Unit 2 Chemistry GCSE Double Award
Revision Guide

Elements
The simple building block of all substances that cannot be broken down into anything simpler. Consist of only one type of atom.
An atom consists of a positively charged nucleus (containing protons and neutrons) surrounded by orbits/shells (containing electrons).

Periodic Table
This is used to identify all the known elements.
Group – The column where the element is placed.
Period – The row where the element is placed.
Some elements can show properties of both metals and non-metals. These elements are located in groups 3, 4 and 5. They are called **METALLOIDS/SEMI-CONDUCTORS**. An example is Silicon (Si).

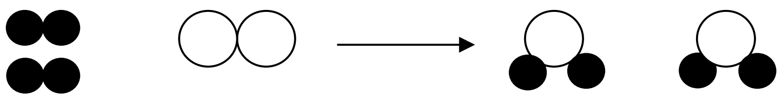
As we go down certain groups, there are trends/patterns in their properties e.g. as we descend group 1, the melting point decreases.
Elements in the same group will have **SIMILAR** properties e.g. group 1 metals all react in water and produce an alkali, as we go down the reactivity of the metal increases.

Metals/Non-metals

Metals	Non-metals
Located on the left hand side of the table	Located on the right hand side
Conduct heat	Does not conduct heat
Conduct electricity	Does not conduct electricity
Malleable (Hammered into shape)	Brittle
High melting point (except Mercury)	Low melting point
High density	Low density
Shiny (Lustre)	Dull

Compounds
Compounds consist of two or more different elements/atoms that have joined together.
The atoms have been **rearranged** during a chemical reaction; however **no** atoms have been gained or lost.

E.g. Hydrogen + Oxygen → Water



Word equation: Hydrogen + Oxygen → Water

Symbol equation: $2H_2 + O_2 \rightarrow 2H_2O$

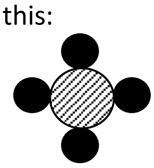
Drawing substances using a key

We can use a key, which represents an atom of an element, to draw molecules of compounds.
Key points

- The atoms **MUST** be touching.
- If you have two different atoms, the element with the most atoms will **SURROUND** the element on its own.

E.g. Methane (CH_4) ● = Hydrogen; ◍ = Carbon

A molecule of methane will look like this:



Ions
Ions are formed when an atom needs to gain or lose electrons to form a full outer shell.
Positive ions = Atoms losing electrons.
Negative ions = Atoms gaining electrons.

Metal atoms form POSITIVE ions; Non-metal atoms form NEGATIVE ions.

The charge of the ion tells us how many electrons the atom has gained or lost e.g.

- Magnesium ion is Mg^{2+}, so the Magnesium atom has LOST 2 electrons for a full outer shell.

Writing formulae
We can use ions to write chemical formulae of compounds.
Key rule
The charges of the ions must balance before writing the formula.

E.g. Sodium Oxide
Sodium ion = Na^{+}
Oxide ion = O^{2-}

We need two sodium ions to balance the charge of the oxide ion.
The chemical formula for sodium oxide is Na_2O.

Brackets are used in chemical formulae when you need two or more ions that contain two different elements.
E.g. Ammonium Sulphate
Ammonium ion = NH_4^{+}
Sulphate ion = SO_4^{2-}

We would need two ammonium ions to balance the charges.

The chemical formula for ammonium sulfate is $(NH_4)_2SO_4$.

You will be given a table of ions in the examination paper.

We can use the chemical formulae of a compound to identify:
- The elements present in the compound
- The number of atoms for each element
- The total number of atoms in the compound

E.g. Water (H_2O) contains two elements (Hydrogen and Oxygen). There are **two** atoms of Hydrogen (H_2) and **one** atom of oxygen (O). In total, there are **three** atoms in water.

Ammonium sulfate, $(NH_4)_2SO_4$, contains four elements (Nitrogen, Hydrogen, Sulfur & Oxygen). By multiplying the number outside the brackets with the number of atoms inside, we have 2 nitrogen atoms, 8 hydrogen atoms, 1 sulfur and 4 oxygens. In total, there are 15 atoms (2+8+1+4=15).

The Atom
Consists of a positively charged nucleus (containing protons and neutrons) surrounded by orbits/shells (containing electrons).

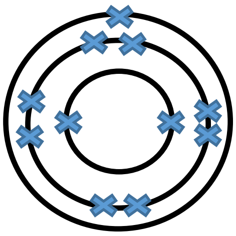
Particle	**Mass**	**Charge**	**Location**
Proton	1	+1	Nucleus
Neutron	1	0 (Neutral)	Nucleus
Electron	Negligible	-1	Shells/Orbits

Atoms have no charge due to the same number of protons and electrons.

Ions
Atoms that have gained or lost electrons to obtain a full outer shell.
Positive ion – Lost electrons, Negative ion – Gained electrons.
The value of the ion depends on the number of electrons gained/lost.

