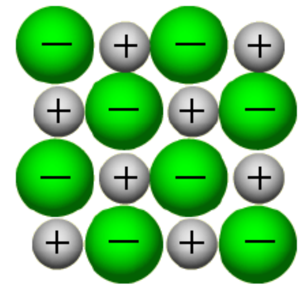
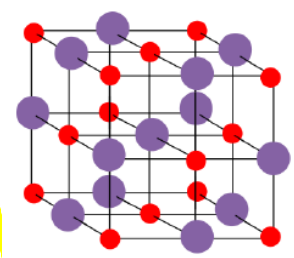
Year 11 Chemistry Revision

Topic 1: Bonding

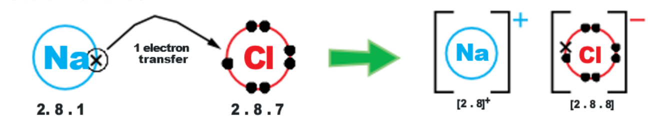




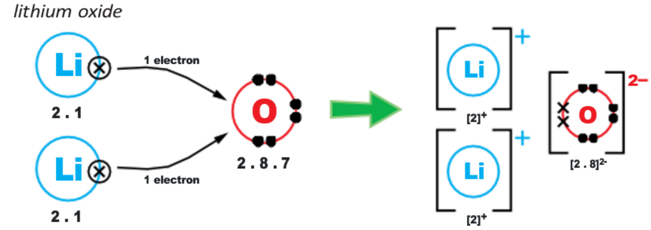
Ionic Bonding:

- Ion – a charged particle. Has different numbers of protons and electrons.
  - » Positive ion – has more protons (+) than electrons (–) – CATIONS
  - » Negative ion – has more electrons (–) than protons (+) – ANIONS
- Ionic bonding – the bonding between metal and non-metal atoms.
- Ionic bonds form when electrons transfer from a metal to a non- metal atom so that both atoms achieve full outer shells.

sodium chloride

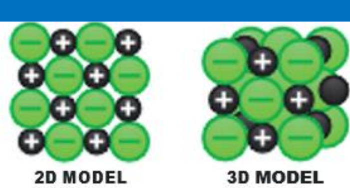


lithium oxide



Giant Ionic Structures

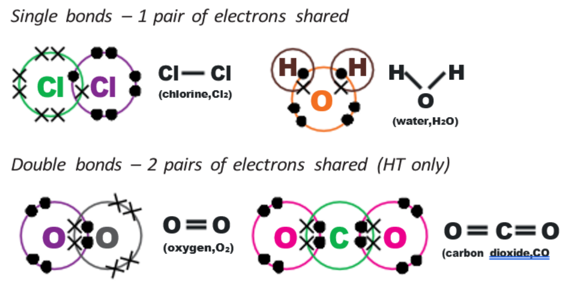
- Ionic compounds have giant ionic structures.
- Ionic lattice – the regular arrangement of the ions in ionic structures.
- *The oppositely charged ions attract each other in a regular pattern.*



High melting and boiling points – due to the strength of the bonds between **ions.**

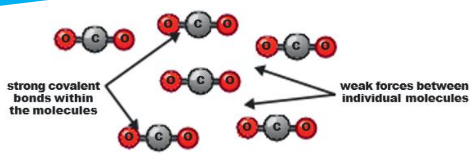
Conduct electricity when dissolved or molten – only then are the ions free to move to carry the charge.

Higher tier:
NaCl has a lower melting point than MgO even though they are both ionic compounds. This is due to the charge being higher on Mg^{2+} and O^{2-} ions (compared to Na^{+} and Cl^{-}). Higher charge = stronger bond = more energy needed to break.



Covalent Bonding:

- Covalent bonding – the bonding between non-metal atoms.
- Covalent bonds form when the atoms share electrons so that both atoms achieve full outer shells.



Simple Molecular Structures

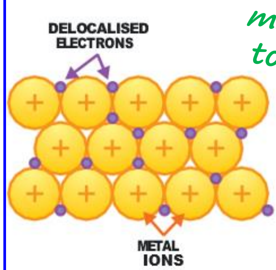
Simple molecules consist of a few atoms held together by covalent bonds. Hydrogen, water and carbon dioxide are examples of simple molecular structures.

Properties of simple molecular compounds

- Low melting and boiling points – due to the weak intermolecular forces between the molecules.
- Do not conduct electricity – no free electrons to carry the electric current

Metallic Bonding:

- Metallic bonding – when metal atoms bond together.
- Metals have giant structures of regularly arranged atoms.

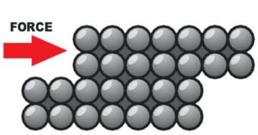


The electrons from the outer shells of the atoms are delocalised – meaning they are free to move through the whole structure.

- By sharing delocalised electrons – strong metallic bonds are formed.
- The strength of a metallic bond is due to the force of attraction between the metal ions (+) and the delocalised

Metallic Properties

- Conduct electricity – the delocalised electrons carry electrical charge through the structure.
- Conduct heat – the delocalised electrons and closely packed ions transfer energy through the structure by conduction.
- Malleable and ductile



The layers of metal ions are able to slide over each other when hammered or stretched

- High melting and boiling points – large amounts of energy are needed to break the strong metallic bonds in melting / boiling.

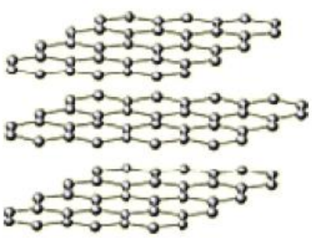
Higher Tier – the melting and boiling points increase as you move across any period of the Periodic table, because there are more delocalised electrons increasing the attraction between the ions and the free electrons (stronger bonds).

Giant Covalent Structures:

- Giant covalent structures consist of lots of atoms held together by covalent bonds
- They are arranged into giant lattices, which are extremely strong because of the large number of bonds in the structure.

Diamond

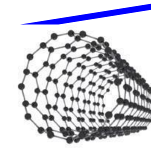
- Each carbon atom is bonded to 4 other carbon atoms – hard and high melting point.
  - Does not conduct electricity.
  - Used in drill bits, glass cutting, gemstones.



Graphite

- Each carbon atom is bonded to 3 other carbon atoms. The fourth electron becomes delocalized between the layers.
- Graphite has a layered structure with weak forces between the layers
  - Conducts electricity – delocalised electrons between layers carries charge.
  - Used in pencils and lubricants – layers can slide over each other.

Allotropes = different forms of the same element, with a different structure.



Nanotubes

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.

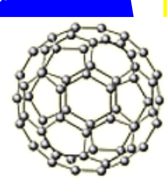
They are proposed to be used in small electronic circuits.

Fullerenes

These are cages and tubes with different number of carbon atoms.

Buckminsterfullerene is one type of fullerene. Its molecules are spherical and contain 60 carbon atoms.

Fullerenes may be used for *drug delivery systems* in the body, in *lubricants* and as *catalysts*.



Smart materials have properties that react to changes in their environment. This means that one of their properties can be changed by an external condition, such as temperature, light, pressure, pH or electricity. This change is reversible and can be repeated many times.

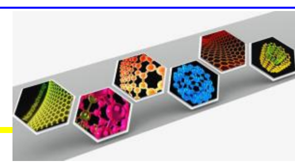
Smart materials are:
1. Reversible
2. React to environment

Shape-memory alloys / polymers
If shape-memory alloy is bent out of shape, when it is heated above a certain temperature it will return to its original shape.
This property makes it useful for making spectacle frames..

Thermochromic materials change colour as the **temperature** changes. These are used on contact thermometers.

Hydrogels are used to make soft contact lenses, nappies, wound dressings and drug delivery systems.
They are used because they can absorb/expel water and swell/shrink (up to 1000 times their volume) due to changes in pH, temperature, salt concentration, etc.

