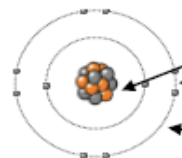


Atoms, elements and compounds

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



Central nucleus	Contains protons and neutrons
Electron shells	Contains electrons

Name of Particle	Relative Charge	Relative Mass	Electronic shell	Max number of electrons
Proton	+1	1	1	2
Neutron	0	1	2	8
Electron	-1	Very small	3	8
			4	2

Electronic structures
Relative electrical charges of subatomic particles

7 Li 3	Mass number	<i>The sum of the protons and neutrons in the nucleus</i>	
	Atomic number	<i>The number of protons in the atom</i>	Number of electrons = number of protons

AQA Chemistry
C1: Atomic Structure & The Periodic Table

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

The development of the model of the atom

James Chadwick	<i>Provided the evidence to show the existence of neutrons within the nucleus</i>
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Rutherford's scattering experiment	<i>A beam of alpha particles are directed at a very thin gold foil</i>		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.

Mixtures	<i>Two or more elements or compounds not chemically combined together</i>	Can be separated by physical processes.
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Chemical equations	<i>Show chemical reactions - need reactant(s) and product(s) energy always involves and energy change</i>	Law of conservation of mass states the total mass of products = the total mass of reactants.
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Method	Description	Example
Filtration	<i>Separating an insoluble solid from a liquid</i>	To get sand from a mixture of sand, salt and water.
Crystallisation	<i>To separate a solid from a solution</i>	To obtain pure crystals of sodium chloride from salt water.
Simple distillation	<i>To separate a solvent from a solution</i>	To get pure water from salt water.
Fractional distillation	<i>Separating a mixture of liquids each with different boiling points</i>	To separate the different compounds in crude oil.
Chromatography	<i>Separating substances that move at different rates through a medium</i>	To separate out the dyes in food colouring.

Word equations	<i>Uses words to show reaction</i> reactants → products magnesium + oxygen → magnesium oxide	Does not show what is happening to the atoms or the number of atoms.
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Symbol equations	<i>Uses symbols to show reaction</i> reactants → products $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$	Shows the number of atoms and molecules in the reaction, these need to be balanced.
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Relative atomic mass

Isotopes	<i>Atoms of the same element with the same number of protons and different numbers of neutrons</i>	^{35}Cl (75%) and ^{37}Cl (25%) Relative abundance = $(\% \text{ isotope 1} \times \text{mass isotope 1}) + (\% \text{ isotope 2} \times \text{mass isotope 2}) \div 100$ e.g. $(25 \times 37) + (75 \times 35) \div 100 = 35.5$
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Alkali metals: 1, 2
 Halogens: 3, 4, 5, 6, 7
 Noble gases: 0

H	Transition metals																He						
Li	Be																	B	C	N	O	F	Ne
Na	Mg																	Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr						
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe						
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn						
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?												

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.

The Periodic table

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 to the left of this line, non metals to the right

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.

Metals and non metals

AQA Chemistry
C1: Atomic Structure & The Periodic Table

Group 7

Group 1

Alkali metals

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.

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

Group 0

Transition metals (Chemistry only)

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

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. $Cl_2 + H_2 \rightarrow 2HCl$
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. $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$

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.

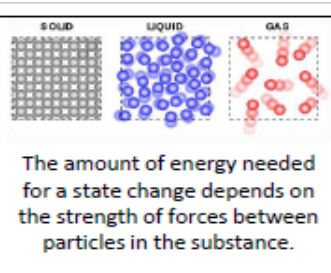
Compared to group 1	<ul style="list-style-type: none"> Less reactive Harder Denser Higher melting points 	<ul style="list-style-type: none"> Cu^{2+} is blue Ni^{2+} is pale green, used in the manufacture of margarine
Typical properties	<ul style="list-style-type: none"> Many have different ion possibilities with different charges Used as catalysts Form coloured compounds 	<ul style="list-style-type: none"> Fe^{2+} is green, used in the Haber process Fe^{3+} is reddish-brown Mn^{2+} is pale pink



Ionic	<i>Particles are oppositely charged ions</i>	Occurs in compounds formed from metals combined with non metals.
Covalent	<i>Particles are atoms that share pairs of electrons</i>	Occurs in most non metallic elements and in compounds of non metals.
Metallic	<i>Particles are atoms which share delocalised electrons</i>	Occurs in metallic elements and alloys.

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

<i>High melting and boiling points</i>	Large amounts of energy needed to break the bonds.
<i>Do not conduct electricity when solid</i>	Ions are held in a fixed position in the lattice and cannot move.
<i>Do conduct electricity when molten or dissolved</i>	Lattice breaks apart and the ions are free to move.

Properties of ionic compounds

Chemical bonds

The three states of matter

Good conductors of electricity

Delocalised electrons carry electrical charge through the metal.

Good conductors of thermal energy

Energy is transferred by the delocalised electrons.

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.

