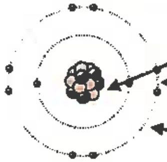
AQA GCSE Atomic structure and periodic table part 1

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

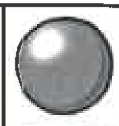
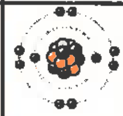
Mass number	*The sum of the protons and neutrons in the nucleus*	
Atomic number	*The number of protons in the atom*	Number of electrons = number of protons

Mixtures

Mixtures	*Two or more elements or compounds not chemically combined together*	Can be separated by physical processes.

Method	Description	Example
Filtration	*Separating an insoluble solid from a liquid*	To get sand from a mixture of sand, salt and water.
Crystallisation	*To separate a solid from a solution*	To obtain pure crystals of sodium chloride from salt water.
Simple distillation	*To separate a solvent from a solution*	To get pure water from salt water.
Fractional distillation	*Separating a mixture of liquids each with different boiling points*	To separate the different compounds in crude oil.
Chromatography	*Separating substances that move at different rates through a medium*	To separate out the dyes in food colouring.

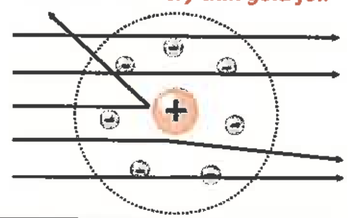
The development of the model of the atom

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.

Chemical equations

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.

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

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AQA GCSE Atomic structure and periodic table part 2

The Periodic table

Alkali metals · Transition metals · Halogens · Noble gases

1	2											3	4	5	6	7	0
H																	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	?	?	?						

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

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. $4Na + O_2 \rightarrow 2Na_2O$
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. $2Na + Cl_2 \rightarrow 2NaCl$

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

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*	

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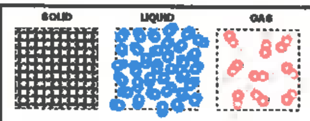
AQA BONDING, STRUCTURE AND THE PROPERTIES OF MATTER 1

Chemical bonds

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

The three states of matter

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

Properties of ionic compounds

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

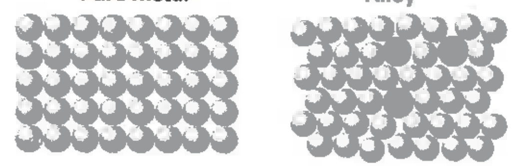
Metals as conductors

Good conductors of electricity	Delocalised electrons carry electrical charge through the metal.
Good conductors of thermal energy	Energy is transferred by the delocalised electrons.

Properties of metals and alloys

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.

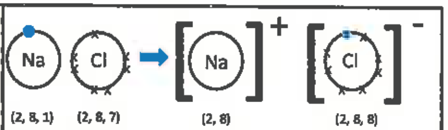
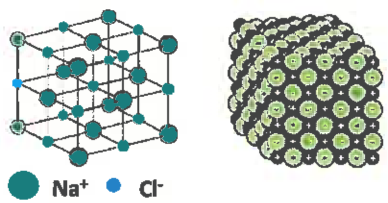
Alloys	*Mixture of two or more elements at least one of which is a metal*	Harder than pure metals because atoms of different sizes disrupt the layers so they cannot slide over each other.



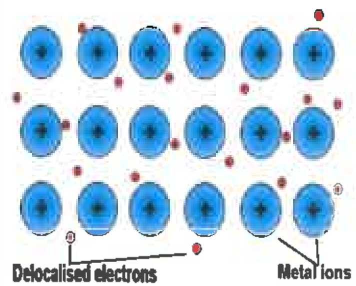
Ionic bonding

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

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*

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.

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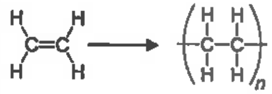
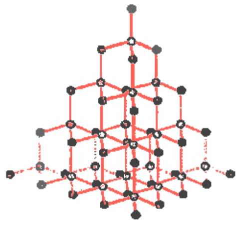
AQA BONDING, STRUCTURE AND THE PROPERTIES OF MATTER 2

Polymers

Very large molecules	*Solids at room temperature*	Atoms are linked by strong covalent bonds.	

