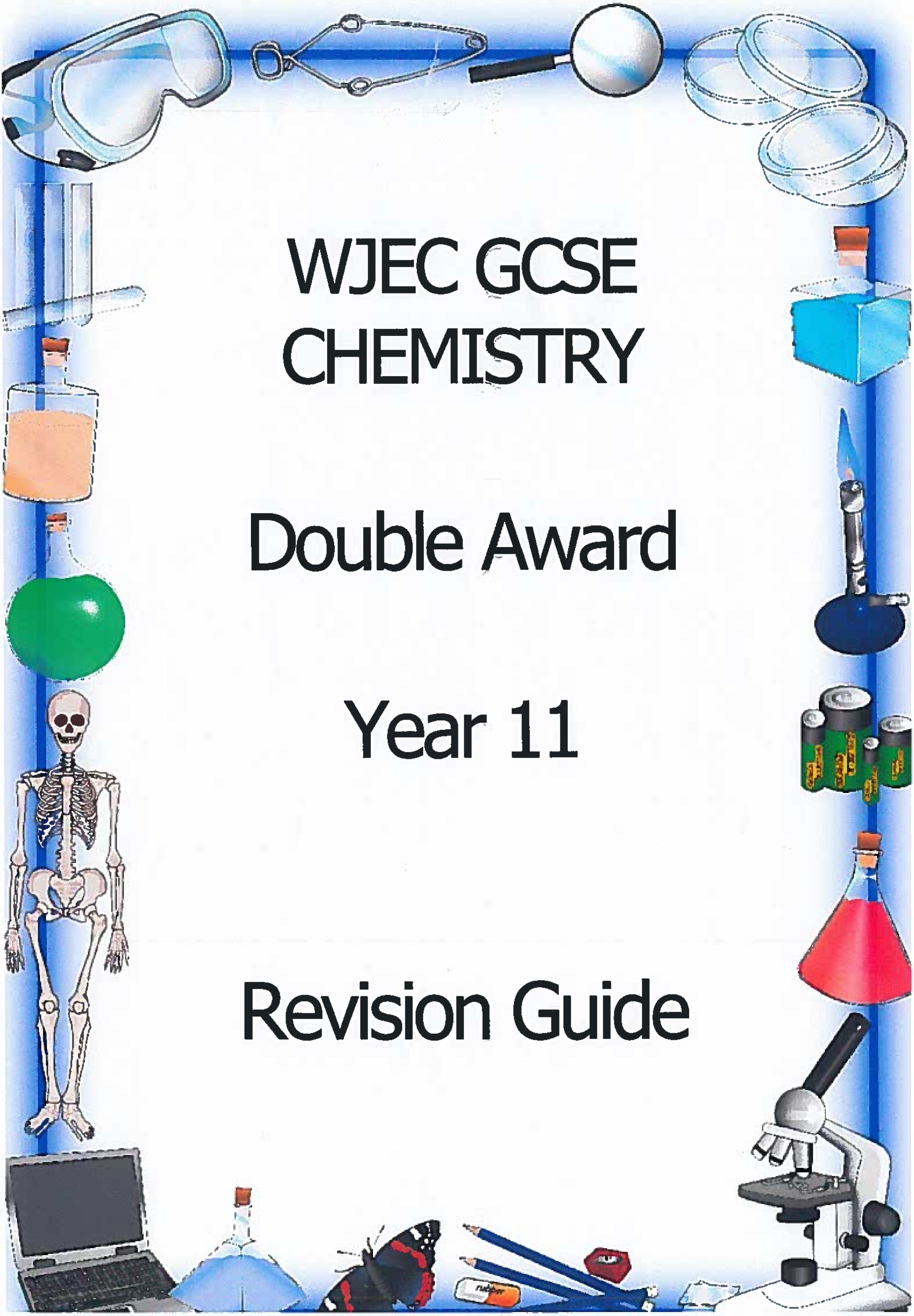


WJEC GCSE CHEMISTRY

Double Award

Year 11

Revision Guide

GCSE SCIENCE

Double Award

Year 11

REVISION GUIDE

CHEMISTRY TOPICS

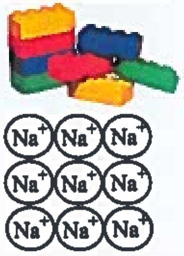
Topic 1:

Bonding, Structure and Properties

C2

Elements

Element

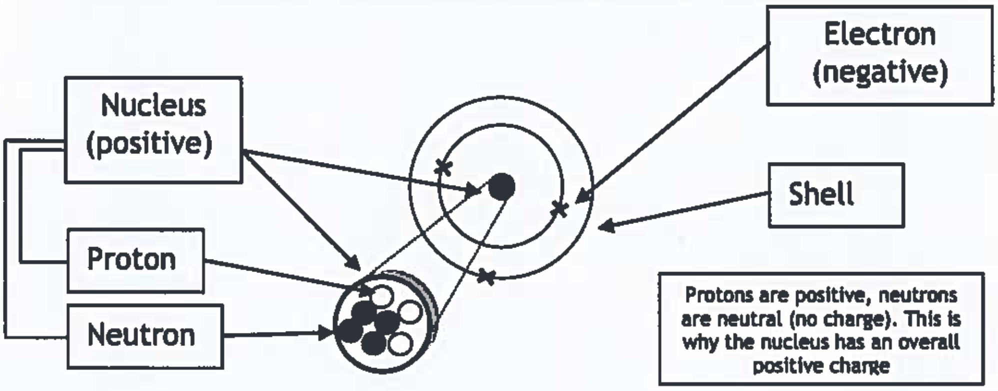


Elements are the building blocks of all substances. They cannot be broken down into simpler substances by chemical means

An Element contains only one type of atom

Atom

Each atom has negatively charged electrons orbiting a positively charged nucleus



Electron (negative)

Nucleus (positive)

Shell

Proton

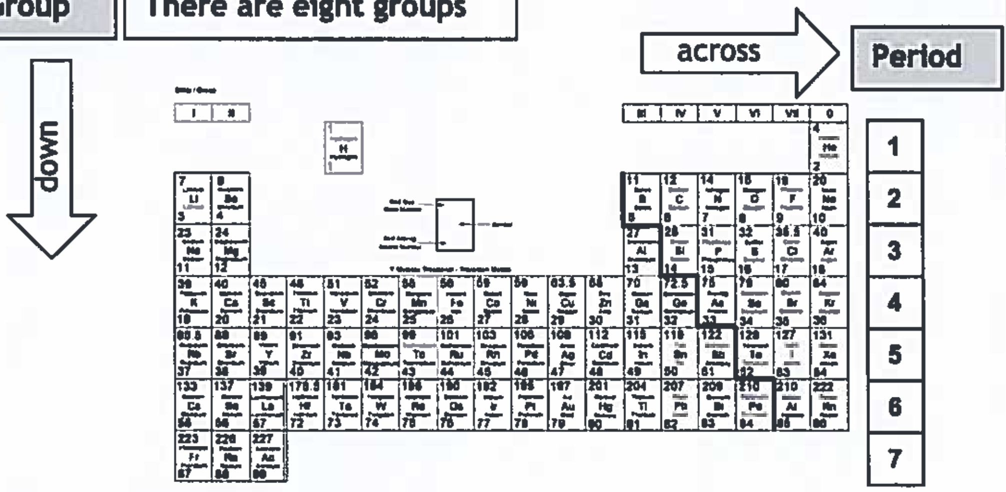
Neutron

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



Period

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

Materials and Bonding

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.

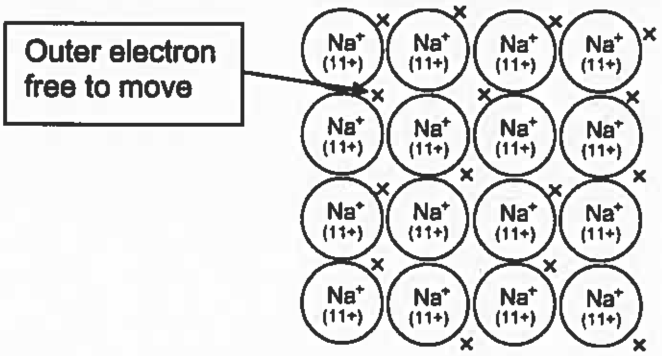
The strength of bond in a metal is the force of attraction between the metal ions and free moving electrons.

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

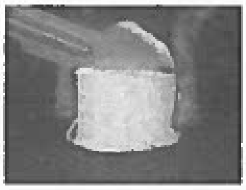
C4

Metallic bonding

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



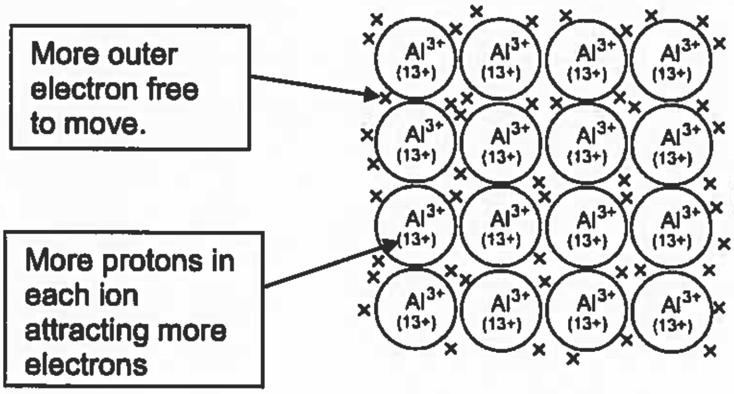
Sodium metal



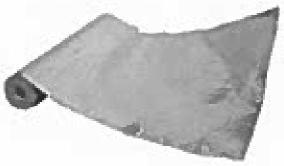
These electrons allow metals to conduct electricity and conduct heat.

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.

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



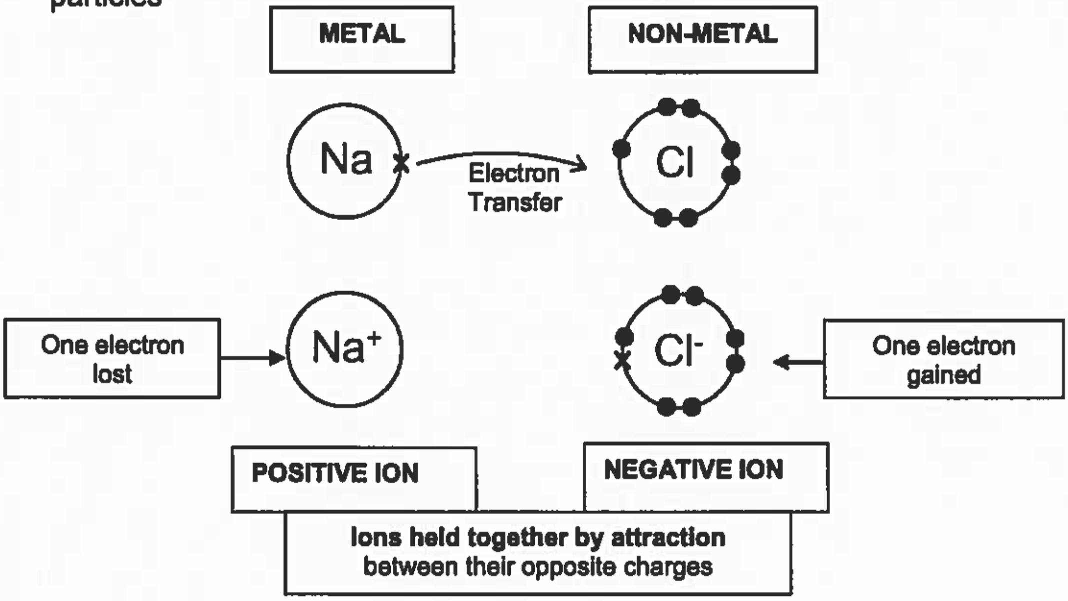
C5

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 gasses) for the two particles



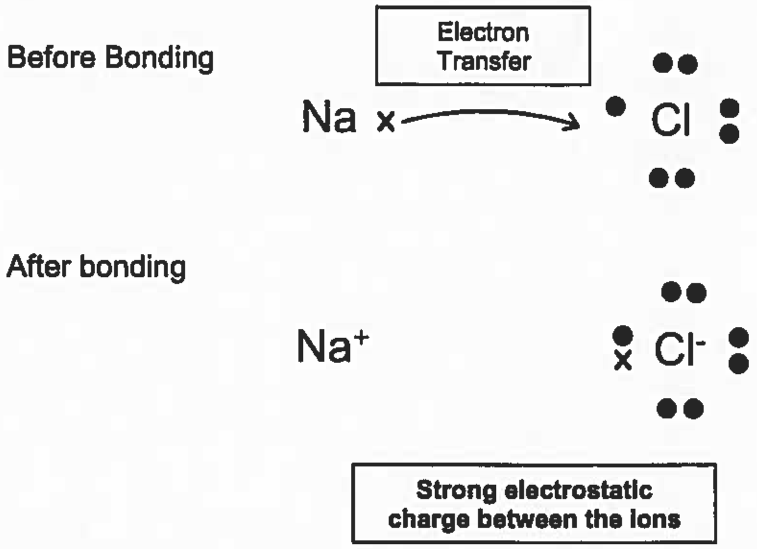
Ionic Bonding

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.

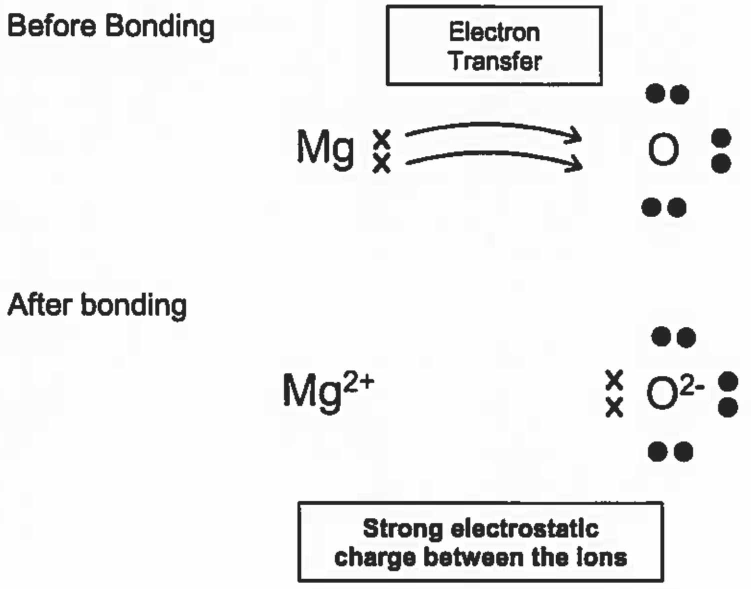
Sodium Chloride (NaCl)

Group 1 with Group 7

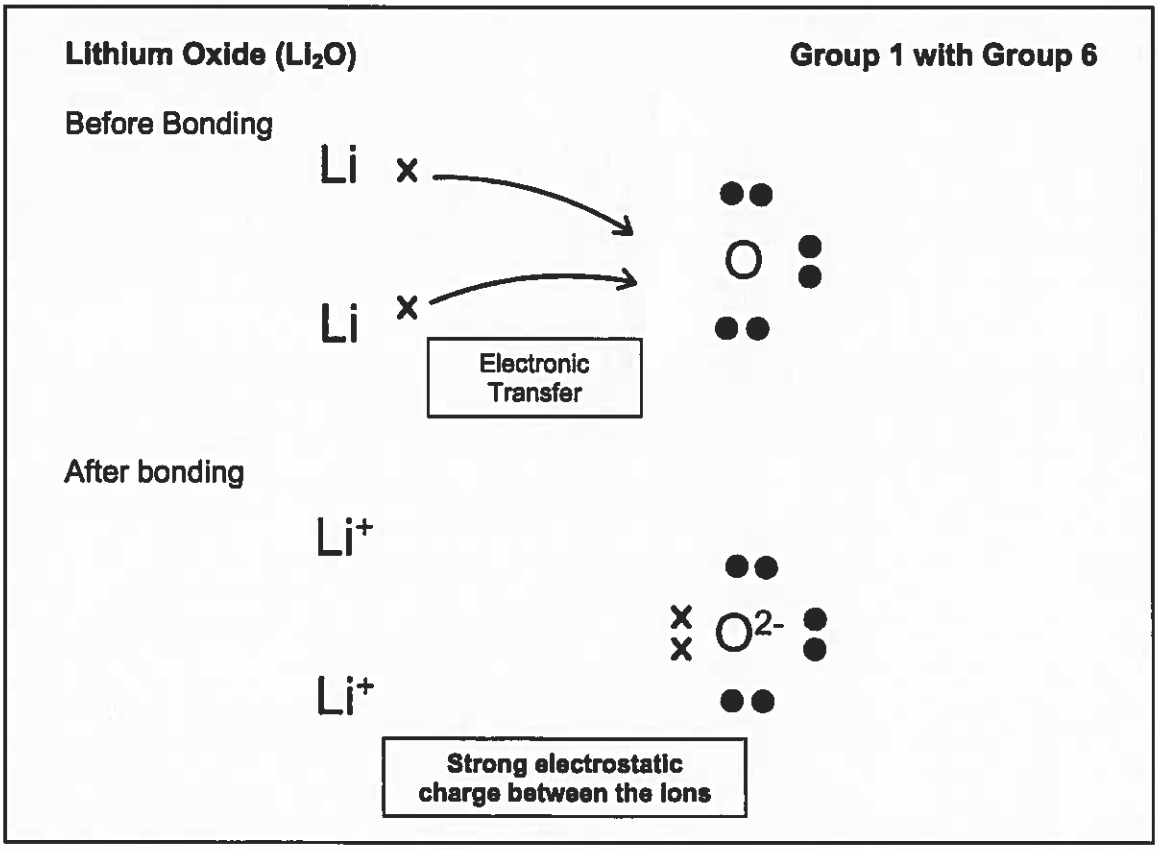


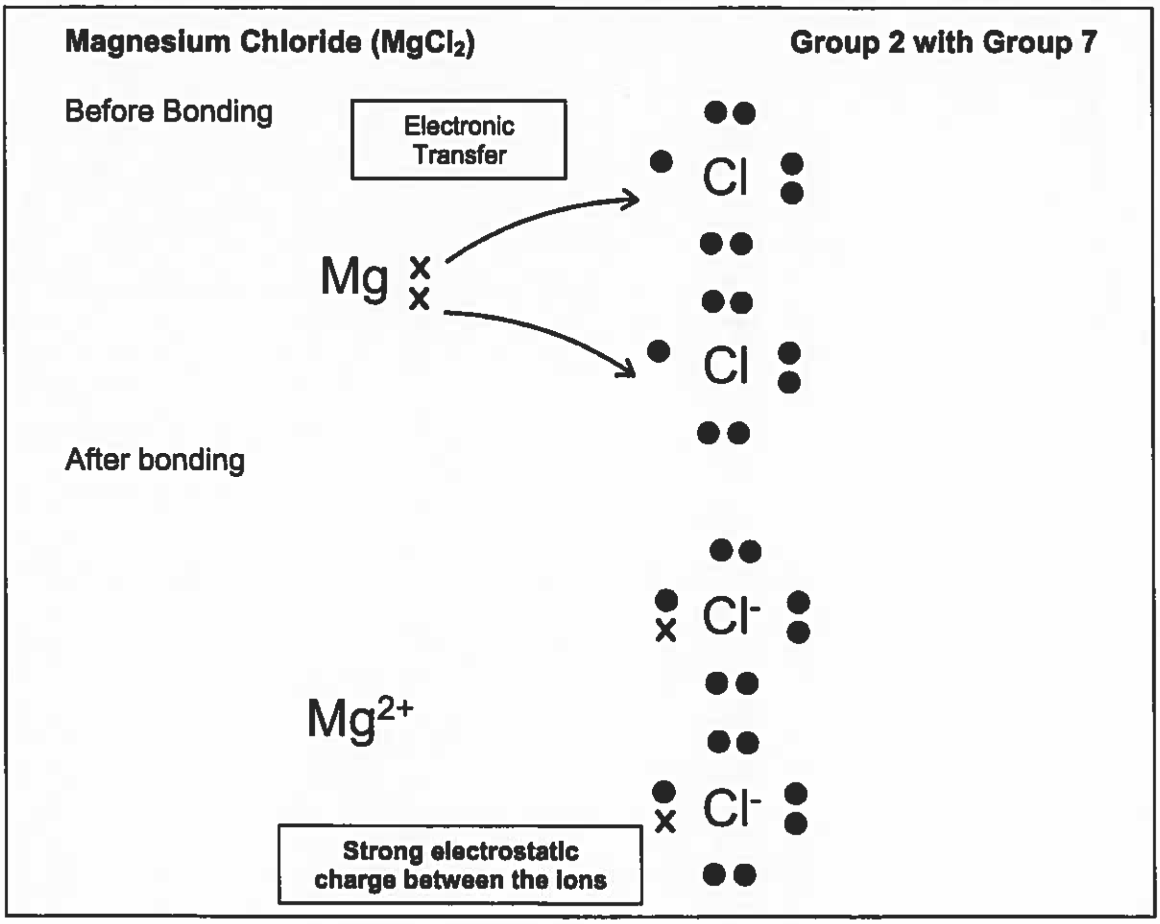
Magnesium oxide (MgO)

