Atomic Structure Part 1

Keyword	Definition
Atom	The smallest part of an element that can exist.
Atomic Number	Number of protons in an element.
Compound	A substance with two or more elements chemically combined in fixed proportions and represented by formulae using the symbols of the atoms from which they were formed.
Electron	A negatively charged subatomic particle. Has a very small relative atomic mass. Orbits the nucleus of an atom.
Element	A substance comprising only one type of atom. Each atom has the same number of protons.
Energy level (shells)	Shells surrounding the nucleus that the electrons occupy.
Ion	An atom or group of atoms which has gained or lost electrons to become charged.
Isotope	A group of atoms with the same number of protons but different numbers of neutrons.
Mixture	Two or more elements or compounds that are not chemically combined to one another.
Neutron	A neutrally charged subatomic particle. Has a relative atomic mass of 1. In the nucleus of an atom.
Nucleus	Found in the centre of an atom. Comprises protons and neutrons (except hydrogen, which is just one proton and no neutrons).
Proton	A positively charged subatomic particle. Has a relative atomic mass of 1. In the nucleus of an atom.
Relative Atomic Mass	The number of protons and neutrons in an atom.
State symbols	Represented in brackets after a chemical formula, they show the physical state of the chemical. (s), (l), (g), (aq)

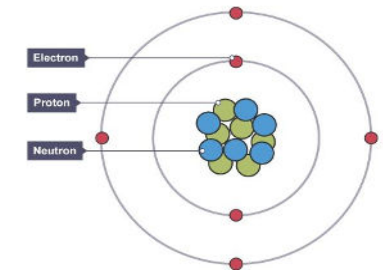
Isotopes

- Atoms of the same element must have the same number of protons.
- They may have **different numbers of neutrons.** These are known as **isotopes**.
- e.g. Hydrogen:

Isotope	Protons	Neutrons	Ar
${}^1_1H$	1	0	1
${}^2_1H$	1	1	2
${}^3_1H$	1	2	3

The Atom

- Has a central nucleus surrounded by electrons arranged in shells (or energy levels).
- They are very small, with a radius of about 0.1 nm (1×10^{-10} m).
- The nucleus comprises of protons and neutrons.
- The nucleus is less than 1/10,000 the overall size of the atom (1×10^{-14}m).



- The remaining space between the electron and nucleus is empty.

The Subatomic Particles: Protons, Neutrons and Electrons

Particle	Atomic Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	very small	-1

Calculating Numbers of Subatomic Particles

The symbol of an atom has two numbers:

- The larger number is the relative atomic mass (Ar). →
- The smaller number is the atomic (proton) number. →

23
Na
11

Number of protons = the atomic number.
Number of neutrons = Ar – the atomic number.
Number of electrons = the atomic number.

Example, sodium (Na):
Number of protons = 11
Number of neutrons = 23 – 11 = 12
Number of electrons = 11

Chemical Equations

- During a chemical reaction, atoms or ions separate from one another and re-join in different ways.
- Represented as **word** or **symbol** equations.
- **Reactants** react together. They go on the **left** of an equation.
- **Products** are made in a chemical reaction. They go on the **right** of an equation.
- Reactants ⟶ Products

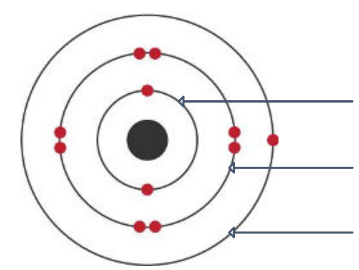
Electronic Structure

- Electrons occupy the **shells**, or **energy levels**, around the nucleus.
- Electrons occupy the lowest energy levels, nearest the nucleus, first.
- When the shell is full, electrons occupy the next shell.

Energy level (shell)	Maximum number of electrons
First	2
Second	8
Third	8

Example: Sodium

- Atomic number = 11. Therefore, 11 electrons.
- Electrons 1 and 2 go into the first energy level.
- Electrons 3, 4, 5, 6, 7, 8, 9, 10 go into the second energy level.
- The remaining electron 11 goes into the third energy level.
- Can be represented as a diagram:



Or written as 2, 8, 1.
You only need to know about the first 20 elements.

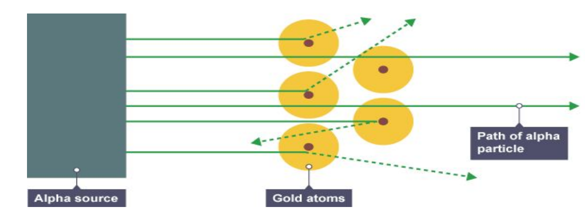
Models of the Atom

The Plum Pudding Model – J.J. Thompson

The atom was a sphere of positive charge with negative electrons embedded in it.

The Nuclear Model – Ernest Rutherford

- Carried out the **gold foil experiment** (along with Geiger and Marsden)
- Fired **positively charged** alpha particles at a thin sheet of gold foil.
- Expected all particles to go straight through.
- Most did, but some were scattered in different directions.



Rutherford concluded that:

- The mass of an atom is concentrated in the centre of the atom – the **nucleus**.
- The nucleus is **positively charged**.

Electron Shells – Niels Bohr

- Determined that **electrons are held in shells**, or energy levels. That orbit the nucleus at specific distances.
- Determined that the nucleus of positive charge could be subdivided into smaller particles called **protons**, each with the same amount of positive charge.

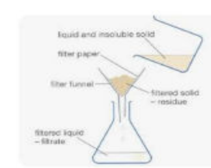
Neutrons – James Chadwick

- Discovered other particles in the nucleus with a mass but no charge.
- Named these particles **neutrons**.

Separating Mixtures

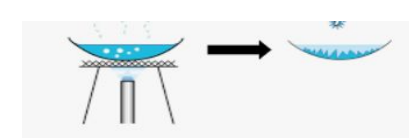
Filtration

Separates an insoluble solid from a liquid.



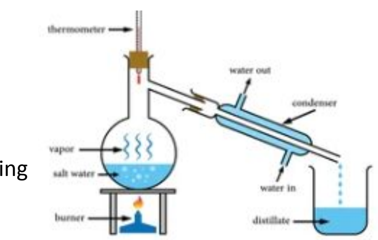
Crystallisation

Separates a soluble solid from a solution.



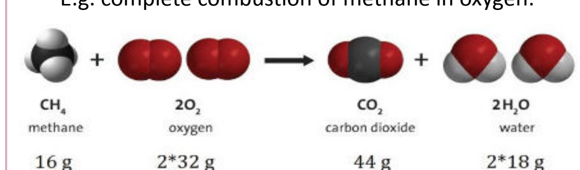
Distillation

Separates mixtures of liquids of different boiling points.



Balanced Chemical Equations

- Relies on the **Law of Conservation of Mass** – no atoms are lost or made during a chemical reaction.
- E.g. complete combustion of methane in oxygen:



- On both sides of the equation there are the same number of each atom; 1 C, 4 Os, 4 Hs.

Balancing Equations

- A balanced chemical equation has the **same number of atoms of each element on both sides** of the equation.

1. Check to see if there are an equal number of atoms of each element on both sides. E.g. $N_2 + H_2 \longrightarrow NH_3$
2. There are two nitrogen atoms on the left but only one on the right, so put a big 2 on the left of the NH_3. $N_2 + H_2 \longrightarrow 2NH_3$
3. Check again. There are two hydrogen atoms on the left but (2 × 3) = 6 on the right, so put a big 3 in front of the H_2. $N_2 + 3H_2 \longrightarrow 2NH_3$
4. Check again to see if there are equal numbers of each element on both sides. (Two nitrogen atoms and six hydrogen atoms).

Useful YouTube links:

AS | Isotopes | Electron Config. | Dev. the atom

Atomic Structure Part 2 – The Periodic Table

Keyword	Definition
Atomic Number	Number of protons in an element.
Chemical property	Describes the ability of a substance to undergo a specific chemical change
Electron	A negatively charged subatomic particle. Has a very small relative atomic mass. Orbits the nucleus of an atom.
Element	A substance comprising only one type of atom. Each atom has the same number of protons.
Energy level (shells)	Shells surrounding the nucleus that the electrons occupy.
Group	A vertical column in the periodic table.
Intermolecular forces	Forces that exist between molecules
Nucleus	Found in the centre of an atom. Comprises protons and neutrons (except hydrogen, which is just one proton and no neutrons). Has a positive charge.
Period	A row in the periodic table
Physical property	A characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include colour, density, hardness, melting and boiling points
Relative Atomic Mass	The number of protons and neutrons in an atom.

Development of the Periodic Table

- Prior to the discovery of Protons, Electrons and Neutrons, efforts were made to order the elements according to their atomic weight.
- Early attempts at a periodic table were incomplete. Many elements were unknown, meaning that elements were placed in groups with other elements that were not similar.
- Russian Scientist **Dmitri Mendeleev** published his first periodic table in 1869.
- He arranged elements in order of **increasing atomic weight**, but importantly **left gaps for elements that were yet to be discovered**.
- Elements with **similar chemical properties then were arranged in vertical groups**.
- This meant that some elements were not in order of their atomic weight. This was later explained by the discovery of **isotopes.**
- Elements with properties predicted by Mendeleev were discovered and filled the gaps.

Useful YouTube links:

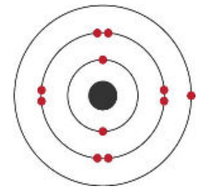
Group 1 Elements | Group 7 Elements | Noble Gases

Electron Structure and the Periodic Table

- With a bit of practice, the electron structure of an element can be quickly derived from the Periodic Table.
- The **period** of an element indicates the **number of electron shells**.
- The **group** of an element indicates the **number of electrons in the outer** shell.

*The **first period** only has two elements*
*- therefore, the first electron shell **can only accommodate two electrons**.*
*The **second and third periods** have eight elements*
*- therefore, the second and third shells **can only accommodate eight electrons.***

Example: Sodium, atomic number = 11, number of electrons = 11.
Sodium is in **period 3** (the 3rd row)
so has **three electron shells**.
The first two shells will be filled
with two and eight electrons, respectively.
Sodium is in **group 1**, so has
one electron in its outer shell.

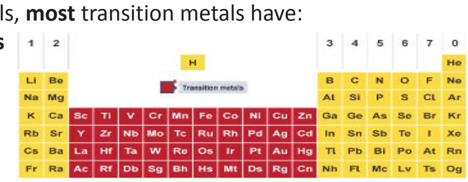


Note: You only need to know about the first twenty elements.

Transition metals – Chemistry only

You must be able to describe the difference when compared to Group 1 metals using examples of chromium, manganese, iron, cobalt, nickel, copper.
They have the same **physical properties** as all metals:
- conduct electricity in solid and liquid states.
- shiny when freshly cut.

Compared to other metals, **most** transition metals have:
- higher **melting points**
- higher **densities**
- greater strength
- greater hardness



Reaction with oxygen and water:
- Less reactive than group 1 elements, reacting slowly or not at all.

Reaction with halogens:
- Some transition metals react slowly with halogens.

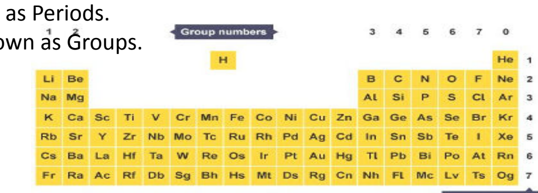
Ion Formation:
- Transition metals form different ions.
  Copper, for instance can form Cu^{+} and Cu^{2+} ions.

Compound formation:
- Most other elements form white compounds. Transition metals form coloured compounds.

Catalysts
- Many transition metals are useful as catalysts.

The modern Periodic Table

Has rows known as Periods.
Has columns known as Groups.



Elements with similar chemical properties are in the same group.
Elements placed in **order of increasing atomic number** (the number of protons in an atom of that element).

Group 1 – The Alkali Metals

Li, Na, K, Rb, Cs, Fr

Physical Properties:
- Are **soft** – get softer going down group.
- Have relatively **low melting points**.
- Have **low densities** – float on water – increases going down group.

Reactivity increases going down group 1
This is because:
- The outer shell electrons are further away from the nucleus.
- The force of attraction between the outer shell electron and nucleus decreases
- Shielding of attraction (between positive nucleus and outer shell negative electron) increases with more shells.
- The outer shell electron can be lost more easily.

Chemical Reactions
- All group 1 metals have one electron in their outer shell, so all react in similar ways.

Reactions with water:
- **Form a metal hydroxide if** reacted with water.

e.g. sodium + water → sodium hydroxide + hydrogen
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
Metal hydroxides are bases – they dissolve in water to form an alkali solution – hence the group 1 metals are called the alkali metals.

You must be able to describe the reactions of Lithium, sodium, and potassium with water:

Element	Observation when added to water
Lithium	Fizzes steadily; slowly becomes smaller until it disappears
Sodium	Fizzes rapidly; melts to form a ball; quickly becomes smaller until it disappears
Potassium	Burns violently with sparks and a lilac flame; quickly melts to form a ball; disappears rapidly, often with a small explosion

Reactions with Oxygen
- Form a **metal oxide** with reacted with oxygen in the air. e.g.
  sodium + oxygen → sodium oxide
  $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
- Alkali metals react vigorously when heated and placed in a oxygen jar.

Reactions with Chlorine
- Form a **metal chloride** when reacted with chlorine. e.g.
  sodium + chlorine → sodium chloride
  $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
- Alkali metals react vigorously with chlorine and form a white solid.

Group 7 – The Halogens

F, Cl, Br, I, At, Ts

Physical Properties:
- Exist as simple **diatomic** molecules e.g., F_2, Cl_2, Br_2. Each atom is bound by a single covalent bond to one other atom.

Melting/Boiling points increase down group:
This is because
- Atoms are larger.
- Intermolecular forces between atoms increase.
- More energy is required to overcome the stronger intermolecular forces.

Element	Colour	State
Chlorine	Pale green	Gas
Bromine	Brown	Liquid
Iodine	Purple-black	Solid

Reactivity decreases going down group 7.
This is because:
- The outer shell electrons are further away from the nucleus.
- The force of attraction between the outer shell electron and nucleus decreases.
- An extra electron, required to give the atom a full outer shell, is more difficult for the atom to gain

Chemical Reactions
- All group 7 elements have seven electrons in their outer shell, so all react in similar ways.

Reactions with metals:
React with metals to form salts.
e.g., sodium + chlorine → sodium chloride
$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Reactions with non-metals:
- Reacts with hydrogen to form a hydrogen halide.
  e.g., hydrogen + chlorine → hydrogen chloride
  $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
- Hydrogen halides are gases at room temperature but dissolve in water to form acids.

Displacement Reactions
- More reactive halogens displace less reactive halogens from their compound.

e.g., chlorine + potassium iodide → potassium chloride + iodine
$Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$

Group 0 – The Noble Gases

He, Ne, Ar, Kr, Xe, Rn, Og

- Have a full outer shell so are **unreactive**; they are inert elements.
- They exist as **single atoms**.
- Going down the group boiling point increases because:
  - Atoms are larger.
  - Intermolecular forces between atoms increase.
  - Therefore more energy is required to overcome the forces.

