WJEC GCSE CHEMISTRY

Triple Award

Year 11

Revision Guide

GCSE Chemistry

Year 11

REVISION GUIDE

CHEMISTRY TOPICS

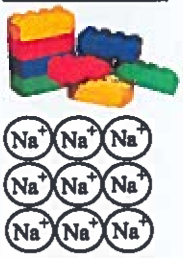
Topic 1:

Bonding, Structure and Properties

C2

Elements

Element

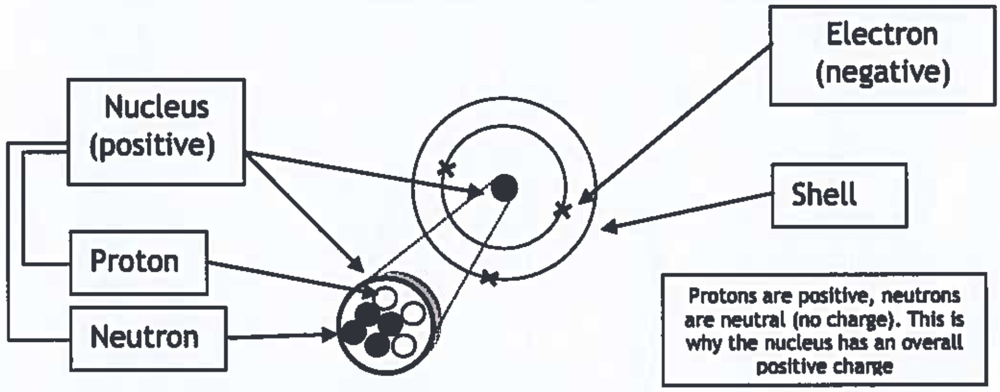


Elements are the building blocks of all substances. They cannot be broken down into simpler substances by chemical means

An Element contains only one type of atom

Atom

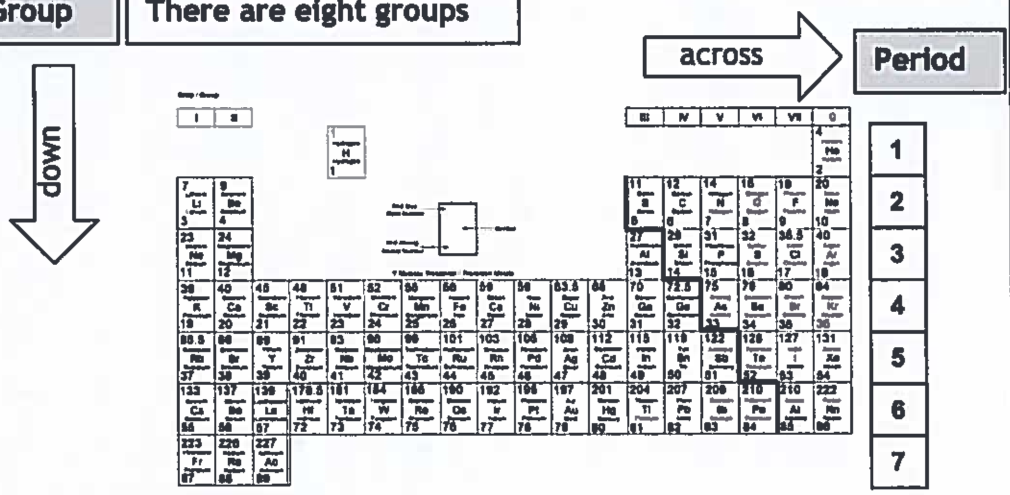
Each atom has negatively charged electrons orbiting a positively charged nucleus



Protons are positive, neutrons are neutral (no charge). This is why the nucleus has an overall positive charge

The Periodic Table - Basics

Group There are eight groups



Describing Position
Sodium is in Group 1, Period 3
Helium is in Group 0, Period 1
Beryllium is in Group 2, Period 2

Materials

The uses of everyday materials depends on their properties.

Aluminium

Low density

Ceramic

Hardness, strength

Diamond

Hardness, lusture

Aluminium Oxide

Low friction and wear

Teflon

Low friction high Mpt

Glass

Hardness transmits light

The properties of all materials, are determined by:-

the types of atoms present,
the types of bonding between the atoms,
and the way the atoms are packed together

Metals are giant structures with free electrons

Metallic bonds are strong, so metals can maintain a regular structure and usually have high melting and boiling points.

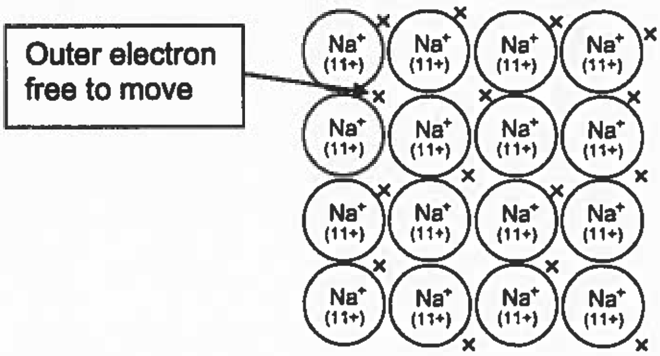
Outer shell electrons of metals are free to move.

The strength of bond in a metal is the force of attraction between the metal ions and free moving electrons.

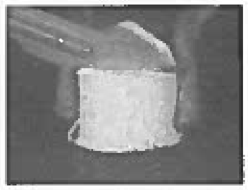
More free electrons and more protons in the ions increase the strength of a metal.

Metallic bonding

Free electrons allow electricity to be carried as well as heat energy.



Sodium metal

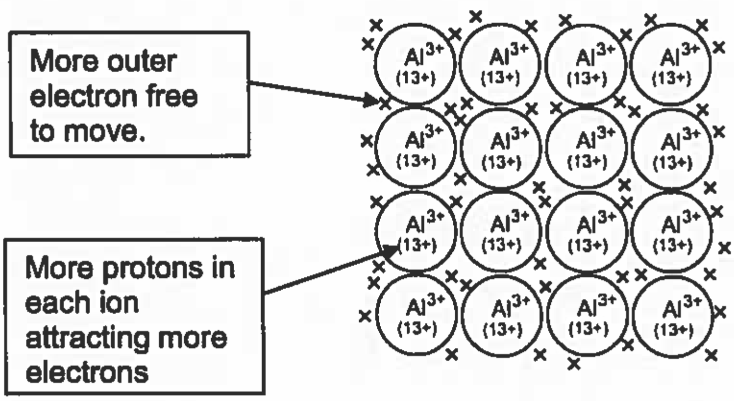


These electrons allow metals to conduct electricity and conduct heat.

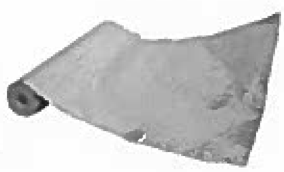
Metals are also malleable (hit into shape) and ductile (drawn into wires) because the free electrons allow the metal atoms to slide over each other.

Higher Tier:

Having more free electrons in the outer shell e.g. Aluminium compared to sodium above and more protons in each nucleus the forces of attraction for the free electrons is greater. This makes the metal stronger.



Aluminium metal

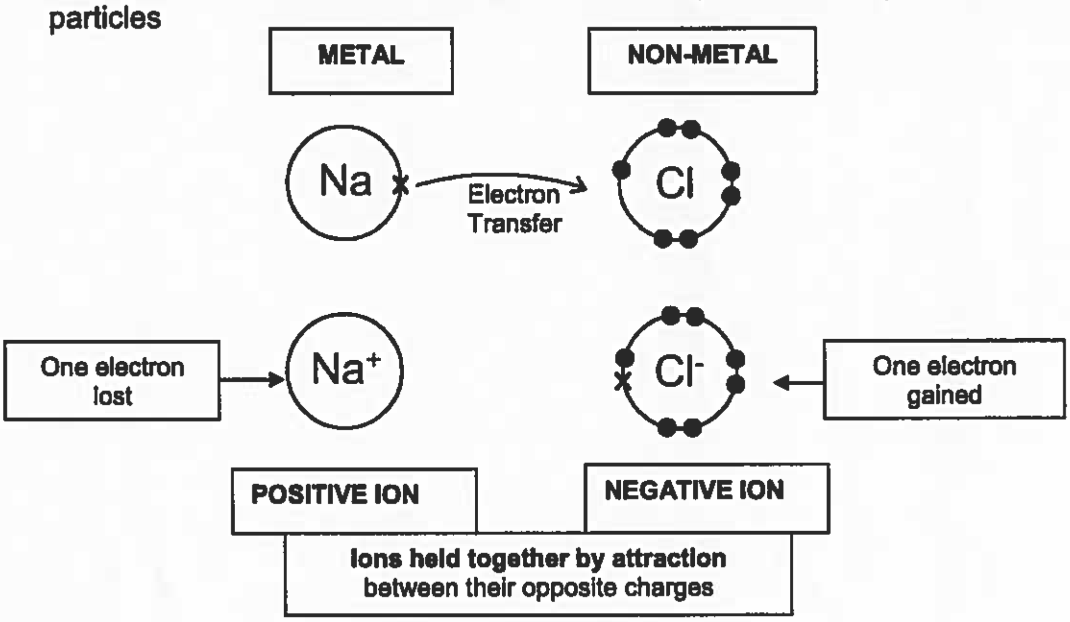


The Ionic Bond

Ionic Compounds

When a chemical reaction takes place new bonds are formed. Ionic compounds form by the **transfer of electrons** from metal to non-metal atom. Charged particles called **ions** are formed

e.g.
When sodium chloride (NaCl) forms, one electron is transferred from sodium to chlorine. This will form a **full stable outer shell** (like noble gases) for the two particles



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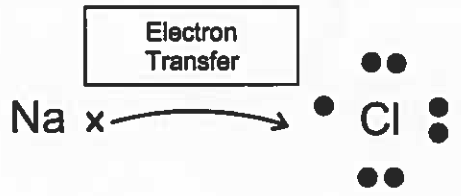
Ionic Bonding

Here is the information needed to draw ionic bonding diagrams.

Check the ionic charges by using the ions table at the back of the examination paper.

Sodium Chloride (NaCl) **Group 1 with Group 7**

Before Bonding

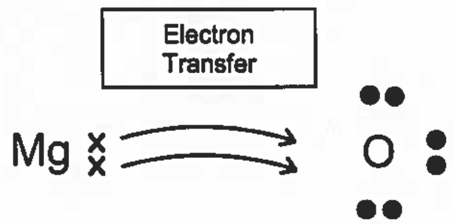


After bonding

Strong electrostatic charge between the ions

Magnesium oxide (MgO) **Group 2 with Group 6**

Before Bonding



After bonding

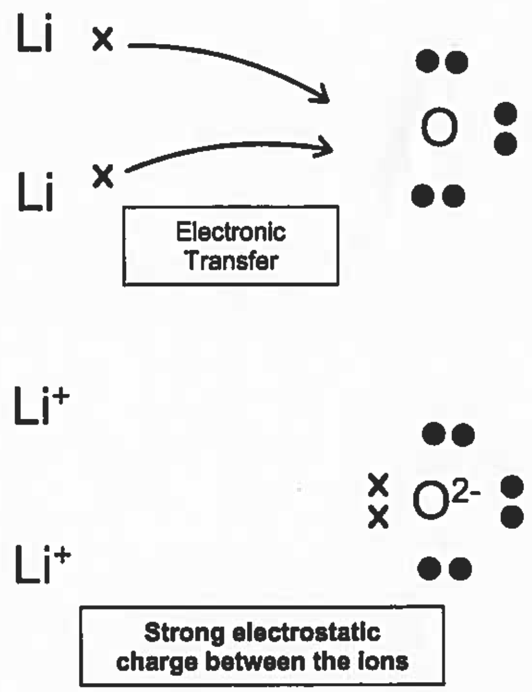
Strong electrostatic charge between the ions

MgO has a higher melting point than NaCl as the charges are greater – there is more attraction between ions

Lithium Oxide (Li_2O) **Group 1 with Group 6**

Before Bonding

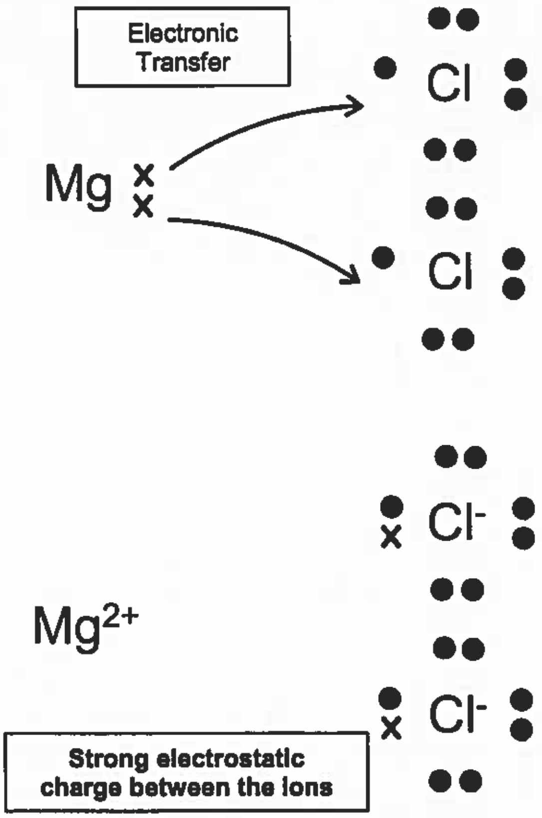
After bonding



Magnesium Chloride ($MgCl_2$) **Group 2 with Group 7**

Before Bonding

After bonding



Using ions to create formulae

Lithium = Li^+

Sodium = Na^+

Potassium = K^+

Magnesium = Mg^{2+}

Calcium = Ca^{2+}

oxide = O^{2-}

sulfide = S^{2-}

chloride = Cl^-

bromide = Br^-

iodide = I^-

Sodium Chloride

Na^+ Cl^-

Ions cancel

$NaCl$

Magnesium Oxide

Mg^{2+} O^{2-}

Ions cancel

MgO

Lithium Oxide

Li^+ O^{2-}

Li^+

Ions cancel

Li_2O

Magnesium Chloride

Mg^{2+} Cl^-

Cl^-

Ions cancel

$MgCl_2$

Hydroxide= OH^-

Sulfate= SO_4^{2-}

Carbonate = CO_3^{2-}

Nitrate = NO_3^-

Sodium Hydroxide

Na^+ OH^-

Ions cancel

$NaOH$

Magnesium Hydroxide

Mg^{2+} OH^-

OH^-

Ions cancel

$Mg(OH)_2$

Two sets of OH^- (brackets used)

Sodium Carbonate

Na^+ CO_3^{2-}

Na^+

Ions cancel

Na_2CO_3

Calcium Carbonate

Ca^{2+} CO_3^{2-}

Ions cancel

$CaCO_3$

Quick method

Lithium Oxide

Li^+ O^{2-}

Li_2O

Sodium Carbonate

Na^+ CO_3^{2-}

Na_2CO_3

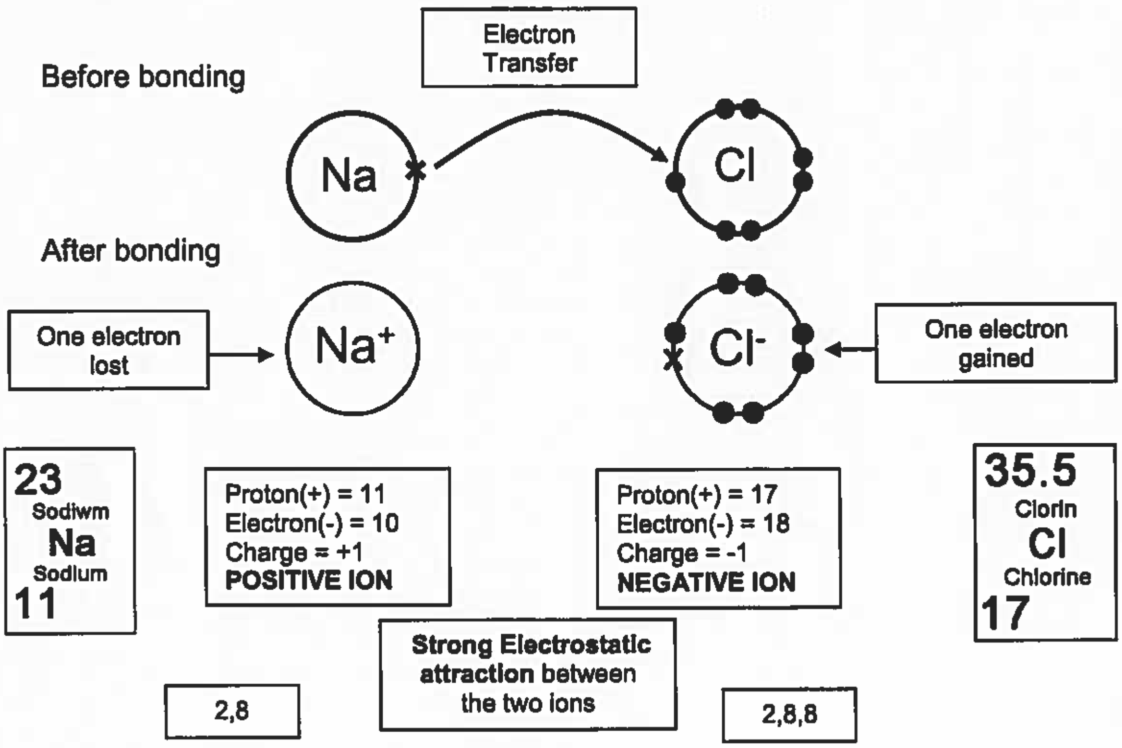
Bonding and Structure

When a chemical reaction occurs new bonds are formed. The can form by the **transfer of electrons** or by the **sharing of electrons.**

Ionic Bonding

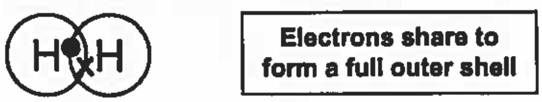
Charged particles called **ions** are formed when electrons are transferred between atoms during chemical bonding.

e.g.
When sodium chloride (NaCl) forms, one electron is transferred to chlorine. This will form a **full stable outer shell** (like noble gasses) for the two particles ('atom').