Photochromic materials change colour according to different **lighting** conditions or changing **light intensity**. They are used for security markers that can only be seen in ultraviolet light.

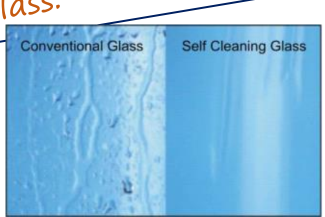


Nanoscience
Nanoparticles range in size from 1 nm to 100 nm and are far too small to see with a microscope. They have remarkable properties that are **different from the same substance in bulk**. They are already being used in consumer products but there are some **uncertainties about their safety long term.**

Nanoparticles have a very large surface area compared with their volume, so they are often able to react very quickly. This makes them useful as catalysts to speed up reactions.

Nano silver = antibacterial, antiviral and antifungal.
Uses include socks, plasters, fridge lining.

Nano titanium dioxide = absorbs and reflects UV light, transparent.
Uses include sun cream and self cleaning glass.

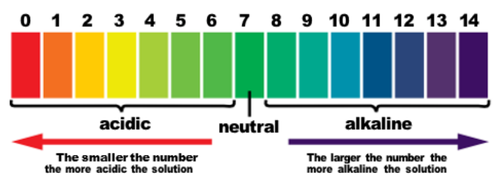


Nanomaterials currently used have been tested to ensure that they cause no damage to individuals or the environment, but that their **long-term effects** are as yet unknown. Some people have expressed concern that nano-scale silver (deodorants) and titanium dioxide (sun screens) are applied to the skin and can therefore be easily absorbed into the body. While it has been shown that these uses are safe in the short term, there is no certainty that exposure over many years will not result in problems.

Topic 2: Acids

Indicators and the pH Scale:

- Indicators are substances that change colour when they are added to acids and alkalis.
- Litmus is the most well-known indicator.
- It turns red in acid and blue in alkalis.
- Universal indicator is most commonly used in the laboratory. When added to a solution, it changes to a colour that shows the pH of the solution.



Acidic = pH < 7
Neutral = pH 7
Alkaline = pH > 7

Acids:

Acids produce hydrogen ions, H^+, when they dissolve in water, e.g. for hydrochloric acid:

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Whilst some acids and alkalis are dangerous, others, such as those in vinegar or lemon juice, can be sprinkled on our food! Those that are dangerous have the corrosive hazard warning symbol on their containers.

Bases:

A base is chemically opposite to an acid. A base that dissolves in water is called an alkali.

Alkalis:

Alkalis produce hydroxide ions, OH^-, when they dissolve in water, e.g. for sodium hydroxide:

$NaOH(aq) \rightarrow Na^+(aq) + OH^-$

Acids and Carbonates:

Acids will react with carbonates to make a salt, water and carbon dioxide gas.

acid + carbonate → salt + water + carbon dioxide

The carbon dioxide causes bubbling during the reaction. The reaction is exothermic.

Acids and Metals:

Acids will react with metals to make a salt and hydrogen gas. The hydrogen causes bubbling during the reaction. The reaction is exothermic.

acid + metal → salt + hydrogen

hydrochloric acid + zinc → zinc chloride + hydrogen

$2HCl + Zn \rightarrow ZnCl_2 + H_2$

The more reactive the metal, the faster the reaction is, resulting in more bubbling and a bigger temperature rise.

Acids and Alkalis/Bases:

Acids react with alkalis and bases to make a salt and water.

Alkalis are commonly metal hydroxides.

acid + alkali → salt + water

nitric acid + sodium hydroxide → sodium nitrate + water

$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$

Bases are commonly metal oxides.

acid + base → salt + water

sulfuric acid + zinc oxide → zinc sulfate + water

$H_2SO_4 + ZnO \rightarrow ZnSO_4 + H_2O$

The reactions of acids with alkalis and bases are exothermic.

Neutralisation:

A neutralisation reaction happens when an acid and an alkali 'cancel each other out'. The reaction always produces a salt and water.

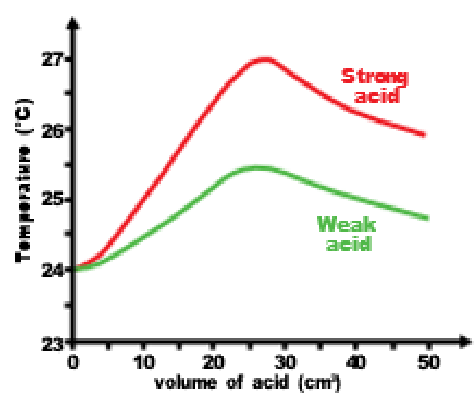
acid + alkali → salt + water

The H+ ions from the acid react with the OH− ions from the alkali to form water. (HT only)

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Naming salts

1. The metal gives the first part of the name
2. The acid gives the second part of the name:
   - Hydrochloric = __ chloride
   - Nitric = ___ nitrate
   - Sulfuric = ___ sulfate



Strong and Weak Acids (TRIPLE only):

- Strong acids fully dissociate (ionise) in water whereas weak acids only partially dissociate.
- Strong acids have high numbers of H+ ions in their solutions, so low pH values.
- Strong acids react more quickly than weak acids.

Note – acid strength and concentration are not the same thing! (TRIPLE ONLY)

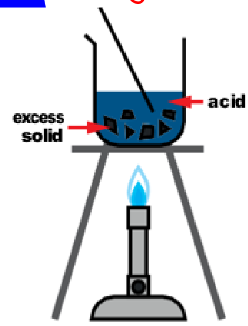
» The strength of an acid is a measure of the degree of its dissociation.

» The concentration of an acid is a measure of the number of moles of acid in 1 dm3 of solution.

Preparation of a salt from a metal or insoluble base/carbonate:

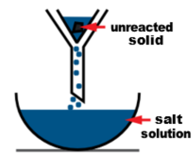
- Metals, bases and certain metal carbonates, are insoluble in water.
- There are three stages in the preparation.

Stage 1



- Excess metal/base/carbonate is added to the acid to make sure all the acid has reacted and been used up.
- Heating and stirring help the process.
- For metals and metal carbonates, the fizzing stops when all the acid has been used up.

Stage 2



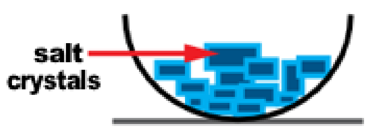
The mixture is filtered using a filter funnel and filter paper.

The excess solid remains in the filter paper. The salt solution passes through into the evaporating basin.

Stage 3

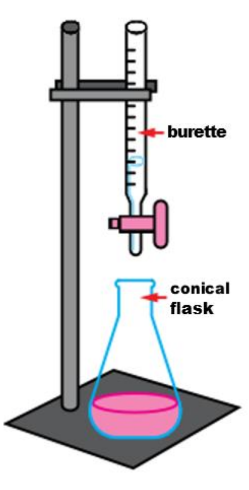
The salt crystals are collected from the solution by evaporation. The solution is heated to evaporate the water. The size of the crystals produced depends on the rate of evaporation.

- Large crystals – evaporate water slowly near a radiator or window ledge.
- Small crystals – use a Bunsen to evaporate 2/3 of the water quickly, before allowing to crystallise naturally.



Preparation of a salt from an alkali or soluble carbonate:

- An alkali is a soluble base.
- A titration is used to produce a salt from an acid and alkali.
- This is an outline method for carrying out a titration in which an acid is added to an alkali.



1. Measure exactly 25 cm^3 of alkali into a clean conical flask.
2. Add a few drops of indicator to the flask.
3. Place the flask onto a white tile.
4. Fill the burette with acid.
5. Slowly add the acid from the burette to the alkali until the indicator changes colour. This is the endpoint of the reaction.
6. Record the volume of acid added to the flask. Repeat steps 1-6 without using the indicator and adding the same volume of acid from the burette.

TRIPLE ONLY:

Preparing Insoluble Salts (TRIPLE only):

- An insoluble salt is made when solutions of two soluble salts react in a precipitation reaction.
- There are three stages in the method.