Bonding, Structure and the Properties of Matter

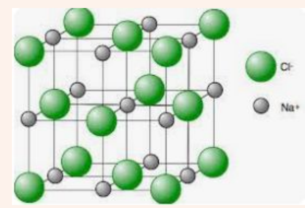
Keyword	Definition
Alloy	A mixture of two or more elements, including at least one metal
Allotropes	Different forms of the same element
Covalent Bond	A strong type of bond in which atoms share their outer shell electrons
Delocalised electrons	Electrons which are not associated to any particular atom or bond
Delocalised electrons	Electrons which are not associated to any particular atom or bond
Giant Covalent Structure	A lattice structure containing very many atoms, each bonded to other atoms by covalent bonds.
Giant Ionic Lattice	A 3D structure of strong electrostatic forces acting in all directions between oppositely charged ions
Intermolecular Force	A weak force of attraction between small molecules
Ion	An atom that has gained or lost electrons to become charged
Ionic Bond	A strong, electrostatic force of attraction between oppositely charged ions.
Metallic Bond	A lattice structure of closely packed metal ions surrounded by delocalized electrons
Nanoparticle	Structures that range from 1 to 100 nanometres in size and contain only a few hundred atoms
Polymer	A large covalent molecule formed from many smaller molecules
Small Molecule	A covalent molecule that contains only a few atoms

Ionic Substances

Bonding: generally a metal with a non-metal

- Metals lose electrons in their outer shell, they become positively charged ions
- Non-metals gain electrons in their outer shell, they become negatively charged ions
- Represented using dot and cross diagrams:

Giant ionic lattice:



- The oppositely charged ions are held together by strong electrostatic forces of attraction.
- The forces act in all directions.

Properties:

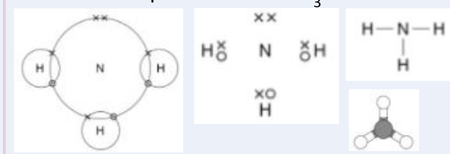
Compounds have high melting and boiling points. **Reason:** Ionic lattices have many strong bonds, so a large amount of energy is required to break them.

Do not conduct electricity unless molten or dissolved in water. **Reason:** Ions must be free to move so charge can flow

Covalent Substances

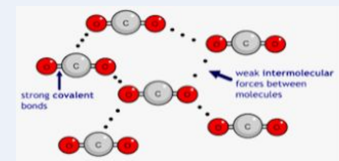
Bonding: between non-metals

- A shared pair of outer-shell electrons
- Represented using dot and cross diagrams in any of the following ways. For example ammonia NH_3

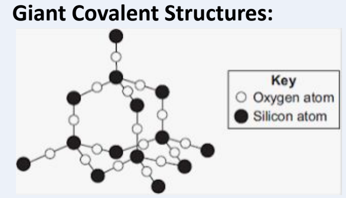


Small Molecules:

- Strong covalent bonds within molecules
- Weak intermolecular forces between the molecules



Giant Covalent Structures:

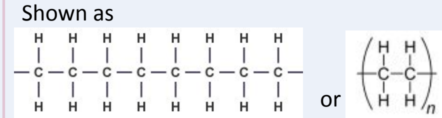


- Many atoms
- All atoms covalently bonded to others

Polymers:

- Have very large molecules with many atoms bonded in a long chain.

Shown as



Properties:

Giant covalent structures have very high melting points. **Reason:** A high amount of energy is needed to overcome so many strong covalent bonds between atoms

Small molecules have low melting/boiling points, making them gases or liquids at room temperature. **Reason:** The weak intermolecular forces are overcome when small molecules melt or boil. This requires a low amount of energy

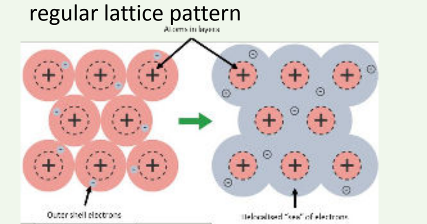
Polymers are solids at room temperature, as they have moderately high melting points **Reason:** polymers are long molecules, and the intermolecular forces are relatively strong in larger molecules

Covalent compounds do not conduct electricity. **Reason:** These compounds do not have a charge, or delocalised electrons, so charge cannot flow through them.

Metallic Substances

Bonding: metals

- Outer shell electrons of atoms are delocalised – free to move through the whole structure.
- This gives a giant structure of ions in a regular lattice pattern

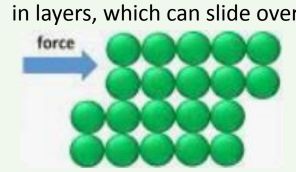


Properties

Most metals have high melting and boiling points. **Reason:** Metallic bonding is very strong

Very good electrical and thermal conductors **Reason:** Delocalised electrons carry charge and transfer thermal energy through the metal structure.

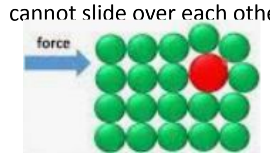
Can be bent and shaped. **Reason:** atoms are in layers, which can slide over each other



Alloys

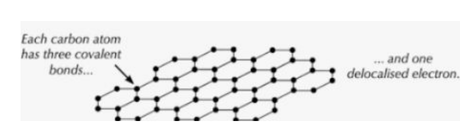
- Metals mixed with at least one other element.
- Improves properties such as hardness, strength, cost and corrosion-resistance

Alloys are often harder than pure metals **Reason:** layers in alloys are disrupted, so cannot slide over each other



Graphene

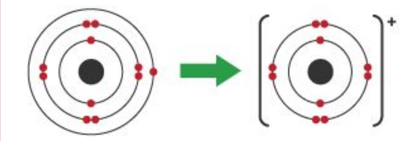
- A single sheet of graphite.
- 3 strong carbon covalent bonds per carbon atom.
- 1 delocalised electron per carbon atom.
- 1 atom thick
- Thermal and electrical conductor.



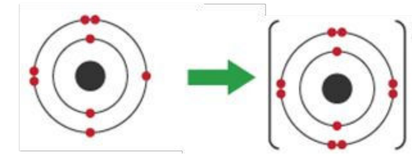
- Useful as a composite or in electronics.

Recap: Ions

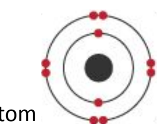
- Atoms gain or lose electrons to form ions
- Metals lose electrons:



- Non-metals gain electrons:

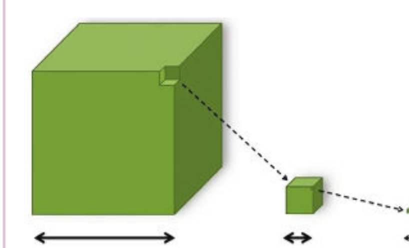


- Atoms gain or lose electrons in order to achieve a full outer shell.
- Atoms with a full outer shell have the same electron structure as a noble gas:



Nanoparticles

- Structures 1 – 100 nm in size
- Contain only a few hundred atoms
- Used in medicine, electronics, cosmetics and sun creams, as deodorants, and as catalysts
- Has a high surface area to volume ratio so only small quantities are needed.
- As the side of a cube decreases by a factor of 10; the surface area to volume ratio increases by a factor of 10 so has different properties to them in bulk.

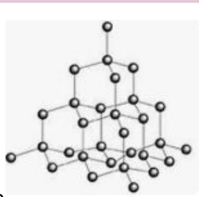


Useful YouTube links:

Bonding in Carbon

Diamond - properties

- Very hard
- High melting point
- Doesn't conduct electricity

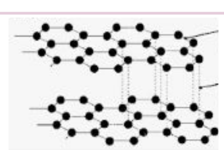


Structure

- Giant covalent structure
- 4 covalent bonds per carbon atom – so a large amount of energy is needed to break all the bonds in a diamond

Graphite - properties

- Soft and slippery
- Conducts electricity between the layers

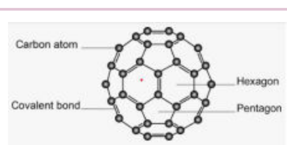


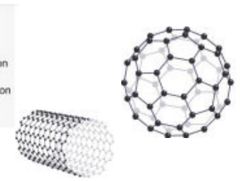
Structure

- 3 strong carbon covalent bonds
- Layers of hexagonal shapes. The layers have very weak forces of attraction between them, so can easily slide over each other
- 1 delocalised electron per atom to carry charge.

Fullerenes

- Molecules of carbon with hollow shapes
- Shapes usually made of hexagonal rings
- Buckminsterfullerene (C60) is a spherical fullerene
- Nanotubes are cylindrical fullerenes
- Properties make them useful for nanotechnology, electronics and materials





Quantitative Chemistry

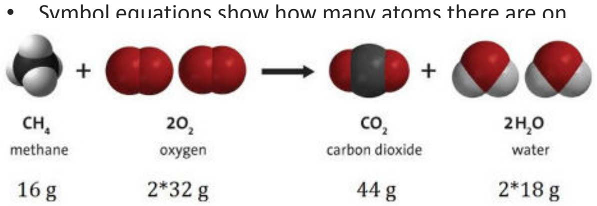
Keyword	Definition
Atom economy	A measure of the efficiency of a chemical reaction in producing a desired product.
Concentration	The amount of solute dissolved in a volume of solvent.
Conservation of mass	The total mass of the products formed in a reaction is equal to the total mass of the reactants.
Limiting reactant	The reactant that determines the amount of product that is formed.
Mole	The amount of substance in the Ar or Mr of a substance (in grams). One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance. The unit is mol.
Reactants	The elements and compounds reacting.
Products	The elements and compounds produced from chemical reactions.
Relative atomic mass	The relative atomic mass (Ar) is the average mass of atoms of an element relative to the mass of an atom of carbon-12 (which is given a mass of 12).
Relative formula mass (M_r)	The total of the relative atomic masses, added up in the ratio shown in the chemical formula of a substance.
Uncertainty	Is the degree of error in a measurement.
Yield	The amount of desired product obtained in a chemical reaction

Conservation of mass and relative formula masses

- To calculate the **relative formula mass (Mr)** of a molecule, you add up all the relative atomic masses of the atoms in the formula.

e.g., CO_2 = (1 x C) + (2 x O) = 12 + (2 x 16) = 44

- The law of **Conservation of mass** states that no atoms are lost or made during a chemical reaction, so the mass of the **products** equals the mass of the **reactants**.
- Symbol equations show how many atoms there are on

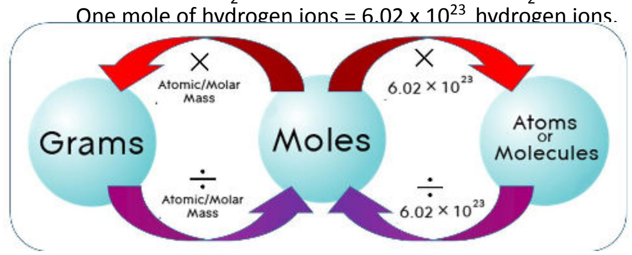


- An equation may seem to have lost or gained mass, but this will be because a reactant or product is a gas.
- An **open system** is where chemicals can enter or escape.

Moles (Higher tier)

- The mass of **one mole** of a substance in **grams** is the the **relative formula mass** of that substance, e.g., the formula mass of CO_2 is 44, so one mole of carbon dioxide is 44 g.
- The mole is a standard measurement of any chemical substance. One mole of any substance contains 6.02 x 102^3 particles, this is known as **Avogadro's constant**.

One mole of silver = 6.02 x 10^{23} atoms of silver
One mole of CO_2 = 6.02 x 10^{23} molecules of CO_2
One mole of hydrogen ions = 6.02 x 10^{23} hydrogen ions.



Concentrations

- The concentration is the number of solute molecules in a certain volume. A concentrated solution has a lot of solute dissolved in a volume of solvent. Weak solutions have little solute dissolved.
- The concentration of a solution can be shown in **g/dm^3** or **mol/dm^3**.
- The concentration of a solution can be calculated using either the amount of dissolved solute in moles (mol) or the volume of solution in cubic decimetres (dm^3).
- 1 dm^3 = 1000 cm^3 e.g., 250 cm^3 is 0.25 dm^3.
- **Molarity (M)** is the concentration usually expressed in mol/dm^3.
- To find out an unknown concentration you need to divide the number of moles by the volume of solution.

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

Limiting reactants (Higher)

- A reaction finishes when one of the reactants is all used up. The other reactant has nothing left to react with, so some of it is left over
- Use a table to calculate the limiting reactant.

e.g., 90g of sodium and 90g of chlorine are reacted together, Which is the limiting reactant, and with the amounts used what is the maximum mass of sodium chloride that can be produced?

	2Na	+ Cl_2	→ 2 NaCl
Mass(ive)	90 g	90 g	2.54 x 58.5 = 148.5 g
M_r	23	2 x 35.5 = 71	58.5
Mole	90 / 23 = 3.91	90 / 71 = 1.27 limiting	1.27 x 2 = 2.54
Ratio	2	1	2

- The moles of reactant need to be in the ratio of the balanced equation.

Percentage Yield (Chemistry only]

The **theoretical yield** is the maximum possible mass of a product that can be made in a chemical reaction.

Percentage Yield = $\frac{\text{Mass of product actually made}}{\text{Maximum theoretical mass of product}}$ x100

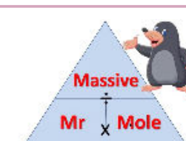
The industrial aim of any company is to have a reaction with as high a percentage yield as possible to maximise profits and minimise costs and waste

Reasons why the product might be less than 100%:

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

Gas equations (Chemistry only)

- At a given **temperature** and **pressure**, one mole of any gas occupies the **same volume**:
- The **molar volume** is the volume occupied by one mole of any gas. At room temperature and pressure this occupies a volume of **24 dm^3**.



GCSE Chemistry Equation Sheet

These are the equations you will need to learn for Chemistry GCSE exam papers

All:

Relative formula mass = Relative atomic mass of all atoms in compound added together

Concentration (g/dm^3) = $\frac{\text{mass (g)}}{\text{volume } (dm^3)}$

Higher tier (all)

Moles = $\frac{\text{Mass (g)}}{\text{Relative formula mass}}$

Number of particles in a substance = Moles x Avogadro's number

Chemistry

Percentage yield = $\frac{\text{Mass of product actually made}}{\text{Maximum theoretical mass of product}}$ X 100

Atom economy = $\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula masses of all reactants from equation}}$ X 100

Chemistry Higher

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

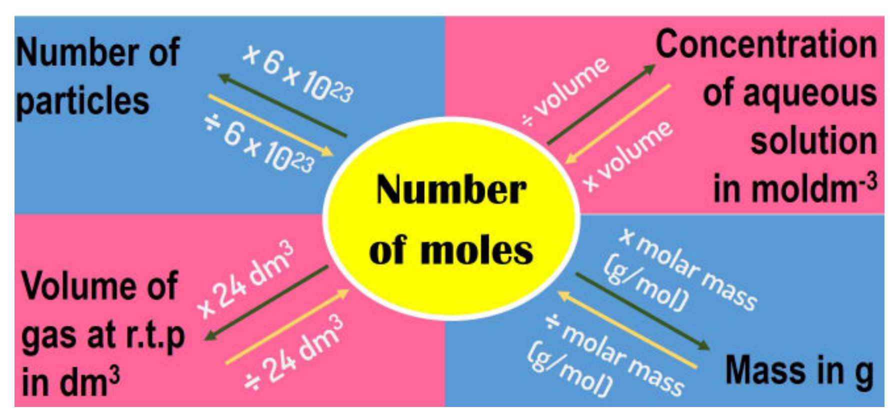
Volume of gas at room temperature and pressure (dm^3) = moles x 24

Atom economy (Chemistry only)

- In some chemical reactions some atoms in the reactants may not end up in the desired product. They instead form other products, and so are regarded as by-products.
- The atom economy of a reaction is a measure of the amount of starting materials that end up as the **desired products**.