Covalent Bond

When hydrogen gas(H_2) forms electrons are shared between two atoms to form a molecule. There is **no charge** on molecules.



Covalent Bonds

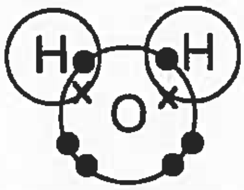
Hydrogen (H_2)

Electrons share to form a full outer shell

Chlorine (Cl_2)

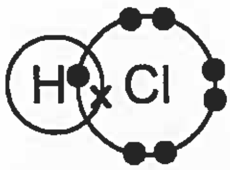
Electrons share to form a full outer shell

Water (H_2O)



Electrons share to form a full outer shell

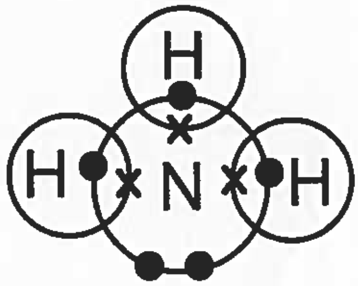
Hydrogen Chloride (HCl)



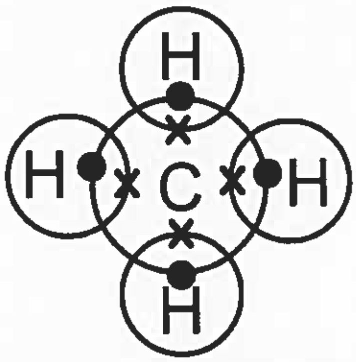
Electrons share to form a full outer shell

Covalent Bonding

Ammonia (NH_3)

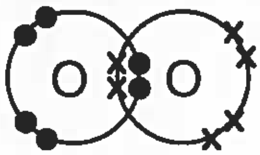


Methane (CH_4)

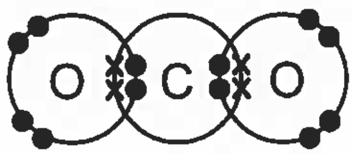


Covalent examples with double bonds (Higher Tier)

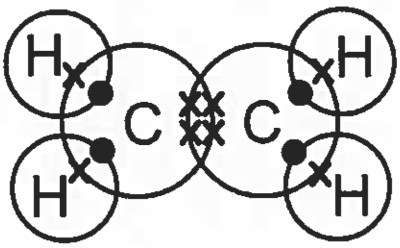
Oxygen (O_2)



Carbon Dioxide (CO_2)

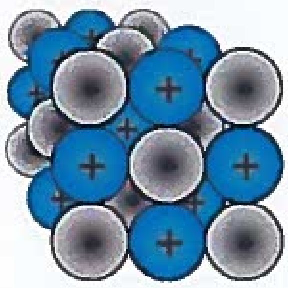


Ethene (C_2H_4)



Simple and Giant structures

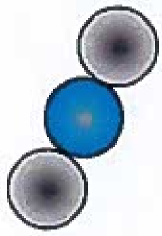
Giant ionic structure(e.g. sodium chloride, magnesium oxide),



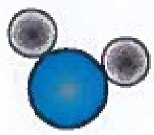
High Melting and Boiling points
Solubility = Dissolved in water

Simple molecular structure (e.g. carbon dioxide, water)

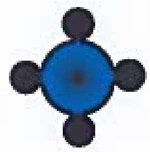
Low Melting and Boiling points
Solubility = Dissolved in water



CO_2



H_2O (water)

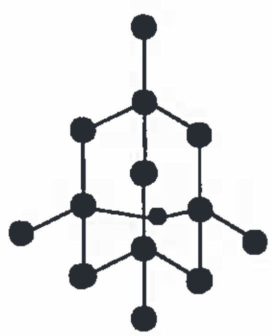


CH_4 (methane)

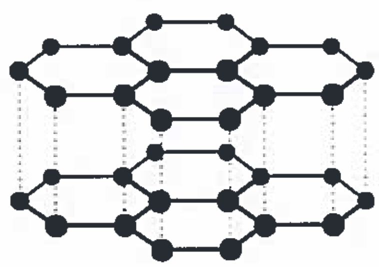
As the forces between molecules are weak the melting and boiling points are low

Giant covalent structure (e.g. diamond, graphite),

Very High Melting and Boiling points
Solubility = Does not dissolve in water

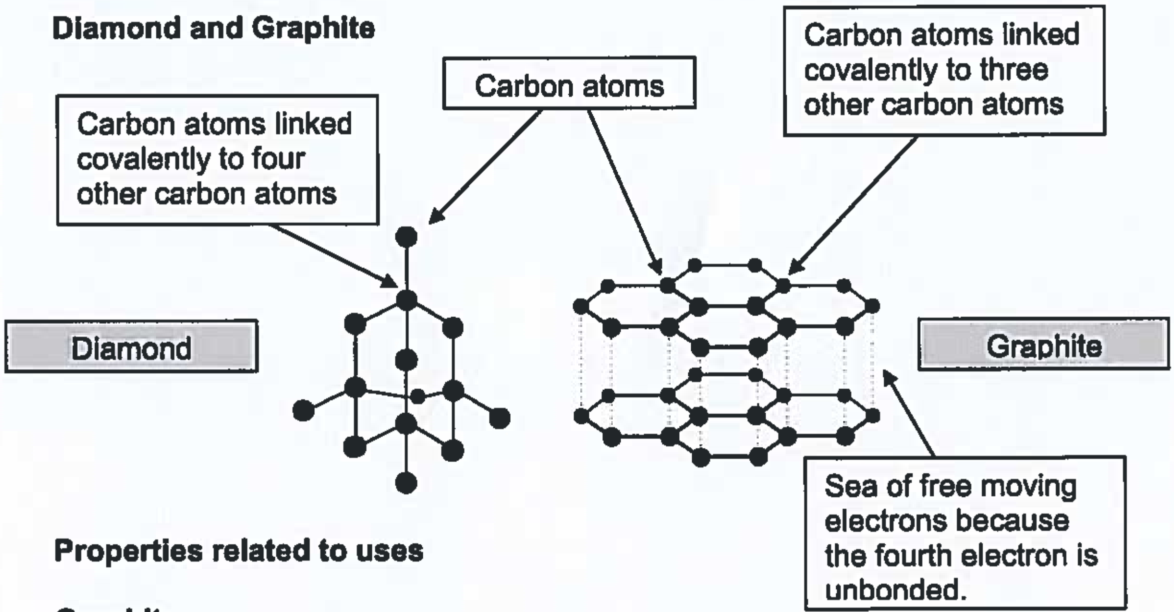


Diamond



Graphite

Diamond and Graphite



Properties related to uses

Graphite

Appearance	Grey/black shiny solid
Hardness	very soft
Conductivity	Conducts electricity
Melting point	Very high over 3600°C

Uses	Pensil lubricants

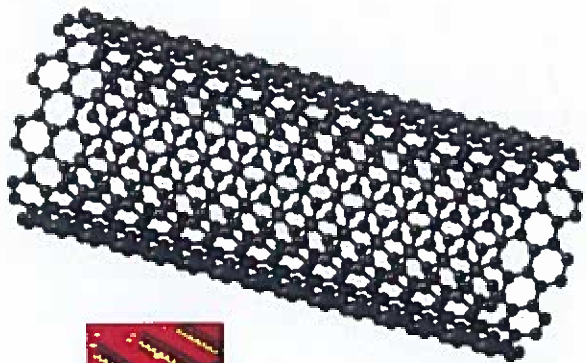
Diamond

Appearance	Transparent/crystalline
Hardness	very hard
Conductivity	Electrical insulator
Melting point	Very high over 3500°C

Uses
Gemstones
Glass cutting, Drill bits

Carbon nanotubes

- They are roles of carbon hexagons similar to graphite.
- They conduct electricity / used in semi-conductors
- The have a very small diameter which is about 10,000 times less than a human hair.
- They are extremely strong.
- Very low density
- They are proposed to be used in **small electronic circuits**

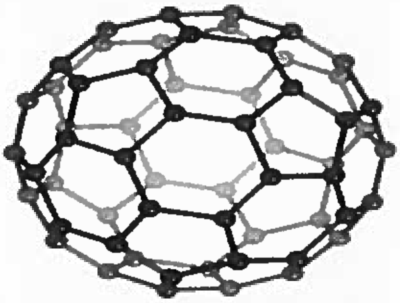


More Giant Covalent Structures

Fullerenes

- Fullerenes are another example of Giant Covalent structures containing **strong covalent bonds** between the carbon atoms.
- Fullerenes are **CAGE STRUCTURES** made completely out of **CARBON ATOMS.**

- **BUCKMINSTERFULLERENE** is the best known example



  of a fullerene. It contains 60 carbon atoms (formula C_{60}) and its molecules are spherical (ball-shaped).

- **Uses of fullerenes** are still being researched and developed:
  - To deliver and target drugs to different sites in the body in medical treatments.
  - Lubricants (e.g. to reduce friction make different parts of machinery move more easily).
  - Catalysts (to increase the rate specific reactions by decreasing the activation energy required for a reaction to take place).

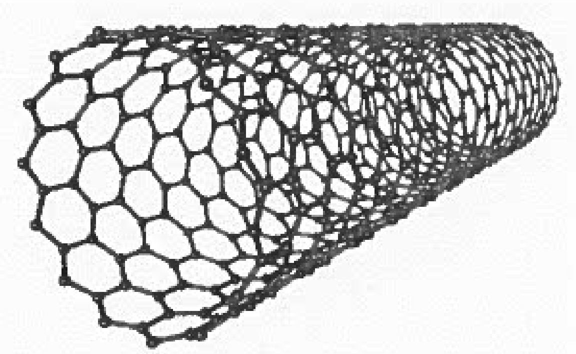
Carbon nano-tubes

- Carbon nano-tubes are made from single layers from a graphite structure called graphene.
- The single graphite layers are rolled into hollow tubes but the covalent bonds between the carbon atoms are very strong.
- Carbon nano- tubes are a lot smaller structures than carbon fibres, so they are <u>not</u> used to make carbon frames of bikes or tennis rackets (which can be made out of carbon fibres).

- Carbon nano-tubes have important useful properties:

  1. Good conductors/semiconductors of electricity because there are free electrons which are not used in covalent bonding and are able to move to conduct an electric current, e.g in touch screen devices like tablets and mobile phones.

  2. Very high strength but very low density (lightweight for their size). The tubes are 10,000 times smaller than human hair, but stronger than steel. Their high strength allows them to be used in bullet proof vests.



Carbon nano tube: single layer of carbon atoms (graphene layer) rolled into a hollow cylinder

Graphene

- Graphene is another example of a Giant Covalent structure containing strong covalent bonds between the carbon atoms.

- Evidence shows that graphene is the strongest material ever tested and the best electrical conductor.

- Some scientists and engineers believe that graphene will change and transform future technology, e.g. aerospace, automotive, electronics, energy storage, coatings & paints and communications. There are not any graphene products available yet to buy or sell (not yet commercially available.)

Nanoscience

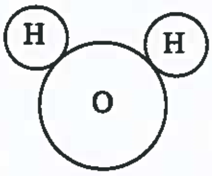
Scientists have a great interest in the nano range because the properties of materials can be different than when they are at a larger scale. The properties change from 100 nm downwards.

Comparing sizes in nanometre scale

80,000 nm wide is a piece of hair

7,000 nm o wide is 1 red blood cell

0.3 nm wide is 1 molecule of water



Many new materials are possible with this technology of building materials from atoms.

Uses which are made from nanotechnology

In sterilising sprays.

Silver particles of nano size are sprayed to kill bacteria, fungus and viruses

In fridges

A layer of silver atoms kill bacteria, fungus and viruses.

Nanoscience

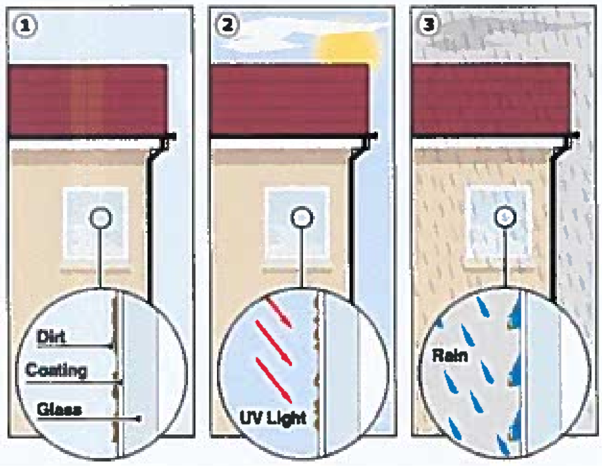
The new properties of these materials will allow people to create many new products.

Sun screen

There are nano particles in sun screens to prevent ultraviolet radiation damage to skin cells causing cancer.

Nano-sized TiO_2 and ZnO are used, they absorb and reflect UV light. Being transparent is appealing to customers

Self-cleaning glass



Self-cleaning glass is coated with nano-scale TiO_2 particles, which are hydrophobic (water repellent), dirt breaks down in sunlight and is washed away by rainwater.

Dangers with nano particles

Higher Tier

Although there are major benefits to nanoscience , nano particles could potentially harm humans and the environment.

Environmental and human experiments have to be performed on nano particles before they can be released commercially

As nano particles are so small and light the can move in the atmosphere. They can also move in rivers. These are methods by which nano particles can enter the body.

Dangerous nano particles can enter the blood stream

Smart Materials

Smart Materials

The term *smart material* has been given to a range of **modern** materials.

A variety of smart materials exist which can change shape and colour, retain shape after bending and can expand greatly with different liquids.

This means that their **properties** change.

The materials properties change with a change in the surroundings, such as changes in **temperature, light, pH.**

Thermochromic Paint

This smart material has the ability to change colour with a change in temperature.

The boat seen in this t shirt appears because thermochromic paint has been used. Under cold conditions the pigments are white, but when heated in warm weather, or if the person becomes warmer the pigments change colour to reviles a picture of a boat.

Photochromic Paint

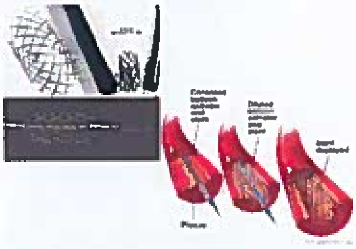
This material has the ability to change colour with a change in the light strength.

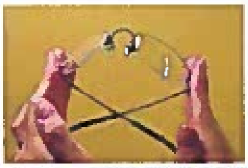
The sunglass lenses become darker when exposed to strong light and become lighter in weak light

Smart Materials

Shape memory alloy

This smart material is a mixture of metals (alloy) that retains its original shape when heated





A mixture of nickel and titanium make up the alloy called **NiTi or nitinol.**

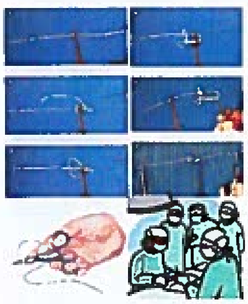
This metal can be bent into any shape at low temperature, but when heated it can remember its original shape so it bends back very quickly. It can be used as a **coffeepot thermostat.**

Stents are metal structures that can be inserted in veins to prevent them from sticking together. The stents are cooled to below 37C so they change shape and become thinner, when inserted into the vein it warms up to body temperature and changes shape to open the vein.