<u>Stage 1 – Mixing</u> The two soluble salt solutions are mixed.

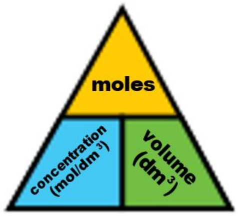
<u>Stage 2 – Filtration</u> The insoluble precipitate is separated from the mixture by filtration. The precipitate stays behind in the filter paper, while the solution passes through.

<u>Stage 3 – Washing and drying</u> Water cannot dissolve the precipitate – it is insoluble – but it can wash off any remaining impurities. The filter paper is then removed, opened out and the precipitate dried in an oven.

Titration Calculations (TRIPLE only):

- The concentration of a solution is the number of moles per dm^3.
- The following equation is used to calculate the concentration of a solution

NOTE: divide by 1000 to convert from cm^3 to dm^3



$$\text{Concentration } (mol/dm^3) = \frac{\text{number of moles}}{\text{volume } (dm^3)}$$

Remember you might also need to use last year's moles equation:

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

Moles calculations examples

1:1 Mole Ratio

15.0 cm³ of 0.1 mol/dm³ hydrochloric acid was needed to neutralise 25.0 cm³ of sodium hydroxide solution. What is the concentration of the sodium hydroxide solution?

$HCl + NaOH \rightarrow NaCl + H_2O$

Step 1: Convert all volumes to dm³

HCl 15.0 cm³ = 15.0 ÷ 1000 = 0.015 dm³

NaOH 25.0 cm³ = 25.0 ÷ 1000 = 0.025 dm³

Step 2: Calculate the number of moles of the substance where the volume and concentration are known

number of moles HCl = concentration × volume

number of moles of HCl = 0.1 × 0.015 = 0.0015 mol

Step 3: Calculate the unknown concentration

This is a 1:1 reaction, i.e. 1 part of HCl reacts with 1 part of NaOH. We say that 0.0015 mol of acid reacts with 0.0015 mol of alkali.

concentration of NaOH = number of moles ÷ volume

concentration of NaOH = moles ÷ volume = 0.0015 ÷ 0.025 = 0.06 mol/dm³

2:1 Mole Ratio

20.0 cm of 0.5 mol/dm³ sulfuric acid was needed to neutralise 25.0 cm³ of sodium hydroxide solution. What is the concentration of the sodium hydroxide solution?

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

Step 1: Convert all volumes to dm³

H_2SO_4 20.0 cm³ = 20.0 ÷ 1000 = 0.020 dm³

NaOH 30.0 cm³ = 30.0 ÷ 1000 = 0.030 dm³

Step 2: Calculate the number of moles of the substance where the volume and concentration are known

number of moles H_2SO_4 = concentration × volume

number of moles of H_2SO_4 = 0.5 × 0.020 = 0.01 mol

Step 3: Calculate the unknown concentration

This is a 2:1 reaction, i.e. 2 parts of NaOH react with 1 part of H_2SO_4. We can say that 0.01 mol of acid will react with 0.02 mol of alkali.

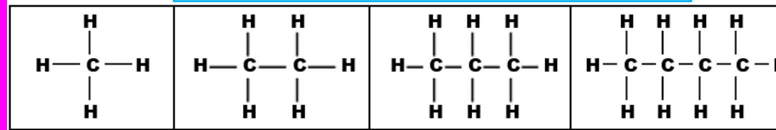
concentration of NaOH = number of moles ÷ volume

concentration of NaOH = moles ÷ volume = 0.02 ÷ 0.025 = 0.8 mol/dm³

Topic 3: Crude oil

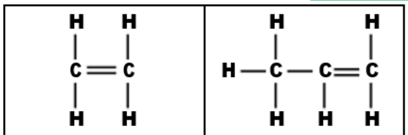
Alkanes:
Alkanes contain single bonds between the carbon atoms only and are said to be saturated. The general formula for alkanes is C_nH_{2n+2}. The names of alkanes end with '-ane'.

Alkanes are fairly unreactive, they combust well.



methane	ethane	propane	butane
CH_4	C_2H_6	C_3H_8	C_4H_{10}

The alkenes contain a double covalent bond between two carbon atoms and are said to be unsaturated. The general formula for alkenes is C_nH_{2n}. The names of alkenes end with '-ene'.

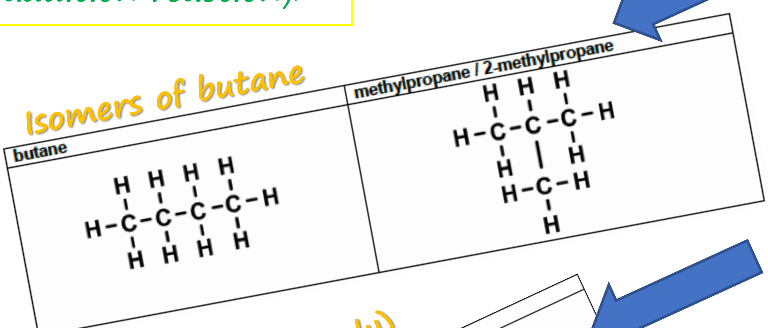


ethene	**propene**
C_2H_4	C_3H_6

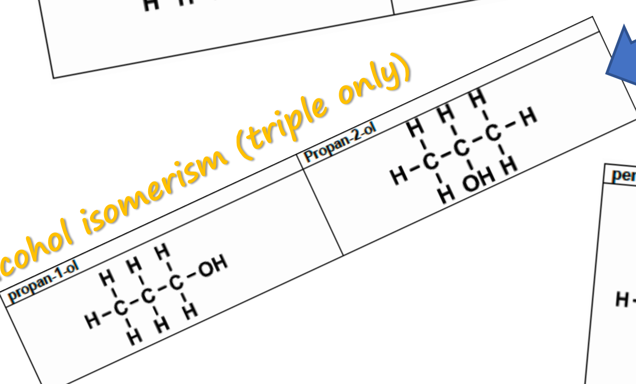
As a result of the **double bond** the alkenes are **very reactive molecules**, the double bond can be broken to form single bonds with other atoms (**addition reaction**).

Higher tier
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.

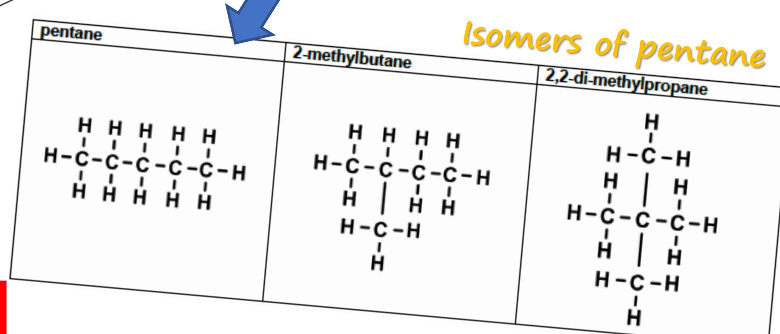
Isomers of butane



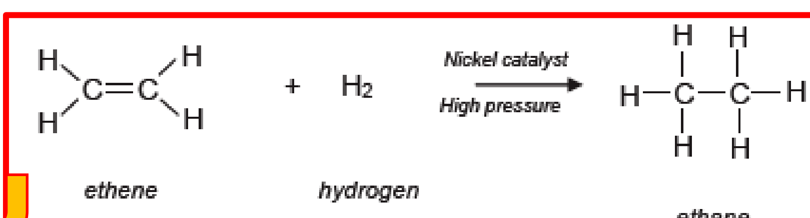
Alcohol isomerism (triple only)



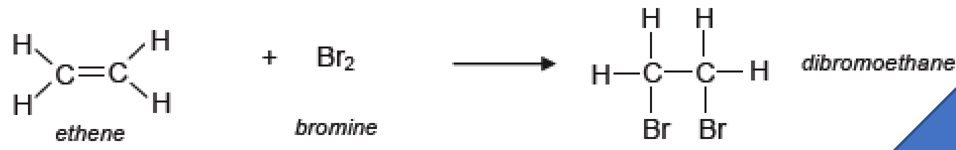
Isomers of pentane



Addition reactions
Two atoms can be added across the C=C bond in an unsaturated compound thus forming a saturated compound. One atom is added to each of the carbon atoms involved.