AQA Chemistry
C2: Bonding, Structure & The Properties of Matter

Ionic bonding

Metals as conductors

Properties of metals and alloys

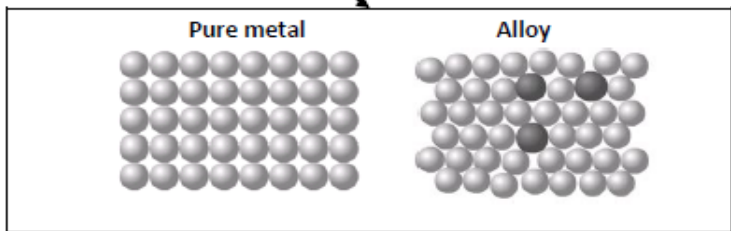
Alloys

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.

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

Metallic bonding



Dot and cross diagram

(2, 8, 1) (2, 8, 7) → (2, 8) (2, 8, 8)

Giant structure

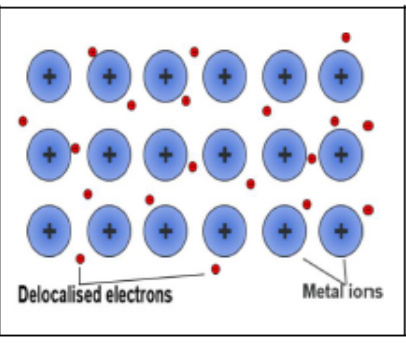
● Na⁺ ● Cl⁻

Ionic compounds

Structure

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

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.



M_r	<i>The sum of the relative atomic masses of the atoms in the numbers shown in the formula</i>	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.	$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ $\downarrow \quad \downarrow \quad \downarrow$ $48\text{g} + 32\text{g} = 80\text{g}$ $80\text{g} = 80\text{g}$
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The reactant that is completely used up	<i>Limits the amount of product that is made</i>	Less moles of product are made.
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Mass appears to increase during a reaction	<i>One of the reactants is a gas</i>	Magnesium + oxygen \rightarrow magnesium oxide
Mass appears to decrease during a reaction	<i>One of the products is a gas and has escaped</i>	Calcium carbonate \rightarrow carbon dioxide + calcium oxide

Mass changes when a reactant or product is a gas

Relative formula mass (M_r)

Limiting reactants (HT only)

Chemical measurements

Whenever a measurement is taken, there is always some uncertainty about the result obtained

Can determine whether the mean value falls within the range of uncertainty of the result

1. Calculate the mean
2. Calculate the range of the results
3. Estimate of uncertainty in mean would be half the range

Example:
 1. Mean value is 46.5s
 2. Range of results is 44s to 49s = 5s
 3. Time taken was $46.5\text{s} \pm 2.5\text{s}$

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.

Conservation of mass	<i>No atoms are lost or made during a chemical reaction</i>	Mass of the products equals the mass of the reactants.
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AQA Chemistry
C3: Quantitative Chemistry

Balanced symbol equations	<i>Represent chemical reactions and have the same number of atoms of each element on both sides of the equation</i>	$\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ <p>Subscript \swarrow \searrow Normal script</p> <p>Subscript numbers show the number of atoms of the element to its left. Normal script numbers show the number of molecules.</p>
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Conservation of mass and balanced symbol equations

Moles (HT only)

Amounts of substances in equations (HT only)

Using moles to balance equations (HT only)

The balancing numbers in a symbol equation can be calculated from the masses of reactants and products

Convert the masses in grams to amounts in moles and convert the number of moles to simple whole number ratios.

Chemical amounts are measured in moles (mol)	<i>Mass of one mole of a substance in grams = relative formula mass</i>	One mole of $\text{H}_2\text{O} = 18\text{g}$ ($1 + 1 + 16$) One mole of $\text{Mg} = 24\text{g}$
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Avogadro constant	<i>One mole of any substance will contain the same number of particles, atoms, molecules or ions.</i>	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
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<i>Number of moles = $\frac{\text{mass (g)}}{A_r}$ or $\frac{\text{mass (g)}}{M_r}$</i>	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. $\frac{4.7}{98} = 0.05 \text{ mol}$ <p style="text-align: right;">\leftarrow (M_r of H_2SO_4)</p>
---	--

Chemical equations show the number of moles reacting and the number of moles made

$\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$
One mole of magnesium reacts with two moles of hydrochloric acid to make one mole of magnesium chloride and one mole of hydrogen

If you have a 60g of Mg, what mass of HCl do you need to convert it to MgCl_2 ?

A_r : $\text{Mg} = 24$ so mass of 1 mole of $\text{Mg} = 24\text{g}$
 M_r : $\text{HCl} (1 + 35.5)$ so mass of 1 mole of $\text{HCl} = 36.5\text{g}$

So 60g of Mg is $60/24 = 2.5$ moles

Balanced symbol equation tells us that for every one mole of Mg, you need two moles of HCl to react with it.

So you need $2.5 \times 2 = 5$ moles of HCl

You will need $5 \times 36.5\text{g}$ of HCl = 182.5g



<p>A measure of the amount of starting materials that end up as useful products</p>	<p>Atom economy = $\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula mass of all reactants from equation}} \times 100$</p>	<p>High atom economy is important or sustainable development and economic reasons</p>
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Calculate the atom economy for making hydrogen by reacting zinc with hydrochloric acid:

$$\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$$

M_r of $\text{H}_2 = 1 + 1 = 2$
 M_r of $\text{Zn} + 2\text{HCl} = 65 + 1 + 1 + 35.5 + 35.5 = 138$

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.