Group 2 with Group 6



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


Lithium Oxide (Li₂O)
Group 1 with Group 6
Before Bonding
Li
Li
O
Electronic Transfer
After bonding
Li⁺
Li⁺
O²⁻
Strong electrostatic charge between the ions


Magnesium Chloride (MgCl₂)
Group 2 with Group 7
Before Bonding
Electronic Transfer
Mg
Cl
Cl
After bonding
Cl⁻
Mg²⁺
Cl⁻
Strong electrostatic charge between the ions

Using ions to create formulae

Lithium = Li^+

Sodium = Na^+

Potassium = K^+

Magnesium = Mg^{2+}

Calcium = Ca^{2+}

chloride = Cl^-

oxide = O^{2-}

bromide = Br^-

sulfide = S^{2-}

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

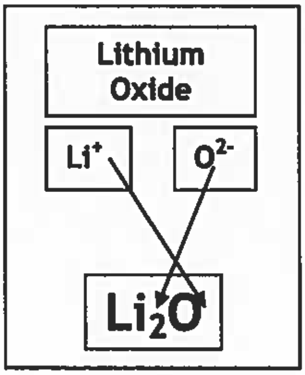
Calcium Carbonate

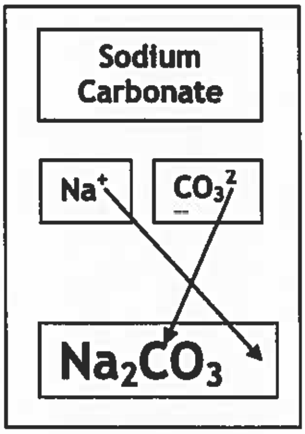
Ca^{2+} CO_3^{2-}

Ions cancel

$CaCO_3$

Quick method





C9

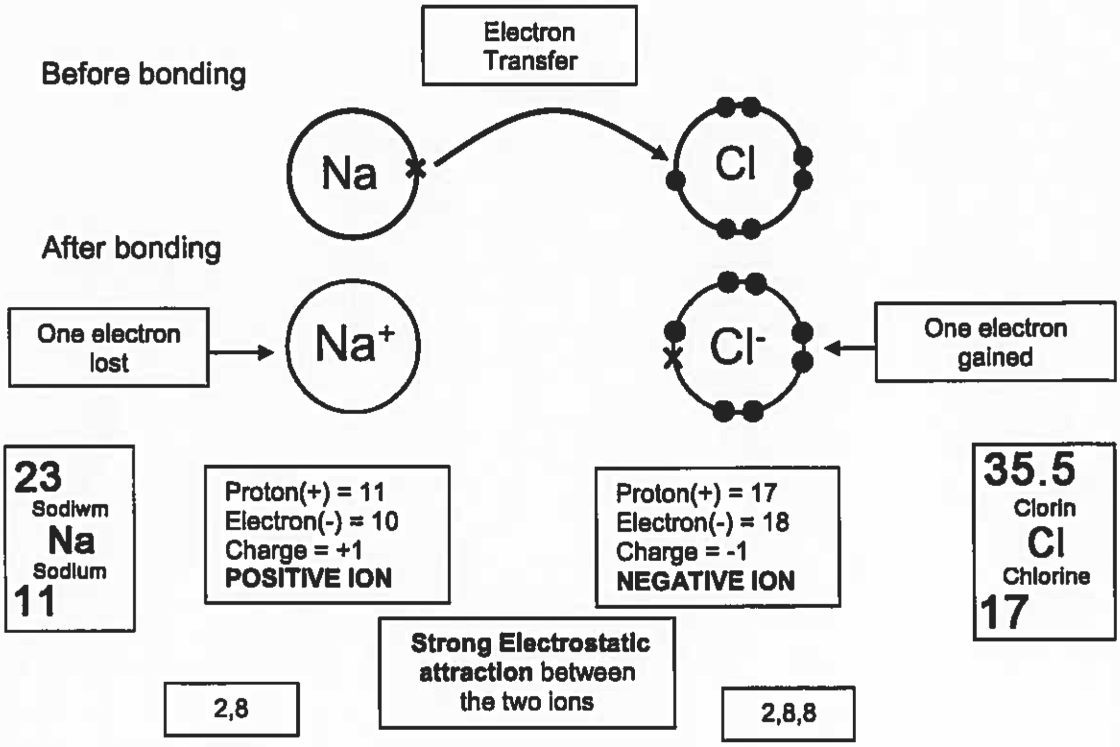
Bonding and Structure

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

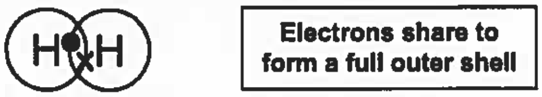
Charged particles called **ions** are formed when electrons are transferred between atoms during chemical bonding.

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').



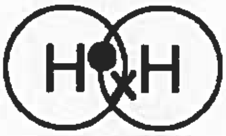
Covalent Bond

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



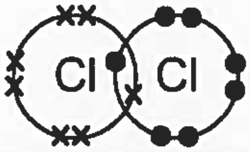
Covalent Bonds

Hydrogen (H_2)



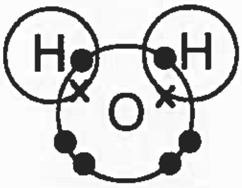
Electrons share to form a full outer shell

Chlorine (Cl_2)



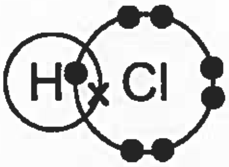
Electrons share to form a full outer shell

Water (H_2O)



Electrons share to form a full outer shell

Hydrogen Chloride (HCl)

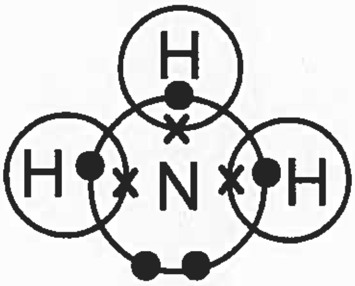


Electrons share to form a full outer shell

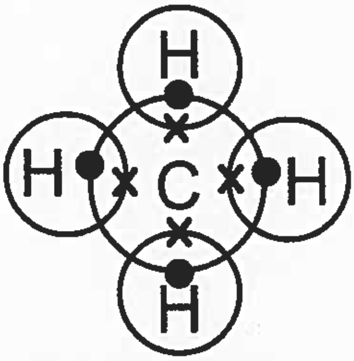
C11

Covalent Bonding

Ammonia (NH_3)

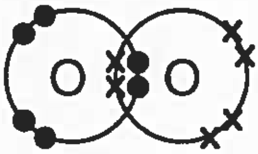


Methane (CH_4)

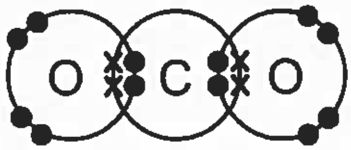


Covalent examples with double bonds (Higher Tier)

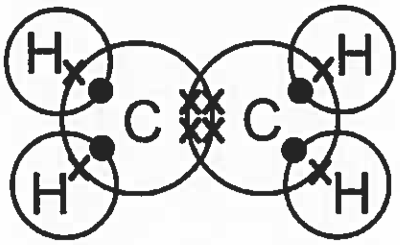
Oxygen (O_2)



Carbon Dioxide (CO_2)

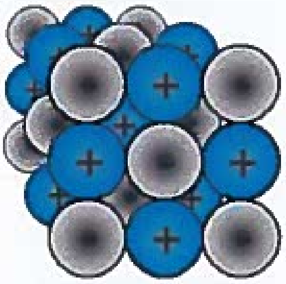


Ethene (C_2H_4)



Simple and Giant structures

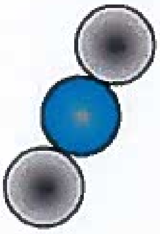
Giant ionic structure(e.g. sodium chloride, magnesium oxide),



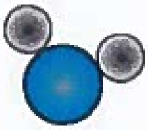
High Melting and Boiling points
Solubility = Dissolved in water

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

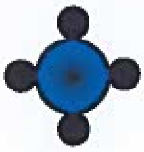
Low Melting and Boiling points
Solubility = Dissolved in water



CO_2



H_2O (water)

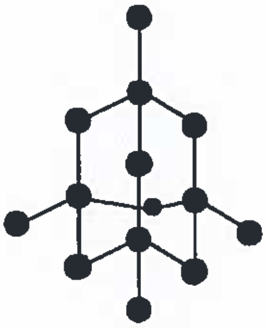


CH_4 (methane)

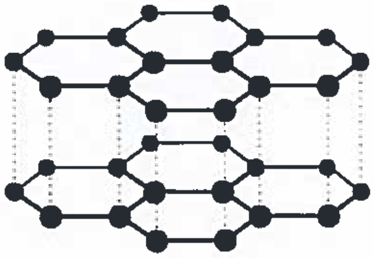
As the forces between molecules are weak the melting and boiling points are low

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

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

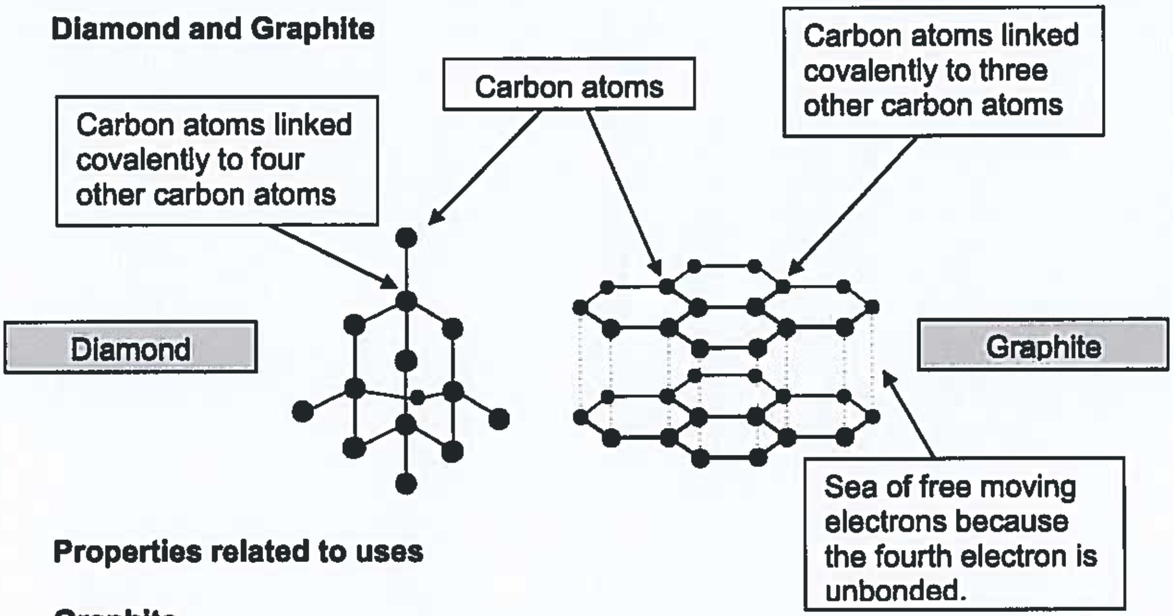


Diamond



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

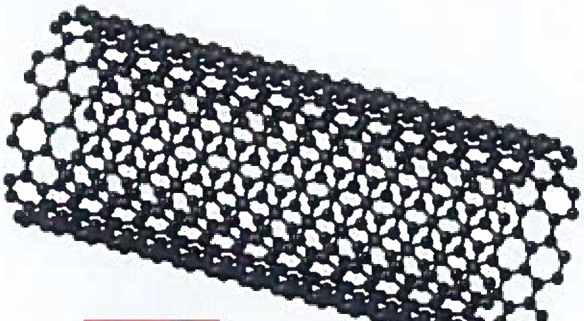
Uses	Pensil lubricants

Diamond

		Uses
Appearance	Transparent/crystalline	Gemstones
Hardness	very hard	Glass cutting, Drill bits
Conductivity	Electrical insulator	
Melting point	Very high over 3500°C	

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**

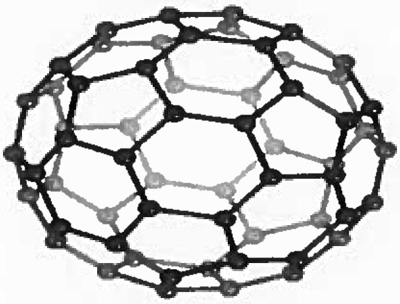


More Giant Covalent Structures

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

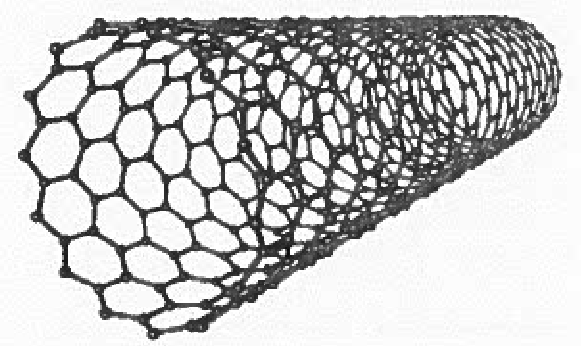
- Carbon nano-tubes are made from single layers from a graphite structure called graphene.
- The single graphite layers are rolled into hollow tubes but the covalent bonds between the carbon atoms are very strong.
- Carbon nano- tubes are a lot smaller structures than carbon fibres, so they are not used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

C15a

- 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

<u>Graphene</u>

- 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.

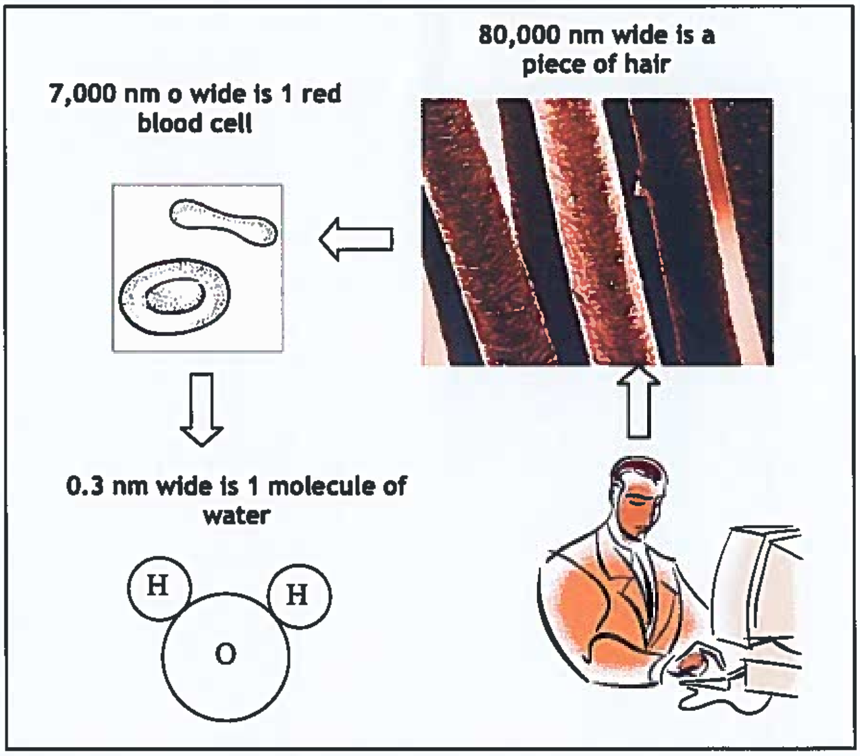
- Some scientists and engineers believe that graphene will change and transform future technology, e.g. aerospace, automotive, electronics, energy storage, coatings & paints and communications. There are not any graphene products available yet to buy or sell (not yet commercially available.)

C15b

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



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

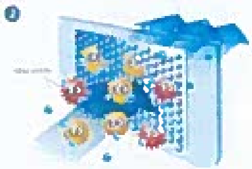
Uses which are made from nanotechnology

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

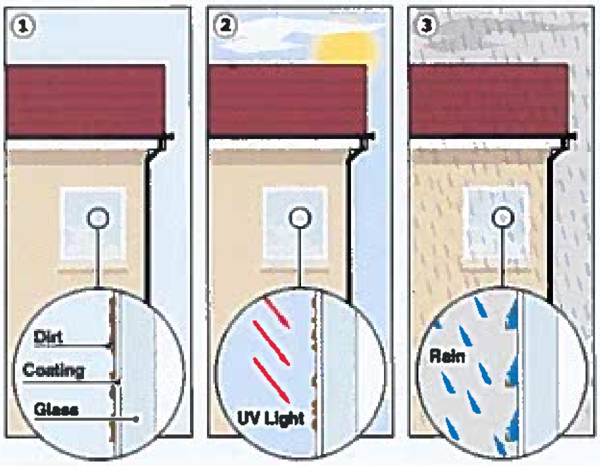
The new properties of these materials will allow people to create many new products.

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

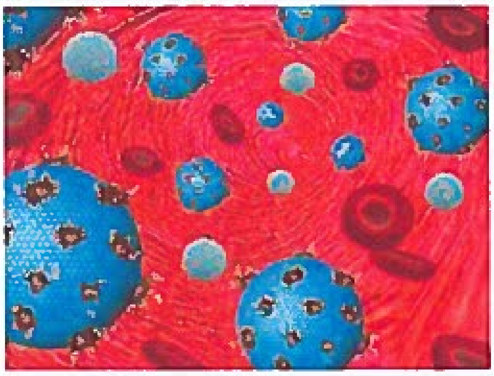
Higher Tier

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.

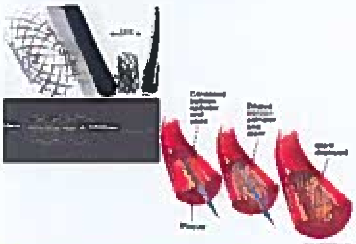
The sunglass lenses become darker when exposed to strong light and become lighter in weak light

C18

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**.

