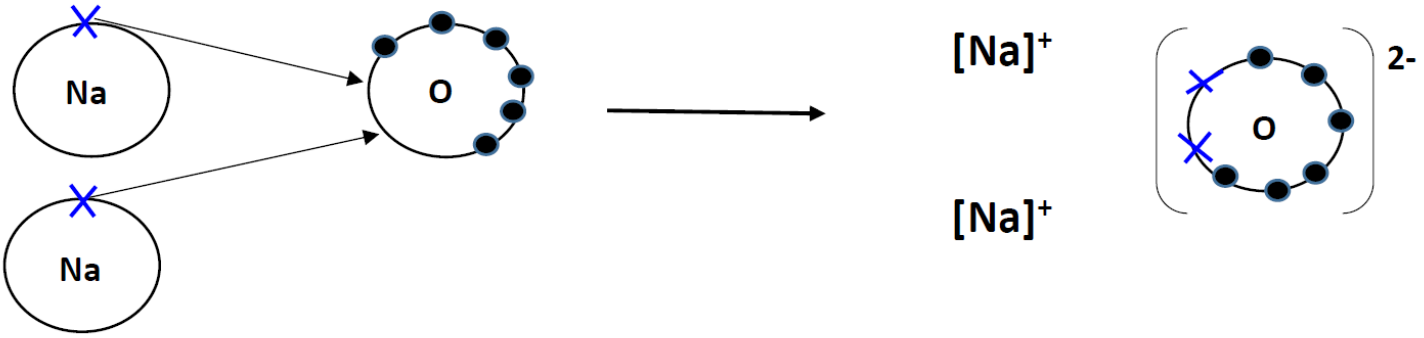
Unit 5 Chemistry GCSE Double Award
Revision Guide

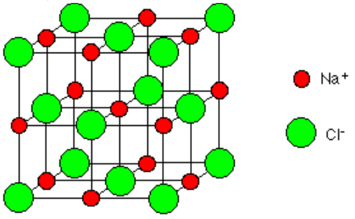
Bonding & Structures

Ionic bonding

When a metal atom transfers electrons to a non-metal atom to form full outer shells.
As a result, ions are formed (metal ion – positive; non-metal ion – negative).
E.g. Sodium Oxide



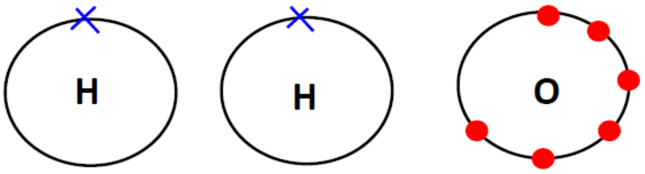
Properties of Giant Ionic compounds:

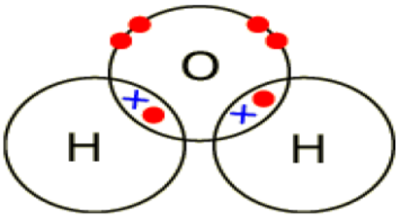


- **High melting point** – Strong forces of attraction between the opposite charged ions (Greater the charge, the stronger the forces between the ions).
- **Conduct electricity when molten** – When molten, the negative ions are free to move, allowing electric current to flow.
- **Soluble in water** – Water molecules can surround the opposite charged ions.

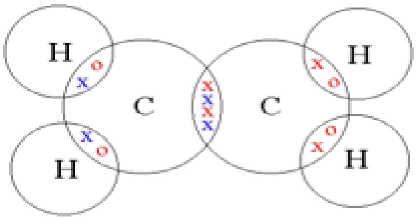
Covalent bonding

When two or more non-metal atoms share electrons to produce a full outer shell. If an element only has one atom, the other elements surround it.
E.g. Water (H_2O)



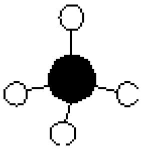


The electrons that are shared between the atoms **MUST** go in pairs. A pair of electrons shared represents a bond.



Ethene has a double bond as two pairs are shared between the carbon atoms.

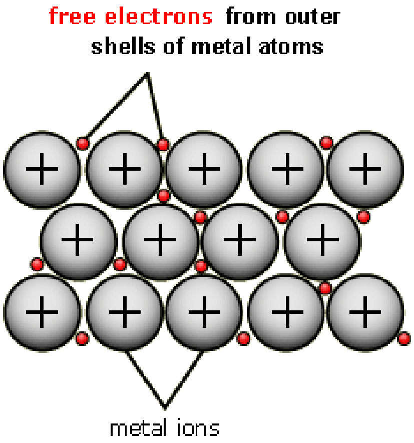
Simple molecular covalent compounds properties:



- **Low melting point** – Weak forces of attraction between the molecules.
- **Does not conduct electricity** – No free moving electrons.
- **Insoluble in water**

Metallic

Structure consists of metal cations (positive atoms) surrounded by free moving (delocalised) electrons.

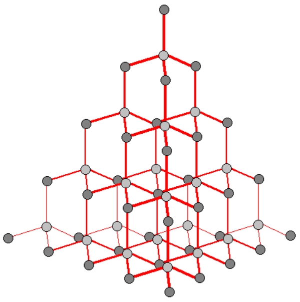


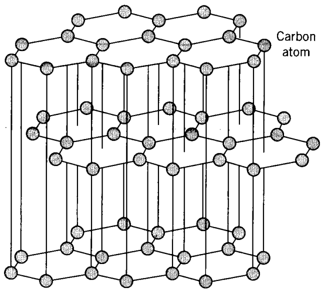
Properties of Metallic elements:

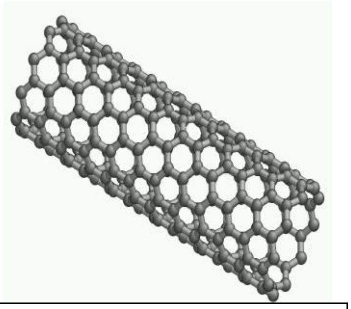
- **High melting point** – Strong forces of attraction between the metal cation and the delocalised electrons (Greater the cation charge, the stronger the force between the cation and electrons).
- **Conduct electricity when solid** – The electrons are free moving, allowing electric current to flow.
- **Malleable & Ductile** – When force is applied, the cations slide over each other with the free electrons keeping the structure intact.