Diamond



Each carbon atom is bonded to four others	Very hard.	Rigid structure.
	Very high melting point.	Strong covalent bonds.
	Does not conduct electricity.	No delocalised electrons.

Properties of small molecules

Usually gases or liquids

Covalent bonds in the molecule are strong but forces between molecules (intermolecular) are weak

Low melting and boiling points.	Due to having weak intermolecular forces that easily broken.
Do not conduct electricity.	Due to them molecules not having an overall electrical charge.
Larger molecules have higher melting and boiling points.	Intermolecular forces increase with the size of the molecules.

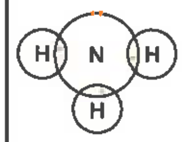
Giant covalent structures

Diamond, graphite, silicon dioxide	*Very high melting points*	Lots of energy needed to break strong, covalent bonds.

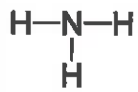
Covalent bonding

Atoms share pairs of electrons

Can be small molecules e.g. ammonia



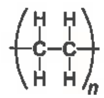
Dot and cross :
+ Show which atom the electrons in the bonds come from
- All electrons are identical



2D with bonds:
+ Show which atoms are bonded together
- It shows the H-C-H bond incorrectly at 90°

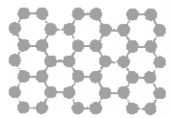
3D ball and stick model:
+ Attempts to show the H-C-H bond angle is 109.5°

Can be giant covalent structures e.g. polymers



Graphene and fullerenes

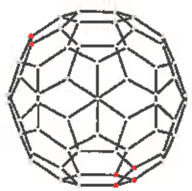
Graphene



Single layer of graphite one atom thick

Excellent conductor.	Contains delocalised electrons.
Very strong.	Contains strong covalent bonds.

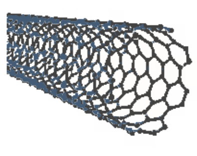
Fullerenes



Buckminsterfullerene, C_{60} First fullerene to be discovered.

Hexagonal rings of carbon atoms with hollow shapes. Can also have rings of five (pentagonal) or seven (heptagonal) carbon atoms.

Carbon nanotubes



Very thin and long cylindrical fullerenes

Very conductive.	Used in electronics industry.
High tensile strength.	Reinforcing composite materials.
Large surface area to volume ratio.	Catalysts and lubricants.

Size of particles and their properties (Chemistry only)

Nanoparticles

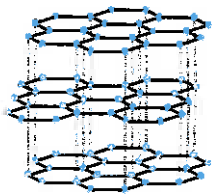
Between 1 and 100 nanometres (nm) in size	1 nanometre (1 nm) = 1×10^{-9} metres (0.000 000 001m or a billionth of a metre).

Use of nanoparticles

Healthcare, cosmetics, sun cream, catalysts, deodorants, electronics.	Nanoparticles may be toxic to people. They may be able to enter the brain from the bloodstream and cause harm.

Graphite

Each carbon atom is bonded to three others forming layers of hexagonal rings with no covalent bonds between the layers



Slippery.	Layers can slide over each other.
Very high melting point.	Strong covalent bonds.
Does conduct electricity.	Delocalised electrons between layers.

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AQA GCSE
QUANTITATIVE CHEMISTRY 1

Relative formula mass (M_r)

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.	$2Mg + O_2 \rightarrow 2MgO$ 48g + 32g = 80g 80g = 80g

Mass changes when a reactant or product is a gas

Mass appears to increase during a reaction	*One of the reactants is a gas*	Magnesium + oxygen → magnesium oxide
Mass appears to decrease during a reaction	*One of the products is a gas and has escaped*	Calcium carbonate → carbon dioxide + calcium oxide

Conservation of mass and balanced symbol equations

Conservation of mass	*No atoms are lost or made during a chemical reaction*	Mass of the products equals the mass of the reactants.
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 / Normal script Subscript numbers show the number of atoms of the element to its left. Normal script numbers show the number of molecules.

