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Topic 4 AQA Chemistry - Chemical Changes-

- Triple Science Content only in purple
- Triple Science and Higher Content Only in blue

Reactivity of metals-

Metal oxides

- **Reduction= loss of oxygen and oxidation= gain of oxygen**
- It is oxidation reactions because the metals gain oxygen
- Metals + oxygen -> metal oxides

The reactivity series

- Metals in order of their reactivity in a reactivity series:

- A more reactive metal can displace a less reactive metals in a compound

Metal- Reaction With Water: **Metal- Reaction with dilute acid:**

Potassium	Violent	Calcium	Very quick
Sodium	Very quick	Magnesium	Quick
Lithium	Quick	Zinc	Fairly slow
Calcium	Slow	Iron Copper	More slow Very Slow

Extraction of metals and reduction with carbon

- most metals are found as compounds e.g MgO except **gold** as it is very unreactve.
- Metals less reactive than carbon can be extracted from their oxides **by reduction with carbon**

reduction involves the loss of oxygen

Reduction with Carbon

Reduction = removal of oxygen.
Oxidation = gain of oxygen.

Metal oxides → reduced (lose oxygen) → metal
Carbon → oxidised (gains oxygen) → carbon dioxide

General word equation:

metal oxide+carbon→metal+carbon dioxide

iron(III) oxide+carbon→iron+carbon dioxide

or in symbols:

$2Fe_2O_3+3C \rightarrow 4Fe+3CO_2$

This is a redox reaction, because:

- The metal oxide is reduced (loses oxygen).
- **The carbon is oxidised (gains oxygen).**

Oxidation and reduction

- **OIL RIG**,- stands for Oxidation Is Loss and Reduction Is Gain (of electrons)
- Writing ionic equations:

$Na \rightarrow Na^+ + e^-$ This shows the oxidation of Sodium- It has lost an electron, and now has a +1 charge.

$Na^+ + e^- \rightarrow Na$ This shows the reduction of Sodium- it has gained an electron, and has gone from a +1 to a 0 (neutral) charge.

<u>Reactions of acids</u>

What Are Acids and Bases?

- **Acids** are substances that **produce hydrogen ions (H^+)** when dissolved in water.
  Example:
  - Hydrochloric acid → $HCl \rightarrow H^+ + Cl^-$
  - Sulfuric acid → $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$
- **Alkalis** are **soluble bases** that produce **hydroxide ions (OH^-)** when dissolved in water.
  Example:
  - Sodium hydroxide → $NaOH \rightarrow Na^+ + OH^-$
- **Neutralisation reaction:**
  Acid + base (or alkali)→ salt+ water

Naming Salts

When an acid reacts, the **metal or base provides the first part** of the salt's name, and the **acid provides the second part**.

Acid	Ion	Salt Ending	Example (with NaOH)
Hydrochloric acid (HCl)	Cl^-	**-chloride**	Sodium chloride
Sulfuric acid (H_2SO_4)	SO_4^{2-}	**-sulfate**	Sodium sulfate
Nitric acid (HNO_3)	NO_3^-	**-nitrate**	Sodium nitrate

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Acids and Metals

General reaction:

Acid + metal→ salt + hydrogen gas

Example:

$2HCl + Mg \rightarrow MgCl_2 + H_2$

Test for hydrogen gas:
→ A **lit splint** makes a **squeaky pop** sound.

Only metals above hydrogen in the reactivity series react with acids to form hydrogen gas (e.g., magnesium, zinc, iron).

Acids and Metal Oxides

Metal oxides are bases, so they react with acids to form **salt + water**.

General reaction:

Acid + metal oxide→ salt + water

Example:

$H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$

Acids and Metal Hydroxides

Metal hydroxides are **alkalis**, so the reaction is **neutralisation**.

General reaction:

Acid + metal hydroxide →salt + water

Example:

$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$

Acids and Metal Carbonates

General reaction:

Acid + metal carbonate→salt + water + carbon dioxide

Example:

$2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$

Test for CO_2:
→ Bubble gas through **limewater** — it turns **cloudy** (calcium carbonate forms)

Required Practical: Making a Soluble Salt (from an Insoluble Base)

Example: Preparing copper sulfate from copper oxide and sulfuric acid.

Steps:

1. **Add excess copper oxide** to warm dilute sulfuric acid (so all acid reacts).
2. **Stir and gently heat** until no more copper oxide dissolves.
3. **Filter** the mixture to remove unreacted solid.
4. **Evaporate** the filtrate slowly to form **blue copper sulfate crystals**.

Equation:

$H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$

Type of reaction: Neutralisation.

Acids, Alkalis and pH Scale

- pH is a measure of **hydrogen ion concentration**.
- Lower pH = higher [H^+].
- The **pH scale** goes from **0 (strong acid)** to **14 (strong alkali)**, with **7 = neutral**.

As pH decreases by 1, [H^+] ions increases 10×

So, a solution of pH 2 is **10 times more acidic** than pH 3.

Strong and Weak Acids

- **Strong acids** (e.g., HCl, H_2SO_4, HNO_3) **completely ionise** in water.
  → All acid molecules release H^+ ions.
- **Weak acids** (e.g., ethanoic acid, citric acid) **partially ionise** in water.
  → Only some acid molecules release H^+ ions.

Example:

CH3COOH⇌ H+ + CH3COO−

Note:

- Both can have the same concentration, but the **strong acid has a lower pH** because it has more free H^+ ions.
- Weak acids are **reversible reactions** (⇌).