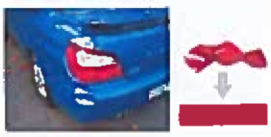
This alloy can also be used in super elastic **spectacle frames**. These retain their original shape after bending them.

Shape memory polymer

This smart material is a form of plastic that can retain its original shape when heated. These **could** be used for:-



Surgical stutures are threads of smart polymer that can tighten to the right tension automatically when heated.

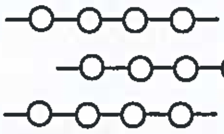


Car bumpers could be made from this material. If the car body such as bumpers were dented, on heating they would regain their original shape.

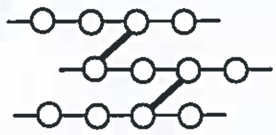
Smart Materials

Polymer gels

This smart material is a form of plastic with cross linkage (see diagram below) that can swell or shrink with different liquids.



No cross linkage across polymer



Cross linkage across polymer

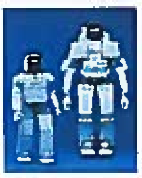
These gels can swell to **1000 times** their volume depending on the **temperature or pH**.

Artificial snow - this smart material expands greatly by adding water. It can also shrink by heating.

Nappies - this smart material is similar to artificial snow and expands greatly when it becomes wet.

Contact Lens - within these lenses there is a smart material which prevents them from drying up. They can then be used for weeks instead of days.

Artificial muscles - gels can be used to swell and shrink creating an artificial muscle.



Robot actuators - gels can be used to swell and shrink creating movement.

Toxic Chemical absorber - gels can be used to block dangerous chemicals in the body.

Topic 2:

Acids, Bases and Salts

C21

Acid Reactions

Sulfuric Acid — H_2SO_4 — Sulfate

Nitric Acid — HNO_3 — Nitrate

Hydrochloric Acid — HCl — Chloride

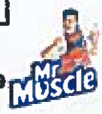
Form salts

Indicator

Universal Indicator (pH) = A substance that changes colour when added to an acidic, alkaline or neutral substance. The colour corresponds to the strength of the acid or alkali (e.g. strong or weak alkali)

0	1	2	3	4	5	6	7	8	9	10	11	12	13	14

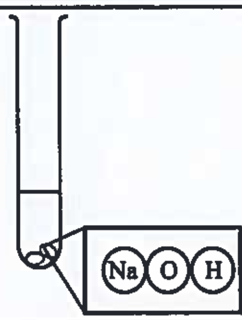
Strong Acid (0–2)	Weak Acid (3–6)	Neutral (7)	Weak Alkali (8–11)	Strong Alkali (12–14)
Sulfuric Acid	Orange	Pure water	Soap	Oven cleaner Sodium Hydroxide



Base — Metal oxide or metal hydroxide

Most are insoluble in water

Alkali — A water soluble base



NEUTRALISATION REACTIONS

1. Acid + Alkali

When the correct amount of acid and alkali are added together a neutral solution is made

ACID + ALKALI ⟶ SALT + WATER

Hydrochloric Acid + Sodium Hydroxide ⟶ Sodium chloride + Water

$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$

Sulfuric Acid + Sodium Hydroxide ⟶ Sodium sulfate + Water

$H_2SO_4\ (aq) + 2NaOH\ (aq) \longrightarrow Na_2SO_4\ (aq) + H_2O\ (l)$

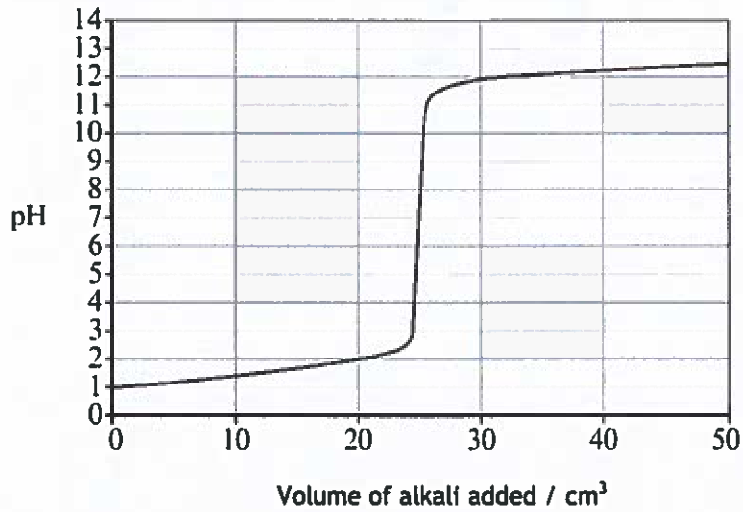
Nitric Acid + Sodium Hydroxide ⟶ Sodium nitrate + Water

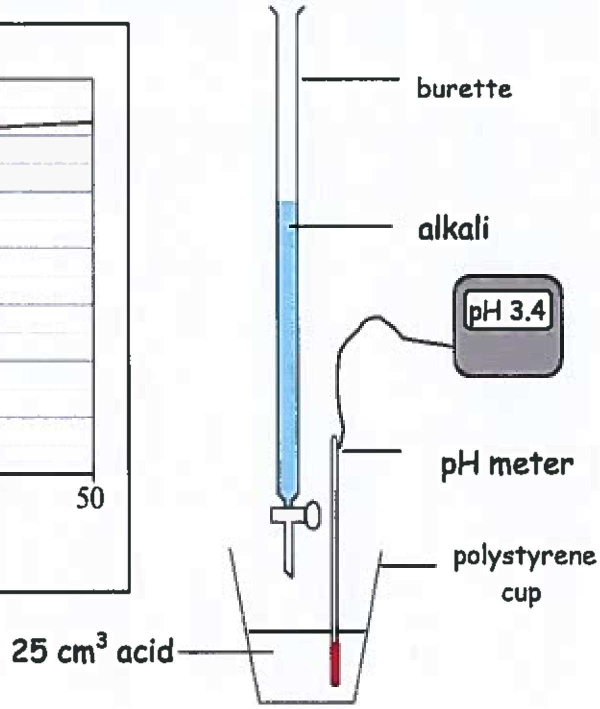
$HNO_3\ (aq) + NaOH\ (aq) \longrightarrow NaNO_3\ (aq) + H_2O\ (l)$

Investigating a Neutralisation Reaction

A pH sensor can be used to monitor a neutralisation reaction; in the reaction below alkali (potassium hydroxide) is added slowly to 25 cm³ acid

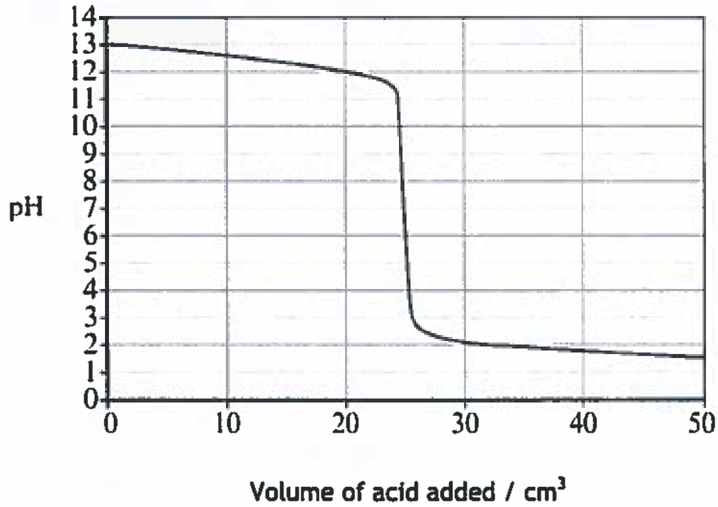
Alkali added to Acid





0 - 24 cm³ - solution is acidic
25.00 cm³ - neutralisation point
26- 50 cm³ - solution is alkaline (too much alkali added)

Acid added to Alkali



0 - 24 cm³ - alkaline
25.00 cm³ - neutralisation point
26- 50 cm³ - acidic

If too much acid (excess) is added the substance will be acidic at the end.

If the correct volume is added (25 cm³) the solution becomes neutral

REMEMBER - All neutralisation reactions are exothermic (heat is released)

3. Acid + Base

ACID + BASE → SALT + WATER

Sulfuric Acid + Copper oxide → Copper sulfate + Water

H_2SO_4 (aq) + CuO (s) → $CuSO_4$ (aq) + H_2O (l)

Hydrochloric Acid + Copper oxide → Copper chloride + Water

2HCl (aq) + CuO (s) → $CuCl_2$ (aq) + H_2O (l)

2. Acid + Carbonate

CO_2 is made in addition to salt and water

ACID + Carbonate → SALT + WATER + CARBON DIOXIDE

Sulfuric Acid + Copper Carbonate → Copper sulfate + Water + Carbon Dioxide

H_2SO_4 (aq) + $CuCO_3$ (s) → $CuSO_4$ (aq) + H_2O (l) + CO_2 (g)

Sulfuric Acid + Sodium Carbonate → Sodium sulfate + Water + Carbon Dioxide

H_2SO_4 (aq) + Na_2CO_3 (s) → Na_2SO_4 (aq) + H_2O (l) + CO_2 (g)

Hydrochloric Acid + Sodium Carbonate → Sodium chloride + Water + Carbon Dioxide

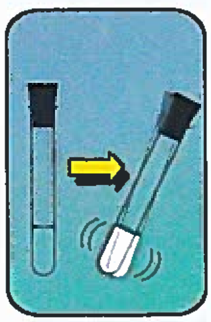
2HCl (aq) + Na_2CO_3 (s) → 2NaCl (aq) + H_2O (l) + CO_2 (g)

Carbonate test

When acid reacts with a carbonate fizzing is observed. Bubbles are seen as CO_2 is a gas

Carbon dioxide test

If clear limewater turns milky there is carbon dioxide present.



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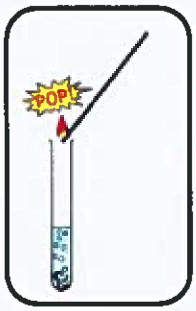
4. Metal + Acid

A reactive metal produces hydrogen with acids

METAL	+	ACID	→	SALT	+	HYDROGEN
Magnesium	+	hydrochloric acid	→	Sodium chloride	+	hydrogen
Mg (s)	+	2HCl (aq)	→	$MgCl_2$ (aq)	+	H_2 (g)
Zinc	+	sulfuric acid	→	zinc sulfate	+	hydrogen
Zn (s)	+	H_2SO_4 (aq)	→	$ZnSO_4$ (aq)	+	H_2 (g)

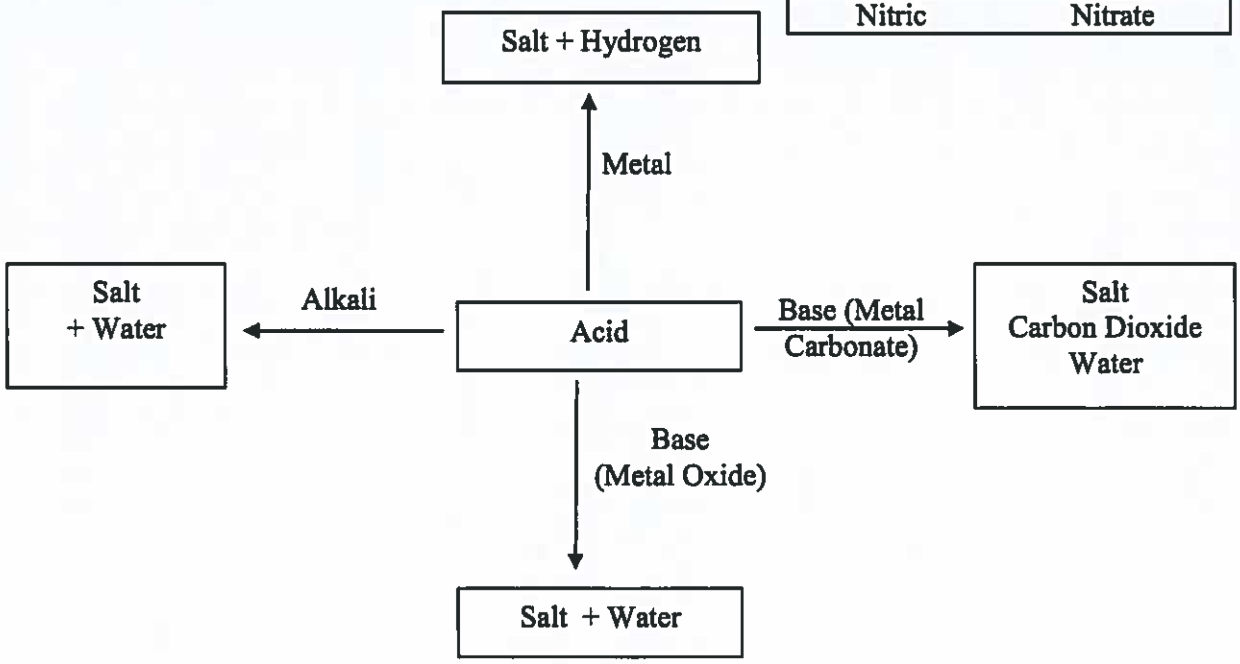
Hydrogen Test

If a lighted splint is placed in hydrogen it will create a squeaky 'pop' sound.



Acid Reactions Summary

Acid	Salt
Hydrochloric	Chloride
Sulfuric	Sulfate
Nitric	Nitrate



C25

Using the Reactivity Series of Metals to explain the reactions of metals with dilute acids

Reactivity Series of Metals

Metal	Reactivity	Reaction of metal with dilute acids
Potassium	Most reactive metal	**All these metals react with dilute acids to form *HYDROGEN* gas because they are more reactive than hydrogen in the Reactivity Series and can displace (take the place of) the hydrogen ion in the acid.**
Sodium	↓	
Calcium		
Magnesium		
Aluminium	Reactivity DECREASES as you descend (go down) this list of metals	
Carbon (non-metal)		
Zinc		
Iron		
Lead	↓	
Hydrogen (non-metal)		
Copper		These metals do not react with dilute acids because they are less reactive than hydrogen and cannot displace the hydrogen ion from the acid.
Silver		
Gold	Least reactive metal	

Which ions are contained in acid and alkalis?

Acids

- **Solutions of acids contain hydrogen ions (shown using the symbol: H^+)**

- **Acids are a source of hydrogen ions (H^+).**

Examples of acids

Acid	*Formula*
hydrochloric acid	HCl
sulphuric acid	H_2SO_4
nitric acid	HNO_3
ethanoic acid	CH_3COOH

Alkalis

- **Solutions of alkalis contain hydroxide ions (shown using the symbol OH^-).**

- **Alkalis are a source of hydroxide ions (OH^-)**

Examples of alkalis

Alkali	*Formula*
sodium hydroxide	$NaOH$
potassium hydroxide	KOH
calcium hydroxide (limewater)	$Ca(OH)_2$

Neutralisation (Higher Tier Only)

- When acids and alkalis react together they form a neutral solution (pH 7). This is called a neutralisation reaction.