Atomic number – Number of **protons**/electrons in an atom.
Mass number – **Total** number of protons and neutrons in an atom.
Number of neutrons = Mass number – Atomic number

Electron structure
First shell, maximum of two electrons; second & third shell, maximum of eight.
E.g. Sodium (Na) – Atomic number 11. Electron configuration – 2, 8, 1



Group number (column) – Number of electrons on the outer shell.
Period number (row) – Number of shells in the configuration.
E.g. Sodium is in group 1 (1 electron on the outer shell) and period 3 (3 shells in the configuration).
Group 0 elements have a full outer shell, which is responsible for its un-reactivity.

Isotopes
Two or more atoms of the same element with the same number of protons (atomic number) but a different number of neutrons (mass number). Example is ^{12}C and ^{14}C.
^{12}C has 6 protons and 6 neutrons, whilst ^{14}C has 6 protons and 8 neutrons.

Separating mixtures methods

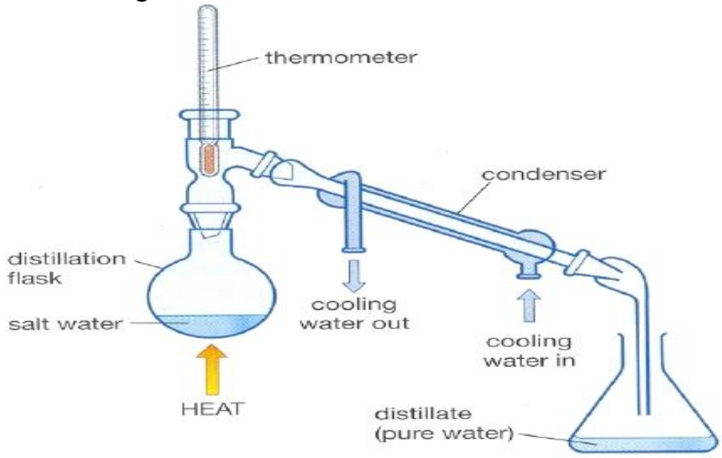
Mixtures contain elements, compounds or both that have NOT chemically joined together. We can separate mixtures by physical process such as:

- Distillation
- Chromatography
- Filtration
- Evaporation

Distillation – Process to separate a mixture of liquids by their **boiling points**.

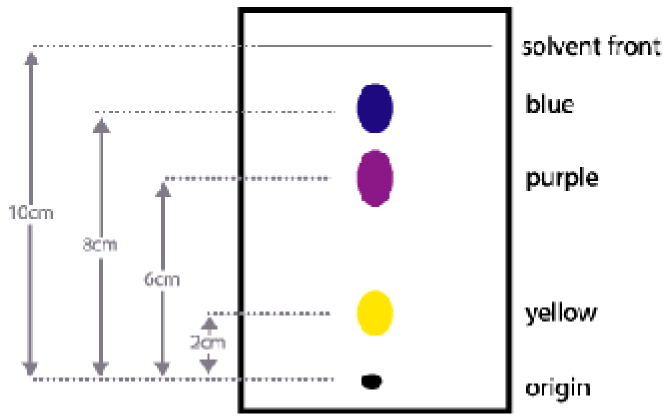
E.g. Water and ethanol

- Water has a b.pt of 100˚C, whilst ethanol has a b.pt of 78˚C.
- By heating the mixture to 80˚C, the ethanol will evaporate first (due to the low b.pt.).
- The ethanol vapour will enter a condenser, which will condense the ethanol vapour into a liquid.
- The condensed ethanol will leave the condenser and collected in a beaker.
- Water will remain in the starting mixture flask.



Chromatography

- Process used to show the solubility of a solute (e.g. ink), within an appropriate solvent, by measuring the height of the pattern on chromatography paper.
- Some solutes have a different solubility in different solvents i.e. some are insoluble in water but soluble in ethanol.
- The higher up the dye on the paper, the more soluble it is.
- If the patterns align/match together, they contain the same dyes.



- On the paper is a pencil line, where the ink spots are placed before putting the paper into the solvent.

Retardation factor (R_f)

The R_f value determines the identity of the substance.

