CHEMISTRY
TRIPLE SCIENCE
REVISION BOOKLET

WORLE COMMUNITY SCHOOL

An Academy

Paper 1 & 2
Higher

SENECA
PIONEER SCHOOL

1 Atomic Structure	Review point RAG					Notes
	1st	2nd	3rd	4th	5th	
	R	R	A	A	G	*Speak to teacher*

1.1 Atoms and Elements								
1.1.1	Elements & Compounds							
1.1.2	Chemical Reactions & Equations							
1.1.3	Mixtures							
1.1.4	Model of the Atom							
1.1.5	Atom Size & Number							
1.1.6	Isotopes							
1.1.7	Isotopes HyperLearning							
1.1.8	Periodic Table							
1.1.9	Noble Gases & Halogens							
1.1.10	Alkali Metals							
1.1.11	Transition Metals							
Apply it! - 1.1 Atoms and Elements		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

2 Chemical bonding		Review point RAG					Notes		
		1st	2nd	3rd	4th	5th			
(HT) - Higher tier only		R	R	A	A	G	Speak to teacher		

2.1 Chemical Bonds									
2.1.1	Types of Bonds **(HT)**								
2.1.2	Ionic Bonds								
2.1.3	Ionic Compounds								
2.1.4	Covalent & Metallic Bonds								
2.1.5	Representing Covalent Bonds								
Apply it! - 2.1 Chemical Bonds		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

2.2 States of Matter									
2.2.1	States of Matter								
2.2.2	Changing State								
2.2.3	Changing State HyperLearning **(HT)**								
Apply it! - 2.2 States of Matter		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

2.3 Chemical properties									
2.3.1	Chemical Properties								
2.3.2	Molecular Forces & Polymers								
2.3.3	Metals & Giant Covalent Structures								
2.3.4	Alloys & Conductors **(HT)**								
2.3.5	Carbon Structures								
2.3.6	Carbon Structures 2 **(HT)**								
2.3.7	Nanotechnology								
2.3.8	Nanotechnology 2 **(HT)**								
Apply it! - 2.3 Chemical Properties		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

3 Quantitative Chemistry	Review point RAG					Notes
	1st	2nd	3rd	4th	5th	
	R	R	A	A	G	*Speak to teacher*

3.1 Chemical Measurements								
3.1.1	Relative Formula Mass							
3.1.2	Measuring Mass							
3.1.3	Equations & Formula **(HT)**							
3.1.4	Moles **(HT)**							
3.1.5	Moles 2 **(HT)**							
3.1.6	Yield							
3.1.7	Atom Economy							
3.1.8	Concentrations & Amounts							
Apply it! - 3.1 Chemical Measurements		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

4 Chemical changes	Review point RAG					Notes
	1st	2nd	3rd	4th	5th	
	R	R	A	A	G	*Speak to teacher*

4.1 Acids, Alkalis & Redox Reactions									
4.1.1	Acids, Alkalis & Redox Reactions								
Apply it! - 4.1 Acids, Alkalis & Redox Reac		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

4.2 Reactivity of Metals									
4.2.1	Displacement Reactions								
4.2.2	Reactivity Tests & Extraction								
Apply it! - 4.2 Reactivity of Metals		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

4.3 Reactions of Acids									
4.3.1	Reactions of Metals with Acids								
4.3.2	Reactions of Metals with Acids 2 **(HT)**								
4.3.3	The pH Scale								
4.3.4	Neutralisation & Titrations								
Apply it! - 4.3 Reactions of Acids		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

4.4 Electrolysis									
4.4.1	Electrolysis								
4.4.2	Electrolysis 2 **(HT)**								
Apply it! - 4.4 Electrolysis		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

5 Energy Changes		Review point RAG					Notes		
		1st	2nd	3rd	4th	5th			
		R	R	A	A	G	Speak to teacher		
5.1 Exothermic & Endothermic Reactions									
5.1.1	Exothermic & Endothermic Reactions								
5.1.2	Reaction Profiles								
5.1.3	Reaction Profiles 2 **(HT)**								
Apply it! - 5.1 Exo. & Endo. Reactions		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8
5.2 Chemical Cells & Fuel Cells									
5.2.1	Chemical Cells								
5.2.2	Chemical Cells 2 **(HT)**								
5.2.3	Fuel Cells								
Apply it! - 5.2 Chemical Cells & Fuel Cells		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

6 Rate and extent of Chemical Change		Review point RAG					Notes		
		1st	2nd	3rd	4th	5th			
		R	R	A	A	G	Speak to teacher		
6.1 Rate of Reaction									
6.1.1	Chemical Reactions & Collisions								
6.1.2	Catalysts								
Apply it! - 6.1 Rate of Reaction		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8
6.2 Reversible Reactions									
6.2.1	Reversible Reactions & Equilibrium								
6.2.2	Equilibrium Position								
Apply it! - 6.2 Reversible Reactions		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

7 Organic Chemistry	Review point RAG					Notes
	1st	2nd	3rd	4th	5th	
	R	R	A	A	G	*Speak to teacher*

7.1 Carbon Compounds									
7.1.1	Crude Oil								
7.1.2	Alkanes								
7.1.3	Fractional Distillation								
7.1.4	Burning Hydrocarbons								
7.1.5	Cracking & Alkenes								
Apply it! - 7.1 Carbon Compounds		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

7.2 Alkenes & Alcohol									
7.2.1	Alkenes								
7.2.2	Alkene Reactions								
7.2.3	Alcohols								
7.2.4	Alcohol Reactions								
7.2.5	Carboxylic Acids								
7.2.6	Synthetic & Naturally Occuring Poly.								
7.2.7	Synthetic & Naturally Occuring Poly. 2								
7.2.8	Proteins								
7.2.9	DNA								
Apply it! - 7.2 Alkenes & Alcohol		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

8 Chemical Analysis		Review point RAG					Notes		
		1st	2nd	3rd	4th	5th			
		R	R	A	A	G	Speak to teacher		

8.1 Purity, Formulations & Chromatography									
8.1.1	Purity & Formulations								
8.1.2	Chromatography								
Apply it! - 8.1 Purity, Formulations & Chro		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

8.2 Identification of Common Gases									
8.2.1	Identification of Common Gases								
8.2.2	Identification of Common Gases 2								
Apply it! - 8.2 Identification of Common Gases		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

8.3 Identifying Ions									
8.3.1	Flame Tests								
8.3.2	Metal Hydroxides								
8.3.3	Carbonates, Halides, Sulfates								
8.3.4	Methodology								
Apply it! - 8.3 Identifying Ions		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

9 Chemistry of the Atmos.	Review point RAG					Notes
	1st	2nd	3rd	4th	5th	
	R	R	A	A	G	Speak to teacher

9.1 The Earth's Atmosphere									
9.1.1	Proportions of Gases in the Atmosphere								
9.1.2	Oxygen & Carbon Dioxide								
9.1.3	Greenhouse Gases								
9.1.4	Greenhouse Gases 2								
9.1.5	Sources of Common Atmospheric Poll								
9.1.6	Sources of Common Atmospheric Poll 2								
Apply it! - 9.1 The Earth's Atmosphere		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

10 Using Resources	Review point RAG					Notes
	1st	2nd	3rd	4th	5th	
	R	R	A	A	G	*Speak to teacher*

10.1 Using the Earth's Resources									
10.1.1	Potable Water								
10.1.2	Potable Water 2								
10.1.3	Natural Resources								
10.1.4	Waste Water								
10.1.5	Alternative Methods of Extracting Metals								
10.1.6	Alternative Methods of Extracting Metals 2								
Apply it! - 10.1 Using the Earth's Resources		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

10.2 Life Cycle Assessments & Recycling									
10.2.1	Life Cycle Assessments								
10.2.2	Recycling								
Apply it! - 10.2 Life Cycle Asses. & Recy.		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

10.3 Using Materials									
10.3.1	Corrosion & Alloys								
10.3.2	Ceramics, Polymers & Composites								
Apply it! - 10.3 Using Materials		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

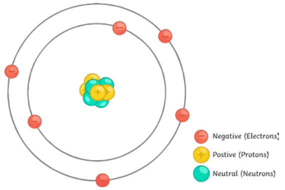
10.4 The Haber Process & NPK Fertilisers									
10.4.1	The Haber Process								
10.4.2	The Haber Process 2								
10.4.3	Fertilisers								
10.4.4	NPK								
Apply it! - 10.4 The Haber Pro. & NPK Fert		Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8

Atomic Structure and the Periodic Table – Foundation and Higher (Separate)

Atoms

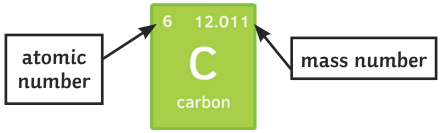
Contained in the nucleus are the **protons** and **neutrons**. Moving around the nucleus are the **electron** shells. They are negatively charged.

Particle	Relative Mass	Charge
proton	1	+1
neutron	1	0
electron	Very small	-1



Overall, atoms have no charge; they have the same number of protons as electrons. An ion is a charged particle - it does not have an equal number of protons to electrons.

Atomic Number and Mass Number



Equations and Maths

To calculate the relative atomic mass, use the following equation:

relative atomic mass (A_r) =

$$\frac{\text{sum of (isotope abundance} \times \text{isotope mass number)}}{\text{sum of abundances of all isotopes}}$$

Balancing Symbol Equations

There must be the same number of atoms on both sides of the equation:

$CH_4 + 4O_2 \rightarrow 2H_2O + CO_2$

C = 1

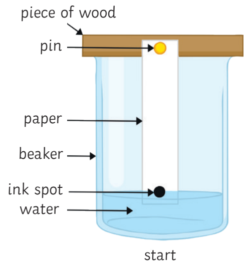
O = 4

H = 4

Mixtures, Chromatography and Separation

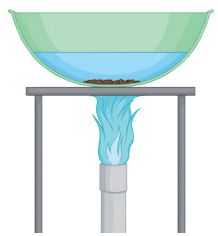
Mixtures – in a mixture there are no chemical bonds, so the elements are easy to separate. Examples of mixtures are air and salt water.

Chromatography – to separate out mixtures.

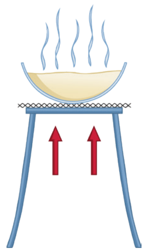


Filtration – to separate solids from liquids.

Evaporation – to separate a soluble salt from a solution; a quick way of separating out the salt.



Crystallisation - to separate a soluble salt from a solution; a slower method of separating out salt.



Separating out salt from rock salt:

1. Grind the mixture of rock salt.
2. Add water and stir.
3. Filter the mixture, leaving the sand in the filter paper
4. Evaporate the water from the salt, leaving the crystals.

Elements

Elements are made of atoms with the same atomic number. Atoms can be represented as symbols.

N = nitrogen **F** = fluorine **Zn** = zinc **Ca** = calcium

Isotopes – an isotope is an element with the **same number of protons** but a **different number of neutrons**. They have the same atomic number, but different mass number.

Isotope	Protons	Electrons	Neutrons
$^{1}_{1}H$	1	1	1 - 1 = 0
$^{2}_{1}H$	1	1	2 - 1 = 1
$^{3}_{1}H$	1	1	3 - 1 = 2

Compounds – a compound is when two or more elements are chemically joined. Examples of compounds are carbon dioxide and magnesium oxide. Some examples of formulas are CO_2, NaCl, HCl, H_2O, Na_2SO_4. They are held together by chemical bonds and are difficult to separate.

Chemical Equations

A chemical reaction can be shown by using a **word equation.**

e.g. magnesium + oxygen → magnesium oxide

On the left-hand side are the reactants, and the right-hand side are the products.

They can also be shown by a **symbol equation.**

e.g. $2Mg + O_2 \rightarrow 2MgO$

Equations need to be **balanced,** so the same number of atoms are on each side. To do this, numbers are put in front of the compounds.

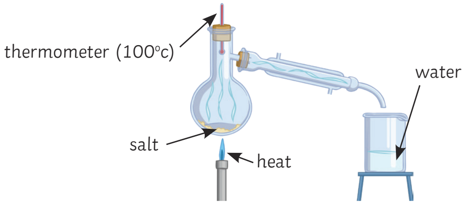
$CH_4 + \mathbf{4}O_2 \rightarrow \mathbf{2}H_2O + CO_2$

Atomic Structure and the Periodic Table – Foundation and Higher (Separate)

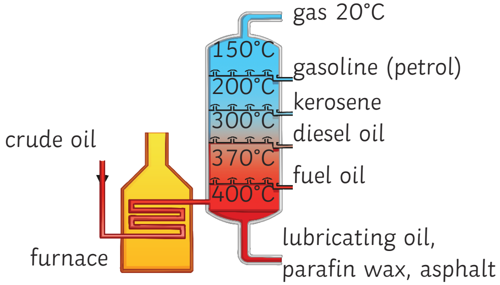
Distillation

To separate out mixtures of liquids.

1. **Simple distillation** – separating a liquid from a solution.



2. **Fractional distillation** – separating out a mixture of liquids. Fractional distillation can be used to separate out crude oil into fractions.



Metals and Non-metals

They are found at the **left** part of the periodic table. Non-metals are at the **right** of the table.

Metals

Are strong, malleable, good conductors of electricity and heat. They bond metallically.

Non-Metals

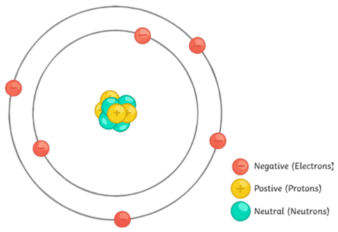
Are dull, brittle, and not always solids at room temperature.

History of the Atom

Scientist	Time	Discovery
John Dalton	start of 19th century	Atoms were first described as solid spheres.
JJ Thomson	1897	Plum pudding model – the atom is a ball of charge with electrons scattered.
Ernest Rutherford	1909	Alpha scattering experiment – mass concentrated at the centre; the nucleus is charged. Most of the mass is in the nucleus. Most atoms are empty space.
Niels Bohr	around 1911	Electrons are in shells orbiting the nucleus.
James Chadwick	around 1940	Discovered that there are neutrons in the nucleus.

Electronic Structure

Electrons are found in shells. A maximum of two in the most inner shell, then eight in the 2nd and 3rd shell. The inner shell is filled first, then the 2nd then the 3rd shell.



Group 7 Elements and Noble Gases

Halogens

The halogens are **non-metals**: fluorine, chlorine, bromine, iodine. As you go down the group they become less reactive. It is harder to gain an extra electron because its outer shell is further away from the nucleus. The melting and boiling points also become higher.

Noble Gases

The **noble gases** (**group 0** elements) include: **helium**, **neon** and **argon**. They are un-reactive as they have full outer shells, which makes them very stable. They are all colourless gases at room temperature.

The boiling points all increase as they go down the group – they have greater intermolecular forces because of the increase in the number of electrons.

Development of the Periodic Table

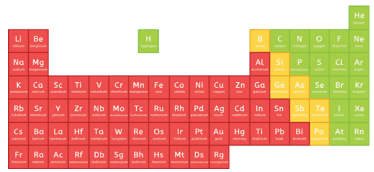
In the early 1800s, elements were arranged by atomic mass. The periodic table was not complete because some of the elements had not been found. Some elements were put in the wrong group.

Dimitri Mendeleev (1869) left gaps in the periodic table. He put them in order of **atomic mass**. The gaps show that he believed there was some undiscovered elements. He was right! Once found, they fitted in the pattern.

The Modern Periodic Table

Elements are in order of **atomic mass/proton number**. It shows where the metals and non-metals are. **Metals** are on the **left** and **non-metals** on the **right**. The **columns** show the **groups**. The **group number** shows the number of **electrons** in the **outer shell**. The rows are **periods** – each period shows another full shell of electrons.

The periodic table can be used to predict the reactivity of elements.



Alkali Metals

The alkali metals (**group** 1 elements) are soft, very reactive metals. They all have **one electron** in their **outer shell**, making them **very reactive**. They are **low density**. As you go down the group, they become more reactive. They get bigger and it is easier to lose an electron that is further away from the nucleus.

They form ionic compounds with non-metals.

The react with water and produce hydrogen.

E.g.

lithium + water → lithium hydroxide + water

$2Li + 2H_2O \rightarrow LiOH + H_2$

They react with chlorine and produce salt.

E.g.

lithium + chlorine → lithium chloride

$2Li + Cl_2 \rightarrow 2NaCl$

They react with oxygen to form metal oxides.

Atomic Structure and the Periodic Table – Foundation and Higher (Separate)

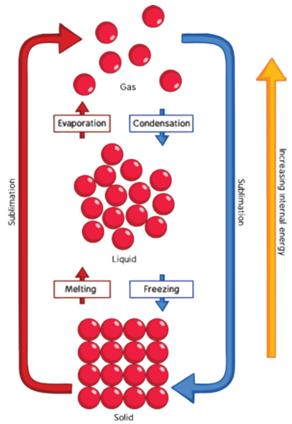
The Transition Metals

The transition metals are a block of elements found between groups 2 and 3 in the middle of the periodic table. Examples of transition metals include copper, nickel and iron with many more included. They have all the properties you would expect metals to have, such as being strong, shiny and conductors of electricity and heat. Transition metals make very good catalysts; this means they speed up a reaction without being used up themselves. Iron is used as a catalyst during the Haber process when making ammonia.

Transition metals can form more than one ion. For example, copper can take the form of Cu^{+}, Cu^{2+} and iron can be Fe^{2+} and Fe^{3+}. The ions are often coloured and the compounds they are found in are also coloured.

																	He helium
Li lithium	Be beryllium						H hydrogen					B boron	C carbon	N nitrogen	O oxygen	F fluorine	Ne neon
Na sodium	Mg magnesium											Al aluminium	Si silicon	P phosphorous	S sulfur	Cl chlorine	Ar argon
K potassium	Ca calcium	Sc scandium	Ti titanium	V vanadium	Cr chromium	Mn manganese	Fe iron	Co cobalt	Ni nickel	Cu copper	Zn zinc	Ga gallium	Ge germanium	As arsenic	Se selenium	Br bromine	Kr krypton
Rb rubidium	Sr strontium	Y yttrium	Zr zirconium	Nb niobium	Mo molybdenum	Tc technetium	Ru ruthenium	Rh rhodium	Pd palladium	Ag silver	Cd cadmium	In indium	Sn tin	Sb antimony	Te tellurium	I iodine	Xe xenon
Cs caesium	Ba barium	La lanthanum	Hf hafnium	Ta tantalum	W tungsten	Re rhenium	Os osmium	Ir iridium	Pt platinum	Au gold	Hg mercury	Ti thallium	Pb lead	Bi bismuth	Po polonium	At astatine	Rn radon
Fr francium	Ra radium	Ac actinium	Rf rutherfordium	Db dubnium	Sg seaborgium	Bh bohrium	Hs hassium	Mt meitnerium	Ds darmstadtium	Rg roentgenium							

AQA GCSE Chemistry (Separate Science) Unit 2: Bonding, Structure and Properties of Matter



The three states of matter are **solid, liquid and gas**.

For a substance to change from one state to another, **energy** must be **transferred**.

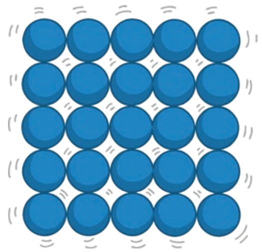
The particles gain energy. This results in the breaking of some of the **attractive forces** between particles during melting.

To evaporate or boil a liquid, more energy is needed to overcome the remaining chemical bonds between the particles.

Note the difference between **boiling** and **evaporation**. When a liquid **evaporates**, particles **leave the surface** of the liquid **only**. When a liquid **boils**, **bubbles** of gas form **throughout** the liquid before rising to the surface and escaping.

The amount of energy needed for a substance to change state is dependent upon the **strength** of the **attractive forces** between particles. The **stronger** the **forces of attraction**, the **more energy** needed to **break them apart**. Substances that have strong attractive forces between particles generally have **higher melting and boiling points**.

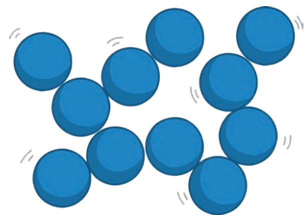
Solid



The particles in a **solid** are arranged in a regular pattern. The particles in a solid **vibrate** in a fixed position and are tightly packed together. The particles in a solid have a **low amount of kinetic energy**.

Solids have a **fixed shape** and are unable to flow like liquids. The particles **cannot be compressed** because the particles are very close together.

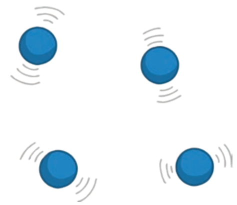
Liquid



The particles in a **liquid** are randomly arranged. The particles in a liquid are able to **move around** each other. The particles in a liquid have a **greater amount** of **kinetic energy** than particles in a **solid**.

Liquids are able to **flow** and can take the shape of the container that they are placed in. As with a solid, liquids **cannot be compressed** because the particles are close together.