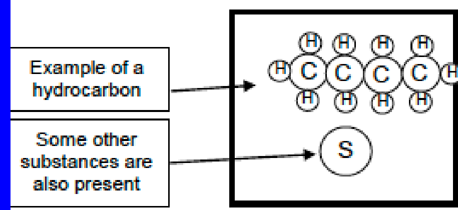
Unsaturated Alkene becomes a saturated Alkane



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

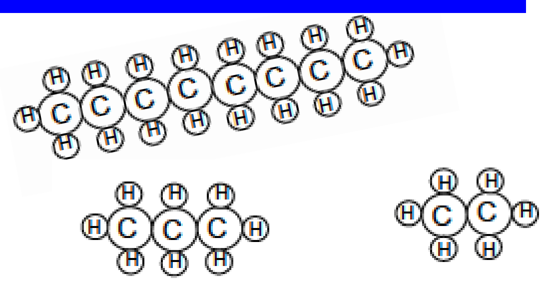
Hydrocarbons:

Hydrocarbons are compounds that contain carbon and hydrogen atoms only. Carbon atoms have the ability to form bonds with other carbon atoms resulting in the formation of hydrocarbon chains. These chains can vary in length and as a result, hydrocarbons have different boiling points. The longer the chain, the higher the boiling point.



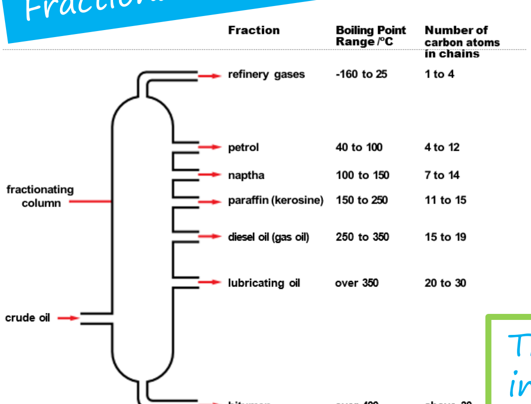
Crude oil:

Crude oil is a complex mixture of hydrocarbons, which can be separated by fractional distillation. It is produced from the remains of dead marine animals and plants that lived around 300 million years ago. When the remains sank to the bottom of the sea, they were covered by sand and other sediments. Layer upon layer of sediments built up over time and pressure and heat caused the remains to break down, forming crude oil.



Crude oil contains a mixture of different sized hydrocarbon chains

Fractional distillation



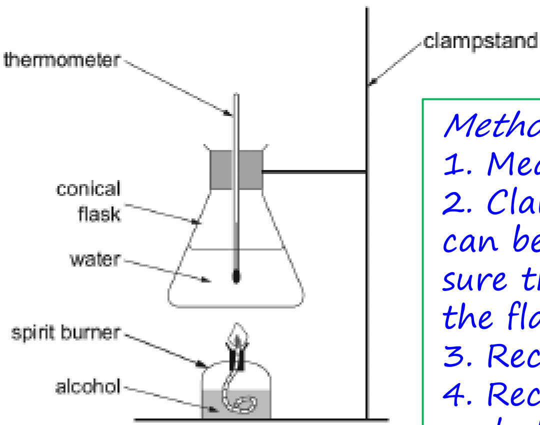
Crude oil is boiled/vaporised before it enters the fractionating column and the hydrocarbons condense at different heights in the column. The lower the boiling point, the higher in the column a compound is collected.
Fractions are mixtures containing hydrocarbon compounds that have similar chain lengths and, therefore, similar boiling points.

The molecules of hydrocarbons are held together by intermolecular forces. Larger molecules have more intermolecular forces, and so more energy is needed to overcome them in order for melting or boiling to occur. These forces also explain why longer chain hydrocarbons are more viscous (i.e. thicker liquids, less easy to pour).

Combustion
Fuels are substances that react with **oxygen** to release useful energy. Most of the energy is released as heat, but light energy is also released.

Hydrocarbons
- The carbon oxidises to **carbon dioxide**

In general, for complete combustion:
hydrocarbon + oxygen → carbon dioxide + water



Determining the heat energy released by a fuel

Method
1. Measure 100cm³ 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.

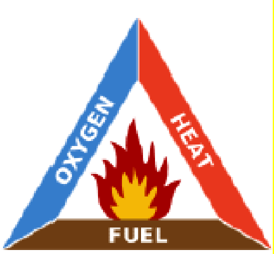
Control variables:
- Volume of water
- Distance from flame to conical flask

Improvements:
- Stirring
- Insulate
- Copper can instead of conical flask

Combustion of hydrogen
Hydrogen burns in oxygen to form water.
hydrogen + oxygen → water
$2H_2 + O_2 \rightarrow 2H_2O$

- **Disadvantages**
- Requires large amounts of electricity to produce hydrogen from water by electrolysis (how is this generated?)
- Storage requires bulky and heavy pressurised containers and is potentially hazardous as it forms an explosive mixture with air.

Advantages
- Produced from water therefore renewable
- water is the only product of its combustion so burning hydrogen does not contribute towards global warming or acid rain.

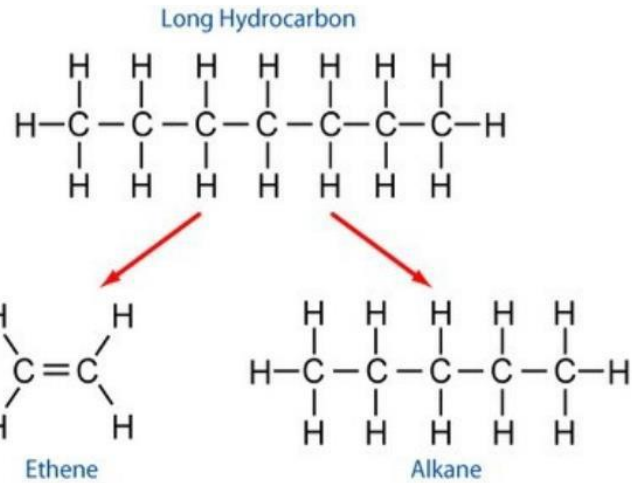


The fire triangle
The fire triangle is a simple way of understanding the factors essential for fire. Each side of the triangle represents one of the **three factors** required for the creation and maintenance of any fire; **oxygen**, **heat** and **fuel**.
Remove any one of these, the triangle is broken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

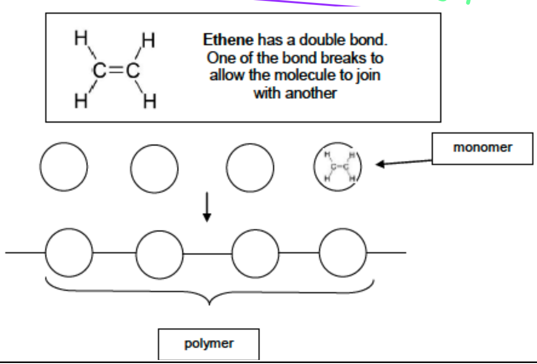
Cracking

The demand for short hydrocarbon molecules is greater than their supply in crude oil, so a reaction called **cracking** is used. **Cracking** converts long alkane molecules into shorter alkanes and alkenes, which are more useful.

At high temperatures, and with help from a catalyst, long hydrocarbon chains are broken down into smaller, more useful hydrocarbons including an **alkene**. One of the most common alkenes to be made is **ethene**.