- **Equation**: **R_f = Height the dye travelled from the pencil line ÷Height of the solvent front**

Calculate the R_f value for the blue dye

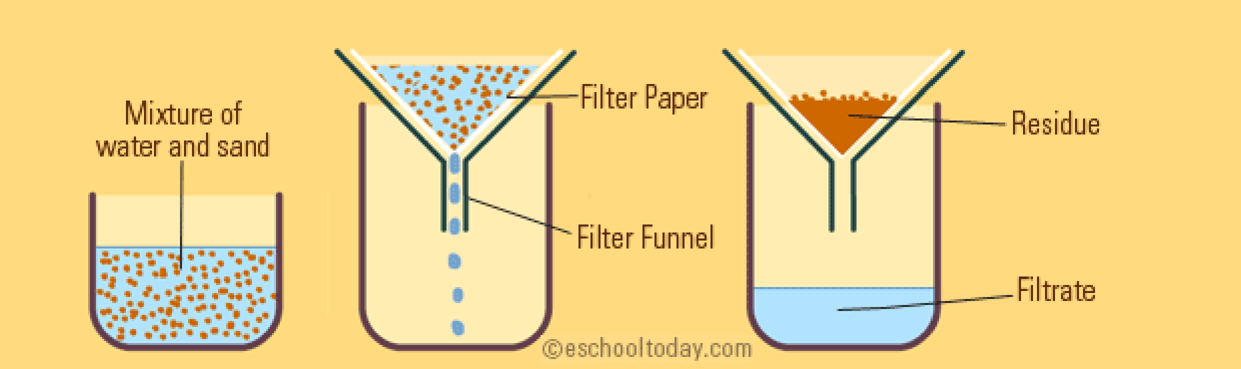
- Height travelled by blue dye = 8 cm
- Height of the solvent front = 10 cm
- R_f = 8 / 10 = 0.8

Filtration

This is the process used to separate insoluble solutes from a solvent/solution.

To carry out the process we require filter paper, filter funnel and a beaker/evaporating dish to collect the pure solution (filtrate).

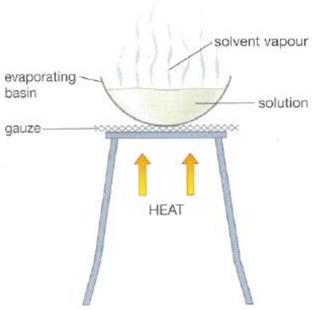
An example is separating sand from sea water.



Evaporation

This is the process used to separate a liquid from a soluble solid/solute.

This method is used if we do not require the liquid e.g. salt and water.



Group 1 Elements
Collective Name = Alkali metals

In the lab, alkali metals are stored in oil – Prevent water and oxygen from getting to the metal.

Reaction with oxygen (O_2)
Produce a white solid and a coloured flame is observed. The metal tarnishes quicker as we descend the group.
Colours: Lithium (Li) – Red; Sodium (Na) – Yellow/Orange; Potassium (K) – Lilac.
Word equation: Lithium + Oxygen → Lithium Oxide
Symbol equation: $\underline{\mathbf{4}}Li + O_2 \rightarrow \underline{\mathbf{2}}Li_2O$

Reaction with water (H_2O)
Produce a colourless metal hydroxide solution and hydrogen gas.

Metal	Observations
Lithium	Floats, fizzes, melts, moves on the surface.
Sodium	Floats, fizzes, melts **into a sphere**, moves on the surface.
Potassium	Floats, fizzes, melts, moves on the surface, **self-ignites** (a lilac flame is seen).

Word equation: Sodium + Water → Sodium Hydroxide + Hydrogen
Symbol equation: $\underline{\mathbf{2}}Na_{(s)} + \underline{\mathbf{2}}H_2O_{(l)} \rightarrow \underline{\mathbf{2}}NaOH_{(aq)} + H_{2(g)}$

Safety aspects of reaction with alkali metals
- Use small pieces (large pieces cause an explosion);
- Wear gloves/handle with tweezers (moisture from the hands);
- Use a safety screen (to prevent splashing of alkali solution).

Reaction with halogens e.g. chlorine (Cl_2) – Produce a white solid and a coloured flame is observed.
- Lithium (Li) – Red; Sodium (Na) – Yellow/Orange; Potassium (K) – Lilac.
- Word equation: Potassium + Chlorine → Potassium Chloride
- Symbol equation: $\underline{\mathbf{2}}K + Cl_2 \rightarrow \underline{\mathbf{2}}KCl$

Testing for metal ions in compounds – Flame test
Use nichrome wire and place the sample on the wire.
Insert the sample and wire into a Bunsen flame.

Metal	Flame Colour
Lithium (Li)	Red
Sodium (Na)	Yellow-Orange
Potassium (K)	Lilac
Calcium (Ca)	Brick red
Barium (Ba)	Apple green

Reactivity of alkali metals – Increases as we go down the group.

HIGHER TIER ONLY Why?
- As you descend the group, number of shells increases.
- Force of attraction between nucleus and outer electron decreases.
- Easier to lose the outer electron for a full shell.

Group 7 Elements – Collective Name = Halogens
Halogens exist as diatomic molecules (Cl_2, Br_2, I_2)
Chlorine and iodine are obtained from sea water (through electrolysis). However, there is very little iodine available in sea water.

Colours and states of halogens at room temperature

Halogen	State at room temperature	Colour at room temperature
Chlorine (Cl_2)	Gas	Yellow-Green
Bromine (Br_2)	Liquid	Orange-red
Iodine (I_2)	Solid	Grey

As you descend the group, the colour gets darker and the states change from gas-liquid-solid.

