

AQA GCSE Chemistry — Chemistry

Topic 4

Chemical changes

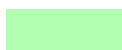
Specification 8462 · AQA

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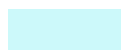
Revision notes for one topic, free to download.

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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HIGHER TIER

Define Oxidation

- the gain of oxygen; **HT ONLY** or the loss of electrons.

Define Redox reaction

- a reaction in which oxidation and reduction happen together.

Define Displacement reaction

- a more reactive metal takes the place of a less reactive metal in a compound.

Define Alkali

- a soluble base; its aqueous solution contains hydroxide ions (OH^-).

Define Neutralisation

- the reaction of an acid with a base or alkali to produce a salt and water.

Define pH scale

- a scale from 0 to 14 measuring acidity or alkalinity; 7 is neutral, below 7 acidic, above 7 alkaline.

Define Concentrated / dilute

- the amount of acid dissolved in a given volume (not the degree of ionisation).

Define Electrolyte

- a molten or dissolved ionic compound whose ions are free to move, so it conducts electricity.

Define Cathode / anode

- the negative electrode (positive ions move to it) / the positive electrode (negative ions move to it).

HIGHER TIER

Define Reduction

- the loss of oxygen; **HT ONLY** or the gain of electrons.

Define Reactivity series

- metals arranged in order of reactivity, from their reactions with water and dilute acids.

Define Acid

- a substance that produces hydrogen ions (H^+) in aqueous solution.

Define Base

- a substance that neutralises an acid to give a salt and water (metal oxides and hydroxides).

Define Salt

- the compound formed when the hydrogen in an acid is replaced by a metal ion.

HIGHER TIER

Define Strong / weak acid

- HT ONLY** an acid that is completely / only partially ionised in aqueous solution.

Define Titration

- CHEMISTRY ONLY** a technique for measuring accurately the volumes of acid and alkali that react together.

Define Electrolysis

- using an electric current to break down an electrolyte into elements.

Define Half equation

- **HT ONLY** an equation showing what happens to one species at one electrode, including the electrons.

4.1 Reactivity of metals

4.1.1 Metal oxides

- Metals react with oxygen to produce metal oxides: *metal + oxygen* → *metal oxide*.
- These are **oxidation** reactions because the metals **gain oxygen**, e.g. $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$.
- The reverse — the **loss of oxygen** — is **reduction**.

4.1.2 The reactivity series

- When metals react with other substances the **metal atoms form positive ions**. The reactivity of a metal is related to its tendency to form positive ions — the more easily it loses its outer electrons, the more reactive it is.
- Metals are ordered in a **reactivity series** from their reactions with **water** and **dilute acids**. The non-metals hydrogen and carbon are often included, as dividing lines.
- A more reactive metal can displace a less reactive metal from a compound, e.g. $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$ (blue solution fades; brown copper coats the zinc).

Metal (most reactive first)	Cold water	Dilute acid	Extraction method
Potassium	Very vigorous, ignites	Violent — not attempted	Electrolysis
Sodium	Vigorous, melts into a ball	Violent — not attempted	Electrolysis
Lithium	Steady fizzing, floats	Very vigorous	Electrolysis
Calcium	Fizzes, cloudy solution	Vigorous fizzing	Electrolysis
Magnesium	Very slow fizzing	Rapid fizzing	Electrolysis
(Carbon)	—	—	<i>dividing line</i>
Zinc	No reaction	Steady fizzing	Reduction with carbon
Iron	No reaction	Slow fizzing	Reduction with carbon
(Hydrogen)	—	—	<i>dividing line</i>
Copper	No reaction	No reaction	Reduction with carbon

- *metal + water* → *metal hydroxide + hydrogen*; *metal + acid* → *salt + hydrogen*.
- Reactions are limited to **room temperature** — reactions with **steam** are not required.
- You should be able to deduce an order of reactivity from experimental results, e.g. how fast each metal fizzes in the same acid, or which metal displaces which.