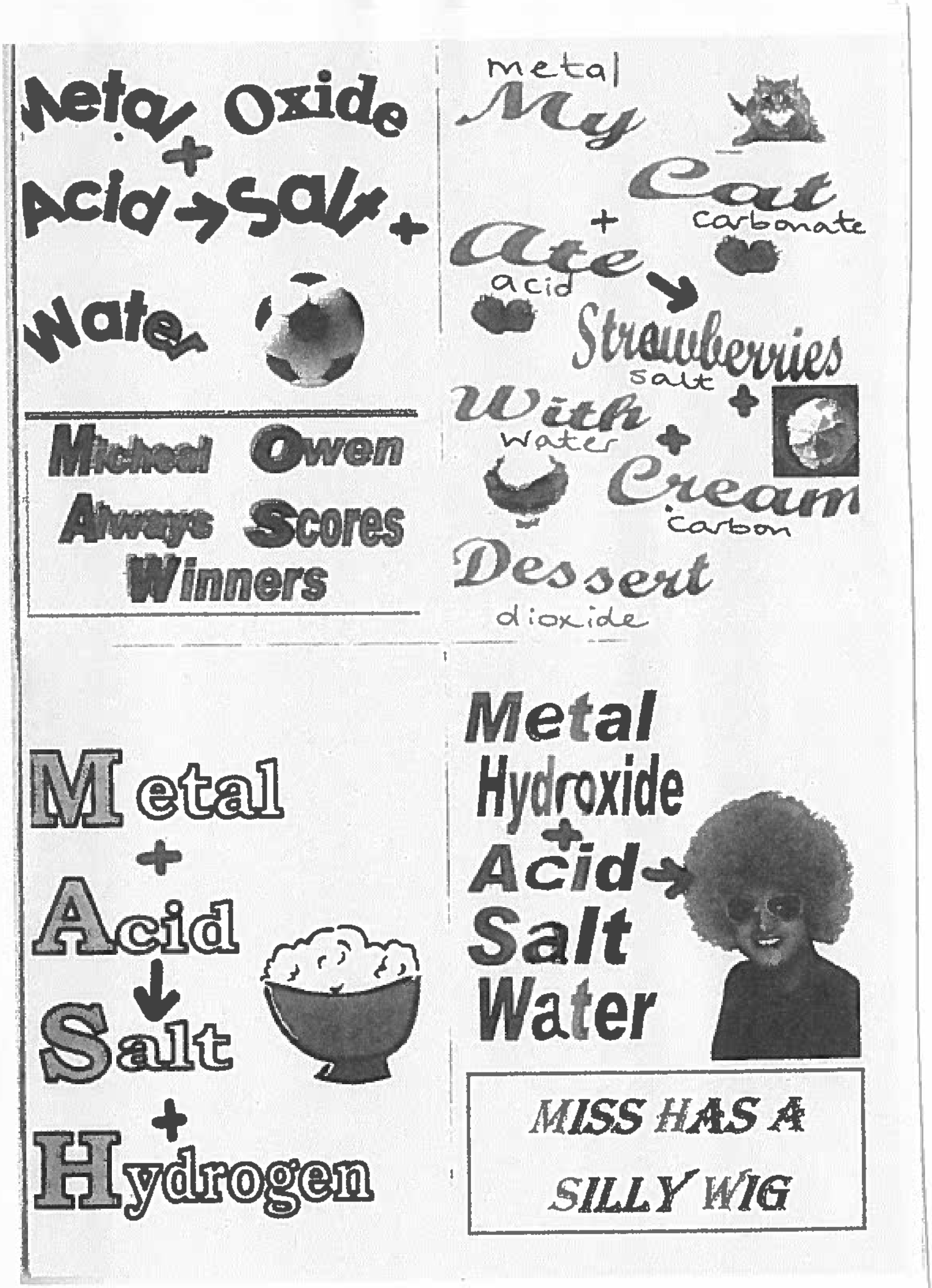
- In a **neutralisation reaction the hydrogen ions (H^+) react with the hydroxide ions (OH^-) to form water (H_2O).** Water has a neutral pH = pH 7.

- The **ionic equation for neutralisation reactions is shown as:**

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

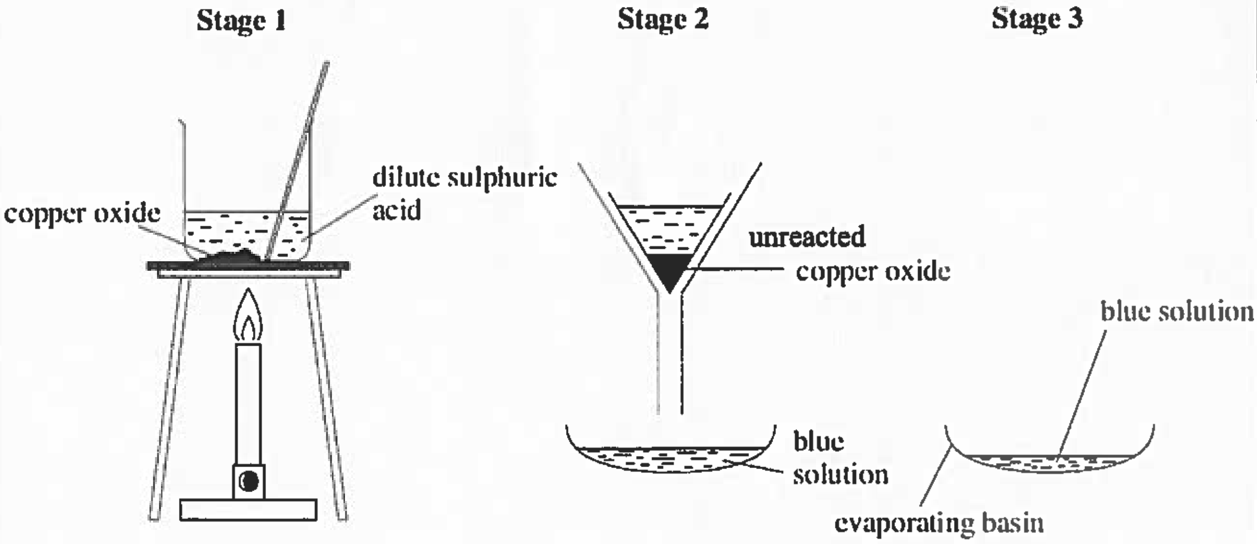
Making Metal Salts

Mnemonics to help you learn the four general equations that can be used to make metal salts



Method of preparing salt crystals

The method below is used to obtain salt from metal oxides and carbonates



Stage 1: Excess base (copper oxide) is added to the dilute acid to make sure all the acid has been reacted and used up. Heat and stirring will assist the process

Stage 2: The excess (unreacted) base is removed by the process of filtration, using a filter funnel and filter paper

Stage 3: Salt is obtained by evaporation - water evaporates and crystals of salt left behind. Water can be evaporated slowly near a window or with additional heating using a Bunsen Burner, 1/3 of the solution should be left behind to evaporate naturally.

Obtaining salt from the metal and acid reaction

The only difference in the method is stage 1 - excess metal is used - to make sure all the acid has been used up

The mole

The mole is a term that describes a specific number - like the word 'dozen' represents the number 12. The mole however is a much larger number 6.02×10^{23} atoms. (6 followed by 23 zeros).

A mole is defined as the number of atoms in exactly 12 grams of ^{12}C. 12 is the mass number of carbon, so one mole of carbon atoms has a mass of 12 grams.

Mole of atoms

The mass of one mole of the atoms of any element is its 'relative atomic mass' in grams

Relative Atomic Mass (RAM)
One mole of Lithium has a mass of 7 g

$^{7}_{3}Li$

Element	R.A.M (Ar)	Mass of 1 mole
H	1	1g
C	12	12g
Ne	20	20g
Mg	24	24g
Ca	40	40g

Moles of molecules

Example: Hydrogen gas exists as H_2 molecules. Because the molecule contains two atoms of hydrogen it has a Relative Molecular Mass (Mr) of 2 (1 x 2). Therefore one mole of hydrogen gas has a mass of 2 g

Molecule	R.M.M (Mr)	Mass of 1 mole
H_2	1 x 2 = 2	2g
O_2	16 x 2 = 32	32g
N_2	14 x 2 = 28	28g
Cl_2	35.5 x 2 = 71	71g

Moles of compounds

To find the mass of one mole of a compound add up the RAMs of the elements of the compound taking account of the formula. The number you get is the Relative Formula Mass which is the Mr of the compound.

For example CO_2

1 x Carbon atom =	1 x 12 =	12
2 x oxygen atoms =	2 x 16 =	32
TOTAL		44

Compound	R.M.M (Mr)	Mass of 1 mole (Molar Mass)
NaCl	23 + 35.5	58.5g
$MgCl_2$	24 + (35.5x2)	95g
$CaCO_3$	40 + 12 + (16x3)	100g

The molar mass of a compound is its relative molecular (formula) mass expressed in grams

Converting Mass into Moles

Moles of atoms = $\frac{\text{mass}}{\text{Mr}}$

Moles of atoms = amount of substance
Mass = mass in grams
Mr = Molecular mass

How many moles of atoms are there in;
4.8 g of carbon ?

moles = $\frac{4.8}{12}$ = 0.4 moles

Ar (C) = 12

How many moles are there in; 640 g of oxygen (O_2) ?

moles = $\frac{640}{32}$ = 0.4 moles

Ar (O) = 16, Mr(O_2) = 16 x 2 = 32

How many moles are there in; 10 g of $CaCO_3$?

moles = $\frac{10}{100}$ = 0.1 moles

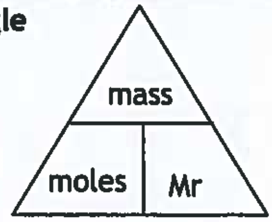
Ar (Ca) = 40, (C) = 12, (O) = 16
Mr $CaCO_3$ = 100

Converting Moles into mass

You can rearrange the equation to form

mass = moles x Mr

If you find rearranging difficult you can use the triangle



What is the mass of 0.05 moles of carbon atoms?

mass = 0.05 x 12 = 0.6 g

Ar (C) = 12

What is the mass of 0.6 moles of chlorine molecules (Cl_2)?

mass = 0.6 x 71 = 42.6 g

Ar (Cl) = 35.5, Mr(Cl_2) = 35.5 x 2 = 71

What is the mass of 0.1 moles of calcium carbonate?

mass = 0.1 x 100 = 10.0 g

Ar (Ca) = 40, (C) = 12, (O) = 16
Mr $CaCO_3$ = 100

It is also possible to calculate the molar mass Mr when mass and the number of moles are known

Mr = $\frac{\text{mass}}{\text{moles}}$

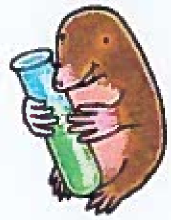
0.5 moles of a compound weighs 80g, calculate its Mr

Mr = $\frac{80}{0.5}$ = 160

Moles and Concentration Higher Tier only

All bottles of solutions in a laboratory must be labelled with its concentration

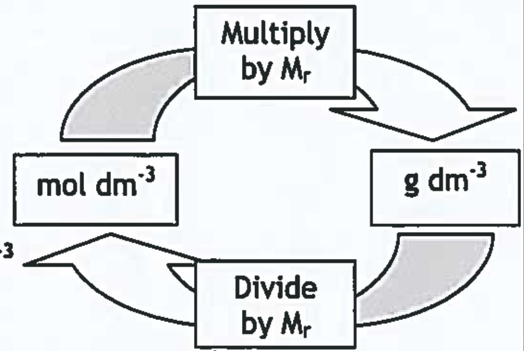
Concentrations are normally stated in mol dm^{-3}



A 1 mol dm^{-3} solution contains 1 mole of the substance dissolved in 1 dm^3 (or 1000 cm^3)

Preparing a 1 mol dm^{-3} solution is easy.
You weigh out its molecular mass and dissolve in 1000 cm^3 water

Compound	R.M.M	Mass in 1 mol dm^{-3} solution
NaCl	23 + 35.5	58.5g

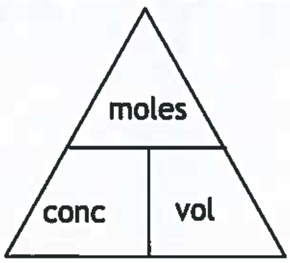


Concentration can sometimes be written as g dm^{-3}
To convert a value from mol dm^{-3} to g dm^{-3}
Multiply by Mr
e.g. 0.01 mol dm^{-3} of NaCl = 0.01 x 58 = **0.58 g dm^{-3}**

To convert a value from g dm^{-3} to mol dm^{-3}
Divide by Mr. e.g. 0.58 g dm^{-3} of NaCl = 0.58 / 58 = **0.01 mol dm^{-3}**

If we know the concentration and volume we can calculate the number of moles in any solution

$$\text{moles} = \text{concentration} \times \frac{\text{volume (in cm}^3\text{)}}{1000}$$



How many moles of sodium chloride are there in 200 cm^3 of a 2.0 mol dm^{-3} solution?

moles = $2.0 \times \frac{200}{1000}$ = **0.4 moles**

What is the concentration of a 100 cm^3 0.05 mole solution?

First we need to rearrange the original equation:

concentration = $\frac{\text{moles}}{\text{volume}/1000}$

$\frac{0.05}{0.1}$ = **0.1 mol dm^{-3}**

A solution of concentration 1 mol dm^{-3} contained 0.25 mole. What was the volume of the solution?

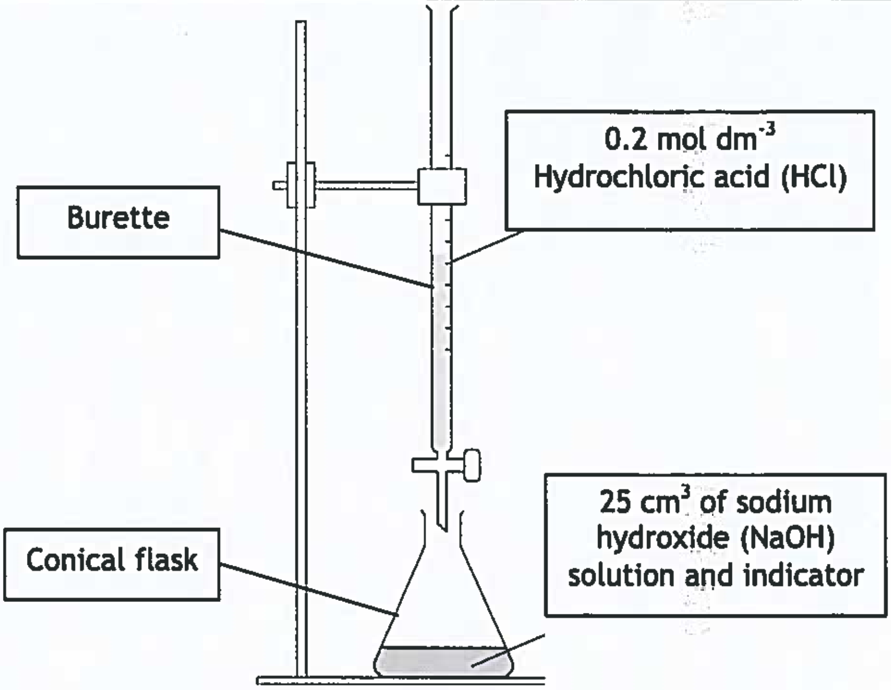
volume = $\frac{\text{moles}}{\text{concentration}}$ = $\frac{0.25}{1}$ = 0.25 dm^3 x 1000 = **250 cm^3**

(x1000 to get the answer in cm^3)

Titrations

You can find out the number of moles that react together in solutions using a technique called **titration**.
In titration experiments, you use a burette to add one solution to another.
You need a way to decide when the reaction is complete - The **end point** of reaction.
Indicators are used to determine end points in acid/alkali titrations.

Example: A student was asked to carry out a titration to check the concentration of the sodium hydroxide using the apparatus below

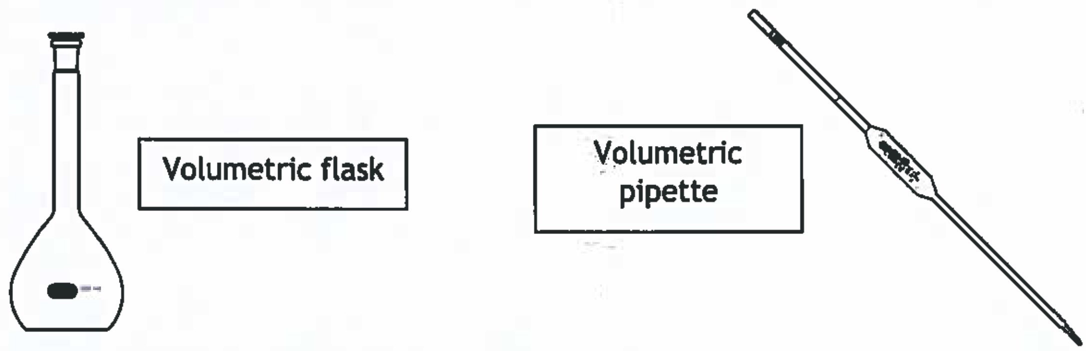


Methyl orange in alkali

End point - neutral

Some practical points:

- When solution are made up - they are made up using a **volumetric flask**
- A **volumetric pipette** is used to accurately measure the solutions that go into the conical flask
- When you fill the burette make sure that you remove the funnel afterwards - it could affect the level of solution in the burette and your results.
- A white tile is placed under the flask so that the colour change is easier to see.