Reaction with alkali metals e.g. Sodium (Na)
Produce a white solid and a coloured flame is observed (see alkali metals notes for colours)
Word equation: Sodium + Bromine → Sodium Bromide
Symbol equation: $\underline{\mathbf{2}}Na + Br_2 \rightarrow \underline{\mathbf{2}}NaBr$

Reaction with iron – Produce a solid and brown fumes are observed.
Word equation: Iron + Chlorine → Iron (III) Chloride
Symbol equation: $\underline{\mathbf{2}}Fe + 3Cl_2 \rightarrow \underline{\mathbf{2}}FeCl_3$ (Use the iron (III) ion for the symbol equation)

Reactivity of halogens – Decreases as we go down the group

Testing for halogen ions in compounds – Silver Nitrate solution ($AgNO_3$)
When the halide compound reacts with silver nitrate, a precipitate (silver halide) is formed.
The colour of the precipitate depends on the halide ion.
Precipitate – An insoluble solid formed when two solutions mix.

Halide ion	Colour of precipitate with silver nitrate	Name of precipitate
Chloride ion (Cl^-)	White ppt	Silver chloride (AgCl)
Bromide ion (Br^-)	Cream ppt	Silver bromide (AgBr)
Iodide ion (I^-)	Yellow ppt	Silver Iodide (AgI)

Word equation: Sodium Chloride + Silver Nitrate → Silver Chloride + Sodium Nitrate
Symbol equation: $NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow AgCl_{(S)} + NaNO_{3\,(aq)}$
Ionic equation: $Cl^-_{(aq)} + Ag^+_{(aq)} \rightarrow AgCl_{(S)}$

HIGHER TIER ONLY
Displacement reactions of halogens
- Halogens of a higher reactivity can displace the halide, of a lower reactivity, within a compound.
- As a result, the solution will change colour.

Halogen	Metal Chloride (MCl)	Metal Bromide (MBr)	Metal Iodide (MI)
Chlorine (Cl_2)	X	Turns orange-brown	Turns brown
Bromine (Br_2)	No reaction	X	Turns brown
Iodine (I_2)	No reaction	No reaction	X

E.g. Lithium Iodide + Chlorine → Lithium Chloride + Iodine (Brown colour solution observed – Iodine)
$\underline{\mathbf{2}}LiI + Cl_2 \rightarrow \underline{\mathbf{2}}LiCl + I_2$

Reactivity of halogens – Decreases as we go down the group. Why?
- As you descend the group, number of shells increases.
- Force of attraction between nucleus and outer electron decreases.
- Harder to attract the required electron for a full shell.

Testing Gases

Gas	Test	Observation
Hydrogen	Lighted splint	Goes pop
Oxygen	Glowing splint	Re-ignites
Carbon Dioxide	Limewater	Goes cloudy/milky

Uses and properties of non-metals

Gases	Use	Properties
Chlorine	Sterilizing water supplies	Kills bacteria (poisonous)
	To make bleach	Kills bacteria
Iodine	Antiseptics	Kills bacteria
Helium	Airships, weather balloons	Low density, unreactive
Argon	Light bulbs, welding	Unreactive
Neon	Advertising signs	Unreactive to electricity, emits coloured light

Calculations

Relative atomic mass (A_r) – The average mass of an atom for the isotopes an element.
Example- Chlorine has two isotopes (^{35}Cl and ^{37}Cl). The A_r value is 35.5 as it has 75% of ^{35}Cl and 25% of ^{37}Cl.

Relative molecular mass (M_r) – The sum of the A_r of the atoms in the numbers shown in its formula.
Example – Carbon Dioxide (CO_2) where A_r C = 12, O = 16. So the M_r of CO_2 = 12 + 16 + 16 = 44

% composition of an element in a compound

Calculate the % of iron present in iron (III) oxide, Fe_2O_3. A_r Fe = 56, O = 16

Step 1 – Calculate the total mass of the element present in the compound.

Total mass of iron = 2 x 56 = 112

Step 2 – Calculate the M_r of the compound.

M_r Fe_2O_3 = (2 x 56) + (3 x 16) = 112 + 48 = 160

Step 3 – Use the equation below to calculate the % mass of the element:

% composition of element = (Total mass of element ÷ M_r of compound) x 100

% composition Fe = (112 ÷ 160) x 100

% composition Fe = 70.0%

Percentage yield

Percentage yield = (Mass of product obtained / Expected mass) x 100

E.g. 12g of magnesium combusted in excess oxygen gas to produce 18g of magnesium oxide.
Calculate the yield of the magnesium oxide.

$$2Mg + O_2 \rightarrow 2MgO$$

A_r Mg = 24, O = 16

Step 1 – Cross out any unnecessary substances that are not in the question (Cross out the oxygen as it is excess).

