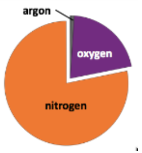
GCSE Chemistry of the atmosphere

Composition and evolution of the atmosphere

Proportions of gases in the atmosphere



Gas	Percentage
Nitrogen	*~80%*
Oxygen	*~20%*
Argon	*0.93%*
Carbon dioxide	*0.04%*

The Earth's early atmosphere

Volcano activity 1st Billion years	*Billions of years ago there was intense volcanic activity*	This released gases (mainly CO_2) that formed to early atmosphere and water vapour that condensed to form the oceans.
Other gases	*Released from volcanic eruptions*	Nitrogen was also released, gradually building up in the atmosphere. Small proportions of ammonia and methane also produced.
Reducing carbon dioxide in the atmosphere	*When the oceans formed, carbon dioxide dissolved into it*	This formed carbonate precipitates, forming sediments. This reduced the levels of carbon dioxide in the atmosphere.

How oxygen increased

Algae and plants	*These produced the oxygen that is now in the atmosphere, through photosynthesis.*	carbon dioxide + water → glucose + oxygen $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
Oxygen in the atmosphere	*First produced by algae 2.7 billion years ago.*	Over the next billion years plants evolved to gradually produce more oxygen. This gradually increased to a level that enabled animals to evolve.

How carbon dioxide decreased

Reducing carbon dioxide in the atmosphere	*Algae and plants*	These gradually reduced the carbon dioxide levels in the atmosphere by absorbing it for photosynthesis.
Formation of sedimentary rocks and fossil fuels	*These are made out of the remains of biological matter, formed over millions of years*	Remains of biological matter falls to the bottom of oceans. Over millions of years layers of sediment settled on top of them and the huge pressures turned them into coal, oil, natural gas and sedimentary rocks. The sedimentary rocks contain carbon dioxide from the biological matter.

CO_2 and methane as greenhouse gases

Greenhouse gases

Carbon dioxide, water vapour and methane	*Examples of greenhouse gases that maintain temperatures on Earth in order to support life*
The greenhouse effect	*Radiation from the Sun enters the Earth's atmosphere and reflects off of the Earth. Some of this radiation is re-radiated back by the atmosphere to the Earth, warming up the global temperature.*

Human activities and greenhouse gases

Carbon dioxide	*Human activities that increase carbon dioxide levels include burning fossil fuels and deforestation.*
Methane	*Human activities that increase methane levels include raising livestock (for food) and using landfills (the decay of organic matter released methane).*
Climate change	*There is evidence to suggest that human activities will cause the Earth's atmospheric temperature to increase and cause climate change.*

Carbon footprints

The total amount of greenhouse gases emitted over the full life cycle of a product/event. This can be reduced by reducing emissions of carbon dioxide and methane.

Global climate change

Effects of climate change
Rising sea levels
Extreme weather events such as severe storms
Change in amount and distribution of rainfall
Changes to distribution of wildlife species with some becoming extinct

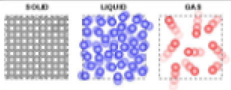
Common atmospheric pollutants

Atmospheric pollutants from fuels

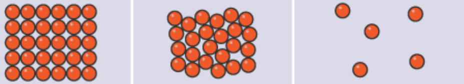
Combustion of fuels	*Source of atmospheric pollutants. Most fuels may also contain some sulfur.*
Gases from burning fuels	*Carbon dioxide, water vapour, carbon monoxide, sulfur dioxide and oxides of nitrogen.*
Particulates	*Solid particles and unburned hydrocarbons released when burning fuels.*

Properties and effects of atmospheric pollutants

Carbon monoxide	*Toxic, colourless and odourless gas. Not easily detected, can kill.*
Sulfur dioxide and oxides of nitrogen	*Cause respiratory problems in humans and acid rain which affects the environment.*
Particulates	*Cause global dimming and health problems in humans.*

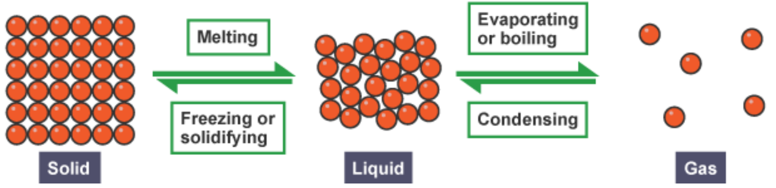
Solid, liquid, gas	*Melting and freezing happen at melting point, boiling and condensing happen at boiling point.*	 The amount of energy needed for a state change depends on the strength of forces between particles in the substance.	(HT only) Limitations of simple model: • There are no forces in the model • All particles are shown as spheres • Spheres are solid	

s	solid
l	liquid
g	gas

State	Solid	Liquid	Gas
Closeness of particles	Very close	Close	Far apart
Arrangement of particles	Regular pattern	Randomly arranged	Randomly arranged
Movement of particles	Vibrate around a fixed position	Move around each other	Move quickly in all directions
Energy of particles	Low energy	Greater energy	Highest energy
2D diagram			

Sublimation/ subliming = when a substance can change directly from solid to gas, or from gas to solid, without becoming a liquid in between

Temperature	Predicted state
Given temperature < melting point	Solid
Given temperature is between melting and boiling points	Liquid
Given temperature > boiling point	Gas



Activation energy	*Chemical reactions only happen when particles collide with sufficient energy*	The minimum amount of energy that colliding particles must have in order to react is called the activation energy.

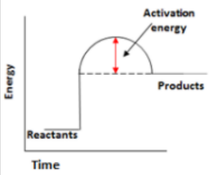
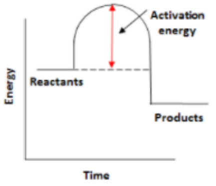
Balanced symbol equations	*Represent chemical reactions and have the same number of atoms of each element on both sides of the equation*	$H_2 + Cl_2 \rightarrow 2HCl$ Subscript numbers show the number of atoms of the element to its left. Normal script numbers show the number of molecules.

Bond energy calculation	Bond breaking: 945 + (3 x 436) = 945 + 1308 = 2253 kJ/mol Bond making: 6 x 391 = 2346 kJ/mol Overall energy change = 2253 - 2346 = -93kJ/mol Therefore reaction is exothermic overall.

Endothermic	*Energy is taken in from the surroundings so the temperature of the surroundings decreases*	• Thermal decomposition • Sports injury packs
Exothermic	*Energy is transferred to the surroundings so the temperature of the surroundings increases*	• Combustion • Hand warmers • Neutralisation

Overall energy change of reaction	*Exothermic*	Energy released making new bonds is greater than the energy taken in breaking existing bonds.
	Endothermic	Energy needed to break existing bonds is greater than the energy released making new bonds.

Breaking bonds in reactants	*Endothermic process*
Making bonds in products	*Exothermic process*

Endo thermic		Products are at a higher energy level than the reactants. As the reactants form products, energy is transferred from the surroundings to the reaction mixture. The temperature of the surroundings decreases because energy is taken in during the reaction.
Exo thermic		Products are at a lower energy level than the reactants. When the reactants form products, energy is transferred to the surroundings. The temperature of the surroundings increases because energy is released during the reaction.

Gas	Test	Positive result
Hydrogen	*Burning splint*	'Pop' sound.
Oxygen	*Glowing splint*	Re-lights the splint.
Chlorine	*Litmus paper (damp)*	Bleaches the paper white.
Carbon dioxide	*Limewater*	Goes cloudy (as a solid calcium carbonate forms).

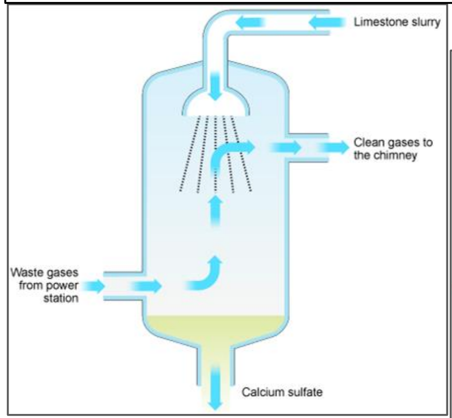
Sterilising agents include chlorine, ozone and UV light.

