

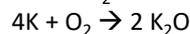
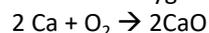
4.4. Revision Checklist: Chemical Changes

Reactivity of metals

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.

Metals react with oxygen to produce metal oxides. The reactions are **oxidation** reactions because the metals **gain oxygen**.

Calcium + Oxygen → calcium oxide



Reactive metals such as magnesium will burn with a flame in oxygen. Less reactive metals like iron will tarnish and change colour without a flame.

Metals can be arranged in order of their reactivity in a reactivity series.

The metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids.

The non-metals hydrogen and carbon are often included in the reactivity series.

Reactive metals such as potassium, sodium and calcium will react with cold water producing bubbles of hydrogen gas

All metals above hydrogen in the reactivity series will react with acids producing bubbles of hydrogen gas.

The more reactive the metal the faster/more vigorous the reaction will be producing hydrogen gas quicker.

Metals below hydrogen in the reactivity series will not react with acids.

The most important metals in the reactivity series can be learnt by using mnemonics like: **Please Stop Calling My Aunty Zebra In The Class- potassium, sodium, calcium, magnesium, aluminium, zinc, iron, tin, copper**

Reactivity series of Metals

Potassium	most reactive
Sodium	
Calcium	
Magnesium	
Aluminium	
Carbon	
Zinc	
Iron	
Tin	
Lead	
Hydrogen	
Copper	
Silver	
Gold	least reactive

Displacement Reactions

A more reactive metal can displace a less reactive metal from a compound

e.g. Copper sulphate + magnesium → Copper + magnesium sulphate
 $\text{CuSO}_4 + \text{Mg} \rightarrow \text{Cu} + \text{MgSO}_4$

Magnesium is more reactive than copper so there is a displacement reaction.

e.g. zinc sulphate + copper → no reaction

Copper is less reactive than zinc so there is no reaction.

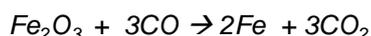
Extracting metals

Unreactive metals such as gold are found in the Earth as the metal itself but most metals are found as compounds that require chemical reactions to extract the metal.

Reduction is defined as **loss of oxygen**

Metals that are **less reactive than carbon** can be extracted from their oxides **by reduction** with carbon

iron oxide is reduced in the blast furnace by reacting with carbon or carbon monoxide to make iron



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

*(in this equation the iron oxide is being **reduced** and the carbon monoxide is causing the reduction)*

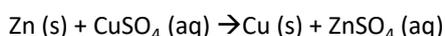
Metals above carbon in the reactivity series are often extracted by **electrolysis** or by displacement reactions with more reactive metals

Oxidation and Reduction

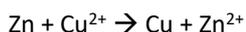
In reactions **oxidation** can be defined as **loss of electrons** and **reduction** as **gain of electrons**

Ionic equations

In displacement reactions, it is only the metal ions that are reacting in terms of electrons being transferred. The non metal ions e.g. sulfate (SO_4^{2-}) ions are 'spectating'.



We can write the equation above as an **ionic** equation which does not include the spectator ion.



In this reaction zinc has lost electrons (to become more positive) and so has **oxidised**

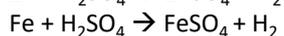
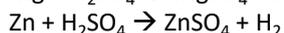
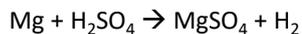
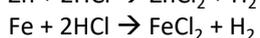
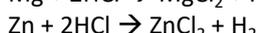
This can be written as a **half equation** $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$

In this reaction copper has gained electrons (to become neutral) and so has **reduced**

This can be written as a **half equation** $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

Reactions of Acid with Metals

Acids react with some metals to produce salts and hydrogen
Metal + acid \rightarrow salt + hydrogen

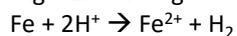
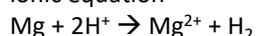


All metals above hydrogen in the reactivity series will react with acids producing bubbles of hydrogen gas.

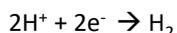
Metals below hydrogen in the reactivity series will not react with acids

Reactions of acid and metals are **redox reactions**

These equations can be written as an ionic equation



They can also be written as a half equations



The magnesium is **oxidised** as it is **losing** electrons

The hydrogen is **reduced** as it is **gaining** electrons

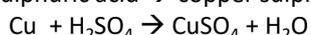
Neutralisation Reactions

Acids are neutralised by alkalis (eg soluble metal hydroxides) and bases (eg insoluble metal hydroxides and metal oxides) to produce salts and water

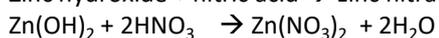
base + acid \rightarrow salt + water

Example:

Copper Oxide + sulphuric acid \rightarrow copper sulphate + water



Zinc hydroxide + nitric acid \rightarrow zinc nitrate + water



Acids are also neutralised by metal carbonates to produce salts, water and carbon dioxide.