Step 2 – Calculate the masses of the substances using their atomic masses.

2 Mg = 2 x 24 = 48

2 MgO = (24 + 16) x 2 = 40 x 2 = 80

Step 3 – Calculate the ratio between the substances to obtain the expected mass

48g into 12 g of Mg (Divide by 4) so the expected mass of magnesium oxide will be 80 ÷ 4 = 20g of magnesium oxide.

Step 4 – Use the expected mass (calculated from step 3) and the mass obtained (from the question) to calculate the yield.

Mass obtained of magnesium oxide = 18g, Expected mass of magnesium oxide = 20g.

Percentage yield of magnesium oxide = (18 / 20) x 100 = 90%

HIGHER TIER ONLY

Formula of a compound using masses/percentages.

E.g. 7.2g of copper oxide contains 6.4g of copper; calculate the formula of copper oxide. A_r Cu = 64, O = 16.

Step 1 – Calculate mass of other element in the formula

Mass of oxygen = 7.2 – 6.4 = 0.8g

	Cu	O
Step 2 – Divide masses by the A_r of the element	6.4 / 64 = 0.10	0.8 / 16 = 0.05
Step 3 – Divide by the smallest answer to obtain number of atoms.	0.10 / 0.05 = 2	0.05 / 0.05 = 1

Formula of copper oxide = Cu_2O

The same methods can be used if percentages of elements are given, rather than masses.

Masses of products & reactants from an equation

E.g. Calculate the mass of iron produced when 480 tonnes of iron (III) oxide is reduced by carbon monoxide.

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

A_r Fe = 56, C = 12, O = 16

Step 1 – Cross out the substances not required in the reaction (3CO & $3CO_2$).

Step 2 – Calculate the M_r of the remaining substances.

Fe_2O_3 = (56 x 2) + (16 x 3) = 160

2Fe = (2 x 56) = 112

160 tonnes of iron (III) oxide produces 112 tonnes of iron.

Step 3 – Calculate the ratio between the masses from step 2 with the mass in the question.

160 → 480 tonnes of iron (III) oxide (x3), so mass of iron formed = 112 x 3 = 336 tonnes.

Moles

The number of moles of a substance can be calculated by the following equation:

Moles = Mass of substance ÷ M_r of substance

How many moles are present in 11g of carbon dioxide, CO_2? A_r C = 12, O = 16.

M_r CO_2 = 12 + (2 x 16) = 44

Moles CO_2 = 11 ÷ 44

Moles CO_2 = 0.25

Converting moles into mass

To convert moles into mass, we re-arrange the moles equation to become:

Mass = Moles x M_r

How much mass does 0.25 moles of CO_2 weigh?

Mass = 0.25 x 44

Mass = 11g

To work out how many atoms/molecules of a substance is present, we multiply the number of moles with Avogadro's constant (6×10^{23}).

Number of atoms/molecules = Moles x Avogadro's constant

How many molecules are present in 11g of carbon dioxide?

We worked out above that there are 0.25 moles of carbon dioxide in 11g.

Number of molecules = $0.25 \times 6\times10^{23}$

Number of molecules = 1.5×10^{23}

Water

Three processes for the treatment of water

- **Sedimentation** – Removal of larger, insoluble solid particles settle to the bottom by gravity.
- **Filtration** – Removal of smaller, insoluble particles through layers of sand/gravel.
- **Chlorination** – Removal of micro-organisms by using chlorine (poisonous) to prevent disease/making it safe to drink.

Conservation of water - Increase in the number of droughts/Water in high demand due to population and industry.

Ways to conserve water

- Use a shower instead of a bath
- Hose-pipe ban
- Wash clothes in a large bulk
- Outside water storage tank to collect rain water.

Desalination of sea water

- Removal of salt from sea water – Using the process of distillation.

Advantages	Disadvantages
• Source of drinking water for places near the sea • The water can be used as a source of cheap energy. • Increase the economy.	• Requires a lot of energy to boil the sea water. • Very expensive

Other methods used for desalination – Vacuum distillation, use of membrane systems.

Hard Water - Caused by soluble calcium (Ca^{2+}) and magnesium (Mg^{2+}) compounds.

Testing the hardness of water

Experiment: Adding soap solution and shake.

- **Soft water** – Lathers easily with small volumes of soap.
- **Hard water** – Forms a scum (white crystals), very little/no lather (only lathers with large volumes of soap).

Permanent & Temporary Hard Water

- Temporary Hard Water – Softens **AFTER BOILING**. Volume of soap used to lather will decrease after boiling.
- Permanent Hard Water – Does not soften after boiling. Volume of soap remains high after boiling.