Potable water

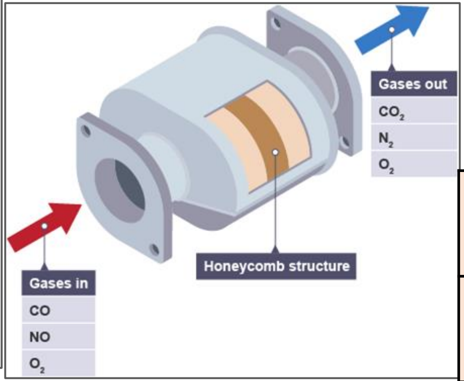
Potable water	*Water of an appropriate quality is essential for life*	Human drinking water should have low levels of dissolved salts and microbes. This is called potable water.
UK water	*Rain provides water with low levels of dissolved substances*	This water collects in the ground/lakes/rivers. To make potable water an appropriate source is chosen, which is then passed through filter beds and then sterilised.
Desalination	*Needs to occur is fresh water is limited and salty/sea water is needed for drinking*	This can be achieved by distillation or by using large membranes e.g. reverse osmosis. These processes require large amounts of energy.

Waste water	*Produced from urban lifestyles and industrial processes*	These require treatment before used in the environment. Sewage needs the organic matter and harmful microbes removed.
Sewage treatment	*Includes many stages*	- Screening and grit removal - Sedimentation to produce sludge and effluent (liquid waste or sewage). - Anaerobic digestion of sludge - Aerobic biological treatment of effluent.

Sulfur scrubbing - reduces pollutants in the atmosphere from factories. Reduces acid rain.

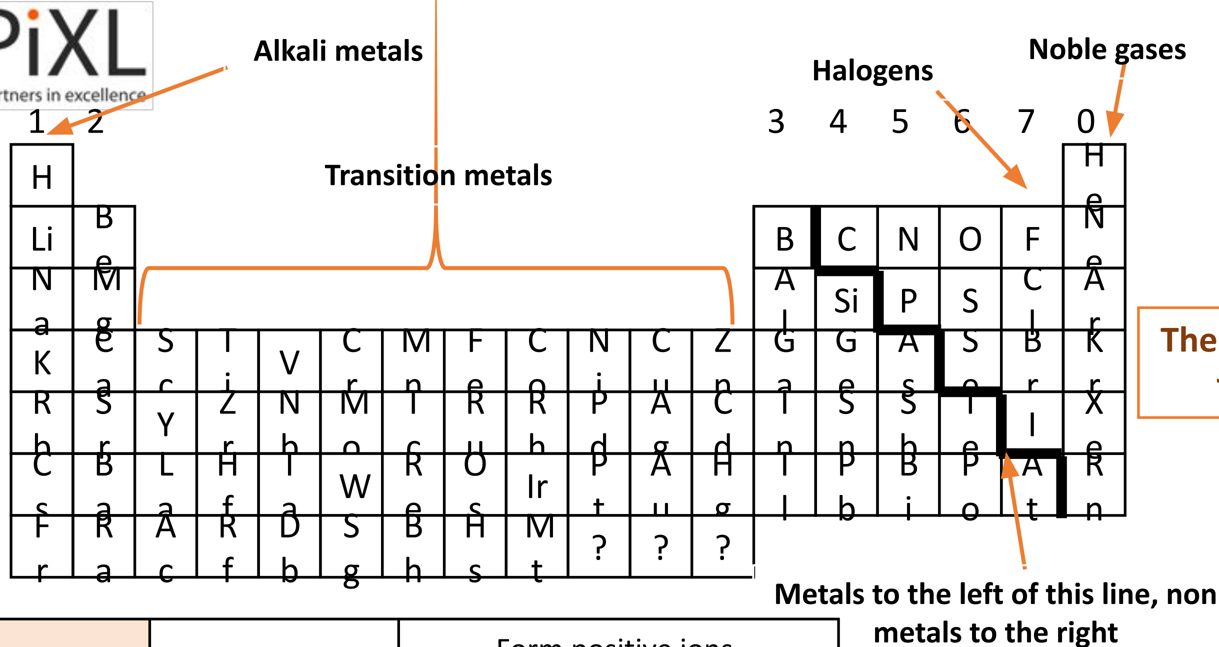


Catalytic converters - help reduce the release of toxic gases from the exhaust pipe of vehicles. They contain platinum and rhodium, which act as catalysts.



Oxidation	*This is when oxygen is gained by a compound during a reaction*	e.g. metals reacting with oxygen, rusting of iron

Chemical equations	*Show chemical reactions - need reactant(s) and product(s) energy always involves and energy change*	Law of conservation of mass states the total mass of products = the total mass of reactants.
Word equations	*Uses words to show reaction* **reactants → products** **magnesium + oxygen → magnesium oxide**	Does not show what is happening to the atoms or the number of atoms.
Symbol equations	*Uses symbols to show reaction* **reactants → products** **$2Mg + O_2 \rightarrow 2MgO$**	Shows the number of atoms and molecules in the reaction, these need to be balanced.



C2 Knowledge Organsier

The Periodic table

Elements arranged in order of atomic number	*Elements with similar properties are in columns called groups*	Elements in the same group have the same number of outer shell electrons and elements in the same period (row) have the same number of electron shells.

Development of the Periodic table

Before discovery of protons, neutrons and electrons	*Elements arranged in order of atomic weight*	Early periodic tables were incomplete, some elements were placed in inappropriate groups if the strict order atomic weights was followed.
Mendeleev	*Left gaps for elements that hadn't been discovered yet*	Elements with properties predicted by Mendeleev were discovered and filled in the gaps. Knowledge of isotopes explained why order based on atomic weights was not always correct.

Metals and non metals

Metals	*To the left of the Periodic table*	Form positive ions. Conductors, high melting and boiling points, ductile, malleable.
Non metals	*To the right of the Periodic table*	Form negative ions. Insulators, low melting and boiling points.

Group 1

Alkali metals	*Very reactive with oxygen, water and chlorine*	Only have one electron in their outer shell. Form +1 ions.
	Reactivity increases down the group	Negative outer electron is further away from the positive nucleus so is more easily lost.

With oxygen	Forms a metal oxide	Metal + oxygen □ metal oxide	e.g. $4Na + O_2$ □ $2Na_2O$
With water	Forms a metal hydroxide and hydrogen	Metal + water □ metal hydroxide + hydrogen	e.g. $2Na + 2H_2O$ □ $2NaOH + H_2$
With chlorine	Forms a metal chloride	Metal + chlorine □ metal chloride	e.g. $2Na + Cl_2$ □ $2NaCl$

Group 7

Halogens	*Consist of molecules made of a pair of atoms*	Have seven electrons in their outer shell. Form -1 ions.
	Melting and boiling points increase down the group (gas □ liquid □ solid)	Increasing atomic mass number.
	Reactivity decreases down the group	Increasing proton number means an electron is more easily gained

With metals	Forms a metal halide	Metal + halogen □ metal halide e.g. Sodium + chlorine □ sodium chloride	e.g. NaCl metal atom loses outer shell electrons and halogen gains an outer shell electron
With hydrogen	Forms a hydrogen halide	Hydrogen + halogen □ hydrogen halide e.g. Hydrogen + bromine □ hydrogen bromide	e.g. $Cl_2 + H_2$ □ $2HCl$
With aqueous solution of a halide salt	A more reactive halogen will displace the less reactive halogen from the salt	Chlorine + potassium bromide □ potassium chloride + bromine	e.g. $Cl_2 + 2KBr$ □ $2KCl + Br_2$

Group 0

Noble gases	*Unreactive, do not form molecules*	This is due to having full outer shells of electrons.
	Boiling points increase down the group	Increasing atomic number.

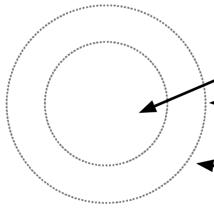
Transition metals (Chemistry only)

Compared to group 1	• *Less reactive* • *Harder* • *Denser* • *Higher melting points*	• *Cu^{2+} is blue* • *Ni^{2+} is pale green, used in the manufacture of margarine* • *Fe^{2+} is green, used in the Haber process* • *Fe^{3+} is reddish-brown* • *Mn^{2+} is pale pink*
Typical properties	• *Many have different ion possibilities with different charges* • *Used as catalysts* • *Form coloured compounds*	

GCSE Atomic structure and periodic table part 1

Atoms, elements and compounds

Atom	*The smallest part of an element that can exist*	Have a radius of around 0.1 nanometres and have no charge (0).
Element	*Contains only one type of atom*	Around 100 different elements each one is represented by a symbol e.g. O, Na, Br.
Compound	*Two or more elements chemically combined*	Compounds can only be separated into elements by chemical reactions.