Atom economy = $\frac{\text{Total } M_r \text{ of the desired product}}{\text{Total Mr of all reactants}}$ x100

- It is important for sustainable development and for economic reasons to use reactions with high atom economy.



Useful YouTube links:

Quantitative chemistry revision

Chemical Changes Reactivity

Keyword	Definition
Bioleaching	Bioleaching uses bacteria to produce leachate solutions that contain metal compounds
Corrosion	The name of the process when a metal reacts with oxygen to form a metal oxide
Ion	A charged atom, i.e., positive or negative charge.
Ore	A rock containing minerals (metal compounds), e.g., Fe_2O_3
Oxidation	Either the gain of oxygen or the transfer (loss) of electrons.
Phytomining	Phytomining uses plants to absorb metal compounds. The plants are harvested and then burned to produce ash that contains metal compounds
Reactivity	How reactive a metal is. The reactivity of metals is arranged into a list called the Reactivity Series.
Reduction	Either the loss of oxygen or the transfer (gain) of electrons.
Rusting	If the metal is iron, corrosion is called rusting. Therefore, rusting is only linked to iron.
Sacrificial protection	Iron can be protected from rusting if it is in contact with a more reactive metal, such as zinc. The more reactive metal oxidises more readily than iron, so it 'sacrifices' itself while the iron does not rust.

Oxidation/reduction in terms of electrons (HT)

Oxidation is the loss of electrons; reduction is the gain of electrons in a chemical reaction.
The mnemonic OILRIG helps us remember this.

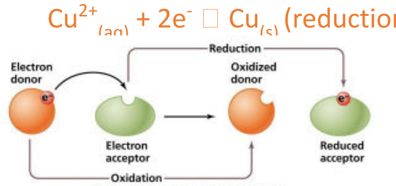
Redox reactions are matched sets: if one species is oxidised in a reaction, another must be reduced. Hence the term redox.
For example:

$$Ca_{(s)} + Cu^{2+}_{(aq)} \rightarrow Ca^{2+}_{(aq)} + Cu_{(s)}$$

This is an example of a balanced ionic equation. It can be split into two half equations:

$$Ca_{(s)} \rightarrow Ca^{2+}_{(aq)} + 2e^- \text{ (oxidation)}$$

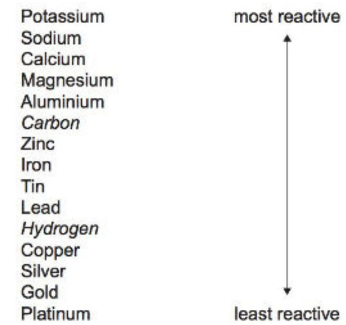
$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \text{ (reduction)}$$



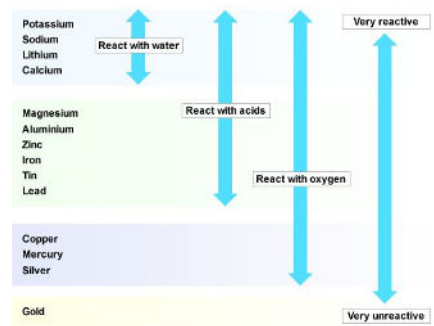
In our example calcium donates two electrons to the copper ion.

The reactivity series:

Oxidation and reduction in terms of oxygen.
Metals react with oxygen to produce metal oxides.
Magnesium + oxygen → magnesium oxide
This is an oxidation reactions because the metals gain oxygen. When metals lose oxygen, it is called a reduction reaction
When metals react with other substances the metal atoms form positive ions. The reactivity of a metal is related to its tendency to form positive ions. Metals can be arranged in order of their reactivity in a reactivity series.

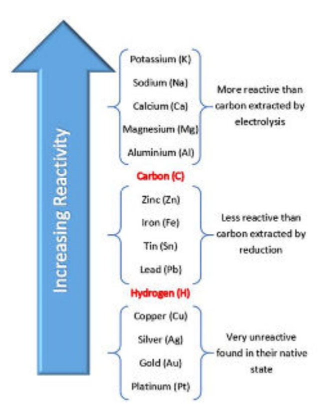


The metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids.



The non-metals hydrogen and carbon are often included in the reactivity series (these will help us with working out reactions and products in other reactions)

Carbon in the reactivity series



If a metal is less reactive than carbon, it can be extracted from its oxide by heating with carbon. The carbon displaces the metal from the compound and removes the oxygen from the oxide.

For example:
Copper oxide + carbon → copper + carbon dioxide

$$2CuO_{(s)} + C_{(s)} \rightarrow 2Cu_{(s)} + CO_{2(g)}$$

This is called a **reduction** reaction because the oxygen has been removed from the copper.

Metal ores:

Ores are minerals found in the Earth's crust. They contain metal compounds, from which metals can be extracted. Gold and silver are examples of metals that are found native. The most reactive metals are extracted by electrolysis, while those towards the middle of the reactivity series can be chemically reduced. In a displacement reaction.

Displacement reactions:

The relative reactivity of metals can be shown by displacement reactions. A more reactive metal will displace a less reactive metal from its compound.

$$A + BC \rightarrow B + AC$$

Calcium + copper sulfate → copper + calcium sulfate

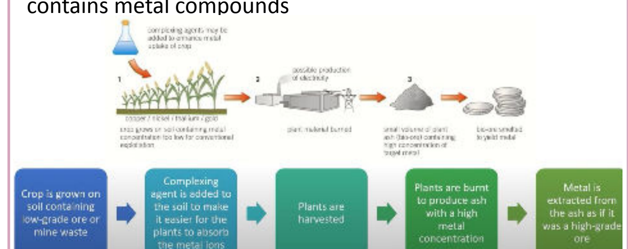
$$Ca_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + CaSO_{4(s)}$$

The calcium is more reactive than the copper so can displace it from its compound.

During a competition reaction, a more reactive metal will remove oxygen from the oxide of a less reactive metal when a mixture of the two is heated.
An example of this is the thermit reaction used in the rail industry to weld rails together on a track.
iron(III) oxide + aluminium → iron + aluminium oxide

Alternative methods of extracting metals (HT) (from Using Resources)

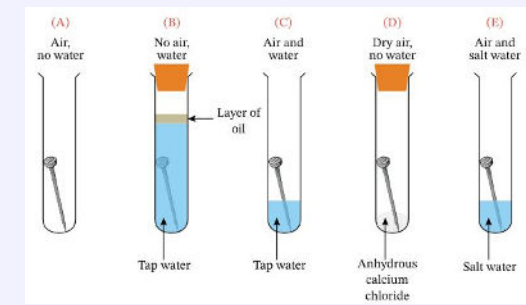
The Earths resources are limited, and sometimes a deposit of a metal ore is not economically viable to extract it by the usual quarrying methods. Alternative methods of extraction include:
Phytomining
Phytomining uses plants to absorb metal compounds. The plants are harvested and then burned to produce ash that contains metal compounds



Bioleaching
Bioleaching uses bacteria to produce leachate solutions that contain metal compounds. Low concentrations of metal ore are not a problem for bacteria because they simply ignore the waste which surrounds the metal ores, attaining extraction yields of over 90% in some cases. The bacteria gain energy by breaking down the metal ores into the leachate solution containing metal compounds.
The metal compounds can be processed to obtain the metal. For example, copper can be obtained from solutions of copper compounds by displacement using scrap iron or by electrolysis.

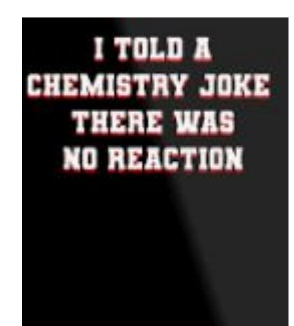
Rusting and Corrosion (Chemistry only) (from Using Resources)

Corrosion is the destruction of materials by chemical reactions with substances in the environment.
Corrosion is the name given to the process when a metal reacts with oxygen in the air to form a metal oxide.
Corrosion is the reverse of the process used to extract the metal.
If the metal is iron, corrosion is called rusting. The iron reacts with both oxygen and water to form iron oxide (rust).
In general, the more reactive a metal is the more quickly it corrodes,
Corrosion can be prevented by applying a coating such as grease, painting or electroplating to prevent the oxygen and water from reaching the metal.



The iron nail in test tube E will begin to rust first, followed by the iron nails in test tubes C, B, and A. The iron nail in test tube D should be the last to start rusting (if at all).
Aluminium
Aluminiums surface layer reacts with oxygen and thus forms an oxide coating that protects the metal from further corrosion (unless this layer becomes damaged).
Sacrificial Protection
Some coatings are reactive and contain a more reactive metal to provide a sacrificial protection.
For example:
Ships hulls are usually made from mild steel (an alloy with small quantities of carbon and manganese to allow it to be rolled during processing and not crack). However, iron rusts quicker in the salt water of the seas Therefore zinc blocks are attached to the hull of the ship. As these are more reactive than the steel hull, they will react to prevent the hull from rusting

Zinc blocks



Useful YouTube links:

Metals

Chemical Changes Electrolysis

Keyword	Definition
Anode	The positively charged electrode.
Aqueous	Dissolved in water.
Cathode	The negatively charged electrode.
Cryolite	A solvent for aluminum oxide that lowers the overall melting point of the electrolyte.
Discharged	Ions gaining or losing electrons at the electrodes (oxidised or reduced).
Electrode	An electrical conductor that carries an electric current into an electrolyte.
Electrolyte	A substance that conducts electricity when molten or in an aqueous solution.
Half equation [HIGER]	Either of two equations that describe each half of a REDOX reaction.
Inert electrode	Electrodes that are not degraded during electrolysis.
OILRIG	Oxidation is loss of electrons; reduction is gain of electrons.
Spectator ion	An ion that is not oxidised or reduced at the electrodes (does not participate in the reaction).
Voltage	Voltage. is a measure of the difference in electrical energy between two parts of a circuit.

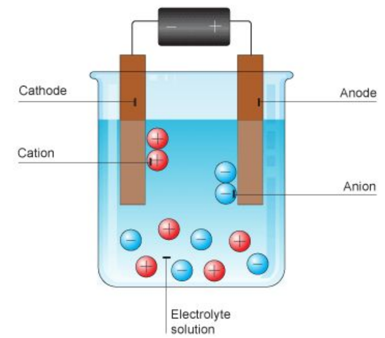
Cells and Batteries (Chemistry only)
(from Using Resources)

A **cell** is the proper word for what we regularly call a battery. Cells contain **chemicals** in the form of a **solid metal** (**electrode**) and an **ionic solution** (the **electrolyte**), which **react** and **generate electricity** by releasing electrons.
The **voltage** produced by a cell depends on the **type of electrode** and **electrolyte** used. You can make a simple cell by **connecting two different metals in contact with an electrolyte solution**.

A **battery** is **two cells joined together** in series to **increase the voltage**.
Batteries can be **rechargeable** or **non-rechargeable**.
In non-rechargeable cells or batteries (such as alkaline batteries), the **reactants** will eventually be **used up** and the reaction will stop.
With rechargeable cells and batteries, an **external electric current reverses the reaction** once the reactants have been used up. This ensures that the **reaction can keep going** and the battery can continue to supply electricity

Introduction to electrolysis

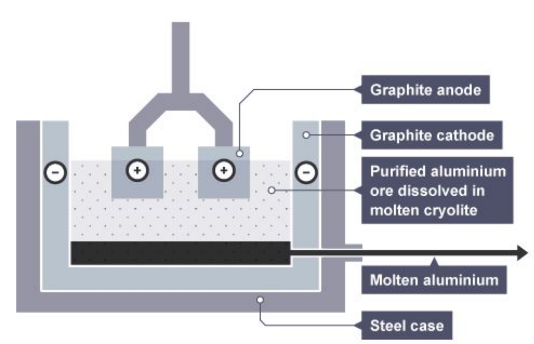
- When an **ionic** substance is **melted** or **dissolved**, the ions are free to move about within the liquid or solution.
- These liquids and solutions can conduct electricity and are called **electrolytes**.



- The **electrodes** induce a current in the electrolyte. The **anode** is positively charged, and the **cathode** is negatively charged.
- The ions move towards the electrodes: negative ions are attracted to the positive anode and and positive ions are attracted to the negative cathode.

Extraction of Aluminium

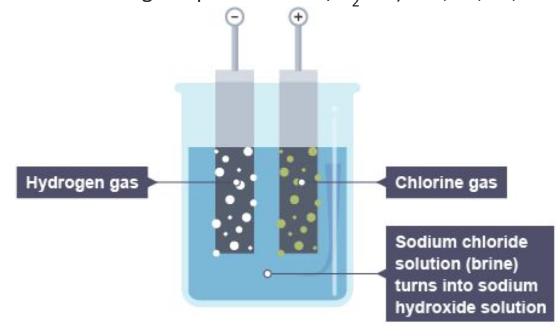
- Aluminium is manufactured by the electrolysis of a molten mixture of **aluminium oxide** and **cryolite**.
- Cryolite is added to lower the melting point and, therefore, energy cost.
- Aluminium forms at the negative electrode (cathode) and oxygen at the positive electrode (anode).
- The positive electrode (anode) is made of **carbon** (graphite), which reacts with the oxygen to produce carbon dioxide and so must be continually replaced.



- The balanced half equation is: $Al^{3+} + 3e^{-}$ □ Al (because three negatively charged electrons are needed to balance the three positive charges on the aluminium ion).
- Oxide ions lose electrons at the positive electrodes and are oxidised to oxygen gas:
  $2O^{2-}$ □ $O_2 + 4e^{-}$ (oxidation – lose electrons)
- The oxygen gas reacts with the carbon anode to form CO_2 gas:
  $O_{2(g)} + C_{(s)}$ □ $CO_{2(g)}$

Electrolysis of a solution ($NaCl_{(aq)}$)

If an aqueous solution of sodium chloride is electrolysed, two products are formed at the electrodes.
Sodium chloride dissolved in water leads to 4 ions being available during the process. $NaCl/H_2O$ --) Na^+, Cl^-, H^+, OH^-



1. The **Chloride (Cl^-)** ions are attracted to the anode (positive electrode) where they **lose an electron** (oxidised) and pair with another chlorine atom to form **chlorine gas (Cl_2)**.
2. The **Hydrogen (H^+)** ions are attracted to the cathode (negative electrode) where they each **gain an electron** (reduced) and pair with another hydrogen atom to form **hydrogen gas (H_2)**.
3. Sodium (Na^+)ions react with **hydroxide (OH^-)** ions to form an alkaline solution of sodium hydroxide.

Important Rules:

At the cathode: If an element is less reactive than hydrogen, then it will be produced at the cathode. If it is more reactive than hydrogen, then hydrogen gas will be formed at the cathode. e.g., sodium is more reactive than hydrogen, so is not produced at the cathode.

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

At the anode: If the negative ion from the ionic compound is a halide (e.g., Cl^- or Br^-), then that element is produced. If the negative ion is a complex ion (e.g., NO_3^-, SO_4^{2-}, CO_3^{2-}), then oxygen is produced from the hydroxide ion present instead.

Half equations (HT)

- Half equations represent the **REDOX** reactions happening at the electrodes.