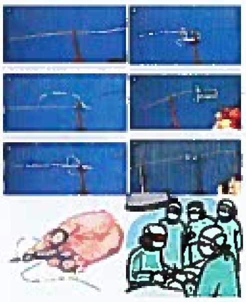
This metal can be bent into any shape at low temperature, but when heated it can remember its original shape so it bends back very quickly. It can be used as a **coffeepot thermostat**.

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:-



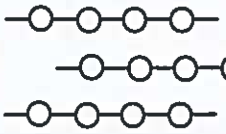
Surgical stutures are threads of smart polymer that can tighten to the right tension automatically when heated.

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.

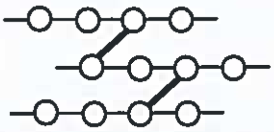
Smart Materials

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

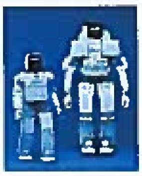
These gels can swell to **1000 times** their volume depending on the **temperature or pH.**

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

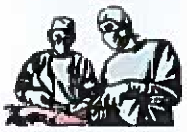
Nappies - this smart material is similar to artificial snow and expands greatly when it becomes wet.

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.

C20

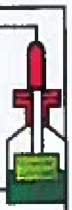
Topic 2:

Acids, Bases and Salts

Acid Reactions	**Sulfuric Acid**	**Nitric Acid**	**Hydrochloric Acid**
	H_2SO_4	**HNO_3**	**HCl**
Form salts	**Sulfate**	**Nitrate**	**Chloride**

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

$\underbrace{0 \; 1 \; 2}_{}$ Strong Acid — **Sulfuric Acid**

$\underbrace{3 \; 4 \; 5 \; 6}_{}$ Weak Acid — **Orange**

$\underbrace{7}_{}$ Neutral — **Pure water**

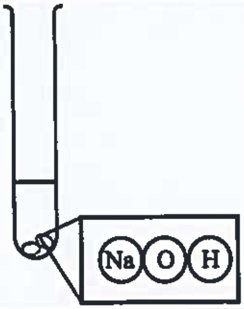
$\underbrace{8 \; 9 \; 10 \; 11}_{}$ Weak Alkali — **Soap**

$\underbrace{12 \; 13 \; 14}_{}$ Strong Alkali — **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

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

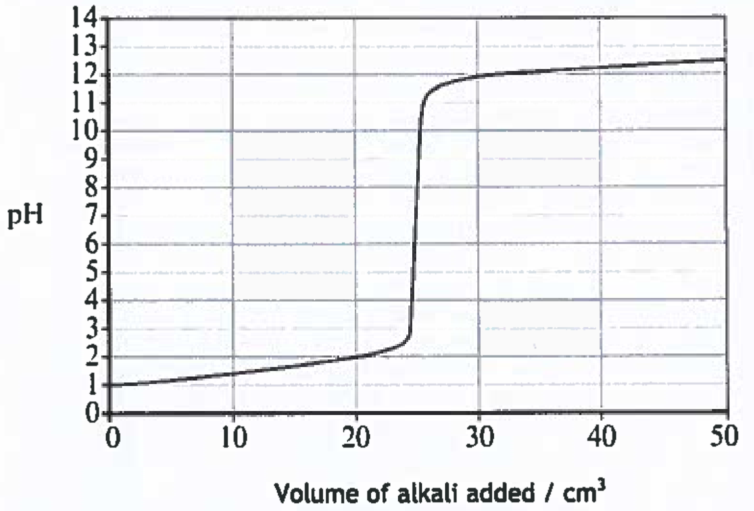
Nitric Acid + Sodium Hydroxide ⟶ Sodium nitrate + Water

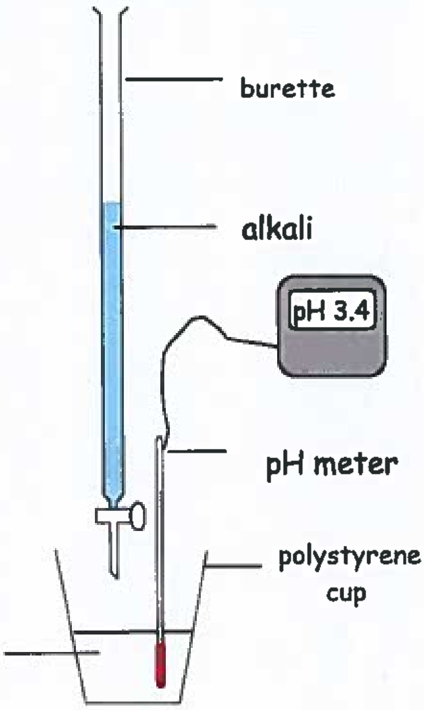
$HNO_3\ (aq) + NaOH\ (aq) \longrightarrow NaNO_3\ (aq) + H_2O\ (l)$

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 cm^3 acid

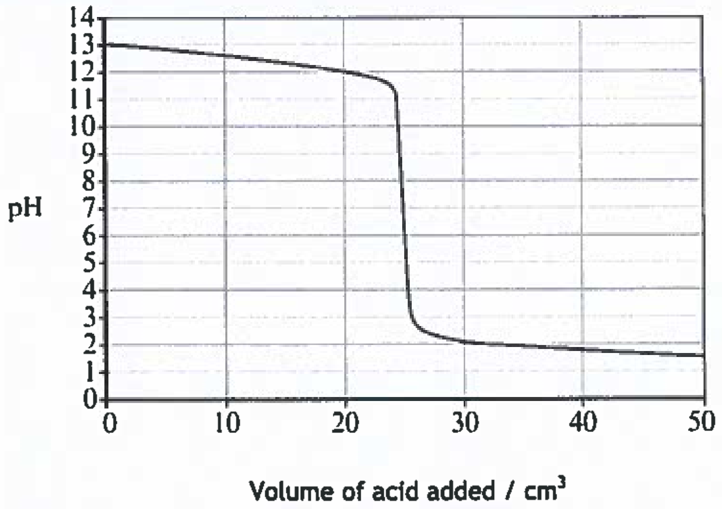
Alkali added to Acid





0 - 24 cm^3 - solution is acidic
25.00 cm^3 - neutralisation point
26- 50 cm^3 - solution is alkaline (too much alkali added)

Acid added to Alkali



0 - 24 cm^3 - alkaline
25.00 cm^3 - neutralisation point
26- 50 cm^3 - acidic

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

If the correct volume is added (25 cm^3) the solution becomes

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

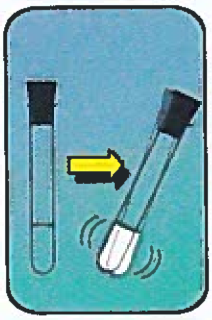
2HCl (aq) + Na_2CO_3 (s) → 2NaCl (aq) + H_2O (l) + CO_2 (g)

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.



4. Metal + Acid

A reactive metal produces hydrogen with acids

METAL + ACID → SALT + HYDROGEN

Magnesium + hydrochloric acid → Sodium chloride + hydrogen

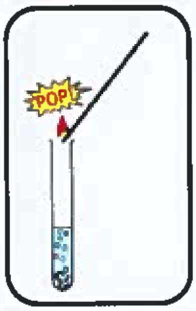
$Mg\ (s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$

Zinc + sulfuric acid → zinc sulfate + hydrogen

$Zn\ (s) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2\ (g)$

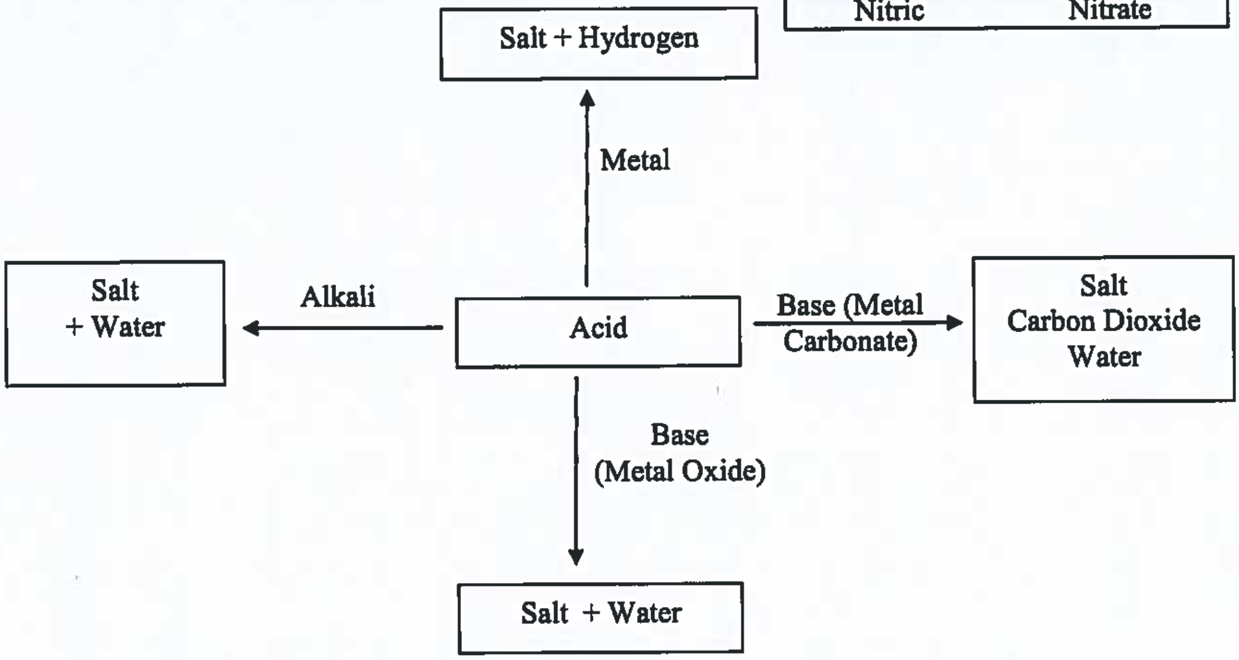
Hydrogen Test

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	Reactivity DECREASES as you descend (go down) this list of metals	
Calcium		
Magnesium		
Aluminium		
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)

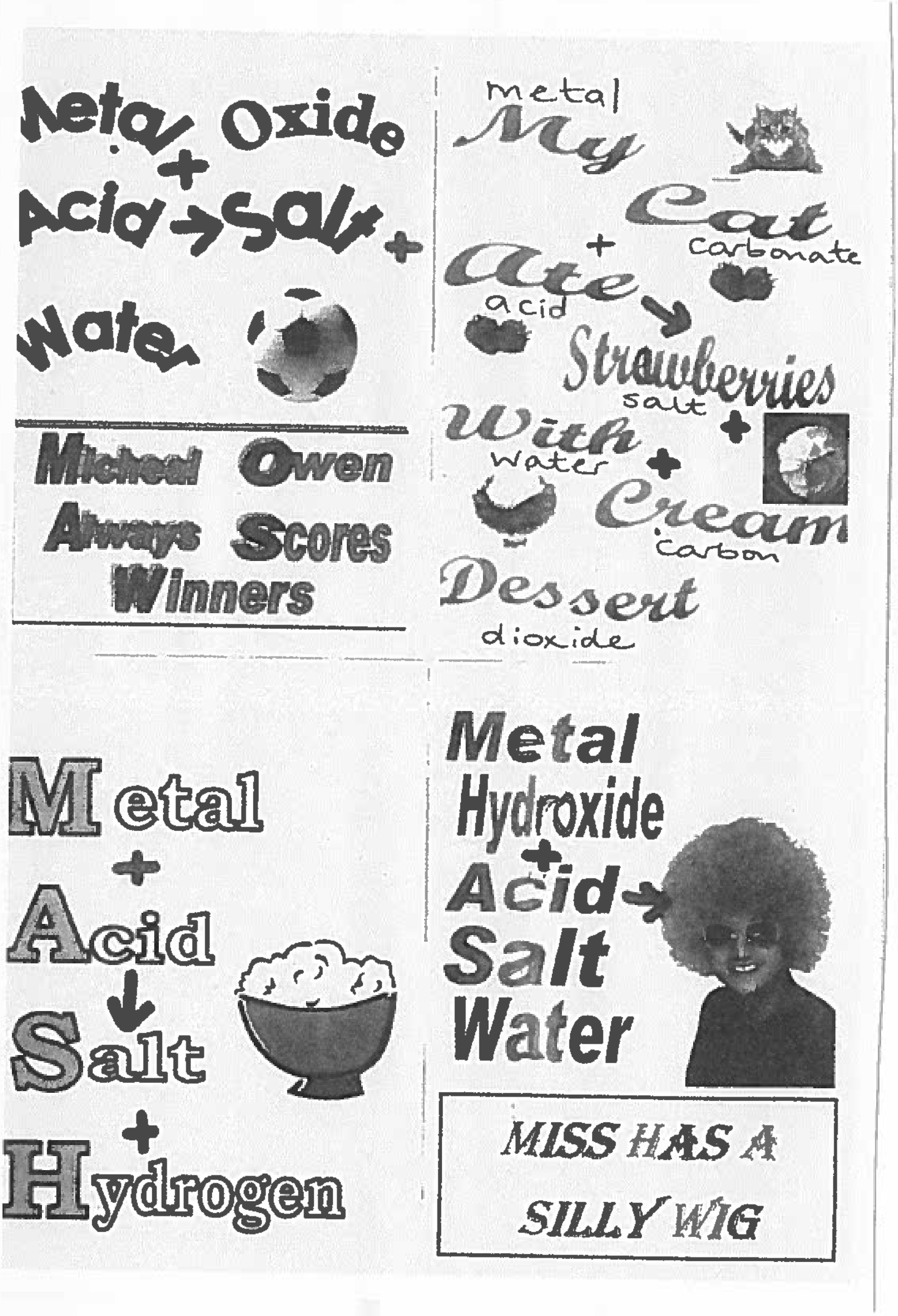
- When acids and alkalis react together they form a neutral solution (pH 7). This is called a neutralisation reaction.
- In a **neutralisation reaction the hydrogen ions (H^+) react with the hydroxide ions (OH^-) to form water (H_2O).** Water has a neutral pH = pH 7.
- The **ionic equation for neutralisation reactions is shown as:**

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_20_{(l)}$$

C27

Making Metal Salts

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



Metal + Oxide
Acid → Salt + Water

Micheal Owen Always Scores Winners

metal
My Cat
carbonate
+ Ate
acid
→ Strawberries
salt
+ With
water
+ Cream
carbon
Dessert
dioxide

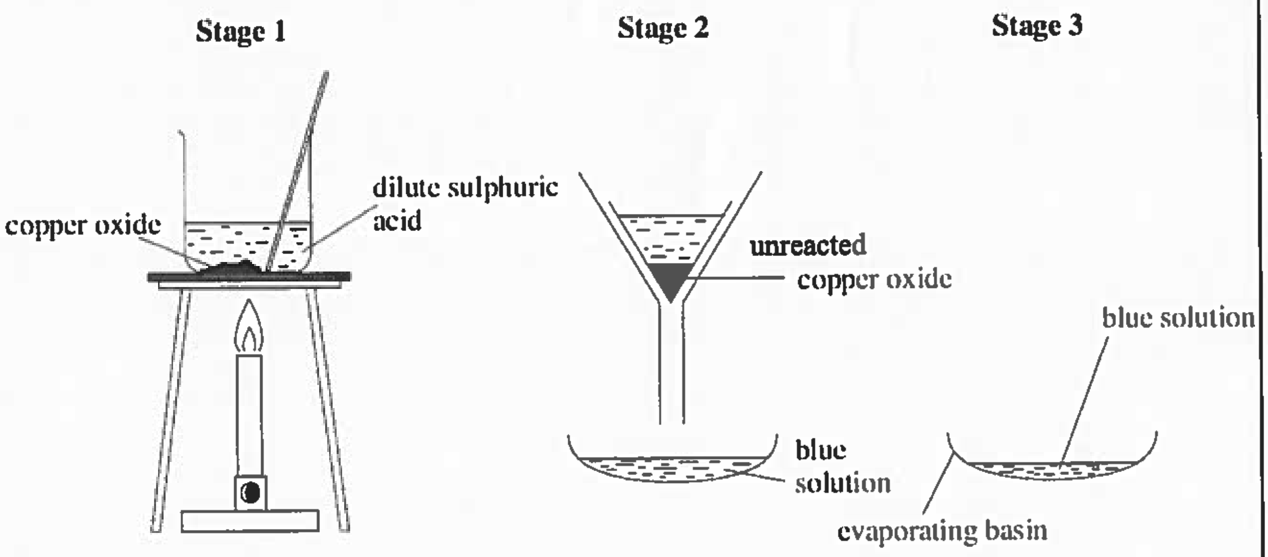
Metal
+
Acid
↓
Salt
+
Hydrogen

Metal
Hydroxide
+
Acid →
Salt
Water

MISS HAS A SILLY WIG

Method of preparing salt crystals

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



Stage 1: <u>Excess</u> 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 <u>filtration</u>, using a filter funnel and filter paper

Stage 3: Salt is obtained by <u>evaporation</u> - 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

C29

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 x 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

$^{7}_{3}Li$

Relative Atomic Mass (RAM)

One mole of Lithium has a mass of 7 g

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

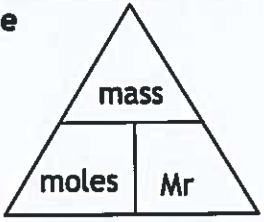
Ar (Ca) = 40, (C) = 12, (O) = 16
Mr $CaCO_3$ = 100

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

Higher Tier only

Moles and Concentration

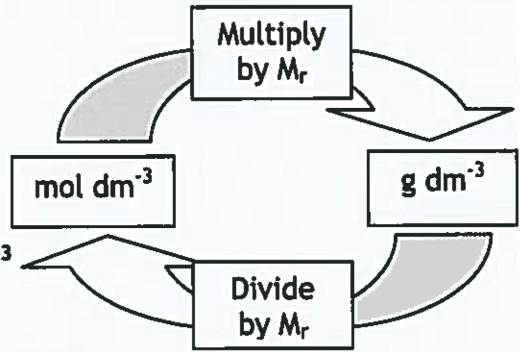
All bottles of solutions in a laboratory must be labelled with its concentration

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

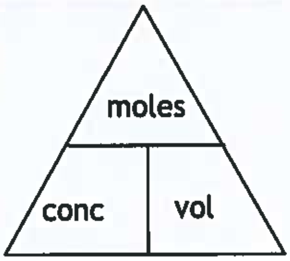


Concentration can sometimes be written as g dm^{-3}
To convert a value from mol dm^{-3} to g dm^{-3}
Multiply by Mr
e.g. 0.01 mol dm^{-3} of NaCl = 0.01 x 58 = **0.58 g dm^{-3}**

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

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



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

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

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

First we need to rearrange the original equation:

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

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

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

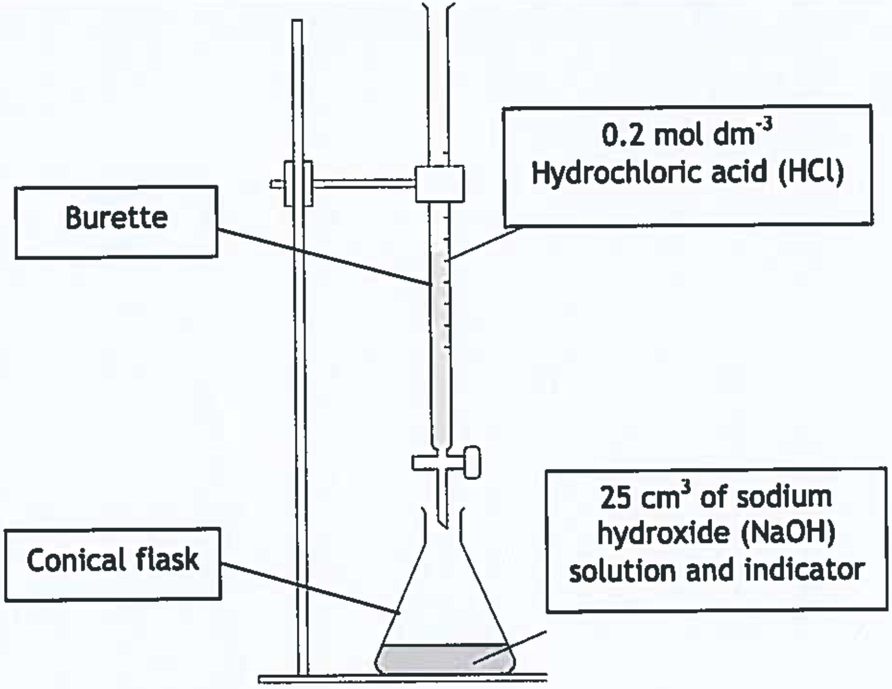
$$\text{volume} = \frac{\text{moles}}{\text{concentration}} = \frac{0.25}{1} = 0.25\ \text{dm}^3 \times 1000 = \mathbf{250\ cm^3}$$

(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

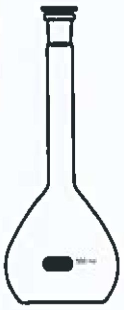


Methyl orange in alkali

End point - neutral

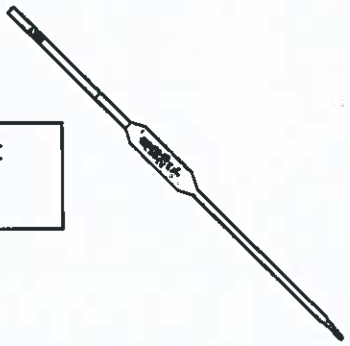
Some practical points:

- 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.