Diamond, Graphite and Nanotubes (Giant covalent structure)

These are allotropes of carbon (different forms of structure from the same element).







Properties	Diamond	Graphite	Nanotube
Structure	Large tetrahedral arrangement (like a pyramid)	Layers of hexagonal carbon rings, held together by weak forces between the layers.	Single layers, from the graphite structure, rolled into a tube.
Melting/Boiling point	Very high (strong covalent bonds between the atoms)	Very high (strong covalent bonds between the atoms)	Very high (strong covalent bonds between the atoms)
Conducts electricity	No (no free electrons)	Yes (free moving electrons)	Yes (free moving electrons)
Hardness	Very hard	Brittle (weak forces allow the layers to slide over each other).	Hard
Uses	Cutting tool blades	Lubricants, pencils.	Miniature electric circuits

Smart Materials

A material whose properties change (e.g. colour, shape) due to a change in the surroundings of the environment (e.g. light, temperature, pH). These changes are reversible.

Smart material	Property Change	Environment change	Uses
Thermochromic pigment	Colour	Temperature	Colour on mugs, baby feeding spoons, forehead thermometers.
Photochromic pigment	Colour	Light	Sunglasses lenses
Hydrogels	Absorb/shrink water due to its cross-linked structure	Temperature or pH	Nappies, fake snow, self-watered plants, contact lenses
Shape memory alloys (Nitinol – NiTi)	Regains its original shape	Heat	Coffeepot thermostats, spectacle frames.
Shape memory polymers	Regains its original shape	Heat	Self-repair car bodies, stents for veins, gum shields

Nanoscience

Nano science is a study of particles whose size range lies between 1-100 nm.
A nanometre is 1×10^{-9} m (very, very small).

The properties of a nanoparticle will differ to the properties of the same material on a larger scale.

E.g. Silver is used in jewellery and cutlery as it does not corrode. Silver nanoparticles have anti-bacterial, anti-fungal and anti-viral properties, making it suitable for sterilizing sprays, plasters, inner coating of refrigerators and deodorants.

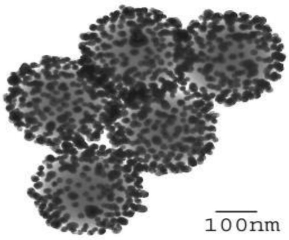
CURAD
silver
Antibacterial
CURAD

Other examples of nanoparticles

- Titanium Dioxide – Used in waterproof sunscreens and self-cleaning windows.

Disadvantages of nanoparticles

- Long term effects of nanoparticles are unknown.
- Nanoparticles are so small; they can be easily inhaled/absorbed into the body.



Reactions with Acids

Substance	pH	Colour with Universal Indicator	Strong or weak
Acid	1 – 6	1 = Red; 6 = Yellow	1 = Strong; 6 =Weak
Alkali	8 – 14	8 = Blue; 14 = Purple	8 = Weak; 14 = Strong
Neutral	7	Green	

Acids contain Hydrogen ions (H^+) whilst alkalis contain hydroxide ions (OH^-). When an acid reacts with an alkali, water is produced. This is an example of **NEUTRALISATION**.

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

Reaction with metal

ACID + METAL → SALT + HYDROGEN

When acids react with metals, it produces a salt and hydrogen gas.
We would observe fizzing and heat is released (an exothermic reaction).

Hydrogen gas is released when metals are more reactive than hydrogen in the reactivity series. The more reactive the metal, the greater the volume of hydrogen gas produced.

Reaction with alkalis and bases

An alkali is a soluble base (metal hydroxide). A base is a metal oxide.

ACID + ALKALI → SALT + WATER
ACID + BASE → SALT + WATER

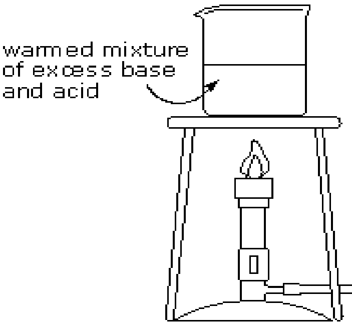
When an acid reacts with an alkali/base, a salt and water are produced. This is an example of **NEUTRALISATION**. Neutralisation is also an exothermic reaction, therefore the temperature will rise.

Reaction with carbonates

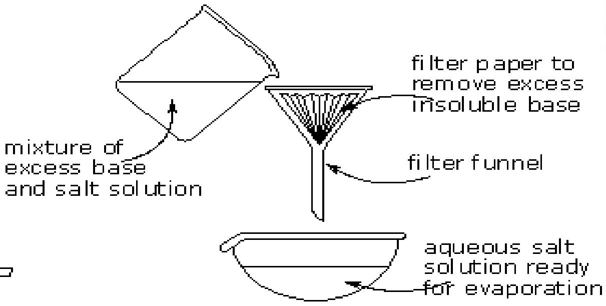
ACID + CARBONATE → SALT + WATER + CARBON DIOXIDE

When an acid reacts with a carbonate, we would observe fizzing (due to carbon dioxide being released). This is also an example of neutralisation.

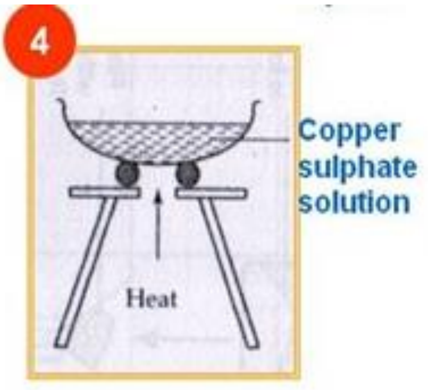
Experimental method of making a salt (acid + metal/base/carbonate)


warmed mixture of excess base and acid

Stage 1


mixture of excess base and salt solution
filter paper to remove excess insoluble base
filter funnel
aqueous salt solution ready for evaporation

Stage 2


4
Copper sulphate solution
Heat

Stage 1

The base/metal/carbonate is added in excess to ensure that ALL of the acid has reacted.
With a metal/carbonate, we would see no more fizzing when all of the acid has reacted.
With a base, the powder will stop dissolving.