Using the example on the left:
At the negative electrode (cathode)
$2H^+_{(aq)} + 2e^-$ □ $H_{2(g)}$

At the positive electrode (anode)
$2Cl^-_{(aq)} + 2e^-$ □ $Cl_{2(g)}$

Using a second example
The electrolysis of sodium nitrate produces hydrogen and oxygen gas (see 'important rules'). The nitrate ion (NO_3^-) is a **spectator ion** because it is not involved in the reaction.

At the negative electrode (cathode)
$2H^+_{(aq)} + 2e^-$ □ $H_{2(g)}$ (hydrogen ion is reduced)

At the positive electrode (anode)

$4OH^-_{(aq)}$ □ $2H_2O_{(l)} + O_{2(g)} + 4e^-$ (hydroxide ion is oxidised)

Remember to use OILRIG

Fuel Cells (Chemistry only)
(from Using Resources)

Fuel cells differ from chemical cells in that they generate a **continuous voltage**, providing they have a **fuel supply** (such as **hydrogen**) and **oxygen**. Within the fuel cell, the **fuel is oxidised electrochemically** rather than being burned. This means that the reaction takes place at a **lower temperature** than a typical combustion reaction. Energy is released as **electrical energy**. The **overall reaction** in a hydrogen fuel cell involves the **oxidation of hydrogen** to produce **water**.

Hydrogen + oxygen □ water
$H_2 + 0_2$ □ H_20

Half equations for this reaction (HT Chemistry only)

At the negative electrode $2H_2 + 4OH^-$ □ $4H_20 + 4e-$
At the positive electrode $O_2 + 2H_20 + 4e-$ □ $4OH^-$

Fuel cells vs Cells and Batteries (Chemistry only)

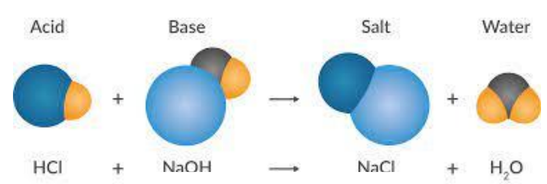
Hydrogen fuel cells are a potential **alternative** to **rechargeable cells and batteries**.
The advantages of hydrogen fuel cells are that they only produce **water** as a product (so **no toxic products** formed). They are also **small** and **easy to maintain**. However, they are **very expensive** to make and require a **constant supply of hydrogen gas** which is **flammable** and **difficult to store**.
Non-rechargeable alkaline batteries are **cheaper** to make but they may **accumulate in landfill sites** once they have run out. These batteries are **recyclable,** but the recycling process is **expensive**.

Chemical Changes
Acids, Alkalis and Bases

Keyword	Definition
Acidity	The concentration of H^+ ions in a solution.
Alkali	Is a soluble base
Alkalinity	The concentration of OH^- ions in a solution.
Aqueous (aq)	Dissolved in water.
Concentration	The number of acid or alkali molecules dissolved in a given volume of water.
Disassociate	The stronger an acid is, the more easily it loses a proton, H^+. Losing a proton means dissociation.
Ion	A charged atom, i.e., positive or negative charge.
Ionisation	How completely an acid ionises in solution, i.e., concentration of H^+ ions
Neutralisation	When hydrogen ions react with hydroxide ions to form water.
Proton donor	A hydrogen ion is also just a proton. Therefore, an acid is a proton donor.
Salt	An ionic compound formed during neutralisation reactions.
Universal indicator	A universal indicator is **a blend of pH indicator solutions designed to identify the pH of a solution** over a wide range of values.

Neutralisation

During a **neutralisation** reaction, hydrogen ions react with hydroxide ions to form water.



$H^+ + OH^-$ □ H_2O

Naming Salts

These depend on the metal and acid being used

Sulfuric acid forms **sulfate** salts, e.g., copper sulfate:

$CuO_{(s)} + H_2SO_{4\,(aq)}$ □ $CuSO_{4\,(aq)} + H_2O_{(l)}$

Nitric acid forms **nitrate** salts, e.g., sodium nitrate:

$Na_2CO_{3\,(s)} + 2HCl_{(aq)}$ □ $2NaCl_{(aq)} + H_2O_{(l)} + CO_{2\,(g)}$

Hydrochloric acid forms **chloride** salts,

e.g., magnesium chloride:

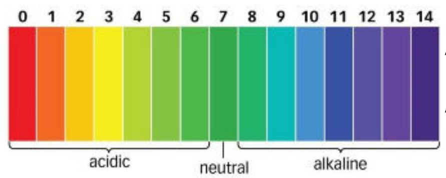
$Mg_{(s)} + 2HCl_{(aq)}$ □ $MgCl_{2(aq)} + H_{2\,(g)}$

Indicators and the pH Scale:

Indicators are substances that change colour when they are added to acids and alkalis.

Litmus is the most well-known indicator. It turns red in acid and blue in alkalis.

Universal indicator is the most used indicator in the laboratory. When added to a solution, it changes to a colour that shows the pH of the solution.



Acidic = pH < 7
Neutral = pH 7
Alkaline = pH > 7

Acids, Bases and Alkalis:

Acids and alkalis are commonly used both in industry and at home.

Acids:
Acids produce hydrogen ions, H^+, when they dissolve in water, e.g., for hydrochloric acid: HCl (aq) ⟶ H^+ (aq) + Cl^- (aq)

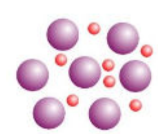
Bases:
A base is chemically opposite to an acid. A base that dissolves in water is called an alkali.

Alkalis:
Alkalis produce hydroxide ions, OH^-, when they dissolve in water,
e.g., for sodium hydroxide: NaOH(aq) ⟶ Na^+ (aq) + OH^- (aq)

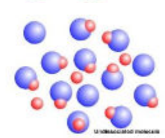
The pH scale (HT only)

The pH scale is a **logarithmic scale**. This means that every decreasing pH number has 10x more hydrogen ions than the last.,
e.g., pH 1 has 10x more hydrogen ions than pH 2, 100x more than pH 3 and 1000x more than pH4 etc.

Strong and weak acids (HT only)

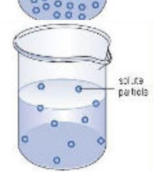


Strong acids completely ionise in solution, i.e., the hydrogen ions completely dissociate.



Weak acids do not completely ionise in solution, i.e., the hydrogen ion does not completely dissociate.

Concentrated acids have many acid molecules dissolved in a volume of water.

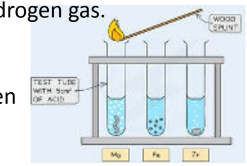


Dilute acids have few acid molecules dissolved in a volume of water.

The strength of an acid is a measure of the degree of its dissociation.
The concentration of an acid is a measure of the number of moles of acid in 1 dm3 of solution.

Acids and Metals:

Acids will react with metals to make a salt and hydrogen gas.



acid + metal ⟶ salt + hydrogen
hydrochloric acid + zinc ⟶ zinc chloride + hydrogen
$2HCl + Zn \longrightarrow ZnCl_2 + H_2$

The hydrogen gas causes bubbling during the reaction.
It can be tested with a lit splint, which will give a squeaky 'pop'
The reaction is exothermic.
The more reactive the metal, the faster the reaction is, resulting in more bubbling and a bigger temperature rise.

Acids and Alkalis/Bases:

Acids react with alkalis and bases to make a salt and water.

Alkalis are commonly metal hydroxides.

acid + alkali ⟶ salt + water
nitric acid + sodium hydroxide ⟶ sodium nitrate + water
$HNO_3 + NaOH \longrightarrow NaNO_3 + H_2O$

acid + alkali ⟶ salt + water
sulfuric acid + sodium hydroxide ⟶ sodium sulfate + water
$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$

acid + alkali ⟶ salt + water
hydrochloric acid + sodium hydroxide ⟶ sodium chloride + water
$HCl + NaOH \longrightarrow NaCl + H_2O$

Bases are commonly metal oxides.

acid + base ⟶ salt + water
nitric acid + zinc oxide ⟶ zinc nitrate + water
$2HNO_3 + ZnO \longrightarrow Zn(NO_3)_2 + H_2O$

acid + base ⟶ salt + water
sulfuric acid + zinc oxide ⟶ zinc sulfate + water
$H_2SO_4 + ZnO \longrightarrow ZnSO_4 + H_2O$

acid + base ⟶ salt + water
hydrochloric acid + zinc oxide ⟶ zinc chloride + water
$2HCl + ZnO \longrightarrow ZnCl_2 + H_2O$

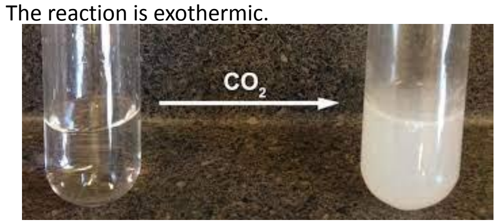
The reactions of acids with alkalis and bases are exothermic.

Acids and Carbonates:

Acids will react with carbonates to make a salt, water and carbon dioxide gas.

acid + carbonate ⟶ salt + water + carbon dioxide
nitric acid + sodium carbonate⟶ sodium nitrate + water + carbon dioxide
$2HNO_3 + Na_2CO_3 \longrightarrow 2NaNO_3 + H_2O + CO_2$

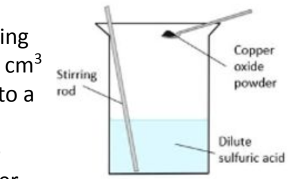
The carbon dioxide causes bubbling during the reaction.
The gas can be tested by bubbling through limewater, if there is a milky white precipitate, carbon dioxide is present.
The reaction is exothermic.



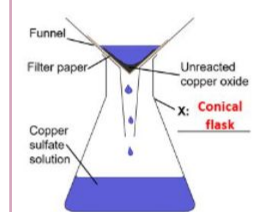
Making Salts
Required practical

Using an insoluble metal oxide base

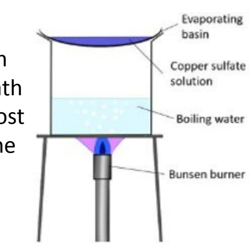
Method



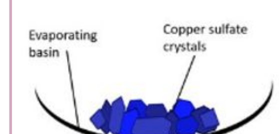
Using a measuring cylinder add 40 cm^3 of sulfuric acid to a beaker.
Gently heat the beaker in a water bath for a couple of minutes
Add a spatula of copper oxide powder to the beaker and stir the solution with a glass rod
Keep adding more copper oxide powder until it no longer disappears (add excess)



Filter the mixture to remove the excess copper oxide, then pour the filtrate (the copper sulfate into an evaporating basin



Place the evaporating basin above a water bath and heat until most of the water in the solution has evaporated.



Leave the evaporating basin in a warm place until the remaining water evaporates, copper sulfate crystals will remain.

This method can be used for other insoluble substances, including metals, hydroxides and carbonates.

Useful YouTube links:

Acid reactions
Making salts

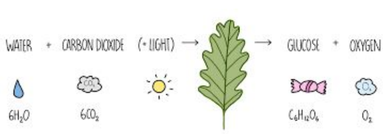
Energy Changes

Keyword	Definition
Activation energy	The minimum amount of energy required to cause atoms or molecules to undergo a chemical change.
Bond energy [HIGHER]	The energy absorbed to break chemical bonds or released from making new bonds.
Conservation of energy	The total energy of a system must remain constant (regardless if it is exothermic or endothermic).
Endothermic	When more energy is absorbed from the surroundings than is transferred to it from a chemical reaction.
Exothermic	When energy is transferred to the surroundings from a chemical reaction.
Products	The elements and compounds produced from chemical reactions.
Reactants	The elements and compounds reacting.
Reaction profile	A diagram showing the change in chemical potential energy for a single chemical reaction.
Reversible reaction	A chemical reaction where the reactants form products that can return back to reactants again.
Thermal decomposition	The breakdown of a compound (decomposition) as a result of heating it.

Cellular respiration

Cellular respiration is an exothermic reaction which is continuously occurring in living cells. The energy transferred supplies all the energy needed for living processes. Respiration in cells can take place aerobically (using oxygen) or anaerobically (without oxygen), to transfer energy.

Photosynthesis



Photosynthesis is an endothermic reaction as it requires light energy to react carbon dioxide and water to produce glucose and oxygen. The light energy required is absorbed by a green pigment called chlorophyll in the leaves. Chlorophyll is located in chloroplasts in plant cells .

Exothermic and endothermic examples

A **hand warmer** works as result of an exothermic reaction. A supersaturated solution of sodium acetate transfers energy to the surroundings as it crystalises. If the hand warmer is heated in boiling water, the reaction reverses back to produce a supersaturated solution again.

Sports injury 'cool packs' work when **citric acid** is reacted with **sodium hydrogencarbonate**. The reaction is so endothermic that it becomes cold (like iced water).

Thermal decomposition.

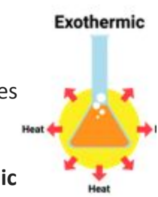
Thermal decomposition reactions are **endothermic** because they are absorbing more energy than is being released from the formation of new products. For example:

copper carbonate □ copper oxide + carbon dioxide

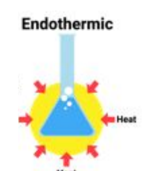
Exothermic and Endothermic reactions

The law of **Conservation of Energy** states that the amount of energy at the end of a chemical reaction is the same as before the reaction takes place.

If a reaction transfers energy to the surroundings, the product molecules must have less energy than the reactant molecules Energy being transferred into the surroundings increases the temperature of the surroundings and is called an **exothermic** reaction.



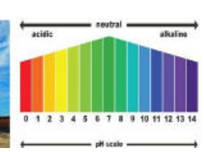
An **endothermic** reaction is one that takes in energy from the surroundings, so the temperature of the surroundings decrease. This means that the product molecules have more energy than the reactant molecules and this energy was absorbed from the surroundings.



Examples of exothermic reactions

Combustion

Oxidation



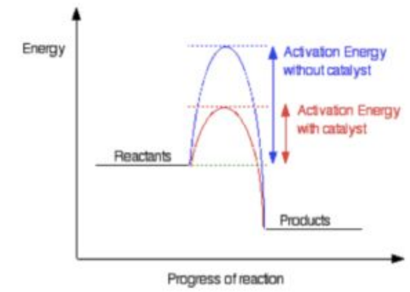
Neutralisation

Example of an endothermic reaction

There is a rapid reaction between citric acid and sodium hydrogencarbonate **giving off lots of carbon dioxide**, but the temperature of the mixture falls.

Reaction profiles

- Chemical reactions can occur only when reacting particles **collide** with sufficient energy.
- The minimum amount of energy that particles must have to react is called the **activation energy**.
- A **reaction profile is** used to show the relative energies of the reactants and products, the activation energy and the overall energy change of a reaction.

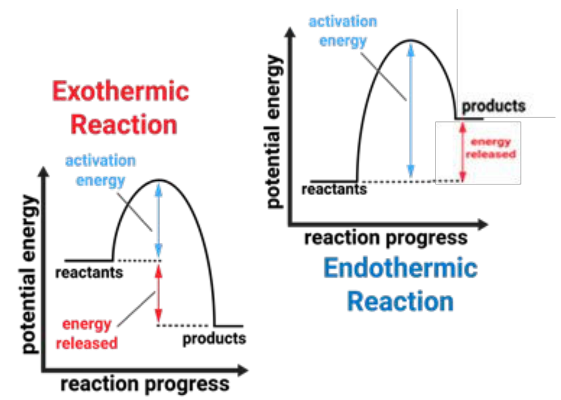


If a **catalyst** is used in a reaction the **activation energy** and **pathway** of the reaction is altered. Catalysts can:

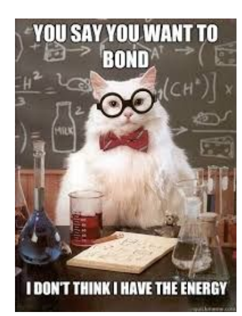
1. **Lower** the activation energy of a reaction.
2. Change the pathway of the reaction so that it proceeds at a **faster rate**.