Gas



The particles in a **gas** are randomly arranged. The particles in a gas are able to **move around very quickly** in all directions. Of the three states of matter, gas particles have the **highest amount** of **kinetic energy**.

Gases, like liquids, are able to **flow** and can fill the container that they are placed in. The particles in a gas are **far apart** from one another which allows the particles to move in any direction.

Gases can be **compressed**; when squashed, the particles have empty space to move into.

Limitations of the Particle Model (HT only)

The chemical bonds between particles are not represented in the diagrams above.

Particles are represented as solid spheres – this is not the case. Particles like atoms are mostly empty space. Particles are not always spherical in nature.

State Symbols

In chemical equations, the three states of matter are represented as symbols:

solid **(s)**
liquid **(l)**
gas **(g)**
aqueous **(aq)**

Aqueous solutions are those that are formed when a substance is dissolved in water.

Identifying the Physical State of a Substance

If the given temperature of a substance is **lower** than the **melting point**, the physical state of the substance will be **solid**.

If the given temperature of the substance is **between** the **melting point and boiling point**, the substance will be a **liquid**.

If the given temperature of the substance is **higher** than the **boiling point**, the substance will be a **gas**.

Formation of Ions

Ions are charged particle. They can be either positively or negatively charged, for example Na^+ or Cl^-.

When an element loses or gains electrons, it becomes an ion.

Metals **lose** electrons to become **positively charged**.

Non-metals **gain** electrons to become **negatively charged**.

Group 1 and 2 elements **lose** electrons and group 6 and 7 elements **gain** electrons.

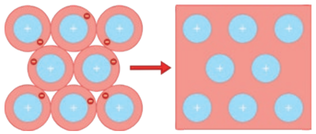
Group	Ions	Element Example
1	+1	$Li \rightarrow Li^+ + e^-$
2	+2	$Ca \rightarrow Ca^{2+} + 2e^-$
6	-2	$Br + e^- \rightarrow Br^-$
7	-1	$O + 2e^- \rightarrow O^{2-}$

Metals and Non-metals

Metals are found on the **left-hand side** of the **periodic table**. Metals are strong, shiny, malleable and good conductors of heat and electricity. On the other hand, non-metals are brittle, dull, not always solids at room temperature and poor conductors of heat and electricity. **Non-metals** are found on the **right-hand side** of the **periodic table**.

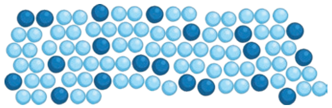
Metallic Bonding

Metallic bonding occurs between **metals only**. Positive metal ions are surrounded by a **sea of delocalised electron**s. The ions are tightly packed and arranged in rows.



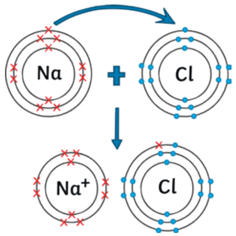
There are strong electrostatic forces of attraction between the positive metal ions and negatively charged electrons.

Pure metals are too soft for many uses and are often mixed with other metals to make alloys. The mixture of the metals introduces different-sized metal atoms. This **distorts the layers** and **prevents them from sliding over one another**. This makes it harder for alloys to be bent and shaped like pure metals.



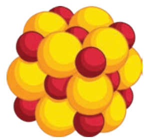
Ionic Bonding

Ionic bonding occurs between a metal and a non-metal. Metals lose electrons to become positively charged. Opposite charges are attracted by electrostatic forces – an ionic bond.



Ionic Compounds

Ionic compounds form structures called giant lattices. There are **strong electrostatic forces of attraction** that **act in all directions** and act between the **oppositely charged ions** that make up the giant ionic lattice.



Properties of Ionic Compounds

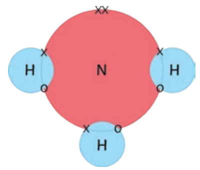
- High melting point – lots of energy needed to overcome the electrostatic forces of attraction.
- High boiling point
- **Cannot conduct electricity** in a **solid** as the ions are not free to move.
- Ionic compounds, when **molten** or in **solution**, can **conduct electricity** as the ions are free to move and can carry the electrical current.

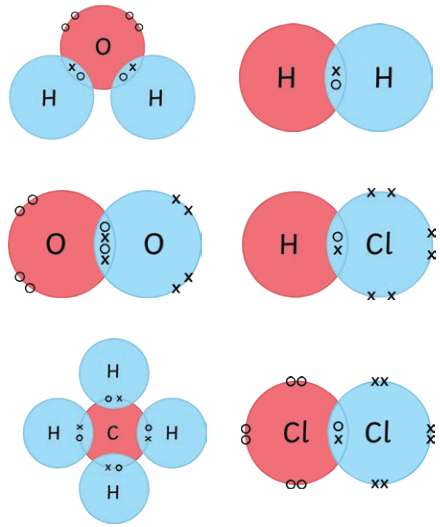
Covalent Bonding

Covalent bonding is the sharing of a pair of electrons between atoms to gain a full outer shell. This occurs between **non-metals only**. Simple covalent bonding occurs between the molecules below. Simple covalent structures have **low melting and boiling points** – this is because the **weak intermolecular forces** that hold the molecules together break when a substance is heated, not the strong covalent bonds between atoms. They **do not conduct electricity** as they do not have any free delocalised electrons.

Dot and cross diagrams are useful to show the **bonding** in **simple molecules**. The **outer electron shell** of each atom is represented as a **circle**, the circles from each atom overlap to show where there is a **covalent bond**, and the electrons from each atom are either drawn as **dots or crosses**. There are **two different types** of **dot and cross diagram** – one with a circle to represent the outer electron shell and one without.

You should be able to draw the dot and cross diagrams for the following simple covalent structures: chlorine, oxygen, nitrogen, water, ammonia, hydrogen chloride and methane.

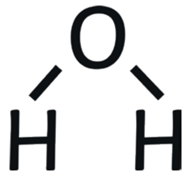




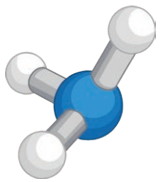
AQA GCSE Chemistry (Separate Science) Unit 2: Bonding, Structure and Properties of Matter

Structural Formulae

In this type of diagram, the element symbol represents the type of atom and the straight line represents the covalent bonding between each atom.



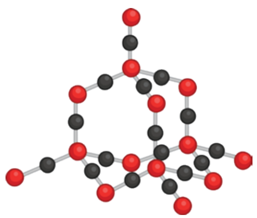
The structure of small molecules can also be represented as a 3D model.



Giant Covalent Structure – Diamond

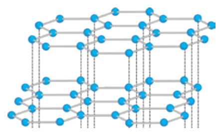
Each **carbon** atom is **bonded** to **four** other carbon atoms, making diamond very strong. Diamond has a high melting and boiling point. **Large** amounts of **energy** are needed to break the strong covalent bonds between each carbon atom. Diamond **does not conduct** electricity because it has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

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



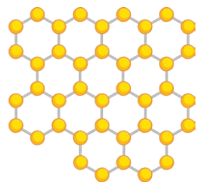
Giant Covalent Structure – Graphite

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



Giant Covalent Structure – Graphene

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



Nanoscience

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

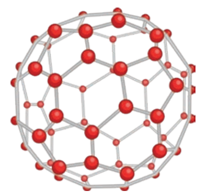
Name of Particle	Diameter
nanoparticle	1–100nm
fine particles ($PM_{2.5}$)	100–2500nm
coarse particles (PM_{10})	2500–10000nm

Polymers

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

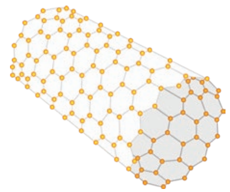
Fullerenes and Nanotubes

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



Buckminsterfullerene has the formula C_{60}

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



Possible Risks of Nanoparticles

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

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

Relative Formula Mass (M_r)

The **relative atomic mass (A_r)** of an element is an elements relative mass compared to the mass of an atom of carbon-12. A_r values are given in the periodic table.

The **relative formula mass (M_r)** of a compound is the **sum** of all the relative atomic masses (A_r) of the atoms in the formula.

Example 1: hydrochloric acid (HCl) consists of one hydrogen atom (A_r 1) and one chlorine atom (A_r 35.5).

The M_r of HCl = 1 + 35.5 = 36.5

Example 2: sulfuric acid (H_2SO_4) consists of two hydrogen atoms (A_r 1), one sulfur atom (A_r 32) and four oxygen atoms (A_r 16).

The M_r of H_2SO_4 = (1 × 2) + 32 + (16 × 4) = 98

Neither A_r or M_r values have any units.

Law of Conservation of Mass

The **law of conservation of mass** states that during a chemical reaction, no atoms are lost or made.

For example: $2Mg + O_2 \longrightarrow 2MgO$

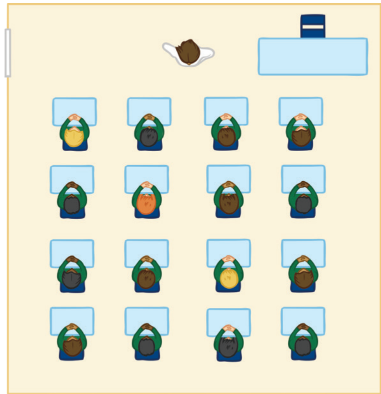
In a chemical reaction, mass is never lost or gained. What **goes in** must **come out**. The **total mass of the reactants** at the beginning of the chemical reaction **equals** the **total mass of the products** made at the end of the reaction.

For example, imagine if we used building bricks to represent the atoms in a chemical reaction: atoms, like building bricks, can be completely rearranged. However, the total mass of the atoms will stay the same. Rearranging the building blocks in different structures takes a little **energy**, just like in a chemical reaction.

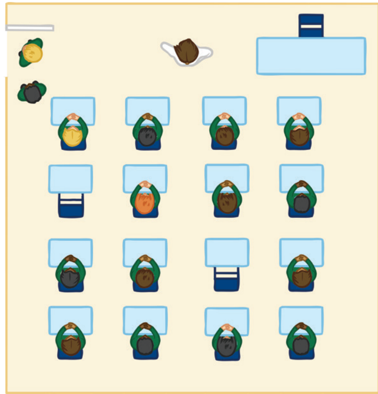
Reactions in Closed and Non-Enclosed Systems

If a reaction occurs in a **closed system**, the **mass** in a chemical reaction will remain **constant**.

In an **non-enclosed system, changes in mass can occur,** such as when a gas is released. It is important to remember that **no atoms are created or destroyed,** they are just **rearranged**. If a gas escapes a non-enclosed system, the total mass will look as if it has decreased. Similarly, if a gas is gained, the total mass will look as if it has increased. However, the **total mass** will **remain the same** if the mass of the gas is included in the reaction calculation.



In this **closed system** (the classroom), the mass in the reaction remains constant. As the system is a closed one, no children are allowed to leave or enter.



In this **non-enclosed system** (the classroom), the mass in the reaction can look as if it has changed as children are allowed to leave the classroom at any time.

Uncertainty

Whenever a measurement is made, there is always some degree of **uncertainty** about the result. Uncertainty is a **measure** of the **variability** in scientific data.

Uncertainty can be measured by considering the **resolution** of the scientific equipment being used or from the **range** of a set of scientific data.

There are two types of errors: **random error** and **systematic error.**

Random errors may be caused by **human error** such as a poor technique when taking measurements or by **equipment** that is **faulty**. For example, three mass balances all showing different mass values for the same object. Random errors are **random** and not something that can be predicted.

Systematic errors are errors that are produced **consistently**: if the experiment is repeated, the **same error** will occur. For example, not taring a mass balance properly or problems with the experimental method.

$$\text{uncertainty} = \frac{\text{range of results}}{2}$$

The **range** is the difference between the **largest** and **smallest** value.

For example, student A carried out a practical to determine how much dilute sulfuric acid is needed to react with exactly $50.0cm^3$ of a sodium hydroxide solution.

Repeat	1	2	3	Mean
Volume of H_2SO_4 needed to react with $50.0cm^3$ of NaOH.	23.13	24.00	23.56	23.56

Calculate the range:

range = 24.00 – 23.13 = $0.87cm^3$

Calculate the uncertainty of the mean:

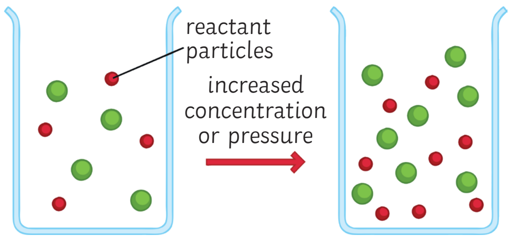
uncertainty = 0.87 ÷ 2 = $0.44cm^3$

The mean with uncertainty:

23.56 ± $0.44cm^3$

Concentration of Solutions

Concentration is a **measure** of the amount of a **substance** in a **volume** of liquid. The higher the concentration, the more particles of a substance are present in the solution.



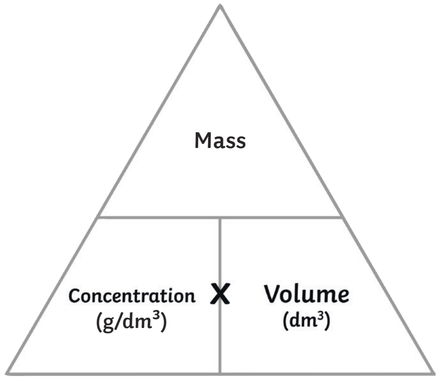
In chemistry, there are two ways to measure the concentration of a solution. This can be done by calculating the **mass** of the substance in grams or by calculating the number of **moles**.

In order to calculate concentration, you must be working in dm^3.

If it is not, it may mean that you need to do a conversion.

$cm^3 \longrightarrow dm^3$ = ÷ 1000

$m^3 \longrightarrow dm^3$ = × 1000



Calculate the **concentration** of a solution with a mass of 2.15g and a volume of $5dm^3$.

concentration = mass ÷ volume

concentration = 2.15g ÷ $5dm^3$

concentration = $0.43g/dm^3$

Calculate the **mass** of sodium chloride that you would need to dissolve in $400cm^3$ of water to make a $20g/dm^3$ volume solution.

mass = concentration × volume

convert $cm^3 \longrightarrow dm^3$

$400cm^3$ ÷ 1000 = $0.40dm^3$

mass = $20g/dm^3$ × $0.40dm^3$ = 8g

Calculate the **volume** of liquid required to add to 8.80g of a solid to make $42g/dm^3$ solution.

volume = mass ÷ concentration

volume = 8.80g ÷ $42g/dm^3$

concentration = $0.210dm^3$

The Mole – Higher Tier Only

When we talk about moles, we are not talking about the moles that live underground.

A **mole** (mol) is a **measurement** that is used in chemistry.

Example 1

Look at this reaction:

$2Mg + O_2 \longrightarrow 2MgO$

The reaction shows that **two moles** of magnesium react with oxygen to produce **two moles** of magnesium oxide. Using moles in a **balanced symbol equation** shows the **ratio** of **reactants** to **products**.

Avogadro's Constant

1 mole = 6.02×10^{23}

The number is known as **Avogadro's constant or Avogadro's number** and is named after the Italian scientist Amedeo Avogadro. The mole is abbreviated to **mol**.

This number is very important and one that you should remember. The mass of one mole of a substance in grams is equal to its relative formula mass. For example, one mole of carbon-12 has a mass of 12g

A mole is the amount of a substance that contains **6.02×10^{23}** particles of that substance. The particles could be atoms, molecules, ions or electrons.

For example, 1 mole of carbon will contain the same number of atoms **(6.02×10^{23})** as you would have molecules in 1 mole of water.

AQA GCSE Chemistry (Separate Science) Unit 3: Quantitative Chemistry

Calculating the Number of Particles

The number of particles can be calculated using Avogadro's constant if the number of moles is known.

In chemistry, Avogadro's constant is given the symbol N_A. To calculate the number of particles in a substance, the following equation can be used:

$N = n \times N_A$

N = the number of particles in a substance

n = the number of moles (mol)

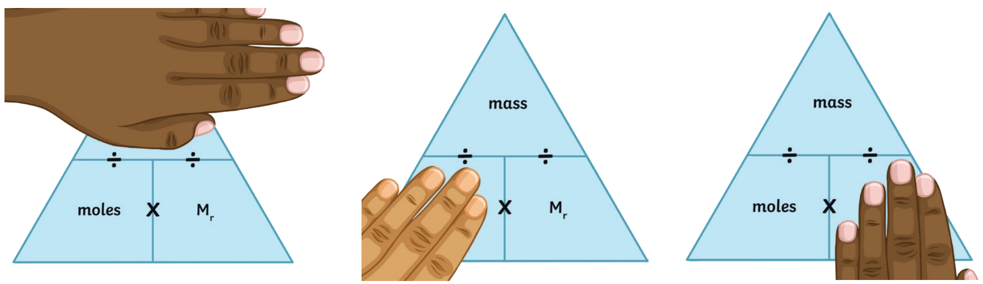
N_A = Avogadro's constant 6.02×10^{23}

For example, calculate the number of helium molecules in 10 mol of helium.

$N = n \times N_A$

N = $10 \times (6.02 \times 10^{23}) = 6.022 \times 10^{24}$

Calculating Moles, Mass and M_r



Calculating Moles, Mass and M_r

Calculate the number of **moles** in 330g of K_2S.

K_2S consists of two potassium atoms (A_r 39) and one sulfur atom (A_r 32).

Calculate the M_r of the compound =
$(39 \times 2) + 32 =$ **110**

moles = mass ÷ M_r

moles = 330 ÷ 110 = 3 moles

Calculate the **mass** of 0.9 moles of $Fe(NO_3)_3(H_2O)_9$.

Calculate the Mr of the compound.

$(16 \times 3) + 14 = 62$

$62 \times 3 =$ **186**

$(1 \times 2) + 16 = 18$

$18 \times 9 =$ **162**

$56 + 186 + 162 =$ **404**

mass = moles × M_r

mass = 0.9 × 404 = 363.6g

Relative Atomic Mass (A_r)

iron (Fe) = 56

oxygen (O) = 16

nitrogen (N) = 14

hydrogen (H) = 1

Amount of Substances in Equations – Higher Tier Only

How do we know the masses involved in the equation?

To work out the masses involved, write in the relative atomic mass (A_r) for an element and the relative formula mass (M_r) for a compound.

Example

Step 1: Write down the **balanced** symbol equation. $2Mg + O_2 \longrightarrow 2MgO$

Step 2: Write in the relative atomic and relative formula masses for the **reactants** and **products** involved in the chemical reaction.

$(2 \times 24) + (2 \times 16) \longrightarrow 2 \times (24 + 16)$

$48g + 32g \longrightarrow 80g$

Masses in Equations

$Fe + \mathbf{2}HCl \longrightarrow FeCl_2 + H_2$

$56 + 2\,(1 + 35.5) \longrightarrow (56 + 2(35.5) + (1 \times 2))$

$56g + 73g \longrightarrow 127g + 2g$

One mole of iron **reacts** with two moles of hydrochloric acid to **produce** one mole of iron chloride and one mole of hydrogen.

Calculate the **mass of water** made when burning **300g of methane**.

Step 1: Balance the equation.

$CH_4 + \mathbf{2}O_2 \longrightarrow \mathbf{2}H_2O + CO_2$

Step 2: Write down the relative formula mass of each compound.

$(12 + (1 \times 4)) + 2(16 \times 2) \longrightarrow 2((1 \times 2) + 16) + (12 + (16 \times 2))$

$16g + 64g \longrightarrow 36g + 44g$

Relative Atomic Mass (A_r)

Carbon (C) = 12

oxygen (O) = 16

hydrogen (H) = 1

We know from the equation that **16g of methane** reacts to produce **36g of water**.

The question asks us to calculate the mass of water made when burning **300g of methane**.

$\frac{\text{known mass}}{M_r} \times M_r \text{ of unknown mass}$ $\qquad \frac{300}{16} \times 36 = 675g$

AQA GCSE Chemistry (Separate Science) Unit 3: Quantitative Chemistry

Limiting Reactants

A chemical reaction ends once one of the **reactants** is used up. The other reactants have nothing to react with and so some are left over.

The **limiting reactant** is the reactant that is **completely used up** in a chemical reaction. This reactant is the one that determines the amount of product that is made.

The reactant in **excess** is the one that is left over and could further react if there was another reactant to react with.

The **amount of product** that is produced during a chemical reaction is **dependent** upon the **amount of the limiting reactant**.

Calculating the maximum mass of a product formed during a chemical reaction can be done by the following:

a. Writing a balanced equation.

b. Calculating the mass (g) of the limiting reactant.

c. The A_r and M_r of the product and limiting reactant.

Determine the **maximum mass of hydrogen** that can be produced when 36g of magnesium (Mg A_r 24) reacts completely with excess hydrochloric acid (HCl) to produce magnesium chloride ($MgCl_2$) and hydrogen (H_2).