Stage 2

Filtration is used to remove the excess metal/base/carbonate from the solution.

Stage 3

The solution is heated to ensure that **SOME** of the water has evaporated. It is left to cool, allowing crystals of the salt to appear.

Naming the salt

To name the salt, we would need to know two parts:

- The metal used in the reaction.
- The type of acid used in the reaction.

Name of acid	Formula of acid	Name of salt
Hydrochloric	HCl	Metal **CHLORIDE**
Sulphuric	H_2SO_4	Metal **SULPHATE**
Nitric	HNO_3	Metal **NITRATE**

Example

Copper Oxide + Sulphuric Acid → Copper Sulphate + Water

$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

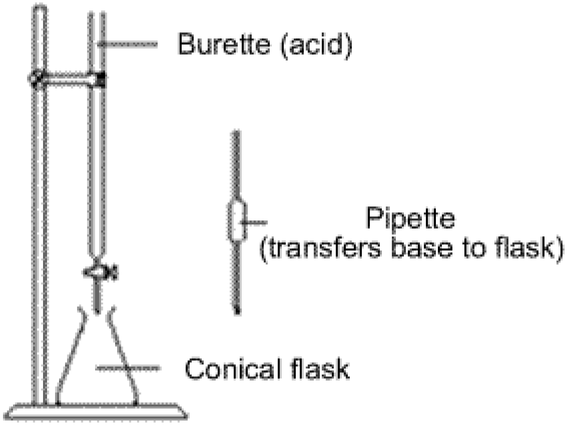
Testing Gases

Gas	Test	Observation
Hydrogen	Lighted splint	Goes pop
Carbon Dioxide	Limewater	Goes cloudy/milky

Testing Ions

Ion	Test	Observation	Product
Carbonate (CO_3^{2-})	Add acid	Fizzing, limewater goes cloudy	Carbon Dioxide (CO_2)
Sulfate (SO_4^{2-})	Add Barium Chloride solution	White precipitate	Barium Sulfate ($BaSO_4$)
Chloride (Cl^-)	Add Silver Nitrate solution	White precipitate	Silver Chloride ($AgCl$)
Bromide (Br^-)	Add Silver Nitrate solution	Cream precipitate	Silver Bromide ($AgBr$)
Iodide (I^-)	Add Silver Nitrate solution	Yellow precipitate	Silver Iodide (AgI)
Lithium (Li^+)	Flame test	Red flame	
Sodium (Na^+)	Flame test	Yellow-orange flame	
Potassium (K^+)	Flame test	Lilac flame	
Calcium (Ca^{2+})	Flame test	Brick-Red flame	
Barium (Ba^{2+})	Flame test	Apple Green flame	

Titrations



Titrations are used when an acid reacts with an alkali to produce a salt.
A pipette is used to measure an exact volume e.g. 25 cm^3 of an alkaline reactant.
The second reactant (acid) is poured into a burette through a funnel.

The alkali is added to a conical flask and 2-3 drops of an indicator is added e.g. phenolphthalein.

The acid is added to the flask, usually 1 cm^3 at a time, whist the flask is swirled to ensure that all of the alkali reacts.

When the colour of the indicator begins to get lighter/change, the acid is added dropwise, swirled until the indicator disappears/shows neutral. This is known as the end-point.

Record the volume of acid added from the burette.

Repeat the experiment without adding indicator to the alkali and add the measured volume of acid from the burette.

Pour the neutral solution to an evaporating dish and evaporate some of the water.

Usually, titrations are carried out 2-3 times to attain repeatable results.

Titrations are also used to determine the concentration of an unknown reactant.
To calculate the concentration, we would use the following calculation:

Concentration of acid = (Volume of alkali x Concentration of alkali) ÷ Volume of acid
Or
Concentration of alkali = (Volume of acid x Concentration of acid) ÷ Volume of alkali

Example
During a titration experiment, 25.0 cm^3 of sodium hydroxide reacted with 20.0 cm^3 hydrochloric acid, whose concentration was 0.50 mol dm^{-3}. Calculate the concentration of the alkali.

Concentration of alkali = (Volume of acid x Concentration of acid) ÷ Volume of alkali

Concentration of alkali = (20.0 x 0.5) ÷ 25.0
Concentration of alkali = 0.40 mol dm^{-3}

The higher the volume of a reactant, the less concentrated the reactant would be.

Metals

Properties of metals

Metals
Located on the left hand side of the table
Conduct heat
Conduct electricity
Malleable (Hammered into shape)
High melting point (except Mercury)
High density
Shiny (Lustre)

Displacement/Competition reactions

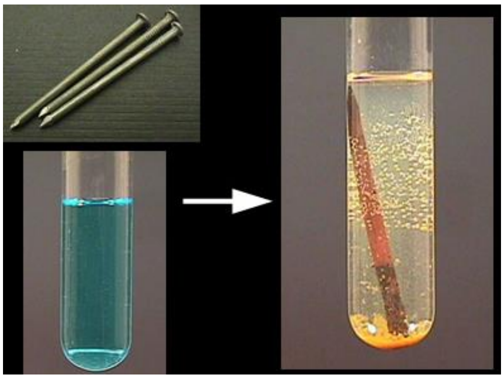
Displacement reactions occur when a more reactive metal takes the substance away from a compound containing a less reactive metal.

As a result, we will observe:

- A change in temperature
- Colour changes in the solution and around the metal

Example

Copper Chloride and iron



Copper Chloride will change colour from blue to orange (iron (III) chloride).

As copper is formed, a brown solid will be formed around the silver iron metal.

Iron + Copper Chloride → Iron (III) Chloride + Copper

$$2Fe_{(s)} + 3CuCl_{2\,(aq)} \rightarrow 2FeCl_{3\,(aq)} + 3Cu$$

Displacement reactions between metals are examples of oxidation and reduction, in terms of electrons (OIL RIG).