Moles (HT only)

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)

Limiting reactants (HT only)

The reactant that is completely used up	*Limits the amount of product that is made*	Less moles of product are made.

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 46.5s ±2.5s

Concentration of solutions

Measured in mass per given volume of solution (g/dm³)	*Conc. = $\frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$*	HT only Greater mass = higher concentration. Greater volume = lower concentration.

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.*

Amounts of substances in equations (HT only)

Chemical equations show the number of moles reacting and the number of moles made	*$Mg + 2HCl \rightarrow MgCl_2 + H_2$* *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 : Mg = 24 so mass of 1 mole of Mg = 24g M_r : HCl (1 + 35.5) so mass of 1 mole of HCl = 36.5g 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 2.5x2 = 5 moles of HCl You will need 5 x 36.5g of HCl= 182.5g

AQA QUANTITATIVE CHEMISTRY 2

Atom economy

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 = $^2/_{138} \times 100$

= $^2/_{138} \times 100$ = 1.45%

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

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

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

Concentration (mol/dm³) = amount (mol) / volume (dm³)

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

35/0.5 = 70 g/dm³

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 12.20cm³ 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 mol/dm³ x (24/1000) dm³ = 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 12.20cm³ of sulfuric acid is (0.012/2) = 0.006 mol of sulfuric acid

Calculate the concentration of sulfuric acid in mol/ dm³

0.006 mol x (1000/12.2) dm³ =0.49mol/dm³

Calculate the concentration of sulfuric acid in g/ dm³:

H_2SO_4 = (2x1) + 32 + (4x16) = 98g

0.49 x 98g = 48.2g/dm³

Percentage yield

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.*

$CaCO_3 \rightarrow CaO + CO_2$

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

Yield is the amount of product obtained

It is not always possible to obtain the calculated amount of a product

- The reaction may not go to completion because it is reversible.
- Some of the product may be lost when it is separated from the reaction mixture.
- Some of the reactants may react in ways different to the expected reaction.

Percentage yield is comparing the amount of product obtained as a percentage of the maximum theoretical amount

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

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

Equal amounts of moles or gases occupy the same volume under the same conditions of temperature and pressure

The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmospheric pressure) is 24 dm³

No. of moles of gas = $\frac{\text{vol of gas (dm}^3\text{)}}{24\text{dm}^3}$

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

M_r : (4 x 12) + (10 x 1) = 58

11.6/58 = 0.20 mol

Volume = 0.20 x 24 = 4.8 dm³

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 mol

M_r = 6 / 0.2 = 30

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

AQA Chemical Changes 1

Reactions of acids

Reactivity of metals

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

Oxidation and reduction in terms of electrons (HT ONLY)

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

Neutralisation of acids and salt production

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

sodium hydroxide + hydrochloric acid → sodium chloride + water

calcium carbonate + sulfuric acid → calcium sulfate, + carbon dioxide + water

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.

Metal oxides

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

Reactions of acids and metals

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.

Reactions with acids	*metal + acid → metal salt + hydrogen*	magnesium + hydrochloric acid → magnesium chloride + hydrogen zinc + sulfuric acid → zinc sulfate + hydrogen

Acids react with some metals to produce salts and hydrogen.

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.

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

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

The reactivity series

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

Metal		Symbol
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

AQA Chemical Changes 2

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 though 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.

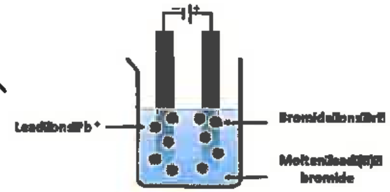
Electrolysis of aqueous solutions

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.

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: $2Br^- \rightarrow Br_2 + 2e^-$

Reactions of acids

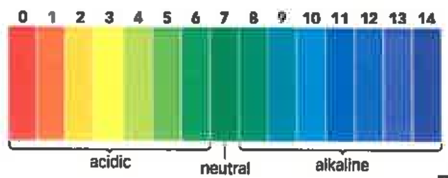
Strong and weak acids (HT ONLY)

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.