The reactivity series, with carbon and hydrogen as dividing lines

Reactivity series	Reaction with water or dilute acid	How the metal is obtained
Potassium (K)	reacts with cold water	More reactive than carbon Extracted by electrolysis of the molten compound. This uses large amounts of energy, so it is expensive.
Sodium (Na)	reacts with cold water	
Lithium (Li)	reacts with cold water	
Calcium (Ca)	reacts with cold water	
Magnesium (Mg)	reacts with dilute acid	
Carbon (C)	non-metal, for comparison	Carbon is the dividing line
Zinc (Zn)	reacts with dilute acid	Less reactive than carbon Extracted from its oxide by reduction with carbon: heat the metal oxide with carbon. Reduction is loss of oxygen. e.g. $2\text{CuO} + \text{C} \rightarrow 2\text{Cu} + \text{CO}_2$
Iron (Fe)	reacts slowly with acid	
Hydrogen (H)	non-metal, for comparison	
Copper (Cu)	no reaction with acid	
Gold (Au)	no reaction with acid	

Reactivity = how easily the metal atom forms a positive ion. Metals above hydrogen react with dilute acids. A more reactive metal will displace a less reactive metal from its compound.

Gain of oxygen = oxidation; loss of oxygen = reduction.

HT only — Electron loss = oxidation, gain = reduction.

ANALOGY

The reactivity series is a competition for a partner. Oxygen (or the sulfate, or the acid) is the partner and the metals are queuing up for it. A metal higher in the series wants the partner more, so it takes it off a metal lower down — that is displacement. Copper never takes the partner from magnesium.

HIGHER TIER

4.1.4 HT ONLY Oxidation and reduction in terms of electrons

- **HT ONLY** Oxidation Is Loss of electrons, Reduction Is Gain of electrons — OIL RIG.
- **HT ONLY** You must write ionic equations for displacement reactions and identify which species are oxidised and which are reduced in a given equation or half equation.

4.1.3 Extraction of metals and reduction

- Unreactive metals such as gold are found in the Earth as the metal itself; most metals are found as **compounds** that need chemical reactions to extract the metal.
- Metals less reactive than carbon can be extracted from their oxides by reduction with carbon — the carbon takes the oxygen away. **Reduction involves the loss of oxygen.**
- $2\text{CuO} + \text{C} \rightarrow 2\text{Cu} + \text{CO}_2$: the copper oxide is reduced (loses oxygen); the carbon is **oxidised** (gains oxygen).
- Metals more reactive than carbon must be extracted by **electrolysis** (4.3.3).
- Knowledge is limited to **reduction of oxides using carbon**; process details are not required, but you must interpret or evaluate a given extraction process and identify what is oxidised and reduced in terms of gain or loss of oxygen.

4.2 Reactions of acids

HIGHER TIER

4.2.1 Reactions of acids with metals

- Acids react with some metals to produce salts and hydrogen: *acid + metal* → *salt + hydrogen*. You see **fizzing**, and the gas gives a **squeaky pop** with a lighted splint.
- Limited to magnesium, zinc and iron with **hydrochloric** and **sulfuric** acids:
- $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$; $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$; $\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2$
- **HT ONLY** These are **redox reactions**: the metal atoms **lose electrons** (oxidised), the hydrogen ions **gain electrons** (reduced).
- **HT ONLY** $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ and $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$, giving the ionic equation $\text{Mg} + 2\text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2$