Volumetric flask

Volumetric pipette



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

C34

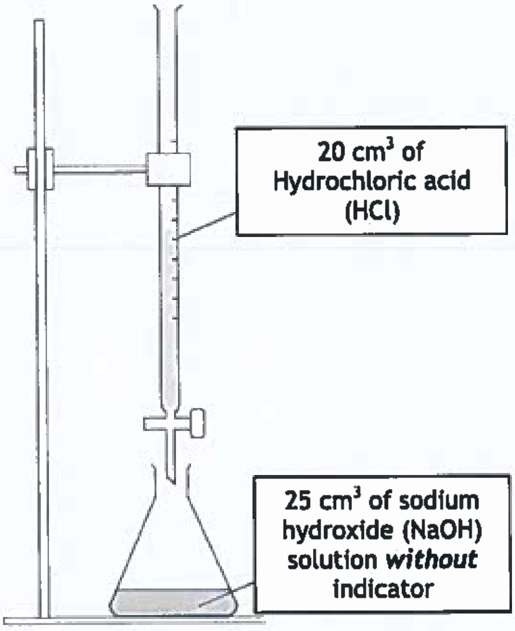
Using titration to prepare pure solutions

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.

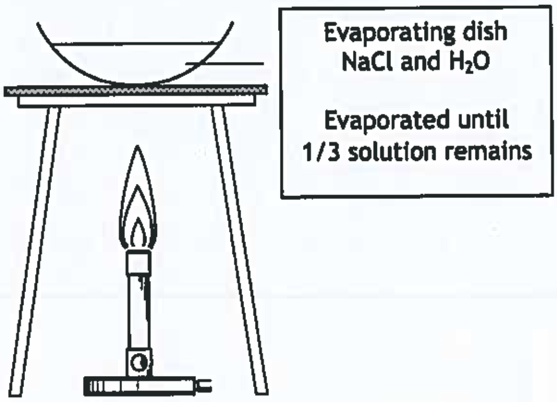
$$NaOH + HCl \longrightarrow NaCl + H_2O$$

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

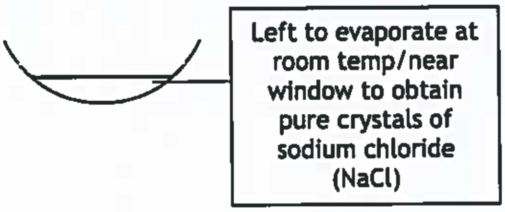
Step 1: Neutralisation



Step 2: Evaporation



Step 3: Evaporate slowly to dryness



C35

Topic 3:

Metals and their Extraction

C36

Extraction of Metals

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

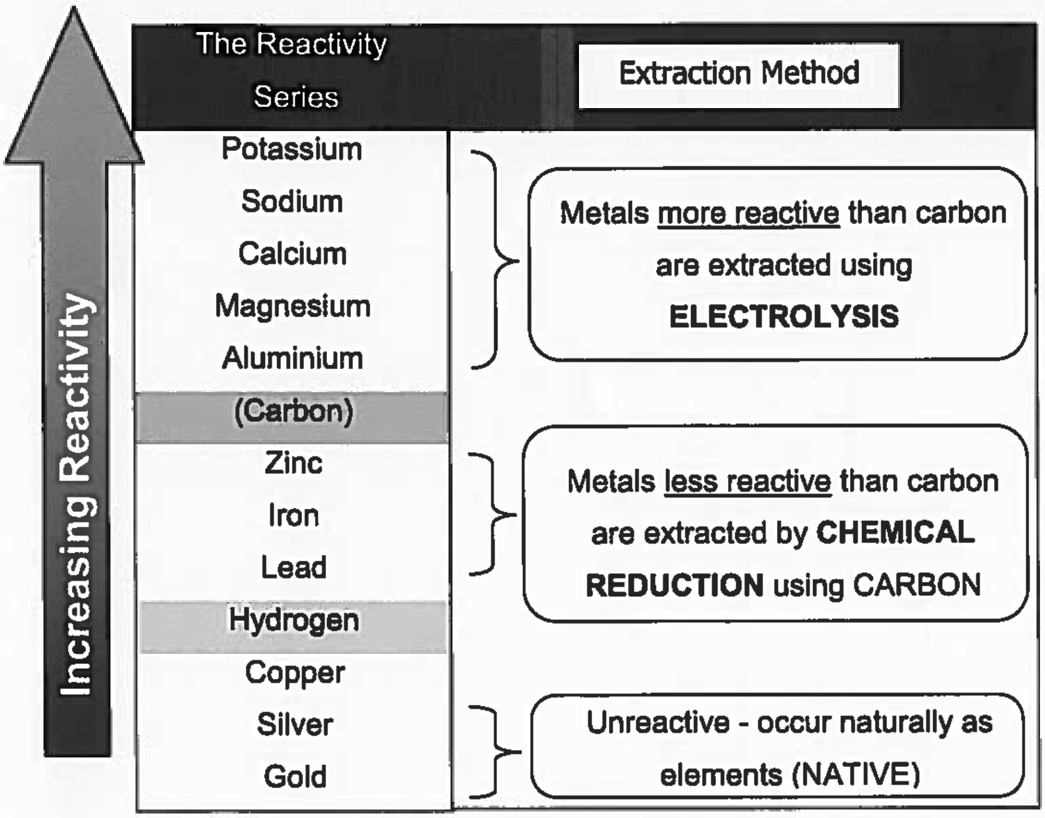
Ore	**Formula**	**Metal extracted**
Bauxite	Al_2O_3	Aluminium
Haematite	Fe_2O_3	Iron

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



The Reactivity Series	Extraction Method
Potassium	Metals more reactive than carbon are extracted using **ELECTROLYSIS**
Sodium	
Calcium	
Magnesium	
Aluminium	
(Carbon)	
Zinc	Metals less reactive than carbon are extracted by **CHEMICAL REDUCTION** using CARBON
Iron	
Lead	
Hydrogen	
Copper	
Silver	Unreactive - occur naturally as elements (NATIVE)
Gold	

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

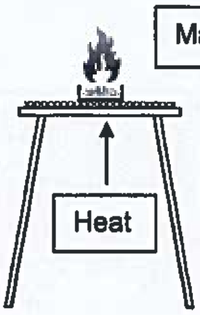
Displacement Reactions

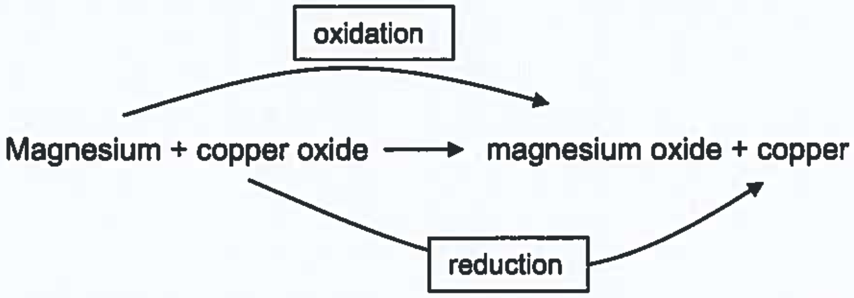
Reduction is the loss of oxygen from a compound

Oxidation is the gain of oxygen to form a compound

Magnesium and copper oxide

$Mg + CuO \longrightarrow MgO + Cu$




oxidation

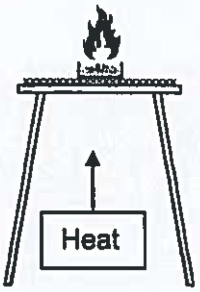
Magnesium + copper oxide → magnesium oxide + copper

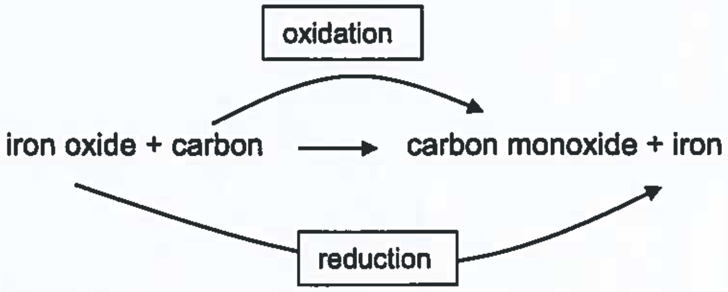
reduction

Blast Furnace Reaction

Iron oxide and carbon monoxide

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



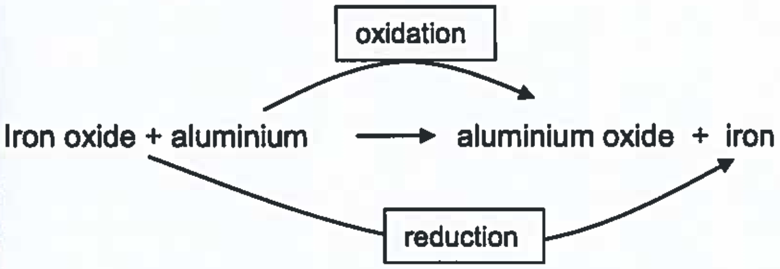

oxidation

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

C38

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

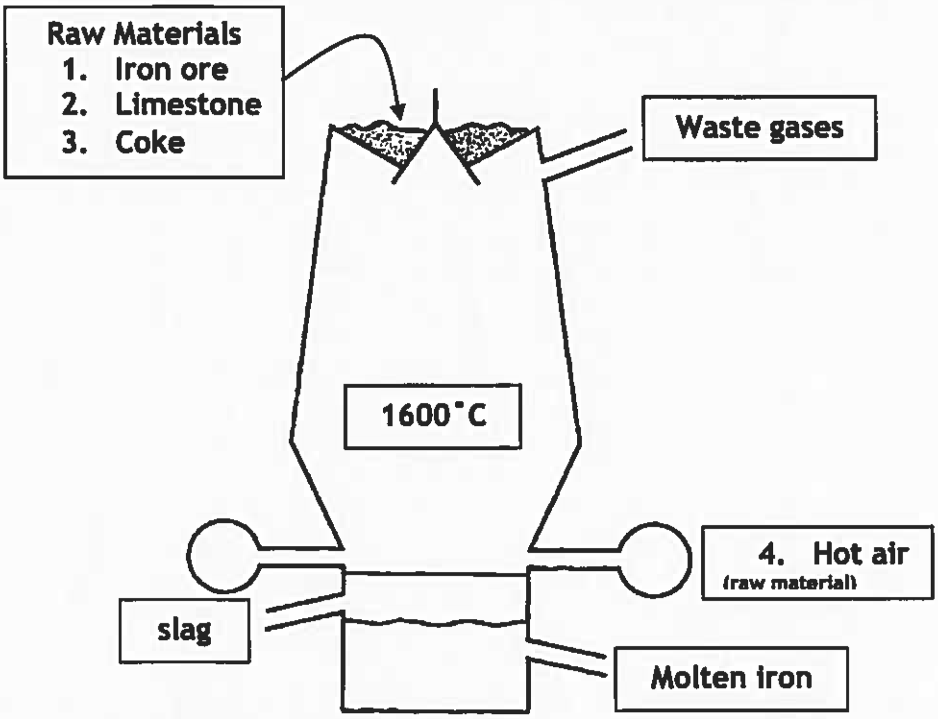
Iron ore - the source of iron

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

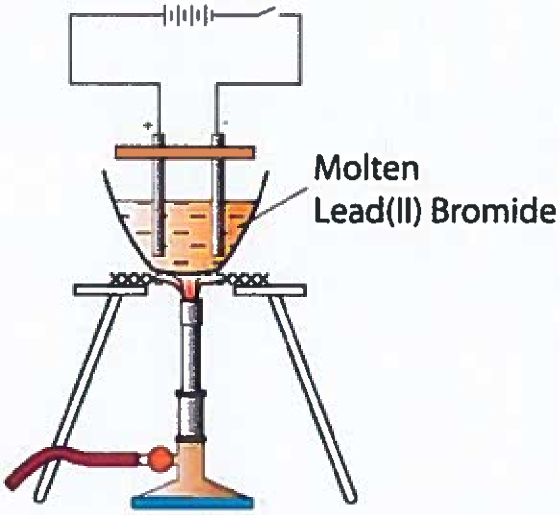
$$Fe_2O_3 + \boxed{3}CO \longrightarrow \boxed{2}Fe + \boxed{3}CO_2$$

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).

C40

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.**

C42

- **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:

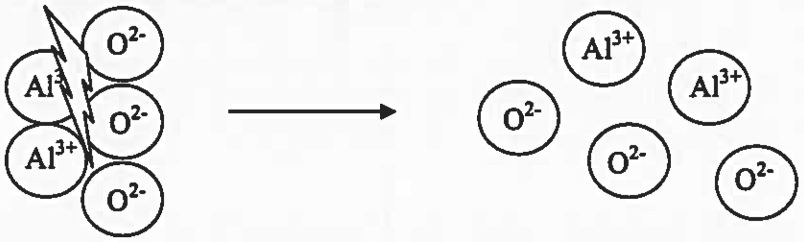
Pb^{2+} + $2e^-$ → Pb
Lead ions gain 2 electrons--------------------→ lead atoms
(reduced)

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

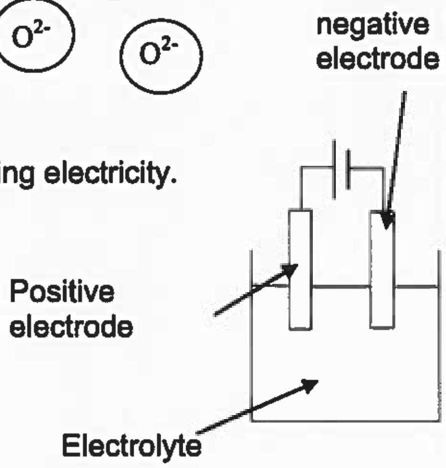
$2Br^-$ – $2e^-$ → Br_2
Bromide ions lose 2 electrons →bromine molecules
(oxidised)

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

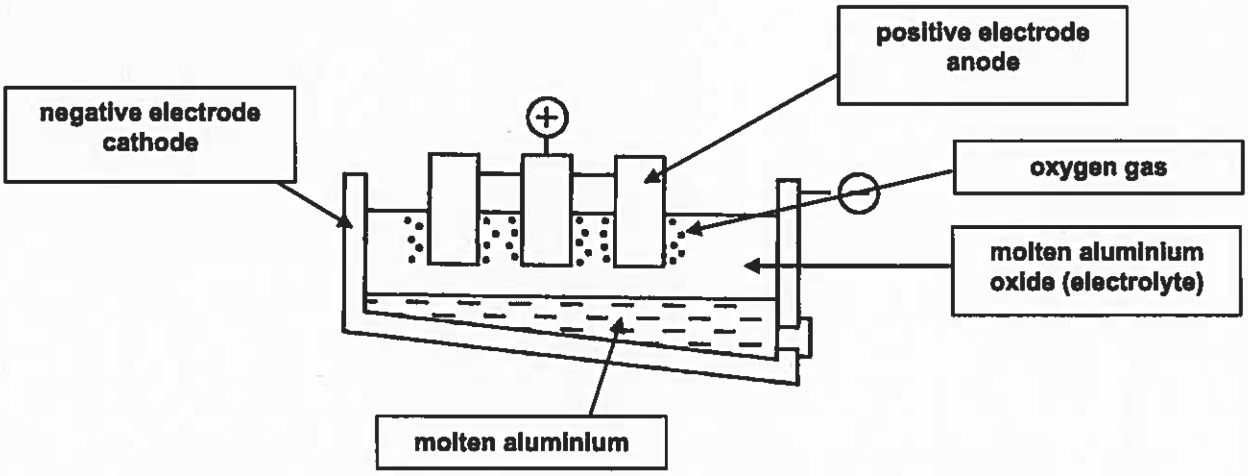
---

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.



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)

C46

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 $CuCl_2$

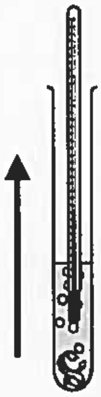
Topic 4:

Chemical Reactions and Energy

C50

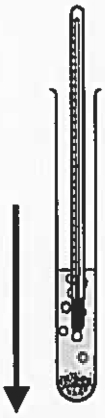
Energy in reactions

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

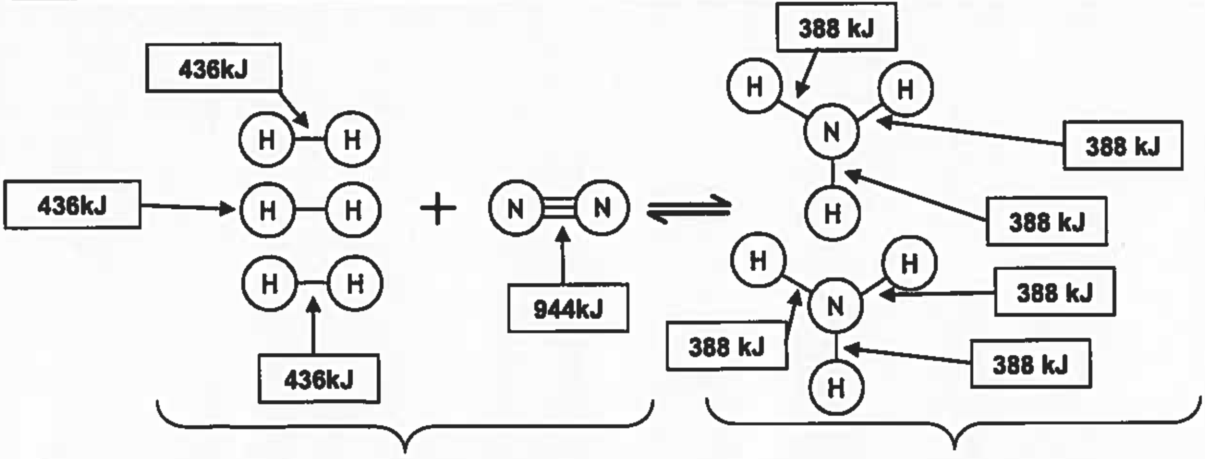
Hydrogen + Nitrogen ⇌ Ammonia

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

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)

C51

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.