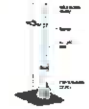
Acids	*Acids produce hydrogen ions (H^+) in aqueous solutions.*
Alkalis	*Aqueous solutions of alkalis contain hydroxide ions (OH^-).*

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

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

Titrations (Chemistry only)

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

	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.

Calculating the chemical quantities in titrations involving concentrations in mol/dm^3 and in g/dm^3 (HT ONLY):

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

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 g/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

AQA GCSE Energy changes

Types of reaction

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

The energy change of reactions (HT only)

Reaction profiles	*Show the overall energy change of a reaction*

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

Overall energy change of a 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.

Bond energy calculation

Calculate the overall energy change for the forward reaction

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

Bond energies (in kJ/mol): H-H 436, H-N 391, N≡N 945

Bond breaking: 945 + (3 x 136) = 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.

Fuel cells (Chemistry only)

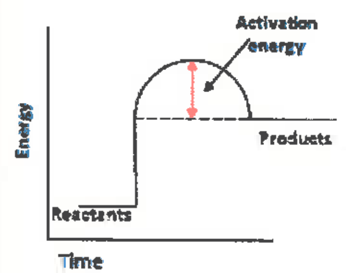
Ionic half equations	Negative electrode:	Positive electrode:
	$2H_2 (g) + 4OH^- (aq) \rightarrow 4H_2O (l) + 4e^-$	$O_2 (g) + 2H_2O (l) + 4e^- \rightarrow 4OH^- (aq)$

Hydrogen fuel cells		
	Word equation: *hydrogen + oxygen → water*	Symbol equation: $2H_2 + O_2 \rightarrow 2H_2O$
	Advantages: • No pollutants produced • Can be a range of sizes	Disadvantages: • Hydrogen is highly flammable • Hydrogen is difficult to store

Reaction profiles

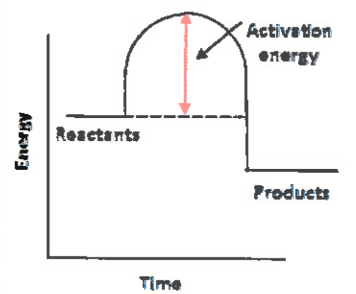
Activation energy	*Chemical reactions only happen when particles collide with sufficient energy*
	The minimum amount of energy that colliding particles must have in order to react is called the activation energy.

Endothermic



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.

Exothermic



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.

Cells and batteries (Chemistry only)

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

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

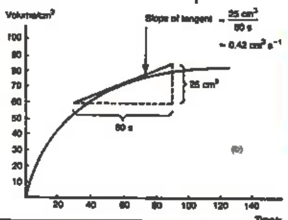
AQA GCSE The rate and extent of chemical change

Rate of reaction

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

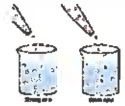
Calculating rates of reactions

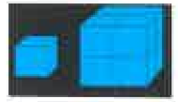
Quantity	Unit
Mass	Grams (g)
Volume	cm^3
Rate of reaction	Grams per cm^3 (g/cm^3) HT: moles per second (mol/s)



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.*



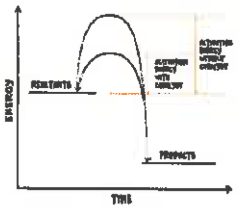


Collision theory and activation energy

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

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 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 ⇌ C + D
The direction	The direction of reversible reactions can be changed by changing conditions: A + B ⇌ C + D (heat →, ← cool)

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 ⇌ Anhydrous copper sulfate + Water (endothermic →, ← exothermic)

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.