Atom economy

Concentration of a solution is the amount of solute per volume of solution

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

What is the concentration of a solution that has 35.0g of solute in 0.5dm³ of solution?

$35/0.5 = 70 \text{ g/dm}^3$

Using concentrations of solutions in mol/dm³ (HT only, chemistry only)

AQA Chemistry
C3: Quantitative Chemistry

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.

$2\text{NaOH(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + 2\text{H}_2\text{O(l)}$

It takes 12.20cm³ of sulfuric acid to neutralise 24.00cm³ of sodium hydroxide solution, which has a concentration of 0.50mol/dm³.

Calculate the concentration of the sulfuric acid in mol/dm³:

$0.5 \text{ mol/dm}^3 \times (24/1000) \text{ dm}^3 = 0.012 \text{ mol}$ of NaOH
 The equation shows that 2 mol of NaOH reacts with 1 mol of H_2SO_4 , so the number of moles in 12.20cm³ of sulfuric acid is $(0.012/2) = 0.006 \text{ mol}$ of sulfuric acid

Calculate the concentration of sulfuric acid in mol/dm³
 $0.006 \text{ mol} \times (1000/12.2) \text{ dm}^3 = 0.49 \text{ mol/dm}^3$

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.

$$\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$$

M_r of $\text{CaCO}_3 = 40 + 12 + (16 \times 3) = 100$
 M_r of $\text{CaO} = 40 + 16 = 56$
 100g of CaCO_3 would make 56 g of CaO
 So 200g would make 112g

Percentage yield

Use of amount of substance in relation to volumes of gases (HT only, chemistry only)

Calculate the concentration of sulfuric acid in g/dm³:

$$\text{H}_2\text{SO}_4 = (2 \times 1) + 32 + (4 \times 16) = 98 \text{ g}$$

$$0.49 \times 98 = 48.2 \text{ g/dm}^3$$

<p>Yield is the amount of product obtained</p>	<p style="color: orange; font-style: italic;">It is not always possible to obtain the calculated amount of a product</p>	<p>The reaction may not go to completion because it is reversible.</p> <p>Some of the product may be lost when it is separated from the reaction mixture.</p> <p>Some of the reactants may react in ways different to the expected reaction.</p>	<p>Equal amounts of moles or gases occupy the same volume under the same conditions of temperature and pressure</p>	<p style="color: orange; font-style: italic;">The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmospheric pressure) is 24 dm³</p>	<p>No. of moles of gas = $\frac{\text{vol of gas (dm}^3\text{)}}{24 \text{ dm}^3}$</p>
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Percentage yield is comparing the amount of product obtained as a percentage of the maximum theoretical amount

$$\% \text{ Yield} = \frac{\text{Mass of product made} \times 100}{\text{Max. theoretical mass}}$$

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 = $8/10 \times 100 = 80\%$

What is the volume of 11.6 g of butane (C_4H_{10}) gas at RTP?

$M_r : (4 \times 12) + (10 \times 1) = 58$

$11.6/58 = 0.20 \text{ mol}$

Volume = $0.20 \times 24 = 4.8 \text{ dm}^3$

6g of a hydrocarbon gas had a volume of 4.8 dm³. Calculate its molecular mass.

1 mole = 24 dm³, so $4.8/24 = 0.2 \text{ mol}$

$M_r = 6 / 0.2 = 30$

If 6g = 0.2 mol, 1 mol equals 30 g



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

HT ONLY: Reactions between metals and acids are redox reactions as the metal donates electrons to the hydrogen ions. This displaces hydrogen as a gas while the metal ions are left in the solution.

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:
 $Cu^{2+} + 2e^{-} \rightarrow Cu$

Reactions with acids	$metal + acid \rightarrow metal\ salt + hydrogen$	magnesium + hydrochloric acid \rightarrow magnesium chloride + hydrogen zinc + sulfuric acid \rightarrow zinc sulfate + hydrogen
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Acids react with some metals to produce salts and hydrogen.

Reactions of acids and metals

Extraction using carbon

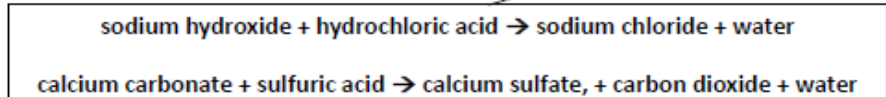
Metals less reactive than carbon can be extracted from their oxides by reduction.

For example:
zinc oxide + carbon \rightarrow zinc + carbon dioxide

Acid name	Salt name
Hydrochloric acid	Chloride
Sulfuric acid	Sulfate
Nitric acid	Nitrate

Oxidation and reduction in terms of electrons (HT ONLY)

Neutralisation of acids and salt production



Neutralisation

Acids can be neutralised by alkalis and bases

An alkali is a soluble base e.g. metal hydroxide.
A base is a substance that neutralises an acid e.g. a soluble metal hydroxide or a metal oxide.

Reactions of acids

Reactivity of metals

The reactivity series

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.