Oxidation **I**s **L**oss of Electrons (More positive)

Reduction **I**s **G**ain of Electrons (Less positive)

If we focus on the metals and their ions only, the chloride ion does not change so we call them 'Spectator Ions'.

$$Fe + Cu^{2+} \rightarrow Fe^{3+} + Cu$$

The iron has gone from 0 (no charge) to 3+, it has **LOST electrons** so it is **OXIDATION**.

The copper has gone from 2+ to 0, it has **GAINED electrons**, so it is **REDUCTION**.

Metals Extraction

There are very few metals found naturally e.g. Gold. Most metals are found in rocks, where the metal has combined with other elements. These rocks are called ORES.

To extract the metal from its ore, we need to consider the reactivity of the metal.

If the metal is **MORE** reactive than CARBON, the process of ELECTROLYSIS is used.
If the metal is **LESS** reactive than CARBON, the process of REDUCTION is used.

Reduction

This is the process used when the metal in the ore is LESS reactive than carbon. As a result, the carbon takes the oxygen away from the metal ore.

E.g. Copper Oxide + Carbon → Copper + Carbon Monoxide.

$$CuO + C \rightarrow Cu + CO$$

The copper oxide has been **REDUCED (loses oxygen),** whilst the carbon has been **OXIDISED (gained oxygen)**.

The most common example used in industry is the extraction of iron (Blast Furnace).

Blast Furnace

Four raw materials

- Iron (III) Oxide – Provides the iron metal.
- Coke (Carbon) – Acts as a fuel; produces carbon monoxide (reducing agent).
- Limestone (Calcium Carbonate) – Removes the impurities from the ore.
- Hot air (Oxygen) – To heat up the furnace by burning the coke.

Stages of reaction

1. Coke reacts with the oxygen to produce carbon dioxide.

   Carbon + Oxygen → Carbon Dioxide

   $$C + O_2 \rightarrow CO_2$$

2. More coke is added to the furnace, reducing the carbon dioxide into carbon monoxide.

   Carbon + Carbon Dioxide → Carbon Monoxide

   $$C + CO_2 \rightarrow 2CO$$

3. Carbon monoxide takes the oxygen (reduces) from the iron oxide to form molten iron and carbon dioxide.

   Iron (III) Oxide + Carbon Monoxide → Iron + Carbon Dioxide

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

4. The limestone undergoes thermal decomposition. It breaks down into calcium oxide (quicklime) and carbon dioxide.

   Calcium Carbonate → Calcium Oxide + Carbon Dioxide

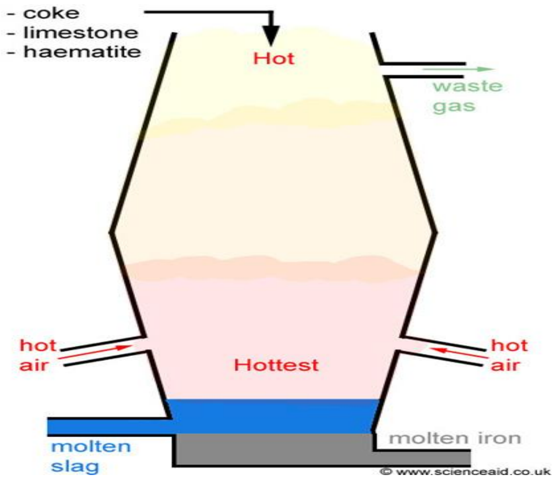
   $$CaCO_3 \rightarrow CaO + CO_2$$

   The calcium oxide (an alkali) neutralises the silicon dioxide (an acid) to produce calcium silicate, which is a layer of impurities called SLAG.

   Calcium Oxide + Silicon Dioxide → Calcium Silicate (Slag)

   $$CaO + SiO_2 \rightarrow CaSiO_3$$

   This is found on top of the molten iron. It is used for road surfacing.



Electrolysis

This is the process used to extract metals, from its ore, that are **MORE** reactive than carbon.

Electrolysis – Breaking down compounds using electric current.

The most common example of electrolysis is the extraction of aluminium metal.

Key terms

Electrolyte – The liquid that contains free moving ions.
Electrode – A rod that allows electricity to flow.
Cathode – The negative electrode.
Anode – The positive electrode.

The aluminium ore (alumina, aluminium oxide) is dissolved in cryolite.
Cryolite allow the ions to flow and to lower the temperature of the reaction.

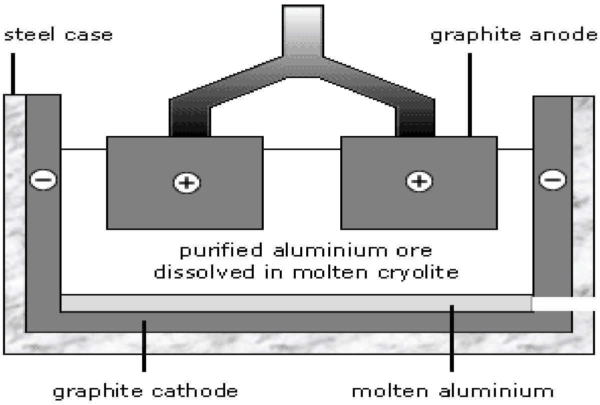
Stages

1. The positive aluminium ions (**Al^{3+}**) are attracted to the negative cathode.
   The ions gain 3 electrons to form molten aluminium metal (Al). This is **REDUCTION** (gaining electrons).
   Equation:

   $$\mathbf{Al^{3+} + 3e^{-} \rightarrow Al}$$

2. The negative oxide ions (**O^{2-}**) are attracted to the positive anode.
   The ions lose electrons to form oxygen gas. This is **OXIDATION** (losing electrons).
   We require two oxide ions to form one molecule of oxygen gas, so 4 electrons are lost in total.
   Equation:

   $$\mathbf{2O^{2-} \rightarrow O_2 + 4e^{-}}$$



Requirements for an aluminium extraction plant

- Near a power station – Uses a LARGE supply of electric current.
- Near the coast – Import the aluminium ore/To generate electricity by hydro electric
- Near a motorway/railway line – To transport the aluminium metal.
- Near a town/city – To employ workers.
- Away from the centre of town – To avoid possible explosions, air/noise pollution.

Recycling

This process is very important to sustain the quantity of iron and aluminium.