4.2.2 Neutralisation of acids and salt production

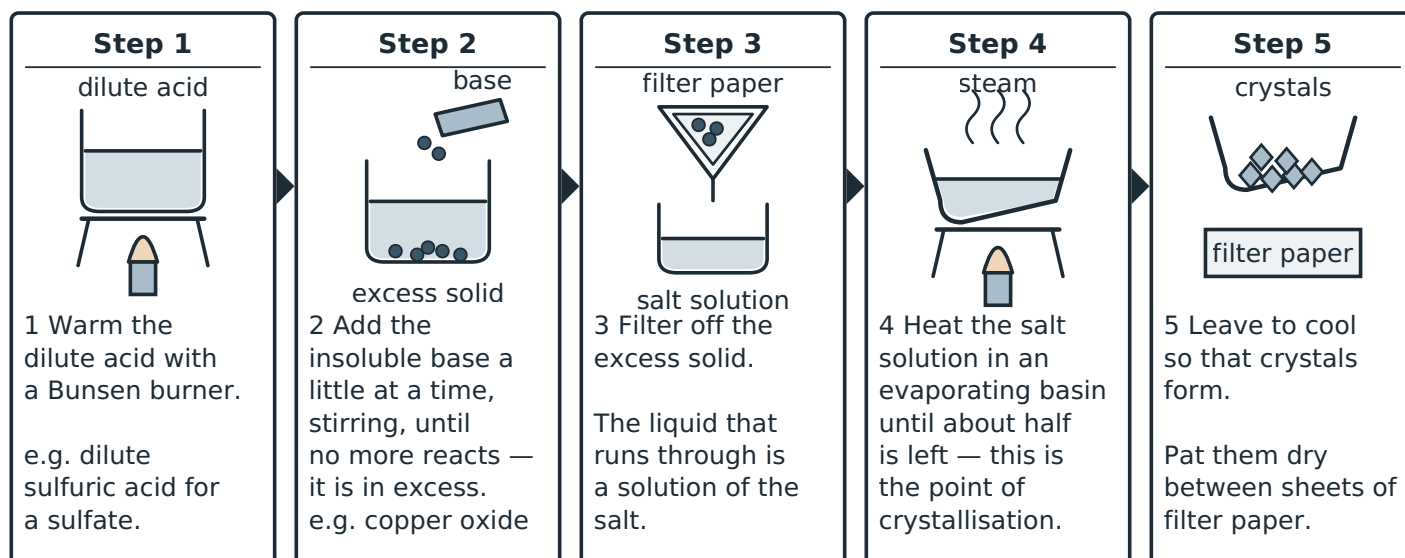
Acid reacts with...	Products	Example
Metal	salt + hydrogen	$\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$
Metal oxide (base)	salt + water	$\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$
Metal hydroxide (base or alkali)	salt + water	$\text{NaOH} + \text{HNO}_3 \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}$
Metal carbonate	salt + water + carbon dioxide	$\text{CuCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O} + \text{CO}_2$

- **Alkalis** are soluble bases (e.g. soluble metal hydroxides); **bases** also include insoluble metal hydroxides and metal oxides.
- The salt depends on **the acid used** and on the positive ion in the base, alkali or carbonate: hydrochloric acid produces chlorides, nitric acid produces nitrates, sulfuric acid produces sulfates.
- Deduce salt formulae by **balancing the charges** of the ions (Cl^- , NO_3^- , SO_4^{2-} with Na^+ , Cu^{2+} , Zn^{2+}): copper nitrate is **$\text{Cu}(\text{NO}_3)_2$** , zinc sulfate **ZnSO_4** . You should be able to predict the products from given reactants.

4.2.3 Soluble salts

- Soluble salts can be made from acids by reacting them with solid insoluble substances — metals, metal oxides, hydroxides or carbonates.
- The solid is added to the acid until no more reacts, and the excess solid is filtered off to produce a solution of the salt, which can be crystallised to produce the solid salt.
- Adding the solid in excess guarantees all the acid is used up, so no acid contaminates the crystals.

Making a pure, dry sample of a soluble salt: react, filter, evaporate, crystallise



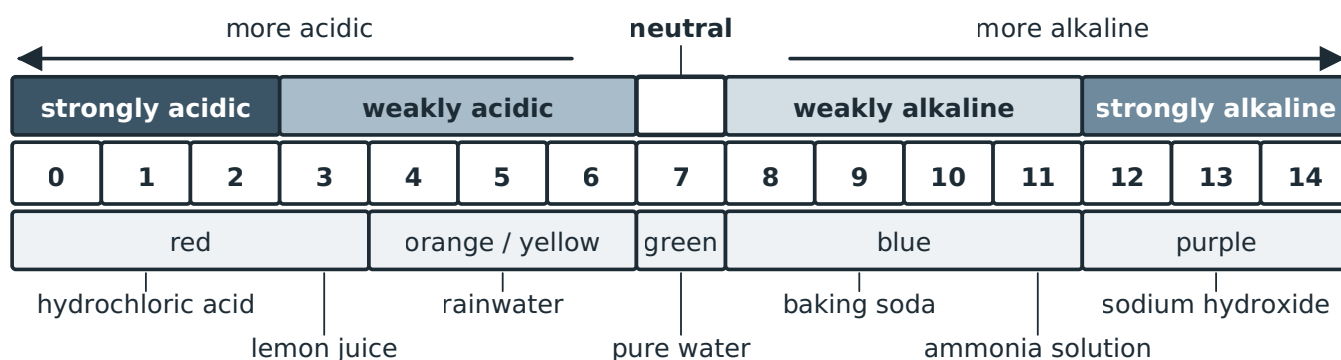
Required practical 1 a pure, dry sample of a soluble salt from an insoluble oxide or carbonate.

