AQA GCSE Chemistry (Combined Science) Unit 4: Chemical Changes Knowledge Organiser

The Reactivity Series

Here's a mnemonic to help you learn the order:

purple (potassium)
slime (sodium)
can (calcium)
make (magnesium)
a (aluminium)
careless (carbon)
zebra (zinc)
insane (iron)
try (tin)
learning (lead)
how (hydrogen)
camels (copper)
surprise (silver)
gorillas (gold)

- potassium
- sodium
- calcium
- magnesium
- aluminium
- carbon →
- zinc
- iron
- tin
- lead
- hydrogen →
- copper
- silver
- gold
- platinum

The reactivity series is a league table for metals. The more reactive metals are near the top of the table with the least reactive near the bottom. In chemical reactions, a more reactive metal will displace a less reactive metal.

Reactions of Metals with Water

Metals, when reacted with water, produce a metal hydroxide and hydrogen.

lithium + water ⟶ lithium hydroxide + hydrogen

$2Li + 2H_2O \longrightarrow 2LiOH + H_2$

The more reactive a metal is, the faster the reaction.

Reactions of Metals with Dilute Acid

Metals, when reacted with acids, produce a salt and hydrogen.

Sodium + hydrochloric acid ⟶ sodium chloride + hydrogen

$2Na + 2HCl \longrightarrow 2NaCl + H_2$

Metals that are below hydrogen in the reactivity series **do not** react with dilute acids.

Reactions of Acids

The general formula for the reaction between an acid and a metal is:
acid + metal ⟶ salt + hydrogen

For example: hydrochloric acid + sodium ⟶ sodium chloride + hydrogen

$2HCl + 2Na \longrightarrow 2NaCl + H_2$

When an acid reacts with an alkali, a neutralisation reaction takes place and a salt and water are produced.

The general formula for this kind of reaction is as follows:

acid + alkali ⟶ salt + water

hydrochloric acid + sodium hydroxide ⟶ sodium chloride + water

$HCl + NaOH \longrightarrow NaCl + H_2O$

Naming Salts

The first part comes from the metal in the metal carbonate, oxide or hydroxide. The second part of the name comes from the acid that was used to make it.

For example, sodium chloride.

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

Redox Reactions (Higher Tier Only)

When metals react with acids, they undergo a redox reaction. A **redox reaction** occurs when both **oxidation** and **reduction** take place at the **same time.**

For example:

$2H^+ + Ca \longrightarrow Ca^{2+} + H_2$

The ionic equation can be further split into two half equations.

$Ca \longrightarrow Ca^{2+} + 2e^-$

Oxidation is loss of electrons.

$2H^+ + 2e^- \longrightarrow H_2$

Reduction is gaining of electrons.

Reactions with Bases

The general formula for the reaction between an acid and a metal oxide is:
acid + metal oxide ⟶ salt + water

sulfuric acid + copper oxide ⟶ copper sulfate + water

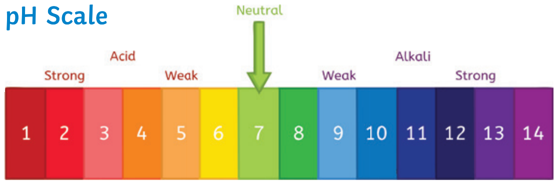
$H_2SO_4 + CuO \longrightarrow CuSO_4 + H_2O$

Reactions with Carbonates

The general formula for the reaction between an acid and a carbonate is:
acid + carbonate ⟶ salt + water + carbon dioxide

hydrochloric acid + calcium carbonate ⟶ calcium chloride + water + carbon dioxide

pH Scale



In aqueous solutions, acids produce H^+ ions and alkalis produce OH^- ions.

Neutral solutions are pH7 and are neither acids or alkalis.

For example, in neutralisation reactions, hydrogen ions from an acid react with hydroxide ions from an alkali to produce water:

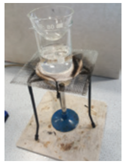
$H^+ + OH^- \longrightarrow H_2O$

Making Soluble Salts

1. Make a saturated solution by stirring copper oxide into the sulfuric acid until no more will dissolve.

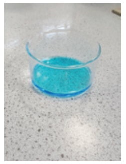
2. Filter the solution to remove the excess copper oxide solid.