Metal carbonate + acid \rightarrow salt + water + carbon dioxide

Example



The particular salt produced in any reaction between an acid and a base or alkali depends on:

- the acid used
 - hydrochloric (HCl) acid produces **chlorides**,
 - nitric acid (HNO_3) produces **nitrates**,
 - sulfuric acid (H_2SO_4) produces **sulfates**
- the metal in the base or alkali.

Making Soluble salts

Soluble salts can be made from acids by reacting them with **solid insoluble** substances, such as metals, metal oxides, hydroxides or carbonates.

- not all metals are suitable, some are too reactive and others are not reactive enough (eg. Copper is too unreactive and sodium too reactive)

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. Salt solutions can be crystallised to produce solid salts.

Required Practical

Detailed method for forming a salt from a solid base (or metal)

- Measure 50 cm³ acid in beaker
- Warm acid
- Add spatulas of solid base (or metal)
- stir
- Keep adding solid until no more reacts and base is in excess
- Filter off excess solid base with filter funnel and filter paper
- Pour solution into evaporating basin
- Heat using a Bunsen burner to crystallisation point
- Leave to crystallise
- Dry crystals on filter paper

pH Scale and Neutralisation

Acids produce **hydrogen ions (H⁺)** in aqueous solutions.

Aqueous solutions of **alkalis** contain **hydroxide ions (OH⁻)**.

pH Scale

The pH scale, from 0 to 14, is a measure of the acidity or alkalinity of a solution, and can be measured using universal indicator or a pH probe.

Base and Alkalis

Bases neutralise acids

Metal oxides and hydroxides are bases .

Soluble hydroxide bases are called **alkalis**

A solution with pH 7 is neutral.

Aqueous solutions of acids have pH values of less than 7.

Aqueous solutions of alkalis have pH values greater than 7.

Universal indicator colours

Red pH 1-2 strong acid

Orange/yellow pH3-6 weak acid

Green pH7 neutral

blue pH8-11 weak alkali

Purple pH 12-14 strong alkali

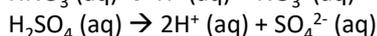
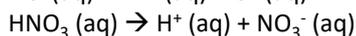
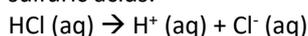
Neutralisation

In neutralisation reactions between an acid and an alkali, hydrogen ions react with hydroxide ions to produce water.

This reaction can be represented by the ionic equation: **H⁺(aq) + OH⁻(aq) → H₂O(l)**.

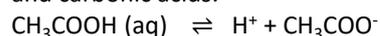
Strong Acids

A strong acid is **completely ionised** in aqueous solution. Examples of strong acids are hydrochloric, nitric and sulfuric acids.



Weak Acids

A weak acid is only **partially ionised** in aqueous solution. Examples of weak acids are ethanoic, citric and carbonic acids.



Ethanoic acid ethanoate ion

Less than 1% of ethanoic acid molecules will ionise in this reversible reaction.

For a given concentration of aqueous solutions, the stronger an acid, the lower the pH.

Dilute and Concentrated Acids

A concentrated acid will have more moles of acid per unit volume than a dilute acid.

A concentrated acid is made into a dilute acid by adding water.

Dilute and concentrated does not mean the same as strong and weak

pH and Concentration

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

If an acid is **diluted 10 times** its pH will increase by **one unit** e.g. if a solution of hydrochloric acid has a concentration of 0.01 mol/dm³ it will have a pH of 2. If it is then diluted a 100 times it will have a pH of 4

Concentration of H ⁺ (aq) ions in mol/dm ³	pH
1.0	0.0
1.0 × 10 ⁻¹	1.0
1.0 × 10 ⁻²	2.0
1.0 × 10 ⁻³	3.0
1.0 × 10 ⁻⁴	4.0
1.0 × 10 ⁻⁵	5.0
1.0 × 10 ⁻⁶	6.0

Electrolysis

When an ionic compound is melted or dissolved in water, the **ions are free to move** about within the liquid or solution. These liquids and solutions are able to conduct electricity and are called **electrolytes**.

Electrolysis is the passing an **electric current** through **ionic substances** that are **molten or in solution** to **breaks them down** into elements. Ions are discharged at the electrodes producing elements

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).

Electrolysis of molten salts

When a simple ionic compound is electrolysed in the molten state using inert electrodes, the salt splits and the metal ion moves to the negative electrode and the negative ion moves to the positive electrode
e.g. if molten lead bromide is electrolysed, the lead will form at the negative electrode and bromine will form at the positive electrode

At the negative electrode (cathode)

At the negative electrode, positively charged ions **gain electrons** to become metal atoms.

This is classed as reduction. **Reduction** is gaining electrons

$\text{Na}^+ (\text{l}) + \text{e}^- \rightarrow \text{Na} (\text{s})$ (sodium ions become sodium atoms)

$\text{Cu}^{2+} (\text{l}) + 2\text{e}^- \rightarrow \text{Cu} (\text{s})$

$\text{Al}^{3+} (\text{l}) + 3\text{e}^- \rightarrow \text{Al} (\text{s})$

At the positive electrode (anode)

At the positive electrode, negatively charged ions **lose electrons**.