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

number of moles = mass ÷ A_r

number of moles = 36 ÷ 24 = 1.5 mol

From the equation, 1 mol of magnesium forms 1 mol of hydrogen. Therefore, 1.5 mol of magnesium forms 1.5 mol of hydrogen.

mass of hydrogen = M_r × number of moles

= 2 × 1.5

= 3g

Balancing Equations

By using the masses of the products and reactants, it is possible to work out the balancing numbers in an equation.

For example, 12g of magnesium (Mg A_r 24) reacts with 8g of oxygen (O_2 M_r 32) to produce magnesium oxide (MgO M_r 40). Determine the balanced symbol equation for the reaction.

Calculate the amount of each of the reactants.

Mg = 12 ÷ 24 = 0.5 mol

O_2 = 8 ÷ 32 = 0.25 mol

Divide both values by the smaller amount.

Mg = 0.5 ÷ 0.25 = 2

O_2 = 0.25 ÷ 0.25 = 1

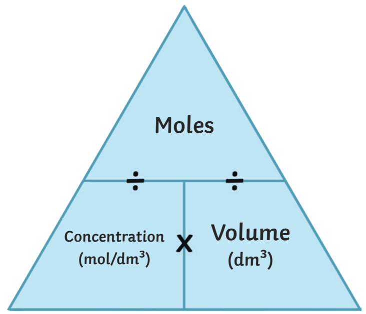
The equation shows that on the left-hand side of the equation, 2 mol of the reactant (Mg) reacts with 1 mol of oxygen. Using this information, it is then possible to balance the rest of the equation in the normal way.

$$2Mg + O_2 \longrightarrow 2MgO$$

Calculating Concentrations

The concentration of a solution can have the units g/dm^3 or mol/dm^3.

Concentration can be calculated using the mass of dissolved solute or the volume of the solvent or solution in dm^3.



Example:

Student A dissolved 1 mol of sodium hydroxide in 4dm^3 of water. Determine the concentration of the sodium hydroxide solution he made.

concentration = 1 mol ÷ 4dm^3

concentration = 0.25mol/dm^3

Converting between Units

To convert between g/dm^3 and mol/dm^3, the relative formula mass of the solute is used.

Multiply by the M_r to convert from **mol/dm^3 to g/dm^3**.

Divide by the M_r to convert from **g/dm^3 to mol/dm^3**.

Example:

Determine the concentration of 0.8mol/dm^3 sodium hydroxide **(M_r 40)** solution in g/dm^3.

concentration = 0.8 × 40 = 32g/dm^3

Volumes of Solutions

By rearranging the concentration equation, it is possible to calculate the amount of a solute in a given volume of solution if the concentration is known.

amount of solute (mol) = concentration (mol/dm^3) × volume (dm^3)

Example:

Determine the amount of 0.2mol/dm^3 sodium hydroxide in 75cm^3 of solution.

Step 1: Convert the volume to dm^3.

75cm^3 = 75.0 ÷ 1000 = 0.075dm^3

Step 2: amount of solute (mol) = concentration (mol/dm^3) × volume (dm^3)

= 0.2mol/dm^3 × 0.075dm^3

= 0.015 mol

Calculating the Mass

Using the example above, calculate the mass of sodium hydroxide (M_r 40) in 75cm^3 of solution.

mass = amount × M_r

mass = 0.015 mol × 40

mass = 0.6g

Percentage Yield – Chemistry Only

The percentage yield can be calculated from the following equation.

$$\text{percentage yield} = \frac{\text{actual mass of product made}}{\text{maxium theoretical mass of product}} \times 100$$

The **theoretical yield** is **the maximum mass** that can be made during a chemical reaction. The law of conservation states that during a chemical reaction, no atoms are lost or made. It's not always possible to obtain the maximum calculated amount of product.

The loss of product may be due to some of the product being lost when filtered. Some of the reactants may not react as expected and so may not produce enough product. The reaction may be a reversible one and as a consequence, the reaction may not go to completion.

Example:

1.8g of copper sulfate crystals are made during a chemical reaction. The theoretical yield for this reaction is **2.0g**. Calculate the percentage yield of copper sulfate.

$$\text{percentage yield} = \frac{1.8\text{g}}{2.0} \times 100$$

percentage yield = 90%

Atom Economy – Chemistry Only

The percentage atom economy can be calculated from the following equation.

$$\text{atom economy} = \frac{\text{relative formula mass of desired product from equation}}{\text{sum of relative formula masses of all reactants from equation}}$$

The **atom economy** is a measure of the amount of starting materials (reactants) that end up as **useful products**. It is important for sustainable development and for economic reasons to use reactions with **high atom economy**. However, not all atoms end up as the desired product and may form other products. We call these **byproducts**.

Example:

When glucose (M_r 180) is fermented, ethanol (M_r 46) is produced.

$C_6H_{12}O_6 \longrightarrow 2CH_3CH_2OH + 2CO_2$

Calculate the atom economy for this reaction.

$$\text{atom economy} = \frac{2 \times 46}{180} \times 100$$

atom economy = 51.1%

Reaction Pathways – Higher Tier Only

There is often more than one way to make a substance. Reaction pathways **describe** the **reactions** that have taken place to form the **desired product**. Choosing a particular pathway is dependent upon a number of factors:

1. percentage yield
2. atom economy
3. rate of reaction
4. position of the equilibrium
5. usefulness of any byproducts

The raw materials needed for a particular reaction may affect its chosen pathway. For example, crude oil is a **non-renewable resource**; the resource will run out if we continue to use it. However, plant sugars are **renewable** and can be replenished as long as other plants are replanted.

Ethanol can be made through the fermentation of glucose or the hydration of ethene.

Method of Ethanol Production	Percentage Yield (%)	Atom Economy (%)	Rate of Reaction
fermentation	15	51.1	low
hydration	95	100	high

The hydration of ethene has a 100% atom economy; all atoms react to form the desired product. On the other hand, fermentation has an atom economy of 51.1%. However, its rate of reaction is low in comparison to the hydration method and only has a percentage yield of 15%. Therefore, hydration is the best method for making ethanol.

A byproduct of the fermentation process is carbon dioxide. The gas is sold to fizzy drinks manufacturers to provide the bubbles for some well known fizzy drinks. As the byproduct produced is one that can be useful, it means that the atom economy can be increased to 100%.

Ethene hydration is a reversible reaction. The position of the equilibrium lies to the left. Therefore, only 5% of the ethene supplied to the reaction is actually converted to ethanol. A 95% yield is achieved by recirculating the unreacted ethene.

Avogadro's Law – Higher Tier Only

When the temperature and pressure stay the same, Avogadro's law states that different gases that have the same volume contain equal numbers of molecules.

For example, 1 mol of methane gas occupies the same volume as 1 mol of argon gas.

$H_2 + Cl_2 \longrightarrow 2HCl$

When hyrdogen and chlorine react, hydrogen chloride is produced. In terms of the molar ratio, 10cm^3 of hydrogen reacts completely with 10cm^3 of chlorine. Therefore, the ratio between hydrogen and chlorine is 1:1.

The molar ratio between hydrogen and hydrogen chloride is 1:2. For example, 10cm^3 of hydrogen reacts to produce 20cm^3 of hydrogen.

Molar Gas Volume

The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmosphere pressure) is 24dm^3 (**24 000 cm^3**).

To calculate a known volume of a gas:

volume = amount in mol × molar volume

For example, determine the volume of 0.55 mol of carbon monoxide at room temperature and pressure.

volume = amount in mol × molar volume

volume = 0.55 × 24

= 13.2dm^3

Calculating the Amount of Gas

By **rearranging the equation,** it is possible to calculate the amount of a gas in moles.

For example, determine the amount of hydrogen gas that occupies 198cm^3 at room temperature and pressure.

$$\text{amount in mol} = \frac{\text{volume}}{\text{molar volume}}$$

$$\text{amount in mol} = \frac{198}{24\ 000}$$

amount in mol = 0.0083 mol

Calculating a Volume from a Mass

When 3.5g of sodium reacts with water it produces sodium hydroxide and hydrogen gas.

$2Na + 2H_2O \longrightarrow 2NaOH + H_2$

1. Determine the molar amount of sodium (A_r 23).

$$\text{amount in mol} = \frac{\text{mass}}{\text{atomic mass}}$$

$$\text{amount in mol} = \frac{3.5}{23}$$

amount in mol = 0.15 mol

2. Determine the molar amount of hydrogen.

The molar ratio of **sodium to hydrogen**, according to the balanced symbol equation, is **2:1**.

Therefore, 0.15 mol of sodium produces 0.075 mol of hydrogen.

3. Determine the volume of hydrogen.

volume = amount in mol × molar volume

volume = 0.075 × 24dm^3

= 1.8dm^3

4. Calculating the mass from a volume.

Sodium reacts with chlorine to produce sodium chloride.

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

5. Determine the mass of sodium chloride (Mr 58.5) that can be produced from 685cm^3 of chlorine.

amount of chlorine = 685cm^3 ÷ 24 000 = 0.029 mol

From the equation, the mole ratio between chlorine and sodium chloride is l:2. Therefore, 0.029 moles of chlorine would produce (0.029 × 2) = 0.058 mol.

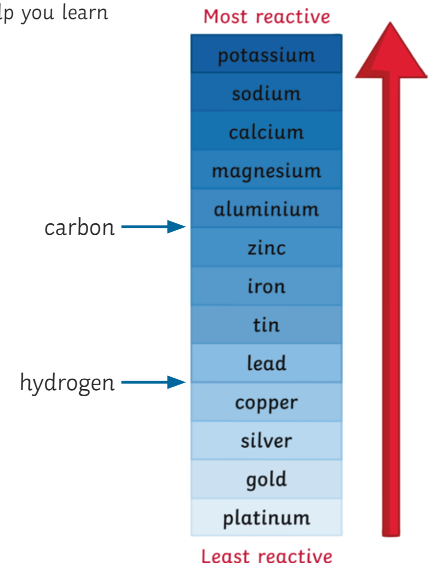
mass of sodium chloride = 0.058 × 58.5 = 3.393g

AQA GCSE Chemistry (Separate Science) Unit 4: Chemical Changes

The Reactivity Series

Here's a **mnemonic** to help you learn the order.

purple (potassium)
slime (sodium)
can (calcium)
make (magnesium)
a (aluminium)
careless (carbon)
zebra (zinc)
insane (iron)
try (tin)
learning (lead)
how (hydrogen)
camels (copper)
surprise (silver)
gorillas (gold)



The reactivity series is a league table for metals. The **more reactive** metals are near the **top** of the table with the **least reactive** near the **bottom**. In chemical reactions, a more reactive metal will displace a less reactive metal.

Reactions of Metals with Water

Metals, when reacted with water, produce a metal hydroxide and hydrogen.

lithium + water → lithium hydroxide + hydrogen

$2Li + 2H_2O \rightarrow 2LiOH + H_2$

The more reactive a metal is the faster the reaction.

Reactions of Metals with Dilute Acid

Metals, when reacted with acids, produce a **salt** and **hydrogen**.

Sodium + hydrochloric acid → sodium chloride + hydrogen

$2Na + 2HCl \rightarrow 2NaCl + H_2$

Metals that are below hydrogen in the reactivity series **do not** react with dilute acids.

Reactions of Acids

The general formula for the reaction between an acid and a metal is:

acid + metal → salt + hydrogen

For example: hydrochloric acid + sodium → sodium chloride + hydrogen

$2HCl + 2Na \rightarrow 2NaCl + H_2$

When an acid reacts with an alkali, a neutralisation reaction takes place and a salt and water are produced.

The general formula for this kind of reaction is acid + alkali → salt + water

hydrochloric acid + sodium hydroxide → sodium chloride + water

$HCl + NaOH \rightarrow NaCl + H_2O$

Naming Salts

The first part comes from the metal in the metal carbonate, oxide or hydroxide. The second part of the name comes from the acid that was used to make it.

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

For example, sodium chloride.

Redox Reactions (Higher Tier Only)

When metals react with acids, they undergo a redox reaction. A **redox reaction** occurs when both **oxidation** and **reduction** take place at the **same time**.

For example:

$2H^+ + Ca \rightarrow Ca^{2+} + H_2$

The ionic equation can be further split into two half equations.

$Ca \rightarrow Ca^{2+} + 2e^-$

Oxidation is loss of electrons.

$2H^+ + 2e^- \rightarrow H_2$

Reduction is gaining of electrons.

Reactions with Bases

The general formula for the reaction between an acid and a metal oxide is:

acid + metal oxide → salt + water

sulfuric acid + copper oxide → copper sulfate + water

$H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$

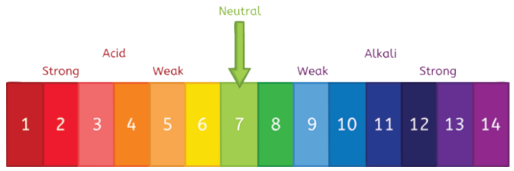
Reactions with Carbonates

The general formula for the reaction between an acid and a carbonate is:

acid + carbonate → salt + water + carbon dioxide

hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide

pH Scale



In aqueous solutions, acids produce H^+ ions and alkalis produce OH^- ions.

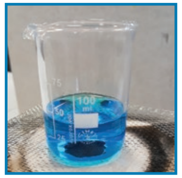
Neutral solutions are pH7 and are neither acids nor alkalis.

For example, in neutralisation reactions, hydrogen ions from an acid react with hydroxide ions from an alkali to produce water:

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

Making Soluble Salts

1. Make a saturated solution by stirring copper oxide into the sulfuric acid until no more will dissolve.

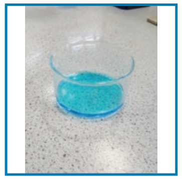


2. Filter the solution to remove the excess copper oxide solid.

3. Half fill a beaker with water and set this over a Bunsen burner to heat the water. Place an evaporating dish on top of the beaker.

4. Add some of the solution to the evaporating basin and heat until crystals begin to form.

5. Once cooled, pour the remaining liquid into a crystallising dish and leave to cool for 24 hours.



6. Remove the crystals with a spatula and pat dry between paper towels.

Strong and Weak Acids (Higher Tier Only)

A **strong** acid **completely dissociates** in a solution. For example: $HCl \rightarrow H^+ + Cl^-$

Hydrochloric acid is able to completely dissociate in solution to form hydrogen and chloride ions.

Examples of strong acids include nitric acid (HNO_3) and sulfuric acid (H_2SO_4).

Weak acids in comparison only **partially dissociate**.

For example, acetic acid partially dissociates to form a hydrogen and acetate ion.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

The **double arrow** symbol indicates that the reaction is **reversible**.

The Process of Electrolysis

Electrolysis is the **splitting up** of an ionic substance using **electricity**.

On setting up an electrical circuit for electrolysis, two **electrodes** are required to be placed in the electrolyte. The electrodes are **conducting rods**. One of the rods is connected to the **positive** terminal and the other to the **negative** terminal.

The **electrodes** are **inert** (this means they do not react in the reaction) and are often made from **graphite** or platinum.

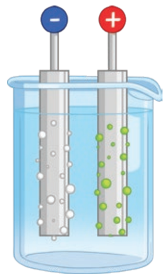
During the process of electrolysis, **opposites attract**. The positively-charged ions will be attracted toward the negative electrode. The negatively-charged ions will be attracted towards the positive electrode.

When ions reach the electrodes, the charges are lost and they become elements.

The **positive** electrode is called the **anode**.

The **negative** electrode is called the **cathode**.

Electrolysis of Aqueous Solutions



Gases may be given off or metals deposited at the electrodes. This is dependent on the reactivity of the elements involved.

If the metal is **more reactive** than **hydrogen** in the reactivity series, then **hydrogen** will be **produced** at the **negative cathode**. At the **positive anode**, negatively charged ions **lose** electrons. This is called **oxidation** and you say that the ions have been oxidised.

Using Electrolysis to Extract Metals

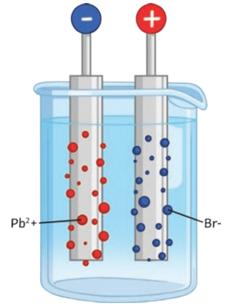
Metals are extracted by electrolysis if the metal in question reacts with carbon or if it is too reactive to be extracted by reduction with carbon. During the extraction process, large quantities of energy are used to melt the compounds.

Aluminium is manufactured by the process of electrolysis. Aluminium oxide has a high melting point and melting it would use large amounts of energy and increase the cost of the process. Therefore, molten **cryolite** is added to aluminium oxide to lower the melting point and thus reduce the cost.

Electrolysis of Molten Ionic Compounds – Lead Bromide

Lead bromide is an **ionic** substance. Ionic substances, when solid, are **not** able to conduct electricity. When molten or in solution, the ions are free to move and are able to carry a charge.

The **positive lead** ions are attracted toward the **negative cathode** at the same time as the **negative bromide** ions are attracted toward the **positive anode**.



Oxidation **i**s the **l**oss of electrons and **r**eduction **i**s the **g**aining of electrons. **OIL RIG (Higher Tier Only).**

We represent what is happening at the electrodes by using **half equations (Higher Tier Only).**

The lead ions are attracted towards the negative electrode. When the **lead ions** (Pb^{2+}) reach the cathode, each ion **gains two electrons** and becomes a neutral atom. We say that the lead ions have been **reduced**.

$Pb^{2+} + 2e^- \rightarrow Pb$

The bromide ions are attracted towards the positive electrode. When the **bromide ions** (Br-) reach the anode, each ion **loses one electron** to become a neutral atom. Two bromine atoms are then able to bond together to form the **covalent** molecule Br_2.

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

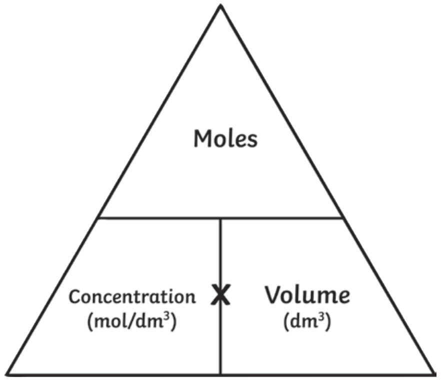
Titration Method (Chemistry Only)

1. Using the pipette and pipette filler, measure $25cm^3$ sodium hydroxide solution and pour into a conical flask.
2. Add several drops of phenolphthalein to the sodium hydroxide solution.
3. Swirl the flask and the mixture should be pink.
4. Place the conical flask on a white tile.
5. Place the burette into its stand, ensuring the tap is closed. Using the funnel, fill the burette with sulfuric acid to the $0cm^3$ line. Should you go above this line, open the tap and allow the excess to run off into a beaker.
6. Once the burette is correctly filled, place over the conical flask.
7. Carefully open the tap so the acid flows slowly into the conical flask. Swirl the flask and look for the indicator changing from pink to colourless.
8. Continue adding the acid to the flask until the indicator is permanently colourless.
9. Record the total volume of acid added to the sodium hydroxide in the results table.
10. Repeat the experiment twice more.

Titration Method (Chemistry Only)

Using the results from a titration experiment, it is possible to calculate the concentration of a solution or the volume of solution required to neutralise an acid or alkali.

Worked Example



In a titration, $20cm^3$ of $1.0mol/dm^3$ sulfuric acid reacted with $25cm^3$ of sodium hydroxide. What was the concentration of sodium hydroxide?

Write out the symbol equation for the reaction.

$NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

Check that the equation is balanced.

$\mathbf{2}NaOH + H_2SO_4 \rightarrow Na_2SO_4 + \mathbf{2}H_2O$

To convert cm^3 to dm^3, just divide by 1000.

Draw a table like the one below and fill it in with the information that you know from the question.

	Acid (H_2SO_4)	Alkali (NaOH)
number of moles		
concentration mol/dm^3	1.0	
volume (dm^3)	0.02	0.025

As the values for the **concentration** and **volume** of the acid are known, it is possible to now work out the **number of moles** of H_2SO_4.

number of moles = concentration × volume

number of moles = 1.0 × 0.02 = **0.02 moles**

From the balanced symbol equation, we know that there is double the amount of NaOH compared to H_2SO_4, therefore to calculate the number of moles of the alkali, we double the number of moles of the acid. 0.02 × 2 = **0.04 moles**.

	Acid (H_2SO_4)	Alkali (NaOH)
number of moles	0.02	0.04
concentration mol/dm^3	1.0	
volume (dm^3)	0.02	0.025

The question asks you to calculate the **concentration** of sodium hydroxide. As the number of moles and volume is now known, it is possible to calculate the concentration.

concentration = number of moles ÷ volume

concentration = 0.04 ÷ 0.025

concentration = **$1.6mol/dm^3$**

AQA GCSE Chemistry (Separate Science) Unit 5 Energy Changes Knowledge Organiser

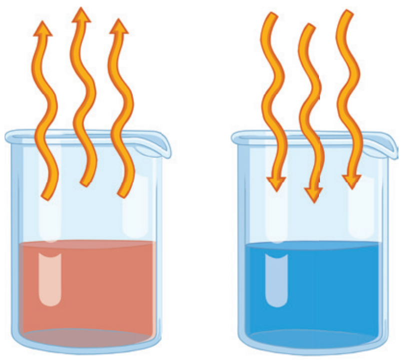
Exothermic and Endothermic Reactions

When a chemical reaction takes place, **energy** is involved. Energy is transferred when chemical **bonds are broken** and when new **bonds are made.**

Exothermic reactions are those which involve the transfer of energy **from the reacting chemicals to** the surroundings. During a practical investigation, an exothermic reaction would show an i**ncrease in temperature** as the reaction takes place.

