, so at the same concentration it has a lower pH than a weak acid.

HIGHER TIER

Dilute and concentrated, weak and strong (3.7–3.8)

HT ONLY Explain the terms dilute and concentrated, with respect to amount of substances in solution (3.7) — how much solute is dissolved in a given volume: concentrated = a lot per dm^3 , dilute = little. **HT ONLY** Explain the terms weak and strong acids, with respect to the degree of dissociation into ions (3.8) — how far the acid dissociates into ions in solution:

- A **strong acid** (hydrochloric, nitric, sulfuric) is **completely dissociated**: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$.
- A **weak acid** (ethanoic, citric, carbonic) is only **partially dissociated**, most staying as molecules: $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$.
- At the **same concentration** a strong acid has a **higher H^+ concentration**, so a **lower pH**, and reacts faster.

WATCH OUT

"Strong" is not "concentrated". Concentration is *how much* acid is in the water; strength is *how much of it has split into ions*. A concentrated weak acid and a dilute strong acid can have the same pH.

Bases, alkalis and neutralisation (3.9–3.10, 3.13–3.14)

- A **base** is any substance that reacts with an acid to form a salt and water only (3.9) — metal oxides and hydroxides.
- **Alkalis are soluble bases** (3.10) — sodium hydroxide, potassium hydroxide, ammonia solution. All alkalis are bases; not all bases are alkalis, as most metal oxides and hydroxides are insoluble.
- A **neutralisation reaction** is a reaction between an acid and a base (3.13). In acid–alkali neutralisation, hydrogen ions (H^+) from the acid react with hydroxide ions (OH^-) from the alkali to form water (3.14):
$$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$$

ANALOGY

Neutralisation as cancelling out. Treat H^+ ions as +1 counters and OH^- as -1 counters: each pair that meets cancels to zero, giving water. Match them exactly and nothing is left over (pH 7); too few -1 counters and the pile stays positive, too many and it goes negative. The other two ions are spectators — they become the salt.

The reactions of acids and the salts they make (3.11–3.12)

Explain the general reactions of aqueous solutions of acids with metals, metal oxides, metal hydroxides and metal carbonates to produce salts (3.11):

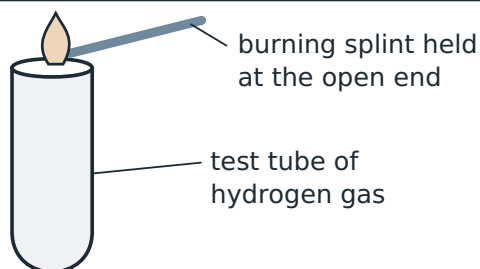
Reaction	Products	Example
a acid + metal	salt + hydrogen	$\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$
b acid + metal oxide	salt + water	$\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$
c acid + metal hydroxide	salt + water	$\text{NaOH} + \text{HNO}_3 \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}$
d acid + metal carbonate	salt + water + carbon dioxide	$\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$

The **acid decides the salt**: hydrochloric acid (HCl) makes **chlorides**, sulfuric (H_2SO_4) **sulfates**, nitric (HNO_3) **nitrates**.

Chemical tests for the gases (3.12). a Hydrogen — a lighted splint burns it with a **squeaky pop**. b Carbon dioxide — bubbled through **limewater**, it **turns milky**.

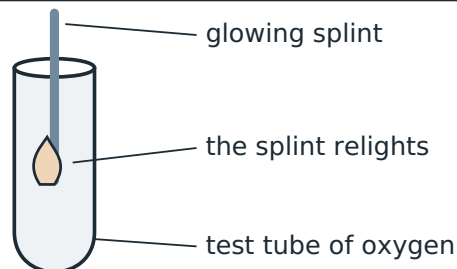
Chemical tests for hydrogen, oxygen and carbon dioxide

Test for hydrogen



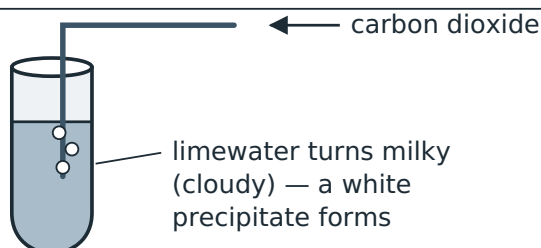
Hydrogen burns rapidly with a 'pop' sound — a squeaky pop.