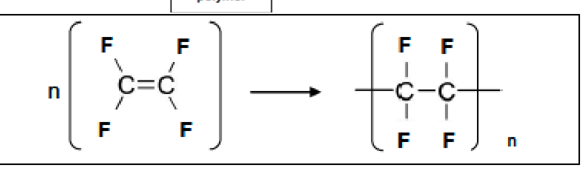
Ethene is a small reactive molecule, a monomer. If many ethene molecules are linked together it is called polythene which is used to make many plastics.



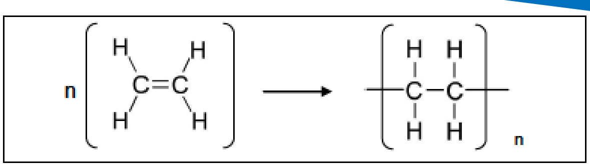
Creating Plastics

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

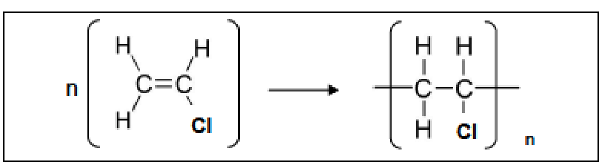
The process whereby monomers link to create a polymer is **polymerisation**:



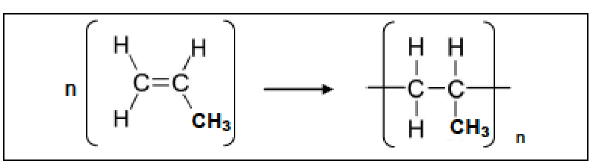
Tetrafluoroethene — Poly(tetrafluoroethene)



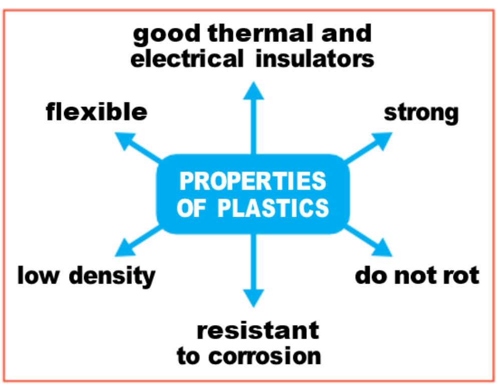
ethene — polythene



Vinyl chloride — Polyvinylchloride



propene — Polypropene



Landfill:

☺ Cheap and easy

☹ Plastic doesn't biodegrade, landfill sites full

Combustion:

☺ Can produce energy from waste

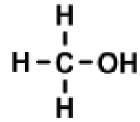
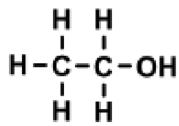
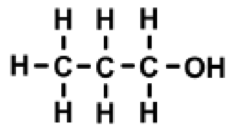
☹ Produces toxic fumes and CO_2

Recycling

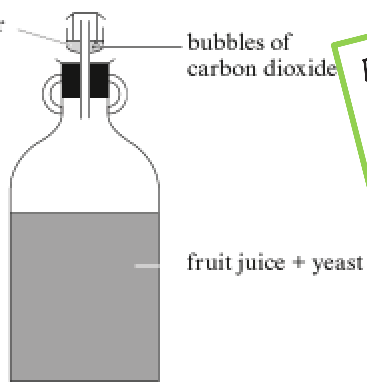
☺ Conserves natural resources

☹ Expensive

Alcohols (Triple only):

Name	Molecular Formula	Structural formula
Methanol	CH_3OH	
Ethanol	C_2H_5OH	
Propanol	C_3H_7OH	

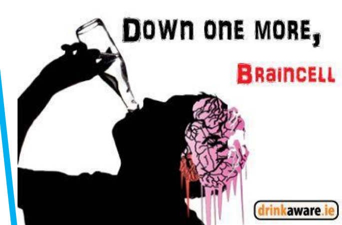
Alcohols contain the functional group –OH, and this is responsible for their properties. The general formula for alcohols is $C_nH_{2n+1}OH$. The names of alcohols end with '-ol'.



To obtain ethanol from the mixture, yeast is removed by **filtering**. Then the ethanol and water (and some sugar) mixture is **distilled**.

Ethanol can be made in a process called fermentation. Yeast contains an enzyme which breaks down sugar, making ethanol and carbon dioxide. It needs to happen in anaerobic conditions (when deprived of oxygen) and at about 35°C.

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$



There are social and health problems associated with 'binge drinking' and misuse of alcohol over a long period. Tax raised from the sale of alcoholic drinks generates significant revenue for the government.

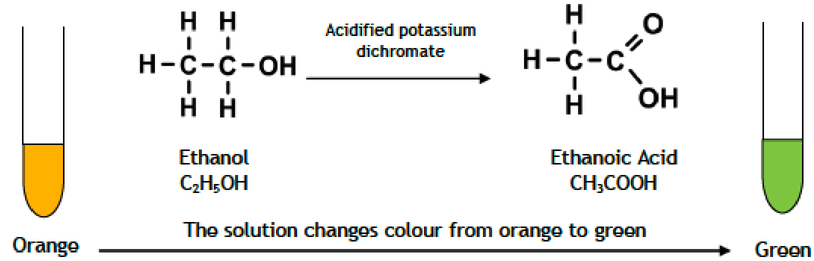
On the other hand, a significant amount of public money is spent each year in treating alcohol-related illnesses and in dealing with various other issues resulting from alcohol abuse.

Ethanol as a fuel

- Ethanol can be produced from the fermentation of plants such as sugar cane.
- Bioethanol produces only carbon dioxide and water as waste products.
- Bioethanol is carbon neutral because the carbon dioxide released during fermentation and combustion is equivalent to the amount removed from the atmosphere while the crop is growing.
- Bioethanol is also renewable.
- Less sulfur dioxide will be formed which prevents acid rain forming.
- Some critics warn of deforestation, and land being grabbed from food crops. This will increase food poverty as food prices are forced up.

Oxidation of alcohols

Alcohols are oxidised to carboxylic acids with acidified potassium dichromate:



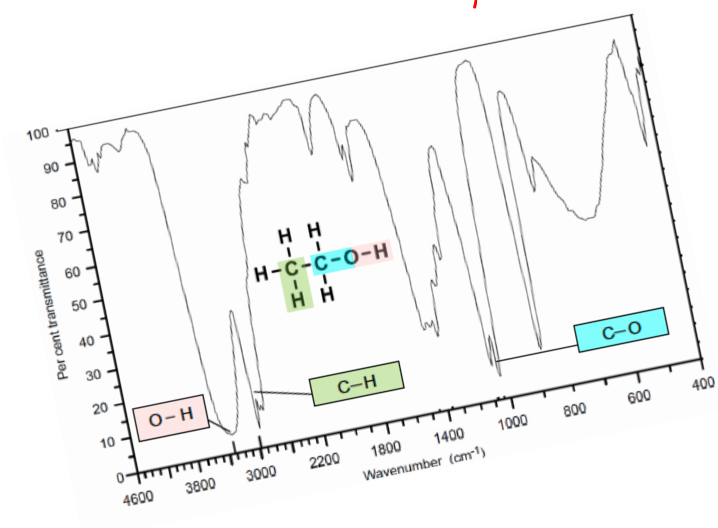
Infrared Spectroscopy

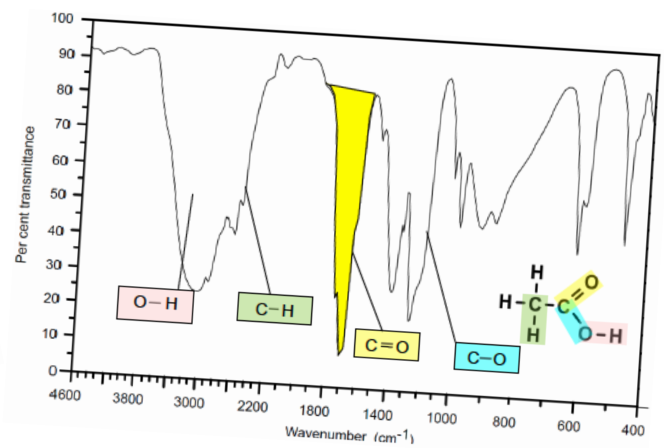
Infrared spectroscopy is used to ***identify*** the presence of ***certain bonds*** in organic molecules. All bonds in a molecule vibrate; ***different bonds*** will vibrate at ***different frequencies***.