AQA GCSE Organic chemistry 1

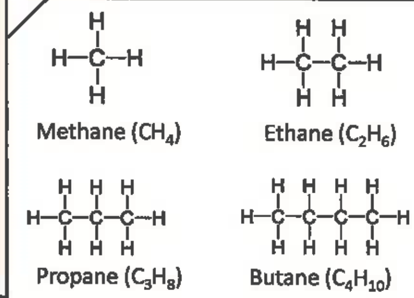
Carbon compounds as fuels and feedstock

Carbon compounds as fuels and feedstock

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}

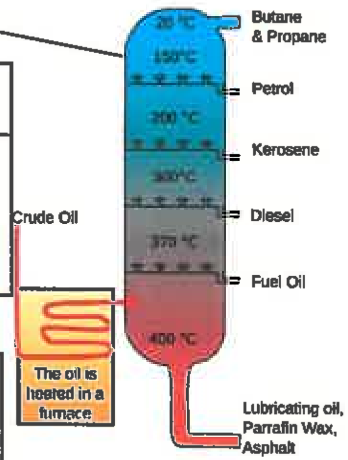
Display formula for first four alkanes



Fractional distillation and petrochemicals

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.

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 and alkenes

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 that alkanes and react with bromine water. Bromine water changes from orange to colourless in the presence of alkenes.*

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.

Decane → pentane + propene + ethane

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

Alkenes and uses as polymers	*Used to produce polymers. They are also used as the starting materials of many other chemicals, such as alcohol, plastics and detergents.*
Why do we crack long chains?	*Without cracking, many of the long hydrocarbons would be wasted as there is not much demand for these as for the shorter chains.*

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.

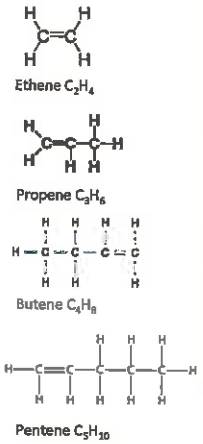
Complete combustion of methane:
Methane + oxygen → carbon dioxide + water + energy
$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(l)$

Boiling point (temperature at which liquid boils)	*As the hydrocarbon chain length increases, boiling point increases.*
Viscosity (how easily it flows)	*As the hydrocarbon chain length increases, viscosity increases.*
Flammability (how easily it burns)	*As the hydrocarbon chain length increases, flammability decreases.*

AQA GCSE Organic chemistry 2 (CHEMISTRY ONLY)

Reactions of alkenes and alcohols

Structure and formula of alkenes

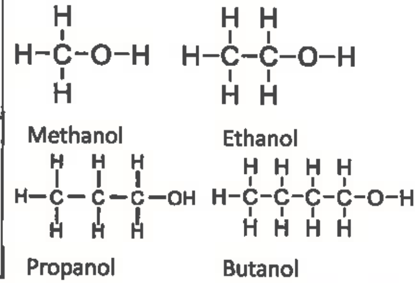


Alkenes	*Hydrocarbons with a double carbon-carbon bond.*	
Unsaturated	*Alkenes are unsaturated because they contain two fewer hydrogen atoms than their alkane counterparts.*	
General formula for alkenes	*C_nH_{2n}*	

Reactions of alkenes

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.

Alcohols



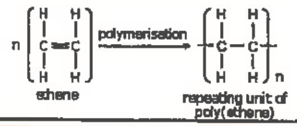
Functional group	*-OH* *For example: CH_3CH_2OH*	Methanol, ethanol, propanol and butanol are the first four of the homologous series.
Alcohol reactions	*Alcohols react with sodium, air and water.*	Alcohols and sodium: bubbling, hydrogen gas given off and salt formed. Alcohols and air: alcohols burn in air releasing carbon dioxide and water. Alcohols and water: alcohols dissolve in water to form a neutral solution.
Fermentation	*Ethanol is produced from fermentation.*	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.

Carboxylic acids

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).

Synthetic and naturally occurring polymers

Addition polymerisation

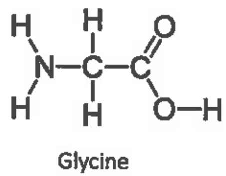
Polymers	*Alkenes are used to make polymers by addition polymerisation.*	Many small molecules join together to form polymers (very large molecules).
Displaying polymers	*In addition polymers, the repeating unit has the same atoms as the monomer.*	It can be displayed like this: 

Condensation polymerisation (HT only)

Condensation polymerisation	*Condensation polymerisation involves monomers with two functional groups*	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.

