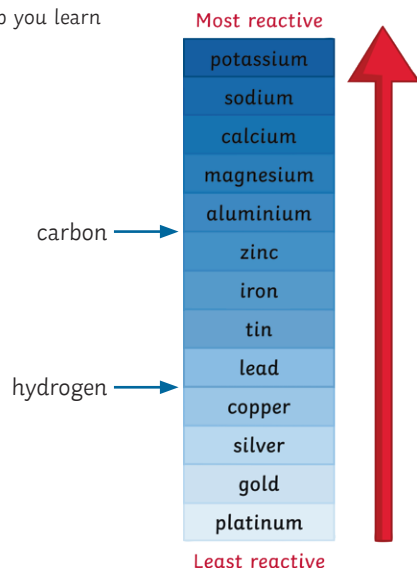


AQA GCSE Chemistry (Separate Science) Unit 4: Chemical Changes

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)



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



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



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:



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



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 acid + alkali → salt + water

hydrochloric acid + sodium hydroxide → sodium chloride + water



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.

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

For example, sodium chloride.

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**.

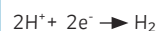
For example:



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



Oxidation is loss of electrons.



Reduction is gaining of electrons.

Reactions with Bases

The general formula for the reaction between an acid and a metal oxide is:



sulfuric acid + copper oxide → copper sulfate + water



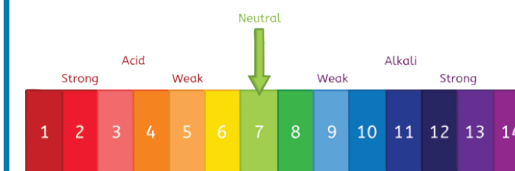
Reactions with Carbonates

The general formula for the reaction between an acid and a carbonate is:



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

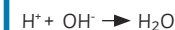
pH Scale



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

Neutral solutions are pH7 and are neither acids nor alkalis.

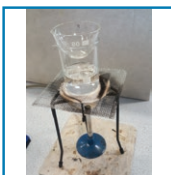
For example, in neutralisation reactions, hydrogen ions from an acid react with hydroxide ions from an alkali to produce water:



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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)

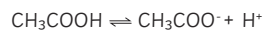
A **strong** acid **completely dissociates** in a solution. For example: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

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.



The **double arrow** symbol indicates that the reaction is **reversible**.

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.

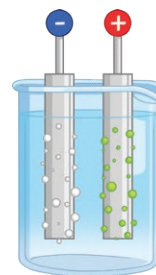
The **positive** electrode is called the **anode**.

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

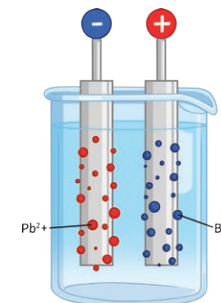
Metals are extracted by electrolysis if the metal in question reacts with carbon or if it is too reactive to be extracted by reduction with carbon. During the extraction process, large quantities of energy are used to melt the compounds.

Aluminium is manufactured by the process of electrolysis. Aluminium oxide has a high melting point and melting it would use large amounts of energy and 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 is the **loss** of electrons and **reduction** is the **gaining** 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**.



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 .



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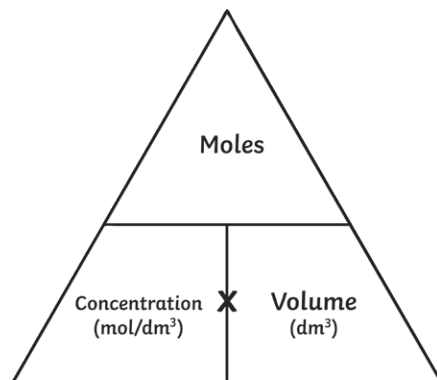
Titration Method (Chemistry Only)

- Using the pipette and pipette filler, measure 25cm³ sodium hydroxide solution and pour into a conical flask.
- Add several drops of phenolphthalein to the sodium hydroxide solution.
- Swirl the flask and the mixture should be pink.
- Place the conical flask on a white tile.
- Place the burette into its stand, ensuring the tap is closed. Using the funnel, fill the burette with sulfuric acid to the 0cm³ line. Should you go above this line, open the tap and allow the excess to run off into a beaker.
- Once the burette is correctly filled, place over the conical flask.
- Carefully open the tap so the acid flows slowly into the conical flask. Swirl the flask and look for the indicator changing from pink to colourless.
- Continue adding the acid to the flask until the indicator is permanently colourless.
- Record the total volume of acid added to the sodium hydroxide in the results table.
- Repeat the experiment twice more.

Titration Method (Chemistry Only)

Using the results from a titration experiment, it is possible to calculate the concentration of a solution or the volume of solution required to neutralise an acid or alkali.

Worked Example



In a titration, 20cm³ of 1.0mol/dm³ sulfuric acid reacted with 25cm³ of sodium hydroxide. What was the concentration of sodium hydroxide?

Write out the symbol equation for the reaction.



Check that the equation is balanced.



To convert cm³ to dm³, just divide by 1000.

Draw a table like the one below and fill it in with the information that you know from the question.

	Acid (H ₂ SO ₄)	Alkali (NaOH)
number of moles		
concentration mol/dm ³	1.0	
volume (dm ³)	0.02	0.025

As the values for the **concentration** and **volume** of the acid are known, it is possible to now work out the **number of moles** of H₂SO₄.

number of moles = concentration × volume

$$\text{number of moles} = 1.0 \times 0.02 = \mathbf{0.02 \text{ moles}}$$

From the balanced symbol equation, we know that there is double the amount of NaOH compared to H₂SO₄, therefore to calculate the number of moles of the alkali, we double the number of moles of the acid. 0.02 × 2 = **0.04 moles**.

	Acid (H ₂ SO ₄)	Alkali (NaOH)
number of moles	0.02	0.04
concentration mol/dm ³	1.0	
volume (dm ³)	0.02	0.025

The question asks you to calculate the **concentration** of sodium hydroxide. As the number of moles and volume is now known, it is possible to calculate the concentration.

concentration = number of moles ÷ volume

$$\text{concentration} = 0.04 \div 0.025$$

$$\text{concentration} = \mathbf{1.6 \text{ mol/dm}^3}$$