- It saves energy (Only need to melt the metal).
- It saves money.
- Conserves the ore (Prevents it from running out).
- Less carbon dioxide emissions (reduces global warming).

Uses of metals

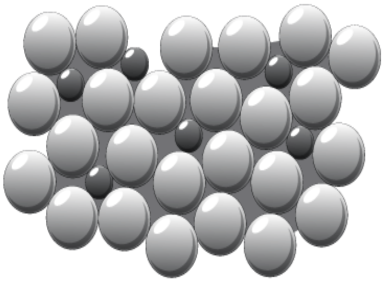
Aluminium (Al)	Copper (Cu)	Titanium (Ti)
Aeroplanes (Low density)	Electric wiring (Conducts electricity, ductile)	Aeroplanes (Low density)
Window frames, cans (does not corrode)	Jewellery (Colourful, malleable)	Hip replacement parts (Strong, does not corrode)
Saucepans (Conducts heat)	Saucepans (Conducts heat)	
Overhead cables (Conduct electricity, low density)	Water pipes (Does not corrode)	

Alloys

An alloy is a mixture of metals. An alloy will have different properties to the pure metal.

Examples:

Iron is a soft metal. By adding carbon to the metal, an alloy of steel is formed. Steel is stronger than iron.
This is due to the different sized atoms, distorting the structure, making it difficult for the atoms to slide over.
Stainless steel is an alloy of steel and chromium. Chromium prevents the steel from corrosion (rusting).



Calculations

Relative molecular mass (M_r) – The sum of the A_r of the atoms in the numbers shown in its formula.
Example – Carbon Dioxide (CO_2) where A_r C = 12, O = 16. So the M_r of CO_2 = 12 + 16 + 16 = 44

% composition of an element in a compound

Calculate the % of iron present in iron (III) oxide, Fe_2O_3. A_r Fe = 56, O = 16
Step 1 – Calculate the total mass of the element present in the compound.
Total mass of iron = 2 x 56 = 112
Step 2 – Calculate the M_r of the compound.
M_r Fe_2O_3 = (2 x 56) + (3 x 16) = 112 + 48 = 160
Step 3 – Use the equation below to calculate the % mass of the element:

% composition of element = (Total mass of element ÷ M_r of compound) x 100

% composition Fe = (112 ÷ 160) x 100
% composition Fe = 70.0%

Percentage yield

Percentage yield = (Mass of product obtained / Expected mass) x 100

E.g. 12g of magnesium combusted in excess oxygen gas to produce 18g of magnesium oxide.
Calculate the yield of the magnesium oxide.

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

A_r Mg = 24, O = 16

Step 1 – Cross out any unnecessary substances that are not in the question (Cross out the oxygen as it is excess).

Step 2 – Calculate the masses of the substances using their atomic masses.
2 Mg = 2 x 24 = 48
2 MgO = (24 + 16) x 2 = 40 x 2 = 80

Step 3 – Calculate the ratio between the substances to obtain the expected mass:
48g into 12 g of Mg (Divide by 4) so the expected mass of magnesium oxide will be 80 ÷ 4 = 20g of magnesium oxide.

Step 4 – Use the expected mass (calculated from step 3) and the mass obtained (from the question) to calculate the yield.

Mass obtained of magnesium oxide = 18g, Expected mass of magnesium oxide = 20g.

Percentage yield of magnesium oxide = (18 / 20) x 100 = 90%

HIGHER TIER ONLY

Formula of a compound using masses/percentages.

E.g. 7.2g of copper oxide contains 6.4g of copper; calculate the formula of copper oxide. A_r Cu = 64, O = 16.

Step 1 – Calculate mass of other element in the formula
Mass of oxygen = 7.2 – 6.4 = 0.8g

	Cu	O
Step 2 – Divide masses by the A_r of the element	6.4 / 64 = 0.10	0.8 / 16 = 0.05
Step 3 – Divide by the smallest answer to obtain number of atoms.	0.10 / 0.05 = 2	0.05 / 0.05 = 1

Formula of copper oxide = Cu_2O
The same methods can be used if percentages of elements are given, rather than masses.

Masses of products & reactants from an equation

E.g. Calculate the mass of iron produced when 480 tonnes of iron (III) oxide is reduced by carbon monoxide.

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

A_r Fe = 56, C = 12, O = 16

Step 1 – Cross out the substances not required in the reaction (3CO & $3CO_2$).

Step 2 – Calculate the M_r of the remaining substances.
Fe_2O_3 = (56 x 2) + (16 x 3) = 160
2Fe = (2 x 56) = 112
160 tonnes of iron (III) oxide produces 112 tonnes of iron.

Step 3 – Calculate the ratio between the masses from step 2 with the mass in the question.
160 → 480 tonnes of iron (III) oxide (x3), so mass of iron formed = 112 x 3 = 336 tonnes.

Moles

The number of moles of a substance can be calculated by the following equation:

Moles = Mass of substance ÷ M_r of substance

How many moles are present in 11g of carbon dioxide, CO_2? A_r C = 12, O = 16.
M_r CO_2 = 12 + (2 x 16) = 44
Moles CO_2 = 11 ÷ 44
Moles CO_2 = 0.25

Converting moles into mass

To convert moles into mass, we re-arrange the moles equation to become:

Mass = Moles x M_r

How much mass does 0.25 moles of CO_2 weigh?
Mass = 0.25 x 44
Mass = 11g

To work out how many atoms/molecules of a substance is present, we multiply the number of moles with Avogadro's constant (6×10^{23}).

Number of atoms/molecules = Moles x Avogadro's constant

How many molecules are present in 11g of carbon dioxide?
We worked out above that there are 0.25 moles of carbon dioxide in 11g.

Number of molecules = $0.25 \times 6 \times 10^{23}$
Number of molecules = 1.5×10^{23}

Titration Calculations

To carry out titration calculations, we would use the following equation:

Moles = Volume x Concentration

Volume is usually measured in cm^3, for the equation we need dm^3. We divide the volume by 1000.
25.0 cm^3 = 0.025 dm^3 (25 ÷ 1000)
Concentration is measured in mol dm^{-3}.