This is classed as **oxidation**. Oxidation is losing electrons

$2\text{Cl}^- (\text{l}) \rightarrow \text{Cl}_2 (\text{g}) + 2\text{e}^-$ (chloride ions becomes chlorine)

$2\text{Br}^- (\text{l}) \rightarrow \text{Br}_2 (\text{l}) + 2\text{e}^-$ (bromide ions becomes bromine)

$2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$ (iodide ions becomes iodine)

$2\text{O}^{2-} (\text{l}) \rightarrow \text{O}_2 (\text{g}) + 4\text{e}^-$ (oxide ions becomes oxygen)

OIL RIG can help remember that **Oxidation is Loss of electrons:**
Reduction is Gain of electrons

Extraction by electrolysis

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.

Aluminium is manufactured by the electrolysis of a molten mixture of aluminium oxide and cryolite. The mixture has a lower melting point than pure aluminium oxide.

Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current. This makes Aluminium an expensive metal

Aluminium forms at the negative electrode (cathode) and oxygen at the positive electrode (anode).

Half equation at negative electrode: $\text{Al}^{3+} (\text{l}) + 3\text{e}^- \rightarrow \text{Al} (\text{s})$ (reduction)

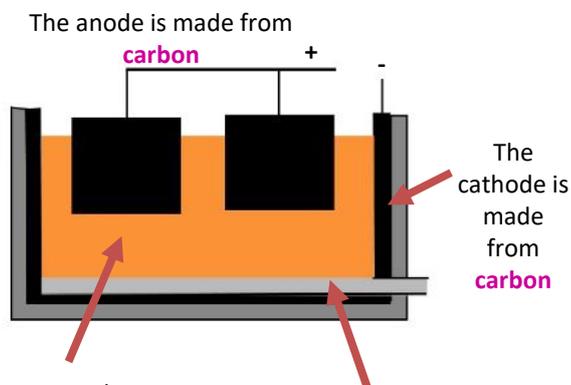
Half equation at positive electrode $2\text{O}^{2-} (\text{l}) \rightarrow \text{O}_2 (\text{g}) + 4\text{e}^-$ (oxidation)

The **positive** electrode (anode) is made of **carbon**, which reacts with the **oxygen to produce carbon dioxide** and so must be continually replaced

Cryolite is used in this process to lower the melting point of the aluminium oxide which therefore reduces the energy costs.

The electrolyte contains molten aluminium oxide dissolved in molten cryolite.

The aluminium ions are attracted to the cathode, gain three electrons to form liquid aluminium.



Electrolysing aqueous solutions

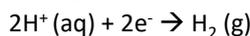
If an aqueous solution is electrolysed, using inert electrodes, the ions discharged depend on the reactivity of the elements involved.

In aqueous solutions there is a mixture of ions: H^+ and OH^- ions are present in addition to the ions from the salt e.g. in copper chloride solution there are H^+ , OH^- (from water) Cu^{2+} , Cl^- (from the salt)

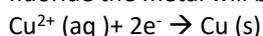
The negative electrode

At the negative electrode (cathode) the **less reactive** element **discharges**.

In **aqueous** solutions of **reactive** metal salts where the metal is above hydrogen in the reactivity series, the metal will not therefore be evolved at the cathode. (e.g. **sodium** chloride, **calcium** fluoride) **Hydrogen gas** will be evolved at the cathode instead.



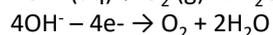
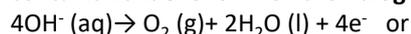
In **aqueous** solutions of **unreactive** metal salts e.g. **copper** chloride or **silver** fluoride the metal will be evolved at the cathode.



This happens because in the aqueous solution water molecules break down producing hydrogen ions and hydroxide ions that are discharged.

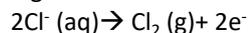
The positive electrode

At the positive electrode (anode), **oxygen** is produced unless the solution contains **halide ions** when the **halogen** is produced.



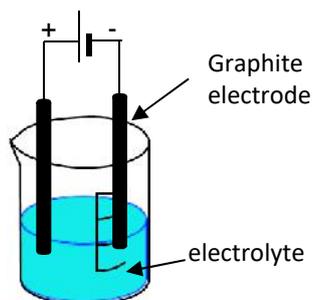
If a halide ion is present then the halogen is produced

e.g.



Required practical electrolysis

The products given off at the electrodes will depend on factors above



- If aqueous copper sulfate is used copper metal would form at the cathode (negative electrode) and the mass of the electrode would increase as copper forms on it.
- The blue colour of the electrolyte would fade as the copper ions are removed from the solution and reduced to copper atoms, so decreasing the concentration of copper ions in solution.
- Oxygen gas would form at the anode