Central nucleus	Contains protons and neutrons
Electron shells	Contains electrons

Electronic structures

Electronic shell	Max number of electrons
1	2
2	8
3	8
4	2

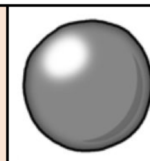
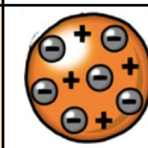
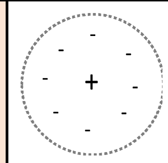
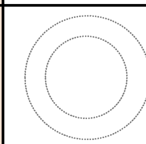
Relative electrical charges of subatomic particles

Name of Particle	Relative Charge	Relative Mass
Proton	+1	1
Neutron	0	1
Electron	-1	Very small

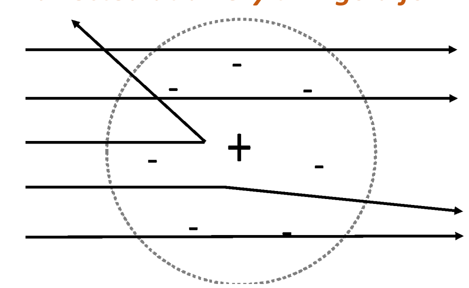
7
Li
3

Mass number	*The sum of the protons and neutrons in the nucleus*	
Atomic number	*The number of protons in the atom*	Number of electrons = number of protons

The development of the model of the atom

Pre 1900		*Tiny solid spheres that could not be divided*	Before the discovery of the electron, John Dalton said the solid sphere made up the different elements.
1897 'plum pudding'		*A ball of positive charge with negative electrons embedded in it*	JJ Thompson 's experiments showed that showed that an atom must contain small negative charges (discovery of electrons).
1909 nuclear model		*Positively charge nucleus at the centre surrounded negative electrons*	Ernest Rutherford's alpha particle scattering experiment showed that the mass was concentrated at the centre of the atom.
1913 Bohr model		*Electrons orbit the nucleus at specific distances*	Niels Bohr proposed that electrons orbited in fixed shells; this was supported by experimental observations.

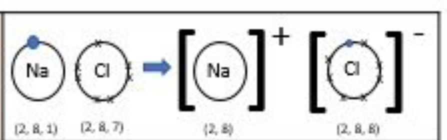
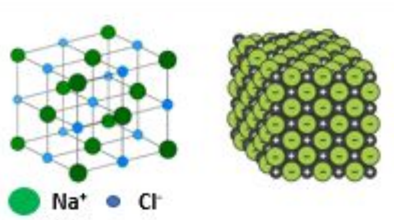
James Chadwick	*Provided the evidence to show the existence of neutrons within the nucleus*

Rutherford's scattering experiment	*A beam of alpha particles are directed at a very thin gold foil* 	Most of the alpha particles passed right through. A few (+) alpha particles were deflected by the positive nucleus. A tiny number of particles reflected back from the nucleus.

Relative atomic mass

Isotopes	*Atoms of the same element with the same number of protons and different numbers of neutrons*	**^{35}Cl (75%) and ^{37}Cl (25%)** Relative abundance = (% isotope 1 x mass isotope 1) + (% isotope 2 x mass isotope 2) ÷ 100 **e.g. (25 x 37) + (75x 35) ÷ 100 = 35.5**

Ionic compounds

Dot and cross diagram	
Giant structure	

Structure	• *Held together by strong electrostatic forces of attraction between oppositely charged ions* • *Forces act in all directions in the lattice*

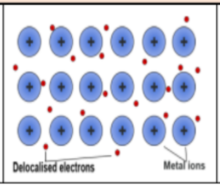
Ionic bonding

Electrons are transferred so that all atoms have a noble gas configuration (full outer shells).	*Metal atoms lose electrons and become positively charged ions*	Group 1 metals form +1 ions Group 2 metals form +2 ions
	Non metals atoms gain electrons to become negatively charged ions	Group 6 non metals form -2 ions Group 7 non metals form -1 ions

C3 - Chemicals of the natural environment

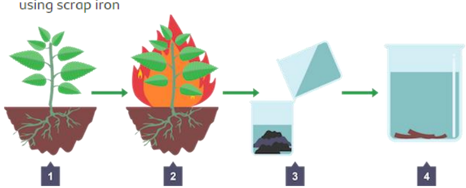
Metals as conductors	
High melting and boiling points	This is due to the strong metallic bonds.
Pure metals can be bent and shaped	Atoms are arranged in layers that can slide over each other.
Good conductors of electricity	Delocalised electrons carry electrical charge through the metal.
Good conductors of thermal energy	Energy is transferred by the delocalised electrons.

Oxidation **I**s **L**oss (of electrons)
Reduction **I**s **G**ain (of electrons)

Metallic bonding		
Giant structure of atoms arranged in a regular pattern		Electrons in the outer shell of metal atoms are delocalised and free to move through the whole structure. This sharing of electrons leads to strong metallic bonds.

Phytomining is used to extract copper from soil containing its ore.

1. Plants are used to absorb metal compounds such as copper(II) compounds
2. The plants are harvested, then burned to produce ash, which contains the metal compounds
3. An acid is added to the ash to produce a solution containing dissolved metal compounds (leachate)
4. Copper can be obtained from these solutions by displacement using scrap iron



Metals form positive ions when they react	*The reactivity of a metal is related to its tendency to form positive ions*	The reactivity series arranges metals in order of their reactivity (their tendency to form positive ions).
Carbon and hydrogen	*Carbon and hydrogen are non-metals but are included in the reactivity series*	These two non-metals are included in the reactivity series as they can be used to extract some metals from their ores, depending on their reactivity.
Displacement	*A more reactive metal can displace a less reactive metal from a compound.*	Silver nitrate + Sodium chloride → Sodium nitrate + Silver chloride

Extraction using carbon	
Metals less reactive than carbon can be extracted from their oxides by reduction.	For example: zinc oxide + carbon → zinc + carbon dioxide

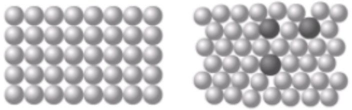
Extraction of metals and reduction	Unreactive metals, such as gold, are found in the Earth as the metal itself. They can be mined from the ground.

	most reactive	
potassium		K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

Bioleaching = Certain **bacteria** can break down low-grade ores to produce an **acidic solution** containing copper ions. The solution is called a **leachate** and the process is called **bioleaching**.

Alloys	*Mixture of two or more elements at least one of which is a metal*	Harder than pure metals because atoms of different sizes disrupt the layers so they cannot slide over each other.

Pure metal | Alloy



Word equations	*Uses words to show reaction* reactants → products magnesium + oxygen → magnesium oxide	Does not show what is happening to the atoms or the number of atoms.
Symbol equations	*Uses symbols to show reaction* reactants → products $2Mg + O_2 \rightarrow 2MgO$	Shows the number of atoms and molecules in the reaction, these need to be balanced.

Displacement in solutions

A more **reactive** metal can **displace** a less reactive metal from its **compounds**. For example, magnesium is more reactive than copper. It displaces copper from copper sulfate **solution**:

Magnesium + copper sulfate → magnesium sulfate + copper

$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

Electrolysis

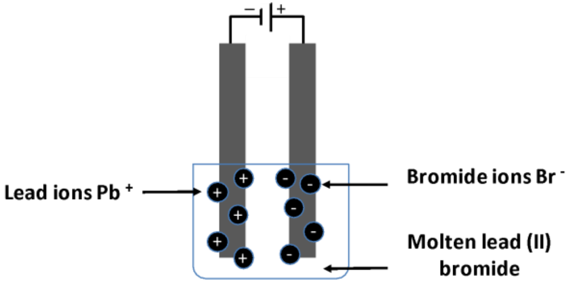
Process of electrolysis	*Splitting up using electricity*	When an ionic compound is melted or dissolved in water, the ions are free to move. These are then able to conduct electricity and are called electrolytes. Passing an electric current through electrolytes causes the ions to move to the electrodes.
Electrode	*Anode* *Cathode*	The positive electrode is called the anode. The negative electrode is called the cathode.
Where do the ions go?	*Cations* *Anions*	Cations are positive ions and they move to the negative cathode. Anions are negative ions and they move to the positive anode.

Extracting metals using electrolysis	
	Metals can be extracted from molten compounds using electrolysis.
	This process is used when the metal is too reactive to be extracted by reduction with carbon.
	The process is expensive due to large amounts of energy needed to produce the electrical current. *Example: aluminium is extracted in this way.*

The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.