Exothermic and endothermic reaction profiles

Endothermic reaction profiles show that a greater proportion of energy is absorbed by the reactants.



Exothermic reaction profiles show that a greater proportion of energy is released to the surroundings from the formation of new products .



Useful YouTube links:

Endo/Exo

Bond energies (HT)

Bond energies [HIGHER].

The energy change in a reaction can be calculated using **bond energies**. A bond energy is the amount of **energy** needed to break one **mole** of a particular **covalent bond**.

Different bonds have different bond energy values. NB. These are given to you in the exam when they are needed for a calculation.

Calculating the energy change in a reaction:

Energy is needed to **break bonds** which is absorbed from the reaction surroundings, so bond breaking is an **endothermic** process.
The opposite occurs for **forming bonds** as it releases energy back to the surroundings in an **exothermic** process.

Example of calculating the changes

$H_2 + Cl_2 \longrightarrow 2HCl$

H-H + H-H □ H-Cl + H-Cl

Bond	Bond energy (kJ/mole)
H-H	436
Cl-Cl	243
H-Cl	432

1. add together all the bond energies for the bonds in the **reactants**;
2. add together all the bond energies for the bonds in the **products;**
3. **energy change = energy in – energy out**.

A negative overall energy change means the reaction is exothermic
A positive overall energy change means the reaction is endothermic

Step	Calculation
Calculate the energy in (breaking bonds)	436 + 243 = 679 kJ/mole
Calculate the energy out (making bonds)	432 + 432 = 864 kJ/mole
Calculate the Overall result energy change	679 – 864 = - 185 kJ/mole

Energy change is negative. Energy is lost too the surroundings. The reaction is exothermic

Chemistry Particles – Year 9

Keyword	Definition
Activation Energy	The minimum energy that colliding particles must have for a reaction to occur
Catalyst	A substance that speeds up a reaction without changing the products or being used up
Closed System	A reaction in which nothing is allowed to enter or leave the reaction container.
Collision Theory	The theory that chemical reactions require reactant particles to collide with enough energy to react
Concentration	A measure of how much of a substance is dissolved in a solution.
Rate of Reaction	A measure of how quickly a reactant is used up, or a product is formed
Surface Area to volume ratio	A measure of the exposed area of a substance compared to the volume of the substance
Turbidity	How cloudy a liquid appears due to tiny suspended particles
Volume	The amount of 3D space an object takes up

Rate of reaction

The rate of a chemical reaction is either:

- How quickly a reactant is used up
- How quickly a product is formed

It can be calculated using the following equations:

$$mean\ rate\ of\ reaction = quantity\ of\ reactant\ used \div time\ taken$$

…or…

$$mean\ rate\ of\ reaction = quantity\ of\ product\ formed \div time\ taken$$

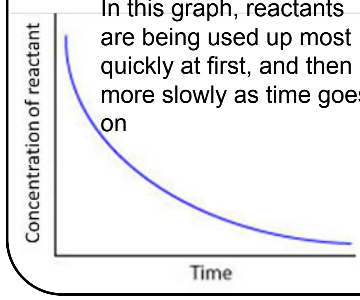
The quantity of reactants or products can be measured using grams, in cm^3, or in moles.

Therefore, rate can be measured in g/s, cm^3 /s or mol/s

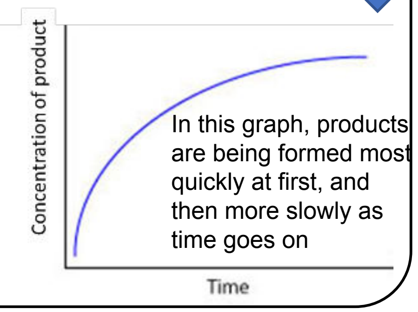
Rate Graphs

Graphs can be used to show how rate changes over time – the steeper the graph, the higher the rate.

The gradient of the graph at any point is equal to the rate of reaction at that point. The steepness of the graph also links to the rate of the reaction. Steep = fast.



In this graph, reactants are being used up most quickly at first, and then more slowly as time goes on



In this graph, products are being formed most quickly at first, and then more slowly as time goes on

Required Practical:

Observing changes in appearance
Measuring volume of gas produced

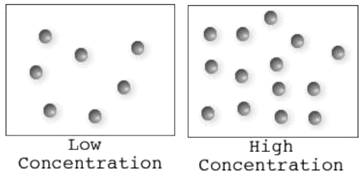
Mathematical Skills

Be prepared to:

- *Recognise and use expressions in decimal form*
- *Use ratios, fractions and percentages*
- *Translate information between graphs and numbers*
- *Draw and interpret graphs to determine rate of reaction*
- *Plot graphs using provided data*
- *Determine the gradient and intercept of linear graphs*
- *Calculate areas of cubes and rectangles*
- *Calculate surface areas and volumes of cubes*

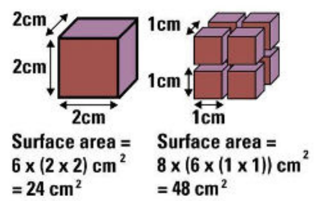
Concentration/Pressure

- Increasing the concentration of a reactant solution or the pressure of a reacting gas increases the rate of a reaction.
- This is because reactant particles are closer together, and so collide with each other more often.



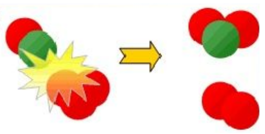
Surface Area

- Breaking lumps of a solid reactant into smaller pieces exposes more of the reactant – the overall volume stays the same.
- This increases the surface area-to-volume ratio
- The 2 diagrams below show the same volume but with a different surface area



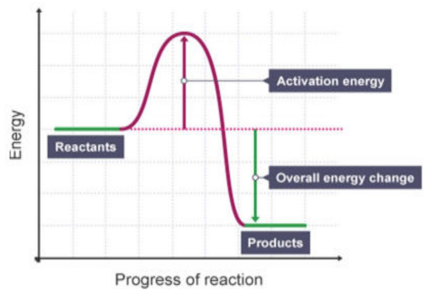
Factors affecting reaction rate

To increase or decrease the rate of a reaction, we consider collision theory. This states that for reactions to occur, the reacting particles must collide with enough energy for a reaction to take place. The more collisions there are with sufficient energy, the faster the reaction.



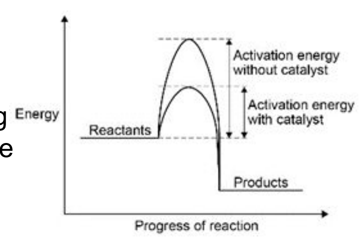
Energy profiles

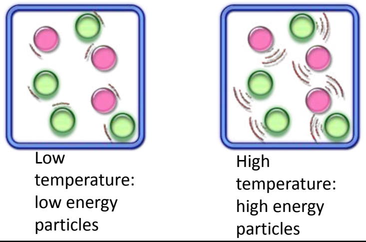
- Activation energy: difference between reactants and the peak of the curve
- Overall energy change – difference between reactants and products



Catalysts

- Catalysts change the rate of a chemical reaction but are not used up.
- Different reactions need different catalysts
- In biology, enzymes act as catalysts
- Catalysts work by providing an alternate pathway for the reaction with a lower activation energy.





Temperature

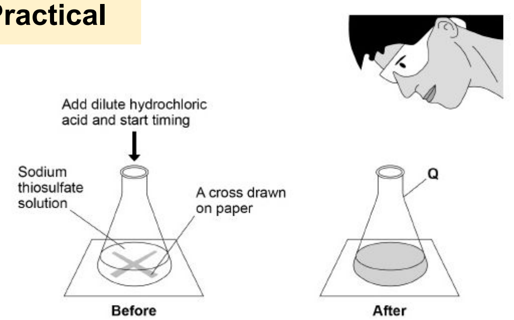
Increasing the temperature of a reaction has 2 effects:

- Increases the frequency of collisions
- Increases the energy of the collisions

Therefore, at higher temperatures, there are more collisions that have enough energy for a reaction to take place

Change in turbidity method

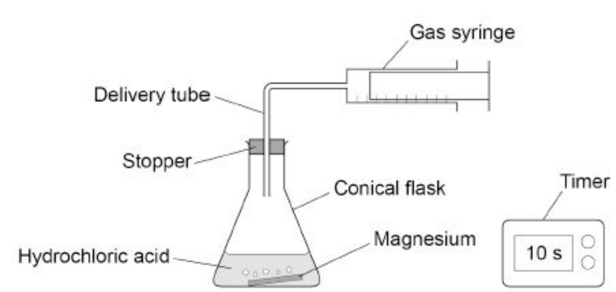
Required Practical



1. Put 50 cm^3 of sodium thiosulfate solution into a container.
2. Put the container on a cross drawn on a piece of paper.
3. Add 5 cm^3 of dilute hydrochloric acid and start timing.
4. Stop timing when the cross can no longer be seen.
5. Record the time taken.
6. Repeat steps 1-5 with different concentrations of sodium thiosulfate.

Gas collection method

Required Practical



1. Pour 50 cm^3 of hydrochloric acid into a conical flask.
2. Add a piece of magnesium.
3. Insert stopper and delivery tube and start a timer.
4. Collect the gas produced in a gas syringe.
5. Record the volume of gas produced every 20 seconds for 2 minutes.
6. Repeat steps 1 to 5 with higher concentrations of hydrochloric acid.

Linked Modules: AQA

- Atomic Structure
- Energy Changes
- Chemical Changes
- Organic Chemistry
- Using Resources

Useful YouTube links

Rates of Reaction

Required Practical

Rate of Reaction Part 2

Keyword	Definition
Anhydrous	A substance containing no water.
Catalyst	Speeds up the rate of a chemical reaction without changing the end product or getting used up in the process.
Endothermic	These are reactions that **take in energy from the surroundings.**
Exothermic	These are reactions that **transfer energy to the surroundings.**
Haber process	an industrial process for producing ammonia from nitrogen and hydrogen, using an iron catalyst at high temperature and pressure.
Le Chateliers Principle	A principle stating that if a condition (such as a change in pressure, temperature, or concentration of a reactant) is applied to a system in equilibrium, the equilibrium will shift to try to counteract the change of the condition.
Nitrate deficiency (separate Biology)	Plants use nitrates as a supply of nitrogen to make protein for healthy growth
Phosphorus deficiency (A level)	Plants make phospholipids for cell membranes.
Potassium deficiency (for info)	Chlorosis (yellowing of leaves) and stunted growth

Production and use of NPK fertlisers.

(from the Using resources unit).

NPK fertilisers are compounds (formulations) that contain **nitrogen, potassium** and **phosphorus** they are used to increase crop yields. An advantage of artificial fertilisers is that they can be **designed** for specific needs. (Link to Biology and deficiencies in plants). The ions they produce are soluble.

Ammonia is used to produce fertilisers as follows:

Ammonia + Hydrochloric acid □ Ammonium chloride

Ammonia + Sulfuric acid □ Ammonium sulfate

Ammonia + Nitric acid □ Ammonium nitrate

It can also be oxidised to form nitic acid.

The Earth's crust also contains **useful minerals** which are useful raw materials for making fertilisers

Phosphate rocks contain potassium chloride and potassium sulfate as they provide potassium. The rock is **insoluble** in water, so it is usually reacted with acid to produce useful compounds with water soluble compounds:

With **nitric acid,** phosphoric acid and **calcium nitrate** are produced.

The phosphoric acid is neutralised with ammonia forming **ammonium phosphate**

With **sulfuric acid**, a mixture of **calcium phosphate and calcium sulfate** is produced - called single superphosphate

With **phosphoric acid**, **calcium phosphate** is produced - this is known as triple superphosphate

Calculating the rate of reaction (from Rates Part 1).

The rate of reaction can be calculated using:

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

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

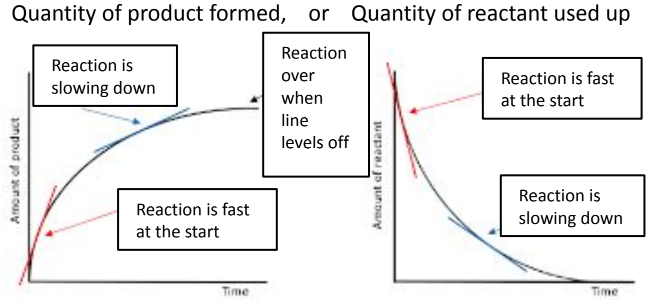
The quantity of reactant or product can be measured by the mass in grams or by a volume in cm^3

The units of rate of reaction may be given as g/s or cm^3/s

Interpreting rate of reaction graphs

Either as

Quantity of product formed, or Quantity of reactant used up



Using a tangent to interpret rate of reaction graphs

The steeper the line the faster the rate of reaction. In the examples above, the red tangent line is steeper than the blue tangent line.

Calculating the rate using a tangent (Higher only)

Step 1: Draw a line that touches the curve at the point you wish to measure the rate (20 seconds in example below)

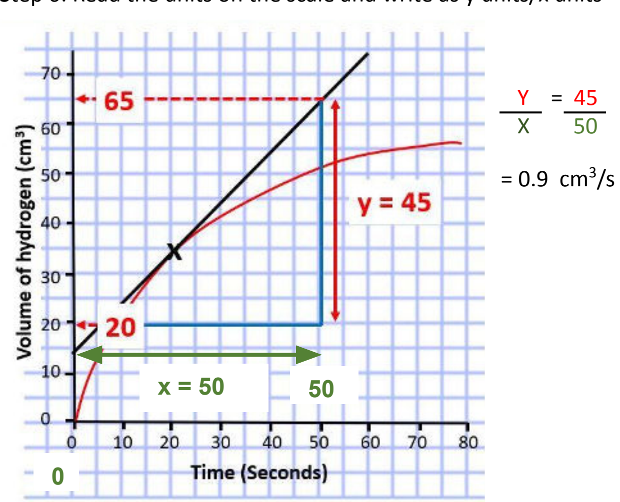
Step 2: Form a triangle that touches the tangent line at the top and the bottom, using values that are easy to read from the scale.

Step 3: Write down the values for the x axis part of the triangle and the y axis part of the triangle.

Step 4: Calculate the values for x and y.

Step 5: Insert the values as y/x and calculate.

Step 6: Read the units on the scale and write as y units/x units



Reversible reactions

In some chemical reactions, the products of the reaction can react to produce the original reactants. Such reactions are called reversible reactions, These can be shown as

$$A + B \rightleftarrows C + D$$

The direction of reversible reactions can be changed by changing the conditions.

$$\text{Ammonium chloride} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftarrows}} \text{ammonia + hydrogen chloride}$$

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

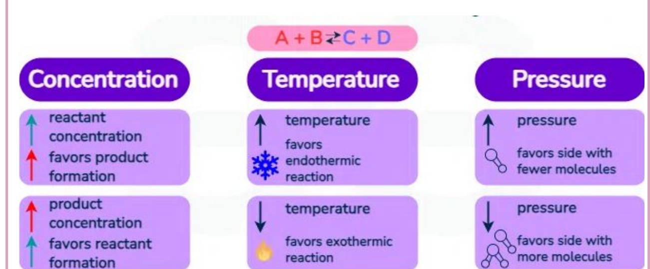
$$\underset{\text{(blue)}}{\text{Hydrated copper sulfate}} \underset{\text{exothermic}}{\overset{\text{endothermic}}{\rightleftarrows}} \text{anhydrous copper sulfate} + \text{water} \;\text{(white)}$$

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

Reversible reactions (Higher Tier only)

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

If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change. The effects of changing conditions on a system at equilibrium can be predicted using Le Chatelier's Principle.