Sample	Volume of soap to lather before boiling (cm^3)	Volume of soap to lather after boiling (cm^3)	Type of water
A	10.0	2.0	Temporary
B	10.0	10.0	Permanent
C	2.0	2.0	Soft
D	10.0	7.0	Mixture of temporary & permanent

Ways to soften permanent hard water

- **Ion exchanger**
  - A tube containing sodium ions within the resin (beads).
  - When hard water is poured, the calcium ions will attach to the resin, releasing the sodium ions into the water.
  - Over time, all the sodium ions, from the resin, will be released. So a concentrated solution of sodium chloride (containing sodium ions) will regenerate the resin.
- **Adding washing soda** (sodium carbonate) crystals
  - The calcium ions will join with the carbonate to form a precipitate of calcium carbonate (limescale).
  - **Equation**: $Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow CaCO_{3\,(s)}$
- **Boiling** (Only temporary hard water)
  - Temporary hard water contains calcium hydrogen carbonate which produces limescale after boiling.

Advantages & Disadvantages of Hard Water

Advantages	Disadvantages
• Strengthens teeth and bones (due to calcium present) • Reduces heart disease	• Build-up of limescale in the hot water pipes. • Electrical appliances will be less efficient due to the limescale build-up. • Use more soap to produce lather.

Fluoridation
In some areas of the UK, fluoride is added into the water supply. We can obtain fluoride from toothpaste and mouthwash.

Advantages
Fluoride prevents tooth decay and strengthens the enamel of teeth.
Disadvantages
Fluoride is carcinogenic (causes stomach cancer) and discolours teeth.

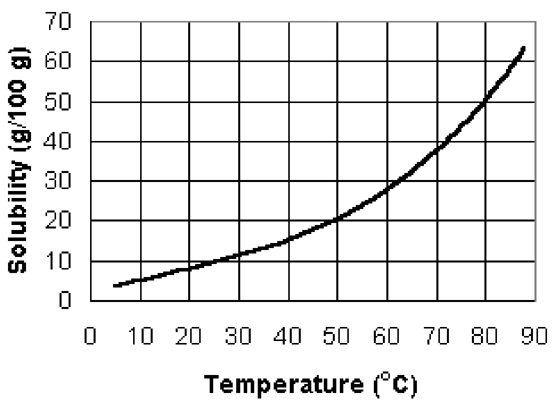
Obtaining data
Scientists obtained data by carrying out a survey on school children.
The data is reliable as only children present in school were surveyed, absentees were not included.

You cannot compare the effects of fluoridated water in two different areas due to other factors could be responsible for the data e.g. lifestyle, economic background.
The water company tend to be biased towards fluoridation and provides necessary data.

<u>Solubility Curves</u>
- Method used to determine how much solute can dissolve in a specific volume of solvent at various temperatures.

Factors that can affect solubility
- Temperature of the solvent
- Mass of solvent
- Type of solute

Using the solubility curve
We can use the curve to determine the mass of crystals formed when a solvent cools.
- Mass of crystals = Mass of solute dissolved at high temp – Mass of solute at low temp



Calculate the mass of crystals formed from 80˚C to 50˚C
- Solubility at 80˚C = 50g; Solubility at 50˚C = 20g.
- Mass of crystals = 50 – 20 = 30g

Ever-changing Earth

Plate Tectonics

The Earth's crust (lithosphere) is broken down into large pieces of rocks called Plates.
The plates can move towards or away from each other.

Alfred Wegener

A scientist called Alfred Wegener had a theory that all of the Earth's continents were once joined, forming a large continent called Pangaea. Over time, the continents drifted apart to its current position. This theory is called "Continental Drift".

Evidence

To support his theory, Alfred Wegener collected the following evidence:

- Outline of continents fitting together like a jig-saw e.g. Africa and South America.
- Similar rocks discovered on different continents.
- Similar fossils discovered on different continents.

However, his theory was not accepted by scientists because he could not explain how the continents moved.

In 1960, scientists discovered that convection currents, within the mantle, caused the plates to move. Therefore, Wegener's theory was accepted.

Plate tectonics today

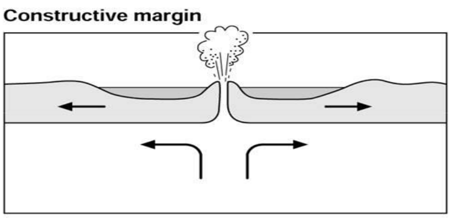
Scientists today have discovered mountain ridges deep on the ocean floor and magnetic properties in the rocks, causing the plates to slide past each other.

Constructive, Destructive and Conservative boundaries

On the boundary of plates, we would discover earthquakes, volcanoes and mountains.
There are three types of boundaries, depending on the movement of the plates.

Constructive boundary

- Plates moving away from each other.
- Forming a gap, allowing magma to escape.
- Magma cools and solidifies to form igneous rock.
- Iceland was formed as a result.