copper oxide + sulfuric acid → copper sulfate + water Use a water bath to evaporate.

4.2.4 The pH scale and neutralisation

- Acids produce hydrogen ions (H^+) in aqueous solutions; aqueous solutions of alkalis contain hydroxide ions (OH^-).
- The pH scale, from 0 to 14, is a measure of the acidity or alkalinity of a solution, measured with **universal indicator** (or a wide range indicator — match the colour to a chart for an *approximate* pH) or with a **pH probe**, which gives an accurate numerical value.
- pH 7 is neutral; acids are below 7 and alkalis above 7. Use the scale to identify a solution as acidic or alkaline.

The pH scale and universal indicator colours

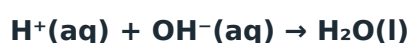


Universal indicator (a wide range indicator) gives these colours. A pH probe is more accurate.

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.

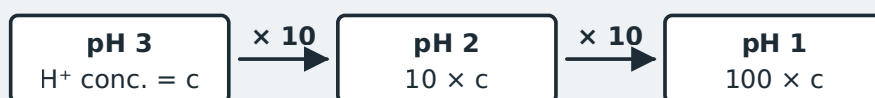
Neutralisation



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.

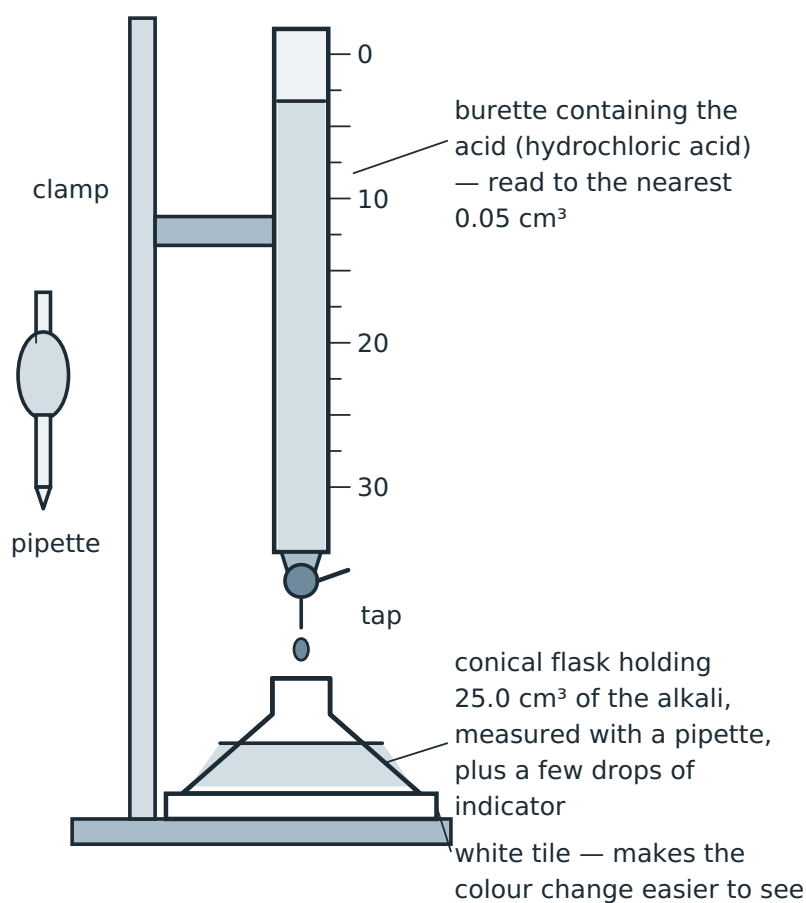


A strong acid is fully ionised, so at the same concentration it has a lower pH than a weak acid.