Examples of exothermic reactions include **combustion, respiration and neutralisation** reactions. Hand-warmers and self-heating cans are examples of everyday exothermic reactions.

Endothermic reactions are those which involve the transfer of energy **from the surroundings to** the reacting chemicals. During a practical investigation, an endothermic reaction would show a **decrease in temperature** as the reaction takes place.



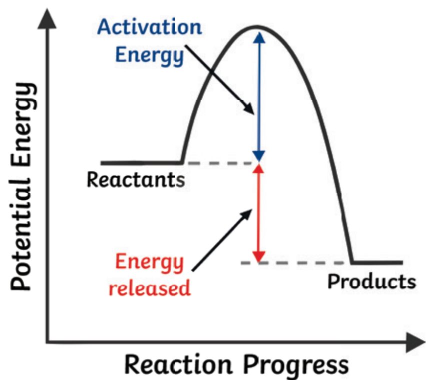
Examples of endothermic reactions include the **thermal decomposition** of calcium carbonate.
Eating **sherbet** is an everyday example of an endothermic reaction. When the sherbet dissolves in the saliva in your mouth, it produces a cooling effect. Another example is **instant ice packs** that are used to treat sporting injuries.

Reaction Profiles – Exothermic

Energy level diagrams show us what is happening in a particular chemical reaction. The diagram shows us the **difference in energy** between the reactants and the products.

In an exothermic reaction, the **reactants** are at a **higher** energy level than the products.

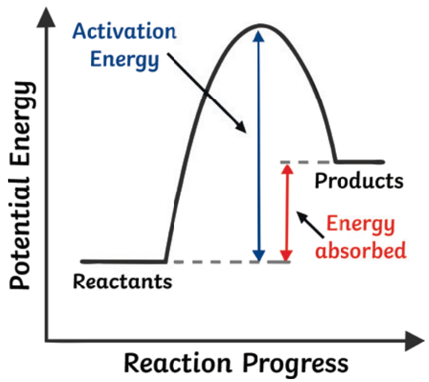
In an **exothermic** reaction, the difference in energy is **released** to the surroundings and so the **temperature** of the surroundings **increases**.



Reaction Profiles – Endothermic

In an **endothermic** reaction, the **reactants** are at a **lower** energy level than the products.

In an **endothermic** reaction, the difference in energy is **absorbed** from the surroundings and so the **temperature** of the surroundings **decreases**.



Activation Energy – the minimum amount of energy required for a chemical reaction to take place.

Catalysts – increase the rate of a reaction. Catalysts provide an alternative pathway for a chemical reaction to take place by **lowering** the activation energy.

Bond Making and Bond Breaking

In an **endothermic** reaction, energy is needed to break chemical bonds. The **energy change (ΔH)** in an endothermic reaction is **positive.**

You may also find, in some textbooks, **ΔH** referred to as the **enthalpy change.**

In an **exothermic** reaction, energy is needed to form chemical bonds. The **energy change (ΔH)** in an exothermic reaction is **negative.**

Bond energies are measured in **kJ/mol.**

Calculations Using Bond Energies (Higher Tier Only)

Bond energies are used to calculate the change in energy of a chemical reaction.

Calculate the change in energy for the reaction: $2H_2O_2 \longrightarrow 2H_2O + O_2$

The first step is to write the symbol equation for the reaction. Once you have done this, work out the bonds that are breaking and the ones that are being made.

2H–O–O–H ⟶ 2H–O–H + O=O

Bond	Bond Energy kJ/mol
H-O	464
O-O	146
O=O	498

On the **left-hand side** of the equation, the **bonds are breaking.**

There are two **O–H** bonds and one **O–O** bond.

So 464 + 146 + 464 = 1074

There are two moles of H_2O_2 therefore the answer needs to be multiplied by two.

So 1074 × 2 = 2148

On the **right-hand** side of the equation, the **bonds are made.**

There are two **H–O** bonds

So 464 + 464 = 928

Two moles of H_2O are made therefore the answer needs to be multiplied by two.

So 928 × 2 = 1856

There is also one **O=O** bond with a bond energy of 498

So 1856 + 498 = 2354

ΔH = sum (bonds broken) – sum (bonds made)

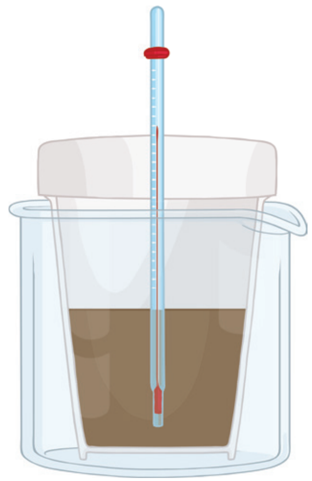
ΔH = 2148 – 2354 = -206 kJ/mol

The reaction is exothermic as ΔH is negative.

Required Practical

Aim

To investigate the variables that affect temperature changes in reacting solutions, e.g. acid plus metals, acid plus carbonates, neutralisations and displacement of metals.



Equipment

- polystyrene cup
- measuring cylinder
- thermometer
- $250cm^3$ glass beaker
- measuring cylinder
- top pan balance

Method

Reaction between a metal and an acid.

1. Gather the equipment.
2. Place the polystyrene cup inside the beaker. This will prevent the cup from falling over.
3. Using a measuring cylinder, measure out $30cm^3$ of the acid. Different acids such as hydrochloric or sulfuric acid may be used. Pour this into the polystyrene cup.
4. Record the temperature of the acid using a thermometer.
5. Using a top pan balance, measure out an appropriate amount of the solid (for example, 10g) or use one strip of a metal such as magnesium.
6. Add the solid to the acid and record the temperature. You may choose to record the temperature of the acid and metal every minute for 10 minutes.

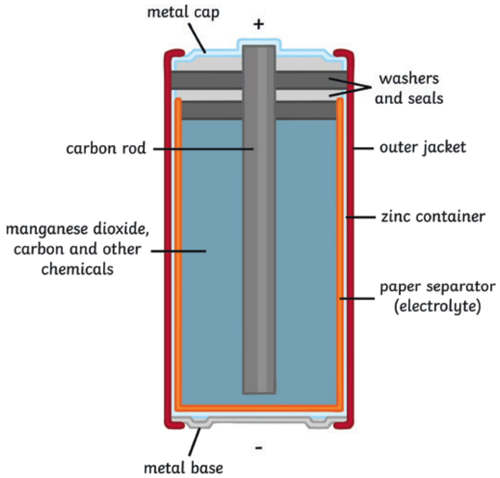
Chemical Cells

A chemical cell converts **chemical energy** into **electrical energy.** More than one cell connected in series is called a battery.

There are two types of chemical cell, **rechargeable** and **non-rechargeable.**

Non-rechargeable cells will produce a **voltage** until the chemicals inside are used up. Once this occurs, the cell is no longer useful and can then be recycled.

Rechargeable cells and batteries can be recharged multiple times. An electrical current is passed backwards through the cell. This works by **reversing** the chemical reactions and the cell or battery can then be used again to produce more electricity. Mobile phones contain rechargeable batteries.

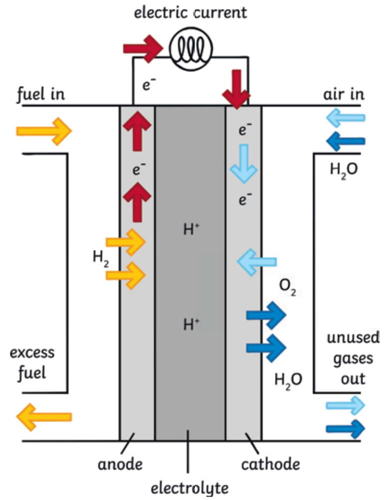


Fuel Cells

Fuels cells work differently to chemical cells in that they need to be supplied with a fuel and oxygen. The constant supply of these two ingredients will allow a fuel cell to produce a voltage continuously.

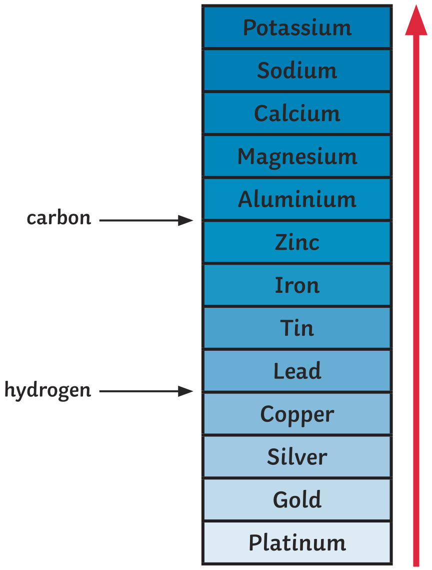
Inside the fuel cell, hydrogen is **oxidised** electrochemically; the fuel is **not combusted.** This allows the reaction to take place at a lower temperature.

Hydrogen-oxygen fuel cells are an alternative to rechargeable batteries and cells as the only product that is produced is water.



Voltage

The voltage of a cell is affected by the combination of metals used inside it. The bigger the difference in the **reactivity** of the two metals, the bigger the **voltage** produced. For example, if the metals used inside the cell are magnesium and zinc, then the voltage produced will be **small** as the two metals are **close together** in the **reactivity series.** By comparison, if magnesium and copper are used, then the voltage produced will be **larger** as the metals are **further apart** in the **reactivity series.**



Ionic Equations

hydrogen + oxygen ⟶ water

$2H_2 + O_2 \longrightarrow 2H_2O$

At the **cathode:** $2H_2 + 4OH^- \longrightarrow 4H_2O + 4e-$

At the **anode:** $O_2 + 2H_2O + 4e- \longrightarrow 4OH^-$

In the fuel cell, **oxygen** is being **reduced** (reduction is the gaining of electrons) whilst **hydrogen** is being **oxidised** (oxidation is the loss of electrons). Oxidation and reduction happen simultaneously – this is called a **redox reaction.**

AQA GCSE Chemistry (Separate Science) Unit 6: The Rate and Extent of Chemical Change

Calculating Rates of Reactions

Reactions happen at **varying rates**. For example, a firework exploding is a fast reaction whereas a piece of iron rusting would take place over a longer period of time.

The **rate of a chemical reaction** tells us how quickly a **product is formed** or how quickly a **reactant is used up**.

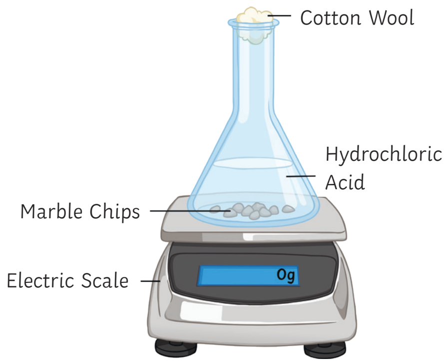
For a chemical reaction to occur, the reactant particles must collide with enough energy. Those collisions that produce a chemical reaction are called successful collisions.

$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time taken}}$$

$$\text{mean rate of reaction} = \frac{\text{quantity of product formed}}{\text{time taken}}$$

Measuring the Mass of a Reaction Mixture

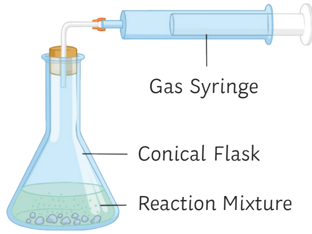
The changing mass of a reaction mixture can be measured during a reaction. This method is particularly useful when gases, such as carbon dioxide, are given off. **Gas escapes during the reaction and the mass of the reaction mixture decreases.** The mass can be measured at regular time intervals.

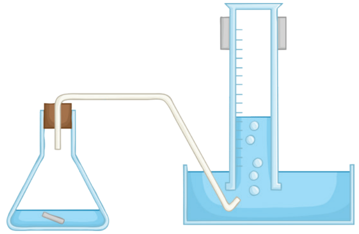


units = g/s or g/min

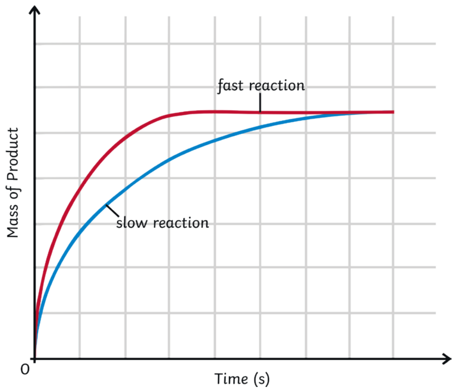
Measuring the Volume of a Reaction Mixture

The changing volume of a reaction mixture can be measured during a reaction. This method is particularly useful when gases, such as carbon dioxide, are given off. The gas can be collected and its volume measured at regular time intervals. Different types of measuring equipment can be used to collect the gas such as a gas syringe, measuring cylinder or upside-down burette.





units = cm^3/s or cm^3/min



Graphs are a useful way to **analyse** the results from a rate of reaction investigation. The graph above shows two lines, one red and one blue.

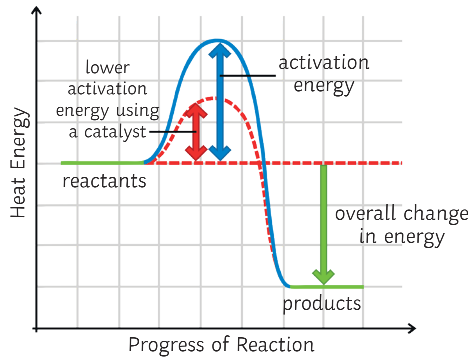
The red line represents a fast reaction and the blue line a slow reaction. We know the fast reaction occurs at a much faster rate as the line is steep. The fast reaction finishes before the slow reaction as the line plateaus sooner.

Factors Affecting the Rate of a Chemical Reaction

- concentration and pressure
- catalyst
- surface area
- temperature

The rate of a chemical reaction will be increased if there are more frequent successful collisions between reactant particles.

Catalyst



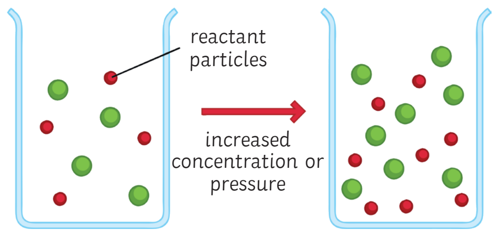
A catalyst is a **substance** that speeds up a chemical reaction without getting used up itself. Catalysts are able to offer an **alternative pathway** at a **lower activation energy**.

Biological catalysts are called **enzymes**.

When a catalyst is used in a chemical reaction (not all reactions have a catalyst that is suitable to use), the **frequency of collisions** is **unchanged**. More **particles** are able to react. The particles have **energy greater** than that of the **activation energy**. Consequently, there is in an **increase** in the **rate successful of collisions**.

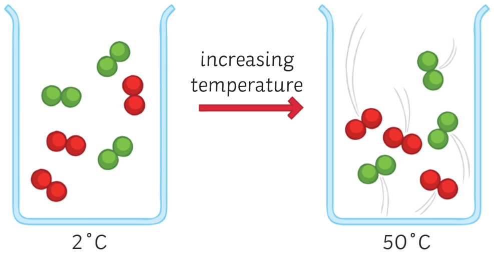
Concentration and Pressure

If the **number of reactant particles** in a given space is **doubled**, there will be **more frequent successful collisions** between reactant particles, therefore, **increasing** the **rate of reaction**.



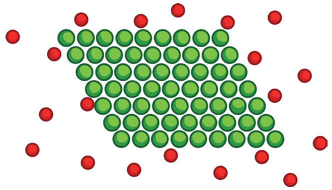
Temperature

When the temperature of the reaction mixture is increased, the reactant particles **gain kinetic energy** and move much more quickly. This results in **more frequent successful collisions** between the reactant particles, therefore, **increasing the rate of the reaction**.



Surface Area

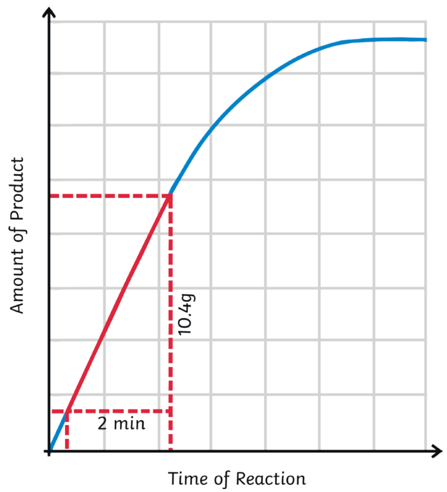
Large lumps of a solid have a **small surface area to volume ratio**. If the solid is broken up into smaller lumps or crushed into a powder, this will increase the surface area to volume ratio.



A larger area of the solid is now exposed to other reactant particles. This increases the frequency of successful collisions thus increasing the rate of reaction.

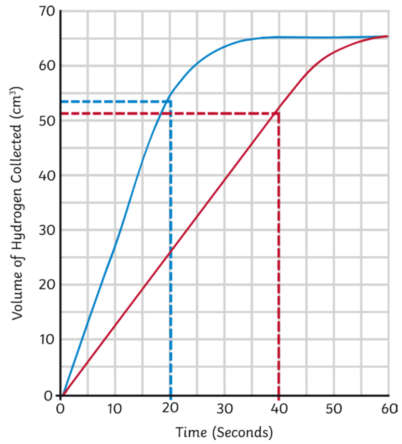
Calculating Gradient (Higher Tier Only) gradient = $\frac{y}{x}$

On the graph, draw construction lines on the part of the graph that has a straight line. Measure the values of x and y.



In the graph below, the gradient of the first line is much steeper than the second line. This indicates that a faster reaction is taking place. Remember, the steeper the line, the faster the reaction.

To calculate the reaction rate at a specific time period, construction lines must first be drawn on the straightest part of the graph.



For the first line, what is the rate of reaction at 20 seconds?

54 ÷ 20 = 2.7cm^3/s

For the second line, what is the rate of reaction at 40 seconds?

52 ÷ 40 = 1.3cm^3/s

Dynamic Equilibrium

In a **closed system** (this means nothing can get in or out), a reversible reaction can reach **dynamic equilibrium**. This is where the **forward** and **reverse reactions** are occurring at the **same rate** and the **concentrations** of all the substances that are reacting remain constant.

Changing Conditions and the Effect on the Position of Equilibrium (Higher Tier Only)

The reaction between nitrogen and hydrogen to make ammonia is an industrial process called the Haber process. It requires a high temperature, high pressure and an iron catalyst.

The symbol equation for the reaction is as follows:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

According to **Le Chatelier's Principle**, the position of equilibrium can be altered by changing the conditions of the reaction i.e. the pressure, concentration and/or the temperature. The **position** of the **equilibrium** will shift to **counteract** any changes made.

Increasing the **temperature** of the reaction in the forward direction (exothermic) will result in the equilibrium shifting in favour of the reverse direction (endothermic) to reduce the temperature.

From the equation, it is clear that on the **left-hand side,** there are **four molecules** and on the **right-hand side,** there are **two molecules**. If the **pressure** in the system were **increased**, the equilibrium **position would shift to the right** as there are fewer molecules. If the pressure in the system were **decreased**, the equilibrium **position would shift to the left** as there are a larger number of molecules.

If the **concentration** of one of the **reactants were increased**, then the equilibrium position would move in **favour of the products**. This would result in more product being produced. If the concentration of the **products was decreased**, equilibrium would shift to **favour the products**. More reactants would react until equilibrium is reached.

Reversible Reactions

A reversible reaction is one in which the **reactants form products**. The products are then able to react together to **reform the reactants**.

For example:
A reacts with B to form C and D.
C and D are able to react to form A and B.

The equation would be as follows (where the **double arrow symbol** represents a **reversible reaction** is taking place):

$A + B \rightleftharpoons C + D$

The **forward reaction** goes to the **left** and the **backwards reaction** goes to the **right**. For example, if the forward reaction is exothermic then the backward reaction will be endothermic. The amount of energy that is transferred is the same for both the forward and reverse reaction.

Hydrated copper sulfate is a blue substance. We say that the copper sulfate is hydrated as it **contains water**. The copper sulfate is heated and the water evaporates leaving a white substance known as **anhydrous** copper sulfate. Anhydrous meaning **no water**.

The word equation for the reaction is as follows:

hydrated copper sulfate ⇌ anhydrous copper sulfate + water

$CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4(s) + H_2O(l)$

The reaction can be reversed when water is added to the anhydrous copper sulfate.

Required Practical 5: Measuring the Production of a Gas

This method outlines one way to carry out an investigation to collect a gas from a chemical reaction.