Infrared spectroscopy characteristic absorption values

Bond	Wavenumber / cm^{-1}
C–O	1000 - 1300
C=C	1620 – 1670
C=O	1650 – 1750
C–H	2800 – 3100
O–H	2500 - 3550

The absorption values will be given in the exam





Topic 4: Metals

Ores are minerals found in the Earth's crust. They contain metal compounds, from which metals can be extracted.

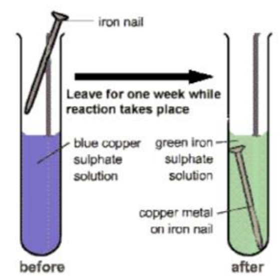
Name of ore	Metal compound
rock salt	sodium chloride
bauxite	aluminium oxide

Potassium
Sodium
Magnesium
Aluminium
(Carbon)
Zinc
Iron Tin
Lead
Copper
Silver
Gold

Reactivity decreasing

Reactivity series:
The method used to extract metals from their ores depends on the reactivity of the metal. The reactivity series allows us to predict how metals will react.

Gold and silver are examples of metals that are found native. The most reactive metals are extracted by electrolysis, while those towards the middle of the reactivity series can be chemically reduced.

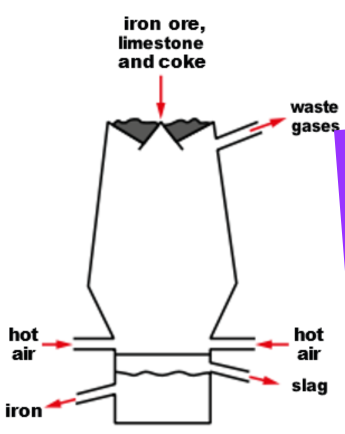


Displacement and competition reactions:
The relative reactivity of metals can be demonstrated by displacement reactions. A more reactive metal will displace a less reactive metal from a solution of one of its salts.

iron + copper(II) chloride → iron chloride + copper

During a competition reaction, a more reactive metal will remove oxygen from the oxide of a less reactive metal when a mixture of the two is heated. An example of this is the thermit reaction used in the rail industry to weld rails together on a track.

iron(III) oxide + aluminium → iron + aluminium oxide



The Blast Furnace

Raw materials
Iron ore: source of iron.
Coke: as a fuel and to produce carbon monoxide for the reduction.
Limestone: to remove impurities (slag formation).
Hot air: provides oxygen so that coke can burn.

Oxygen from the hot air reacts with carbon (coke) to form carbon monoxide. This is an exothermic reaction and heats the furnace.

$2C + O_2 \rightarrow 2CO$

Carbon monoxide then reacts with iron(III) oxide to give iron.

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

Limestone decomposes thermally to form calcium oxide.

$CaCO_3 \rightarrow CaO + CO_2$

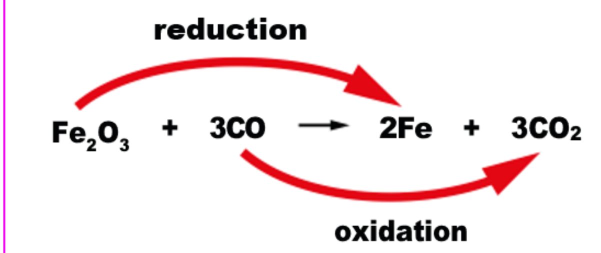
Calcium reacts with silicon dioxide (sand) to form slag.

$CaO + SiO_2 \rightarrow CaSiO_3$

The process in the blast furnace is continuous, with new raw materials added and products removed all the time due to the time and cost associated with getting the furnace up to temperature.

Oxidation and reduction:

Oxidation is the gain of oxygen whilst reduction is the loss of oxygen. For the reaction that happens in the blast furnace, the iron(III) oxide is reduced whilst the carbon monoxide is oxidised.

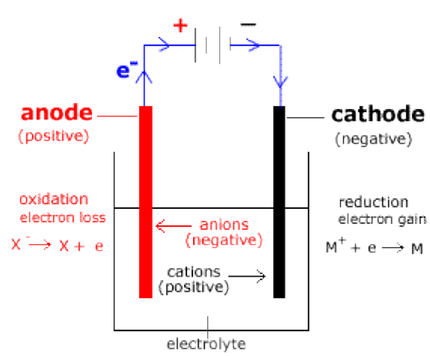


Electrolysis

- Electrolysis is the process of breaking down an ionic compound using electrical energy.
- Electrolyte: liquid* which is broken down by an electrical charge.
- Electrode: graphite rods which carry a current in and out of the electrolyte.
- Cathode: negative electrode.
- Anode: positive electrode.

  * so that the ions are free to move.

The **positive electrode**, called the **anode**, will attract **negatively** charged **non-metal ions**. The non-metal ions **lose electrons** to the anode (this is called **oxidation**) and are discharged as **non-metal atoms** which often combine to form molecules.



The **negative electrode**, called the **cathode**, will attract **positively** charged **metal ions**. The metal ions **gain electrons** from the cathode (this is called **reduction**) and are discharged as **metal atoms**.

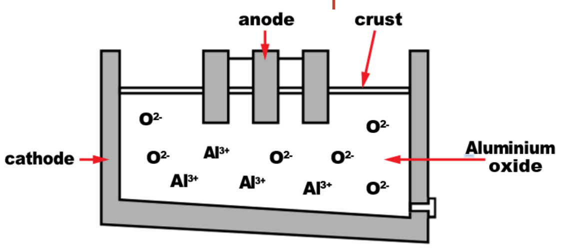
Positive
Anode
Negative
Is
Cathode

Oxidation
Is
Loss (of electrons)
Reduction
Is
Gain (of electrons)

Aluminium Electrolysis:

The electrolysis process can be used to extract aluminium from aluminium oxide.

aluminium oxide → aluminium + oxygen



Alumina (aluminium oxide) dissolves in molten cryolite at a temperature much lower than its melting point, therefore saving energy.

On the cathode: $Al^{3+} + 3é \rightarrow Al$

On the anode: $2O^{2-} - 4é \rightarrow O_2$

The oxygen formed reacts with the carbon anodes, forming carbon dioxide gas and requiring anodes to be replaced frequently.

Alloys:

An alloy is a mixture made by mixing molten metals. Its properties can be modified by changing its composition. Steel is much harder and stronger than iron and is therefore more useful.

Copper

Very good conductor of heat and electricity, malleable and ductile, attractive colour and lustre.

Aluminium

Strong, low density, good conductor of heat and electricity, resistant to corrosion.

Titanium

Hard, strong, low density, resistant to corrosion, high melting point.

Transition metals:

Transition metals are found in the centre of the Periodic Table and they display the typical metallic properties of high melting and boiling points, malleability, high density, good electrical and thermal conductivity. Many transition metals are useful catalysts – e.g. iron in the manufacture of ammonia, platinum in catalytic converters. They can form more than one type of ion, e.g. Fe^{2+}/Fe^{3+} and their compounds are often coloured.

Ion	Colour of compounds/ solutions
Fe^{2+}	pale green
Fe^{3+}	brown
Cu^{2+}	blue

Higher tier:

We can test for the presence of transition metal ions by using a solution of sodium hydroxide.