	Reactions with water	Reactions with acid
Group 1 metals	<i>Reactions get more vigorous as you go down the group</i>	<i>Reactions get more vigorous as you go down the group</i>
Group 2 metals	<i>Do not react with water</i>	<i>Observable reactions include fizzing and temperature increases</i>
Zinc, iron and copper	<i>Do not react with water</i>	<i>Zinc and iron react slowly with acid. Copper does not react with acid.</i>

Metals form positive ions when they react	<i>The reactivity of a metal is related to its tendency to form positive ions</i>	The reactivity series arranges metals in order of their reactivity (their tendency to form positive ions).
Carbon and hydrogen	<i>Carbon and hydrogen are non-metals but are included in the reactivity series</i>	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	<i>A more reactive metal can displace a less reactive metal from a compound.</i>	Silver nitrate + Sodium chloride \rightarrow Sodium nitrate + Silver chloride

- potassium **most reactive** 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
-

Metals and oxygen	<i>Metals react with oxygen to form metal oxides</i>	magnesium + oxygen \rightarrow magnesium oxide $2Mg + O_2 \rightarrow 2MgO$
Reduction	<i>This is when oxygen is removed from a compound during a reaction</i>	e.g. metal oxides reacting with hydrogen, extracting low reactivity metals
Oxidation	<i>This is when oxygen is gained by a compound during a reaction</i>	e.g. metals reacting with oxygen, rusting of iron



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

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.

Electrolysis of aqueous solutions

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

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

Process of electrolysis	<i>Splitting up using electricity</i>	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	<i>Anode Cathode</i>	The positive electrode is called the anode. The negative electrode is called the cathode.
Where do the ions go?	<i>Cations Anions</i>	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

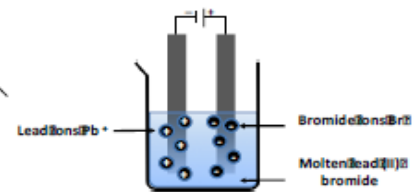
Metals can be extracted from molten compounds using electrolysis.

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.

Higher tier: You can display what is happening at each electrode using half-equations:
 At the cathode: $Pb^{2+} + 2e^{-} \rightarrow Pb$
 At the anode: $2Br^{-} \rightarrow Br_2 + 2e^{-}$



Strong and weak acids (HT ONLY)

Electrolysis

AQA Chemistry
C4: Chemical Changes

Reactions of acids

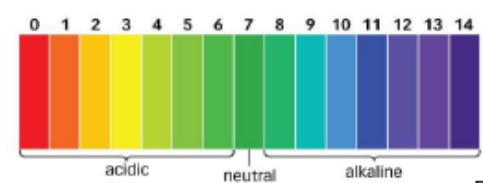
Titration (Chemistry only)

- Use the pipette to add 25 cm³ of alkali to a conical flask and add a few drops of indicator.
- 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.
- 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.

Titration is used to work out the precise volumes of acid and alkali solutions that react with each other.

Soluble 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:
 $H^{+} + OH^{-} \rightarrow H_2O$

Acids	<i>Acids produce hydrogen ions (H⁺) in aqueous solutions.</i>
Alkalis	<i>Aqueous solutions of alkalis contain hydroxide ions (OH⁻).</i>

Calculating the chemical quantities in titrations involving concentrations in mol/dm³ and in g/dm³ (HT ONLY):

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

It takes 12.20cm³ of sulfuric acid to neutralise 24.00cm³ of sodium hydroxide solution, which has a concentration of 0.50mol/dm³.

Calculate the concentration of the sulfuric acid in g/dm³

$$0.5 \text{ mol/dm}^3 \times (24/1000) \text{ dm}^3 = 0.012 \text{ mol of NaOH}$$

The equation shows that 2 mol of NaOH reacts with 1 mol of H₂SO₄, so the number of moles in 12.20cm³ of sulfuric acid is $(0.012/2) = 0.006 \text{ mol of sulfuric acid}$

Calculate the concentration of sulfuric acid in mol/dm³

$$0.006 \text{ mol} \times (1000/12.2) \text{ dm}^3 = 0.49 \text{ mol/dm}^3$$

Calculate the concentration of sulfuric acid in g/dm³

$$H_2SO_4 = (2 \times 1) + 32 + (4 \times 16) = 98 \text{ g}$$

$$0.49 \times 98 \text{ g} = 48.2 \text{ g/dm}^3$$



Endothermic	<i>Energy is taken in from the surroundings so the temperature of the surroundings decreases</i>	<ul style="list-style-type: none"> Thermal decomposition Sports injury packs
Exothermic	<i>Energy is transferred to the surroundings so the temperature of the surroundings increases</i>	<ul style="list-style-type: none"> Combustion Hand warmers Neutralisation

Ionic half equations	Negative electrode: $2\text{H}_2(\text{g}) + 4\text{OH}^-(\text{aq}) \rightarrow 4\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$	Positive electrode: $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$
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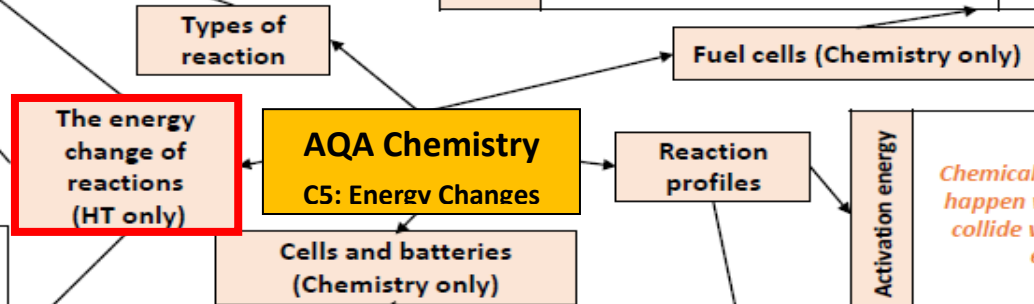
Hydrogen fuel cells	Word equation: <i>hydrogen + oxygen → water</i>	Symbol equation: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
	Advantages: <ul style="list-style-type: none"> No pollutants produced Can be a range of sizes 	Disadvantages: <ul style="list-style-type: none"> Hydrogen is highly flammable Hydrogen is difficult to store