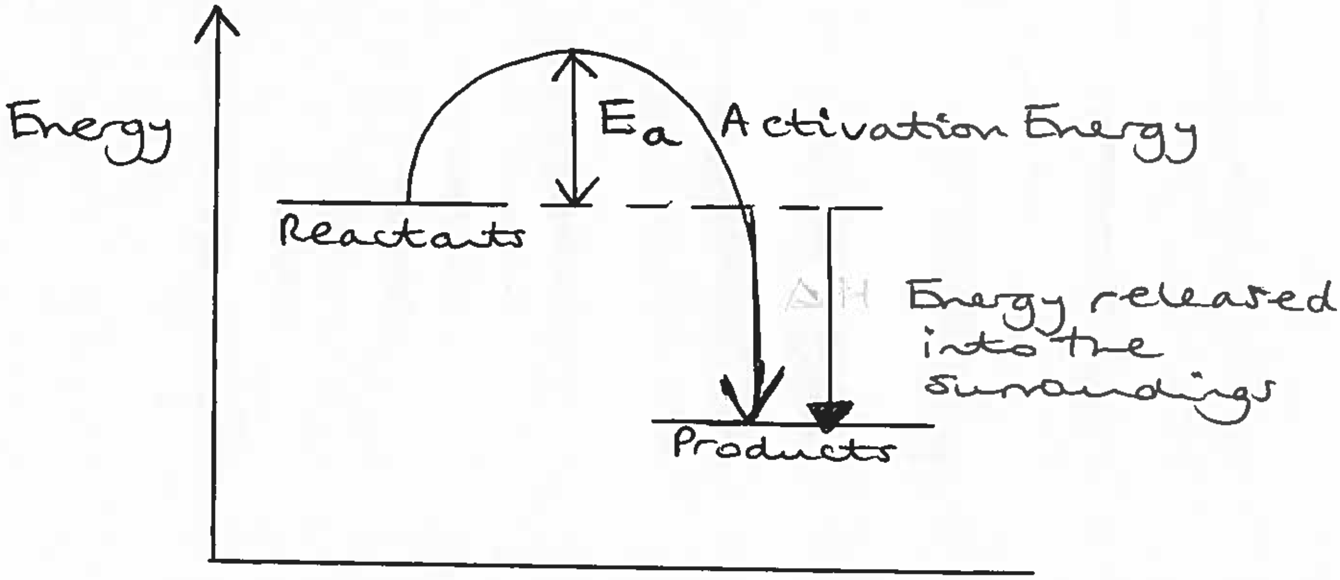
Energy is released when new bonds are formed.

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)

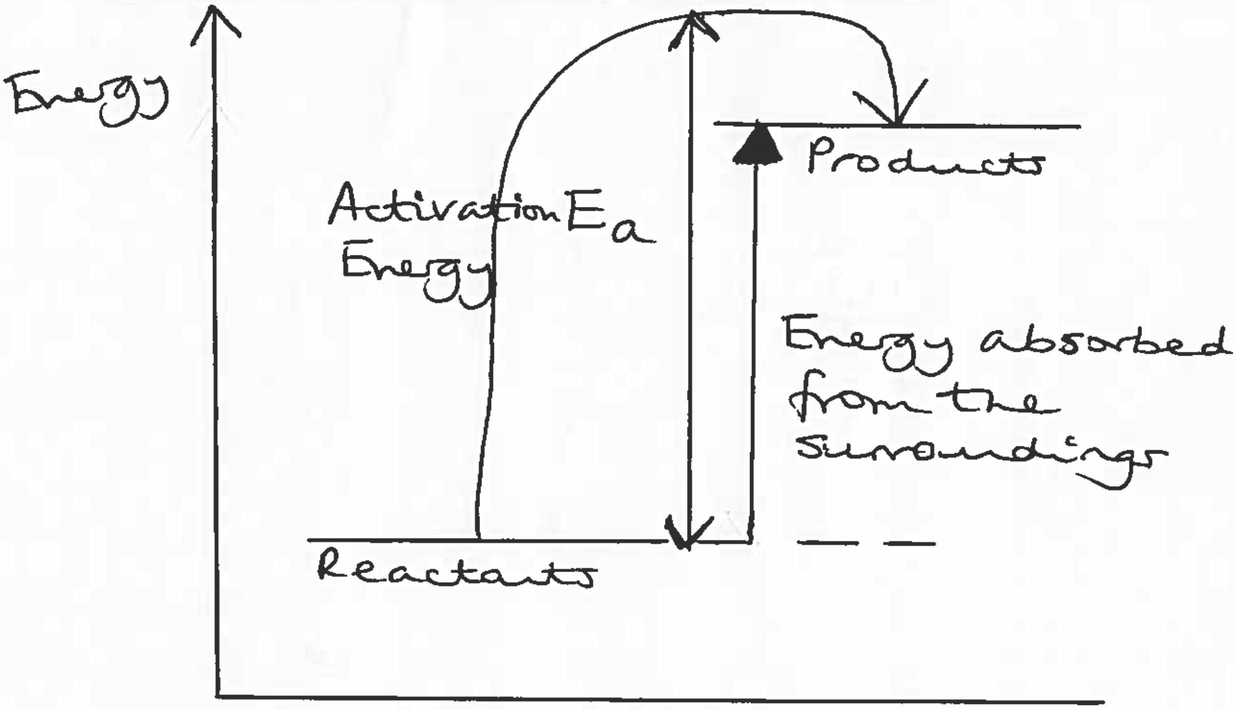
Energy Profile diagram for an Exothermic reaction



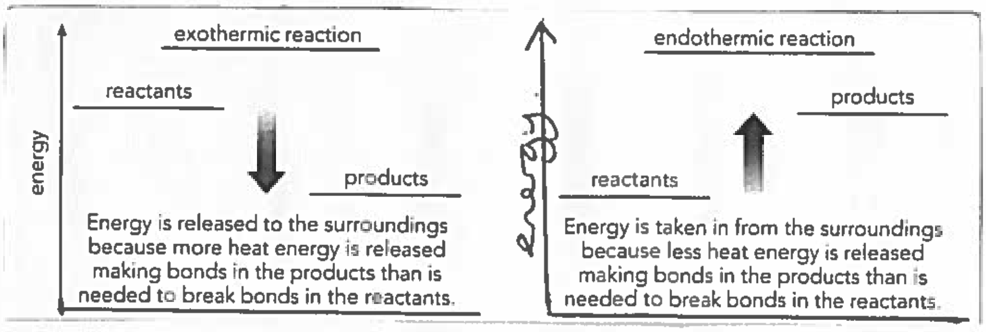
Endothermic reactions

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



Energy is released to the surroundings because more heat energy is released making bonds in the products than is needed to break bonds in the reactants.

Energy is taken in from the surroundings because less heat energy is released making bonds in the products than is needed to break bonds in the reactants.

Topic 5:

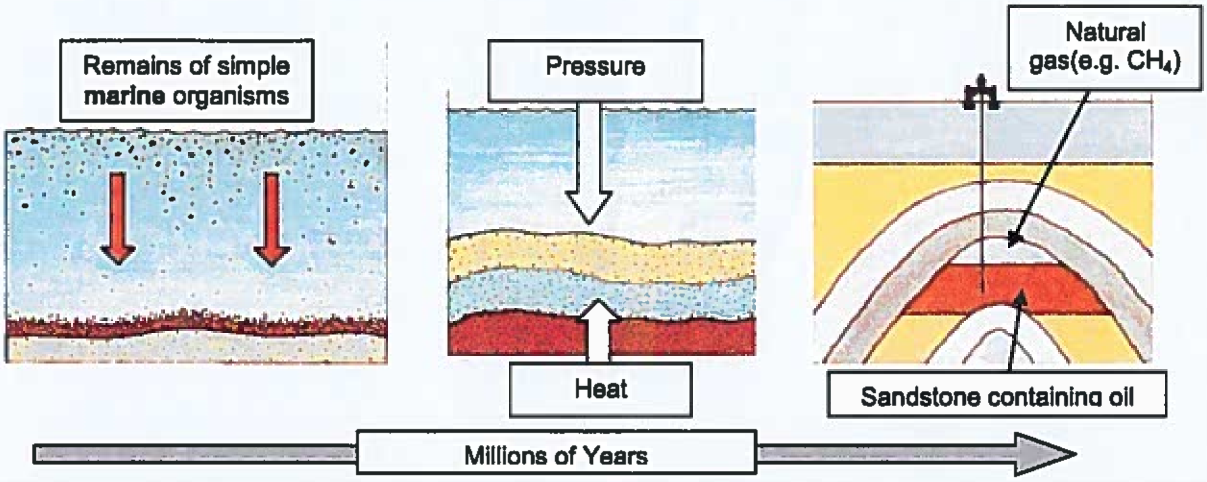
Crude Oil, Fuels and Organic Chemistry

C54

Production and uses of fuels

Crude oil (petroleum)

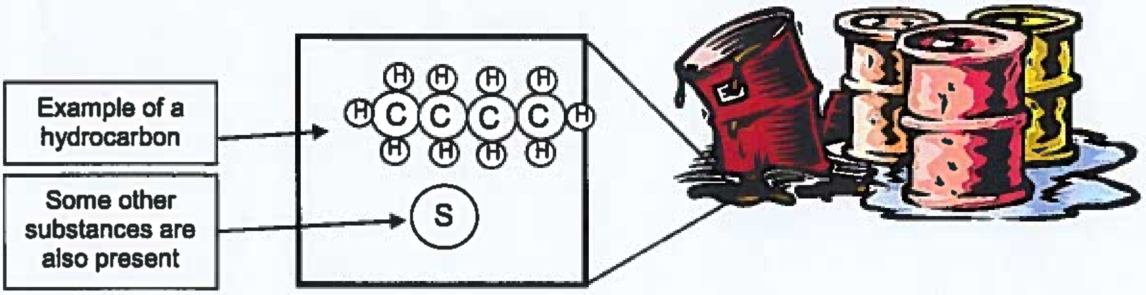
Formed over millions of years from the remains of simple marine organisms



There is a limit to coal, crude oil (petroleum) and natural gas life as they will run out over time – they are **finite** – or **non-renewable**.

Crude oil is a mixture of hydrocarbons

Hydrocarbons are compounds that contain the elements **hydrogen** and **carbon** only.

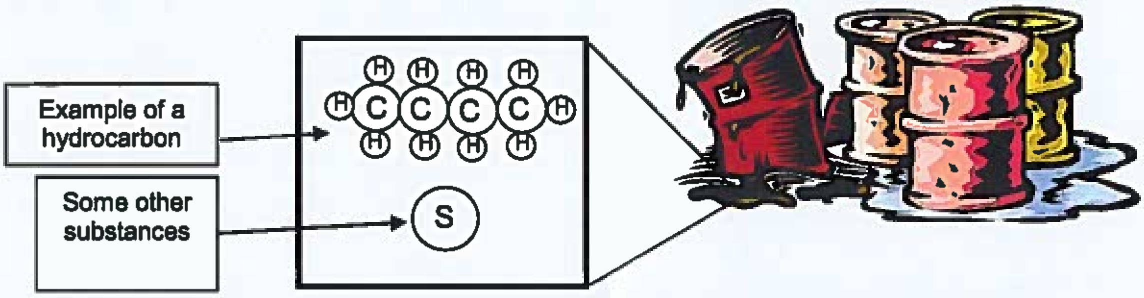


Carbon has the ability to form bonds with other carbon atoms resulting in the formation of carbon atom chains, e.g.

Crude oil contains a mixture of different sized hydrocarbon chains

Fractional Distillation

Crude oil is a mixture of different substances, most of them being hydrocarbons

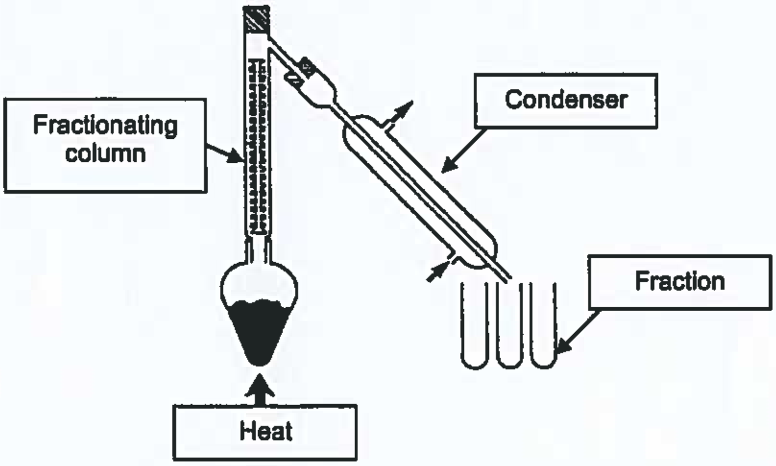


Hydrocarbons are molecules which contain the elements hydrogen and carbon only.

Fractional Distillation – it is possible to separate hydrocarbons by fractional distillation because hydrocarbons boil at different temperature ranges

Fractional Distillation in a laboratory

The reaction is carried out in a fume cupboard as poisonous gases such as sulphur dioxide can form.



As some of the hydrocarbons have similar boiling points a group of them will collect together. **Fraction** is the name given to a group of hydrocarbons that collect this way.

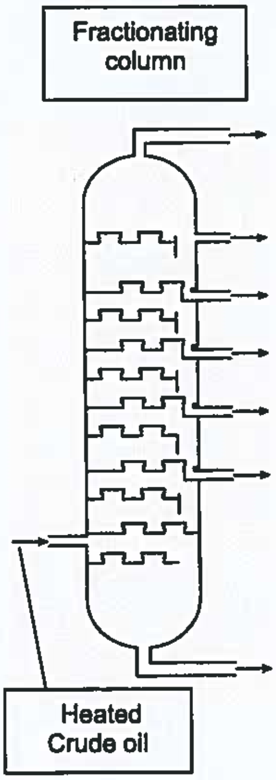
The hydrocarbon boiling point increases with the size of the carbon chain.

In the industrial process the crude oil is vaporized. The vapour is let into the column where it is hot at the bottom an cools up the column. The fractions with shorter chains have lower boiling points and can condense higher up the column. The longer hydrocarbons condense at a lower level in the column.

Fractional Distillation

Crude oil is separated into fractions

The process is called **Fractional Distillation**



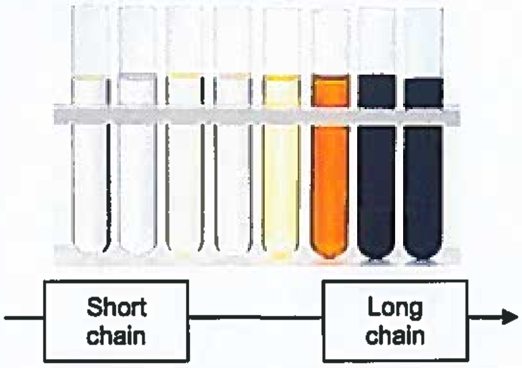
Fraction	Number of carbon atoms in a chain	Boiling point range / ˚C	Use
Gas	1 – 4	-160 to 25	Fuel
Petrol	4 – 12	40 to 100	Car fuel
Naphtha	7 – 14	100 to 150	Chemicals
Kerosene	11 – 15	150 to 250	Jet fuel
Diesel oil	15 – 19	250 to 350	Heating fuel
Lubricating oil	20 – 30	over 350	Car oil
Bitumen	C30 – above	over 400	Road pitch

Crude oil is separated into less complex mixtures, these are called fractions. Fractions contain hydrocarbons with **boiling points in the same range**, e.g. the petrol fraction has hydrocarbons with boiling points in the range 40-100 ˚C

Long chain hydrocarbons are at the **bottom** of the column as they do not boil until a very high temperature.

Some of the fractions are used as **fuels** (e.g. kerosine - aeroplane fuel) others are further processed by **cracking**.

Properties of the fractions



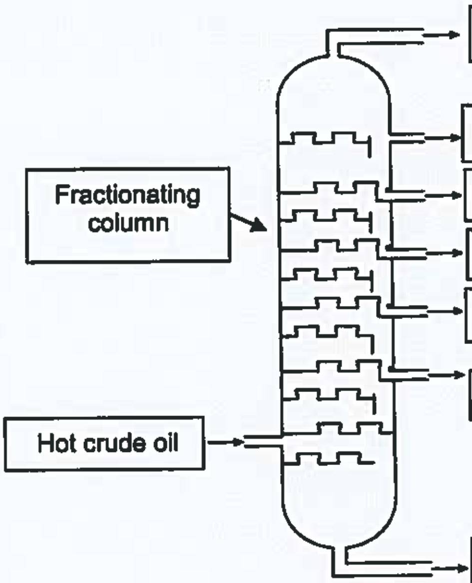
As the length of the chain increases:

1. The colour of the fraction turns from colourless - yellow - brown.
2. They are harder to ignite.
3. They burn dirtier.
4. They get more viscous

Production and uses of fuels

Crude oil is separated into fractions

The process is called **Fractional Distillation**



Fraction	No. of Carbon atoms in chain	Boiling point range / ˚C
Gas	1 – 4	-160 to 25
petrol	4 – 12	40 to 100
naphtha	7 – 14	100 to 150
paraffin/kerosine	11 – 15	150 to 250
diesel oil	15 – 19	250 to 350
Lubricating oil	20 – 30	over 350
Bitumen	C30 – and above	over 400

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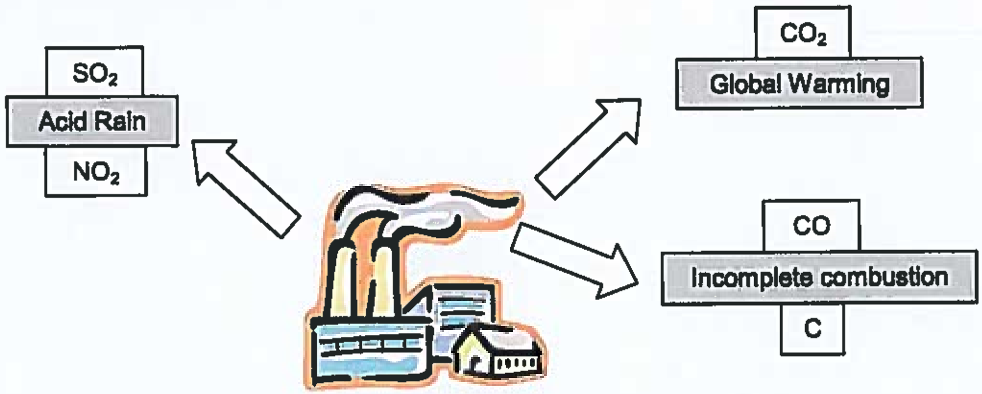
Some of the fractions are used as **fuels** (e.g. Kerosine - aeroplane fuel) others are further processed by **cracking**. (see next page)

Problems with burning fossil fuels



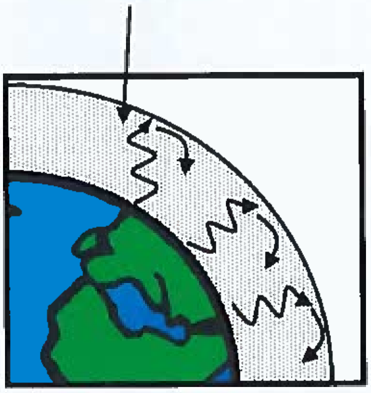
Atmosphere

Global Warming

- There is evidence to suggest that the Earth is warming but scientists do not all agree on the cause of this.