Electrolysis of aqueous solutions

At the negative electrode	Metal will be produced on the electrode if it is less reactive than hydrogen. Hydrogen will be produced if the metal is more reactive than hydrogen.
At the positive electrode	Oxygen is formed at positive electrode. If you have a halide ion (Cl^-, I^-, Br^-) then you will get chlorine, bromine or iodine formed at that electrode.



Ionic half equations (HT only)		
For displacement reactions	*Ionic half equations show what happens to each of the reactants during reactions*	For example: The ionic equation for the reaction between iron and copper (II) ions is: $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$ The half-equation for iron (II) is: $Fe \rightarrow Fe^{2+} + 2e^-$ The half-equation for copper (II) ions is: $Cu^{2+} + 2e^- \rightarrow Cu$

Ecological footprint

Metal ore is found in the Earth's crust and the extraction and processing of the metal ore has an impact on the environment. The **ecological footprint** helps designers and industry consider how the impact from mining for metal ore can be minimised. There are 6 aspects that form the ecological footprint. These are

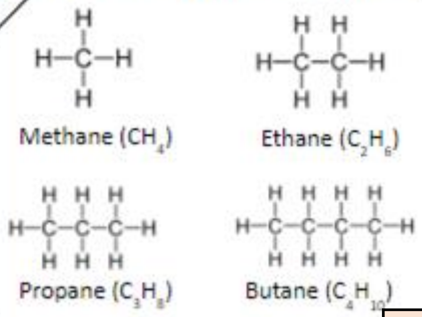
- **Sustainability** - Ore is a finite resource, which means it is **non-renewable**. Recycling metals can help ensure they are available for future generations.
- **Extraction and erosion of landscape** - As ore is found in the Earth's crust, extraction involves clearing the land so that mining can take place. This can lead to **deforestation**, soil erosion and loss of habitat for wildlife.
- **Processing** - Ore has to be removed from rocks and other minerals, which requires a lot of energy to process.
- **Transportation** - Ore is usually found in remote areas that have limited or no transport links. A transport infrastructure often needs to be built to allow heavy goods vehicles to transport the ore to a processing site.
- **Wastage** - Processing leads to a large amount of waste material that often need to be processed again before being disposed of.
- **Pollution** - Extraction of the ore involves the use of heavy equipment that produces noise pollution. The heavy goods vehicles used to transport the ore release carbon dioxide and other gases into the environment.

The law of conservation of mass states that no **atoms** are lost or made in a chemical reaction. Instead, the atoms join together in different ways to form **products**.

Crude oil, hydrocarbons and alkanes

Crude oil	***A finite resource***	Consisting mainly of plankton that was buried in the mud, crude oil is the remains of ancient biomass.
Hydrocarbons	***These make up the majority of the compounds in crude oil***	Most of these hydrocarbons are called alkanes.
General formula for alkanes	C_nH_{2n+2}	For example: C_2H_6 C_6H_{14}

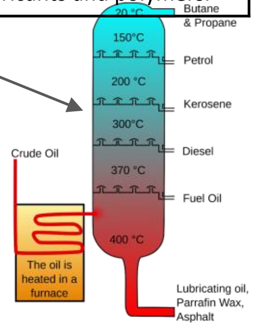
Display formula for first four alkanes



Hydrocarbon chains	***In oil***	Hydrocarbon chains in crude oil come in lots of different lengths.
	Boiling points	The boiling point of the chain depends on its length. During fractional distillation, they boil and separate at different temperatures due to this.

Cracking	*The breaking down of long chain hydrocarbons into smaller chains*	The smaller chains are more useful. Cracking can be done by various methods including catalytic cracking and steam cracking.
Catalytic cracking	*The heavy fraction is heated until vaporised*	After vaporisation, the vapour is passed over a hot catalyst forming smaller, more useful hydrocarbons.
Steam cracking	*The heavy fraction is heated until vaporised*	After vaporisation, the vapour is mixed with steam and heated to a very high temperature forming smaller, more useful hydrocarbons.

Fractions	***The hydrocarbons in crude oil can be split into fractions***	Each fraction contains molecules with a similar number of carbon atoms in them. The process used to do this is called fractional distillation.
Using fractions	***Fractions can be processed to produce fuels and feedstock for petrochemical industry***	We depend on many of these fuels; petrol, diesel and kerosene. Many useful materials are made by the petrochemical industry; solvents, lubricants and polymers.



Alkanes to alkenes	*Long chain alkanes are cracked into short chain alkenes.*
Alkenes	*Alkenes are hydrocarbons with a double bond (some are formed during the cracking process).*
Properties of alkenes	*Alkenes are more reactive than alkanes and react with bromine water. Bromine water changes from orange to colourless in the presence of alkenes.*

Combustion	During the complete combustion of hydrocarbons, the carbon and hydrogen in the fuels are oxidised, releasing carbon dioxide, water and energy.

Decane $\rightarrow$ pentane + propene + ethane

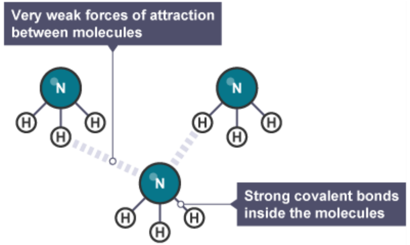
$C_{10}H_{22} \rightarrow C_5H_{12} + C_3H_6 + C_2H_4$

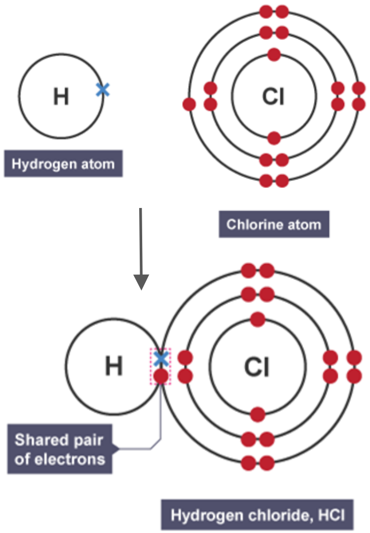
Complete combustion of methane:

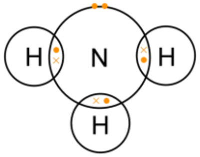
Methane + oxygen $\rightarrow$ carbon dioxide + water + energy

$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2\ H_2O\,(l)$

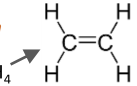
Covalent bonding - forms when two non-metal atoms share a pair of electrons.





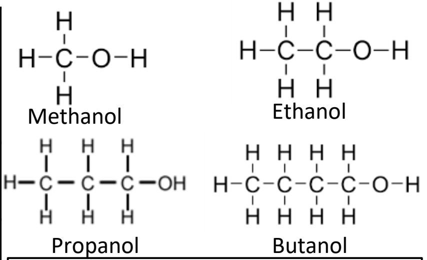


Dot and cross :
+ Show which atom the electrons in the bonds come from
- All electrons are identical

Alkenes	*Hydrocarbons with a double carbon-carbon bond.*
Unsaturated	*Alkenes are unsaturated because they contain two fewer hydrogen atoms than their alkane counterparts.*
General formula for alkenes	*C_nH_{2n}* Ethene C_2H_4 

Functional group	*Alkenes are hydrocarbons in the functional group C=C.*	The functional group of an organic compound determined their reactions.
Alkene reactions	*Alkenes react with oxygen in the same way as other hydrocarbons, just with a smoky flame due to incomplete combustion.*	Alkenes also react with hydrogen, water and the halogens. The C=C bond allows for the addition of other atoms.

Functional group	*-COOH* *For example: CH_3COOH*	Methanoic acid, ethanoic acid, propanoic acid and butanoic acid are the first four of the homologous series.
Carboxylic acid reactions	*Carboxylic acids react with carbonates, water and alcohols.*	Carboxylic acids and carbonates: These acids are neutralised by carbonates
		Carboxylic acids and water: These acids dissolve in water.
		Carboxylic acids and alcohols: The acids react with alcohols to form esters.
Strength (HT only)	*Carboxylic acids are weak acids*	Carboxylic acids only partially ionise in water. An aqueous solution of a weak acid with have a high pH (but still below 7).



The empirical formula of a **compound** is the **simplest whole number ratio** of **atoms** of each **element** in the compound. It is determined using data from experiments and therefore **empirical**.