Reaction profiles	<i>Show the overall energy change of a reaction</i>
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Breaking bonds in reactants	<i>Endothermic process</i>
Making bonds in products	<i>Exothermic process</i>

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

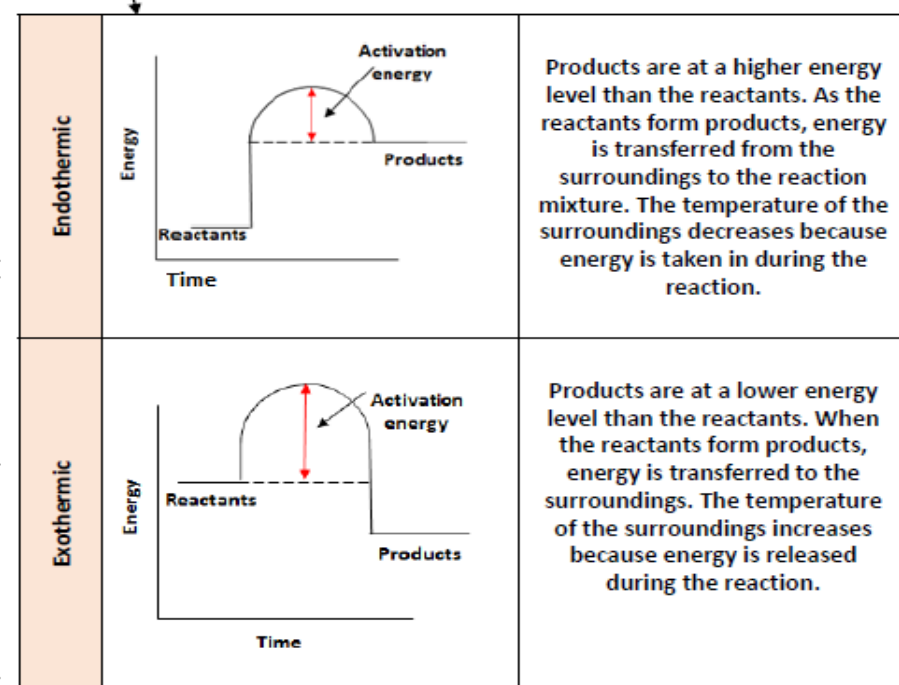
Bond energy calculation	Calculate the overall energy change for the forward reaction $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$
	Bond energies (in kJ/mol): H-H 436, H-N 391, N≡N 945
	Bond breaking: $945 + (3 \times 436) = 945 + 1308 = 2253 \text{ kJ/mol}$
	Bond making: $6 \times 391 = 2346 \text{ kJ/mol}$ Overall energy change = $2253 - 2346 = -93 \text{ kJ/mol}$ Therefore reaction is exothermic overall.



Simple cell	<i>Make a simple cell by connecting two different metals in contact with an electrolyte</i>	Increase the voltage by increasing the reactivity difference between the two metals.
Batteries	<i>Consist of two or more cells connected together in series to provide a greater voltage.</i>	

Non-rechargeable cells	<i>Stop when one of the reactants has been used up</i>	Alkaline batteries
Rechargeable cells	<i>Can be recharged because the chemical reactions are reversed when an external electrical current is supplied</i>	Rechargeable batteries

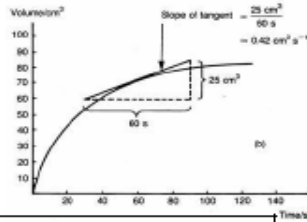
Activation energy	<i>Chemical reactions only happen when particles collide with sufficient energy</i>	The minimum amount of energy that colliding particles must have in order to react is called the activation energy.
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Rate of chemical reaction	<i>This can be calculated by measuring the quantity of reactant used or product formed in a given time.</i>	Rate = $\frac{\text{quantity of reactant used}}{\text{time taken}}$ Rate = $\frac{\text{quantity of product formed}}{\text{time taken}}$
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Quantity	Unit
Mass	Grams (g)
Volume	cm ³
Rate of reaction	Grams per cm ³ (g/cm ³) HT: moles per second (mol/s)



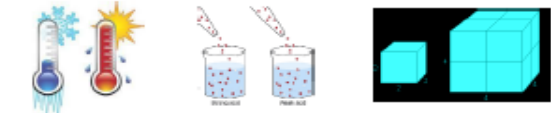
Calculating rates of reactions

Rate of reaction

Factors affecting rates

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

Collision theory and activation energy

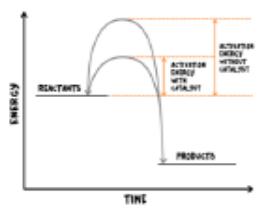


Collision theory	<i>Chemical reactions can only occur when reacting particles collide with each other with sufficient energy.</i>	Increasing the temperature increases the frequency of collisions and makes the collisions more energetic, therefore increasing the rate of reaction.
Activation energy	<i>This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

AQA Chemistry C6: The Rate and extent of chemical change

Reversible reactions and dynamic equilibrium

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.