Many think that it is due mainly to increased levels of carbon dioxide in the atmosphere as a result of the combustion of fossil fuels and deforestation.

As a result the carbon cycles has been imbalanced



Higher level of carbon dioxide causes the temperature of the Earth to rise. Heat can not escape as easily

The effects of global warming

Global warming can cause :-

1. Changing weather patterns e.g. drier, hotter summers in some parts of the world leading to drought.
2. Flooding due to increase rainfall in some areas
3. Quicker melting of ice caps and glaciers
4. Rising sea levels

Carbon capture
Scientists are thinking of storing the CO_2 produced by burning fossil fuels under the sea or underground in geological formations

C59

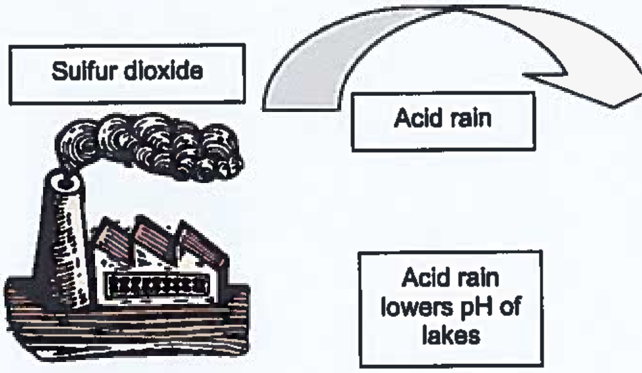
Acid Rain

In fuels such as oil and petrol there are **impurities** (i.e. oil is not pure hydrocarbons), compounds such as sulphur and nitrogen are present.

When these burn they form **polluting gases**, such as **sulfur dioxide** and **oxides of nitrogen**.

Acid rain forms when sulfur dioxide is released from factories. Acid rain forms when **sulfur dioxide** reacts with **rain** to form **sulfuric acid**.

It kills plants (forests) and aquatic life such as fish. It also damages buildings and statues made of limestone (calcium carbonate) and metals e.g. bridges.



Natural rain water has a pH of 5.5

Acid rain has a pH range of 2-4

Sulfur Scrubbing

The process of removing sulphur dioxide from exhaust flue gases of fossil fuel powered plants

Combustion of Fuels

Fuels are compounds are called **hydrocarbons**. They contain **carbon** and **hydrogen** atoms only bonded together.

Examples of fuels: crude oil, natural gas (methane), coal, petrol, diesel, ethanol, kerosene

Combustion (burning) fuels requires oxygen from the air. The hydrocarbons in the fuel react with the oxygen.

Burning hydrocarbons makes the products carbon dioxide and water.

fuel + oxygen (in air) → carbon dioxide + water

Example:

methane + oxygen → carbon dioxide + water

$CH_{4\,(g)}$ + $2O_{2\,(g)}$ → CO_2 (g) + $2H_2O_{(l)}$

All combustion reactions are exothermic. Heat energy is given out into the surroundings. The temperature increases.

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Determination of the amount of energy released by a fuel

Introduction

Fuels react with oxygen when they burn, releasing energy. You will burn four different alcohols and compare the energy they give off.

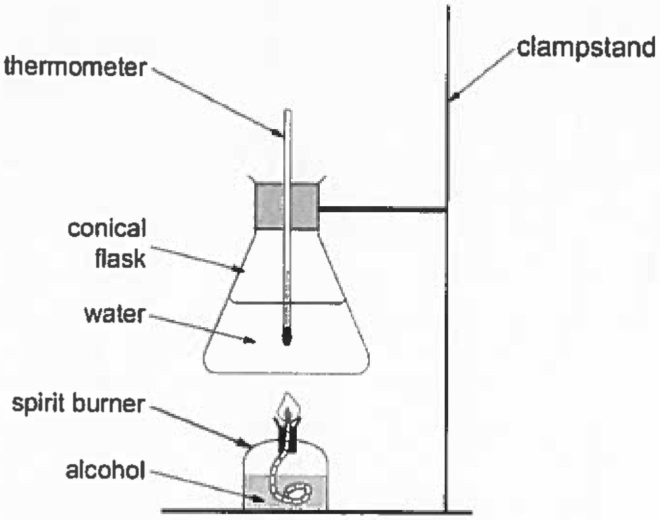
alcohol + oxygen ⟶ carbon dioxide + water

Apparatus

clamp stand, clamp and boss
250 cm^3 conical flask
100 cm^3 measuring cylinder
thermometer

Access to:
electronic balance ± 0.01 g
4 × spirit burners containing methanol, ethanol, propanol, butanol

Diagram of Apparatus



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Method

1. Measure 100 cm^3 of water into the conical flask.
2. Clamp the flask at a suitable height so the spirit burner can be placed below it (as shown in the diagram - make sure that the thermometer does not touch the bottom of the flask).
3. Record the temperature of the water.
4. Record the mass of the spirit burner (including the lid and alcohol).
5. Place the spirit burner under the conical flask and light it.
6. Allow the burner to heat the water until the temperature rises by about 40 °C. Record the temperature of the water.
7. Extinguish the flame carefully and record the mass of the burner.
8. Repeat steps 1-7 with each of the other alcohols.

Risk Assessment

Hazard	Risk	Control measure
Methanol is harmful and highly flammable	May set light to / burn individuals or equipment Vapour can cause irreversible damage	Work in a well ventilated lab Wear eye protection and ensure work station is clear
Ethanol is highly flammable	May set light to / burn individuals or equipment	Work in a well ventilated lab Wear eye protection and ensure work station is clear
Propanol is highly flammable and an irritant	May set light to / burn individuals or equipment	Work in a well ventilated lab Wear eye protection and ensure work station is clear
Butanol is highly flammable and harmful if swallowed	Vapour may irritate respiratory system and may irritate skin if spilt	Work in a well ventilated lab Wear eye protection and ensure work station is clear Rinse immediately if spilt on skin

C63

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Results

Alcohol	Initial mass of burner (g)	Final mass of burner (g)	Mass of fuel burnt (g)	Initial temperature (°C)	Final temperature (°C)	Temperature increase (°C)	Energy released per gram (J)

Analysis

1. Calculate the temperature rise for each alcohol

2. Calculate the mass of each alcohol burnt

3. Calculate the energy released by burning 1g of each alcohol using the equation below:

$$\text{energy released from alcohol per gram (J)} = \frac{\text{mass of water (g)} \times \text{temperature increase (}^{o}\text{C)} \times 4.2}{\text{mass of alcohol}}$$

C64

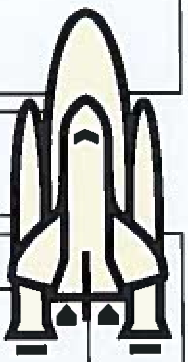
Hydrogen as a fuel

Hydrogen burns in air to make only water. The reaction is exothermic and produces a lot of energy. [exothermic – releases energy]

Hydrogen + Oxygen → Water

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

Advantages and Disadvantages of Hydrogen as a fuel



The Chevrolet Sequel car

Hydrogen is a rocket fuel.

It is also used to power hydrogen fuel cell cars.

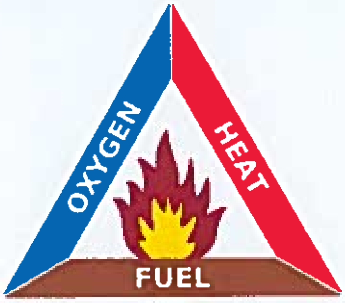
Fuel cells were invented by a Welshman Sir William Grove in 1839. It is only recently that they have been used to power cars. The technology has benefits and drawbacks.

Advantages	**Disadvantages**
Only water is produced and no carbon dioxide released – therefore it does not contribute to global warming.	Large amount of electricity needed to produce hydrogen in the first place
Does not contribute to acid rain	Storage requires bulky and heavy pressurised containers
	Safe storage is also important as hydrogen makes an explosive mixture with air

NOTE: In order for the process to remain green Hydrogen must be made by the electrolysis of water using renewable energy (solar/wind)

The Fire Triangle

The fire triangle is a simple way of understanding the factors essential for fire. Each side of the triangle represents one of the three factors required for the creation and maintenance of any fire; **oxygen, heat and fuel.** Remove any one of these, the triangle is broken and the fire is stopped.



Removing Heat
Heat can be removed by the addition of something to reduce it. Water is used to put housefires and bonfires out.

Removing Oxygen
Cover things that are burning with foam, carbon dioxide or a fire blanket to remove the air supply.

- A fire blanket is used to extinguish a chip pan fire or a person on fire.
- Carbon dioxide powder is used to put out indoor fires, chemical and electrical fires.
- Foam is used to extinguish aircraft fire.

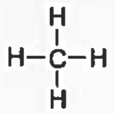
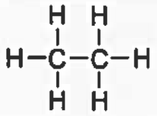
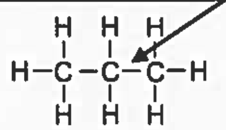
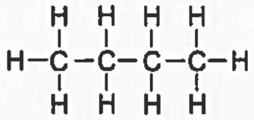
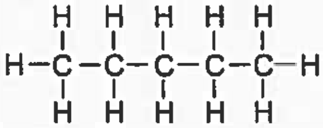
Removing Fuel
Without fuel a fire will stop.
Switch off the electrical or gas supply,
Fire-breaks are used to put forest fires out. This is when a section of trees is cleared deliberately to remove the fuel.

C66

Basic Organic Chemistry

Alkanes

These are hydrocarbons with **single covalent bonds** between the carbon atoms. They are referred to as **saturated hydrocarbon** for this reason. Alkanes have the general formula C_nH_{2n+2}

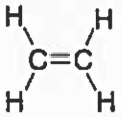
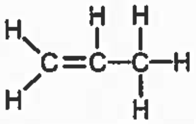
Name	Formula	Structural Formula
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	
Butane	C_4H_{10}	
Pentane	C_5H_{12}	

Single bond

Alkanes are fairly unreactive, they combust well only.

Alkenes

When there are **double bonds** between two carbon atoms the name given to the group is **alkenes**. For this reason they are described as **unsaturated** molecules. Alkenes have the general formula C_nH_{2n}

Name	Formula	Structural formula
Ethene	C_2H_4	
Propene	C_3H_6	

As a result of the double bond the alkenes are very reactive molecules, the double bond can be broken to form single bonds with other atoms (addition reaction).

C67

Cracking and Addition Polymerisation

Cracking

At high temperature long hydrocarbon chains are broken down into smaller, more useful hydrocarbons.

This can create **ethene**.

$$\text{Decane} \longrightarrow \text{Octane} + \text{Ethene}$$

Octane: Shorter more useful

Ethene: Used to make the plastic polythene

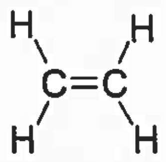
Ethene is a **small reactive molecule**, a **monomer**

If many ethene molecules are linked together it is called polythene which is used to make many plastics

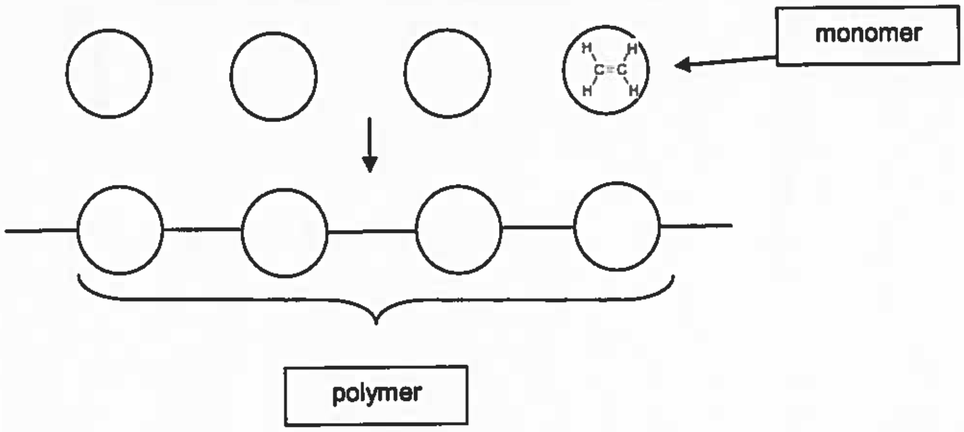
Creating Plastics

When small **reactive** molecules such as ethene react together in a chemical reaction a long chain molecule called a **polymer** is formed.

Monomer is the name given to small reactive organic molecule



Ethene has a double bond. One of the bond breaks to allow the molecule to join with another



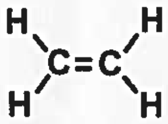
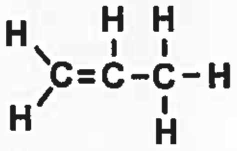
The process whereby **monomers link** to create a polymer is **polymerisation.**

The type of polymerisation that happen here is **addition polymerisation** as there is only **one product** formed

Alkenes

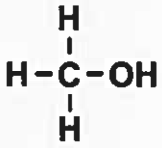
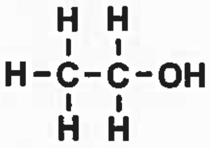
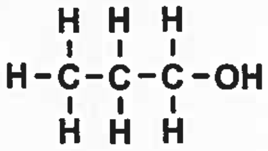
An alkene is an organic compound that contains a double bond between two carbon atoms. For this reason an alkene is referred to as an **unsaturated hydrocarbon**.

The general formula for an alkene – C_nH_{2n}

Name	**Molecular Formula**	**Structural formula**
Ethene	C_2H_4	
Propene	C_3H_6	

Alcohols

The general formula for a simple alcohol – $C_nH_{(2n+1)}OH$

Name	**Molecular Formula**	**Structural formula**
Methanol	CH_3OH	 $\begin{array}{ccccc} & & H & & \\ & & \vert & & \\ H & - & C & - & OH \\ & & \vert & & \\ & & H & & \end{array}$
Ethanol	C_2H_5OH	 $\begin{array}{ccccccc} & & H & & H & & \\ & & \vert & & \vert & & \\ H & - & C & - & C & - & OH \\ & & \vert & & \vert & & \\ & & H & & H & & \end{array}$
Propanol	C_3H_7OH	 $\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & \vert & & \vert & & \vert & & \\ H & - & C & - & C & - & C & - & OH \\ & & \vert & & \vert & & \vert & & \\ & & H & & H & & H & & \end{array}$

Functional group

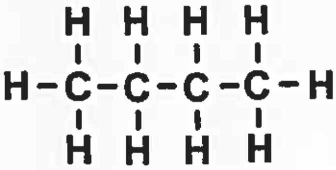
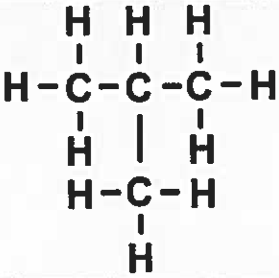
Functional groups are groups of atoms within molecules that are responsible for the characteristic chemical reactions of those molecules. The same functional group will undergo the same or similar chemical reaction(s) regardless of the size of the molecule it is a part of. The functional group for alkenes is the double bond between the carbon atoms C=C. The functional group for alcohols is the –OH group.

Isomers

In organic chemistry, isomers are molecules with the same molecular formula (i.e. the same number of atoms of each element), but different structural or spatial arrangements of the atoms within the molecule.

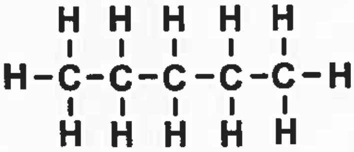
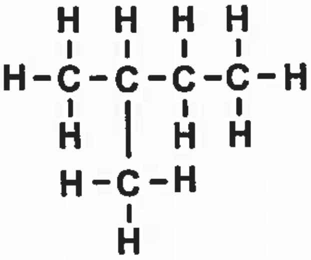
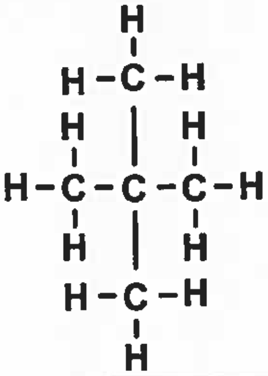
Isomer – has the same molecular formula but has a different structure.

Structural Isomers of Butane

butane	methylpropane / 2-methylpropane
	

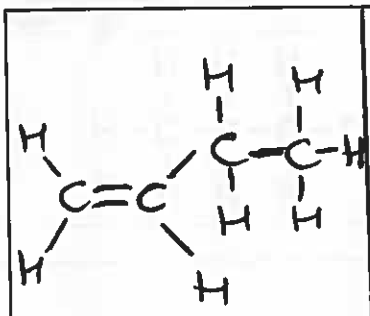
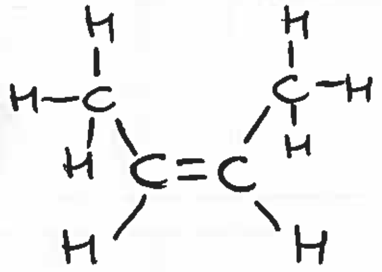
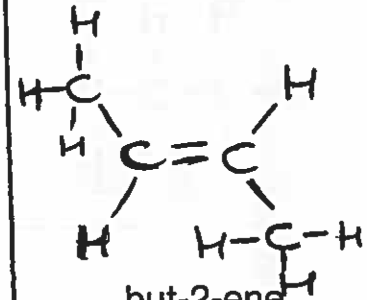
Note: You will not be required to remember the names of the isomers

Structural Isomers of Pentane

pentane	2-methylbutane	2,2-di-methylpropane
		

Isomerism in alkenes (Higher Tier only)

Isomers of the alkene C_4H_8 (molecular formula) : 3 isomers

 but-1-ene	 but-2-ene	 but-2-ene

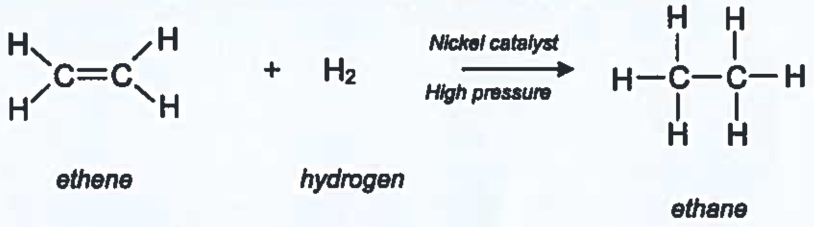
These diagrams above show the **structural formula** of the isomers.

C70

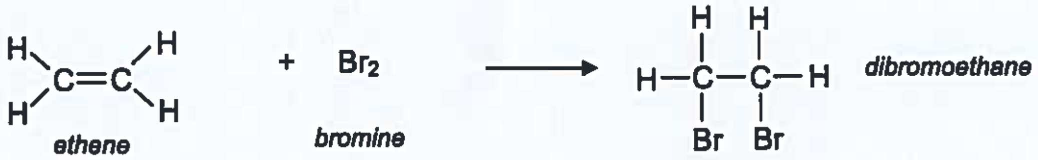
Alkenes - Addition Reactions

Reaction with Hydrogen (Hydrogenation) (Higher Tier)

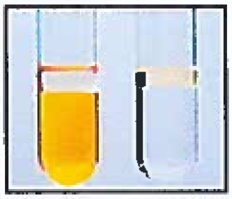
Can you write the equation for propene?