Using as an example:

$$N_2\,(g) + 3\,H_2\,(g) \underset{\text{endothermic}}{\overset{\text{exothermic}}{\rightleftarrows}} 2\,NH_3\,(g)$$

A summary of the effects of changing conditions can be seen in the table below.

Change/Stress on System	Direction of Shift	Reason for Shift
Increase temperature	Left / reverse	To remove the extra heat
Decrease temperature	Right/ forward	To replace the lost heat
Add more N_2 or H_2	Right/ forward	To decrease the extra reactant
Add more NH_3	Left / reverse	To decrease the extra product
Increase pressure	Right/ forward	To decrease the moles of gas
Decrease pressure	Left / reverse	To increase the moles of gas

The commercially used conditions for the Haber process are related to the availability and cost of raw materials and energy supplies, control of equilibrium position and rate.

The Haber Process (Chemistry only) (from the Using resources unit).

Ammonia is made using the **Haber Process**

The reactants are hydrogen (from natural gas) and nitrogen (from the air).

1 H_2 and N_2 gases are pumped into the compressor through pipes

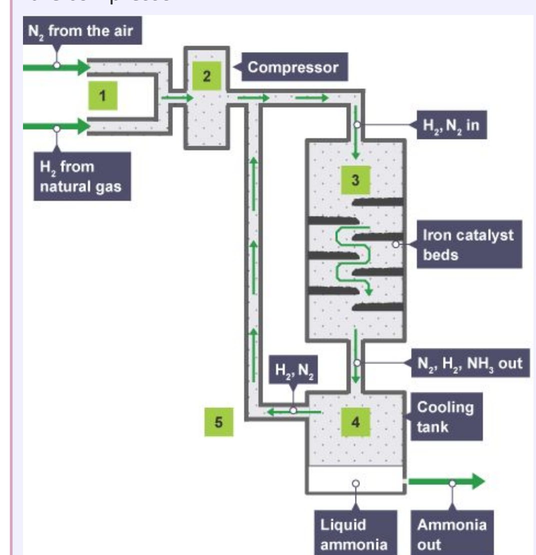
2 The gases are compressed to about 200 atmospheres inside the compressor

3 The pressurised gases are pumped into a tank containing layers of iron catalyst beads at a temperature of 450°C. Some of the hydrogen and nitrogen react to form ammonia in the following reversible reaction:

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

4 Unreacted H_2 and N_2 and the product NH_3 pass into a cooling tank. The ammonia is liquefied (as it has a lower boiling point) and is removed.

5 The unreacted H_2 and N_2 gases are recycled back into the compressor.

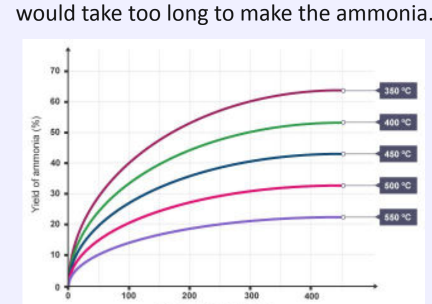


The compromise made in Industry

Due to the Haber process being a reversible reaction, the yield of ammonia can be changed by changing the pressure or temperature of the reaction

Increasing the pressure of the reaction increases the yield of ammonia. However, if the pressure is made too high, the equipment needed to safely contain the reaction becomes very expensive.

Increasing the temperature of the reaction actually decreases the yield of ammonia in the reaction. This means that we could get a bigger yield of ammonia with a lower temperature. However, if the temperature is too low, the rate of the reaction would be so slow that it would take too long to make the ammonia.



Organic Chemistry - Crude oil

Keyword	Definition
Alkane	A saturated hydrocarbon with the general formula CnH2n+2.
Combustion	Burning a fuel in oxygen.
Cracking	The reaction used in the oil industry to break down large hydrocarbons into smaller, more useful ones.
Empirical formula	A formula giving the proportions of the elements present in a compound but not the actual numbers or arrangement of atoms
Fractional distillation	A way to separate liquids from a mixture of liquids by boiling off the substances at different temperatures, then condensing and collecting the liquids.
Fractions	Hydrocarbons with similar boiling points separated from crude oil.
Hydrocarbons	A compound containing only hydrogen and carbon.
Organic	Compounds that contain carbon covalently bonded to other elements, typically, hydrogen, oxygen and nitrogen.
Polymers	A substance made from very large molecules made up of many repeating units (monomers).
Saturated hydrocarbon	A hydrocarbon with only single bonds between its carbon atoms.
Thermal decomposition	The breakdown of a compound by heating it.
Unsaturated hydrocarbon	A hydrocarbon whose molecules contain at least one carbon-carbon double bond.

Hydrocarbons

Organic molecules containing carbon and hydrogen, only. Hydrocarbons differ in the number carbon atoms that make up the molecule, e.g. C_2H_4 or C_3H_8

Hydrocarbon molecules can be represented in several ways: **Empirical formula, Structural formula or Atomic diagram**

Complete Combustion

Excess oxygen available. Products are carbon dioxide and water only. For example:

$CH_4 + 2O_2$ □ $CO_2 + 2H_2O$

Incomplete Combustion

Insufficient oxygen available. Products include carbon monoxide, carbon (seen visibly as soot) and water. For example:

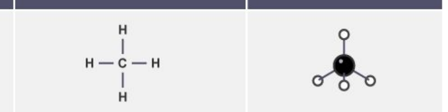
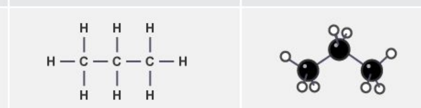
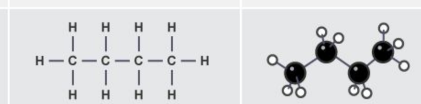
$C_2H_6 + O_2$ □ $CO + C + 3H_2O$

Crude Oil

- Finite (it will run out) resource found in rocks underground.
- Remains of plankton, buried in sediments millions of years ago.
- Mixture of many different compounds, mostly hydrocarbons.
- Hydrocarbons are molecules containing only hydrogen and carbon atoms.

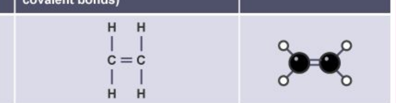
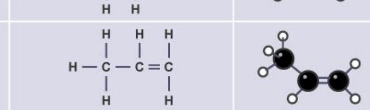
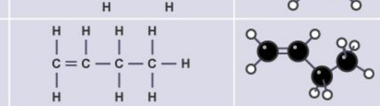
Alkanes

- General formula for alkanes is $C_2H_{2n+2.}$
- Only single bonds between each carbon atom.
- Each molecule is saturated (no bonds spaces left).

Alkane	Molecular formula	Structural formula	Ball-and-stick model
Methane	CH_4		
Ethane	C_2H_6		
Propane	C_3H_8		
Butane	C_4H_{10}		

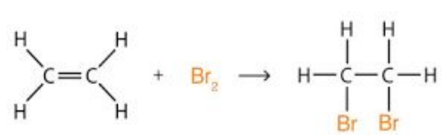
Alkenes – Chemistry only

- General formula for alkenes is C_2H_{2n}
- One double bond per molecule.
- Each molecule is unsaturated, having a double bond between two carbon atoms.

Alkene	Molecular formula	Structure (showing all the covalent bonds)	Ball-and-stick model
Ethene	C_2H_4		
Propene	C_3H_6		
Butene	C_4H_8		

Testing for alkenes

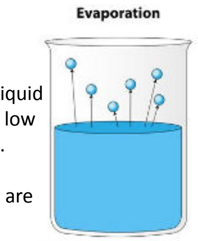
Alkenes can be identified using bromine water. It is orange (or yellow/brown) if alkenes are not present, however will turn colourless in the presence of alkenes



ethene + bromine □ 1,2-dibromoethane

The bromine breaks the double bond and covalently bonds with carbon.

Volatility



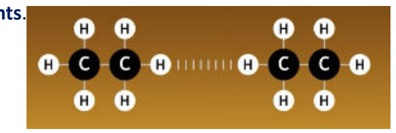
How easily a liquid evaporates at low temperatures. Short-chain hydrocarbons are very volatile.

Viscosity

How much a liquid resists flowing. Longer-chain hydrocarbons are more viscous than shorter-chain hydrocarbons.

Alkanes and boiling point

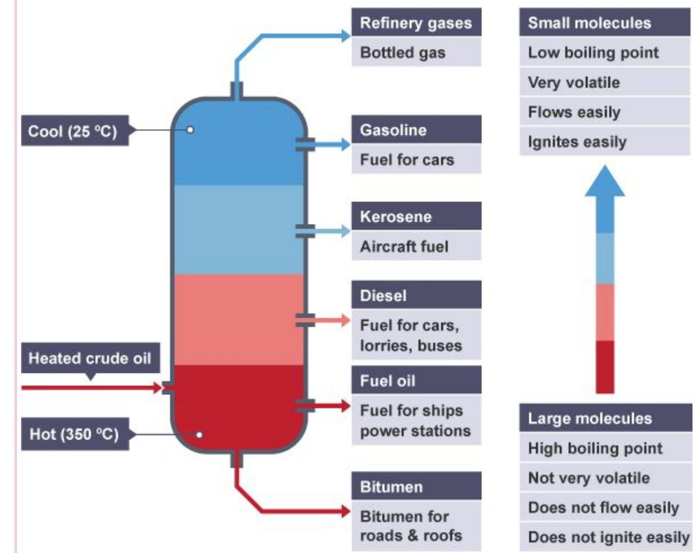
Shorter-chain alkanes have **weaker intermolecular forces** than longer chain hydrocarbons, therefore, have **lower boiling points**.



Fractional distillation

Crude oil is a mixture, so can be separated by heating as the different fractions have different boiling points.

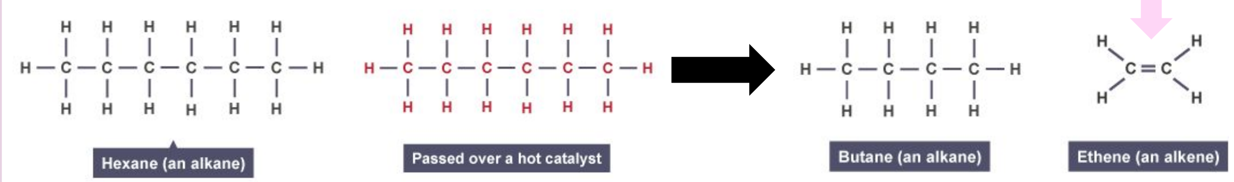
The crude oil is vaporised on entering the column. There is a temperature gradient inside the column, hotter at the bottom and cooler at the top. Individual fractions condense at the level corresponding to their boiling point.



Cracking Hydrocarbons:

- Shorter-chain (C5-C8) fractions are used to make petrol.
- The demand for shorter-chain fractions is greater than the availability of these fractions.
- Longer-chain fractions can be cracked to meet some of this demand.
- Alkanes that are 'cracked' undergo a **Thermal Decomposition** reaction.
- Cracking always produces at least one alkene, for example:

A double bond is formed because there are not enough hydrogen atoms to form single bonds on each molecule.

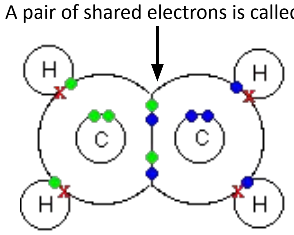


Covalent bonding reminder from the Bonding Unit we studied in Year 9

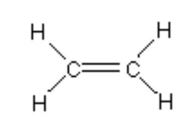
Covalent bonds between atoms are **strong** and the **intermolecular forces** between molecules are **weak**.

- Carbon has four electrons on the outer shell (needs 8 electrons to complete it)
- The double (covalent) bond is **pair of shared electrons**.
- All of the shared electrons between hydrogen and each carbon atom give each carbon atom a full outer shell.

A pair of shared electrons is called a **double bond**.



or



Useful YouTube links:

Crude oil | Cracking | Alkenes

Chemical Analysis – Common content

Keyword	Definition
Chromatography	A technique to separate a mixture of chemical substances into its compounds.
Chromatogram	A chromatogram is essentially the output of a chromatography run.
Formulation	A mixture that has been designed to produce a useful product
Impurity	When there is another substance that make a pure substance impure.
Insoluble	A substance that does not dissolve in a solvent.
Mobile phase	.The solvent
Purity	How pure or impure a substance is.
Soluble	A substance that dissolves in a solvent.
Solvent	The liquid that can dissolve a solute into it.
Stationary phase	The chromatography paper
Retention factor	A measurement from chromatography: it is the distance a spot of substance has been carried above the baseline divided by the distance of the solvent front.

Everyday versus Chemistry language

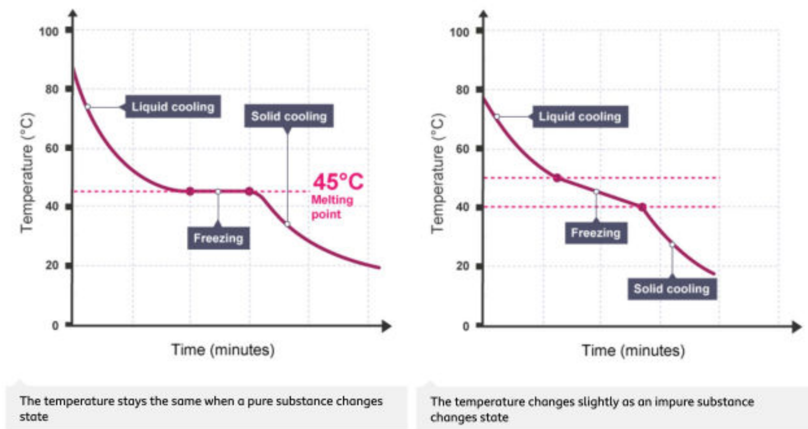
- In everyday language, a pure substance can mean a substance that has had nothing added to it, so it is unadulterated and in its natural state, e.g., pure milk
- In Chemistry a pure substance consists only of one element or compound.

Common Chromatography mistakes?

- Drawing start line in pen – pen is soluble. Start line must be in pencil
- Having solvent above the start line – dye will dissolve in solvent –start line must be above the surface of the solvent.
- Having a lid on the chromatography beaker will cause the results to be wrong – NO!

Purity

- An element contains atoms with the same atomic number.
- A compound contains two or more elements chemically joined together.
- A mixture contains two or more different substances that are not chemically joined together. The different substances in a mixture can be elements and/or compounds.
- In Chemistry a pure substance consists only of one element or compound.
- A mixture consists of two or more different substances, not chemically joined together.
- Pure substances have a sharp melting point, but mixtures melt over a range of temperatures.