During a titration experiment, 25.0 cm^3 of sodium hydroxide reacted with hydrochloric acid, whose concentration was 0.50 mol dm^{-3}.

$$NaOH + HCl \rightarrow NaCl + H_2O$$

The titrations for the volume of hydrochloric acid used to neutralise the sodium hydroxide are shown below:

	Titre 1	**Titre 2**	**Titre 3**
Volume of acid used (cm^3)	26.55	26.70	26.65

Calculate the concentration of the sodium hydroxide.

Step 1 – Calculate the mean volume of the titrations (Use those that are within 0.25cm^3 of each other)
Volume of hydrochloric acid = (26.55 + 26.70 + 26.65) ÷ 3
Mean volume = 26.63 cm^3

Step 2 – Calculate the moles of acid used in the titration.
Moles acid = Volume of acid x Concentration of acid
Moles acid = (26.63 ÷ 1000) x 0.50
Moles acid = 0.0133

Step 3 – Using the symbol equation for the reaction, look for the moles ratio between the two substances (Moles = Number before the formula)
Moles ratio NaOH : HCl is 1:1 The moles of the acid = Moles of NaOH
Moles NaOH = 0.0133

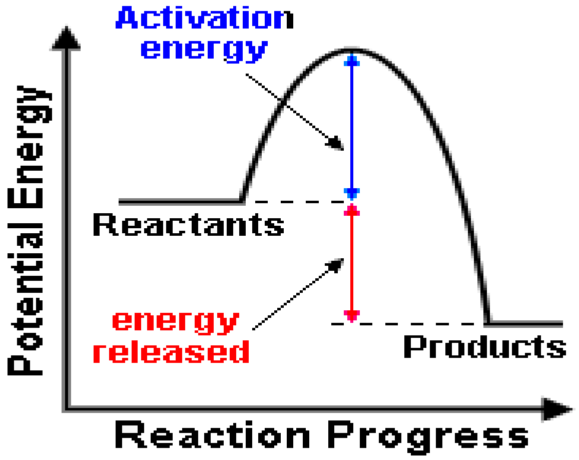
Step 4 – Re-arrange the moles equation to calculate the concentration of the alkali.
Concentration NaOH = Moles NaOH ÷ Volume NaOH
Concentration NaOH = 0.0133 ÷ (25.0 ÷ 1000)
Concentration NaOH = 0.53 mol dm^{-3}

As the volume of sodium hydroxide is less than the hydrochloric acid, the concentration of sodium hydroxide would be greater than the acid.

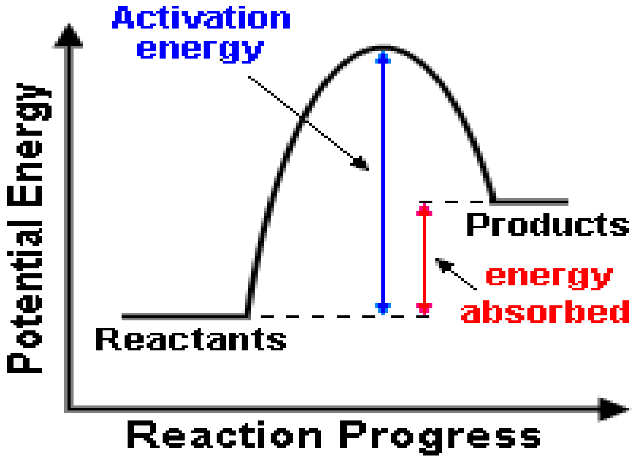
Energy Changes

During chemical reactions, the temperature of the reaction may change.
If the temperature INCREASES, energy is released/exits to the surroundings. This is an EXOTHERMIC reaction.
If the temperature DECREASES, energy is absorbed from the surroundings. This is an ENDOTHERMIC reaction.

Energy profile diagrams


Activation energy
Reactants
energy released
Products
Potential Energy
Reaction Progress

Exothermic reaction

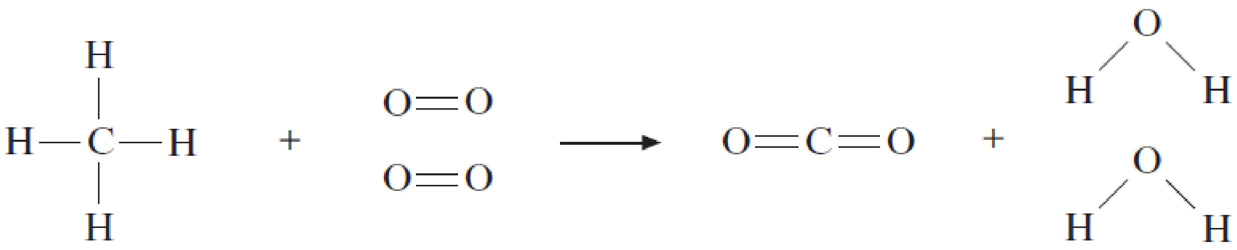

Activation energy
Products
energy absorbed
Reactants
Potential Energy
Reaction Progress

Endothermic reaction

Activation energy – The **minimum** amount of energy required for particle to collide successfully.
Energy released & energy absorbed = Overall energy change

Bond energies

Energy is required to break the bond within a molecule and to form new bonds. The value of the energy is identical.
We can use the bond energies to determine whether the reaction will be exothermic or endothermic.
E.g.


H
H—C—H
H
+
O=O
O=O
O=C=O
+
O
H H
O
H H

Bond	Energy involved (kJ mol^{-1})
C-H	413
O=O	498
C=O	805
O-H	464

Step 1 – Calculate the energy taken in to break all of the reactants (left hand side) bonds.
4 C-H bonds and 2 O=O bonds Energy = (4 x 413) + (2 x 498) = 1652 + 996 = 2648 kJ mol^{-1}

Step 2 – Calculate the energy released to form all of the products (right hand side) bonds.
2 C=O bonds and 4 O-H bonds Energy = (2 x 805) + (4 x 464) = 1610 + 1856 = 3466 kJ mol^{-1}

Step 3 – Energy of the reaction = Energy of the reactants – Energy of the products
Energy of the reaction = 2648 – 3466 = -818 kJ mol^{-1}

If the energy value of the reaction is **NEGATIVE**, the reaction is **EXOTHERMIC** (Most likely answer!!!)
If the energy value of the reaction is POSITIVE, the reaction is ENDOTHERMIC.