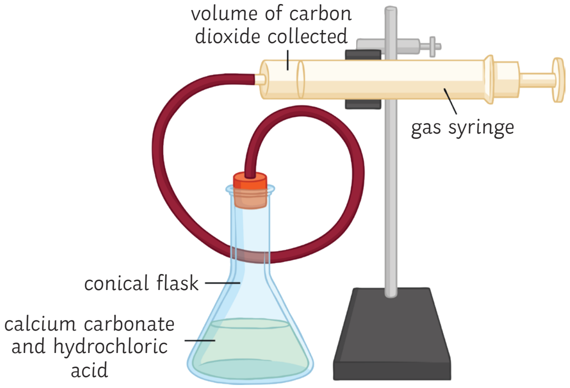
The practical involves changing the concentration of hydrochloric acid and measuring the volume of carbon dioxide gas produced when the acid reacts with calcium carbonate.

The word equation for the reaction is as follows:

calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

The symbol equation for the reaction is:

$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$



Method

Step 1 – Clamp a gas syringe to a retort stand using a boss and clamp. Ensure the syringe is a quarter of the way from the top of the stand. Place the delivery tube to the end of the gas syringe.

Step 2 – Measure out 50ml of hydrochloric acid using a measuring cylinder and pour into a conical flask.

Step 3 – Using a top pan balance, measure out 0.5g of powdered calcium carbonate and place in the conical flask.

Step 4 – Immediately connect the bung and delivery tube to the conical flask. Start the stopwatch.

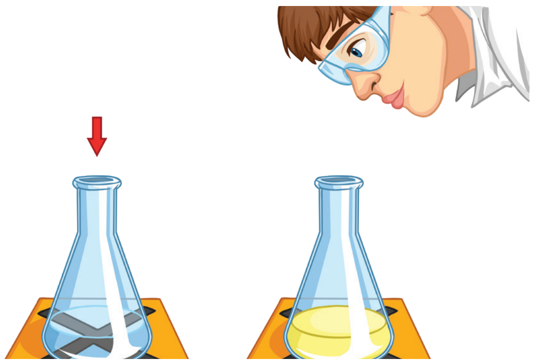
Step 5 – Record the volume of carbon dioxide gas produced every 10 seconds.

Step 6 – When the reaction has finished and there are no more bubbles of gas being produced, clean the equipment and repeat using four other different concentrations of hydrochloric acid.

When analysing the results from the practical investigation, plot a graph of Time (s) against Volume of Gas Produced (cm^3). Draw a curve of best fit through the points. A graph should be plotted for each concentration of acid.

Calculate the mean rate of reaction (cm^3/s) for each concentration of acid used. This can be calculated by dividing the total mass of gas produced (cm^3) by the reaction time (s).

Required Practical 5: Investigating a Change in Colour



This method outlines one way to carry out an investigation into the effect of increased temperature on the rate of a reaction.

The word equation for this reaction is as follows:

sodium thiosulfate + hydrochloric acid → sodium chloride + water + sulfur dioxide + sulfur

The symbol equation for this reaction is:

$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2 + S$

The reaction between sodium thiosulfate and hydrochloric acid produces a **precipitate**. **Sulfur** is responsible for the formation of the precipitate. A precipitate is a **solid** that is formed in a solution. It is the formation of this precipitate that causes the reaction mixture to become **cloudy**; the cloudiness is a way to measure the **reaction time**.

Method

Sodium thiosulfate from three different temperatures may be used, for example, ice cold, room temperature and hot.

Step 1 – Place a black cross on a white tile.

Step 2 – Using the first temperature, measure out $35cm^3$ of sodium thiosulfate using a measuring cylinder. Place the liquid in a conical flask and position over the black cross on the white tile.

Step 3 – Measure out $5cm^3$ of water and $10cm^3$ of hydrochloric acid in separate measuring cylinders.

Step 4 – Pour the water and acid into the conical flask.

Step 5 – Pour the measured amount of sodium thiosulfate into the conical flask and immediately start the stopwatch.

Step 6 – Look down through the conical flask to the black cross below. When the black cross is no longer visible, stop the stopwatch and record the results in a table.

Step 7 – Repeat the steps with the remaining temperatures of sodium thiosulfate.

Crude Oil

Hydrocarbons are compounds that are made up of the elements **hydrogen** and **carbon** only.

Crude oil is a **non-renewable resource**, **a fossil fuel**. Crude oil is made up of a mixture of compounds, most of which are long- and short-chain hydrocarbons.

Most of the compounds in crude oil are hydrocarbons called **alkanes**. The alkanes form a **homologous series**. This is a family of hydrocarbons that all share the **same general formula** and have **chemical properties** that are **similar**.

Alkanes are held together by **single bonds**.

The general formula for an alkane is C_nH_{2n+2}.

They differ from the neighbouring alkane with the addition of a CH_2.

Alkanes are **saturated hydrocarbons**. This means that all their bonds are taken up and they cannot bond to any more atoms.

Alkanes have **similar chemical properties** but have **different physical properties** due to differences in chain length. The longer the chain, the higher the boiling point of the hydrocarbon.

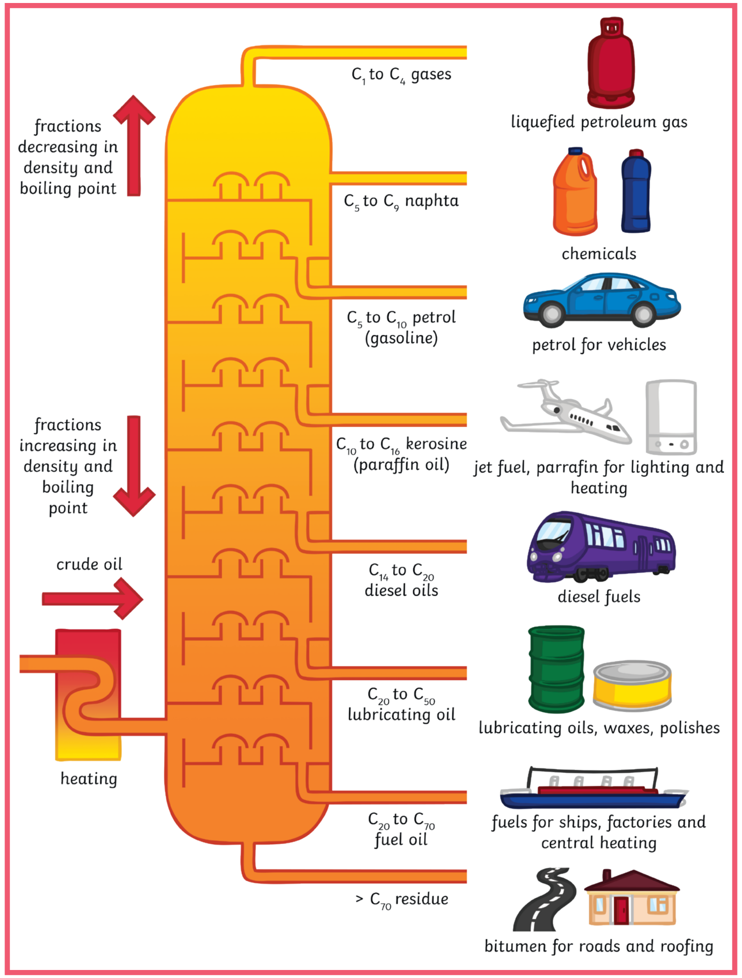
The first four alkanes are: methane, ethane, propane and butane.

A mnemonic to help you remember the order of the alkanes: **m**ice **e**at **p**aper **b**ags.

Fractional Distillation

Fractional distillation is used to **separate** a mixture of long-chain hydrocarbons in crude oil into smaller, more useful fractions.

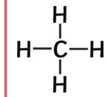
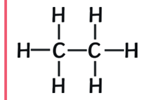
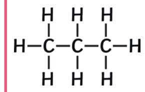
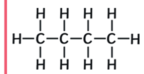
Hydrocarbons have different boiling points depending on their chain length. **Each fraction contains hydrocarbons of a similar chain length**. These fractions will boil at different temperatures due to the difference in sizes of the molecules. The different parts of crude oil are called fractions because they are a small part of the original mixture.



Crude oil is heated and enters at all column called a **fractioning column**. The column is **hot at the bottom** and decreases in temperature toward the top. As the crude oil is heated, it begins to evaporate and its vapours begin to rise up through the column. These vapours condense at the different fractions.

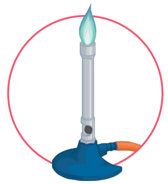
Short-chain hydrocarbons are found at the **top** of the column. This is because shorter chain molecules are held together by **weak intermolecular forces** resulting in low boiling points. These shorter chain hydrocarbons leave the column as gas.

Long-chain hydrocarbons are found at the bottom of the column and are held together by **strong intermolecular forces**, resulting in high boiling points.

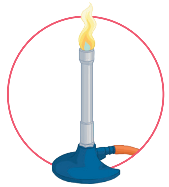
Name of Alkane	Structural Formula	Molecular Formula
methane		CH_4
ethane		C_2H_6
propane		C_3H_8
butane		C_4H_{10}

Combustion

Complete combustion occurs when there is **enough oxygen** for a fuel to burn. A hydrocarbon will react with oxygen to produce carbon dioxide and water.



Incomplete combustion occurs when there **isn't enough oxygen** for a fuel to burn. The products in this reaction are water and poisonous **carbon monoxide**.



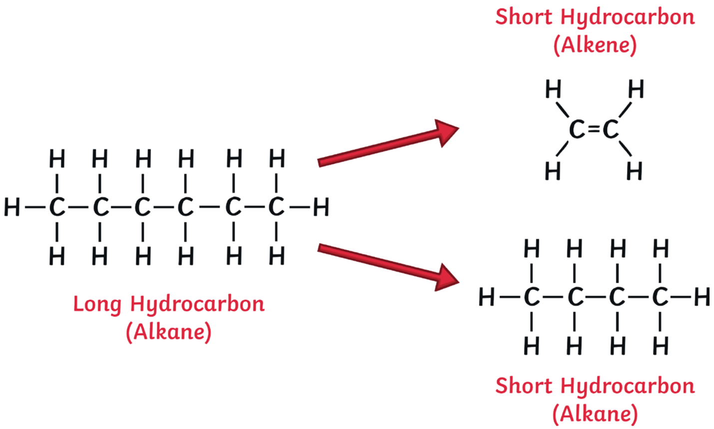
Cracking

Cracking is an example of a **thermal decomposition reaction. Long-chain** hydrocarbons can be **broken** down into **shorter**, more useful hydrocarbon chains.

Cracking can be carried out with a catalyst in **catalytic cracking** or with steam in **steam cracking**.

Catalytic cracking involves heating a hydrocarbon to a high temperature (550˚C) and passing over a hot catalyst.

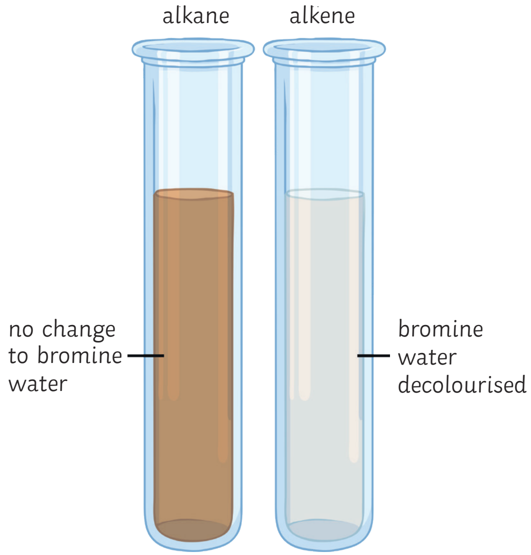
Cracking of a long-chain hydrocarbon **produces** a **short-chain alkane** and an **alkene**.



Test for Alkanes

Bromine, when added to an **alkane**, will **remain brown/ orange**. Alkanes are saturated hydrocarbons, they have no double bonds which could be broken to accept the bromine molecule and so remain orange.

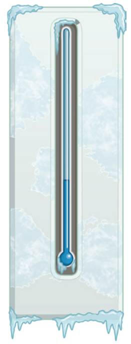
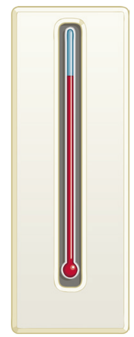
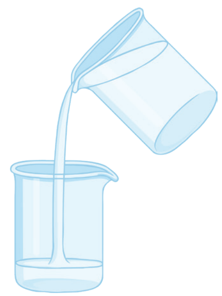
Bromine, when added to an **alkene**, will **change from brown/orange to colourless**. This is because alkenes are unsaturated hydrocarbons. The double bond breaks and the bromine molecule is accepted.



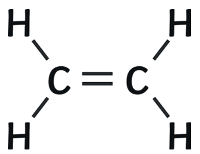
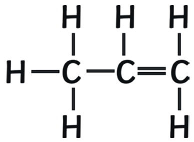
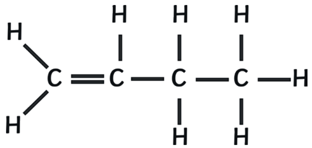
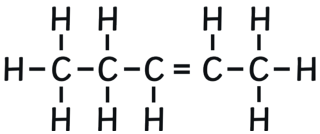
Making Polymers

The fractional distillation of crude oil and cracking produces an array of hydrocarbons that are **key** to our **everyday lives.**

Alkenes are used to produce plastics such as poly(ethene) which is used to make plastic bags, drinks bottles and dustbins. Poly(propene), another polymer, forms very strong, tough plastic.

Short-Chain Molecules	Increasing Chain Length	Long-Chain Molecules
	As chain length increases, the **boiling point** of the hydrocarbon chains also increases.	
thin 	**Viscosity** describes how easily a substance can flow e.g. treacle is very viscous.	thick 
	**Flammability** is a measure of how easily a substance burns.	

Alkenes (Chemistry Only)

Name of Alkene	Structural Formula	Molecular Formula
ethene		C_2H_4
propene		C_3H_6
butene		C_4H_8
pentene		C_5H_{10}

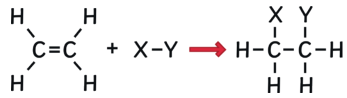
Alkenes are another type of hydrocarbon that is double bonded. The general formula for an alkene is $\mathbf{C_nH_{2n}}$.

Alkenes are **unsaturated hydrocarbons**. In a chemical reaction, the double bond of the alkenes can break. This allows other molecules to bond to it. Note that alkenes all have the suffix '**ene**'.

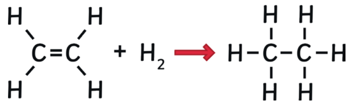
Reactions of Alkenes (Chemistry Only)

Alkenes, just like alkanes, also undergo **combustion** reactions. Alkenes rarely combust completely and tend to undergo **incomplete combustion**. When burning in the air, they produce a smoky flame.

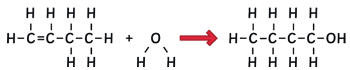
Alkenes have the functional group **C=C**. This double bond between the carbon atoms is able to undergo an addition reaction. This means that the double bond can break and will accept another molecule. **Alkanes** are **unable** to take part in **addition reactions** as their functional group is **C-C**. This means the bond cannot break in order to accept a new molecule.



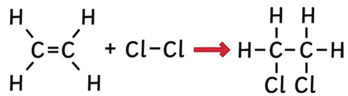
Alkenes are able to **react** with **hydrogen** in an addition reaction called **hydrogenation**. This particular reaction **requires** a **catalyst**.



Alkenes can also **react** with **water** to **produce** an **alcohol**. This is called a **hydration reaction**. The type of compound produced contains a hydroxyl group (-OH), this compound is an alcohol. The reaction **requires** a **high temperature** (300˚C) and a **catalyst**.

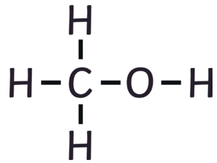
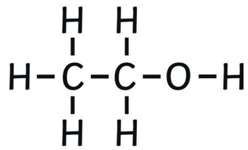
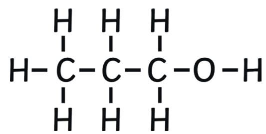
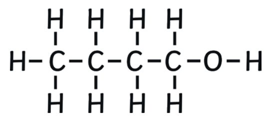


Addition reactions also occur with the **group 7 elements**, the **halogens**. The reaction is called a **halogenation reaction**. When an **alkene reacts** with a **halogen,** an **alkyl halide** is produced.



Alcohols (Chemistry Only)

Alcohols all belong to the **same homologous group**. This is a group of organic compounds that have the same functional group (**-OH, hydroxyl group**) and that have similar chemical properties but different physical properties to each other. Note that alcohols all have the suffix '**ol**'.

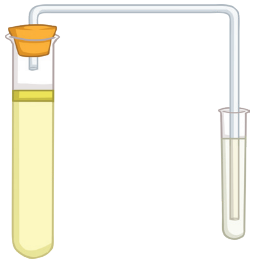
Name of Alcohol	Structural Formula	Molecular Formula	Uses
methanol		CH_3OH	chemical feedstock
ethanol		C_2H_5OH	alcoholic drinks, fuels and solvents
propanol		C_3H_7OH	fuels and solvents
butanol		C_4H_9OH	fuels and solvents

Fermentation

The alcohol that is found in beers, wines and spirits is called ethanol. Ethanol isn't just used in alcoholic drinks, it can also be used as a fuel in vehicles. Ethanol is made through the process of **fermentation**.

Fermentation is an **anaerobic process** and this means that it occurs **without oxygen**.

glucose ⟶ ethanol + carbon dioxide



The fermentation process requires yeast, sugar and water, a warm temperature between 25-35˚C and a reaction vessel that will allow **carbon dioxide** to **escape** but not allow oxygen to get in.

The enzymes needed for fermentation are provided by a single-celled fungus called **yeast**. If the temperature of the reaction mixture is too **cold**, the **fermentation** process will happen very **slowly** or not at all.

If the reaction mixture containing the yeast becomes too **hot**, the **enzymes** may become **denatured** and the process of **fermentation** will **stop**.

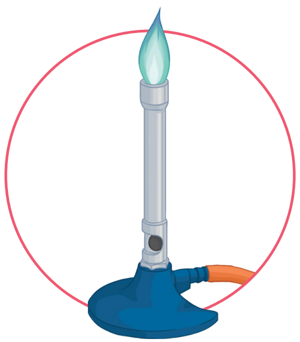
If **oxygen** is allowed to enter the reaction vessel, the **ethanol** will **oxidise** and form ethanoic acid making the drink taste of **vinegar**.

Combustion

Complete combustion occurs when there is **enough oxygen** for a fuel to burn. An alcohol will react with oxygen to produce carbon dioxide and water.

ethanol + oxygen ⟶ carbon dioxide + water

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$



Reactions with Sodium Metal

When dropped into **ethanol**, **sodium** produces **sodium ethoxide** and **hydrogen gas**. Methanol, propanol and butanol all undergo a similar reaction with sodium.

The word equation for this reaction is:

sodium + ethanol ⟶ sodium ethoxide + hydrogen

The symbol equation for this reaction is:

$$2Na + 2C_2H_5OH \longrightarrow 2C_2H_5ONa + H_2$$

Oxidation of Alcohol

Alcohols can be **oxidised** to produce a carboxylic acid. Carboxylic acids are a family of compounds that contain the functional group **–COOH** (carboxyl group). Note that carboxylic acids have the suffix '**oic acid**'. The carboxylic acids have varying physical properties but similar chemical properties.

Oxidation can mean a number of different things: the loss of electrons, the addition of oxygen or the removal of hydrogen. In a chemical equation, the oxidising agent is represented as **[O]**, this symbol means **oxygen from the oxidising agent**.

ethanol + oxidising agent ⟶ ethanoic acid + water

$$CH_3CH_2OH + 2[O] \longrightarrow CH_3COOH + H_2O$$

The equation can also be written in two stages. The first stage shows the formation of **ethanal** and the second stage shows its oxidation.

Stage 1: $CH_3CH_2OH + [O] \longrightarrow CH_3CHO + H_2O$

Stage 2: $CH_3CHO + [O] \longrightarrow CH_3COOH$

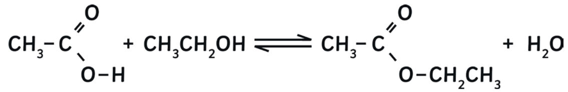
Carboxylic Acids

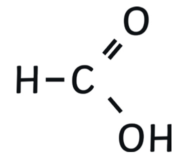
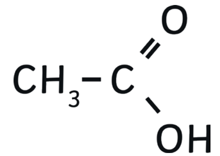
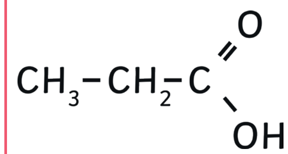
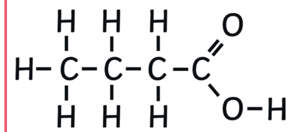
Carboxylic acids are able to react with bases to produce a salt and water. They are also able to react with carbonates to produce a salt, water and carbon dioxide.