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

blue precipitate

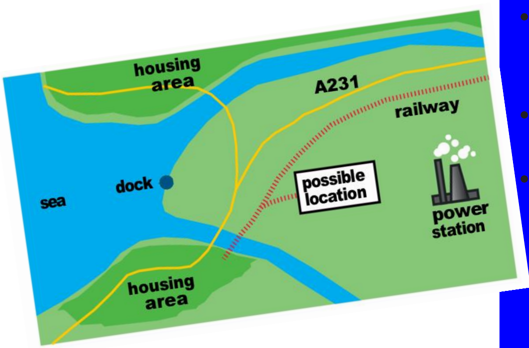
$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$

green precipitate

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

orange/brown precipitate

Aluminium plants:

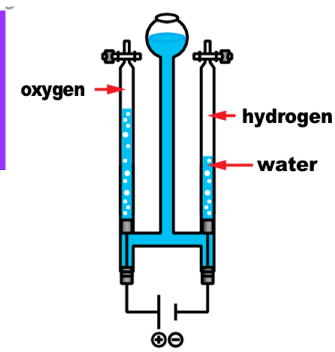


Factors which impact location:

- Location near the coast in order to import raw materials
- A site away from built up areas
- A town or city within commuting distance to accommodate the workforce
- Good transport links for transporting the product to buyers
- A direct electricity supply in the case of aluminium.

Electrolysis of water (TRIPLE ONLY)

Hydrogen gas is collected on the cathode while oxygen gas is collected on the anode.



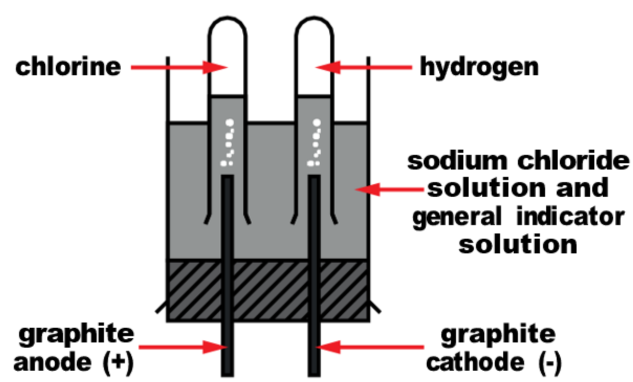
On the cathode: $2H^+ + 2é \rightarrow H_2$
On the anode: $2OH^- \rightarrow O_2 + 2H^+ + 4é$
The volume of the hydrogen formed is twice the volume of the oxygen formed, as there are two hydrogen atoms for every oxygen atom in a molecule of water (H_2O).

Electrolysis of aqueous solutions (TRIPLE ONLY)

In aqueous solution there are H^+ and OH^- ions as well as the ions from the dissolved salt. Metals lower in the reactivity series than hydrogen are formed on the cathode. In the case of copper(II) chloride solution, the products are copper metal and chlorine gas.

On the cathode: $Cu^{2+} + 2é \rightarrow Cu$

On the anode: $2Cl^- \rightarrow Cl_2 + 2é$

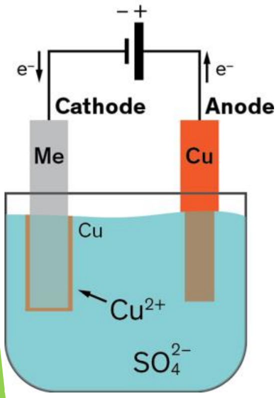


But when the dissolved salt includes ions from metals that are higher than hydrogen in the reactivity series, hydrogen gas is formed rather than the metal.

On the cathode: $2H^+ + 2é \rightarrow H_2$

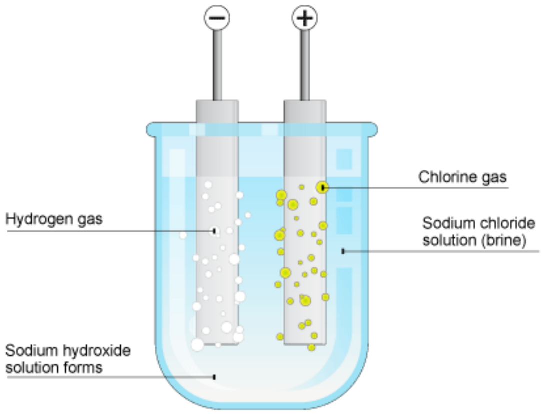
On the anode: $2Cl^- \rightarrow Cl_2 + 2é$

Electroplating (TRIPLE ONLY)
Electrolysis is used to electroplate objects. This is useful for **coating a cheaper metal** with a more expensive one, such as copper or silver.



How it works
- The negative electrode should be the object that is to be electroplated
- The positive electrode should be the metal that you want to coat the object with
- The electrolyte should be a solution of the coating metal, such as its metal nitrate or sulfate

Manufacturing sodium hydroxide (TRIPLE ONLY)
Useful substances can be obtained by the **electrolysis of sodium chloride solution.**
During electrolysis:
- **chlorine gas** forms at the anode (positive electrode)
- **hydrogen gas** forms at the cathode (negative electrode)
- a solution of **sodium hydroxide** forms.



Topic 5: Chemical reactions and energy

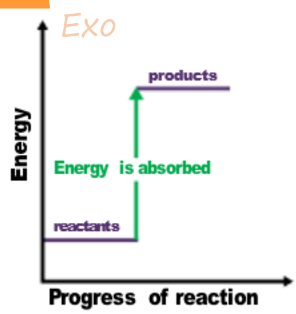
- Almost every chemical reaction is accompanied by an energy change.
- When a chemical reaction happens, energy is transferred to or from the surroundings.
- Energy changes can be explained by examining the changes in chemical bonding during a reaction.
- Energy changes are used to classify reactions as exothermic or endothermic.

Energy Profiles:

- An energy level diagram shows whether a reaction is exothermic or endothermic.
- The diagram shows the energy in the reactants and products, and the difference in energy between them.

Endothermic Reactions:

- An endothermic reaction happens when energy is taken in from the surroundings.
- During an endothermic reaction, the temperature of the surroundings decreases.

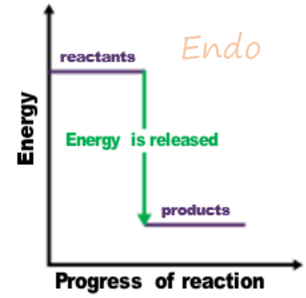


- The products are at a lower energy level than the reactants.
- Energy is given out to the surroundings.
- The downwards arrow shows that energy is given out.

Exothermic Reactions:

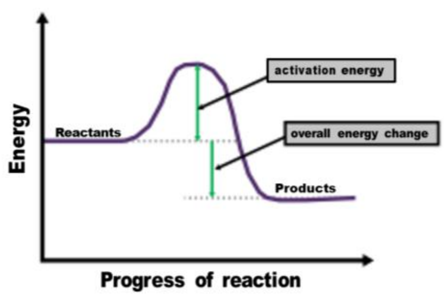
- An exothermic reaction happens when energy is transferred to the surroundings.
- During an exothermic reaction, the temperature of the surroundings increases.

- The products are at a lower energy level than the reactants.
- Energy is given out to the surroundings.
- The downwards arrow shows that energy is given out.



Reaction Profiles:

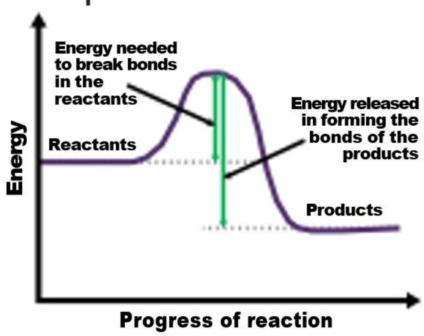
- A reaction profile shows how the energy of the chemicals changes during a reaction.
- A reaction profile also includes the activation energy – the minimum energy needed to start a reaction.
- The activation energy is shown as a 'hump' in the line.