Oil

Formation of crude oil

- Decayed sea organisms
- Died millions of years ago

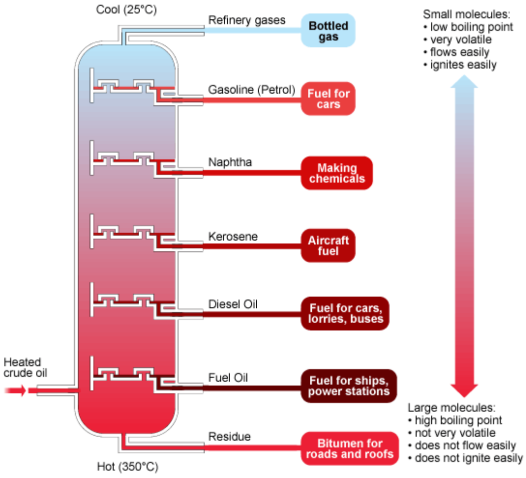
Crude oil is a mixture of hydrocarbons. A hydrocarbon is a compound containing **HYDROGEN** and **CARBON** atoms only.
We can obtain many useful products from crude oil such as petrol, diesel and bitumen.
A process called Fractional Distillation is used to get these products from crude oil.

Fractional distillation of crude oil

- Crude oil is a mixture of hydrocarbons with various sizes and boiling points.
- The crude oil is evaporated in a furnace. The vapour enters the tower and condenses into different fractions.

Fractions –Mixtures containing hydrocarbons of similar chain lengths, resulting in similar boiling points.

- As we descend the tower, the boiling point of the fractions increases and the chain lengths get bigger.


Cool (25°C)
Refinery gases
Bottled gas
Gasoline (Petrol)
Fuel for cars
Naphtha
Making chemicals
Kerosene
Aircraft fuel
Diesel Oil
Fuel for cars, lorries, buses
Heated crude oil
Fuel Oil
Fuel for ships, power stations
Residue
Bitumen for roads and roofs
Hot (350°C)
Small molecules:
• low boiling point
• very volatile
• flows easily
• ignites easily
Large molecules:
• high boiling point
• not very volatile
• does not flow easily
• does not ignite easily

Examples of fractions

Fractions	Colour	Viscosity (Ease to evaporate)	Easy to ignite	Cleanliness of burn	Use
Petroleum Gases	Colourless	Very viscous	Very easy	Clean flame	Home heating fuel
Petrol	Colourless	Very viscous	Very easy	Clean flame	Fuel for cars
Naphtha	Pale yellow	Viscous	Easy	Clean	Making chemicals
Kerosene	Pale yellow	Viscous	Easy	Clean	Fuel for aeroplanes
Diesel	Yellow	Not viscous	Difficult	Smokey	Fuel for lorries
Lubricating Oil	Orange	Not viscous	Difficult	Very smoky	Lubricant to reduce friction
Fuel Oil	Orange-brown	Does not evaporate	Very difficult	Lots of smoke	Fuel for ships
Bitumen	Black	Will not evaporate	Will not ignite	Extremely smoky	Road surfacing

As we descend the tower:

- The colour darkens
- Viscosity decreases
- Gets harder to ignite
- Cleanliness of the flame decreases

Problems with crude oil

- When combusted (burnt), it contributes to global warming as it releases carbon dioxide.
- Some fractions may contain sulphur, which will contribute to acid rain when burnt.
- Crude oil is non-renewable; no more oil will be produced.
- Great demand for crude oil increases the cost of living.

Cracking

- Breaking down large hydrocarbon chains into small, more useful ones.
- For cracking to occur, a lot of heat is required and a catalyst.
- We would produce a smaller alkane (used for petrol to meet demand) and a monomer (small, reactive molecule used to make plastics).

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

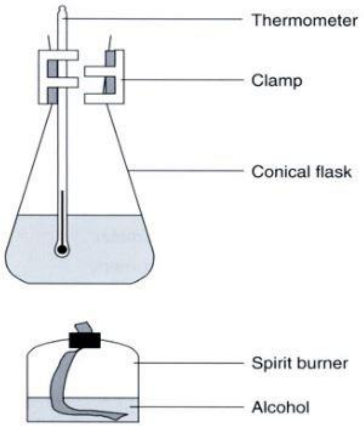
$$C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$$

Combustion of fuels

When fuels (mostly hydrocarbon fuels) combust in a plentiful supply of oxygen, carbon dioxide and water are produced.

Fuel + Oxygen → Carbon Dioxide + Water

To measure the energy released during a combustion reaction, we would use the method of calorimetry.



Independent Variable	Dependent Variable	Control Variables
• The mass of fuel burnt	• The energy released from the fuel / The temperature of the water	• The volume of water • The height between the wick and the flask

To calculate the energy released by the fuel, we would use the following equation:

Energy released = (Volume/mass of water x 4.18 x Change in temperature) ÷ Mass of fuel burnt

The units for energy released will be in Joules per gram (J/g).
Usually, energy is measured in kJ/mol so:
To convert J into kJ, divide by 1000.
To convert g into moles, divide the mass of fuel burnt by its M_r (molecular mass).

The value attained will be **LOWER** than the actual value because:

- Not all the heat from the fuel goes directly to the water.
- Heat is lost to the surroundings / Draughts may be present.

Fire Triangle

To start a fire, we would require **HEAT**, **FUEL** and **OXYGEN**. This can be shown by the Fire Triangle.

OXYGEN
HEAT
FUEL

To extinguish a fire, we would remove one of these components:

Remove Heat	Remove Oxygen	Remove Fuel
• Water (except electrical & chip pan fires) • Carbon dioxide • Roll on the floor	• Foam • Fire blanket • Carbon Dioxide • Heat proof mat • Beat using fire shovels.	• Fire breaks (Forest fires)

Hydrogen as a fuel

Advantages	Disadvantages
• Releases a lot of energy when combusted. • Only water is produced as a by-product. • Does not contribute to global warming/acid rain. • A renewable source (obtained from water).	• Highly explosive • Difficult to store, high pressure containers. • Requires a lot of energy to produce (Electrolysis) • Uses expensive catalysts (Platinum)

Organic Chemistry

Hydrocarbons – A compound containing hydrogen and carbon atoms **ONLY**. Examples are alkanes and alkenes.