Reaction with Bromine Water (Higher Tier)



This reaction is a way of identifying alkenes. Brown bromine water turns colourless

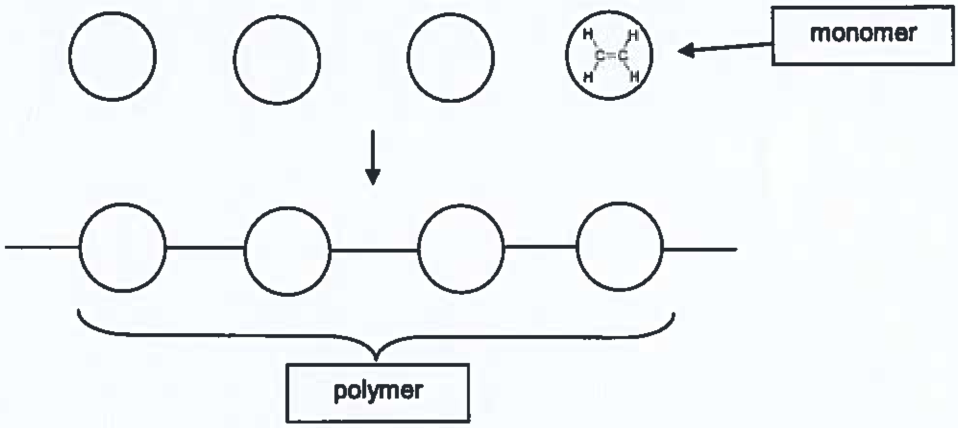


Addition Polymerisation

Creating Plastics

When small **reactive** molecules such as ethene react together in a chemical reaction a long chain molecule is formed called a **polymer**.

Monomer is the name given to small reactive organic molecule

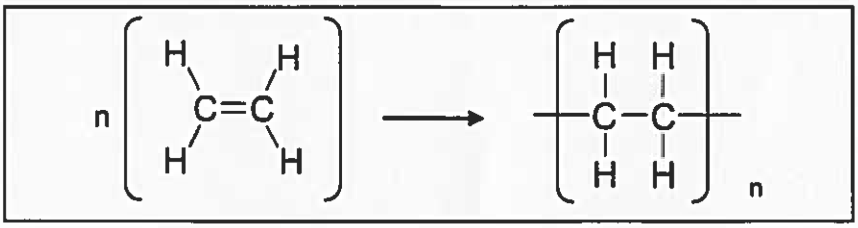


The process whereby **monomers link** to create a polymer is **polymerisation.**

The type of polymerisation that happen here is **addition polymerisation** as there is only **one product** formed

Addition Polymerisation

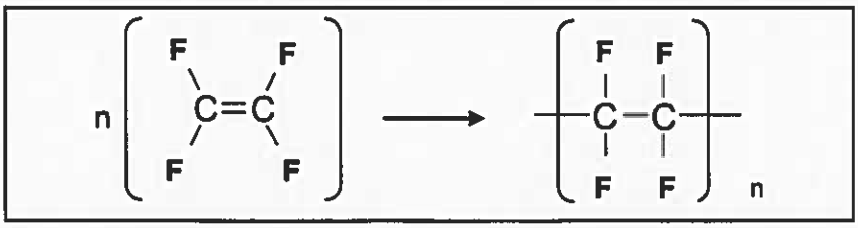
The process of making **poly(ethene)** is an example of addition polymerisation. The unsaturated monomers used are **ethene.**



ethene polythene

$_n(H_2C{=}CH_2) \longrightarrow -\!(\,CH_2 - CH_2\,)\!-_n$

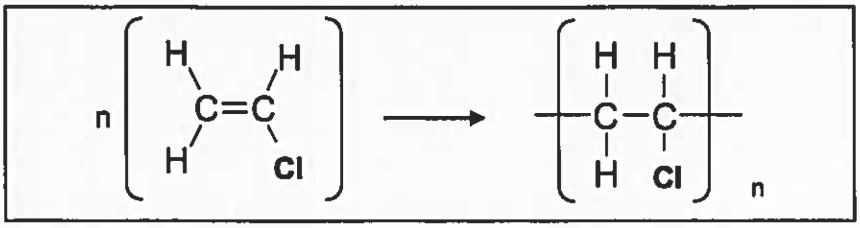
PTFE / Poly(tetrafluoroethene) /Teflon



Tetrafluoroethene Poly(tetrafluoroethene)

$_n(F_2C{=}CF_2) \longrightarrow -\!(\,CF_2 - CF_2\,)\!-_n$

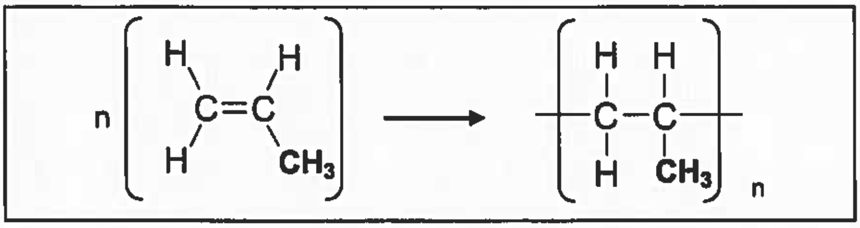
Polyvinyl chloride (PVC)



Vinyl chloride Polyvinylchloride

$_n(H_2C{=}CHCl) \quad -\!(\,CH_2 - CHCl\,)\!-_n$

Polypropene



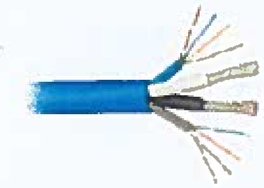
propene Polypropene

$_n(H_2C{=}CHCH_3) \quad -\!(\,CH_2 - CHCH_3\,)\!-_n$

C72

Properties of Plastics

There are many types of plastics, all made by polymerisation, e.g. polythene, PVC, PTFE (Teflon) and polystyrene.



Electrical insulator / flexible

Strong / low density

Thermal insulator

Strong / low density

Transparent / flexible

Plastics versus traditional materials

Plastics are used widely in place of natural materials such as paper and iron

PVC plastic is used to make water pipes/guttering because they are light, do not rust like iron, cheaper and last longer

Polythene is used to make plastic bags in place of paper as they are stronger, do not rip and are waterproof

The disadvantages of plastics are that the do not rot i.e. the do not decompose (takes hundreds of years) and fill landfill sites.

With heat some plastics **melt** easily ☠

If plastics burn they form **poisonous gases**

Recycling waste plastic:

1. reduces the amount of waste but equally importantly
2. conserves crude oil reserves and
3. requires less energy than making new plastics

Properties and Uses of Plastics (polymers)

Polymer plastic	Properties	Uses
polythene also called poly(ethene)	flexible, cheap, good insulator, low density	bags, plastic drinks bottles, cling film food wrapping, insulation for electrical wires
poly(propene)	flexible, shatterproof	ropes, crates, buckets, bowls
poly(vinylchloride) pvc also called poly(chloroethene)	Tough, cheap, good thermal and electrical insulator, does not rot	window frames, drain pipes, gutters
poly(tetrafluoroethene) PTFE sometimes called Teflon	tough, slippery, does not corrode or rust, good thermal and electrical insulator	non-stick coating for cooking pans, low friction surface for skis, insulation for electrical wires

C74

Year 11 GCSE Double Award Chemistry

and

Year 11 GCSE Chemistry

This section of the Year 11 Revision Guide contains work from Year 10.

You are expected to know and be able to use these ideas in your Year 11 Exams.

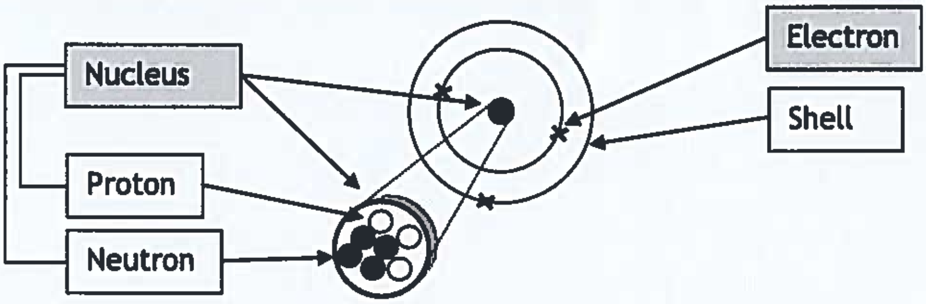
Atomic Structure

Atoms contain a **nucleus** and **electrons**

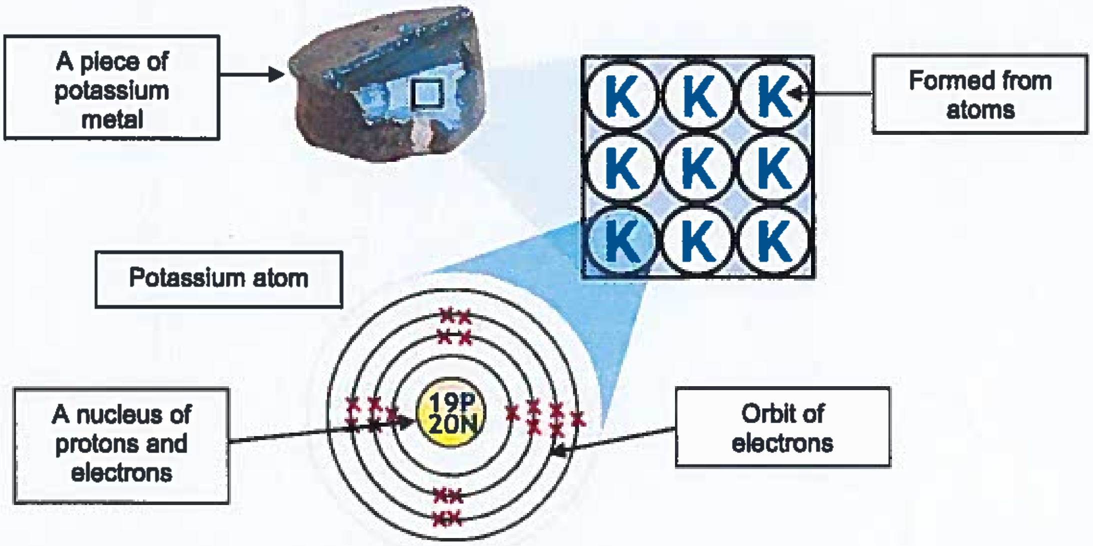
The small central nucleus is made from **protons** and **neutrons**.

Around these are **orbits** (shells) of **electrons**.

Here is a diagram showing an atom of **Lithium**



This diagram shows that a piece of **Potassium** is made up of millions of the same atom.



Atoms of **different** elements are different.

The number of **protons** is always different with different elements.

Element	*Lithium*	*Potassium*
Protons	3	19
Neutrons	4	20
Electrons	3	19

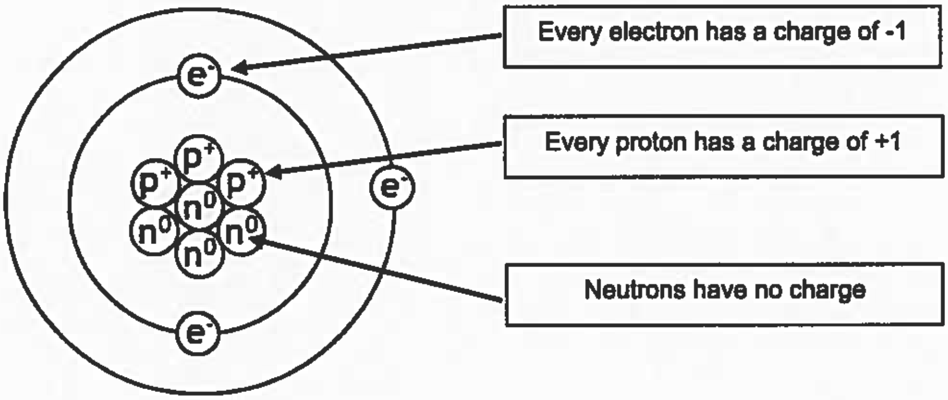
Neutron number for some elements are the same.
Electron number can be the same when the atoms have bonded.

Atomic Structure

Atoms have no charge.

The number of protons (in the nucleus) is always **the same** as the number of electrons (in shells)

Protons are positively charged. (+)
Electrons are negatively charged (-)
Neutrons do not have a charge (0)



Therefore an atom of **lithium** has no charge :- +3p + -3e = **0** **no charge**

Ion has **uneven** number of **protons and electrons**

This happens when an **electron is lost**

Or when an **electron is gained**

The proton number does not change.

Mass and **Charge** of atoms

Here are the relative mass of each particle and their electric charge.

	mass	**charge**
proton	1	+1
electron	0	-1
neutron	1	0

Protons and **neutrons** have similar mass.
Electrons have no mass, or extremely little amount.

Atomic Structure

Atomic Number → $^{7}_{\textcircled{3}}\mathrm{Li}$

Number on the **bottom** which means the number of protons or electrons

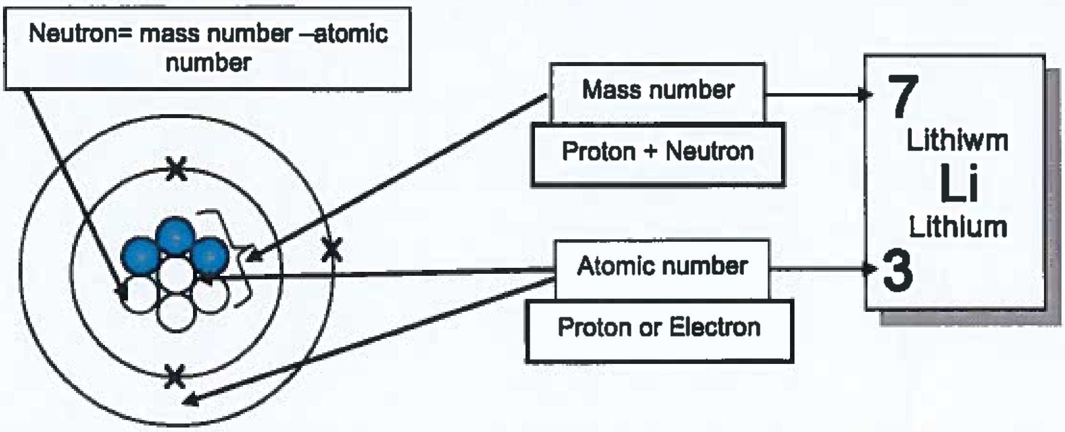
The number increases across the periodic table

Mass Number → $^{\textcircled{7}}_{3}\mathrm{Li}$

Number on the **top** which means the number of protons and neutrons in the nucleus.

Neutron Number

The number of neutrons in an atom is worked out by subtracting the number of protons (Atomic number) from the Mass number.

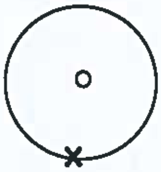


Isotopes

The same element (as it has the same number of protons) but with different number of neutrons (making the mass number different). Hydrogen

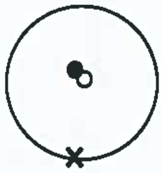
Proton = ○
Electron = X
Neutron = ●

$^{1}_{1}\mathrm{H}$



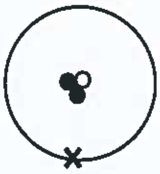
Proton number = 1
Neutron number = 0

$^{2}_{1}\mathrm{H}$



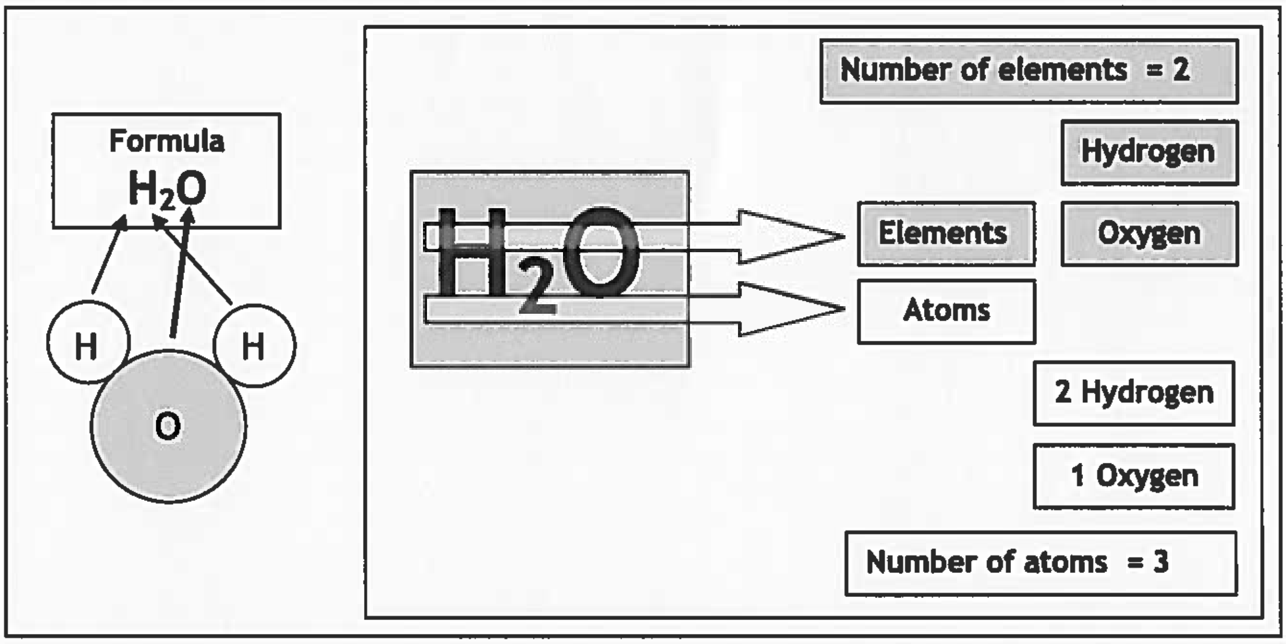
Proton number = 1
Neutron number = 1

$^{3}_{1}\mathrm{H}$



Proton number = 1
Neutron number = 2

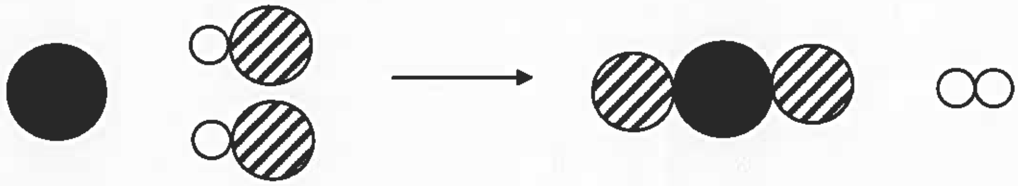
Compounds Substance that contains two or more elements joined together chemically



Compound	Formula	No. of elements	No. of atoms
Sodium Chloride	NaCl	2	2 (1 Na, 1 Cl)
Sodium Hydroxide	NaOH	3	3 (1 Na, 1 O, 1 H)
Sodium Oxide	Na_2O	2	3 (2 Na, 1 O)
Sodium Sulfate	Na_2SO_4	3	7 (2 Na, 1 S, 4 O)
Calcium Carbonate	$CaCO_3$	3	5 (1 Ca, 1 C, 3 O)

Chemical Reactions Atoms are rearranged but none are created or destroyed

e.g.