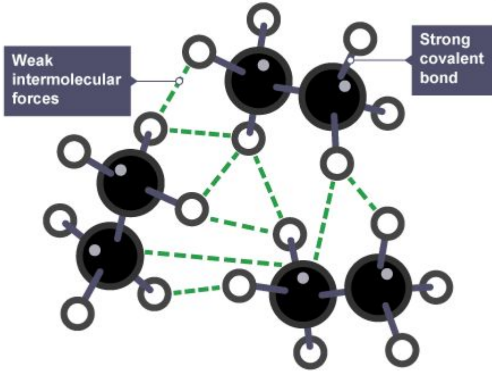
For example, the **molecular formula** of glucose is $C_6H_{12}O_6$ but the empirical formula is CH_2O. This is because we can divide each number in $C_6H_{12}O_6$ by 6 to make a simpler whole number ratio.

M_r	*The sum of the relative atomic masses of the atoms in the numbers shown in the formula*	The sum of the M_r of the reactants in the quantities shown equals the sum of the M_r of the products in the quantities shown.	$2Mg + O_2$ → $2MgO$ 48g + 32g = 80g 80g = 80g

C4 - Material choices

Composite materials	*A mixture of materials put together for a specific purpose e.g. strength*	Soda-lime glass, made by heating sand, sodium carbonate and limestone.
		Borosilicate glass, made from sand and boron trioxide, melts at higher temperatures than soda-lime glass.
		MDF wood (woodchips, shavings, sawdust and resin)
		Concrete (cement, sand and gravel)
Ceramic materials	*Made from clay*	Made by shaping wet clay and then heating in a furnace, common examples include pottery and bricks.
Polymers	*Many monomers can make polymers*	These factors affect the properties of the polymer. Low density (LD) polymers and high density (HD) polymers are produced from ethene. These are formed under different conditions.

Polymers	*Thermosetting*	polymers that do not melt when they are heated.
	Thermosoftening	polymers that melt when they are heated.



Metals can **oxidise** in air. They react with oxygen and form metal oxides. Sodium, for example, is a very **reactive** metal. When sodium is cut or scratched, its freshly exposed shiny surface rapidly turns dull as a thin layer of sodium oxide forms:

Corrosion

sodium + oxygen → sodium oxide

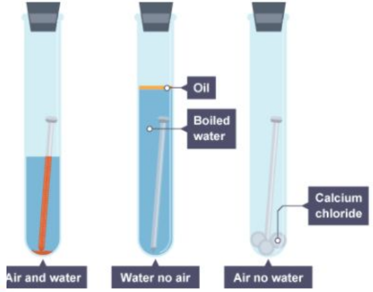
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

Other metals may oxidise more slowly. Gold and other very unreactive metals do not oxidise in air at all.

Corrosion happens when a metal continues to oxidise. The metal becomes weaker over time, and eventually all of it may become metal oxide.

A rusting experiment

The experiment in the diagram shows that both oxygen and water are needed for rusting to happen.



Calcium chloride (in the right-hand test tube) absorbs water

Working properties:

- **strength** - the ability of a material to withstand compression, tension and **shear**, eg in woven fabrics cotton isn't as strong as wool when pulled
- **hardness** - the ability to withstand **impact** without damage, eg pine is easier to dent with an impact than oak; therefore, oak is harder
- **toughness** - materials that are hard to break or snap are tough and can absorb shock, eg Kevlar in bulletproof vests is a very tough material
- **malleability** - being able to bend or shape easily would make a material easily malleable, eg sheet metal such as steel or silver is malleable and can be hammered into shape
- **ductility** - materials that can be stretched are ductile, eg pulling copper into wire shows it is ductile
- **elasticity** - the ability to be stretched and then return to its original shape, eg elastane in swimming costumes is a highly elastic material

Physical properties:

- **absorbency** - the ability to soak up moisture, light or heat, eg natural materials (such as cotton or paper) tend to be more absorbent than man-made materials (such as acrylic or polystyrene)
- **density** - how solid a material is. This is measured by dividing mass (grams) by volume (cm^3), eg lead is a dense material
- **fusibility** - the ability of a material to be heated and joined to another material when cooled, eg webbing is fusible and can be ironed onto fabrics
- **electrical conductivity** - the ability to conduct electricity, eg copper is a good conductor of electricity
- **thermal conductivity** - the ability to conduct heat, eg steel is a good heat conductor, whereas pine is not

Redox reactions - Higher

In terms of electrons:

- **oxidation** is loss of **electrons**
- **reduction** is gain of electrons

Rusting is a complex process. The example below show why both water and oxygen are needed for rusting to occur. They are interesting examples of oxidation, reduction and the use of half equations:

1. iron loses electrons and is oxidised to iron(II) ions: $Fe \rightarrow Fe^{2+} + 2e^-$
2. oxygen gains electrons in the presence of water and is reduced: $½O_2 + 2e^- + H_2O \rightarrow 2OH^-$
3. iron(II) ions lose electrons and are oxidised to iron(III) ions by oxygen: $2Fe^{2+} + ½O_2 \rightarrow 2Fe^{3+} + O^{2-}$

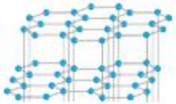
C4 - Material choices

Giant Covalent Structure – Diamond

Each **carbon** atom is **bonded** to **four** other carbon atoms, making diamond very strong. Diamond has a high melting and boiling point. **Large** amounts of **energy** are needed to break the strong covalent bonds between each carbon atom. Diamond **does not conduct** electricity because it has **no free electrons**.

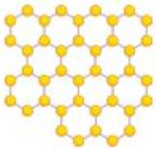
Silicon dioxide (silicon and oxygen atoms) has a similar structure to that of diamond, in that its atoms are held together by **strong covalent bonds**. Large amounts of energy are needed to break the strong covalent bonds therefore silicon dioxide, like diamond, has a high melting and boiling point.

Giant Covalent Structure – Graphite



Graphite is made up of layers of **carbon** arranged in **hexagons**. Each carbon is bonded to **three** other carbons and has **one free delocalised electron** that is able to move between the layers. The layers are held together by weak intermolecular forces. The layers of carbon can slide over each other easily as there are no strong covalent bonds between the layers. Graphite has a high melting point because a lot of energy is needed to break the covalent bonds between the carbon atoms. Graphite can **conduct** electricity.

Giant Covalent Structure – Graphene



Graphene is one layer of graphite. It is very **strong** because of the covalent bonds between the carbon atoms. As with graphite, each carbon in graphene is bonded to three others with one **free delocalised electron**. Graphene is able to **conduct electricity**. Graphene, when added to other materials, can make them even stronger. Useful in electricals and composites.

Nanoscience

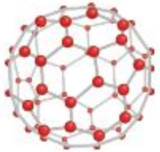
Nanoscience refers to structures that are **1–100nm** in size, of the order of a few hundred atoms. Nanoparticles have a **high surface area to volume ratio**. This means that smaller amounts are needed in comparison to normal sized particles. As the side length of a cube decreases by a factor of 10, the surface area to volume ratio increases approximately

Name of Particle	Diameter
nanoparticle	1–100nm
fine particles ($PM_{2.5}$)	100–2500nm
coarse particles (PM_{10})	2500–10000nm

Polymers

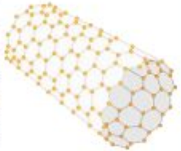
Polymers are long chain molecules that are made up of many smaller units called **monomers**. Atoms in a polymer chain are held together by **strong covalent bonds**. Between polymer molecules, there are **intermolecular forces**. Intermolecular forces **attract** polymer chains towards each other. Longer polymer chains have stronger forces of attraction than shorter ones therefore making stronger materials.

Fullerenes and Nanotubes



Molecules of carbon that are shaped like hollow tubes or balls, arranged in hexagons of five or seven carbon atoms. They can be used to **deliver drugs into the body**.

Buckminsterfullerene has the formula C_{60}



Carbon Nanotubes are tiny carbon cylinders that are very long compared to their width. Nanotubes can conduct electricity as well as strengthening materials without adding much weight. The properties of carbon nanotubes make them useful in electronics and nanotechnology.

Possible Risks of Nanoparticles

As nanoparticles are so **small**, it makes it possible for them to be inhaled and enter the lungs. Once inside the body, nanoparticles may **initiate harmful reactions** and toxic substances could bind to them because of their large surface area to volume ratio. Nanoparticles have many applications. These include medicine, cosmetics, sun creams and deodorants. They can also be used as catalysts.