Concentration vs Strength

- **Concentration** = how much acid is dissolved per dm^3 of solution (mol/dm^3).
- **Strength** = degree of ionisation.

So a concentrated weak acid can have more moles of acid per dm^3 but still fewer free H^+ ions than a dilute strong acid.

Ionic Equation for a neutralisation reaction-

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

Titrations

The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator.

How to carry out a titration:

1. Using dilute Hydrochloric acid, wash the burette, then use water to wash it.
2. Fill burette to $100cm^3$ with the acid with the meniscus' base line directly on the $100cm^3$ line.
3. Use $25cm^3$ pipette to add $25cm^3$ of alkali into a conical flask.
4. Add a few drops of phenolphthalein- pink when alkaline and colourless when acidic.
5. Add acid from burette to alkali until end-point is reached (as shown by indicator)
6. The titre (volume of acid needed to exactly neutralise the acid) is the difference between the first ($100cm^3$) and second readings on the burette
7. Repeat the experiment to gain more precise results and get concordant titres.

Titration calculations

Titration Calculation Video

GCSE Quantitative Chemistry - Amount Of Substances, Calculating Concentration

What is Electrolysis?

Definition:
Electrolysis is the **decomposition (breaking down)** of an **ionic compound** using **electricity**.

- It only works when the **ions are free to move** — meaning the compound must be **molten** or **dissolved in water (aqueous)**.

- The electric current is passed through an **electrolyte** — a liquid that conducts electricity because it contains free ions.

Key components:

Term	Meaning
Electrolyte	The ionic substance (molten or dissolved) that electricity passes through
Electrodes	Conducting rods that allow current to enter/leave the electrolyte (usually made of **inert** material like graphite or platinum)
Anode	The **positive electrode** (attracts negative ions)
Cathode	The **negative electrode** (attracts positive ions)

When electricity passes through the electrolyte:

- **Positive ions (cations)** move towards the **negative electrode (cathode)** → they **gain electrons** → **reduction** occurs.
- **Negative ions (anions)** move towards the **positive electrode (anode)** → they **lose electrons** → **oxidation** occurs.

Remember:

OIL RIG —
Oxidation Is Loss, **Reduction Is Gain** (of electrons)

Half Equations

Half equations show what happens to ions at each electrode.

Example (electrolysis of molten lead bromide, $PbBr_2$):

- **At the cathode (–):**
  Pb2+ + 2e− →Pb
- (Lead ions gain electrons → lead metal formed)

- **At the anode (+):**
  2Br− →Br2 +2e-
- (Bromide ions lose electrons → bromine gas formed)

Electrolysis of molten ionic compounds

- When an **ionic compound** (e.g Magnesium Chloride) is electrolysed in the **molten state** (using inert electrodes), the metal (Magnesium) is produced at the **cathode** and the non-metal (Chlorine) is produced at the **anode**
- this is because the metal is the **positive Mg2+** ions and the non-metal is the **Cl- negative ions**

Electrolysis of Aqueous Solutions (Water-Based)

Aqueous solutions contain **ions from the compound** and **ions from water**:

$H_2O \rightleftharpoons H^+ + OH^-$

So in total, there are usually **four ions** present.

To predict what forms:

1. Look at which ions are present.
2. Apply the **electrolysis rules** below.

Rules for the Cathode (–):

- **Hydrogen gas (H_2)** is produced **if the metal is more reactive than hydrogen**.
- If the metal is **less reactive than hydrogen** (e.g. copper, silver), then the **metal** is produced.

Metal Reactivity	Product at Cathode
More reactive than hydrogen (e.g. Na, Mg, Al)	Hydrogen gas
Less reactive than hydrogen (e.g. Cu, Ag)	Metal deposited

Example:

Electrolysis of copper(II) sulfate solution ($CuSO_4$):

- $Cu^{2+} + 2e^- \rightarrow Cu$ (copper metal forms at cathode)

Rules for the Anode (+):

- **Oxygen gas (O_2)** is produced **unless** a **halide ion (Cl^-, Br^-, I^-)** is present.
- If halide ions are present, the **halogen gas** is produced.

Ion Present	Product at Anode
Cl^-, Br^-, I^-	Chlorine, bromine, or iodine gas
No halide (e.g. SO_4^{2-}, NO_3^-)	Oxygen gas (from OH^-)

Example:
Electrolysis of sodium chloride solution (NaCl (aq)):

- At cathode: $2H^+ + 2e^- \rightarrow H_2$
- At anode: $2Cl^- \rightarrow Cl_2 + 2e^-$
  Products: Hydrogen gas, chlorine gas, and sodium hydroxide remain in solution.

Using Electrolysis to Extract Metals

Metals that are **more reactive than carbon** (e.g. aluminium, magnesium) cannot be extracted by reduction with carbon.
Instead, **electrolysis** is used.

Example: Extraction of Aluminium from Bauxite

- **Ore:** Bauxite (contains aluminium oxide, Al_2O_3)
- **Step 1:** Aluminium oxide is **dissolved in molten cryolite** (Na_3AlF_6) to **lower its melting point** (reduces energy cost).

- **Step 2:** Electricity is passed through the molten mixture.

Reactions:

Electrode	Reaction	Product
Cathode (−)	$Al^{3+} + 3e^- \rightarrow Al$	Aluminium (molten)
Anode (+)	$2O^{2-} \rightarrow O_2 + 4e^-$	Oxygen gas

Note:

- The **oxygen reacts with the carbon anode**, forming **CO_2** → so the anodes need to be **replaced regularly**.
- The process is **very energy-intensive** (expensive).