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

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: $A + B \xrightleftharpoons[\text{cool}]{\text{heat}} C + D$

Energy changes and reversible reactions

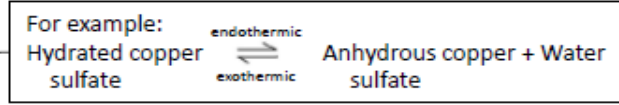
If one direction of a reversible reaction is exothermic, the opposite direction is endothermic. The same amount of energy is transferred in each case.

Catalysts

Equilibrium

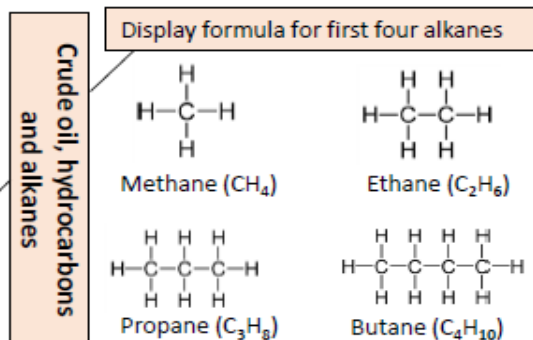
Changing conditions and equilibrium (HT)
The relative amounts of reactants and products at equilibrium depend on the conditions of the reaction.

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.



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.	

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



Fractions	<i>The hydrocarbons in crude oil can be split into fractions</i>	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	<i>Fractions can be processed to produce fuels and feedstock for petrochemical industry</i>	We depend on many of these fuels; petrol, diesel and kerosene. Many useful materials are made by the petrochemical industry; solvents, lubricants and polymers.

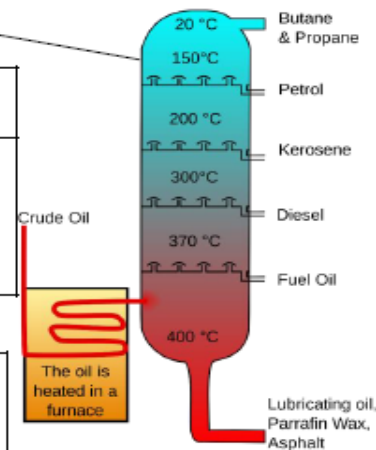
Carbon compounds as fuels and feedstock

**AQA Chemistry
C7: Organic Chemistry**

Carbon compounds as fuels and feedstock

Fractional distillation and petrochemicals

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



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

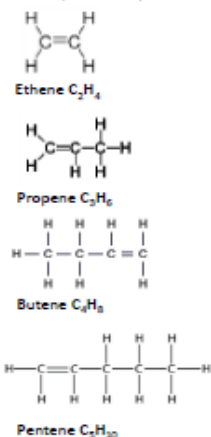
Cracking and alkenes

Properties of hydrocarbons

Combustion	During the complete combustion of hydrocarbons, the carbon and hydrogen in the fuels are oxidised, releasing carbon dioxide, water and energy.
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Cracking	<i>The breaking down of long chain hydrocarbons into smaller chains</i>	The smaller chains are more useful. Cracking can be done by various methods including catalytic cracking and steam cracking.
Catalytic cracking	<i>The heavy fraction is heated until vaporised</i>	After vaporisation, the vapour is passed over a hot catalyst forming smaller, more useful hydrocarbons.
Steam cracking	<i>The heavy fraction is heated until vaporised</i>	After vaporisation, the vapour is mixed with steam and heated to a very high temperature forming smaller, more useful hydrocarbons.

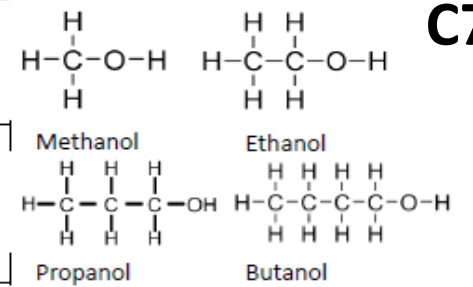
Decane → pentane + propene + ethane $C_{10}H_{22} \rightarrow C_5H_{12} + C_3H_6 + C_2H_4$		Complete combustion of methane: Methane + oxygen → carbon dioxide + water + energy $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	
Alkenes and uses as polymers	<i>Used to produce polymers. They are also used as the starting materials of many other chemicals, such as alcohol, plastics and detergents.</i>	Boiling point (temperature at which liquid boils)	<i>As the hydrocarbon chain length increases, boiling point increases.</i>
Why do we crack long chains?	<i>Without cracking, many of the long hydrocarbons would be wasted as there is not much demand for these as for the shorter chains.</i>	Viscosity (how easily it flows)	<i>As the hydrocarbon chain length increases, viscosity increases.</i>
		Flammability (how easily it burns)	<i>As the hydrocarbon chain length increases, flammability decreases.</i>



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

Structure and formula of alkenes

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



Reactions of alkenes and alcohols

Functional group	-OH <i>For example: CH₃CH₂OH</i>	Methanol, ethanol, propanol and butanol are the first four of the homologous series.
Alcohol reactions	<i>Alcohols react with sodium, air and water.</i>	Alcohols and sodium: bubbling, hydrogen gas given off and salt formed. Alcohols and air: alcohols burn in air releasing carbon dioxide and water. Alcohols and water: alcohols dissolve in water to form a neutral solution.
Fermentation	<i>Ethanol is produced from fermentation.</i>	When sugar solutions are fermented using yeast, aqueous solutions of ethanol are produced. The conditions needed for this process include a moderate temperature (25 – 50°C), water (from sugar solution) and an absence of oxygen.