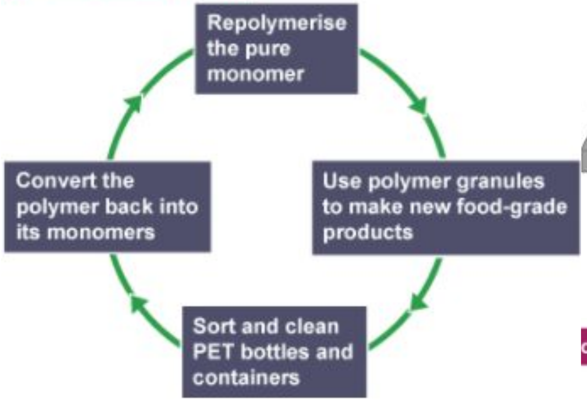
Modern nanoparticles are a relatively new phenomenon therefore it is difficult for scientists to truly determine the risks associated with them.

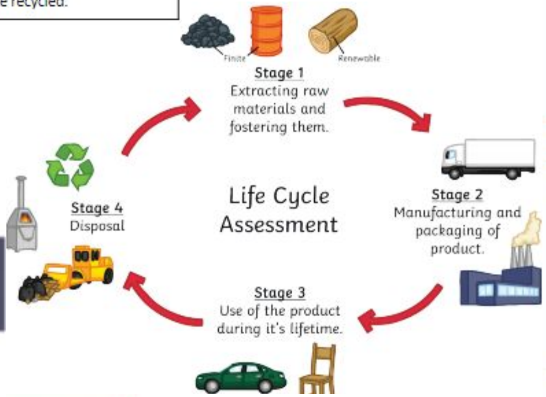
Unit	Decimal	Standard form
1 cm	0.01 m	1×10^{-2} m
1 mm	0.001 m	1×10^{-3} m
1 µm	0.000,001 m	1×10^{-6} m
1 nm	0.000,000,001 m	1×10^{-9} m

C4 - Material choices

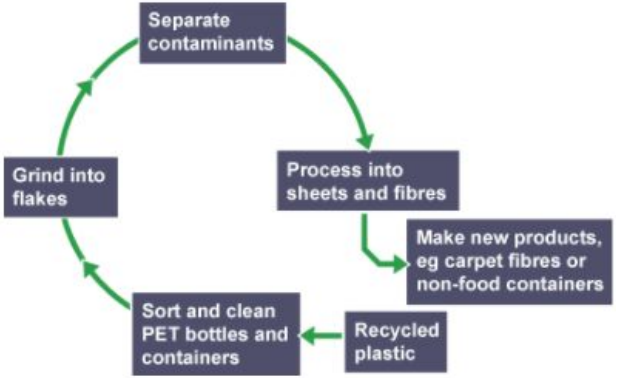
Reduce, reuse and recycle	*This strategy reduces the use of limited resources*	This, therefore, reduces energy sources being used, reduces waste (landfill) and reduces environmental impacts.
Limited raw materials	*Used for metals, glass, building materials, plastics and clay ceramics*	Most of the energy required for these processes comes from limited resources. Obtaining raw materials from the Earth by quarrying and mining causes environmental impacts.
Reusing and recycling	*Metals can be recycled by melting and recasting/reforming*	Glass bottles can be reused. They are crushed and melted to make different glass products. Products that cannot be reused are recycled.

Closed-loop recycling





op recycling



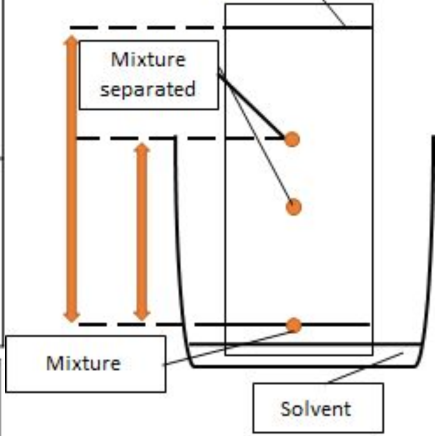
LCAS	*Life cycle assessments are carried out to assess the environmental impact of products*	They are assessed at these stages: - Extraction and processing raw materials - Manufacturing and packaging - Use and operation during lifetime - Disposal
Values	*Allocating numerical values to pollutant effects is difficult*	Value judgments are allocated to the effects of pollutants so LCA is not a purely objective process.

Stage of Life Cycle	Plastic Bag	Paper Bag
Stage 1 – raw material	Uses a finite resource (crude oil). The processes of fractional distillation, cracking and polymerisation all require energy to make crude oil useful.	Made from trees/recycled paper. Making paper from trees requires more energy than recycled paper because trees have to be chopped down. Still uses less energy than making plastics from crude oil.
Stage 2 – manufacture	Cheap to make.	More expensive to make.
Stage 3 – use	Plastic bags have a low environmental impact as they can be used a number of times. In comparison to paper bags, they are much stronger.	Paper bags can only be reused a limited number of times and so have a short lifetime.
Stage 4 – disposal	The downside to plastic bags is that they do not biodegrade easily in landfill. Recycling options are available. If they are not disposed of correctly, plastic bags can have a detrimental impact on the environment and animal habitats.	Paper bags biodegrade easily in landfill sites.

C5 - Chemical analysis

Chromatography

Chromatography	*Can be used to separate mixtures and help identify substances.*	Involves a mobile phase (e.g. water or ethanol) and a stationary phase (e.g. chromatography paper).
R_f Values	*The ratio of the distance moved by a compound to the distance moved by solvent.*	$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$
Pure substances	*The compounds in a mixture separate into different spots.*	This depends on the solvent used. A pure substance will produce a single spot in all solvents whereas an impure substance will produce multiple spots.



Pure substances	*A pure substances is a single element or compound, not mixed with any other substance.*	Pure substances melt and boil at specific temperatures. Heating graphs can be used to distinguish pure substances from impure.

Melting point of a pure substance

Melting point of an impure substance

Flame tests

Element	Colour flames
Lithium	*Crimson*
Sodium	*Yellow*
Potassium	*Lilac*
Calcium	*Orange-red*
Copper	*Green*

Identification of ions

Metal hydroxides (chem only)

Sodium hydroxide	*Is added to solutions to identify metal ions.*
White precipitates	*Aluminium, calcium and magnesium ions form this with sodium hydroxide solution.*
Coloured precipitates	*Copper (II) = blue Iron (II) = green Iron (III) = brown*

Carbonates, halides and sulfates

Carbonates	*React with dilute acids to form carbon dioxide.*
Halide ions	*When in a solution, they produce precipitates with silver nitrate solution in the presence of nitric acid.*
Sulfate ions	*When in a solutions they produce a white precipitate with barium chloride solutions in the presence of hydrochloric acid.*

Instrumental methods

Instrumental methods	*Methods that rely on machines*	Can be used to identify elements and compounds. These methods are accurate, sensitive and rapid.

Flame emission spectroscopy

Flame emission spectroscopy	*An instrumental method used to analyse metal ions.*	The sample solution is put into a flame and the light that is given out is put through a spectroscope. The output line spectrum, can be analysed to identify the metal ions in the solution. It can also be used to measure concentrations.

Chemical amounts are measured in moles (mol)	*Mass of one mole of a substance in grams = relative formula mass*	One mole of H_2O = 18g (1 + 1 + 16) One mole of Mg = 24g

Avogadro constant	*One mole of any substance will contain the same number of particles, atoms, molecules or ions.*	6.02×10^{23} per mole One mole of H_2O will contain 6.02×10^{23} molecules One mole of NaCl will contain 6.02×10^{23} Na^+ ions

Number of moles = $\frac{\text{mass (g)}}{A_r}$ *or* $\frac{\text{mass (g)}}{M_r}$	How many moles of sulfuric acid molecules are there in 4.7g of sulfuric acid (H_2SO_4)? Give your answer to 1 significant figure. $\frac{4.7}{98} = 0.05$ mol 98 ← (M_r of H_2SO_4)

Relative formula mass has the symbol, M_r. To calculate the M_r for a substance:

1. work out how many atoms of each element there are in the chemical formula
2. add together the A_r values for all the atoms of each element present

For example, the formula for carbon dioxide is CO_2. It consists of one carbon atom (A_r = 12) and two oxygen atoms (A_r = 16):

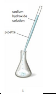
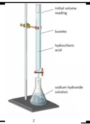
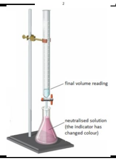
M_r of CO_2 = 12 + 16 + 16 = 44

Concentration of solutions

Measured in mass per given volume of solution (g/dm^3)	*Conc.* = $\frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$	HT only Greater mass = higher concentration. Greater volume = lower concentration.