Test for oxygen



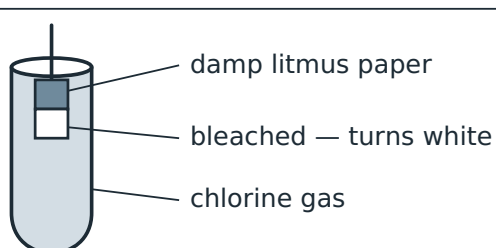
A glowing splint inserted into the tube relights in oxygen.

Test for carbon dioxide



limewater = calcium hydroxide solution
Bubble the gas through limewater, or shake the gas with limewater.

Test for chlorine



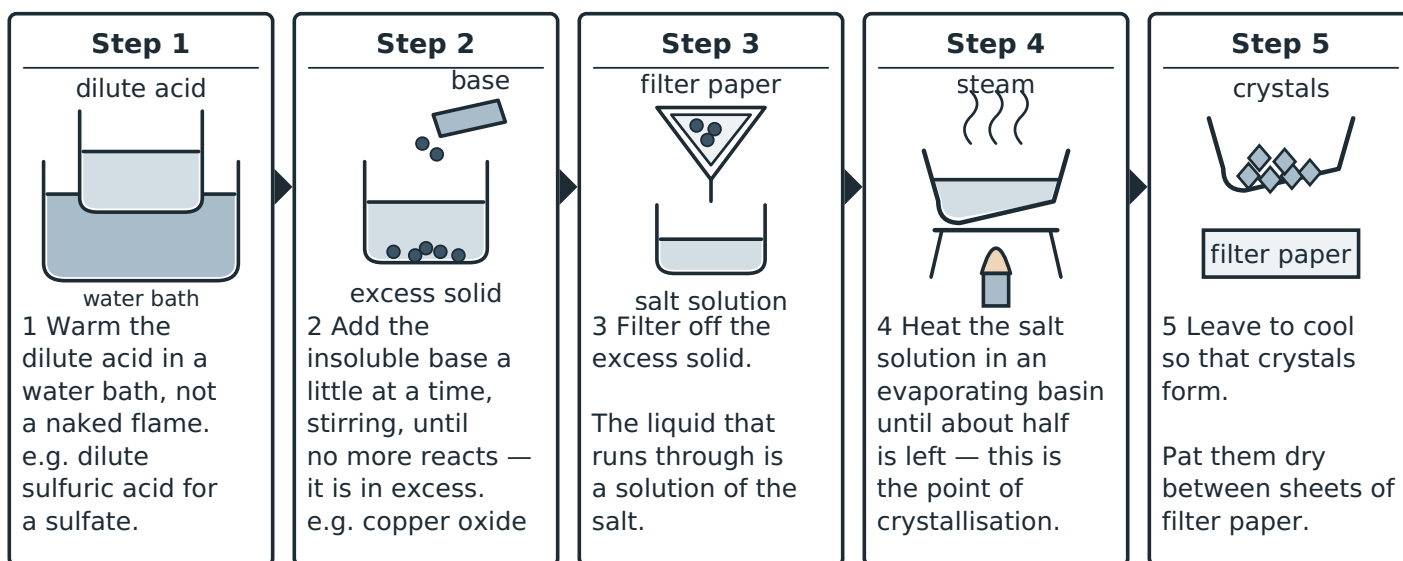
Damp litmus paper is bleached and turns white in chlorine gas.