Amino acids

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



Glycine

DNA and naturally occurring polymers

DNA	*Deoxyribonucleic acid is a large molecule essential for life. DNA gives the genetic instructions to ensure development and functioning of living organisms and viruses.*
DNA structure	*Most DNA molecules are two polymer chains made from four different monomers, called nucleotides. They are in the double helix formation.*
Natural polymers	*Other naturally occurring polymers include proteins, starch and cellulose and are all important for life.*

AQA Chemical analysis

Purity, formulations and chromatography

Pure substances

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.

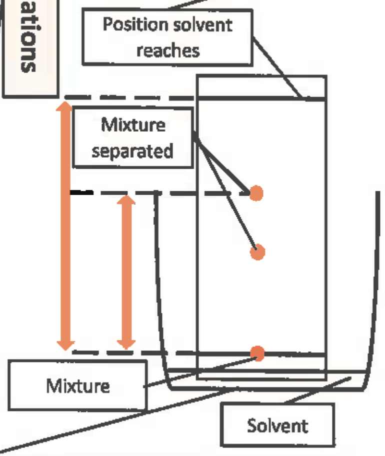
Melting point of a pure substance

Melting point of an impure substance

Formulations

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

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.

Identification of common gases

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).

Identification of ions (CHEMISTRY ONLY)

Flame tests (chem only)

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

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 (chem only)

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.

'Aim For Excellence'

AQA 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.

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.*

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.*

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

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.*

AQA GCSE Using resources 1

Using the Earth's resources and obtaining potable water

Using the Earth's resources and sustainable development

Earth's resources	*Used to provide warmth, shelter, food and transport for humans*	Natural resources and resources from agriculture provide: timber, food, clothing and fuels. Finite resources from the Earth, oceans and atmosphere are processed to provide energy and materials.
Chemistry and resources	*Research and techniques improve agricultural and industrial processes*	These improvements provide new products and improve sustainability.
Plastics	*Normally made using ethene from crude oil*	However, the raw material ethene can also be obtained from ethanol, which can be produced during fermentation. Industries are now starting to use a renewable crop for this process.

Potable water

Sterilising agents include chlorine, ozone and UV light.

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 treatment

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.

Alternative methods of extracting metals (HT)

Metals ores	*These resources are limited*	Copper ores especially are becoming sparse. New ways of extracting copper from low-grade ores are being developed.
Phytomining	*Plants absorb metal compounds*	These plants are then harvested and burned; their ash contains the metal compounds.
Bioleaching	*Bacteria is used to produce leachate solutions that contain metal compounds*	The metal compounds can be processed to obtain the metal from it e.g. copper can be obtained from its compounds by displacement or electrolysis.

Life cycle assessment and recycling

Life cycle assessment

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.

Ways of reducing the use of resources

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.

AQA GCSE Using resources 2 (CHEM ONLY)

Using materials

Corrosion and its prevention

Corrosion	*The destruction of materials by chemical reactions with substances in the environment*	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	*Coatings can be added to metals to act as a barrier*	Examples of this are greasing, painting and electroplating. Aluminium has an oxide coating that protects the metal from further corrosion.
Sacrificial corrosion	*When a more reactive metal is used to coat a less reactive metal*	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.

Alloys are useful materials

Alloys	*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.*
Gold carats	*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.*
Steels	*Alloys of iron, carbon and other metals.*
	High carbon steel is strong but brittle.
	Low carbon steel is softer and easily shaped.
	Steel containing chromium and nickel (stainless) are hard and corrosion resistant.
	Aluminium alloys are low density.

Ceramics, polymers and composites

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

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.

The Haber process and the use of NPK fertilisers

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	*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.

The Haber process – conditions and equilibrium	
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.*