Functional group	-COOH <i>For example: CH₃COOH</i>	Methanoic acid, ethanoic acid, propanoic acid and butanoic acid are the first four of the homologous series.
Carboxylic acid reactions	<i>Carboxylic acids react with carbonates, water and alcohols.</i>	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)	<i>Carboxylic acids are weak acids</i>	Carboxylic acids only partially ionise in water. An aqueous solution of a weak acid with have a high pH (but still below 7).

Carboxylic acids

Addition polymerisation

AQA Chemistry
C7: Organic Chemistry

Synthetic and naturally occurring polymers

Condensation polymerisation (HT only)

Amino acids

Amino acids have two functional groups in a molecule. They react by condensation polymerisation to produce peptides.

$$\begin{array}{c} \text{H} & & \text{H} & & \text{O} \\ | & & | & & || \\ \text{H}-\text{N} & - & \text{C} & - & \text{C} & - & \text{O}-\text{H} \\ | & & | & & \\ \text{H} & & \text{H} & & \end{array}$$

Glycine

DNA and naturally occurring polymers

Polymers	<i>Alkenes are used to make polymers by addition polymerisation.</i>	Many small molecules join together to form polymers (very large molecules).
Displaying polymers	<i>In addition polymers, the repeating unit has the same atoms as the monomer.</i>	It can be displayed like this: $n \left(\begin{array}{c} \text{H} & \text{H} \\ & \\ \text{C} & = & \text{C} \\ & \\ \text{H} & \text{H} \end{array} \right) \xrightarrow{\text{polymerisation}} \left[\begin{array}{c} \text{H} & \text{H} \\ & \\ -\text{C} & - & \text{C}- \\ & \\ \text{H} & \text{H} \end{array} \right]_n$ ethene repeating unit of poly(ethene)

Condensation polymerisation	<i>Condensation polymerisation involves monomers with two functional groups</i>	When these types of monomers react they join together and usually lose small molecules, such as water. This is why they are called condensation reactions.
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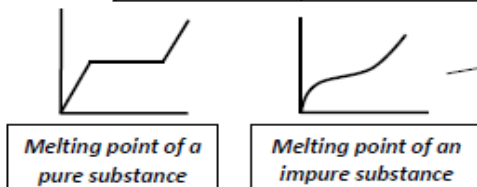
DNA	<i>Deoxyribonucleic acid is a large molecule essential for life. DNA gives the genetic instructions to ensure development and functioning of living organisms and viruses.</i>
DNA structure	<i>Most DNA molecules are two polymer chains made from four different monomers, called nucleotides. They are in the double helix formation.</i>
Natural polymers	<i>Other naturally occurring polymers include proteins, starch and cellulose and are all important for life.</i>



Pure substances	<i>A pure substance is a single element or compound, not mixed with any other substance.</i>	Pure substances melt and boil at specific temperatures. Heating graphs can be used to distinguish pure substances from impure.
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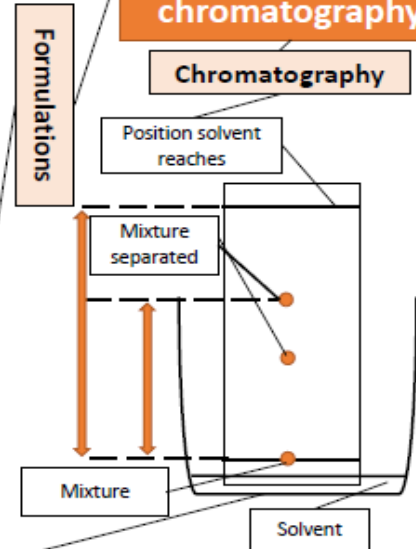
Element	Colour flames
Lithium	<i>Crimson</i>
Sodium	<i>Yellow</i>
Potassium	<i>Lilac</i>
Calcium	<i>Orange-red</i>
Copper	<i>Green</i>

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



Pure substances
Purity, formulations and chromatography

Formulation	<i>A formulation is a mixture that has been designed as a useful product.</i>
How are formulations made?	<i>By mixing chemicals that have a particular purpose in careful quantities.</i>
Examples of formulations.	<i>Fuels, cleaning agents, paints, medicines and fertilisers.</i>



Chromatography	<i>Can be used to separate mixtures and help identify substances.</i>	Involves a mobile phase (e.g. water or ethanol) and a stationary phase (e.g. chromatography paper).
R_f Values	<i>The ratio of the distance moved by a compound to the distance moved by solvent.</i>	$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$
Pure substances	<i>The compounds in a mixture separate into different spots.</i>	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.