4.2.5 CHEMISTRY ONLY Titrations

- **CHEMISTRY ONLY** The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator. Titrations here use strong acids and strong alkalis only (sulfuric, hydrochloric and nitric acids).
- **CHEMISTRY ONLY** Use a **single indicator** with a sharp end-point — **phenolphthalein** (pink in alkali $\rightarrow$ colourless in acid) or **methyl orange** (yellow in alkali $\rightarrow$ red in acid). Universal indicator is unsuitable because its colour changes gradually, giving no sharp end-point.

CHEMISTRY ONLY Titration apparatus and how the end-point is judged



Method

- 1 Use a pipette and pipette filler to put 25.0 cm³ 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 cm³ 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

4.2.6 HT ONLY Strong and weak acids

- **HT ONLY** A strong acid is completely ionised in aqueous solution — hydrochloric, nitric and sulfuric acids ($\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$, essentially every molecule).
- **HT ONLY** A weak acid is only partially ionised in aqueous solution — ethanoic, citric and carbonic acids; only a small fraction of molecules release H^+ .
- **HT ONLY** For a given concentration, the stronger an acid, the lower the pH, because more of it has ionised so the hydrogen ion concentration is higher. **HT ONLY** Strong/weak describes the **degree of ionisation**; concentrated/dilute describes the **amount of substance** in a given volume. They are not the same thing.

WATCH OUT

A weak acid is not the same as a dilute acid, and weak does not mean safe — concentrated ethanoic acid is corrosive. Only the fraction of molecules that have ionised decides strong versus weak.

4.3 Electrolysis

HIGHER TIER

4.3.1 The process of electrolysis

- When an ionic compound is melted or dissolved in water, the ions are free to move within the liquid or solution. These liquids and solutions conduct electricity and are called electrolytes. A solid ionic compound does not conduct — its ions are held in fixed positions.
- Passing an electric current through electrolytes causes the ions to move to the electrodes: positively charged ions move to the negative electrode (the cathode) and negatively charged ions move to the positive electrode (the anode).
- Ions are discharged at the electrodes producing elements. This process is called electrolysis.
- **HT ONLY** Throughout electrolysis you must be able to write half equations for the reactions at the electrodes, and may have to complete and balance supplied half equations.

ANALOGY

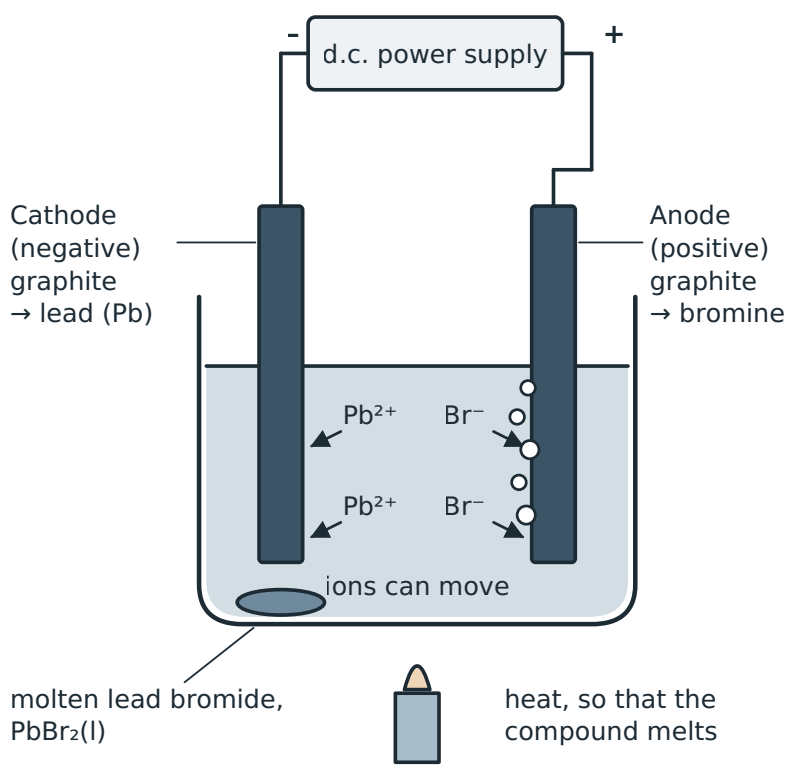
Electrolysis is a two-desk operation. At the negative electrode, electrons are pushed onto the arriving positive ions — they gain electrons, so they are reduced. At the positive electrode, electrons are pulled off the arriving negative ions — they lose electrons, so they are oxidised. The power supply is just the pump moving electrons from one desk to the other.