- In the example the acid is added slowly by means of a burette, the volume of acid needed to change the indicator colour is recorded.
- It is easy to overshoot the end point the first time (turning the mixture acidic rather than neutral); the first titration is normally a practice run.

Results:

	Titration				
	1	**2**	**3**	**4**	**Average**
Volume of Hydrochloric acid added (cm^3)	**23.50 ***	**20.00**	**20.05**	**19.95**	**20.00**

Calculating the average= Titration $\frac{2+3+4}{3} = \frac{20.00 + 20.05 + 19.95}{3} = 20.00\ cm^3$

* Titration 1 not used in average as it is not reliable (practice run)

Calculating the concentration of the sodium hydroxide solution (Higher Tier only)

Step 1: Write the balanced equation for the reaction

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

Step 2: Gather the information

	NaOH	HCl
concentration	?	0.2
volume	25 / 1000	20 / 1000
moles	0.004 (step 4)	0.004 (step 3)

Step 3: Calculate the number of moles of HCl used

moles = concentration x $\frac{\text{volume (in cm}^3\text{)}}{1000}$

Moles (HCl) $= 0.2 \times \frac{20}{1000} = 0.004$ moles

Step 4: Check the mole ratio

One mole of HCl reacts with one mole of NaOH 1:1

As 1: 1 ratio; 0.004 mole of HCl reacts with 0.004 mole NaOH

Step 5: Calculate the concentration of NaOH

Rearranging gives;

conc = $\frac{\text{moles}}{\text{volume/1000}}$

Concentration = $\frac{0.004}{25/1000} = 0.16\ mol\ dm^{-3}$

Example 2: (Higher Tier only)

25.0 cm^3 of sodium hydroxide (NaOH) solution of unknown concentration was titrated with dilute sulfuric acid (H_2SO_4) of concentration 0.050 mol dm^{-3}. 20.0 cm^3 of the acid was required to neutralise the alkali. Find the concentration of the sodium hydroxide solution in mol dm^{-3}

Step 1: Write the balanced equation for the reaction

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$$

Step 2: Gather the information

	2NaOH	H_2SO_4
concentration	?	0.05
volume	25 / 1000	20 / 1000
moles	0.002	0.001

Step 3: Calculate the number of moles of H_2SO_4 used

$$\text{Moles } (H_2SO_4) = 0.05 \times \frac{20}{1000} = 0.001 \text{ moles}$$

Step 4: Check the mole ratio

<u>One mole</u> of H_2SO_4 reacts with <u>two moles</u> of NaOH   1 H_2SO_4 : 2 NaOH

Therefore 0.001 mole H_2SO_4 reacts with 0.002 (0.001 x 2) mole NaOH

Step 5: Calculate the concentration of NaOH

$$\text{Concentration} = \frac{0.002}{25/1000} = 0.08 \text{ mol dm}^{-3}$$

Brain teasers:

1. A student reacted 25 cm^3 of 1 mol dm^{-3} HCl with 25 cm^3 of 1 mol dm^{-3} NaOH, is the solution alkali, acidic or neutral ?
2. A student reacted 25 cm^3 of acid (HCl) with 20 cm^3 alkali (NaOH) to make a neutral solution, which had the highest concentration, the acid or the alkali?
3. A student reacted 10 cm^3 of acid (HCl) with 40 cm^3 alkali (NaOH) to make a neutral solution, which had the weakest concentration, the acid or the alkali?

Answers; 1 neutral, 2 alkali, 3 alkali

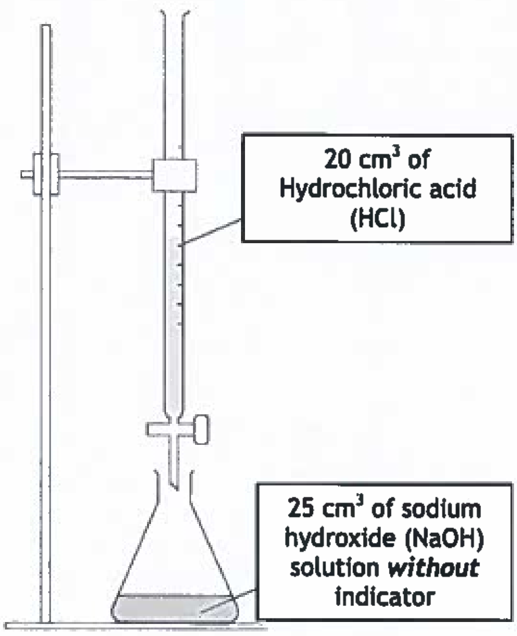
Using titration to prepare pure solutions

Once you have established the end point of a titration you can repeat the experiment without an indicator to obtain a pure sample of the compound.

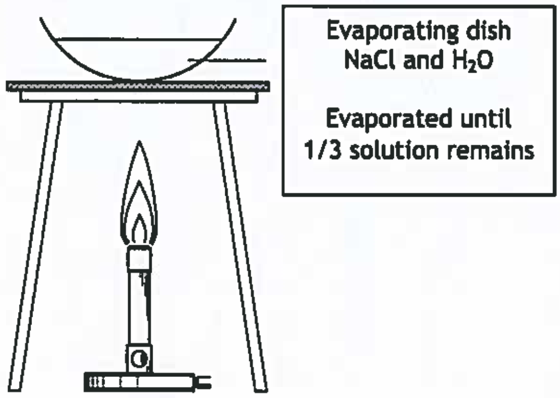
$$NaOH + HCl \longrightarrow NaCl + H_2O$$

In the reaction on page 23, using an indicator it was found that exactly 20.00 cm^3 of the hydrochloric acid was required to neutralise 25.00 cm^3 of the sodium hydroxide alkaline solution. The experiment could be repeated without an indicator to make a pure solution which could be further evaporated to give pure sodium chloride salt

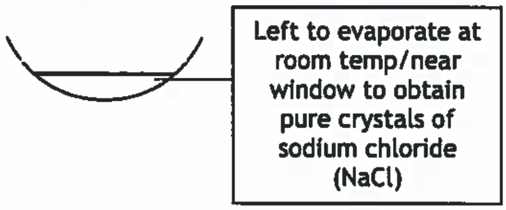
Step 1: Neutralisation



Step 2: Evaporation



Step 3: Evaporate slowly to dryness



Describing Acids

Concentrations of acids

- Any acid (or any solution) can be either **DILUTE** or **CONCENTRATED**. The concentration of an acid (or any solution depends on the amount (e.g. volume) of acid that is dissolved in water.

Strengths of acids

- Some types of acids are **STRONG ACIDS**, for example: hydrochloric acid, sulphuric acid and nitric acid. Their pH values are pH 0, 1 or 2.

- Strong acids have a high degree of ionisation to release the hydrogen ions (H^+) into the solution.

- Other acids are **WEAK ACIDS**, for example ethanoic acid and citric acid. Their pH range is pH 3 to pH 6.

- Weak acids have a lower degree ionisation to release the hydrogen ions (H^+) into the solution

The concentration and the strength of an acid are different ways of describing acids.

Examples

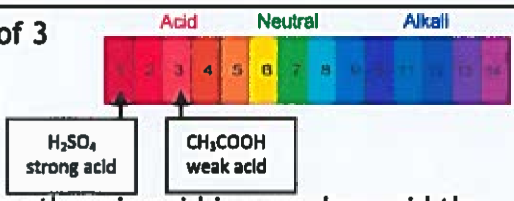
- 0.5 mol/dm^3 hydrochloric acid is a strong acid with a lower concentration of 0.5 mol/dm^3 .
- 2.0 mol/dm^3 hydrochloric acid is a strong acid with a higher concentration of 2.0 mol/dm^3 .

- 0.5 mol/dm^3 ethanoic acid is a weak acid with a lower concentration of 0.5 mol/dm^3 .
- 2.0 mol/dm^3 ethanoic acid is a weak acid with a higher concentration of 2.0 mol/dm^3 .

Comparing the reactions of Ethanoic Acid and Sulfuric Acid

Ethanoic Acid CH_3COOH is a <u>weak acid</u>, it has a pH of 3

Sulfuric Acid H_2SO_4 is a <u>strong acid</u>, it has a pH of 1



Ethanoic acid and sulfuric acid react similarly but as ethanoic acid is a weaker acid the **reactions are slower.**

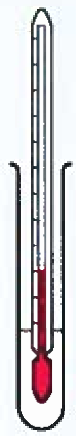
Ethanoic acid forms **ethanoate** salts. Sulfuric acid forms **sulfate** salts

ACID + ALKALI → SALT + WATER

sulfuric acid + sodium hydroxide → sodium sulfate + water

ethanoic acid + sodium hydroxide → sodium ethanoate + water

Observations: temperature rise in both reactions



ACID + BASE → SALT + WATER

sulfuric acid + copper oxide → copper sulfate + water

ethanoic acid + copper oxide → copper ethanoate + water

Observations: temperature rise in both reactions, base dissolves slower in ethanoic acid

ACID + CARBONATE → SALT + WATER + CARBON DIOXIDE

sulfuric acid + sodium carbonate → sodium sulfate + water + carbon dioxide

ethanoic acid + sodium carbonate → sodium ethanoate + water + carbon dioxide

Observations: Bubbles of CO_2 given off in both, more bubbles given off in the sulfuric acid reaction

METAL + ACID → SALT + HYDROGEN

magnesium + sulfuric acid → magnesium sulfate + hydrogen

magnesium + ethanoic acid → magnesium ethanoate + hydrogen

Observations: fizzing due to hydrogen gas in both reactions, more fizzing in the sulfuric acid reaction, Mg reacts faster in sulfuric acid

Conclusion: Both acids react similarly but the rate of reaction for ethanoic acid is slower as it is a weaker acid.

Chemical Analysis – negative ions

Testing for the halide ions

Add dilute nitric acid followed by silver nitrate

Nitric acid is added first to remove impurities. When silver nitrate is then added to a compound that contains a halide ion a precipitate is formed, the colour corresponds to the halide.

Non-metal	Ion	Silver Nitrate test
Chloride	Cl^-	white
Bromide	Br^-	cream
Iodide	I^-	yellow

Example equation

silver nitrate + sodium chloride ⟶ silver chloride + sodium nitrate

$AgNO_3 (aq) + NaCl (aq) \longrightarrow AgCl (s) + NaNO_3 (aq)$

ionic equation: $Ag^+ (aq) + Cl^- (aq) \longrightarrow AgCl (s)$

Testing for a carbonate ion CO_3^{2-}

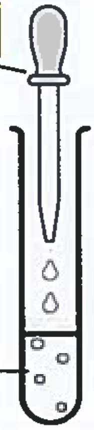
Add acid

When acid reacts with a carbonate **fizzing** is observed. Bubbles are of CO_2 gas which turns limewater milky

Example equations

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

$Na_2CO_3 (aq) + 2HCl (aq) \longrightarrow 2NaCl (aq) + H_2O (l) + CO_2 (g)$



bubbles / fizz

Testing for a sulfate ion SO_4^{2-}

Add dilute hydrochloric acid followed by barium chloride

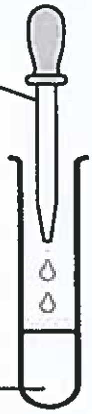
When barium chloride is added to a compound that contains a sulfate ion a white precipitate is seen - the white precipitate is barium sulfate

Example equations

barium chloride + sodium sulfate ⟶ sodium chloride + barium sulfate

$BaCl_2 (aq) + Na_2SO_4 (aq) \longrightarrow 2NaCl (aq) + BaSO_4 (s)$

$Ba^{2+} (aq) + SO_4^{2-} (aq) \longrightarrow BaSO_4 (s)$



white precipitate forms

Topic 3:

Metals and their Extraction

Extraction of Metals

Ores – Metals are found in compounds in rocks which make up the Earth's crust, these are called ores

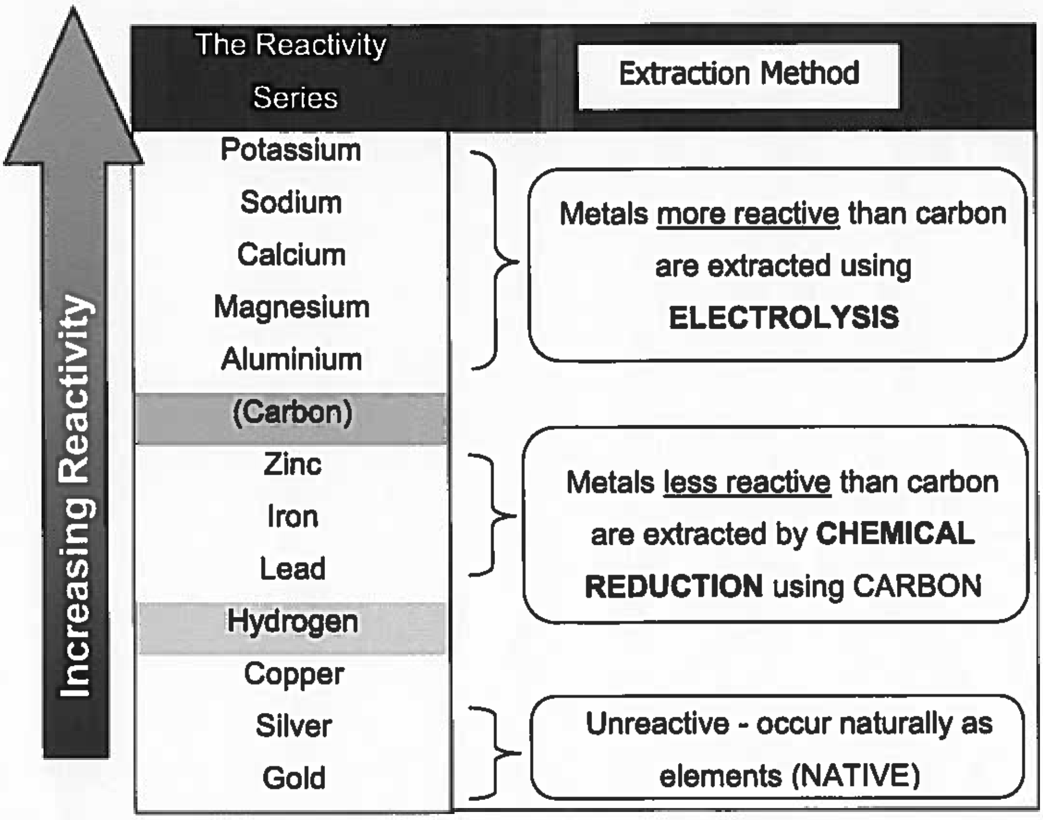
Ore	Formula	Metal extracted
Bauxite	Al_2O_3	Aluminium
Haematite	Fe_2O_3	Iron

Extraction is the term for getting pure metal out of the ore; there are two methods of extracting metals which depend on their reactivity

Reduction is the process of removing oxygen from the ore using carbon

Electrolysis is the process of using electricity to extract a metal

Reactivity Series – metals are placed in order of reactivity by reacting them with oxygen, water and acid. From this data a reactivity series is produced

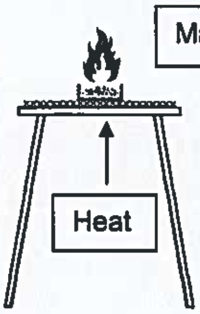


At the top metals naturally bond to oxygen stronger which makes it difficult to remove.