When a carboxylic acid is heated with an alcohol in the presence of an acid catalyst (usually concentrated sulfuric acid), an **ester** is formed. Esters typically smell fruity and are used in perfumes. They have the functional group –**COO**-.

For example:

ethanoic acid + ethanol ⇌ ethyl ethanoate + water



Name of Carboxylic Acid	Structural Formula	Molecular Formula
methanoic acid		HCOOH
ethanoic acid		CH_3COOH
propanoic acid		C_2H_5COOH
butanoic acid		C_3H_7COOH

Carboxylic acids are **acidic** due to the hydrogen in the functional group (COOH). When a carboxylic acid forms a salt, the hydrogen is lost and replaced with a metal.

Carboxylic Acids - Higher Tier Only

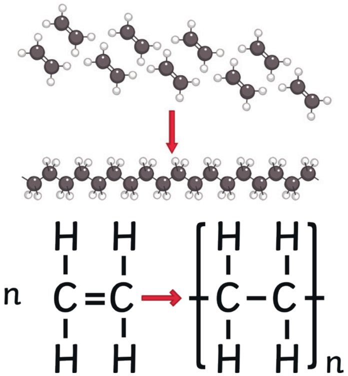
When dissolved in water, carboxylic acids are able to form **acidic solutions**. The pH of the solution is less than 7. They are **weak acids**. Carboxylic acid solutions contain **fewer hydrogen ions** compared with a solution that is the same concentration and contains a strong acid. **Strong acids** are **fully ionised** in solution whereas **weak acids** are only **partially ionised** in solution.

Addition Polymerisation

Addition polymerisation occurs when **two or more monomers** join together to form a **polymer**.
For example, during the polymerisation of ethene, many monomers (single units of ethene) are joined together to make poly(ethene). **Poly** meaning '**many**' (many ethene molecules joined together).

The number of ethene molecules that are joined together could be in the thousands, therefore, when writing the equation the letter 'n' is used to represent the **large number of molecules**.

Notice that the **monomer** of **ethene** has a **double bond**. When it bonds to form **poly(ethene)** the double bond breaks and a **single bond** is formed.



Biological Polymers

DNA (deoxyribonucleic acid) is an example of a **naturally occurring polymer**. DNA is a **double helix** (twisted ladder) and it is made up of two polymer chains that are twisted to form a double helix. The **monomers** of the two polymer chains are called **nucleotides**. The four nucleotides in DNA are called adenine, guanine, cytosine and thyroxine. The nucleotide sequence codes for genes. **Genes** are **sections of DNA** that determine an organisms characteristics.

Proteins are another example of a naturally occurring polymer. Proteins are made from individual **monomer** units called **amino acids**. Proteins have many roles within our bodies; all enzymes are made from proteins.

Plants make the biological polymers **starch** and **cellulose**. They are made up of individual **monomer** units of **sugar** molecules. **Plants** use **starch** as a way to **store energy**. **Cellulose** is used by plants to give the **cell wall strength**.

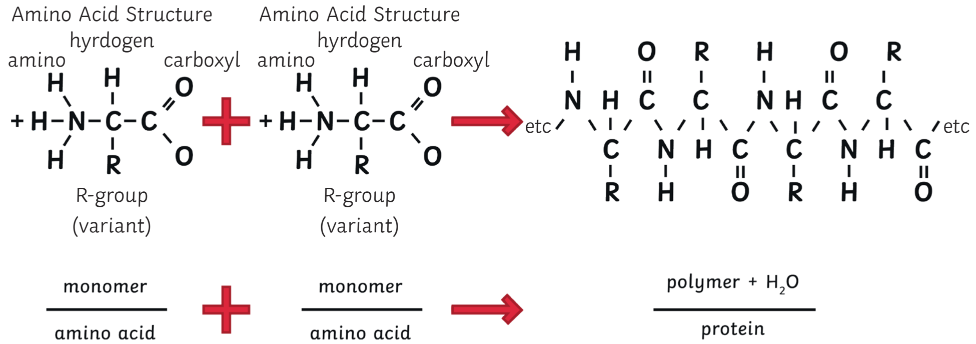
Amino Acids – Higher Tier Only

There are 20 different types of amino acids and when arranged in a particular order, they produce the proteins that are found within our cells.

An amino acid is a molecule that has two functional groups. The amine group (**NH_2**) and the carboxyl group (**COOH**). In between these two functional groups is a single carbon atom with a hydrogen atom bonded to it, along with another group.

Amino acids bond together through the process of a **condensation polymerisation** reaction.

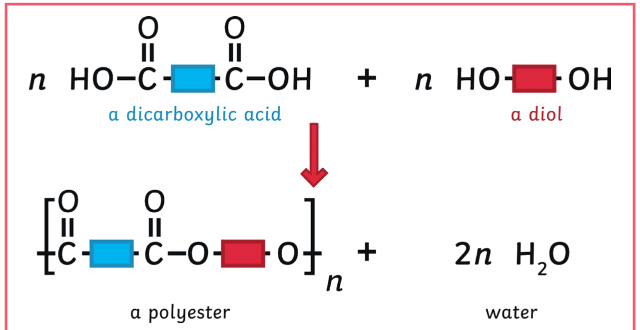
For every **monomer** (amino acid) that is added to the growing chain of the **polymer**, a molecule of **water** is **produced**.



Amino Acids – Higher Tier Only

Addition polymerisation requires the monomers to have a **C=C double bond**. **Condensation polymerisation** does not require a C=C double bond but does need **two functional groups**. When two monomers react, a **water molecule** is usually **produced**.

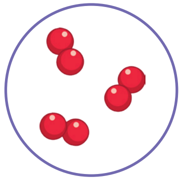
An example of a condensation polymer is polyester. Polyester is made from one **monomer** that has **two hydroxyl groups** and another monomer which has **two carboxylic acid groups**.



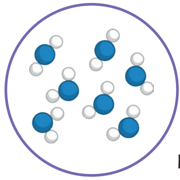
Pure Substances

Pure substances, in chemistry, only contain **one type of element** or **one type of compound**. For example, pure water will just contain water (a compound).

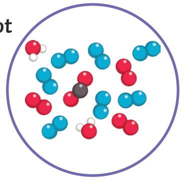
In our everyday language, we use the word 'pure' differently to how it is used in chemistry. Pure can mean a **substance** that has had **nothing else added to it** and is in its natural state. An example of this is pure orange juice. This means that the bottle will just contain orange juice and no other substances.



Elements are made up of **one type of atom**. For example, oxygen is made up of oxygen atoms. Carbon is made up of carbon atoms.



Compounds are **two or more elements** that are **chemically joined** together. For example, NaCl which is sodium chloride.



Mixtures are **two or more elements or compounds** that are **not chemically joined** together. An example of this is a standard cup of coffee. Coffee contains water, milk, coffee and possibly sugar. The components of the cup of coffee are not bonded together.

Pure Substances have a **sharp melting point** compared to **impure** substances which **melt over a range** of temperatures.

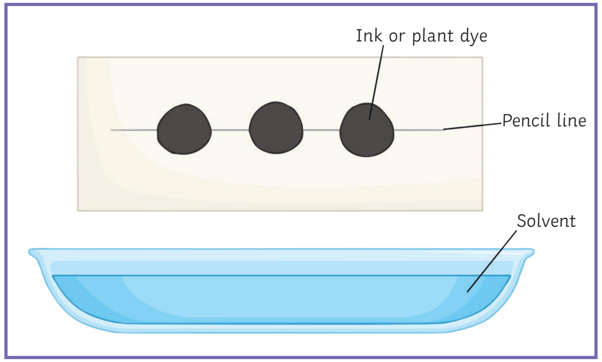
Formulations

Formulations are **mixtures of compounds or substances** that **do not react together**. They do **produce a useful product** with desirable characteristics or properties to suit a particular function.

There are examples of formulations all around us such as medicines, cleaning products, deodorants, hair colouring, cosmetics and sun cream.

Chromatography

Paper chromatography is a separation technique that is used to **separate** mixtures of **soluble substances**. How soluble a substance is determines how far it will travel across the paper.



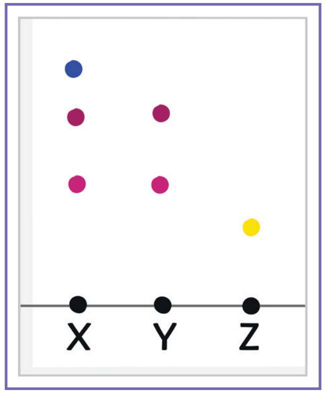
In chromatography, there are **two phases**: the mobile and **stationary** phase.

The **mobile phase moves** through the stationary phase. The **solvent** is the **mobile phase**. It moves through the paper carrying the different substances with it.

The **stationary phase** in paper chromatography is the **absorbent paper**.

Separation of the dissolved substances produces what is called **chromatogram**. In paper chromatography, this can be used to **distinguish** between those substances that are **pure** and those that are **impure**. **Pure substances** have **one spot** on a chromatogram as they are made from a single substance. **Impure substances** produce **two or more spots** as they contain multiple substances.

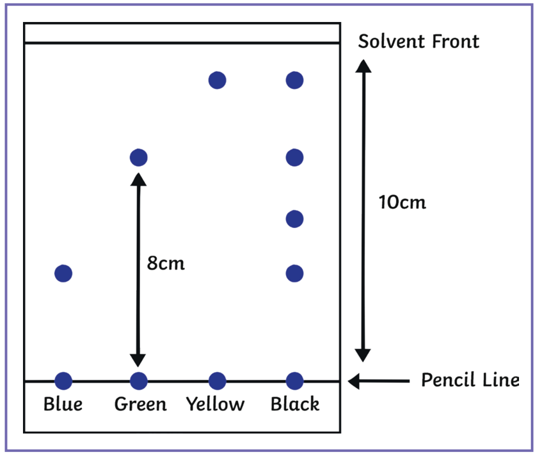
By calculating the **R_f values** for each of the spots, it is possible to identify the unknown substances. Similarly, if an unknown substance produces the **same number** and **colour of spots**, it is possible to match it to a known substance.



R_f Value

$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

Different compounds have different R_f values in different solvents. The R_f values of known compounds can be used to help identify unknown compounds.



Required Practical – Paper Chromatography

Investigate how paper chromatography can be used to separate and distinguish between coloured substances.

Step 1 – Using a ruler, measure 1cm from the bottom of the chromatography paper and mark with a small dot using a pencil. Rule a line across the bottom of the chromatography paper with a pencil, going through the dot you have just made.

Step 2 – Using a pipette, drop small spots of each of the inks onto the pencil line. Leave a sufficient gap between each ink spot so that they do not merge.

Step 3 – Get a container and pour a suitable solvent into the bottom. The solvent should just touch the chromatography paper. The solvent line must not go over the ink spots as this will cause the inks to run into each other.

Step 4 – Place the chromatography paper into the container and allow the solvent to move up through the paper.

Step 5 – Just before the solvent line reaches the top of the paper, remove the chromatogram from the container and allow to dry.

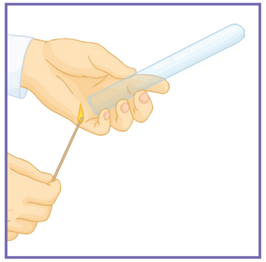
Step 6 – Once the chromatogram has dried, measure the distance travelled by the solvent.

Step 7 – Measure the distance travelled by each ink spot.

Step 8 – Calculate the R_f value. Compare the R_f value for each of the spots of ink.

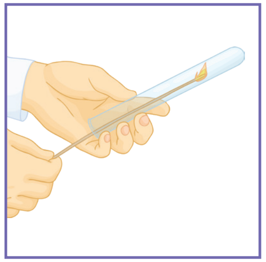
$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

Identification of the Common Gases



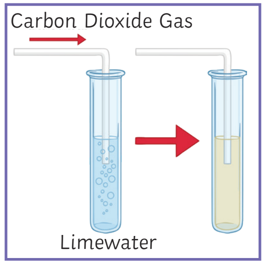
The Test for Hydrogen

Place a burning splint at the opening of a test tube.
If hydrogen gas is present, it will burn rapidly with a **squeaky-pop sound**.



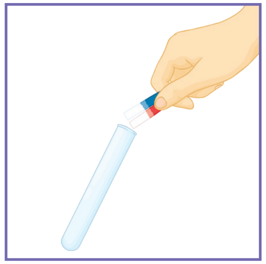
The Test for Oxygen

Place a glowing splint inside a test tube.
The **splint will relight** in the presence of oxygen.



The Test for Carbon Dioxide

Calcium hydroxide (lime water) is used to test for the presence of carbon dioxide.
When carbon dioxide is bubbled through or shaken with limewater, the limewater turns **cloudy**.



The Test for Chlorine

Damp litmus paper is used to test for chlorine gas.
The litmus paper becomes **bleached and turns white**.

Flame Tests

Metal ions when heated produce a variety of flame colours. Flame tests are used to **identify** the **metal ion** that is present; each metal ion produces a different coloured flame.

Step 1 – Dip a wire loop into a sample of the solid compound being tested.

Step 2 – Place the loop into the flame of the Bunsen burner. Ensure that the Bunsen burner is set to a roaring blue flame.

Step 3 – Observe the colour of the flame produced and record it in a table.

Mixtures of ions may cause some flame colours to not be as clear.

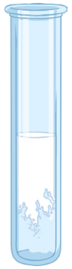
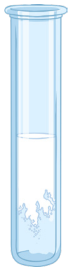
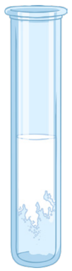
Ion	Colour of the Flame
Li^+	 crimson
Na^+	 yellow
K^+	 lilac
Ca^{2+}	 orange-red
Cu^{2+}	 green

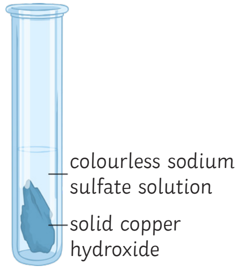
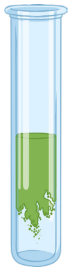
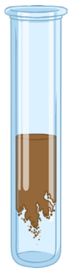
Metal Hydroxides

In order to identify metal ions, **sodium hydroxide solution** is added. Solutions of calcium, magnesium and aluminium all form white precipitates. Only the aluminium hydroxide **precipitate** dissolves in excess sodium hydroxide. Iron (II), iron (III) and copper (II) all form coloured precipitates when sodium hydroxide solution is added.

magnesium sulfate + sodium hydroxide ⟶ magnesium hydroxide + sodium sulfate

$MgSO_4 + 2NaOH \longrightarrow Mg(OH)_2 + Na_2SO_4$

Ion	Colour of the Precipitate Produced
Al^{3+}	white 
Ca^{2+}	white 
Mg^{2+}	white 

Ion	Colour of the Precipitate Produced
Cu^{2+}	blue 
Fe^{2+}	green 
Fe^{3+}	brown 

Ionic Equations

An ionic equation can be used to represent each of the **precipitation** reactions. These equations only show the ions that are involved in the precipitation reaction. The equations do not show the sodium or sulfate ions. This is because these ions are called spectator ions. **Spectator ions** are ions that do not take part in the chemical reaction.

Copper (II)

$Cu^{2+} + 2OH^- \longrightarrow Cu(OH)_2$

Copper has **lost** two **negative charges**, hence why copper is Cu^{2+}. In order to balance out this loss of charges, the copper ion **must gain** two negative charges. These negative charges come in the form of **two OH^- ions**.

Iron (III)

$Fe^{3+} + 3OH^- \quad Fe(OH)_3$

Iron (III) has **lost** three **negative charges**, hence why iron is Fe^{3+}. In order to balance out the loss of charges, the iron ion **must gain** three negative charges. These negative charges come in the form of **three OH^- ions**.

Testing for Carbonate Ions (CO_3^{2-}) Chemistry Only

Place a small volume of limewater into a test tube. In a separate test tube, add a small sample of the **carbonate** and add a few drops of **hydrochloric acid** (acids are a source of H^+ ions) using a pipette. Seal the test tube with a bung connected to a delivery tube; the delivery tube should be placed in the test tube containing the limewater. Bubbles of **carbon dioxide** gas will be produced. The **limewater will turn a milky colour** indicating a positive test for carbon dioxide.

$CO_3^{2-} + 2H^+ \longrightarrow CO_2 + H_2O$

Testing for Sulfate Ions (SO_4^{2-})

Using a pipette, add a few drops of **barium chloride** solution to the sample followed by a few drops of **hydrochloric acid**. A positive result for sulfate ions will produce a white precipitate.

$SO_4^{2-} + Ba^{2+} \longrightarrow BaSO_4$

Testing for Halide Ions (I^-, Br^-, Cl^-)

Using a pipette, add a few drops of dilute **nitric acid** to the sample followed by a few drops of **silver nitrate solution**. Leave it to stand and **observe the colour** of the **precipitate formed**.

Each halide ion produces a different coloured precipitate.

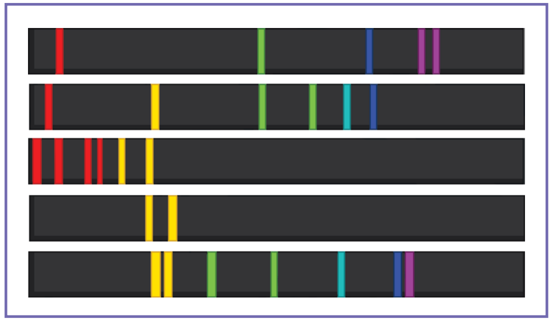
- **Chloride** produces a **white** precipitate.
- **Bromide** ions produce a **cream** precipitate.
- **Iodide** ions produce a **yellow** precipitate.

Flame Emission Spectroscopy

Flame emission spectroscopy is an instrumental method of analysis. The benefits of instrumental methods of analysis are that it is **rapid, accurate and sensitive**. The drawbacks to such methods are that the equipment is often **expensive** and **requires special training** to use.

Flame emission spectroscopy is a technique that is used to **identify** metal ions in solution. The samples that are tested normally include biological fluids and tissues.

How It Works



Step 1 – A sample is heated in a flame.

Step 2 – Electrons in the metal ions are excited by the thermal energy provided from the flame. As a result, the electrons move into a higher energy level.

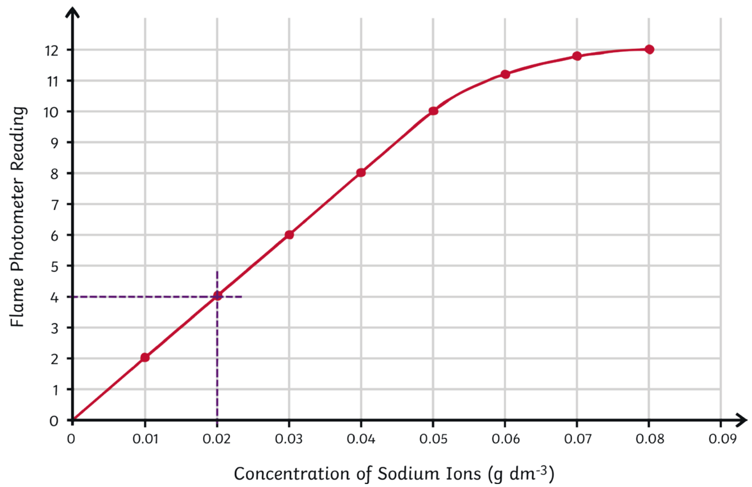
Step 3 – When the electrons fall back into a lower energy level, they release energy in the form of light.

Step 4 – The emitted wavelengths of light are analysed instrumentally.

Step 5 – To identify the metal present, its spectrum is compared with reference spectra from known metal ions.

Above is an example of the spectra produced by flame emission spectroscopy. It looks like a colourful array of lines. **Each metal ion produces a unique emission spectrum**.

Calibration Curve



The readings for different concentrations of metal ions in solutions are taken. These readings are then used to plot a calibration curve.

AQA GCSE Chemistry (Separate Science) Unit 9: Chemistry of the Atmosphere

The Early Atmosphere

Approximately **4.6 billion years ago** the Earth was formed. Scientists have lots of ideas and **theories** about how the atmosphere was produced and the gases within it, but due to the lack of evidence, they cannot be sure.

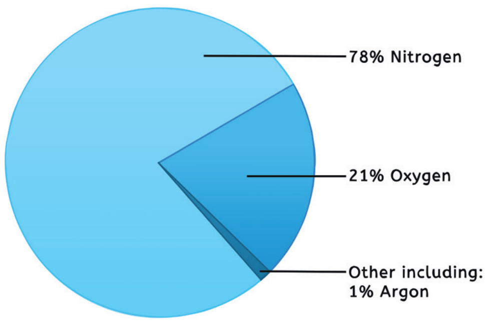
One theory suggested that **intense volcanic activity** released gases that made **Earth's early atmosphere** very similar to that of Mars and Venus. These planet's atmospheres mainly consist of carbon dioxide with little oxygen.

Nitrogen gas would have also been released from volcanoes and would have built up in the atmosphere.