Alkanes – A hydrocarbon containing **SINGLE COVALENT BONDS** between the carbon atoms.

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

General formula of alkanes = C_nH_{2n+2} (n = number of carbon atoms)
Alkanes are **SATURATED** (Only contains single bonds) and burn easily (used in fuels).

Alkenes – A hydrocarbon that contains a **DOUBLE BOND** between the carbon atoms

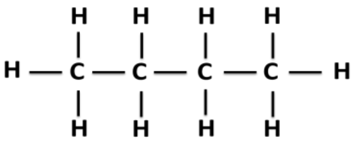
Name of alkene	Formula	Structure
Ethene	C_2H_4	
Propene	C_3H_6	

General formula of alkenes = C_nH_{2n}
Alkenes are **UNSATURATED** (Contains a double bond) and are reactive (due to the double bond being able to break).

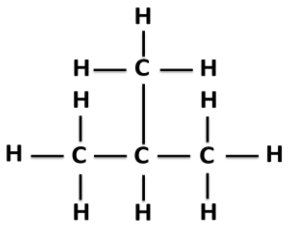
TEST FOR ALKENES – Add Bromine/Bromine Water (Br_2)
Observation – Goes from ORANGE/BROWN to COLOURLESS

Isomers – HIGHER TIER ONLY

Isomers are molecules that have the same molecular formula but the atoms are arranged differently.
Butane (C_4H_{10}) has two isomers:



butane (C_4H_{10})

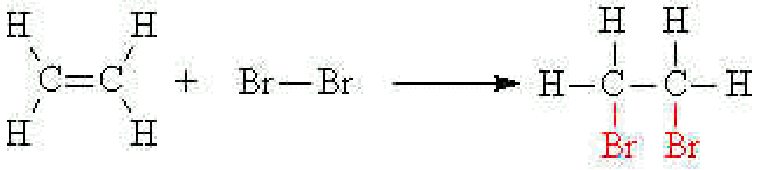


2-methylpropane (C_4H_{10})

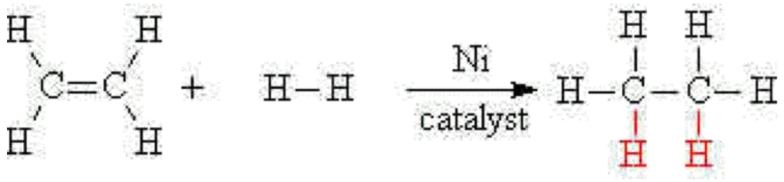
Addition reactions of alkenes

Reaction with bromine (Br_2)

The double bond breaks, allowing one bromine atom to join with each carbon atom of the double bond.



Reaction with hydrogen (H_2) - needs a **nickel catalyst** at 170°C for the reaction.
When hydrogen is added to an alkene, the double bond breaks to form an ALKANE.
Example:



This reaction is also known as **CATALYTIC HYDROGENATION**.

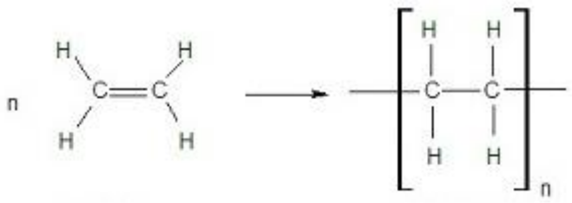
Addition Polymerisation - This involves one type of monomer joining to form a polymer.

How does polymerisation work?

The double bond of the monomer breaks, resulting in spare bonds.

Large number of monomers join onto each other, due to the spare bonds, to form a long chain (the polymer).

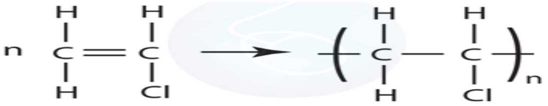
Example: Ethene → Polyethene (Polythene)



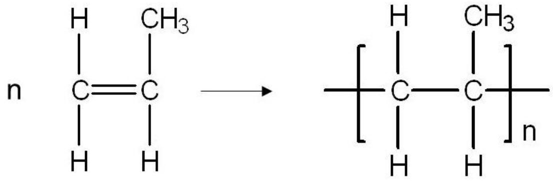
n = large number of monomers (thousands)

Other examples:

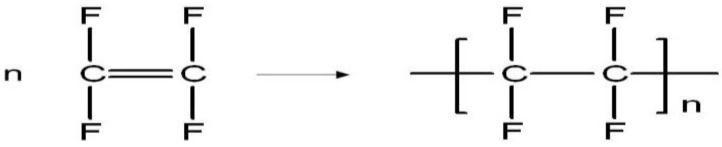
Vinyl Chloride → Poly Vinyl Chloride (PVC)



Propene → Polypropene



Tetra Fluoro Ethene → Poly Tetra Fluoro Ethene (PTFE, Teflon)



Uses and properties of polymers

Polymer	Uses	Properties
Polyethene	Shopping bags	Strong, Waterproof
Polystyrene	Drinking cups	Light weight, insulator
Polyvinyl chloride (PVC)	Window frames	Weather resistant, low density
Polytetra fluoro ethane (PTFE)	Coating frying pans	Non-stick

Comparison to traditional materials

Plastics have replaced a large number of traditional materials e.g. wood, metal and glass for making objects.

Object	Traditional material	Why plastics are better?
Shopping bags	Paper	Stronger, waterproof
Bottles	Glass	Does not break, unreactive
Guttering	Metal (Steel/Iron)	Does not corrode, light weight

Advantages & disadvantages of plastics

Advantages	Disadvantages
Easy to shape	Fill up landfill sites
Light weight	Non-biodegradable
Colourful	Melt easily, releasing toxic gases
Strong	Made from crude oil (non-renewable source)

Recycling plastics

- Conserves the raw material (oil).
- Requires less energy to produce new plastics.
- Reduces the landfill sites.