The temperature stays the same when a pure substance changes state

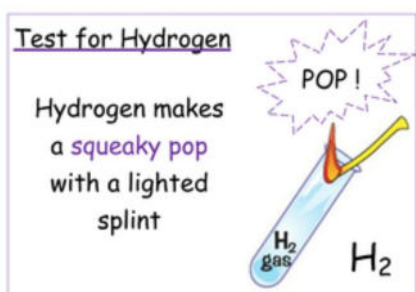
The temperature changes slightly as an impure substance changes state

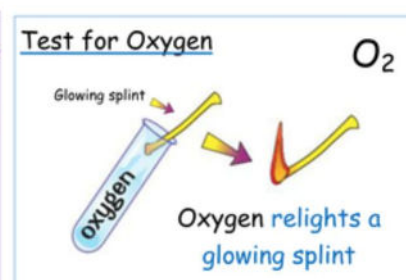
Formulation

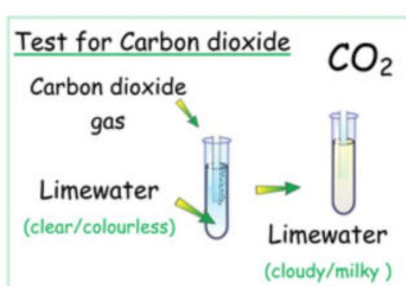
- A formulation is a mixture that has been designed as a useful product.
- Formulations are made by mixing the components in carefully measured quantities to ensure that the product has the required properties.
- Examples of formulations include fuels, paints, medicines, alloys, fertilisers and food.

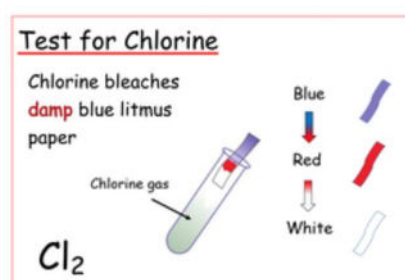
Testing gases

- You can be asked to recall the gas tests in Chemical changes topic to prove what is produced at the electrodes.

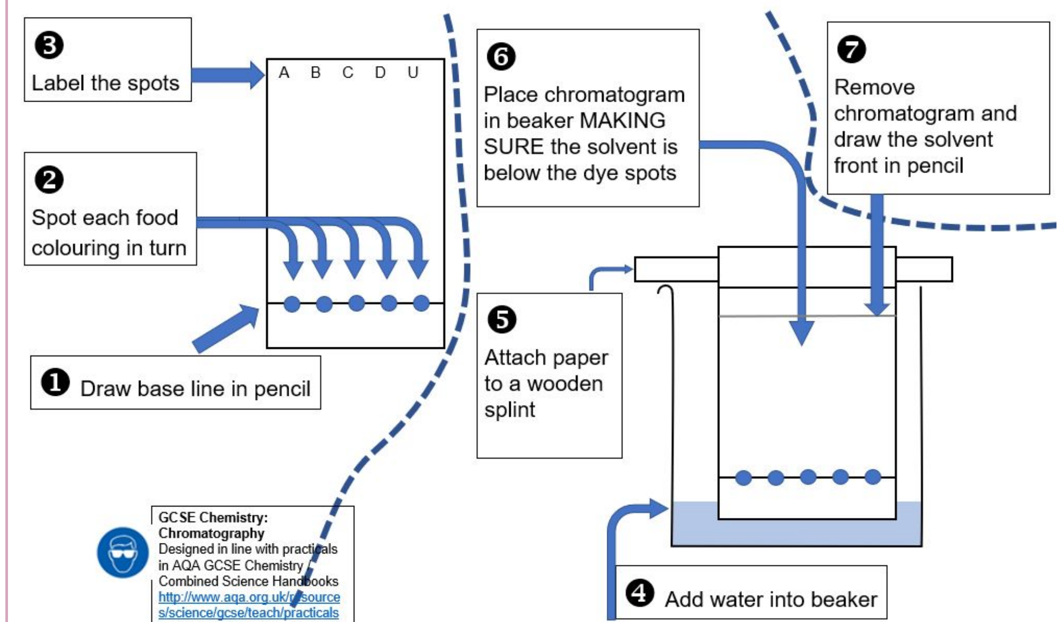








Chromatography Required Practical

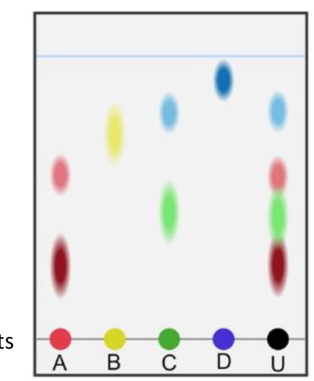


Chromatography Required Practical

- Paper chromatography is a method of separating dissolved substances from one another.
- Separation depends on the distribution of substance between the stationary phase (paper) and mobile phase (solvent such as water or ethanol).
- The compounds in a mixture may separate into different spots depending on the solvent but a pure compound will produce a single spot in all solvents.
- The ratio of the distance moved by a component to the distance moved by the solvent can be expressed as its R_f value.

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

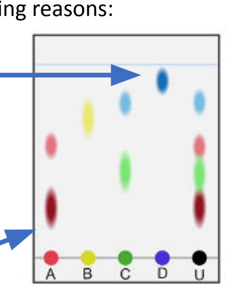
- Different compounds have different R_f values in different solvents



Why do the dyes travel different distances in a solvent?

A dye that travels higher on the chromatography paper, does so for the following reasons:

- It is weakly attracted to the paper (stationary phase)
- It spends a longer time in the mobile phase (solvent)
- So, it travels at a faster speed
- It travels a longer way on the chromatography paper
- Because of these reasons it has a higher Rf value.



Therefore, a dye that does not travel as high on the chromatography paper:

- Is more strongly attracted to the paper (stationary phase)
- It spends a shorter time in the mobile phase (solvent)
- So, it travels at a slower speed
- It doesn't travel as far in cm up the paper
- Therefore, has a lower Rf value

Chromatography revision video

Why do the speeds vary?
Speed = distance/time

Dye A travels 1.5cn in 300s
so speed = 0.05 cm/s
Dye B travels 6cm in 300s
so speed = 0.02 cm/s

Using Resources

Keyword	Definition
Bioleaching [HT]	A method of extracting copper from copper containing compounds using genetically modified bacteria.
Desalination	The process of removing salt from water using distillation or reverse osmosis.
Finite	A resource that will run out (but may be recycled).
Ground Water	Water from rivers, lakes and reservoirs.
Lifecyle assessment	An evaluation of the environmental impact of products, processes or services at different stages in their life cycle (i.e. from original manufacture to becoming a waste product).
Phytomining [HT]	A method of extracting copper from the ash of plants that absorbed copper ions from soil.
Potable water	Water that is safe for humans to drink.
Pure water	Water that does not contain any dissolved substances.
Recycle	Processing waste materials so that they can be used again.
Renewable	A resource that does not run out.
Reverse Osmosis	A method for removing salt from water by using pressure to force water against the concentration gradient.
Sustainable	When a resource can be managed so that it does not run out.

Finite Resources

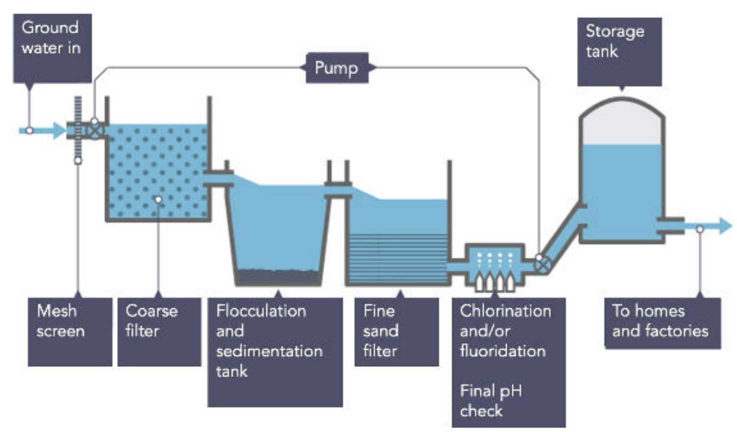
- Many natural resources are **finite** and some are **renewable.**
- All resources need to be managed so that they are **sustainable** and available for future generations.
- Agriculture and industry provide us with products, including:

- Chemistry plays a role in developing new materials that improve the lifecycle of products, are **recyclable** and use less resources.
- **Synthetic** products made in a lab or factory can replace natural products.

Water Purification

- Potable water is obtained through water purification of fresh water.
- The stages of water purification include:
  choosing an appropriate source of fresh water
  passing the water through filter beds
  sterilising using chlorine, fluorine or ozone (to kill pathogens)
- A flocculant may be added to help organic matter 'clump together' and sink to the bottom.

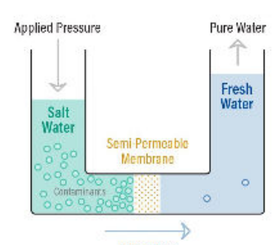
Stages of water treatment



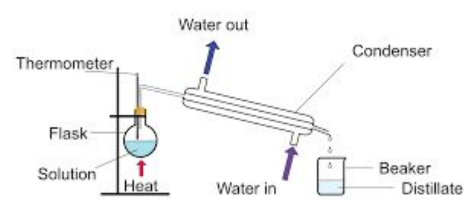
Distillation and Desalinisation

- Where access to fresh water is limited, salt water can be purified using **distillation** and **reverse osmosis**.
- Distillation involves heating the salt water to boiling point, then recondensing the water, leaving behind the salt and other minerals.
- Reverse osmosis involves passing the salt water through a membrane at high pressure, forcing the water through, but leaving behind the salt and other minerals.
- Both methods use lots of energy.

Reverse Osmosis



Distillation



Reduce, reuse, recycle.

The reduction in use, reuse and recycling of materials reduces the use of **limited resources**, **energy**, **waste** and **environmental impacts**.
The following are produced from limited resources (finite):

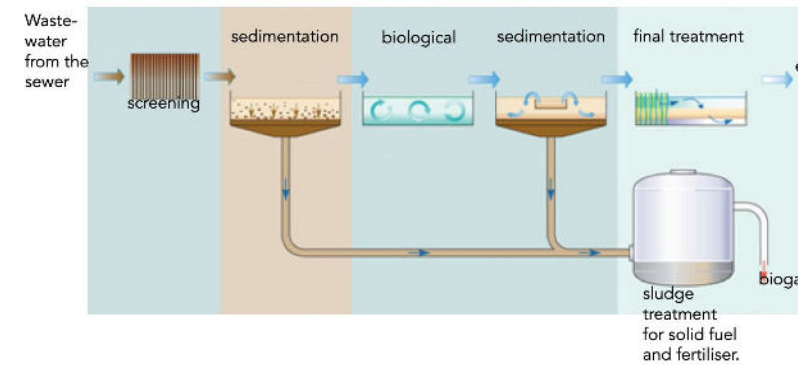
Meta ls **Glas s** **Plasti cs** **Clay (ceramic s)**

Obtaining raw materials from the Earth by **quarrying** and **mining** causes environmental impacts.

Waste Water

- Our lifestyle produces a lot of **wastewater** that must be treated before it is released into the environment.
- Untreated wastewater (raw sewage) and agricultural runoff pose a threat to the wild environment as it may contain harmful chemicals, organic matter and microorganisms.
- Sewage treatment includes:
  - the removal of dirt and grit
  - sedimentation
  - anaerobic digestion of sewage sludge (the heavier sediment)
  - aerobic biological treatment of the effluent (the liquid)

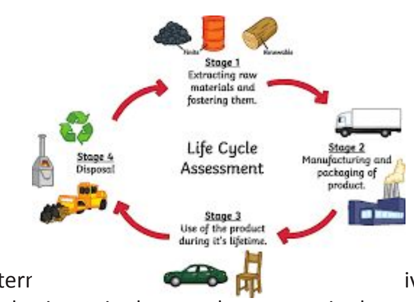
Stages of Sewage Treatment



Life Cycle Assessments (LCAs)

The main stages of a Lifecycle Assessment (LCA) are:

1. Extracting and processing the **raw materials** needed
2. Manufacturing the product and its packaging
3. Using the product during its lifetime
4. Disposing of the product at the end of its useful life.



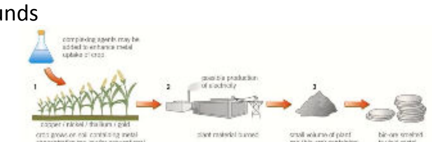
Comparative LCAs can be used to **evaluate** which of two alterr ive impact on the environment. For example, we can compare plastic carrier bags and paper carrier bags. All stages of the LCA should be considered before drawing conclusions as to which product has the lowest environmental impact.

Alternative methods of extracting metals – Higher tier only
Taught in the Chemical Changes unit

The Earths resources are limited, and sometimes a deposit of a metal ore is not economically viable to extract it by the usual quarrying methods. Alternative methods of extraction include:

Phytomining

Phytomining uses plants to absorb metal compounds. The plants are harvested and then burned to produce ash that contains metal compounds



Bioleaching

Bioleaching uses bacteria to produce leachate solutions that contain metal compounds. Low concentrations of metal ore are not a problem for bacteria because they simply ignore the waste which surrounds the metal ores, attaining extraction yields of over 90% in some cases. The bacteria gain energy by breaking down the metal ores into the leachate solution containing metal compounds.
The metal compounds can be processed to obtain the metal. For example, copper can be obtained from solutions of copper compounds by displacement using scrap iron or by electrolysis.

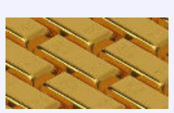
Using Resources

Keyword	Definition
Alloy	A mixture of two or more metallic elements producing a new metal with properties of each metal.
Borosilicate glass	A type of heat-resistant glass made from silicon dioxide and boron trioxide.
Ceramic	Any material made by forming and baking clay at high temperatures.
Composite	A material made from two or more constituent materials that has different properties to the individual materials that make it.
Corrosion	A process of metals reacting with substances in the environment. Rusting of iron and steel is an example.
Lifecycle assessment	An evaluation of the environmental impact of products, processes or services at different stages in their life cycle (i.e. from original manufacture to becoming a waste product).
Matrix	A substance that binds together the fibres that make up a composite material.
Recycle	Processing waste materials so that they can be used again.
Reinforcement	The fibres that are bound together by the matrix to make a composite.
Thermosetting	A plastic (polymer) that cannot be melted and reshaped.
Thermosoftening	A plastic (polymer) that can be melted and reshaped.

Alloys as useful materials – Chemistry only
Taught in the Bonding unit

- Alloying metals takes advantage of the different properties of each metal. Some important examples include:

Bronze
(copper + tin)



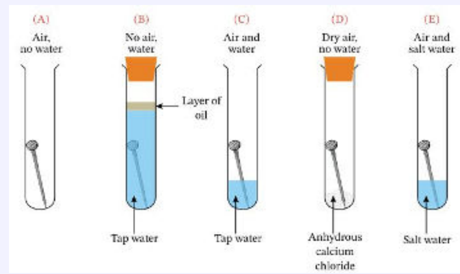
Gold (jewelry)
(silver + copper +zinc)

Steel
(iron + carbon + other)

- **Brass** is an alloy of copper and zinc.
- Gold alloy is measured in **carats**. 24 carat is 100% gold, 18 carat gold is 75% gold.
- **High carbon steel** is extremely strong, but brittle. **Low carbon steel** is softer and more easily shaped.
- **Stainless steel** contains **chromium** and **nickel** and is hard and resistant to corrosion.
- **Aluminum alloys** have a lower density.