Displacement Reactions

Reduction is the loss of oxygen from a compound

Oxidation is the gain of oxygen to form a compound



Magnesium and copper oxide

$Mg + CuO \longrightarrow MgO + Cu$

oxidation

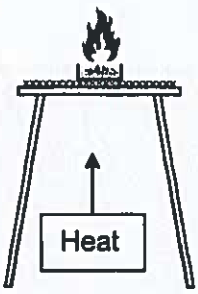
Magnesium + copper oxide ⟶ magnesium oxide + copper

reduction

Blast Furnace Reaction

Iron oxide and carbon monoxide

$Fe_2O_3 + 3CO \longrightarrow 3CO_2 + 2Fe$



oxidation

iron oxide + carbon ⟶ carbon monoxide + iron

reduction

The Thermit Reaction

$Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$

oxidation

Iron oxide + aluminium ⟶ aluminium oxide + iron

reduction

Displacement Examples

Iron and copper chloride	$Fe + CuCl_2 \longrightarrow Cu + FeCl_2$

iron + copper chloride ⟶ copper + iron chloride

iron is more reactive than copper, as a result iron displaces copper

copper and silver nitrate*	$Cu + 2AgNO_3 \longrightarrow 2Ag + Cu(NO_3)_2$

copper + silver nitrate ⟶ silver + copper nitrate

copper is more reactive than silver, as a result copper displaces silver

zinc and copper sulphate*	$Zn + CuSO_4 \longrightarrow Cu + ZnSO_4$

zinc + copper sulfate ⟶ copper+ zinc sulfate

zinc is more reactive than copper, as a result zinc displaces copper

* higher tier only

The Blast Furnace - The extraction of iron

There are 4 raw materials; iron ore, coke, limestone and hot air

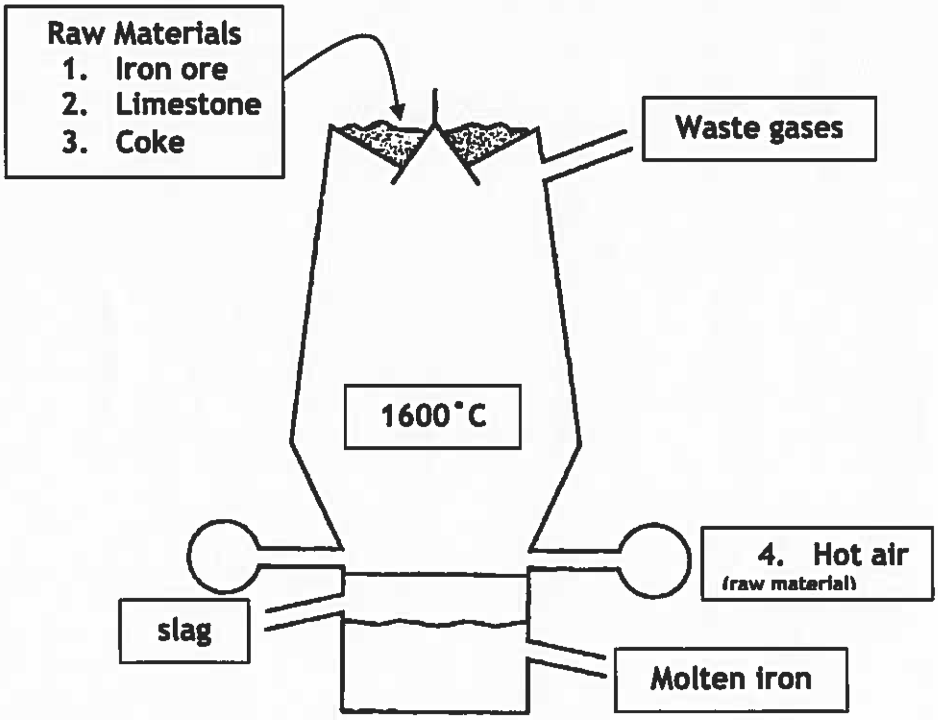
Iron ore - the source of iron

Limestone - to remove impurities. Limestone breaks down and reacts with sand from the rocks to form slag

Coke - a fuel that produces carbon monoxide for the reduction reaction

Hot air - the fourth raw material

Required for coke to burn



Carbon (coke) and oxygen (from the hot air) produce carbon monoxide and gives off heat. Reduction is achieved by Carbon monoxide at a high temperature

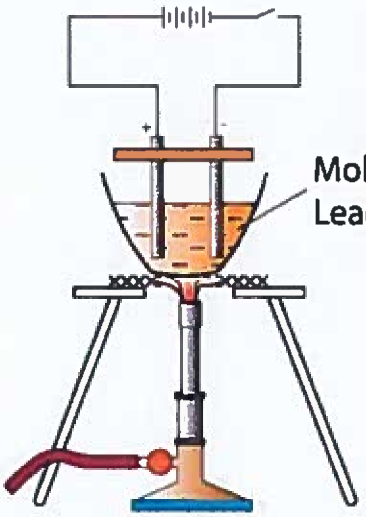
Iron oxide + carbon monoxide ⟶ iron + carbon dioxide

$$Fe_2O_3 + \boxed{3}CO \longrightarrow \boxed{2}\,Fe + \boxed{3}CO_2$$

Getting the furnace up to temperature takes a lot of time and costs a lot. As a result raw materials are constantly added and products removed - the process is continuous.

At the factory in Port Talbot iron ore, limestone and coke are imported from other countries even though they are available in Wales. Using raw materials from Wales is not sustainable due to cost and the effect it could have on the environment (quarrying).

Electrolysis of Lead (II) Bromide



At the negative electrode / cathode

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

At the positive electrode / anode

$2Br^{-} - 2e^{-} \longrightarrow Br_2$

The positive ions Pb^{2+} move towards the cathode where they gain electrons

The negative ions Br^{-} move towards the anode where they lose electrons

Oxidation and Reduction

Oxidation is:

- when a substance **gains oxygen.**
  or
- when a substance **loses electrons** (Remember: **OIL = Oxidation Is Loss).**

Reduction is:

- when a substance **loses oxygen.**
  or
- when a substance **gains electrons** (Remember: **RIG = Reduction Is Gain).**

Examples of oxidation and reduction in the Blast furnace in the Extraction of Iron:

- The carbon monoxide gas is used to **reduce** the iron oxide because the carbon monoxide **removes the oxygen** from the iron oxide. Iron is formed. **The iron oxide is reduced.**
  Remember: The substance that is reduced has oxygen removed from it.

- The **carbon monoxide** (CO) **gas is oxidised** because it **gains oxygen** from the iron oxide. Carbon dioxide gas (CO_2) is formed.

Examples of oxidation and reduction in reactions that do not involve oxygen.

It is also useful to define oxidation and reduction in terms of either losing or gaining electrons because **some reactions do not involve oxygen.**

- **Example: The Electrolysis of lead (II) bromide**

At the negative electrode (cathode) the Pb^{2+} ions are ***reduced*** because they gain electrons to from lead atoms:

Pb^{2+} + $2e^-$ → Pb
Lead ions gain 2 electrons-------------------→ lead atoms
(reduced)

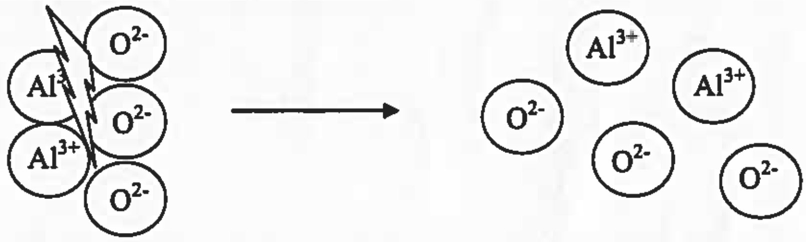
At the positive electrode (anode) the bromide ions (Br^-) are ***oxidised*** because they lose electrons to from bromine atoms(then molecules of bromine):

$2Br^-$ − $2e^-$ → Br_2
Bromide ions lose 2 electrons →bromine molecules
(oxidised)

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

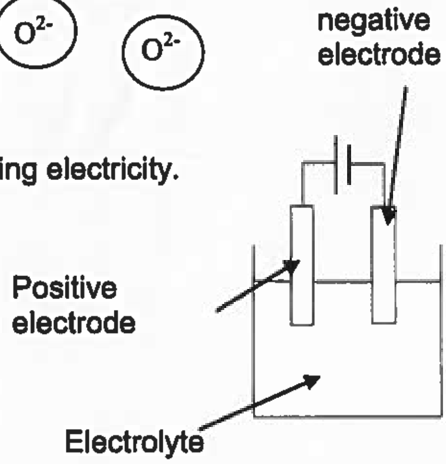
Electrolysis of Aluminium Oxide

Electrolysis is the method used to extract aluminium from aluminium oxide As aluminium is a reactive metal, aluminium oxide is very stable, a more powerful method is needed to break the bonds



Electrolysis is the decomposition of a compound using electricity.

Electrodes carry the current into and out of the molten compound, they are conducting rods. One is positive and the other is negative.



Anode = positive electrode

Cathode = negative electrode

Electrolyte is a solution containing ions.
Must be dissolved or molten to allow ions to move and carry charge

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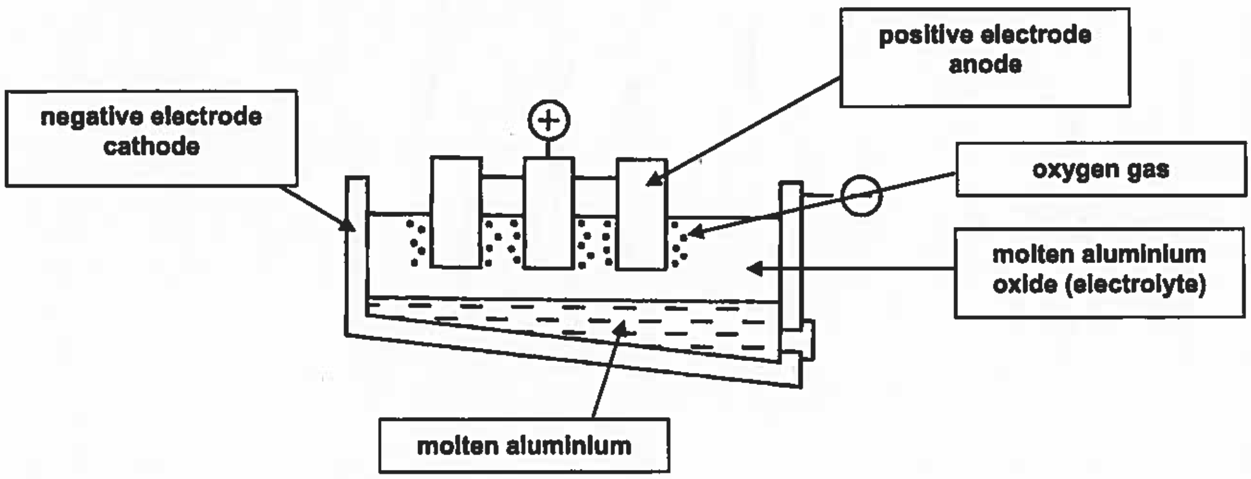
Aluminium Extraction (Separating aluminium oxide to create aluminium)

Electrolyte = molten aluminium oxide (950°C)

Electrodes = Carbon

Both **electrodes** are placed in molten aluminium oxide (electrolyte). This contains ions of aluminium (+ charge) and oxygen (- charge). These are able to move when molten and therefore allow conduction of electricity.



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Electrolysis of Aluminium Oxide

Aluminium ions are attracted to the **negative electrode** (cathode)

Oxygen ions are attracted to the **positive electrode** (anode)

Reaction at the negative electrode **cathode**	aluminium ions+ electrons → aluminium atoms $Al^{3+} + 3e \rightarrow Al$
Reaction at the positive electrode **anode**	Oxide ions – electrons → oxygen molecules $2O^{2-} - 4e \rightarrow O_2$

Aluminium has many uses due to its physical properties $2O^{2-} \rightarrow O_2 + 4e^-$

Uses	Property
Car manufacturing	Resistant to corrosion
Power lines	Electrical conductivity
Saucepans	Heat conduction
Aeroplanes	Low density

Locating aluminium plants

Electrolysis is an expensive process as it **needs a lot electrical energy** constantly. Most are located **next to a power station**

Aluminium is reactive so it needs an enormous amount of electricity to separate it from oxygen. Also it is expensive as it needs **heat** energy to heat up the ore to 1000°C

The energy costs associated with aluminium production are very high and when Wylfa Power Station was decommissioned, Anglesey Aluminium closed. When it was running the plant accounted for around 10-15% of all the electricity used in Wales. Without a power station close by, guaranteeing the supply of electricity, this became unsustainable and the plant closed.

Factories are located **near the coast** as they need **to import the aluminium** ore from abroad.

To increase the lifetime of metal ores such as aluminium oxide and iron oxide it is necessary to **recycle** metals.

Recycling aluminum uses only about 5% of the energy needed to extract it from bauxite and saves waste. Less electrical consumption means less greenhouse gas (CO_2) emissions. The environment is spoilt by quarrying.

Uses of metals

Copper

Copper has many uses due to its physical properties

Uses	Property
Jewellery	Shiny
electrical Wires	Electrical conduction
saucepans	Heat conduction
pipes	Malleability (create sheets)
Electrical wires	Ductility (create wires)

Titanium

Titanium is important as an alloying agent with aluminum, molybdenum, manganese, iron, and other metals. Alloys of titanium are principally used for aircraft and missiles where **lightweight strength** and ability to **withstand extremes of temperature** are important.

Titanium is as strong as steel, but 45% lighter. It is 60% heavier than aluminium, but twice as strong. Does not corrode in water. 1660 °C M.pt

An alloy is a mixture made by mixing molten metals; the properties can be changed by altering the amount of each metal

Steel

Steels are a large family of metals. All of them are **alloys** in which iron is mixed with carbon and other elements. Steels are described as mild, medium- or high-carbon steels according to the percentage of carbon they contain, although this is never greater than about 1.5%.

Type of steel	Percentage of carbon	Strength
Mild steel	Up to 0.25%	hard
Medium carbon steel	0.25% to 0.45%	harder
High carbon steel	0.45% to 1.50%	hardest

The metal in the scissors contains nearly twenty times as much carbon and is many times harder than the steel in a drinking can.

Steel is recycled on a large scale.

Recycling steel saves 50% of the energy used in the extraction of iron.

Recycling helps to conserve iron ore

Recycling cuts down on the emission of greenhouse gases (carbon dioxide)

Aluminium

Properties	Uses
Strong	Window frames and greenhouse frames Aeroplane and car bodies
Low density	High-voltage power cables for electric pylons Window frames and greenhouse frames Drinks cans Aeroplanes and car bodies
Good heat conductor	Cooking saucepans and cooking foil
Good electrical conductor	High-voltage power cables for electric pylons

Transition Metals

- Transition metals are elements found in the **centre of the Periodic Table** (in between Groups 2 and 3).
- Examples are iron, copper , silver, gold, platinum

Properties of Transition Metal elements

They have the 'typical' properties of metallic elements:

- **High melting and boiling points**
- **High density**
- **Good electrical conductivity** (allow an electrical current to pass through them)
- **Good thermal conductivity** (allow heat to pass through them)
- **Malleable** (easily shaped)
- Used as **catalysts** to increase the rate of some chemical reactions, for example:
    - Iron is used as a catalyst in the production of ammonia NH_3
    - Platinum is used as a catalyst in the catalytic converters of car exhaust systems. It helps to speed up the reaction between toxic carbon monoxide and unburned petrol with oxygen from the air to form carbon dioxide and water vapour.
- Transition metal elements can **form more than one type of ion with different charges on the ions** (found in transition metal compounds):

e.g. Iron (Fe) can form 2 ions: Fe^{2+} or Fe^{3+}

Copper (Cu) can form 2 ions: Cu^{+} or Cu^{2+}

Properties of Transition Metal compounds

- Transition metal compounds are often **coloured** and this makes them useful for making coloured pigments in paints, glazes and dyes.

Colours of transition metal ions in solutions and compounds (Higher Tier only)

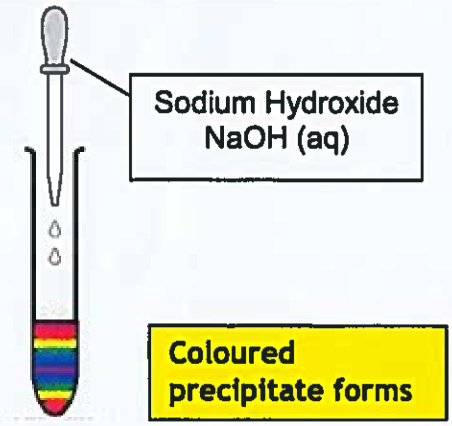
Transition metal ion	*Colour in compound or solution*	*Example of compound or solution*
Iron (II) Fe^{2+}	**pale green**	Iron (II) hydroxide $Fe(OH)_2$ Iron (II) sulphate $FeSO_4$ Iron (II) chloride $FeCl_2$
Iron (III) Fe^{3+}	**brown**	Iron (III) hydroxide $Fe(OH)_3$ Iron(III) oxide Fe_2O_3 Iron (III) chloride $FeCl_3$
Copper (II) Cu^{2+}	**blue**	Copper (II) hydroxide $Cu(OH)_2$ Copper (II) sulphate $CuSO_4$ Copper (II) chloride $CuCl_2$

Chemical Analysis – Precipitation Reaction

Higher Tier only

- Transition metal hydroxides are insoluble in water.
- If a solution of any soluble transition metal compound is mixed with sodium hydroxide then we get a displacement reaction. The sodium is the more reactive metal, and displaces the transition metal from its compound.
- The transition metal hydroxide is formed as a result. As this is insoluble it appears as a solid in the liquid - this is called a **precipitate**

Add Sodium Hydroxide



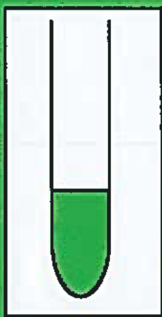
Iron (II) / Fe^{2+}

When we add sodium hydroxide to an iron (II) salt a **dirty green precipitate** is formed

iron (II) sulfate + sodium hydroxide → iron (II) hydroxide + sodium sulfate

$FeSO_4 + 2NaOH \rightarrow Fe(OH)_2 + Na_2SO_4$

$$Fe^{2+} (aq) + 2OH^- (aq) \rightarrow Fe(OH)_2 (s)$$



Iron (III) / Fe^{3+}

When we add sodium hydroxide to an iron (III) salt a **red brown precipitate** is formed

iron (III) nitrate + sodium hydroxide → iron (III) hydroxide + sodium nitrate

$Fe(NO_3)_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaNO_3$

$$Fe^{2+} (aq) + 3OH^- (aq) \rightarrow Fe(OH)_3 (s)$$

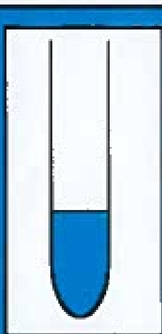
Copper (II) / Cu^{2+}

When we add sodium hydroxide to copper (II) salts a **blue precipitate** is formed

copper (II) sulfate + sodium hydroxide → copper (II) hydroxide + sodium sulfate

$CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$

$$Cu^{2+} (aq) + 2OH^- (aq) \rightarrow Cu(OH)_2 (s)$$



<u>Precipitation Reactions and Spectator Ions</u> (Higher Tier only)

Spectator ions are ions which are <u>not</u> used to form the precipitate in the reaction. They do <u>not</u> take part in the reaction to form the solid precipitate.