Making soluble salts (3.15–3.16)

From an acid and an insoluble reactant (3.15) — a metal oxide, hydroxide or carbonate. Explain why:

- **a Excess of the reactant is added**, so **all** the acid is used up.
- **b The excess reactant is removed by filtration** — insoluble, it stays on the filter paper.
- **c The solution remaining is only salt and water**, so crystallising gives a pure salt.

Preparing a pure, dry sample of a soluble salt from an insoluble base



Core Practical 3.17 a pure, dry sample of a soluble salt from an insoluble oxide or carbonate.
copper oxide + sulfuric acid → copper sulfate + water Use a water bath to evaporate.

From an acid and a soluble reactant (3.16) — an alkali such as sodium hydroxide. Explain why:

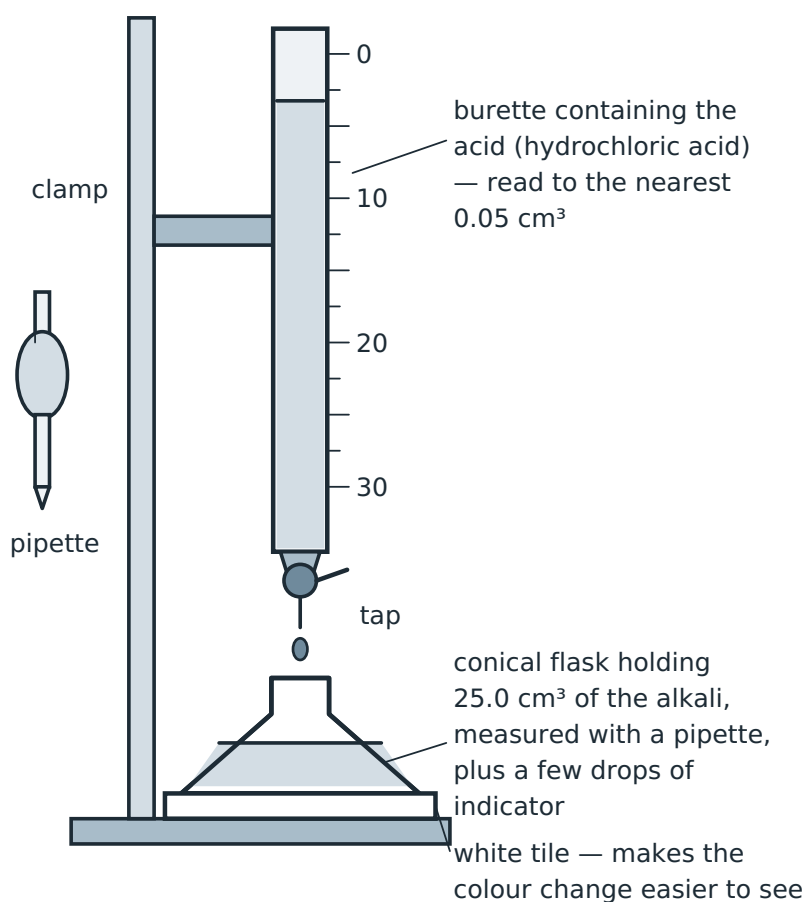
- **a Titration must be used**: excess soluble reactant cannot be seen or filtered off.
- **b The acid and the soluble reactant are then mixed in the correct proportions** found by titration, repeated without indicator.
- **c The solution remaining, after reaction, is only salt and water**, which is crystallised.

Titration (3.18)

Describe how to carry out an acid-alkali titration, using burette, pipette and a suitable indicator, to prepare a pure, dry salt (3.18).