Titration	*If the volumes of two solutions that react completely are known and the concentrations of one solution is known, the concentration of the other solution can be calculated.*	$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$ It takes 12.20cm^3 of sulfuric acid to neutralise 24.00cm^3 of sodium hydroxide solution, which has a concentration of 0.50mol/dm^3. Calculate the concentration of the sulfuric acid in mol/dm^3: 0.5 mol/dm^3 x (24/1000) dm^3 = 0.012 mol of NaOH The equation shows that 2 mol of NaOH reacts with 1 mol of H_2SO_4, so the number of moles in 12.20cm^3 of sulfuric acid is (0.012/2) = 0.006 mol of sulfuric acid Calculate the concentration of sulfuric acid in mol/ dm^3 0.006 mol x (1000/12.2) dm^3 =0.49mol/dm^3

Calculate the concentration of sulfuric acid in g/ dm^3:
H_2SO_4 = (2x1) + 32 + (4x16) = 98g
0.49 x 98g = 48.2g/dm^3

	1. Use the pipette to add 25 cm^3 of alkali to a conical flask and add a few drops of indicator.
	2. Fill the burette with acid and note the starting volume. Slowly add the acid from the burette to the alkali in the conical flask, swirling to mix.
	3. Stop adding the acid when the end-point is reached (the appropriate colour change in the indicator happens). Note the final volume reading. Repeat steps 1 to 3 until you get consistent readings.

Reactions of acids

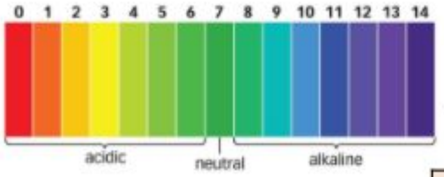
Strong and weak acids (HT ONLY)

Strong acids	*Completely ionised in aqueous solutions e.g. hydrochloric, nitric and sulfuric acids.*
Weak acids	*Only partially ionised in aqueous solutions e.g. ethanoic acid, citric acid.*
Hydrogen ion concentration	*As the pH decreases by one unit (becoming a stronger acid), the hydrogen ion concentration increases by a factor of 10.*

Soluble salts

Soluble salts	*Soluble salts can be made from reacting acids with solid insoluble substances (e.g. metals, metal oxides, hydroxides and carbonates).*
Production of soluble salts	*Add the solid to the acid until no more dissolves. Filter off excess solid and then crystallise to produce solid salts.*

The pH scale and neutralisation



You can use universal indicator or a pH probe to measure the acidity or alkalinity of a solution against the pH scale.

Acids	*Acids produce hydrogen ions (H^+) in aqueous solutions.*
Alkalis	*Aqueous solutions of alkalis contain hydroxide ions (OH^-).*

In neutralisation reactions, hydrogen ions react with hydroxide ions to produce water:

$$H^+ + OH^- \rightarrow H_2O$$

Examples of purification techniques

Mixture	Technique	Reason
Ethanol and water	Fractional distillation	Ethanol is a liquid which dissolves in water
Water from sea water	Simple distillation	Water is a solvent for a number of dissolved solids in sea water
Copper(II) sulfate from copper(II) sulfate solution	Crystallisation	Copper(II) sulfate is a dissolved solid
Chlorophyll from plant leaves	Chromatography	Chlorophyll is one of a number of dissolved pigments in plant leaves
Sand from sea water	Filtration	Sand is an insoluble solid in sea water

Choosing purification techniques

The method used to separate a **mixture** will depend on the **properties** of the substances in the mixture.

Insoluble solid and a liquid

The solid is removed from the liquid using **filtration**. The filter paper traps the particles of **insoluble** solid but lets the liquid through.

Soluble solid from a solution

The solution is warmed to **evaporate** the **solvent**. Once the solution is concentrated, it is left to cool. This allows the **soluble** solid to form by **crystallisation**.

Solvent from a solution

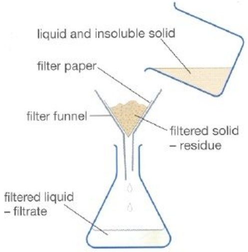
The solvent is removed using **simple distillation**. As the solvent is distilled off, the dissolved solid remains behind.

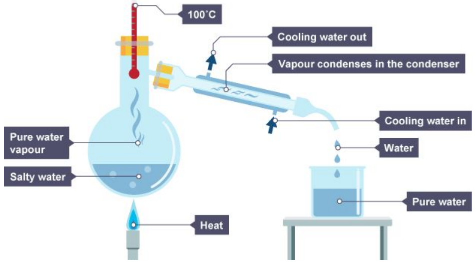
A mixture of liquids

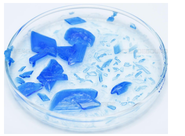
The liquids are heated and, because they have different **boiling points**, they will evaporate at different temperatures. This process is **fractional distillation**.

A mixture of soluble solids from a solution

Paper **chromatography** is used to separate mixtures of soluble substances. These are often coloured substances such as food colourings, inks, dyes or plant pigments.







Reactions of acids and bases

Strong and weak acids (HT ONLY)

Strong acids	*Completely ionised in aqueous solutions e.g. hydrochloric, nitric and sulfuric acids.*
Weak acids	*Only partially ionised in aqueous solutions e.g. ethanoic acid, citric acid.*
Hydrogen ion concentration	*As the pH decreases by one unit (becoming a stronger acid), the hydrogen ion concentration increases by a factor of 10.*

Acids

Hydrochloric acid	**Formula:** HCl **Hydrogen ions formed:** 1 **Anion formed:** Chloride, Cl^-
***Nitric acid**	**Formula:** HNO_3 **Hydrogen ions formed:** 1 **Anion formed:** Nitrate, NO_3^-
***Sulfuric acid**	**Formula:** H_2SO_4 **Hydrogen ions formed:** 2 **Anion formed:** Sulfate, SO_4^{2-}
*****Ions and pH**	The higher the hydrogen ion concentration the lower the pH, the higher the hydroxide ion concentration, the higher the pH.

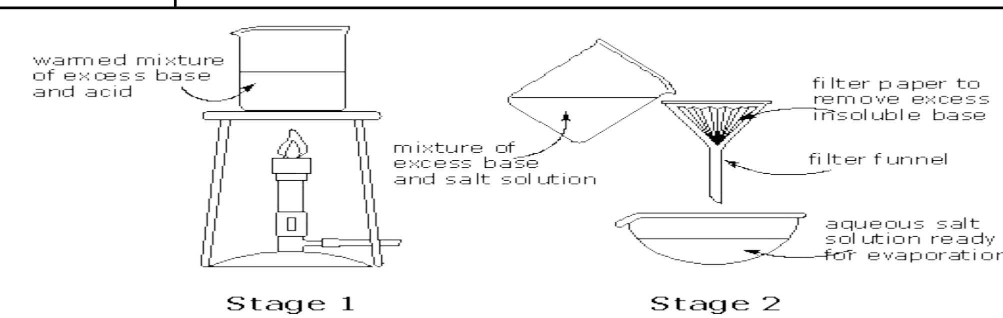
Bases

5. Alkalis and balancing equations

****Bases and alkalis**	A base is a substance that neutralises an acid to form a salt and water. An alkali is a base that is soluble in water.
***Common alkalis**	Sodium hydroxide, NaOH Potassium hydroxide, KOH Calcium hydroxide, $Ca(OH)_2$
***Reaction of alkalis with acids**	Acid + alkali → salt + water Eg: Sodium hydroxide nitric acid → sodium nitrate + water $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

3. Bases and salts

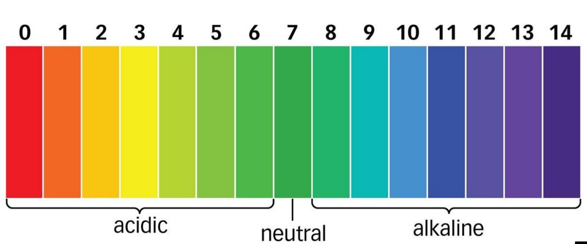
*Base	A substance that neutralises an acid to form a salt and water.
*Salt	A compound formed from the metal cation of a base and the non-metal anion of an alkali.
*Naming salts	Two-part names. First part = the metal from the base, second part = the anion from the acid.
*Acids and their anions	Sulfuric acid □ sulfate Nitric acid □ nitrate Hydrochloric acid □ chloride
**Reaction of metal oxides with acid	Metal oxide + acid □ salt + water E.g. Magnesium oxide + hydrochloric acid □ magnesium chloride + water $MgO(s) + 2HCl(aq)$ □ $MgCl_2(aq) + H_2O(l)$
*Preparing soluble salts	- Gently warm a beaker of acid - Add a spatula of metal oxide and stir until dissolved - Repeat until it no longer dissolves - Filter to remove excess oxide - Allow water to evaporate to produce pure crystals



Soluble salts

Soluble salts	*Soluble salts can be made from reacting acids with solid insoluble substances (e.g. metals, metal oxides, hydroxides and carbonates).*
Production of soluble salts	*Add the solid to the acid until no more dissolves. Filter off excess solid and then crystallise to produce solid salts.*

The pH scale and neutralisation



You can use universal indicator or a pH probe to measure the acidity or alkalinity of a solution against the pH scale.