Spectator ions can be identified by looking at the ionic equations for the reactions. The spectator ions do not change in any way. They are the same ions in both the reactants and products. The ions stay in the same state, eg. aqueous (aq)

Example:

Balanced chemical equation

$FeSO_4\ _{(aq)} + 2NaOH\ _{(aq)} \rightarrow Fe(OH)_2\ _{(s)} + Na_2SO_4\ _{(aq)}$

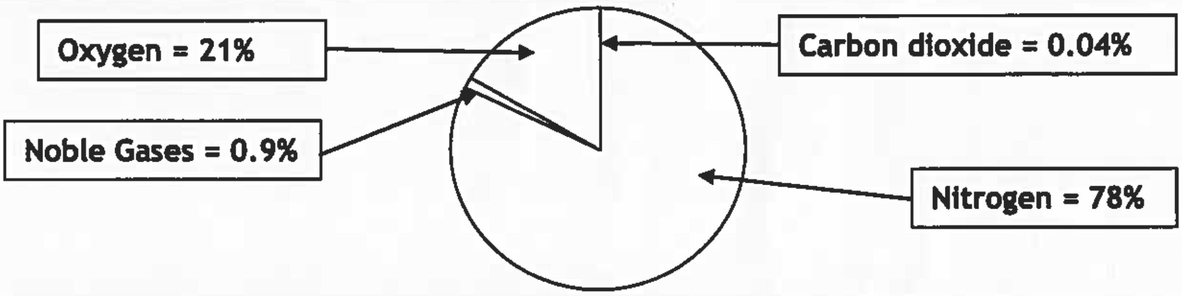
Ionic equation for this reaction

$Fe^{2+}\ _{(aq)} + 2OH^{-}\ _{(aq)} \rightarrow Fe(OH)_2\ _{(s)}$

The sulphate ions in the aqueous solution $(SO_4)^{2-}$ and the hydroxide ions $(OH)^{-}$ in the aqueous solution are not involved in forming the solid precipitate. Both of these ions are the **spectator ions** in the reaction.

Non-metals

Composition of the air

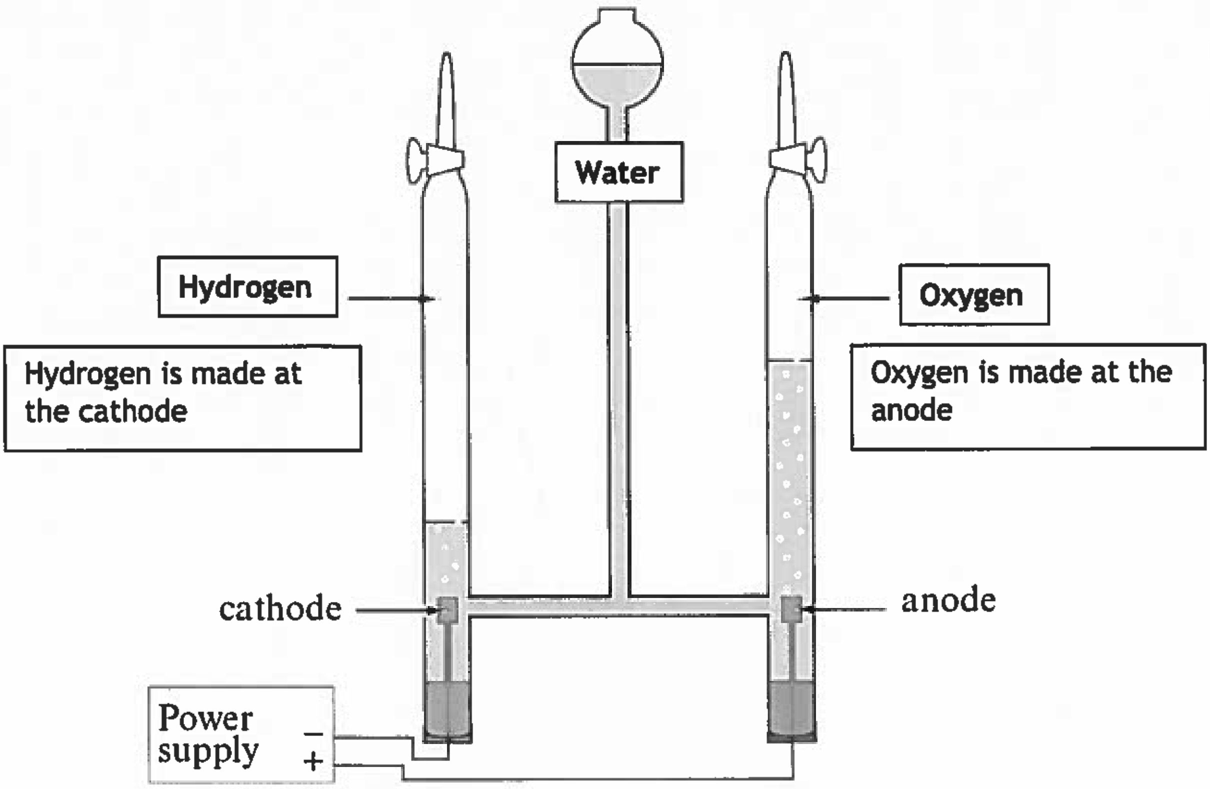


Air as a raw material

Non-metals such as nitrogen, oxygen, neon and argon are obtained from the air.

Electrolysis of water - the Hoffmann Voltameter

Oxygen and hydrogen can be made from the electrolysis of water. The equipment below is used



Twice the volume of Hydrogen as oxygen is made, this is because the formula of water is H_2O.

$$2H_2O_{(l)} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)}$$

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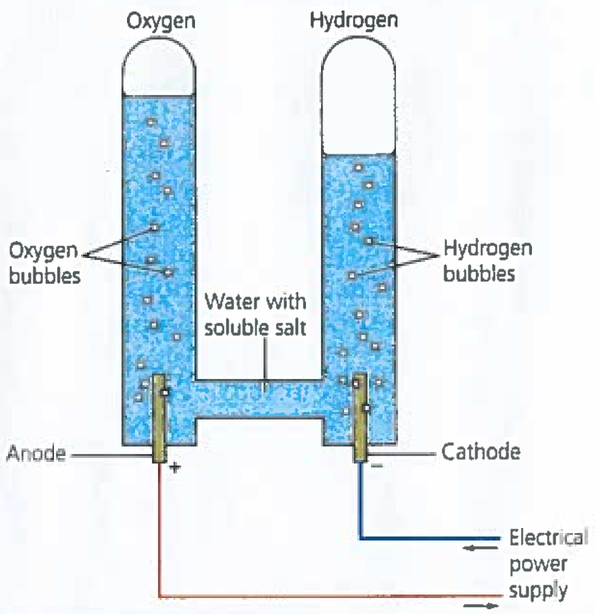
Electrolysis of Water: Reactions at the Electrodes

- **Hydrogen ions (H^+) and hydroxide ions (OH^-) are present in the solution because the water molecules have been split up (electrolysed).**

$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$



Cathode (-) (negative electrode): hydrogen gas is formed: reduction reaction

Positively charged hydrogen ions (H^+) are attracted to the negatively charged cathode electrode.

The hydrogen ion $(H^+)_{(aq)}$ gains one electron to form a hydrogen atom. Then two hydrogen atoms join together to make a ***hydrogen gas*** molecule.

Overall reaction: $2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}$

The hydrogen ions are **reduced (reduction)** because they **gain** one electron each.

Anode (+) (positive electrode): oxygen gas is formed (Higher Tier only): oxidation reaction

Negatively charged hydroxide ions (OH^-) are attracted to the positively charged anode electrode.

Oxygen gas is formed.

$$2OH^-_{(aq)} \rightarrow O_{2\,(g)} + 2H^+_{(aq)} + 4e^-$$

<u>Electrolysis of aqueous solutions</u> ***(Higher & Foundation Tier)***

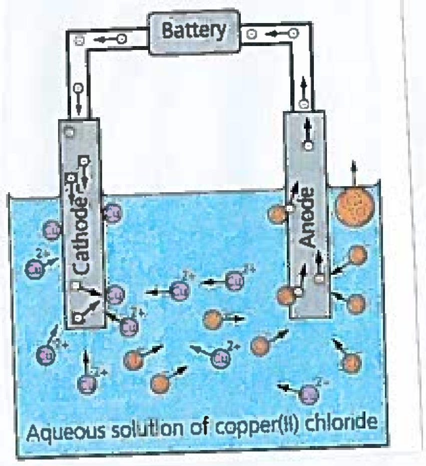
When metal salts dissolve in water, positive and negative ions are released. An **aqueous solution** is formed. In aqueous solutions there are **ions present** from the metal salt and the water.

Example:

The ions present in copper (II) chloride solution are:

- Copper (II) ions $Cu^{2+}_{(aq)}$ from the copper (II) chloride
- Chloride ions $Cl^-_{(aq)}$ from the copper (II) chloride
- Hydrogen ions $H^+_{(aq)}$ from the water
- Hydroxide ions $OH^-(aq)$ from the water

Electrolysis of copper (II) chloride solution



Negative cathode reaction: copper metal is formed

- The positively charged copper ions Cu^{2+} are attracted to the negative cathode.
- The Cu^{2+} ions gain 2 electrons from the cathode and form copper atoms (seen as a layer of salmon pink copper metal) on the surface of the cathode.
- The Cu^{2+} ions are **reduced** because they **gain** electrons (Remember: RIG Reduction Is Gaining electrons).
- Ionic equation for the reaction at the cathode:

$$Cu^{2+}{}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)}$$

Positive anode reaction: chlorine gas is formed

- The negatively charged chloride ions Cl^- are attracted to the positively charged anode.
- Two chloride ions ($2Cl^-$) lose one electron each to form 2 chlorine atoms.
- These two chlorine atoms then combine to form a molecule of chlorine gas (Cl_2).

- The chloride ions Cl^- are **oxidised** because they **lose** electrons (Remember: OIL Oxidation Is Loss).
- Ionic equation for the reaction at the anode:

$$2Cl^-_{(aq)} \rightarrow Cl_{2\,(g)} + 2e^-$$

Why is copper metal formed at the cathode and not hydrogen gas?

The hydrogen ions (H^+) and copper ions (Cu^{2+}) are both positively charged and are attracted to the negative cathode.

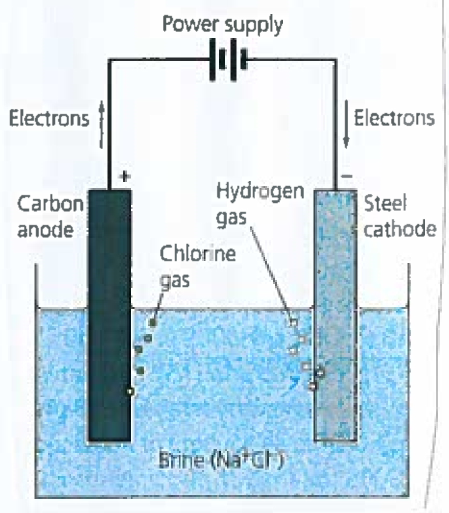
When aqueous solutions are electrolysed, if the METAL in the metal salt solution is LOWER in the Reactivity Series than HYDROGEN, then the METAL IS FORMED AT THE CATHODE.

This is why the copper metal is formed at the cathode instead of hydrogen gas because the copper ions gain electrons more easily than the hydrogen ions so copper is formed at the cathode.

Reactivity Series

Metal	Reactivity	Electrolysis of aqueous solutions containing the metal salts dissolved in water
Potassium	Most reactive metal	***HYDROGEN GAS is formed at the CATHODE when the aqueous solution contains any of these more reactive metal ions.***
Sodium	***Reactivity DECREASES as you descend (go down) this list of metals***	
Calcium		
Magnesium		
Aluminium		
Carbon (non-metal)		
Zinc		
Iron		
Lead		
Hydrogen (non-metal)		
Copper	Least reactive metal	***Copper or silver or gold METAL are formed at the CATHODE.***
Silver		
Gold		

Electrolysis of sodium chloride solution (and brine)
HIGHER TIER ONLY



The ions present in sodium chloride solution are:

- Sodium ions $Na^+_{(aq)}$ from the sodium chloride
- Chloride ions $Cl^-_{(aq)}$ from the sodium chloride
- Hydrogen ions $H^+_{(aq)}$ from the water
- Hydroxide ions $OH^-(aq)$ from the water

Both the positively charged sodium ions and the hydrogen ions are attracted to the negatively charged cathode. There is a competition between these 2 ions.

The hydrogen ions gain electrons more easily than sodium ions so hydrogen gas is produced at the cathode.

Higher Tier only:
When the metal salt dissolved in water (aqueous solution) contains METAL IONS HIGHER in the Reactivity Series than HYDROGEN, then HYDROGEN GAS is formed at the CATHODE.

Cathode (-) (negative electrode): hydrogen gas is formed: reduction reaction

Positively charged hydrogen ions (H^+) are attracted to the negatively charged cathode electrode.

The hydrogen ion (H^+) $_{(aq)}$ gains one electron to form a hydrogen atom. Then two hydrogen atoms join together to make a ***hydrogen gas*** molecule.

Overall reaction: $2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}$

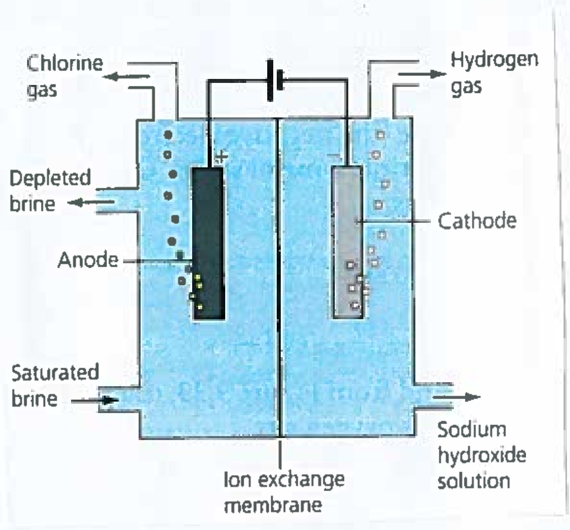
The hydrogen ions are **reduced (reduction)** because they **gain** one electron each.

Positive anode reaction: chlorine gas is formed

- The negatively charged chloride ions Cl^- are attracted to the positively charged anode.
- Two chloride ions ($2Cl^-$) lose one electron each to form 2 chlorine atoms.
- These two chlorine atoms then combine to form a molecule of chlorine gas (Cl_2).
- The chloride ions Cl^- are **oxidised** because they **lose** electrons (Remember: OIL Oxidation Is Loss).
- Ionic equation for the reaction at the anode:

$$2Cl^-_{(aq)} \rightarrow Cl_{2\,(g)} + 2e^-$$

<u>The Chloralkali Process is used to manufacture chlorine, hydrogen and and sodium hydroxide in industry (HIGHER TIER ONLY)</u>



- **Chlorine gas** is manufactured by the **electrolysis of brine (concentrated sodium chloride dissolved in water)** when an electric current is passed through the brine) in the Chloralkali process.