Water vapour in Earth's early atmosphere would have **condensed** to create the **seas and oceans**. Carbon dioxide would have dissolved into the water, decreasing the level in the atmosphere.

Percentage of Gases in the Atmosphere

The pie chart below shows the abundance of each gas in our atmosphere.



How Did the Levels of Oxygen Increase?

2.7 billion years ago, algae first produced oxygen. Gradually over time, the levels of oxygen in our atmosphere increased as plants evolved. This was followed by animals as the levels of oxygen increased to a level that would sustain more complex life.

Oxygen is produced by plants in the process of **photosynthesis**.

$$6CO_2 + 6H_2O \xrightarrow[\text{Chlorophyll}]{\text{Light}} C_6H_{12}O_6 + 6O_2$$

Carbon dioxide + Water → Glucose + Oxygen

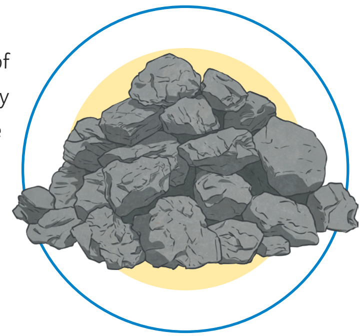
How Did the Levels of Carbon Dioxide Decrease?

Carbon dioxide **dissolves** in water. As water vapour condensed and the oceans and seas formed, the carbon dioxide gas dissolved producing **carbonate compounds**. This process reduced the amount of carbon dioxide in the atmosphere. Carbonate compounds were then **precipitated**: limestone is an example of a sedimentary rock; it has the chemical name calcium carbonate.

Plants in the oceans absorbed **carbon dioxide** gas for **photosynthesis**. The organisms from the food chains that the plants supported were turned into fossil fuels. **Fossil fuels** are **non-renewable** and consist of **coal, crude oil, and gas**, all of which contain carbon.

Crude oil was formed millions of years ago. When aquatic plants and animals died, they fell to the bottom of the sea and got trapped under layers of sand and mud. Over time, the organisms got buried deeper below the surface. The **heat and pressure** rose, turning the remains of the organisms into crude oil or natural gas. Oxidation did not occur due to the lack of oxygen.

Coal is a fossil fuel formed from **giant plants** that lived hundreds of millions of years ago in swamp-like forests. When these plants died, they sank to the bottom of the swamp where dirt and water began to pile on top of them. Over time, pressure and heat increased and the plant remains underwent chemical and physical changes. The oxygen was pushed out and all that remained was coal.



The Human Impact and the Greenhouse Effect

Scientists believe that human activities have resulted in the **increased** amount of greenhouse gases in the atmosphere. Activities such as **farming cattle** and **farming rice** release huge amounts of **methane** into the atmosphere.

Burning **fossil fuels** in cars and power stations releases large amounts of **carbon dioxide**. With large areas of the rainforest being cut down through **deforestation**, the excess carbon dioxide is not being absorbed by photosynthesis.

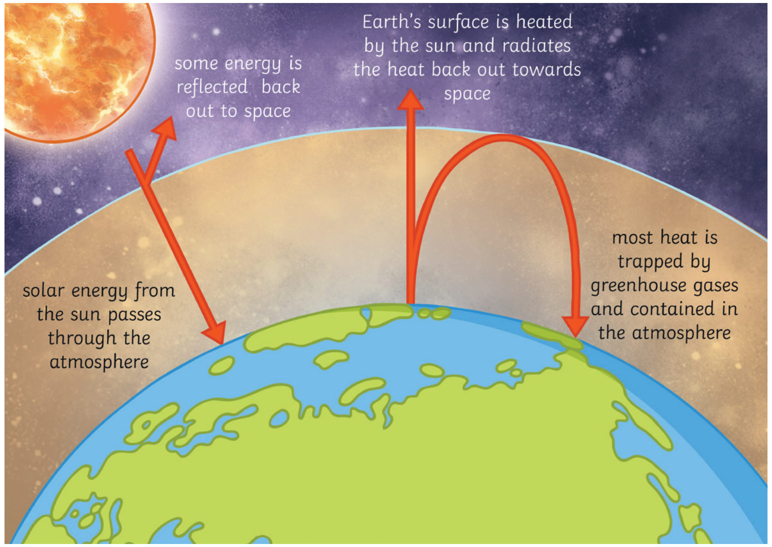
However, not everyone believes that humans are causing the rise in greenhouse gases. Some believe that the rise in global temperatures is associated with cycles of climate change and natural factors.

Climate science is often complicated as there are **difficulties** associated with **predicting future global temperatures**. The media present information that can be biased, inaccurate or lacks substantial evidence.

After reading an article on global warming, consider the trustworthiness of the source by considering these factors:

- Is the research done by an expert in that field and do they have the right skills and qualifications to report on the issue?
- Which organisation is reporting the evidence? If it is a newspaper, some stories are sensationalised in order to sell papers.
- Was the research funded by a legitimate organisation and was it conducted in a non-biased way? Think about the methods that were used to obtain the data and the impact the collection and analysis of this data had on the overall result.

The Greenhouse Effect



A greenhouse is a house made of glass and is commonly used by gardeners to help grow plants and keep them warm. As the sun shines through the greenhouse, the air is heated up and becomes trapped by the glass and is prevented from escaping. During daylight, a greenhouse stays quite warm and this lasts into the night.

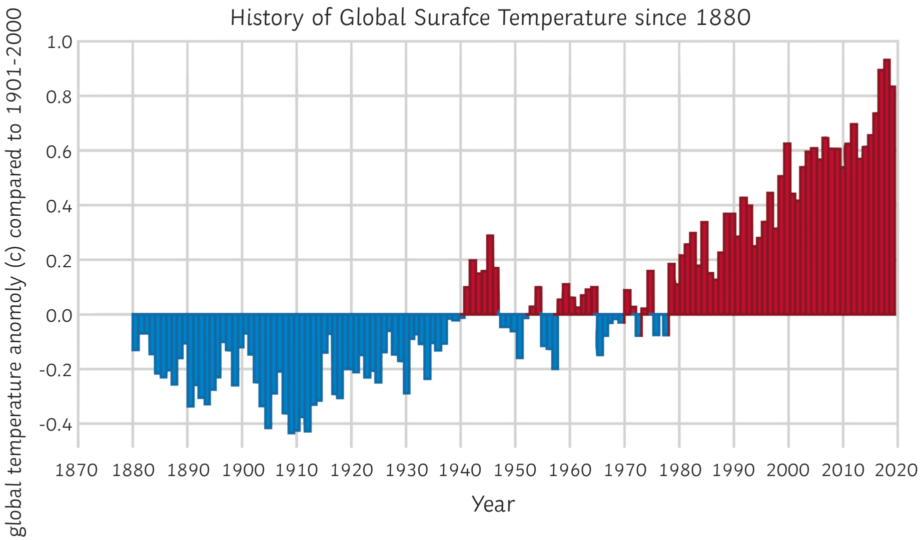
The earth and its atmosphere are very similar to that of a greenhouse. The greenhouse gases in the atmosphere trap the heat and keep the earth warm. The main greenhouse gases are **carbon dioxide**, **water vapour and methane**. During the daylight, the sun warms up the earth's surface. During the night, as the earth begins to cool and release the heat back into the atmosphere, some of the heat is trapped by the greenhouse gases in the atmosphere.

If the **greenhouse effect** becomes too **strong**, the earth will get too warm and melt the Arctic ice. As we burn more fossil fuels, the levels of **carbon dioxide** and the other greenhouse gases **increase** in our atmosphere which makes the greenhouse effect stronger.

What is the Difference Between Climate Change and Global Warming?

Since the Earth was formed over 4.6 billion years ago, its climate has constantly been changing with several ice ages followed by warmer temperatures. Changes in the Sun's energy reaching the Earth and volcanic eruptions were responsible for the changes until about 200 years ago.

Global warming is different to climate change and is used to explain how the earth's climate has warmed up over the past 200 years. Scientists believe that the warming of the climate is due to the activities of humans.

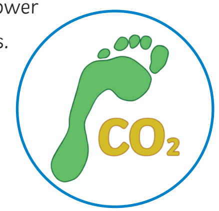


Carbon Footprint

The carbon footprint is the total amount of **carbon dioxide** and other greenhouse gases emitted over the full life cycle of a product, service or event.

An individual's carbon footprint is a calculation of all the activities that that person has taken part in throughout the year.

These activities might involve flying abroad or **travelling** by bus or rail. Each of which might be powered by petrol or diesel. **Heating a home** in winter by using a gas-powered boiler and using electricity to power lights and electronic devices. **Food** also has a **carbon footprint**, for example, beef and rice produces huge amounts of methane when farmed.



Nitrogen

Nitrogen and oxygen react together to make oxides of nitrogen. This occurs inside a **car engine** where there is a high temperature and pressure. Many compounds can be formed when nitrogen reacts with oxygen. The two that are formed inside a car engine are NO and NO_2.

Nitrogen compounds are grouped together with the general formula NO_x. Nitrogen compounds, along with sulfur dioxide, are also responsible for acid rain.

Compounds of nitrogen oxides react in the atmosphere with ultraviolet light from the sun to produce **photochemical smog**. The smog is most noticeable during the morning and afternoon and occurs mainly in densely populated cities.

The presence of smog can have a **major impact on human health**, particularly to those who suffer with **asthma**.

Combustion

Complete combustion occurs when there is **enough oxygen** for a fuel to burn. A hydrocarbon will react with oxygen to produce carbon dioxide and water.

propane + oxygen ⟶ carbon dioxide + water

$C_3H_8 + \mathbf{5}O_2 \longrightarrow \mathbf{3}CO_2 + \mathbf{4}H_2O$

Incomplete combustion occurs when there **isn't enough oxygen** for a fuel to burn. The products in this reaction are water and poisonous **carbon monoxide**. Carbon particles (soot) may also be seen.

ethane + oxygen ⟶ carbon monoxide + water

$\mathbf{2}C_2H_6 + \mathbf{5}O_2 \longrightarrow \mathbf{4}CO + \mathbf{6}H_2O$

Carbon monoxide is a poisonous gas. It is often called the **silent killer** due to it being colourless and odourless. Carbon monoxide works by binding to the **haemoglobin** in your red blood cells. This prevents them from carrying oxygen to the cells around your body. Carbon monoxide detectors are used to detect levels of the gas in the surrounding air and are often placed near gas-powered boilers to detect gas leaks.

Particulate carbon irritates the lining of the lungs making asthma worse and could cause cancer. **Global dimming** is caused by particulates of carbon blocking out the Sun's rays and may reduce rainfall.

Sulfur Dioxide

Sulfur dioxide is an **atmospheric pollutant**. It is a gas that is produced from the burning of **fossil fuels**. Sulfur dioxide is able to dissolve in rainwater and produces **acid rain**. Acid rain causes damage to forests, kills plants and animals that live in aquatic environments, and damages buildings and statues as the acid rain erodes the stone that they are made from.

sulfur + oxygen ⟶ sulfur dioxide

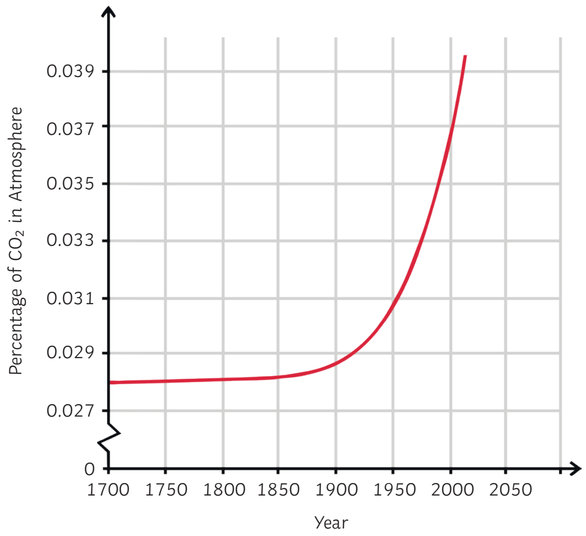
$S + O_2 \longrightarrow SO_2$

Sulfur dioxide can be further oxidised to form sulfur trioxide.

What is the Link Between Carbon Dioxide and Global Warming?

There is a strong correlation between the percentage concentration of carbon dioxide in the atmosphere and increased global temperatures.

The impact of this is that the polar ice caps are melting, sea levels are rising and habitats and rainfall patterns are changing. The impact of which is already being felt around the globe. The consequences of human activity will affect us all.



Sustaining Human Life on Earth

The human **population** is **increasing** rapidly and our use of earth's finite resources has increased. If humans continue to use these resources at the rate at which we are, then we will reach a point where the human population cannot be sustained on earth.

Humans use the **earth's natural resources** for warmth, shelter, food, clothing and transport. Scientists are making **technological advances** in **agricultural** and **industrial processes** to provide food and other products that meet the growing needs of the human population but it is of major importance that this is done in a sustainable way so that our finite resources are not used up.

Earth's Resources

Finite resources are those of which there is a **limited supply**, for example coal, oil and gas. These resources can be used to provide energy but, one day, their supply will run out.

Crude oil is processed through **fractional distillation** and **cracking** to produce many useful materials such as petrol, diesel and kerosene.

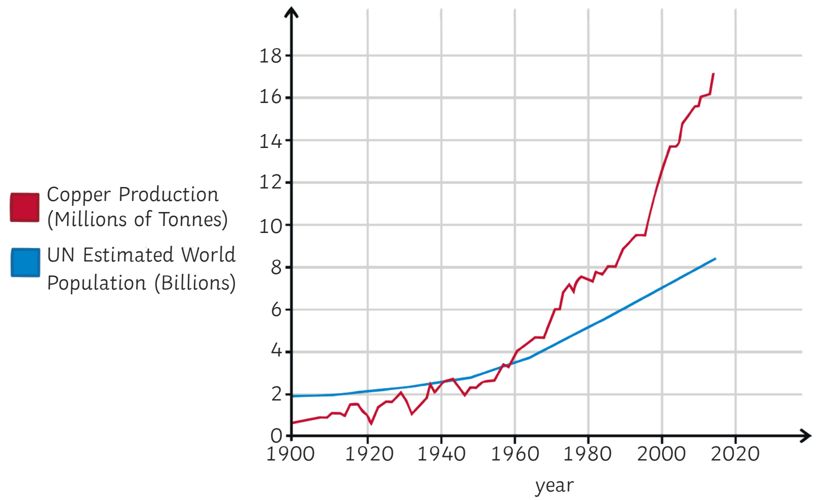
Renewable resources will not run out in the near future because the reserves of these resources are high. Examples of renewable resources include solar energy, wind power, hydropower and geothermal energy.

Haber Process and Copper

Scientists often discover new ways to produce a product; **synthetic methods** of production replace **natural methods**. For example, fertilisers were obtained from manure (a natural resource).

The **Haber process** allowed the synthetic production of **fertilisers** and this enabled **intensive farming** methods to spread across the globe. In turn, this supported the growing human population.

Copper is another resource that has been exploited over time. As the human population has increased since 1900, the demand for copper has also increased. Copper is a finite resource which means that there is a limited supply.



Water

Potable water is water that is **safe to drink**. Potable water is **not pure; dissolved impurities** still **remain** in the water. Pure water is odourless, tasteless and colourless compared to rainfall or water from streams and wells as these **harbour chemicals** such as acid.

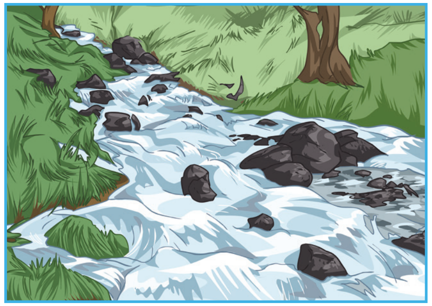
Pure – the **definition** of a pure substance is one that contains only a single type of material that has not been contaminated by another substance.

Potable water must contain **low levels** of microbes and salts for it to be deemed safe to consume. This is because **high levels** of microbes and salts can be harmful to human health.

The methods of making water safe vary depending on where you live. Starting with sea water is harder than starting with fresh water. This is because the **energy cost** of removing large amounts of sodium chloride from seawater is greater.

In the UK, our populations' water needs are met through **rainfall**. During the **summer, water levels** in reservoirs **decrease** and local areas are encouraged to reduce their water usage by swapping baths for showers and they are asked to avoid using hosepipes.

In the UK, **insoluble particles** are **removed** from naturally occurring fresh water by passing it through **filter beds**. **Microbes** are **killed** by **sterilising the water**. Several different sterilising agents are used for potable water. These are chlorine, ozone or ultraviolet light. The right amount of chlorine and ozone gas (O_3) must be used as both are harmful to human health.



Desalination of Sea Water

If fresh water supplies are limited, sea water can undergo a process called **desalination**. This process requires **large amounts of energy**, but can be done by distillation or the use of membranes such as **reverse osmosis**.

Distillation involves **heating** the sea water until it reaches **boiling point**. Once the water is boiling, it will begin to **evaporate**. The steam then rises up where it cools and condenses in a condensing tube. The salt is left behind. The **downside** to this process is the **energy cost** of boiling the water and cooling down the steam sufficiently in the condensing tube. Not all of the water evaporates which leaves behind a **salty wastewater** that can be **difficult to sustainably dispose of** without harming aquatic organisms.

Reverse Osmosis of Salt Water

Osmosis, as you will have learnt in biology, is the **movement of particles** from an area of **high** concentration **to** an area of **low** concentration through a **semi-permeable membrane**.

Reverse osmosis involves **forcing water** through a **membrane** at **high pressure**. Each membrane has tiny holes within it that only allow water molecules to pass through. Ions and other molecules are prevented from passing through the membrane as they are too large to fit through the holes.

The **disadvantage** of this method is that it produces **large amounts of wastewater** and requires the use of **expensive membranes**. Due to a large amount of wastewater produced, the efficiency of this method is very small.

Water Treatment

Before the **wastewater** from industry, agriculture and peoples' homes can be released back into the environment, it must be **treated**.

Pollutants such as human waste contain **high levels of harmful bacteria** and **nitrogen compounds** which can be a **danger to aquatic organisms**.

Industrial and agricultural waste may contain **high levels** of **toxic metal compounds** and **fertilisers and pesticides** which may also damage the ecosystem.

Cleaning sewage requires several steps:

Step 1 – The water must be **screened**. This is where material such as branches, twigs and grit is removed.

Step 2 – The water undergoes **sedimentation**; wastewater is placed in a settlement tank. The heavier solids sink to the bottom and form a sludge whilst the lighter effluent floats on the surface above the sludge.

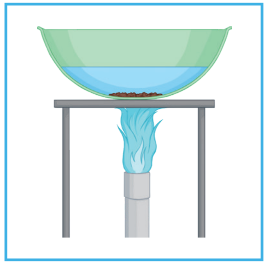
Step 3 - The effluent is then transferred to another tank where the organic matter undergoes **aerobic digestion**. Although not pure, this water can be safely released back into the environment. The sludge is placed in another tank where the organic matter undergoes **anaerobic digestion**. It is broken down to produce fertiliser and methane gas which can be used as an energy resource (fuel).

Required Practical 8 – Analysis and Purification of Water Samples from Different Sources

Analysing the pH of Water Samples

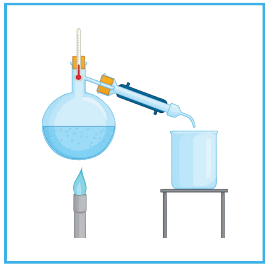
Test the pH of each water sample using a pH meter or universal indicator. If using universal indicator, use a pH colour chart so that you are able to identify the pH of the sample against the colour produced by the indicator.

Analysing the Mass of Dissolved Solids



To measure the mass of dissolved solids in a water sample, measure out 50cm^3 of the sample using a measuring cylinder. Take the mass of an evaporating basin before heating and record the mass in a table. Place the measured amount of water into an evaporating basin and gently heat over a Bunsen burner until all the liquid has evaporated. Once the evaporating basin has cooled, place it on a top pan balance and record its mass. Calculate the mass of the solid left behind.

Distillation of the Water Sample



To distil a water sample, set up your equipment as per the diagram.

Heat the water sample gently using a Bunsen burner. After a short period of time, distilled water should be produced.

Life-Cycle Assessment (LCA)

Life-Cycle Assessments follow the four main stages of the life cycle of a product.

Stage 1 – Extracting the raw materials needed to make the products and then processing them.

At this stage, the energy and environmental costs need to be considered. For example, if the raw material being used is a finite or renewable resource, the energy to extract and transport the raw material should be considered. Environmental factors also play a large part in stage 1 as the extraction of the raw material can leave scars on the landscape and waste products may be produced that could damage local ecosystems.

Life-Cycle Assessment (LCA) (continued)

Stage 2 – Manufacturing and packaging of the product.

The main consideration is how much energy and resources are needed to manufacture the product. Energy may be used in the form of fuel, electricity or chemicals used in the production of the product. In the manufacturing process, there may be pollution and waste products that need to be considered. Transportation of the goods from the factory to the user will have an environmental impact.