Rusting and Corrosion – Chemistry only
Taught in the Chemical Changes Unit

Corrosion is the destruction of materials by chemical reactions with substances in the environment.
Corrosion is the name given to the process when a metal reacts with oxygen in the air to form a metal oxide.
Corrosion is the reverse of the process used to extract the metal.
If the metal is iron, corrosion is called rusting. The iron reacts with both oxygen and water to form iron oxide (rust).
In general, the more reactive a metal is the more quickly it corrodes,
Corrosion can be prevented by applying a coating such as grease, painting or electroplating to prevent the oxygen and water from reaching the metal.



The iron nail in test tube E will begin to rust first, followed by the iron nails in test tubes C, B, and A. The iron nail in test tube D should be the last to start rusting (if at all).

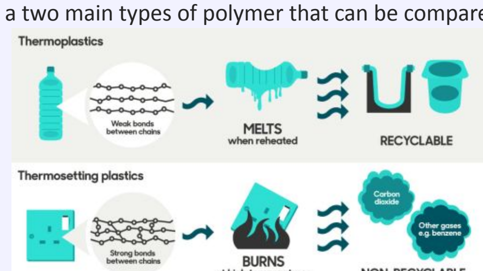
Aluminium
Aluminiums surface layer reacts with oxygen and thus forms an oxide coating that protects the metal from further corrosion (unless this layer becomes damaged).

Sacrificial Protection
Some coatings are reactive and contain a more reactive metal to provide a sacrificial protection.
For example:
Ships hulls are usually made from mild steel (an alloy with small quantities of carbon and manganese to allow it to be rolled during processing and not crack). However, iron rusts quicker in the salt water of the seas Therefore zinc blocks are attached to the hull of the ship. As these are more reactive than the steel hull, they will react to prevent the hull from rusting

Zinc blocks

Polymers – Chemistry only

The properties of a polymer depend on the **monomers** that they are made of.
There a two main types of polymer that can be compared:



Weak **intermolecular forces** between the polymer chains means that Thermosoftening plastics have low melting points.

Ceramics and Glass – Chemistry only

- Most glass we use is soda-lime glass.
- Soda-lime glass is made by heating a mixture of sand, sodium carbonate and limestone.
- Borosilicate glass is made from sand and boron trioxide and melts at a higher temperature than soda lime glass.

Glass	Shock resistance	Melting Point	Uses
Borosilicate	High	High	-Industrial equipment -Laboratory and kitchen glassware -Outdoor lighting
Soda-lime glass	Low	Average	-Windows -rood and drink containers -indoor lighting

- Clay ceramics are made by shaping wet clay, then heating it in a furnace.
- Ceramics have many uses as they are impermeable to water, and they are also heat and chemical resistant.

Ceramic fireplace

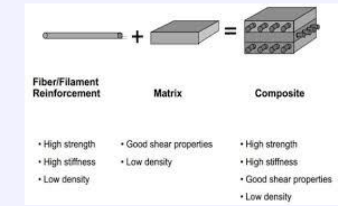
Ceramic mug

Ceramic countertop

Composites – Chemistry only

Composites are made from two materials: a **binder** (or matrix) surrounding and binding together fibers or fragments of another material, which is called the **reinforcement**.

The materials for a composite material are chosen because they have different properties that combine to make a more useful material. MDF is an example.



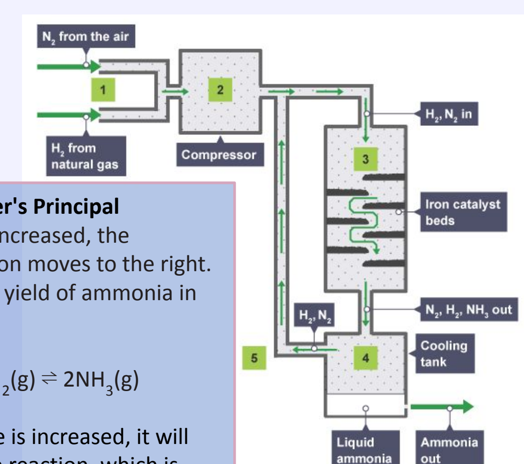
Steel-reinforced concrete is a composite material. It is made by pouring concrete around a mesh of steel cables. When the concrete sets, the material is:

1 strong when stretched (because of the steel)
2 strong when squashed (because of the concrete)

Ammonia and The Haber Process – Chemistry only

- The **Haber process** is used to manufacture ammonia, which can be used to produce nitrogen-based fertilisers. It is a reversible reaction.
- The process involves passing hydrogen and nitrogen gas over an **iron catalyst** at a high temperature (**450 °C**) and a high pressure (**200 atmospheres**).

- A balance of temperature and pressure is chosen as a **compromise** to yield, profit margin and safety.
- Ideal conditions are too slow, dangerous and expensive to maintain.



Revise Le Chatelier's Principal
If the pressure is increased, the equilibrium position moves to the right. This increases the yield of ammonia in this **equilibrium**.

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

If the temperature is increased, it will favour the reverse reaction, which is **endothermic**.

Fertilisers – Chemistry only

Compounds of nitrogen, phosphorus and potassium are used as fertilisers (NPK).
NPK fertilisers can be produced on an industrial level to meet the demands of agriculture.

NPK fertilisers are **formulations** of various salts containing appropriate percentages of the elements.

Salts can be produced in a lab by reacting an acid and alkali using the **titration** method.
For example:

Sulfuric acid + ammonia will make the nitrogen-based salt – **ammonium sulfate**.

Potassium chloride, potassium sulfate and phosphate rock are obtained by mining, but phosphate rock cannot be used directly as a fertiliser.

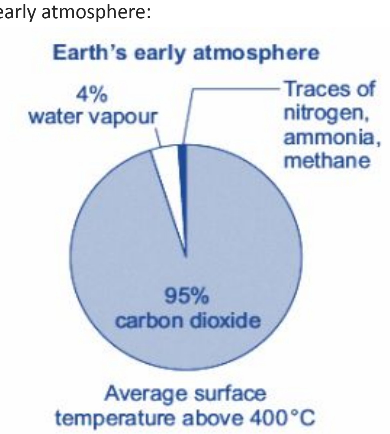
Phosphate rock reacts with
- Sulfuric acid to form single superphosphate
- Phosphoric acid to form triple superphosphate
- Nitric acid to form calcium nitrate and phosphoric acid

Chemistry of the Atmosphere

Keyword	Definition
Atmosphere	The thin layer of gases that surround some planets
Algae	An early organism capable of photosynthesizing
Photosynthesis	A reaction used by plants to produce glucose. Carbon Dioxide is absorbed, and Oxygen produced as a byproduct
Greenhouse gases	Gases in the atmosphere that absorb infra-red radiation
Carbon Footprint	The total amount of CO_2 / greenhouse gases emitted in the life cycle of a product
Particulate	Particles of unburnt carbon that contribute to atmospheric smog, global dimming and effects on human health.
Combustion	Burning fossil fuels in the presence of oxygen.
Acid rain	Produced as a result of oxides of nitrogen and sulfur dissolving in rainwater.
Carbonate	Compound containing CO_3 (a store of carbon in rocks)
Fossil fuel	Fuels containing carbon and hydrogen, e.g. coal, oil and gas.

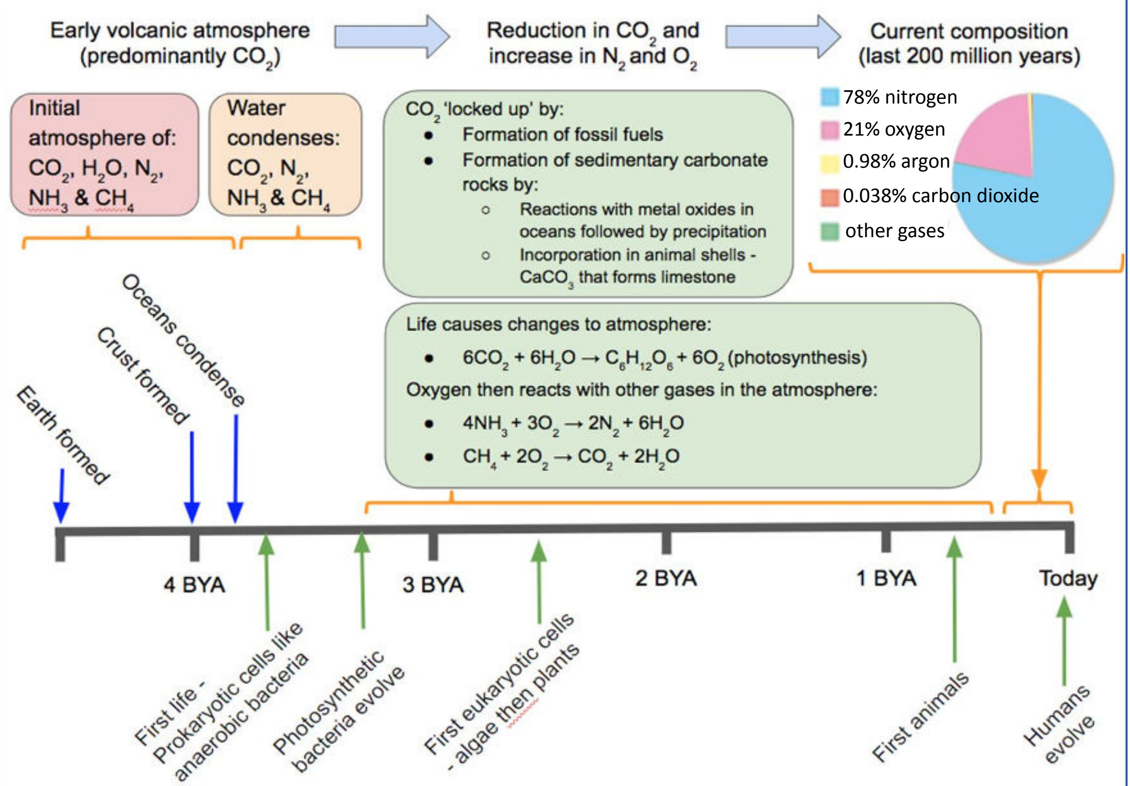
Earth's early atmosphere (4.6 Billion years ago)

- Theories about Earth's early atmosphere have changed over time.
- Evidence of the early atmosphere is limited due to changing over 4.6 billion years.
- A key theory suggests that volcanoes released gases such as: **nitrogen, carbon dioxide, water vapour and small amounts of methane and ammonia.**
- The current theory suggests the following percentages of gases in the early atmosphere:



Our evolving atmosphere

- The Earth's atmosphere has evolved to what it is today and continues to change (in part, due to human activity).



BYA = Billion years ago

Evolving atmosphere

- **Algae** and complex plants produced the oxygen that is now in the atmosphere by **photosynthesis** (see equation in green box).
- The increase in oxygen allowed animal life to evolve and spread all over the planet.
- The remains of **plankton** were deposited in the oceans, covered over and compressed over millions of years producing **fossil fuels** and **natural gas**.

Mathematical Skills:

- *You may be asked to use ratio, fractions and percentages.*
- *Be prepared to evaluate data from graphs and tables.*

Relevant Modules:

AQA

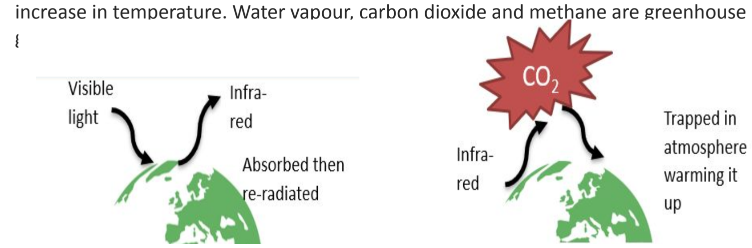
- Organic Chemistry
- Waves

Greenhouse gases

- Carbon dioxide locked-up in fossil fuels as carbon compounds can be **combusted** (burned) to re-release carbon dioxide into the atmosphere: A summary of the reaction is:

  fossil fuel + oxygen carbon dioxide + water

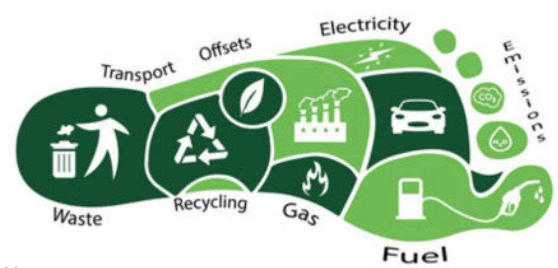
- Greenhouse gases in the atmosphere maintain temperatures on Earth by allowing **short wavelength radiation** to pass through the atmosphere to the Earth's surface but absorb the outgoing **long wavelength radiation** from the Earth causing an increase in temperature. Water vapour, carbon dioxide and methane are greenhouse



Shorter wavelength visible light
Longer wavelength infra red

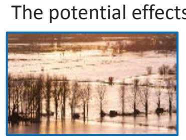
Carbon footprint

- The **carbon footprint** is the total amount of carbon dioxide and other greenhouse gases emitted over the full lifecycle of a product, service or event.
- **Actions** to reduce the carbon footprint include:
  - Increased use of renewable energy sources
  - Energy conservation
  - Carbon capture and storage
  - Carbon taxes
  - Carbon offsetting through tree planting
- **Problems** reducing the carbon footprint include:
  - Lack of public information and education
  - People making lifestyle changes
  - Economic problems in making the right changes
  - Incomplete co-operation between countries
  - Scientists cannot agree over the causes and consequences of global climate change.

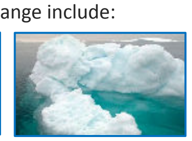


Climate change

- Some human activities increase the amount of greenhouse gases in the atmosphere. These include:
  - **carbon dioxide** from combustion and deforestation;
  - **methane** from farming (digestion in pigs, sheep and sow cows) and decomposition of rubbish in landfill sites.
- The increase in the percentage of carbon dioxide in the atmosphere has increased over the last 100 years with an increase in burning fossil fuels.
- Many scientists believe that human activities will cause the temperature of the atmosphere to increase, resulting in **global climate change**.
- The potential effects of global climate change include:


Flooding

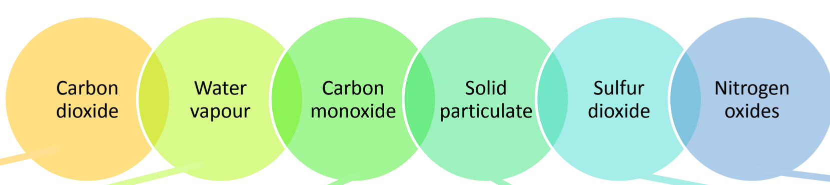

Desertification


Melting ice caps


Drought

Atmospheric pollutants

- Causes of **atmospheric pollution** and their effects on the environment are as follows:



Produced from the combustion of fossil fuels. **Contributes to global warming.**	Produced from the combustion of fossil fuels. **Contributes to global warming.**	Produced from the incomplete combustion of fossil fuels. **It is a toxic gas.**	Produced from the incomplete combustion of fossil fuels. **Causes global dimming.**	Produced from burning coal. Sulfur dioxide contributes to the **formation of acid rain.**	Nitrogen in air reacts with oxygen (at high temperatures) to produce **gases that form acid rain.**

Useful YouTube links:

Atmosphere
Global warming
Reducing CO_2