- This Chloralkali process involves the use of a non-semi permeable ion exchange membrane at the centre of the electrochemical cell. This allows the sodium ions (Na^+ ions) to pass into the second chamber where **sodium hydroxide solution** (a very useful alkali) forms.

- There are 3 useful products formed in this industrial manufacturing electrolysis reaction:

  Hydrogen gas (at the cathode)
  Chlorine gas (at the anode)
  Sodium hydroxide solution

Electroplating

- Electroplating uses **electrolysis to put a layer of one metal on top of (coat) another metal.** For example, gold plated jewellery uses a thin layer of more expensive gold put on the surface of a cheaper metal, e.g. steel.

- Electroplating can also increase metals' resistance to corrosion (nickel plating: layer of nickel on the surface of more reactive metals to protect them) and increase electrical conductivity, e.g. gold plated ear phone connections.

Using Electrolysis in Electroplating to purify Copper

- Copper metal is purified by electrolysis because very pure copper is needed to make electrical conductors (in electric circuits and wires).

- Both electrodes in the electrochemical cell are made from copper.

- The electrolyte used is copper (II) sulphate solution $CuSO_4$.

  Reaction at the negative cathode: layer of copper metal is deposited on the surface of the copper electrode: reduction reaction

  The cathode is a thin piece of very pure copper.

  The positively charged copper ions (Cu^{2+}) are attracted to the negatively charged cathode (electrode) .

  The copper ions (Cu^{2+}) from the copper (II) sulphate solution and impure copper anode **gain electrons (reduction)** to form a layer of copper metal. More pure copper adds to the surface of the cathode from the impure anode.

  The mass of the cathode increases.

  Cathode ionic equation (Higher Tier only):
  $\mathbf{Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu\ (s)}$

Reaction at the positive anode: impure copper electrode dissolves: oxidation reaction

The anode (positive electrode) is a large block of impure copper called an **active electrode**. This impure copper metal (Cu) electrode dissolves to form more copper ions (Cu^{2+}) which go into the copper (II) sulphate solution.

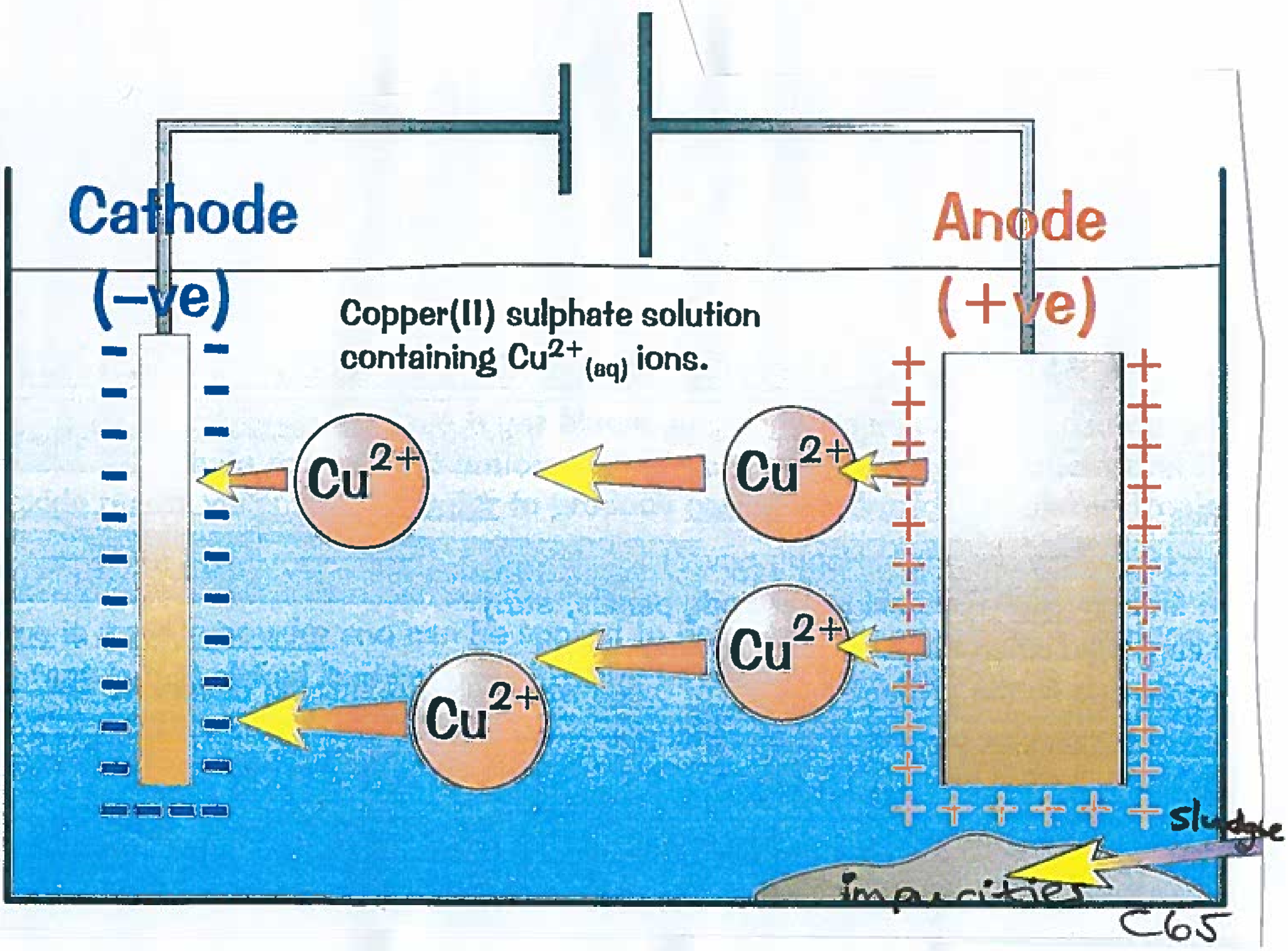
The mass of the impure copper anode decreases as it dissolves away.

The impurities in the copper anode stay as a sludge underneath the dissolving anode.

Anode ionic equation (Higher Tier only):

$$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$$

The copper atoms in the impure electrode lose 2 electrons (oxidation) to form copper ions Cu^{2+} which are attracted to the negatively charged cathode.



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Topic 4:

Chemical Reactions and Energy

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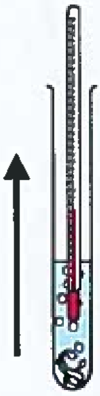
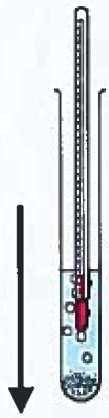
Energy in reactions

Changes in temperature happen often during chemical reactions.



Exothermic reactions

A reaction where temperature rises.
e.g. magnesium and acid



Endothermic Reaction

A reaction where temperature falls.
e.g. ammonium nitrate and acid

Here is a reaction for making ammonia

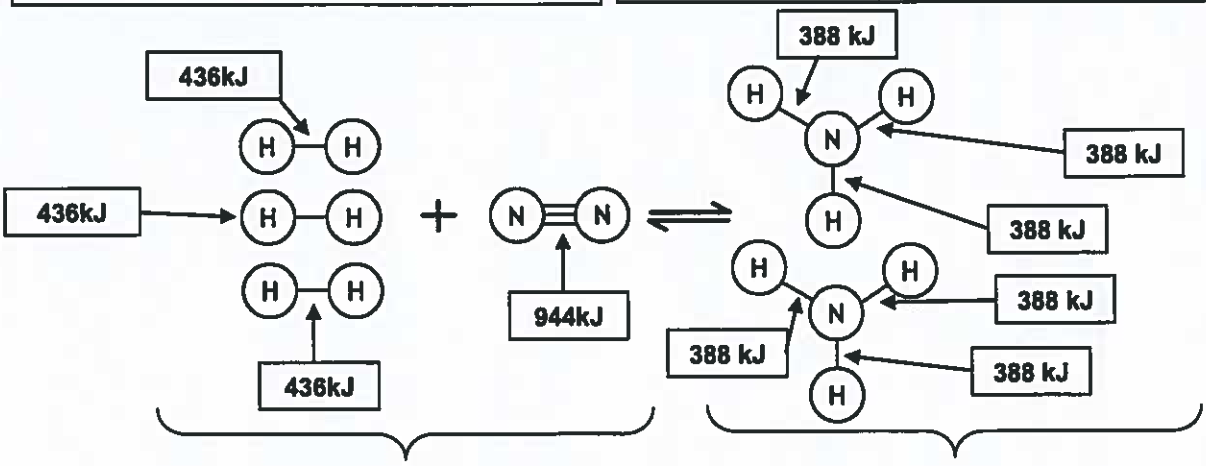
Hydrogen + Nitrogen ⇌ Ammonia

$$3 H_2(g) + N_2(g) \rightleftharpoons 2 NH_3(g)$$

To create the product ammonia its is necessary to break bonds between hydrogen and nitrogen (the reactants).

To **break bonds** energy in **required**.

Creating bonds **releases** energy.



Bonds Broken
436 + 436 + 436 + 944 = 2252kJ

Bonds Made
388 x 6 = 2328kJ

Overall Energy Change = Bonds broken – Bonds made

In an **endothermic** reaction the number would be **positive**.

2252 - 2328 = -76kJ ←

Exothermic Reaction
- Because energy is released (negative number)

Exothermic and Endothermic Reactions

Activation Energy is the minimum (lowest) amount of energy needed to start a chemical reaction. If there is not enough energy to reach the activation energy, the reaction will not happen.

Making and breaking bonds: bond energies

In chemical reactions, the bonds between the reactants are broken and new bonds in the products are formed.

Energy is required to break bonds.

Energy is released when new bonds are formed.

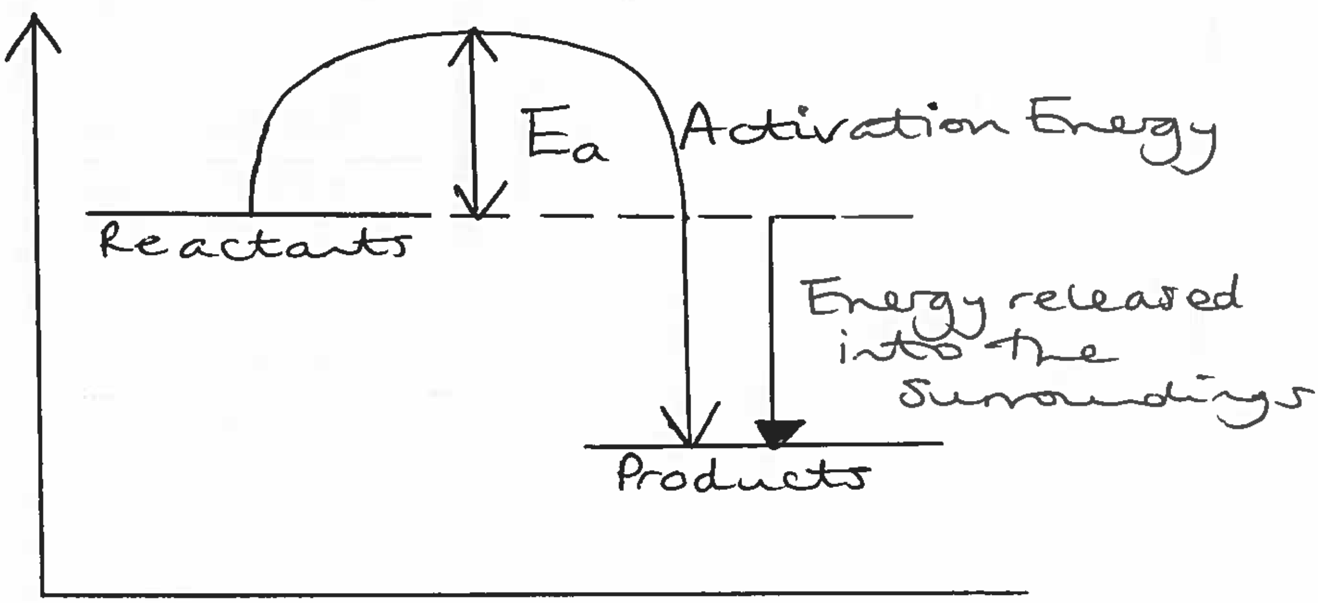
Exothermic reactions

Less energy is needed to break the bonds in the reactants than is released in making bonds in the products. Heat energy is released into the surroundings in an exothermic reaction. The temperature increases.

Examples of exothermic reactions:

- Combustion (e.g. burning fuels)
- Neutralisation (e.g. reaction between an acid and alkali)

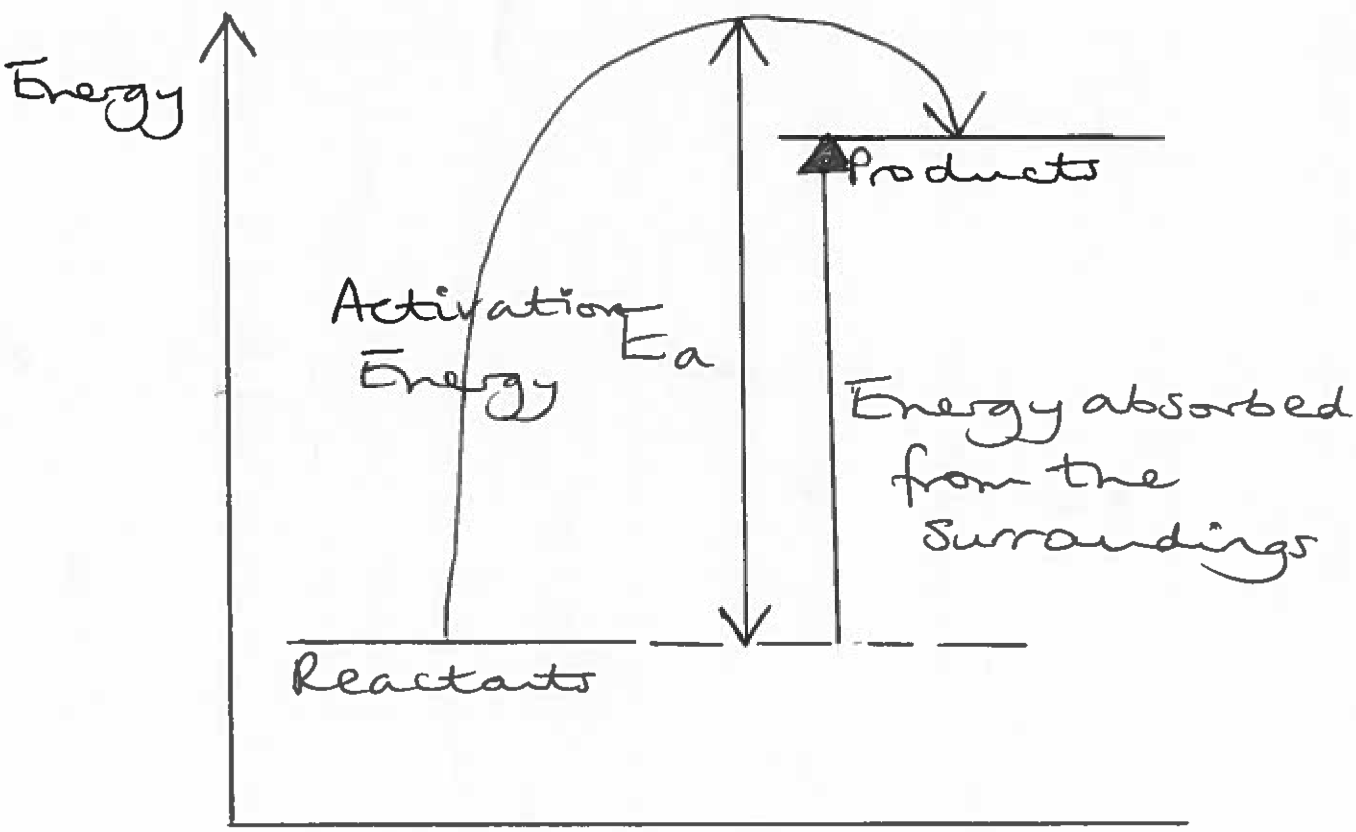
Energy Profile diagram for an Exothermic reaction



Endothermic reactions

More energy is needed to break the bonds in the reactants than is released in making bonds in the products. Heat energy is taken in from the surroundings. The temperature decreases.

Energy Profile diagram for an Endothermic reaction



Key point summary.