During a chemical reaction:

- Energy is needed to break the bonds in the reactants
- Energy is released when the bonds in the products are formed
- The difference between the energy needed to break bonds and the
- energy released when forming bonds determines the type of reaction.

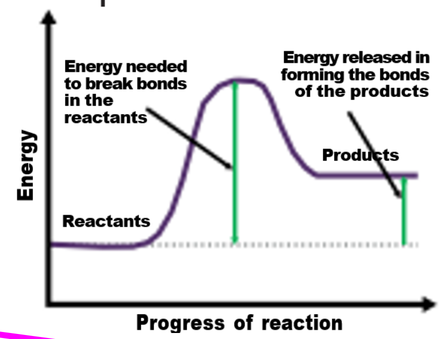
The profile shows that:



Exothermic reactions:

- More energy is released when new bonds are made than is needed to break existing bonds
- The overall energy change is negative – meaning that energy is given out to the surroundings.

The profile shows that:



Endothermic reactions:

- More energy is taken in when the existing bonds are broken than is released in making new bonds
- The overall energy change is positive – meaning that energy is taken in from the surroundings.

Calculating bond energies:

- The energy change in a reaction can be calculated using bond energies.
- A bond energy is the amount of energy needed to break one mole of covalent bonds of a given type.

To calculate the energy change for a reaction:

1. Add together the bond energies for all the bonds in the reactants – this is the 'energy in'.
2. Add together the bond energies for all the bonds in the products – this is the 'energy out'.
3. Overall energy change = energy in – energy out.

Example

Hydrogen and chlorine react to form hydrogen chloride.

H—H + Cl—Cl → H—Cl H—Cl

Bond	Bond energy (kJ/mol)
H—H	436
Cl—Cl	243
H—Cl	432

Energy in = 436 + 243 = 679 kJ/mol
Energy out = (2 × 432) = 864 kJ/mol
Energy change = in – out = 679 – 864 = –185 kJ/mol
The energy change is negative. More energy is given out than is taken in and the reaction is exothermic.
Note – positive energy change = endothermic reaction!

TRIPLE ONLY: REVERSIBLE REACTIONS

Reversible Reactions:

- Reversible Reaction – a reaction that happens in both directions. This means the products of the reaction can react together to produce the original reactants.
- ⇌ – the symbol used to represent a reversible reaction.

If the forward reaction is exothermic, the reverse reaction is endothermic.

The Haber Process:

- The Haber process is used in the industrial production of ammonia.
- Ammonia is a pungent smelling alkaline gas with the chemical formula NH_3.
- Nitrogen gas (from the air) and hydrogen gas (from natural gas) react together to produce ammonia.

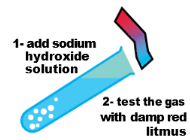
Nitrogen	+	Hydrogen	⇌	Ammonia
$N_{2(g)}$		$3H_{2(g)}$	⇌	$2NH_{3(g)}$

- The ammonia is collected by cooling the reaction mixture so the ammonia condenses into a liquid.
- The unreacted nitrogen and hydrogen are recycled back through the process, so there is no waste.

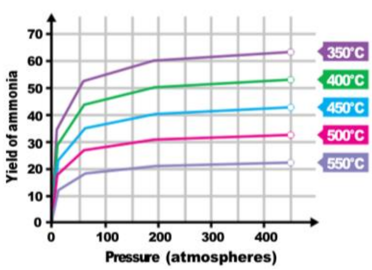
Tests for ammonia gas and ammonium ions

Ammonia gas (NH_3) – **Ammonia gas will change damp red litmus paper blue.**

Ammonium ion (NH_4^+)



The ammonium ions convert into ammonia gas which turns the damp red litmus paper blue.



The graph shows that a ***lower temperature*** and ***higher pressure*** would produce the ***best theoretical yield***

The Chosen Reaction Conditions:

The reaction conditions for the process are a compromise between the yield of production, rate of production, cost and safety.

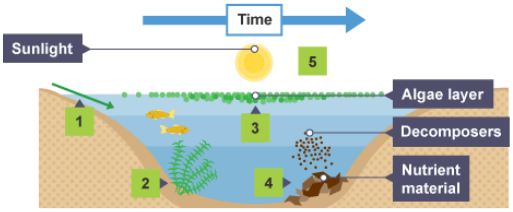
- The rate of production is too slow at a lower temperature.
- A higher temperature is a compromise between yield and rate.
- Operating at higher pressures is expensive. There is also more risk of explosions. A lower pressure is a compromise between yield and cost/safety.
- The catalyst works like any catalyst – speeding up the rate of production, without getting used up. However, over time, it does get poisoned and needs replacing.

Fertilisers:

- The majority of ammonia and sulfuric acid produced is used to make fertilisers.
- Ammonium sulfate – common fertiliser made by neutralising the sulfuric acid* with ammonia or ammonium hydroxide.

ammonia	+	sulfuric acid	→	ammonium sulfate
$2NH_3$		H_2SO_4	→	$(NH_4)_2SO_4$

ammonium hydroxide	+	sulfuric acid	→	ammonium sulfate	+	water
$2NH_4OH$	+	H_2SO_4	→	$(NH_4)_2SO_4$	+	$2H_2O$



1 **Nutrient load up:** excessive nutrients from fertilisers are flushed from the land into rivers or lakes by rainwater.

2 **Plants flourish:** these pollutants cause aquatic plant growth of algae, duckweed and other plants.

3 **Algae blooms, oxygen is depleted:** algae blooms prevent sunlight reaching other plants. The plants die and oxygen in the water is depleted.

4 **Decomposition further depletes oxygen:** dead plants are broken down by bacteria decomposers, using up even more oxygen in the water.

5 **Death of the ecosystem:** oxygen levels reach a point where no life is possible. Fish and other organisms die.

Advantages

- Increases crop yield
- Healthier crops
- Improves soil quality

Advantages and disadvantages of fertilisers:

Disadvantages

- Eutrophication
- Risk of stomach cancer
- Blue baby syndrome

The Contact Process:

- The Contact process is used in the industrial production of sulfuric acid, H_2SO_4.
- The process is in 3 stages. The raw materials are sulfur (stage 1), air (stages 1 + 2) and water (stage 3).

Stage 1: Sulfur buns in air to form sulfur dioxide gas.

sulfur	+	oxygen	→	sulfur dioxide
$S_{(s)}$	+	$O_{2(g)}$	→	$SO_{2(g)}$

Stage 2: Sulfur dioxide reacts with more oxygen to form sulfur trioxide gas:

sulfur dioxide	+	oxygen	⇌	sulfur trioxide
$2SO_{2(g)}$	+	$O_{2(g)}$	⇌	$2SO_{3(g)}$

The reaction in this stage is reversible.
The conditions used are:

» 400 – 500°C
» Atmospheric pressure
» Vanadium(V) oxide catalyst

Stage 3: Sulfur trioxide is dissolved in concentrated sulfuric acid to produce oleum.

sulfur trioxide	+	sulfuric acid	→	oleum
$SO_{3(g)}$	+	$H_2SO_{4(l)}$	→	$H_2S_2O_{7(l)}$

The oleum is then diluted with water to produce sulfuric acid.

oleum	+	Water	→	sulfuric acid
$H_2S_2O_{7(l)}$	+	$H_2O_{(l)}$	→	$2H_2SO_{4(l)}$

Note – adding sulfur trioxide directly to water is too violent!!!

What are the uses of Sulfuric Acid?

a. The acid in a car battery.
b. Making detergents.
c. Metal treatment and anodising.
d. A catalyst.
e. A dehydrating agent.
f. Making fertiliser.
g. Paints and dyes.

Sulfuric acid as a dehydrating agent:

Concentrated sulfuric acid is a dehydrating agent – it removes water from a substance.

In **glucose** – the concentrated sulfuric acid takes away the elements of water leaving **only carbon.**

glucose	→	carbon	+	water
$C_6H_{12}O_6$	→	$6C_{(s)}$	+	$H_2O_{(g)}$