HIGHER TIER

4.3.2 Electrolysis of molten ionic compounds

- When a simple ionic compound (e.g. lead bromide) is electrolysed in the molten state using inert electrodes, the metal (lead) is produced at the cathode and the non-metal (bromine) at the anode.
- Predicting products for binary ionic compounds in the molten state is therefore straightforward: metal at the cathode, non-metal at the anode.
- **Anhydrous zinc chloride** is a safer alternative for practical work (zinc at the cathode, chlorine at the anode).
- **HT ONLY** Molten lead bromide: $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$ (cathode) and $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$ (anode).

Electrolysis of molten lead bromide



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.

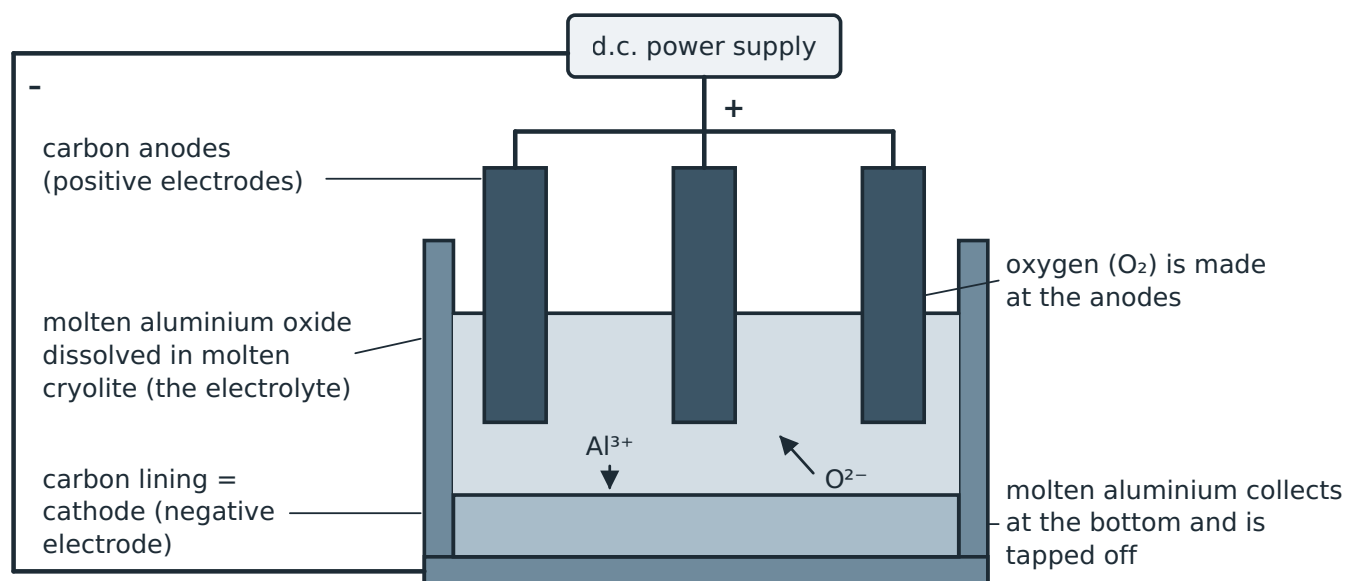
HIGHER TIER

4.3.3 Using electrolysis to extract metals

- Metals can be extracted from molten compounds using electrolysis. Electrolysis is used if the metal is too reactive to be extracted by reduction with carbon, or if the metal reacts with carbon.
- Large amounts of energy are used** — to melt the compounds and to produce the electrical current — so it is expensive.
- Aluminium is manufactured by the electrolysis of a molten mixture of aluminium oxide and cryolite, using carbon as the positive electrode (anode).
- Why a mixture? Aluminium oxide has a very high melting point; dissolving it in molten cryolite lowers the melting point, so less energy is needed and the process is cheaper.
- Why is the positive electrode continually replaced? **Oxygen** produced at the anode reacts with the hot carbon anode to form carbon dioxide, so the anode burns away.

HT ONLY $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$ (cathode) and $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$ (anode).

Extracting aluminium by electrolysis of aluminium oxide dissolved in molten cryolite



Why cryolite is added

Aluminium oxide melts at over 2000 °C. Dissolving it in molten cryolite lowers the melting point, so less energy is needed and it is cheaper. Aluminium is too reactive to be extracted by reduction with carbon.

Why the anodes must be replaced

Oxygen is made at the positive electrodes. It reacts with the hot carbon of the anodes to make carbon dioxide, so the anodes burn away and have to be replaced regularly. The process uses a lot of energy — costly.