1. Use a **pipette** and filler to transfer an exact volume (e.g. 25.0 cm^3) of alkali into a **conical flask**.
2. Add a few drops of a **suitable indicator** — phenolphthalein or methyl orange. **Universal indicator will not do**: its gradual colour change gives no sharp end point.
3. Fill a **burette** with acid using a funnel, below eye level; remove the funnel and record the **initial reading** from the bottom of the meniscus, to 0.05 cm^3 .
4. Stand the flask on a **white tile**; add acid while **swirling**.
5. Near the end point add **dropwise** until the indicator just changes colour permanently. Record the **final reading**; titre = final – initial.
6. Repeat for **concordant titres** (within 0.10 cm^3); take a **mean of the concordant results only**.
7. To make the salt, repeat **without indicator** using the mean titre, then **crystallise**.

Titration apparatus and method



Method

- 1 Use a pipette and pipette filler to put 25.0 cm^3 of alkali into a conical flask.
- 2 Add a few drops of indicator (methyl orange or phenolphthalein).
- 3 Fill the burette with the acid and record the starting reading.
- 4 Add the acid a little at a time, swirling the flask, until the indicator just changes colour — this is the end point.
- 5 Record the final reading:
titre = final reading – starting reading
- 6 Repeat until you have concordant titres.
- 7 Work out the mean of the concordant titres only.

Concordant titres

Titres within 0.10 cm^3 of each other. Ignore the rough titre and any anomalies.

Points to remember

Use a strong acid with a strong alkali (hydrochloric, sulfuric or nitric acid). Universal indicator is not used — its colour changes gradually, not sharply.

HIGHER TIER

NOTE

Where the calculations sit. Statement 3.18 is the *method* only. Concentration in g dm^{-3} is **1.49** (Topic 1); mol dm^{-3} (**5.8C**), the titration core practical (**5.9C**) and titration calculations (**5.10C**) are

CHEMISTRY ONLY, with 5.8C and 5.10C also

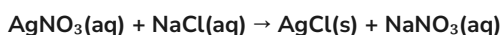
HT ONLY. They are set out here because they go Topic 5 — with the technique.

Solubility rules and insoluble salts (3.19–3.21)

Recall the general rules which describe the solubility of common types of substances in water (3.19).

Rule	Solubility	Exceptions
a Common sodium, potassium and ammonium salts	Soluble	None
b All nitrates	Soluble	None
c Common chlorides	Soluble	silver, lead
d Common sulfates	Soluble	lead, barium, calcium
e Common carbonates, hydroxides	Insoluble	sodium, potassium, ammonium soluble

Predicting a precipitate (3.20). Swap the partners of the two solutions and check each new pair against the rules; the **insoluble** pair is the precipitate. Silver nitrate + sodium chloride gives **silver chloride** (insoluble, rule c) and sodium nitrate:



Preparing a pure, dry sample of an insoluble salt (3.21) — precipitation:

1. Choose **two soluble salts** containing the ions you need — for lead(II) sulfate, lead nitrate and sodium sulfate solutions.
2. Mix them in a beaker and stir; the insoluble salt forms as a **precipitate**.
3. **Filter**; the precipitate is the **residue**.
4. **Wash** the residue with **distilled water** to remove soluble salt solution.
5. **Dry** it in a warm oven or between filter papers.

Suggested practicals — neutralising acids with metal oxides, hydroxides and carbonates; testing for hydrogen and carbon dioxide; preparing an insoluble salt by precipitation.

3.2 Electrolytic processes (3.22–3.31)

Electrolytes and electrolysis (3.22–3.24)

- **Electrolytes are ionic compounds in the molten state or dissolved in water (3.22).** Solids do **not** conduct — the ions are fixed in the lattice; melting or dissolving frees them.
- **Electrolysis is a process in which electrical energy, from a direct current supply, decomposes electrolytes (3.23).** Direct current keeps each electrode's charge fixed.
- **Movement of ions (3.24):** a positive **cations migrate to the negatively charged cathode**, b negative **anions migrate to the positively charged anode** — opposites attract.