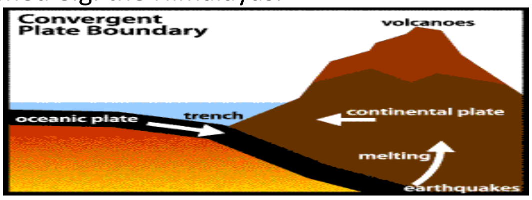


Destructive boundary (Different masses)

- Plates moving towards each other.
- The **DENSEST** plate will go beneath the other plate.
- The plate melts to form magma.
- Magma escapes through volcanic eruptions, cools to form igneous rock.

Destructive boundary (Similar masses)

- Plates moving towards each other.
- Both plates collide and pushing upwards.
- As a result, mountains are formed e.g. the Himalayas.



Conservative boundary

- When two plates slide next to each other in opposite directions, causing friction.
- The land mass does not change.
- As a result, earthquakes occur.
- Example of a conservative boundary: San Andreas Fault, California.

Atmosphere
Today's atmosphere consists of a mixture of four main gases:
- Nitrogen (N_2) – 78%
- Oxygen (O_2) – 21%
- Argon (other noble gases) – 0.9%
- Carbon dioxide – 0.04%

However the atmosphere was entirely different billions of years ago.
History of the atmosphere
- Original source of the atmosphere were volcanoes.
- Carbon dioxide and water vapour made up the majority of the atmosphere. Ammonia was also released.
- Over time, water vapour condensed to form oceans.
- Organisms took in carbon dioxide and produced oxygen (example of photosynthesis).
- Oxygen reacted with the ammonia to produce nitrogen and water.
- Carbon dioxide was trapped in rocks within the ocean, forming carbonates.

Over time, the levels of oxygen increases and the levels of carbon dioxide decreases.

Processes involving oxygen and carbon dioxide
There are three key processes that control the levels of oxygen and carbon dioxide in the atmosphere:
- **Photosynthesis** – Levels of carbon dioxide DECREASES; levels of oxygen INCREASES.
- **Respiration** – Levels of carbon dioxide INCREASES; levels of oxygen DECREASES.
- **Combustion** – Levels of carbon dioxide INCREASES; levels of oxygen DECREASES.

Global Warming
This is an effect when gases traps heat energy from the Earth and the Sun, causing the temperature of the Earth to increase. The main gas responsible for global warming is Carbon Dioxide.

Causes of Global Warming	**Consequences of Global Warming**	**Ways to prevent global warming**
• **INCREASE** in the combustion of fossil fuels - Caused by an increase in industry. • Deforestation – Cutting down trees, resulting in less photosynthesis.	• Ice caps are melting **at a faster rate**. • Sea levels are rising. • Hotter summers and colder winters. • Climate change. • Increase in flooding	• Planting more trees • Use renewable sources of energy e.g. solar, wind, hydro-electric. • Using public transport e.g. cycling. • Carbon-capture.

Acid Rain
When fossil fuels, containing sulfur, are combusted to produce an acidic gas called sulfur dioxide, SO_2.
Sulfur dioxide enters the atmosphere, combines with water vapour to produce acid rain.

Consequences of Acid Rain
- Erodes buildings (containing limestone).
- Turns rivers acidic, killing fish.
- Damaging trees and crops.
- Corrodes metals.

Ways to prevent acid rain
- Use renewable sources of energy e.g. solar, wind, hydro-electric.
- Sulphur scrubbing.

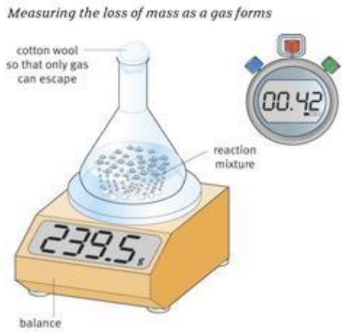
Obtaining gases from the atmosphere
Fractional distillation is used to obtain gases such as nitrogen and oxygen from the atmosphere.
How it works?
- Air enters a cooler, where water and carbon dioxide are removed (melting points above -200°C)
- The air is condensed into a liquid below -200°C.
- It is heated gently as it enters the fractionating tower.
- The lower the boiling point (higher negative value), the higher up the tower it is extracted.

Rates of Reaction

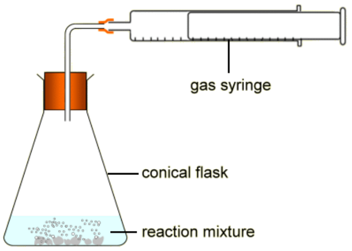
Rate – The measurement of a change over a given time. Rate = 1 / Time

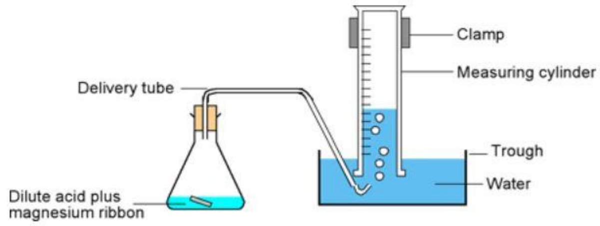
Measurements during an experiment

- The mass loss of a reactant – Use a balance to measure the mass. Cotton wool is used to allow gas to escape.