In neutralisation reactions, hydrogen ions react with hydroxide ions to produce water:

$H^+ + OH^-$ □ H_2O

Acids	*Acids produce hydrogen ions (H^+) in aqueous solutions.*
Alkalis	*Aqueous solutions of alkalis contain hydroxide ions (OH^-).*

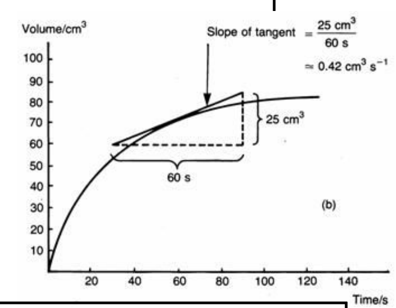
GCSE The rate and extent of chemical change

Rate of reaction

Calculating rates of reactions

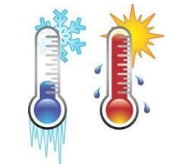
Rate of chemical reaction	***This can be calculated by measuring the quantity of reactant used or product formed in a given time.***	Rate = quantity of reactant used / time taken Rate = quantity of product formed / time taken

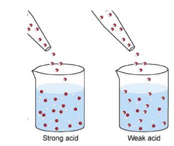
Quantity	Unit
Mass	Grams (g)
Volume	cm^3
Rate of reaction	Grams per cm^3 (g/cm^3) HT: moles per second (mol/s)

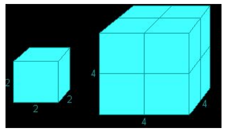


Factors affecting rates

Factors affecting the rate of reaction	
Temperature	***The higher the temperature, the quicker the rate of reaction.***
Concentration	***The higher the concentration, the quicker the rate of reaction.***
Surface area	***The larger the surface area of a reactant solid, the quicker the rate of reaction.***
Pressure (of gases)	***When gases react, the higher the pressure upon them, the quicker the rate of reaction.***



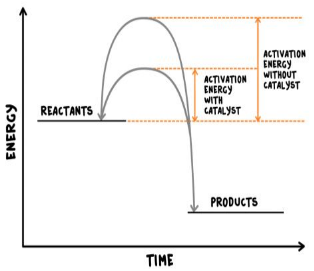




Collision theory and activation energy

Collision theory	***Chemical reactions can only occur when reacting particles collide with each other with sufficient energy.***	Increasing the temperature increases the frequency of collisions and makes the collisions more energetic, therefore increasing the rate of reaction.
Activation energy	***This is the minimum amount of energy colliding particles in a reaction need in order to react.***	Increasing the concentration, pressure (gases) and surface area (solids) of reactions increases the frequency of collisions, therefore increasing the rate of reaction.

Catalysts



If a catalyst is used in a reaction, it is not shown in the word equation.

Catalyst	A catalyst changes the rate of a chemical reaction but is not used in the reaction.
Enzymes	These are biological catalysts.
How do they work?	Catalysts provide a different reaction pathway where reactants do not require as much energy to react when they collide.

Reversible reactions and dynamic equilibrium

Reversible reactions

Reversible reactions	In some chemical reactions, the products can react again to re-form the reactants.
Representing reversible reactions	$A + B \rightleftharpoons C + D$
The direction	The direction of reversible reactions can be changed by changing conditions: $A + B \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} C + D$

Energy changes and reversible reactions

If one direction of a reversible reaction is exothermic, the opposite direction is endothermic. The same amount of energy is transferred in each case.

For example:
Hydrated copper sulfate $\underset{\text{exothermic}}{\overset{\text{endothermic}}{\rightleftharpoons}}$ Anhydrous copper sulfate + Water

Equilibrium

Equilibrium in reversible reactions	When a reversible reaction occurs in apparatus which prevents the escape of reactants and products, equilibrium is reached when the forward and reverse reactions occur exactly at the same rate.

Changing conditions and equilibrium (HT)

The relative amounts of reactants and products at equilibrium depend on the conditions of the reaction.

Le Chatelier's Principles	States that when a system experiences a disturbance (change in condition), it will respond to restore a new equilibrium state.
Changing concentration	If the concentration of a reactant is increased, more products will be formed . If the concentration of a product is decreased, more reactants will react.
Changing temperature	If the temperature of a system at equilibrium is increased: - Exothermic reaction = products decrease - Endothermic reaction = products increase
Changing pressure (gaseous reactions)	For a gaseous system at equilibrium: - Pressure increase = equilibrium position shifts to side of equation with smaller number of molecules. - Pressure decrease = equilibrium position shifts to side of equation with larger number of molecules.

The Haber process and the use of NPK fertilisers

Atom economy Triple

A measure of the amount of starting materials that end up as useful products	Atom economy = $\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula mass of all reactants from equation}} \times 100$	High atom economy is important or sustainable development and economic reasons

Calculate the atom economy for making hydrogen by reacting zinc with hydrochloric acid:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

M_r of H_2 = 1 + 1 = 2

M_r of Zn + 2HCl = 65 + 1 + 1 + 35.5 + 35.5 = 138

Atom economy = $\frac{2}{138} \times 100$

= $\frac{2}{138} \times 100$ = 1.45%

This method is unlikely to be chosen as it has a low atom economy.

HT only:

200g of calcium carbonate is heated. It decomposes to make calcium oxide and carbon dioxide. *Calculate the theoretical mass of calcium oxide made.*

$CaCO_3$ □ $CaO + CO_2$

M_r of $CaCO_3$ = 40 + 12 + (16x3) = 100

M_r of CaO = 40 + 16 = 56

100g of $CaCO_3$ would make 56 g of CaO

So 200g would make 112g

Percentage yield

Percentage yield is comparing the amount of product obtained as a percentage of the maximum theoretical amount	*% Yield = $\frac{\text{Mass of product made}}{\text{Max. theoretical mass}} \times 100$*	A piece of sodium metal is heated in chlorine gas. A maximum theoretical mass of 10g for sodium chloride was calculated, but the actual yield was only 8g. *Calculate the percentage yield.* Percentage yield = 8/10 x 100 =80%

Production and uses of NPK fertilisers

NPK fertilisers	*These contain nitrogen, phosphorous and potassium*	Formulations of various salts containing appropriate percentages of the elements.
Fertiliser examples	*Potassium chloride, potassium sulfate and phosphate rock are obtained by mining*	Phosphate rock needs to be treated with an acid to produce a soluble salt which is then used as a fertiliser. Ammonia can be used to manufacture ammonium salts and nitric acid.

Phosphate rock

Treatment	*Products*
Nitric acid	*The acid is neutralised with ammonia to produce ammonium phosphate, a NPK fertiliser.*
Sulfuric acid	*Calcium phosphate and calcium sulfate (a single superphosphate).*
Phosphoric acid	*Calcium phosphate (a triple superphosphate).*

The Haber process

The Haber process – conditions and equilibrium HIGHER

Pressure	*The reactants side of the equation has more molecules of gas. This means that if pressure is increased, equilibrium shifts towards the production of ammonia (Le Chatelier's principle). The pressure needs to be as high as possible.*
Temperature	*The forward reaction is exothermic. Decreasing temperature increases ammonia production at equilibrium. The exothermic reaction that occurs releases energy to surrounding, opposing the temperature decreases. Too low though and collisions would be too infrequent to be financially viable.*

The Haber process	*Used to manufacture ammonia*	***Ammonia is used to produce fertilisers*** Nitrogen + hydrogen $\rightleftharpoons$ ammonia
Raw materials	*Nitrogen from the air while hydrogen from natural gas*	Both of these gases are purified before being passed over an iron catalyst. This is completed under high temperature (about 450°C) and pressure (about 200 atmospheres).
Catalyst	*Iron*	The catalyst speeds up **both** directions of the reaction, therefore not actually increasing the amount of valuable product.