ANALOGY

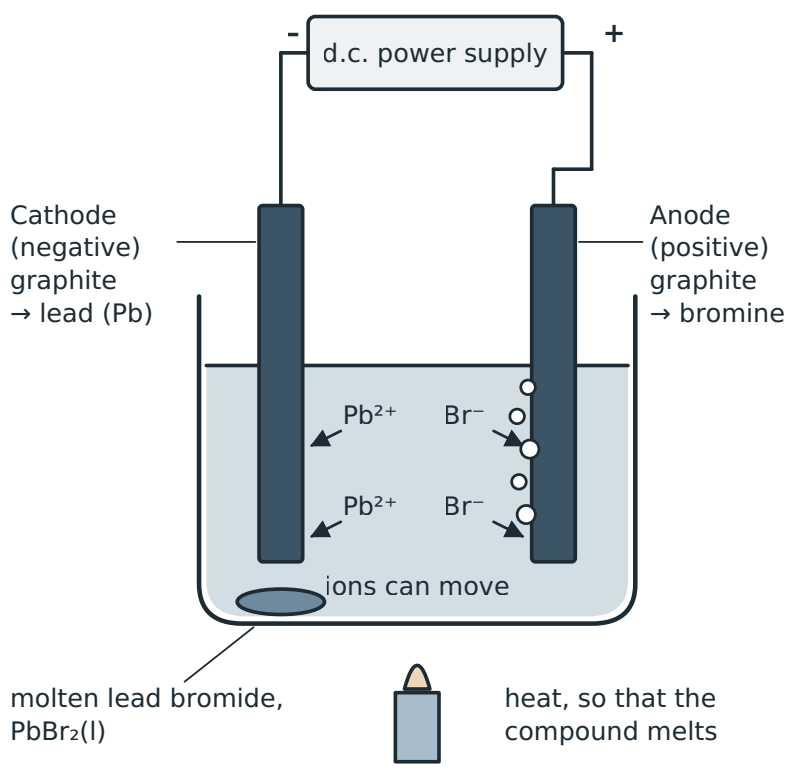
Electrolysis as a sorting office. The electrodes are two magnets of opposite sign in a crowd of ions: the cathode pulls in every positive ion, the anode every negative one. There each ion loses its charge — the cathode hands out electrons, the anode takes them away — and the neutral element is set free. The current does not *make* elements; it sorts ions and pays them in electrons.

Products of electrolysis (3.25–3.26)

Explain the formation of the products in the electrolysis, using inert electrodes, of some electrolytes (3.25).

Molten compounds (3.25e, 3.26). A molten binary ionic compound has only two ions, so you get the **metal at the cathode** and the **non-metal at the anode**: molten lead bromide (demonstration) gives **lead** and **bromine**. Use this to **predict the products of electrolysis of other binary, ionic compounds in the molten state (3.26)** — molten zinc chloride gives zinc and chlorine.

Electrolysis of a molten binary ionic compound



The electrolyte

Molten lead bromide. The electrodes are made of inert graphite so that they do not react with the products.

What happens

Positive Pb^{2+} ions move to the cathode. Negative Br^{-} ions move to the anode. Cathode: grey molten lead (Pb) forms. Anode: orange-brown bromine (Br_2). The compound is broken down into its two elements.

HT only — Half equations

Cathode — reduction (gain of electrons):



Anode — oxidation (loss of electrons):



Why must the compound be molten or dissolved?

In a solid, the ions are held in fixed positions and cannot move, so a solid ionic compound does not conduct. Melting or dissolving it frees the ions to move, making it an electrolyte.