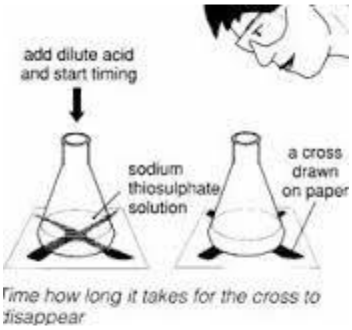


- The volume of gas produced – Use a gas syringe/measuring cylinder to measure the volume.





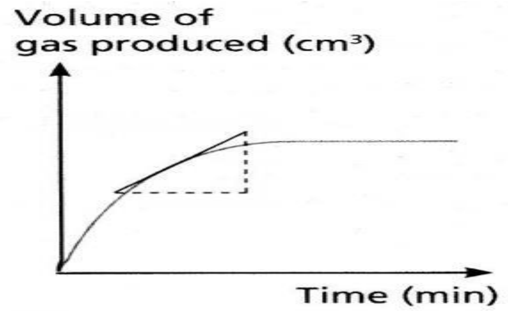
- Formation of a precipitate – Use a light sensor to measure the formation of a precipitate.

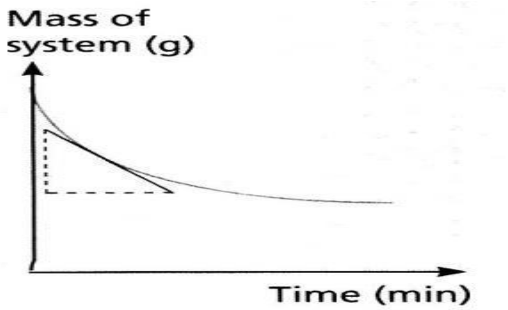


Advantages of using ICT to measure rates

- Continuous readings throughout the experiment
- A graph can be plotted automatically

Rate graphs

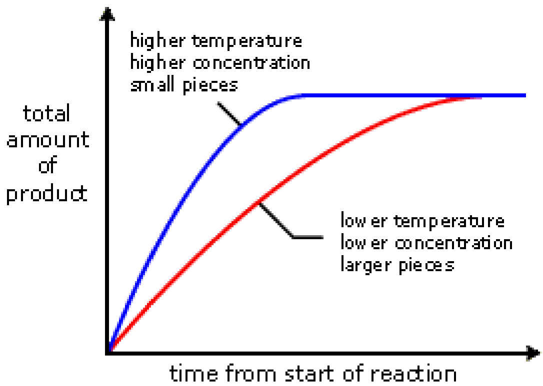




Curve goes horizontal as time goes on – Experiment has finished when the line is flat.

If the rate increases, the graphs will be steeper.

Measure rate on graphs = Change in mass or volume / Change in time



__**Factors that can alter the rate**__

Temperature – Increase in temperature, the rate of reaction increases.

- Particle theory – Particles gain energy/move faster, greater **CHANCE** of a **SUCCESSFUL** collision between the particles.

Concentration (How dilute is the solution) – Increase in concentration (no dilution), the rate increases.

- Particle theory – More particles present **IN A GIVEN VOLUME**, greater **CHANCE** of a successful collision between the particles.

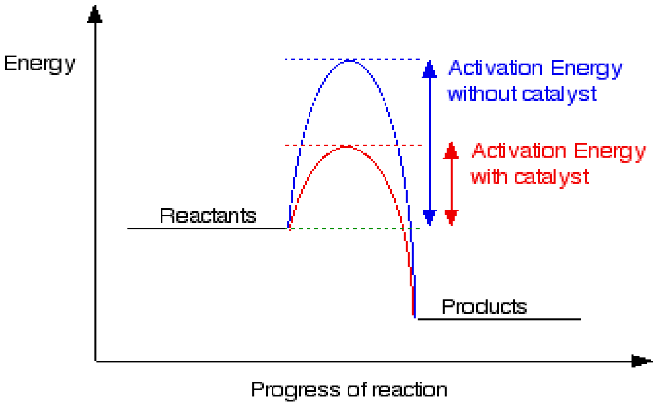
Surface area – Increase in surface area (using powder), the rate increases.

- Particle Theory – More particles exposed, greater **CHANCE** of a successful collision between the particles.

Catalyst - A chemical substance that speeds up the reaction **WITHOUT** being used up.

- Particle theory – Catalyst lowers the activation energy for the reaction, greater CHANCE of a successful collision between the particles.

Rate graphs with catalyst



Benefits of using catalysts

- Use less energy for the reaction to take place.
- Increases the yield of the products.
- Preserves raw materials.

Examples of catalysts – Iron (Making ammonia), Vanadium Oxide (Making Sulphuric acid)