HT only — Half equations

Cathode (reduction):



Anode (oxidation):



Overall:



aluminium oxide is reduced to aluminium

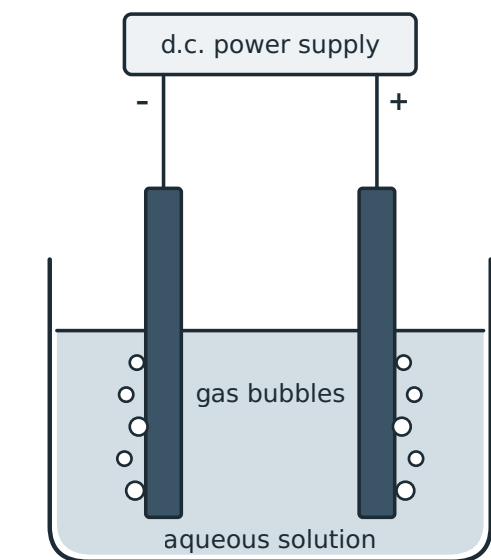
4.3.4 Electrolysis of aqueous solutions

- The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.
- In the aqueous solution, water molecules break down producing hydrogen ions and hydroxide ions that are discharged, so the products are not always the ions of the dissolved compound.

Electrode	Product	Rule	Example
Negative (cathode)	hydrogen	metal more reactive than hydrogen	sodium chloride solution
Negative (cathode)	the metal	metal less reactive than hydrogen	copper sulfate solution
Positive (anode)	the halogen	solution contains halide ions	sodium chloride solution → chlorine
Positive (anode)	oxygen	any other solution (from OH^-)	copper sulfate solution → oxygen

- You should be able to predict the products of the electrolysis of aqueous solutions containing a single ionic compound.
- Gas tests: hydrogen — squeaky pop with a lighted splint; oxygen — relights a glowing splint; **chlorine** — bleaches damp litmus paper.

Electrolysis of aqueous solutions and the rules for predicting the products



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

4.3.5 HT ONLY Representation of reactions at electrodes as half equations

- HT ONLY** At the cathode (negative electrode), positively charged ions gain electrons, so the reactions are reductions.
- HT ONLY** At the anode (positive electrode), negatively charged ions lose electrons, so the reactions are oxidations.

Electrode	Half equation
Cathode (hydrogen)	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$
Cathode (metal)	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
Anode (oxygen)	$4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$ (or $4\text{OH}^- - 4\text{e}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O}$)
Anode (halogen)	$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$

WATCH OUT

Electrons never travel through the electrolyte — **ions** carry the charge through the solution and **electrons** travel through the wires and electrodes. And never leave the electrons out of a half equation: the charges must balance on both sides.