3. Half fill a beaker with water and set this over a Bunsen burner to heat the water. Place an evaporating dish on top of the beaker.



4. Add some of the solution to the evaporating basin and heat until crystals begin to form.

5. Once cooled, pour the remaining liquid into a crystallising dish and leave to cool for 24 hours.



6. Remove the crystals with a spatula and pat dry between paper towels.

Strong and Weak Acids (Higher Tier Only)

A **strong** acid **completely dissociates** in a solution. For example: $HCl \rightarrow H^+ + Cl^-$

Hydrochloric acid is able to completely dissociate in solution to form hydrogen and chloride ions.

Examples of strong acids include nitric acid (HNO_3) and sulfuric acid (H_2SO_4).

Weak acids in comparison only partially dissociate.

For example acetic acid **partially dissociates** to form a hydrogen and acetate ion.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

The **double arrow** symbol indicates that the reaction is **reversible**. Both the forward and reverse reaction occur at the same time and the reaction never goes to completion.

The Process of Electrolysis

Electrolysis is the **splitting up** of an ionic substance using **electricity.**

On setting up an electrical circuit for electrolysis, two **electrodes** are required to be placed in the electrolyte. The electrodes are **conducting rods.** One of the rods is connected to the **positive** terminal and the other to the **negative** terminal.

The **electrodes** are **inert** (this means they do not react in the reaction) and are often made from **graphite** or platinum.

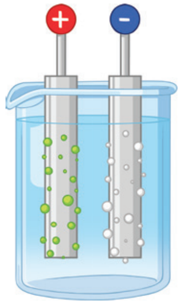
During the process of electrolysis, **opposites attract.** The positively charged ions will be attracted toward the negative electrode. The negatively charged ions will be attracted towards the positive electrode.

When ions reach the electrodes, the charges are lost and they become elements.

The **positive** electrode is called the **anode.**

The **negative** electrode is called the **cathode.**

Electrolysis of Aqueous Solutions



Gases may be given off or metals deposited at the electrodes. This is dependent on the reactivity of the elements involved.

If the metal is **more reactive** than **hydrogen** in the reactivity series, then **hydrogen** will be **produced** at the **negative cathode.** At the **positive anode,** negatively charged ions **lose** electrons. This is called **oxidation** and you say that the ions have been oxidised.

Using Electrolysis to Extract Metals

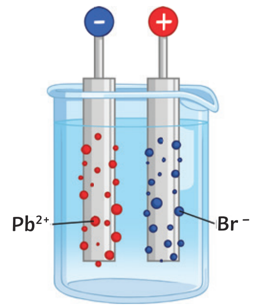
Metals are extracted by electrolysis if the metal in question reacts with carbon or if it is too reactive to be extracted by reduction with carbon. During the extraction process, large quantities of energy are used to melt the compounds.

Aluminium is manufactured by the process of electrolysis. Aluminium oxide has a high melting point and melting it would use large amounts of energy. This would increase the cost of the process, therefore molten **cryolite** is added to aluminium oxide to lower the melting point and thus reduce the cost.

Electrolysis of Molten Ionic Compounds – Lead Bromide

Lead bromide is an **ionic** substance. Ionic substances, when solid, are **not** able to conduct electricity. When molten or in solution, the ions are free to move and are able to carry a charge.

The **positive** lead ions are attracted toward the **negative cathode** at the same time as the **negative bromide** ions are attracted toward the **positive anode.**



Oxidation **i**s the **l**oss of electrons and **r**eduction **i**s the **g**aining of electrons. **OIL RIG** (Higher Tier Only).

We represent what is happening at the electrodes by using **half equations** (Higher Tier Only).

The lead ions are attracted towards the negative electrode. When the **lead ions** (Pb^{2+}) reach the cathode, each ion **gains two electrons** and becomes a neutral atom. We say that the lead ions have been **reduced.**

$Pb^{2+} + 2e^- \rightarrow Pb$

The bromide ions are attracted towards the positive electrode. When the **bromide ions** (Br-) reach the anode, each ion **loses one electron** to become a neutral atom. Two bromine atoms are then able to bond together to form the covalent molecule Br_2.

$2Br^- \rightarrow Br_2 + 2e^-$