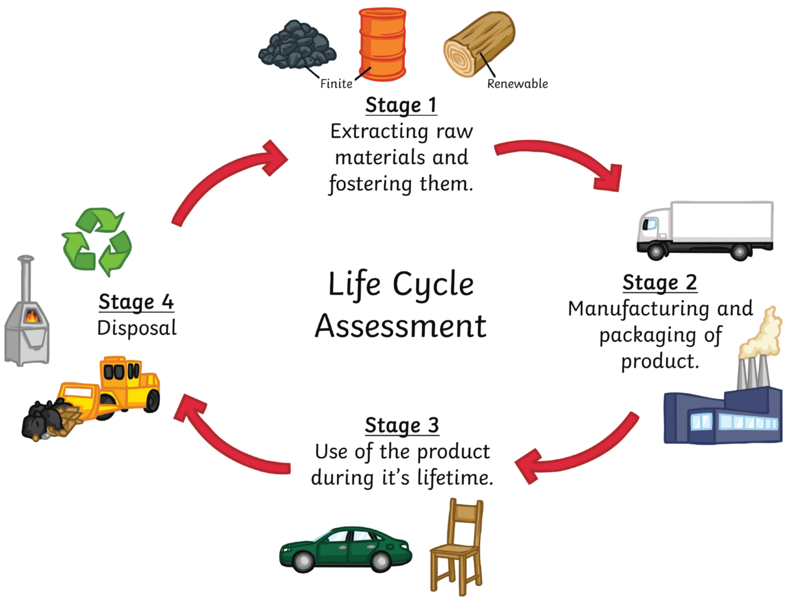
Stage 3 – Use of the product during its lifetime.

The environmental impact of a product during its life depends on the type of product. For example, a car will have a significant impact i.e. it needs to be filled with petrol or diesel, a finite resource, to get to where you are going. A car's engine releases harmful emissions into the atmosphere. On the other hand, a wooden chair may only need minor repairs and is made from a renewable resource.

Stage 4 – Disposal at the end of a product's life.

There are different methods of disposal:

1. Landfill – the product is put in a hole in the ground – high environmental impact.
2. Incineration (organic matter) – burning of the product – low environmental impact.
3. Recycling – for example, batteries contain metal compounds that are not good for the environment. By recycling, it means that no new compounds have to be taken out of the ground.



Comparative LCAs

Comparative LCAs are used to evaluate products and to find which one will have a lower environmental impact.

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

Disadvantages of Comparative LCAs

The disadvantage of **comparative LCAs** is that some parts of it require certain judgements to be made.

Different people have different opinions and this is dependent on who completes the LCA and whether a certain level of bias is added. For example, if the LCA is completed by a company that is manufacturing a specific product, they may only discuss **some** of the environmental impact of their product in the LCA. Accurate numerical values, for example, show a company how much energy has been used in the **manufacturing process** or how much **carbon dioxide** was produced when the goods were transported.

Recycling

Many materials are made from **natural resources** that have **limited supplies**. Reusing items such as glass bottles that only need washing and sterilising saves energy and reduces the environmental impact. Not all products can be reused, some need to be recycled before reuse.

There are both advantages and disadvantages to recycling materials.

Advantages

- Fewer resources such as **mines** and **quarries** are needed to remove raw, finite materials from the ground. For example, copper.
- Crude oil, the raw material used in the production of plastics, does not need to be extracted. This, in turn, **avoids** high energy cost processes such as fractional distillation and cracking. If more items are recycled, less would end up in landfill sites.
- The amount of greenhouse gases would reduce as the energy cost of recycling is a lot **less** than making a new product.

Disadvantages

- Recycling items require collection and transport of the goods to the organisation. This involves using staff, vehicles and the use of fuel.
- Some materials, such as **metals**, can be **difficult to sort**; the amount of sorting is dependent on the purity of the materials or metals and the level of purity required for the final product. For example, copper used in electrical appliances must have a high purity. To achieve this, the copper needs to be processed and then melted down again to make copper wiring.
- Steel that is used in the construction industry does not require such high purity. Often scrap iron is added to the furnace when steel is made. This reduces the need for as much iron ore and reduces the cost of making steel.

Biological Extraction Methods (Higher Tier Only)

Biological methods of extraction are needed as the resources of **metal ores** on earth are in **short supply**. Large scale **copper mining** leaves **scars on the landscape** and produces significant amounts of waste rock that must be disposed of. Biological methods have a lower impact on the environment and make use of waste containing small amounts of copper. The disadvantages of **biological extraction methods** are that they are **slow**, but they do reduce the need to obtain new ore through mining and conserve limited supplies of high-grade ore.

Phytomining

Phytomining involves the use of **plants**. Plants absorb the metal compounds found in the soil. The plants cannot get rid of the copper ions and it builds up in the leaves. The plants are then **harvested, dried** and then placed in a furnace. The ash that is produced from the burning process contains soluble metal compounds that can be extracted. The ash is dissolved in an acid such as hydrochloric or sulfuric and the copper is then extracted by electrolysis or through a **displacement reaction** with iron.

Bioleaching

Bioleaching uses **bacteria** to produce an acidic solution called **leachate** which contains **copper ions**. The disadvantage of bioleaching is that it produces **toxic substances** that are **harmful to the environment**. To process the copper, the copper undergoes a displacement reaction with iron. Iron is cheaper and a **more cost-effective** way of producing copper from the leachate.

Corrosion

Metals can corrode when **exposed** to **oxygen**; they oxidise and can form metal oxides. Some metals oxidise more quickly than others, like sodium, and some such as gold are very unreactive and do not oxidise at all.

Corrosion occurs when a metal continues to oxidise and the metal becomes weaker over time until it eventually becomes a metal oxide.

Rusting occurs when **iron or steel** reacts with **oxygen** in the **air or water**. Rusting is an example of corrosion.

iron + oxygen + water ⟶ hydrated iron(III) oxide

$4Fe + 3O_2 + 6H_2O \longrightarrow 4Fe(OH)_3$

How Can Rusting Be Prevented?

To prevent rusting, oxygen and water must be kept away from the iron or steel.

Storing the metal in an atmosphere containing unreactive argon prevents it from reacting with oxygen.

A substance such as calcium chloride can be used to absorb water vapour and keep the metal dry.

AQA GCSE Chemistry (Separate Science) Unit 10: Using Resources

Barriers to Prevent Rusting

There are several different methods that are used to prevent rusting.

1. painting
2. coating with plastic
3. oiling and greasing

Electroplating

To improve the appearance of metal or to prevent corrosion, a thin layer of a metal can be applied to an object using electrolysis. This process is called **electroplating**.

In electrolysis, there are two electrodes – the **positive anode** (plating metal) and the **negative cathode** (the iron or steel object). The **electrolyte** is the solution that contains the metal ions needed to plate the metal. For example, cutlery made of steel can be electroplated with silver.

Sacrificial Protection

Metals such as iron can be **prevented** from rusting if they are put into contact with **more reactive metals** such as zinc. The reactive metals will react more readily with oxygen whilst iron does not corrode.

We say that the more reactive metal has **'sacrificed'** itself. Once the more reactive metal has corroded away, it can simply be replaced.

Galvanising

Galvanisation is the process of coating iron with zinc. The purpose is to **prevent** oxygen and water reacting with iron and so prevents **rusting**. Zinc acts as a **sacrificial** metal.

Alloys

Name of Alloy	Component Metals	Uses
bronze	copper and tin	bells coins statues
brass	copper and zinc	locks taps instruments door hinges door knobs
gold	Alloyed with other metals such as silver, zinc and copper.	jewellery

Steel Alloys

Steel is an **alloy** made up of **iron** mixed with certain amounts of **carbon**. Different steels have different properties and this determines their use.

- **High-carbon steel** contains a **high proportion of carbon**. This type of steel is **strong and brittle** and is used in the construction industry.
- **Low-carbon steel** contains a **low proportion of carbon** and is softer and more **easily shaped**. This makes it useful for making car body panels.
- **Stainless steel** is made up of **iron** but also the elements **chromium** and **nickel**. It is used for making **cutlery** as it does not rust.

Glass

Glass is made by **melting** a mixture of **sand** (silicon dioxide), **limestone** and **sodium carbonate**. Once it has melted, the molten liquid then cools and solidifies. Glass made with this mixture of ingredients is called soda-lime glass. **Soda-lime glass** is used for window panes, glass jars and bottles.

Glassware that is used in **baking** and in the **laboratory** contains boron trioxide. **Borosilicate glass** has a **higher melting point** than soda-lime glass which makes it better suited to its function where high temperatures are often used.

Ceramics

Ceramics made from **clay** include china, porcelain and brick. Wet clay is shaped and then placed into a furnace where it is heated to a **high temperature**. Crystals form in the clay and join it together.

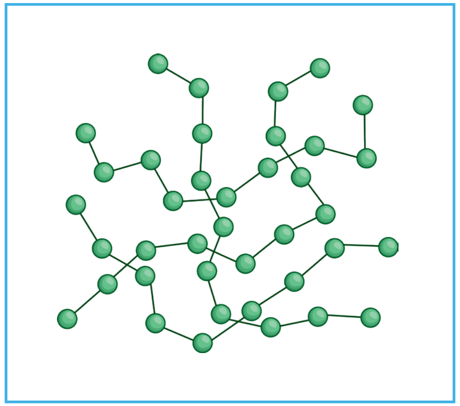
Dinner plates and bowls are made from clay ceramics. Once taken out of the furnace, the ceramics are allowed to cool and are coated with a glaze. This glaze hardens over time and forms a waterproof layer.

Polymers

Polymer properties are dependent upon the **monomer** that it is made from and the conditions in which it was made. For this reason, different polymers have different jobs. For example, **low-density** (LD) and **high-density** (HD) poly(ethene) are made from the monomer ethene using different catalysts and reaction conditions. Low-density poly(ethene) LDPE is flexible and is commonly used in carrier bags and bubble wrap. High-density poly(ethene) HDPE is much stronger, flexible, resists shattering and chemical attack. It is commonly used in plastic bottles, pipes and buckets.

Thermosetting and Thermosoftening Plastics

The polymer chains in **thermosetting plastics** are held together by strong **covalent bonds**. This means that plastics in this group can withstand **higher temperatures** and do not melt when heated – they have **high melting points**. Thermosetting plastics are used to make electrical plugs. Even if there is a fault and the wiring becomes hot, the plastic casing will not melt.



Composite Materials

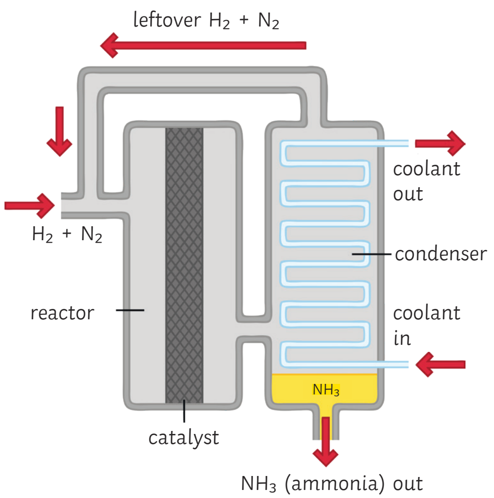
Composites are made up of two materials: a **reinforcement** and a **matrix** which binds the reinforcement together.

Wood is a natural composite. The matrix is **lignin** which is a material that can be found lining the xylem vessels of plants. Wood is reinforced with **cellulose**; in wood, the cellulose fibres are lined up next to each other and this makes the wood stronger in one direction than another. **Chipboard** is a material that can be used for kitchen worktops and doors. Chipboard is made up of **wood chips** (reinforcement) that is randomly arranged and held together by **resin glue** (matrix). This makes it **strong in all directions**.

Fibreglass and carbon fibre reinforced polymer (CFRP) contain fibres that are strong under tension. Fibreglass contains glass fibres and CRFP contains **carbon fibres**, both of the fibre types are used as **reinforcement**. The fibres themselves are flexible but do not easily stretch. The fibres in each of these composite materials are held together by **polymer resin** (matrix) which helps to bind the fibres together making them stiff.

Concrete is such a versatile material and is often used in the construction industry. The strength of concrete can be increased by **reinforcing** it with other materials such as **wire mesh or steel rods**. The compressive strength of **concrete** (matrix) is greater than its tensile strength. This means that it can withstand more force from crushing than it can force under tension. **Steel** (reinforcement), on the other hand, has greater tensile strength. This means that by combining the two materials, one is created that is both strong under tension and strong under compression. This makes reinforced concrete an important material in the construction of large buildings.

Haber Process



The Haber process is used by the chemical industry to synthesise **ammonia**. Ammonia is used in the production of fertilisers, dyes and explosives.

The reaction is a **reversible** one and involves nitrogen reacting with hydrogen to produce ammonia. As the reaction is reversible, some of the ammonia will decompose back into nitrogen and hydrogen.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

The reaction mixture is cooled, the ammonia liquifies and is then removed. The hydrogen and nitrogen that has not reacted is recycled to increase the efficiency of the process. The reaction reaches **dynamic equilibrium** and this is where the rate of the forward reaction occurs at the same rate as the backward reaction.

In the Haber process, nitrogen and hydrogen are pumped through pipes at a pressure of **200 atmospheres**.

Nitrogen is obtained by extraction from the air and hydrogen is obtained from natural gas. The gases are passed through a tank containing a **catalyst** (iron); catalysts speed up the rate of a chemical reaction without getting used up themselves. The gases are heated to **450˚C** as they pass through the tank.

The reaction mixture is allowed to cool and this allows the ammonia to turn from a gas to a liquid. Once this has happened, the ammonia is removed. Any unreacted nitrogen and hydrogen is then recycled.

Fertilisers

Fertilisers contain lots of **mineral ions** that are key to the growth of healthy crops. Plants absorb these minerals through their root hair cells; these mineral ions need to be replaced and so farmers need to add fertiliser to the soil in order to replace the lost mineral ions.

Farmers often use **NPK fertilisers**. These are fertilisers that contain the elements **nitrogen, phosphorus and potassium**.

- Ammonium nitrate - NH_4NO_3 - and ammonium sulfate - $(NH_4)_2SO_4$ - are examples of fertilisers that contain the essential element nitrogen.
- Ammonium phosphate - $(NH_4)_3PO_4$ - contains the elements nitrogen and phosphorus.
- Potassium nitrate - KNO_3 - contains the elements potassium and nitrogen.

Ammonia

Ammonia has the chemical formula NH_3.

Ammonia produces the ammonium ion NH^{4+} when it is involved in neutralisation reactions. Ammonia is an alkali. Oxidation of ammonia produces nitric acid HNO_3; nitric acid is the source of the nitrate ion NO^{3-}.

alkali + acid ⟶ salt

ammonia + nitric acid ⟶ ammonium nitrate

$NH_3 + HNO_3 \longrightarrow NH_4NO_3$

In aqueous solutions:

ammonium hydroxide + nitric acid ⟶ ammonium nitrate + water

$NH_4OH + HNO_3 \longrightarrow NH_4NO_3 + H_2O$

Mining

The raw materials for fertilisers need to be mined. The minerals needed to make fertilisers are extracted from the **earth's crust**.

Potassium chloride and **potassium sulfate** are a source of potassium ions and are used as fertilisers. **Phosphate rock** is **insoluble** and so cannot be used in fertilisers, but it does contain **phosphorus** which when reacted with acid, will produce **soluble compounds**.

Phosphate rock when reacted with **nitric acid** produces calcium nitrate and phosphoric acid.

Phosphate rock when reacted with **sulfuric acid** produces a mixture of calcium sulfate and calcium phosphate which is called single superphosphate.

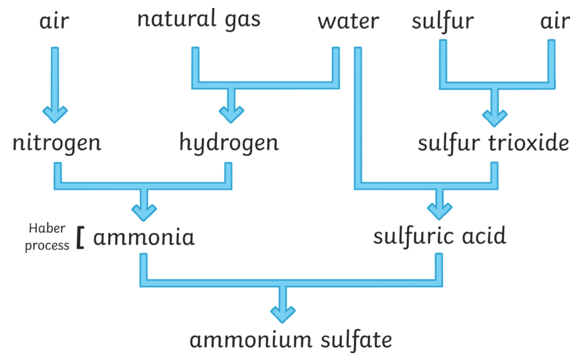
Phosphate rock when reacted with **phosphoric acid** produces calcium dihydrogen phosphate also called triple superphosphate.

Ammonium Sulfate

The salt ammonium sulfate ($(NH_4)_2SO_4$) is used as a fertiliser and is made when ammonia and sulfuric acid react.

ammonia + sulfuric acid ⟶ ammonium sulfate

$2NH_3 + H_2SO_4 \longrightarrow (NH_4)_2SO_4$



Chemical Industry

To make sulfuric acid, sulfur, air and water are needed.

Sulfur first reacts with oxygen to produce sulfur dioxide. The sulfur dioxide further reacts with oxygen at a temperature of 450˚C to produce sulfur trioxide. This in turn reacts with water to produce sulfuric acid.

In the Laboratory

Ammonium sulfate is produced by reacting ammonia solution with sulfuric acid.

Stage 1 – A measured amount of ammonium sulfate solution is poured into a conical flask.

Stage 2 – Two to three drops of the indicator methyl orange is added. The ammonia solution will turn yellow as it is an alkaline.

Stage 3 – The conical flask is placed under a burette containing sulfuric acid. Slowly the sulfuric acid is added to the flask until the indicator turns orange. If the indicator turns red, this means that too much acid has been added.

Stage 4 – Once the solution turns orange, the volume of acid that was added is recorded and the neutral ammonium sulfate solution containing the indicator is discarded.

Stage 5 – The experiment is then repeated with the same volumes of sulfuric acid and ammonia solution, but this time the indicator is not added. The solution is then heated and the water evaporates leaving behind crystals. The crystals left in the evaporating basin are then placed in an oven.

The Advantageous and Disadvantages of Industrial vs Laboratory Method of Fertiliser Production

Industrial Method

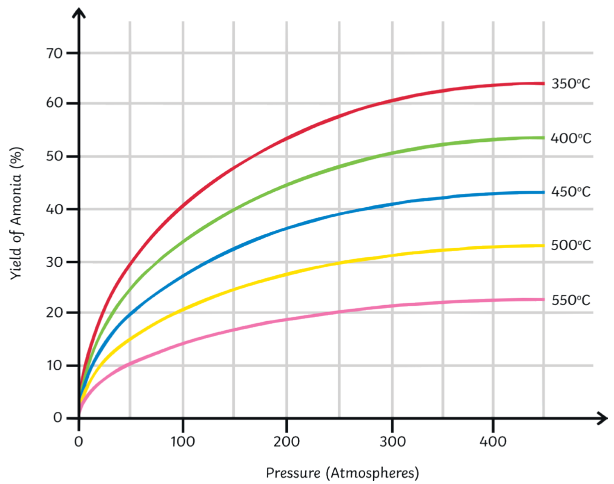
The industrial method of production requires a temperature range between 60-450°C, depending on the stage in the production process. As this is a **continuous method** of production, it requires the use of expensive machinery. The starting materials in this method are acquired from **raw materials** with **large quantities** of fertiliser being made **quickly**. The cost of labour is reduced by using **automated mechanisms** and machines.

Laboratory Method

The laboratory method, on the other hand, is much **slower** and more **labour intensive** and this makes the **running costs high**. The starting materials for this method are purchased directly from a chemical supplier. As this is a batch process, the **equipment** used is **relatively cheap**. A Bunsen burner is used for heating and room temperature is required for the neutralisation stage.

Haber Process – Higher Tier Only

The graph shows that as the temperature increases, the yield of ammonia decreases.



Increasing Temperature

As the temperature increases, the **equilibrium** position moves to the **left** and the **yield** of ammonia **decreases**. Using a low temperature may seem the most sensible option, but if the temperature is too low then the rate of reaction will also be reduced. That is why the temperature that is chosen is a **compromise** between the **yield** and **rate of reaction**.

Increasing Pressure

In a reaction where gas particles are reacting or produced, increasing the pressure will **shift** the **equilibrium** position to the side with the **fewest moles of gas**.

In the Haber process, the right-hand side of the equation has the fewest number of molecules; if the pressure is increased, then the equilibrium position will shift to the right and the yield of ammonia will increase. The disadvantage to using higher pressures are that more expensive equipment is required to cope with the increased pressure and this increases energy costs. The decision here is a **compromise** between **yield** and **cost**.

Catalysts

Catalysts are useful in the Haber process as they **speed** up the rate of reaction in both directions. The time taken for the system to reach **equilibrium** is reduced. A catalyst does not affect the position of the equilibrium or the yield. Using a catalyst allows a low temperature to be used whilst also keeping the yield high.

Reducing Cost

Any unreacted hydrogen and nitrogen are recycled back into the reactor and this reduces the cost of making raw materials. Energy is a large cost. Often, exothermic reactions where energy is released are used to heat up other parts of the process.