AQA Chemistry
C8: Chemical Analysis

Identification of ions (CHEMISTRY ONLY)

Identification of common gases

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

Metal hydroxides (chem only)

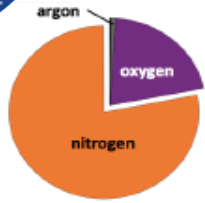
Carbonates, halides and sulfates (chem only)

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

Instrumental methods

Instrumental methods	<i>Methods that rely on machines</i>	Can be used to identify elements and compounds. These methods are accurate, sensitive and rapid.
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Flame emission spectroscopy	<i>An instrumental method used to analyse metal ions.</i>	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.
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Gas	Percentage
Nitrogen	~80%
Oxygen	~20%
Argon	0.93%
Carbon dioxide	0.04%

Proportions of gases in the atmosphere

Algae and plants	<i>These produced the oxygen that is now in the atmosphere, through photosynthesis.</i>	carbon dioxide + water → glucose + oxygen $6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2$
Oxygen in the atmosphere	<i>First produced by algae 2.7 billion years ago.</i>	Over the next billion years plants evolved to gradually produce more oxygen. This gradually increased to a level that enabled animals to evolve.

Volcano activity 1 st Billion years	<i>Billions of years ago there was intense volcanic activity</i>	This released gases (mainly CO ₂) that formed to early atmosphere and water vapour that condensed to form the oceans.
Other gases	<i>Released from volcanic eruptions</i>	Nitrogen was also released, gradually building up in the atmosphere. Small proportions of ammonia and methane also produced.
Reducing carbon dioxide in the atmosphere	<i>When the oceans formed, carbon dioxide dissolved into it</i>	This formed carbonate precipitates, forming sediments. This reduced the levels of carbon dioxide in the atmosphere.

The Earth's early atmosphere

How oxygen increased
How carbon dioxide decreased

Composition and evolution of the atmosphere

AQA Chemistry C9: Chemistry of the atmosphere

Common atmospheric pollutants

Reducing carbon dioxide in the atmosphere	<i>Algae and plants</i>	These gradually reduced the carbon dioxide levels in the atmosphere by absorbing it for photosynthesis.
Formation of sedimentary rocks and fossil fuels	<i>These are made out of the remains of biological matter, formed over millions of years</i>	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.

CO₂ and methane as greenhouse gases

Greenhouse gases

Carbon dioxide, water vapour and methane	<i>Examples of greenhouse gases that maintain temperatures on Earth in order to support life</i>
The greenhouse effect	<i>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.</i>

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

Human activities and greenhouse gases

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

Atmospheric pollutants from fuels

Properties and effects of atmospheric pollutants

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

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

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

Corrosion	<i>The destruction of materials by chemical reactions with substances in the environment</i>	An example of this is iron rusting; iron reacts with oxygen from the air to form iron oxide (rust) water needs to be present for iron to rust.
Preventing corrosion	<i>Coatings can be added to metals to act as a barrier</i>	Examples of this are greasing, painting and electroplating. Aluminium has an oxide coating that protects the metal from further corrosion.
Sacrificial corrosion	<i>When a more reactive metal is used to coat a less reactive metal</i>	This means that the coating will react with the air and not the underlying metal. An example of this is zinc used to galvanise iron.

Corrosion and its prevention

Alloys are useful materials

Alloys	<i>A mixture of two elements, one of which must be a metal e.g. Bronze is an alloy of copper and tin and Brass is an alloy of copper and zinc.</i>
Gold carats	<i>Gold jewellery is usually an alloy with silver, copper and zinc. The carat of the jewellery is a measure of the amount of gold in it e.g. 18 carat is 75% gold, 24 carat is 100% gold.</i>
Steels	<i>Alloys of iron, carbon and other metals.</i>
	<i>High carbon steel is strong but brittle.</i>
	<i>Low carbon steel is softer and easily shaped.</i>
	<i>Steel containing chromium and nickel (stainless) are hard and corrosion resistant.</i>
	<i>Aluminium alloys are low density.</i>

Ceramics, polymers and composites

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

NPK fertilisers	<i>These contain nitrogen, phosphorous and potassium</i>	Formulations of various salts containing appropriate percentages of the elements.
Fertiliser examples	<i>Potassium chloride, potassium sulfate and phosphate rock are obtained by mining</i>	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.

Production and uses of NPK fertilisers

Using materials

AQA Chemistry

C10: Using Resources

The Haber process and the use of NPK fertilisers

Composite materials	<i>A mixture of materials put together for a specific purpose e.g. strength</i>	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	<i>Made from clay</i>	Made by shaping wet clay and then heating in a furnace, common examples include pottery and bricks.
Polymers	<i>Many monomers can make polymers</i>	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.

The Haber process

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

The Haber process – conditions and equilibrium	
Pressure	<i>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.</i>
Temperature	<i>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.</i>

The Haber process	<i>Used to manufacture ammonia</i>	<i>Ammonia is used to produce fertilisers</i> Nitrogen + hydrogen \rightleftharpoons ammonia
Raw materials	<i>Nitrogen from the air while hydrogen from natural gas</i>	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	<i>Iron</i>	The catalyst speeds up both directions of the reaction, therefore not actually increasing the amount of valuable product.