Aqueous solutions. Water supplies H^+ and OH^- ions too, so the electrodes have a choice:

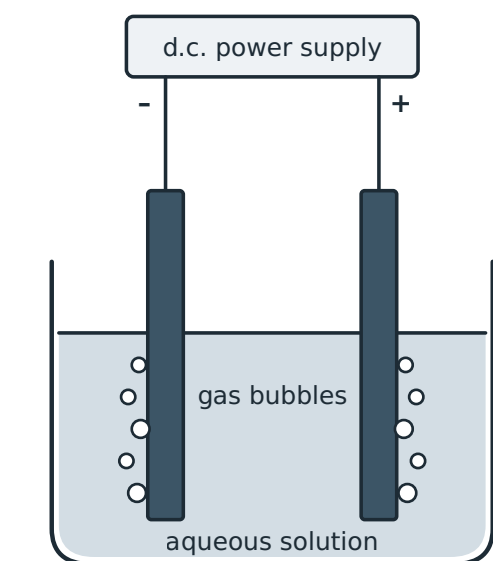
Electrode	Rule	Product
Cathode (–)	Metal more reactive than hydrogen	Hydrogen
Cathode (–)	Metal less reactive (copper, silver, gold)	The metal
Anode (+)	A halide ion (Cl^- , Br^- , I^-) present	The halogen
Anode (+)	No halide present	Oxygen (from OH^-)

The spec's five examples (3.25):

Electrolyte	Cathode	Anode
a Copper chloride solution	Copper (less reactive than H)	Chlorine (halide)
b Sodium chloride solution	Hydrogen (sodium more reactive)	Chlorine (halide)
c Sodium sulfate solution	Hydrogen	Oxygen (no halide)
d Water acidified with sulfuric acid	Hydrogen	Oxygen — 2 : 1 by volume
e Molten lead bromide (demonstration)	Lead	Bromine

Product tests: hydrogen — **squeaky pop**; oxygen — **relights a glowing splint**; chlorine — **bleaches damp litmus paper**.

Electrolysis of aqueous solutions with inert electrodes



Cathode (-) Anode (+)
Inert electrodes (graphite or platinum) are used so they do not react. The electrolyte is a solution of one ionic compound in water.

At the CATHODE (negative electrode)

Is the metal less reactive than hydrogen?

YES

The metal is produced
(e.g. copper, silver)

NO

Hydrogen is produced

HT only — $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ reduction — the ions gain electrons

At the ANODE (positive electrode)

Are halide ions present (Cl^- , Br^- or I^-)?

YES

The halogen is produced
(chlorine, bromine, iodine)

NO

Oxygen is produced

HT only — $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ oxidation — the ions lose electrons

HT only — $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$ when no halide is present

Tests for the gases produced

Hydrogen: hold a lighted splint at the mouth of the tube → burns with a squeaky pop
Oxygen: put a glowing splint into the tube → the splint relights
Chlorine: put damp litmus paper into the tube → the paper is bleached white

Why hydrogen or oxygen can appear

In solution, water molecules break down to give a few H^+ ions and OH^- ions, so these are present as well as the ions of the dissolved compound, and may be discharged instead of them.

HIGHER TIER

Half equations, oxidation and reduction (3.27–3.29)

HT ONLY Write half equations for reactions occurring at the anode and cathode in electrolysis (3.27). **HT ONLY** Explain oxidation and reduction in terms of loss or gain of electrons (3.28), and recall that reduction occurs at the cathode and that oxidation occurs at the anode in electrolysis reactions (3.29).

Copper electrodes and purifying copper (3.30)

Explain the formation of the products in the electrolysis of copper sulfate solution, using copper electrodes, and how this electrolysis can be used to purify copper (3.30).

With **copper (non-inert) electrodes** the electrodes take part:

- **Anode (+)** — the copper **dissolves**: $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$, so it **loses mass**. No oxygen forms, as copper atoms give up electrons more easily than hydroxide ions.
- **Cathode (-)** — Cu^{2+} ions are deposited: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$, so it **gains mass**.
- The solution stays blue: Cu^{2+} ions are replaced as fast as they are removed.

Purifying copper. Make the **impure copper the anode** and a **thin sheet of pure copper the cathode** in copper sulfate solution. Copper dissolves from the anode and deposits as **pure copper** on the cathode; the impurities drop off as **anode sludge**, containing valuable silver and gold. The cathode gains almost exactly the mass the anode loses.