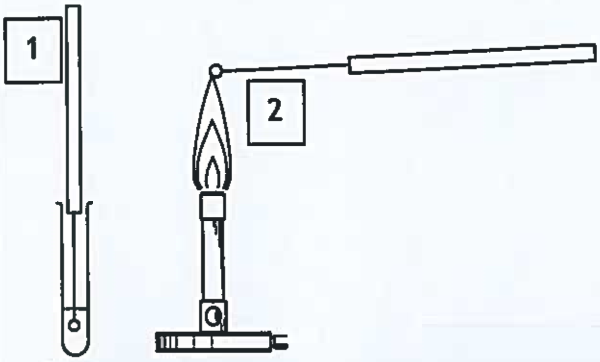
$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

Same number of atoms in reactants and products, atoms are differently arranged.

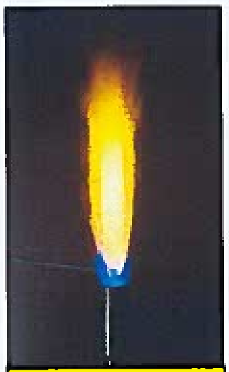
Chemical Analysis – Flame tests

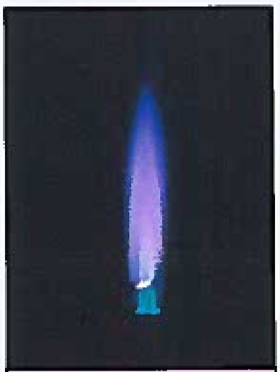
Method

1. Dip a clean wire loop in the sample solution
2. Hold the flame test wire loop at the edge of a Bunsen flame
3. Observe the changed colour of the flame, and decide which metal it indicates
4. Clean the loop in acid, rinse with water and repeat procedure with another sample



Metal	Ion	Flame test
Sodium	Na^+	Yellow-orange
Potassium	K^+	Lilac
Calcium	Ca^{2+}	Brick red
Copper	Cu^{2+}	Green


Sodium


Potassium


Calcium


Copper

Atomic Spectroscopy

This method is used to identify and show the amount (concentration) of specific atoms/ions present in the sample.

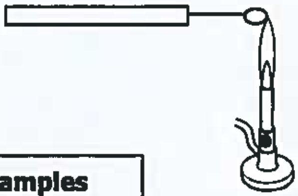
The colour of the light emitted during a flame test corresponds to a specific frequency

The intensity of the emission is measured - this corresponds to the amount of the metal present.

Identifying compounds

Metal	**Flame test**
Lithium	Red
Sodium	Yellow-orange
Potassium	Lilac

Non-metal	**Silver Nitrate test**
Chloride	white
Bromide	cream
Iodide	yellow



Flame Test
(to identify the metal)

Silver Nitrate Test
(to identify non metal ions)

Examples

Lithium Chloride

Sodium Iodide

Potassium Bromide

Red
due to lithiwm

Yellow-orange
due to sodium

Lilac
due to potassium

White precipitate
due to chloride ions

Yellow precipitate
due to iodide ions

Cream precipitate
Due to bromide ions

Higher Tier: Silver Nitrate ionic equation:
e.e. $Ag^+ (aq) + Cl^- (aq) \Rightarrow AgCl (s)$

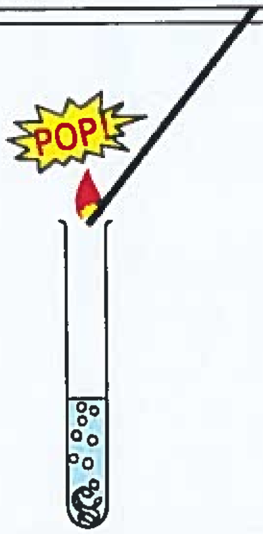
Atomic Spectroscopy (Higher Tier): This method is used to identify and show the amount (concentration) of specific atoms/ions present in the sample.

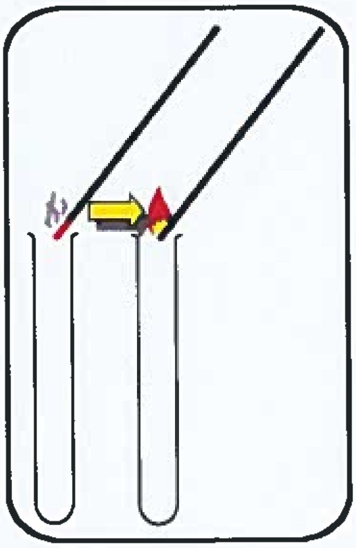
Identifying Hydrogen and oxygen gas

It is possible to test for the gases made by the electrolysis of water

Hydrogen Test

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





Oxygen Test

Oxygen will re-light a glowing splint

C82

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-}

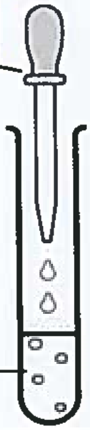
Add acid

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

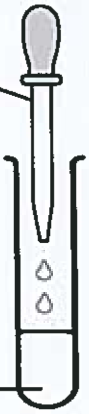
When barium chloride is added to a compound that contains a sulfate ion a white precipitate is seen - the white precipitate is barium sulfate

Example equations

barium chloride + sodium sulfate ⟶ sodium chloride + barium sulfate

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

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

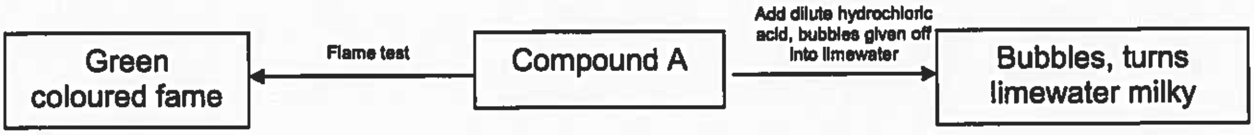


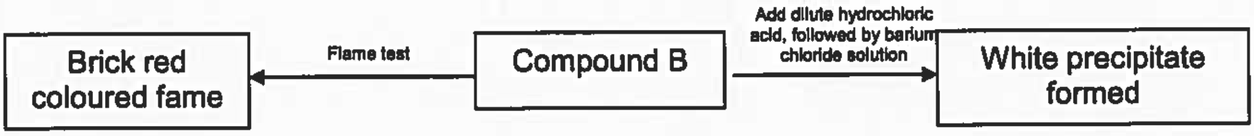
white precipitate forms

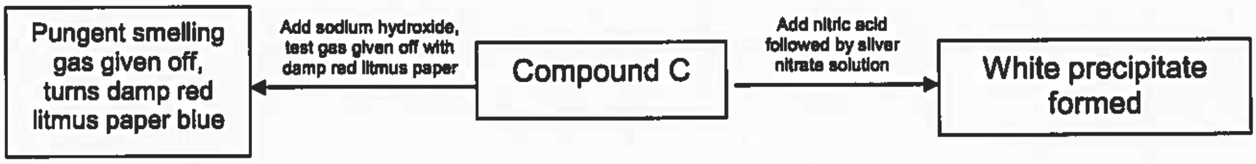
Chemical Analysis – Problem solving

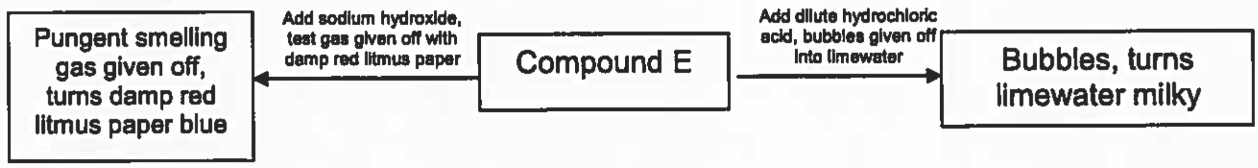
All the tests described so far can be used to identify unknown solutions. The flow charts below show tests carried out on compounds A, B and C and the results of those tests.

NH_4^+ Ca^{2+} Cu^{2+} K^+ Na^+ Cl^- Br^- I^- SO_4^{2-} CO_3^{2-}









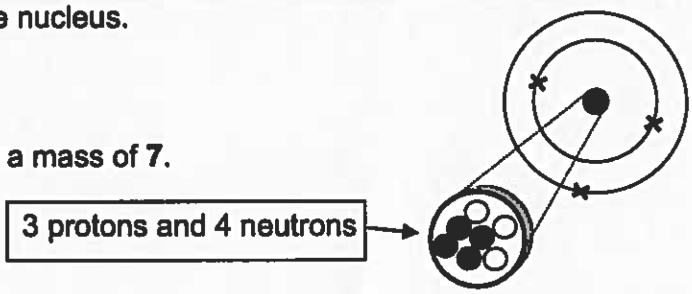
Answers: A is Copper carbonate, B is Calcium sulfate, C is Ammonium chloride, D is Sodium iodide, E is Ammonium carbonate

C84

Chemical Calculations

Every atom has different mass. This is determined by the number of protons and neutrons in the nucleus.

A lithium atom has a mass of **7**.



Relative atomic mass (A_r) is a way of saying how heavy different atoms are compared to each other.

The $\mathbf{A_r}$ of Lithium is 7 and that of Carbon is 12. We use the top number to determine this; this is called the mass number

Relative formula mass or **relative molecular mass (M_r)** is the mass for a compound (e.g. $MgCl_2$) so the masses for each element are

Mass numbers → **Mg Cl Cl**
24 + 35 + 35 = **94**

What is the molecular mass of ammonium sulphate $(NH_4)_2SO_4$?
(N=14, S=32, O=16, H=1)
Calculate $(NH_4)_2$ first = 14+1+1+1+1 = 18 x 2 = 36
S= 32
4 oxygen atoms 16 x 4 = 64
Mr= ***132***

Calculating % composition

After calculating M_r it is possible to calculate % composition, this shows how much of a specific element is in a compound in percentage form

e.g. % Mg in $MgCl_2$ = $\frac{\text{total } M_r \text{ of Mg in } MgCl_2 \times 100}{M_r\, MgCl_2}$

$\frac{24}{94} \times 100 = 25.5\ \%$

C85

Calculating Reacting Masses

By using relative atomic masses and (Ar) and relative molecular masses (Mr) it is possible to calculate how much of a product is produced or how much reactants are needed.

e.g. (product calculation)

What is the mass of **magnesium oxide** is produced when 60g of magnesium is burned in air?

Symbol Equation

$$2Mg + O_2 \rightarrow 2MgO$$

Mr =	2×24 48	2 (24+16) 80
Therefore	48g (or tonnes)	will produce 80g
	1g	80 ÷ 48 = 1.67g
	60g	will produce 60 × 1.67 = 100.2g

e.g. (reactant calculation)

What is the mass of **magnesium** needed to produce 90g of magnesium oxide?

$$2Mg + O_2 \rightarrow 2MgO$$

Mr =	2×24 48	2 (24+16) 80
Therefore	48g (or tonnes)	will produce 80g
Or	*80g of MgO will be produced with 48g of Mg*	
	1g	48÷80=0.6g
	90g	will produce 90 × 0.6 = 54g

Calculations

Determining the formula of a compound from experimental data

When 4 g of copper oxide is reduced in a steam of hydrogen, 3.2 g of copper remains.

1. Work out how much oxygen was contained in the compound

$$4 - 3.2 = 0.8 \text{ g}$$

	Cu	O
Divide with Ar	$\frac{3.2}{64}$	$\frac{0.8}{16}$
Divide with smallest	$\frac{0.05}{0.05}$	$\frac{0.05}{0.05}$
Whole number	1	1
	1 Cu	1 O

Formula = CuO

Example 2

Find the formula of iron oxide produced when 44.8g of iron react with 19.2g of oxygen. (Ar Fe = 56 and O = 16)

	Fe	O
Mass	44.8	19.2
Divide with Ar	44.8÷56	19.2÷16
	0.8	1.2
Divide with the smallest value	0.8÷0.8	1.2 ÷0.8
	1	1.5

A formula must have whole numbers therefore

	2	3

Formula = Fe_2O_3

Calculating reactants or product masses

	Reactants				Products		
	NaOH	+	HCl	⇨	NaCl	+	H_2O
	23+16 + 1		1 + 35		23+35		1+1+16
	40		36		58		18
	$\underbrace{\qquad\qquad}$ 76				$\underbrace{\qquad\qquad}$ 76		
Units	g / tones				g / tones		

Calculating the percentage yield

When we want to create a chemical, the aim is to work carefully and to produce the maximum amount possible.

The amount formed or yield is calculated in percentage. It is very unlikely that 100% yield will be achieved e.g. some might be stuck in filter paper, evaporating dish, the product might react with the air.

Example

Magnesium metal dissolves in hydrochloric acid to form magnesium chloride.

$\mathbf{Mg_{(s)}}$ **+** $\mathbf{2HCl_{(aq)}}$ $\mathbf{MgCl_{2(aq)}}$ **+** $\mathbf{H_{2(g)}}$

24g **95g**

(24 + 35.5 + 35.5)

(a) What is the **maximum theoretical mass** of magnesium chloride which can be made from 12g of magnesium?

12g 95/2 = **47.5g**

(b) If only **47.0g** of purified magnesium chloride was obtained after crystallising the salt from the solution, what is the % yield of the salt preparation?

$$\textbf{\% yield} = \frac{\text{actual amount obtained x 100}}{\text{maximum possible}}$$

$$\textbf{\% yield} = \frac{47.0 \text{ x } 100}{47.5} = \textbf{98.9\%} \text{ (to 1 decimal place)}$$

C88

Examination Reference Materials

You are given a copy of the following in your examination paper:

- **The Periodic Table of the Elements**
- **Formulae for some Common Ions.**

You do NOT need to learn these two tables.

You are expected to be able to use the information that they contain in the examination questions.

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FORMULAE FOR SOME COMMON IONS

POSITIVE IONS		NEGATIVE IONS	
Name	**Formula**	**Name**	**Formula**
Aluminium	Al^{3+}	Bromide	Br^-
Ammonium	NH_4^+	Carbonate	CO_3^{2-}
Barium	Ba^{2+}	Chloride	Cl^-
Calcium	Ca^{2+}	Fluoride	F^-
Copper(II)	Cu^{2+}	Hydroxide	OH^-
Hydrogen	H^+	Iodide	I^-
Iron(II)	Fe^{2+}	Nitrate	NO_3^-
Iron(III)	Fe^{3+}	Oxide	O^{2-}
Lithium	Li^+	Sulphate	SO_4^{2-}
Magnesium	Mg^{2+}		
Nickel	Ni^{2+}		
Potassium	K^+		
Silver	Ag^+		
Sodium	Na^+		

Avogadro's Number, L = 6×10^{23}

PERIODIC TABLE OF ELEMENTS

1	2						Group					3	4	5	6	7	0
								$^{1}_{1}\mathrm{H}$ Hydrogen									$^{4}_{2}\mathrm{He}$ Helium
$^{7}_{3}\mathrm{Li}$ Lithium	$^{9}_{4}\mathrm{Be}$ Beryllium											$^{11}_{5}\mathrm{B}$ Boron	$^{12}_{6}\mathrm{C}$ Carbon	$^{14}_{7}\mathrm{N}$ Nitrogen	$^{16}_{8}\mathrm{O}$ Oxygen	$^{19}_{9}\mathrm{F}$ Fluorine	$^{20}_{10}\mathrm{Ne}$ Neon
$^{23}_{11}\mathrm{Na}$ Sodium	$^{24}_{12}\mathrm{Mg}$ Magnesium											$^{27}_{13}\mathrm{Al}$ Aluminium	$^{28}_{14}\mathrm{Si}$ Silicon	$^{31}_{15}\mathrm{P}$ Phosphorus	$^{32}_{16}\mathrm{S}$ Sulphur	$^{35}_{17}\mathrm{Cl}$ Chlorine	$^{40}_{18}\mathrm{Ar}$ Argon
$^{39}_{19}\mathrm{K}$ Potassium	$^{40}_{20}\mathrm{Ca}$ Calcium	$^{45}_{21}\mathrm{Sc}$ Scandium	$^{48}_{22}\mathrm{Ti}$ Titanium	$^{51}_{23}\mathrm{V}$ Vanadium	$^{52}_{24}\mathrm{Cr}$ Chromium	$^{55}_{25}\mathrm{Mn}$ Manganese	$^{56}_{26}\mathrm{Fe}$ Iron	$^{59}_{27}\mathrm{Co}$ Cobalt	$^{59}_{28}\mathrm{Ni}$ Nickel	$^{64}_{29}\mathrm{Cu}$ Copper	$^{65}_{30}\mathrm{Zn}$ Zinc	$^{70}_{31}\mathrm{Ga}$ Gallium	$^{73}_{32}\mathrm{Ge}$ Germanium	$^{75}_{33}\mathrm{As}$ Arsenic	$^{79}_{34}\mathrm{Se}$ Selenium	$^{80}_{35}\mathrm{Br}$ Bromine	$^{84}_{36}\mathrm{Kr}$ Krypton
$^{86}_{37}\mathrm{Rb}$ Rubidium	$^{88}_{38}\mathrm{Sr}$ Strontium	$^{89}_{39}\mathrm{Y}$ Yttrium	$^{91}_{40}\mathrm{Zr}$ Zirconium	$^{93}_{41}\mathrm{Nb}$ Niobium	$^{96}_{42}\mathrm{Mo}$ Molybdenum	$^{99}_{43}\mathrm{Tc}$ Technetium	$^{101}_{44}\mathrm{Ru}$ Ruthenium	$^{103}_{45}\mathrm{Rh}$ Rhodium	$^{106}_{46}\mathrm{Pd}$ Palladium	$^{108}_{47}\mathrm{Ag}$ Silver	$^{112}_{48}\mathrm{Cd}$ Cadmium	$^{115}_{49}\mathrm{In}$ Indium	$^{119}_{50}\mathrm{Sn}$ Tin	$^{122}_{51}\mathrm{Sb}$ Antimony	$^{128}_{52}\mathrm{Te}$ Tellurium	$^{127}_{53}\mathrm{I}$ Iodine	$^{131}_{54}\mathrm{Xe}$ Xenon
$^{133}_{55}\mathrm{Cs}$ Caesium	$^{137}_{56}\mathrm{Ba}$ Barium	$^{139}_{57}\mathrm{La}$ Lanthanum	$^{179}_{72}\mathrm{Hf}$ Hafnium	$^{181}_{73}\mathrm{Ta}$ Tantalum	$^{184}_{74}\mathrm{W}$ Tungsten	$^{186}_{75}\mathrm{Re}$ Rhenium	$^{190}_{76}\mathrm{Os}$ Osmium	$^{192}_{77}\mathrm{Ir}$ Iridium	$^{195}_{78}\mathrm{Pt}$ Platinum	$^{197}_{79}\mathrm{Au}$ Gold	$^{201}_{80}\mathrm{Hg}$ Mercury	$^{204}_{81}\mathrm{Tl}$ Thallium	$^{207}_{82}\mathrm{Pb}$ Lead	$^{209}_{83}\mathrm{Bi}$ Bismuth	$^{210}_{84}\mathrm{Po}$ Polonium	$^{210}_{85}\mathrm{At}$ Astatine	$^{222}_{86}\mathrm{Rn}$ Radon
$^{223}_{87}\mathrm{Fr}$ Francium	$^{226}_{88}\mathrm{Ra}$ Radium	$^{227}_{89}\mathrm{Ac}$ Actinium															

Key: