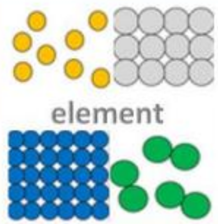
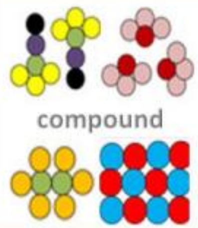
Knowledge Organiser – 5.1 Atomic structure & the periodic table

5.1.1.1 Atoms, elements & compounds

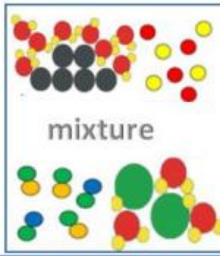
An **Atom** is the **smallest part of an element that can exist.**



Made from only one type of atom e.g. Gold is made of only gold atoms



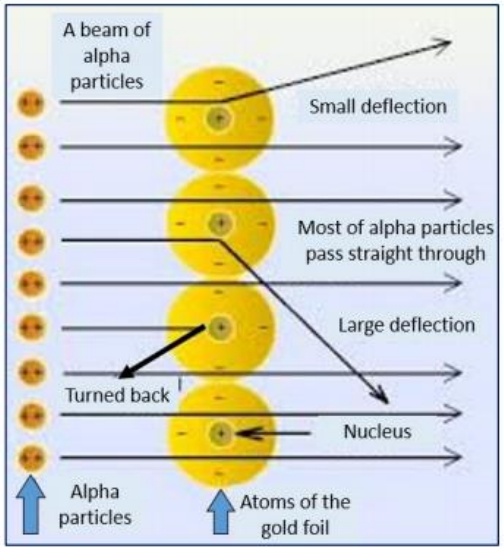
Two or more elements chemically bonded (joined) together. Separated by chemical reactions



Two or more elements that are NOT chemically combined together. Separated by physical processes

5.1.1.3 The development of the model of the atom

Ideas about atoms have changed over time. Scientists developed new atomic **models** as they gathered new experimental evidence.

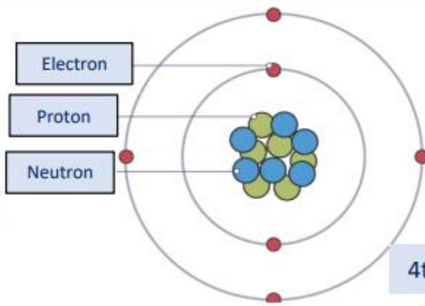


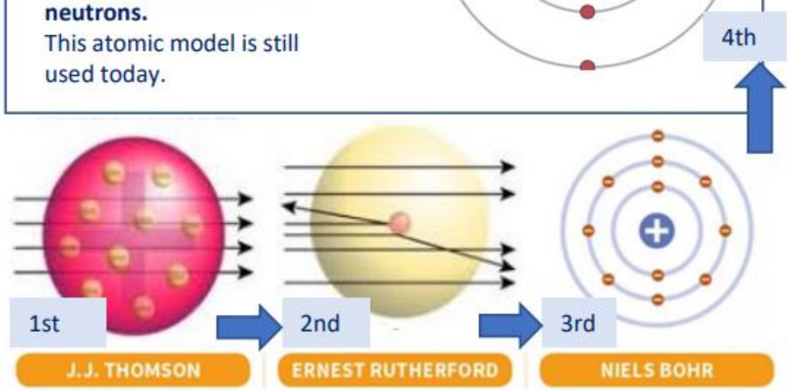
Ernest Rutherford's Gold scattering experiment

- **Positive alpha particles** fired at gold leaf
- **Most passed straight through** suggesting the atom was mainly empty space
- **Some deflected at angles** suggesting the presence of electrons
- **Some bounced straight back** suggesting a positive nucleus repelled the alpha particles.

The Nuclear Model

In 1932 James Chadwick found evidence for the existence of particles in the nucleus with mass but no charge.
These particles are called **neutrons.**
This atomic model is still used today.





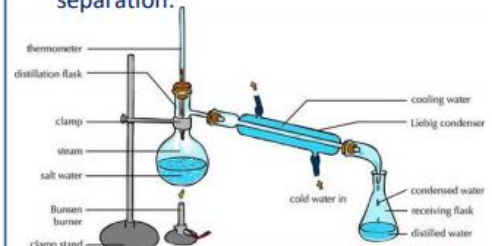
1904- **electron** discovered, placed in a sphere of positive charge (the **plum pudding model**)

1911 - gold scattering experiment discovered mass was concentrated in a central positive **nucleus** (the nuclear model)- further experiments led to discovery of **protons**

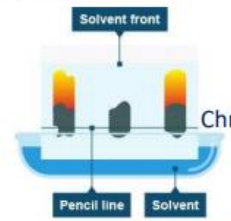
1913 - Suggested electrons orbit the nucleus in **shells.** The shells are at certain distances from the nucleus.

5.1.1.2 Mixtures

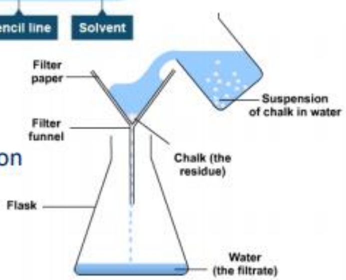
- Mixtures can be **separated by physical processes** such as **filtration, crystallisation, simple distillation, fractional distillation & chromatography**.
- These physical processes **do not involve chemical reactions** and **no new substances are made.**
- Examples of the specified processes of separation:



Simple distillation



Chromatography



Filtration

Crystallisation

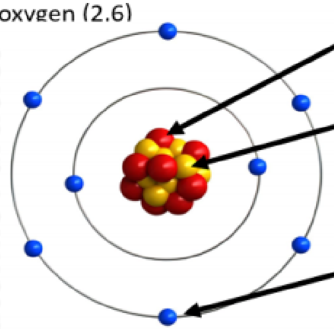
Knowledge Organiser – 5.1 Atomic structure & the periodic table

5.1.1.4 Relative electrical charges of subatomic particles & 5.1.1.7 Electron structure

Innermost, lowest energy level, **shell** has **2 electrons**.
Next shell 8, next shell 8. (2,8,8)

Electronic structure an be shown as a diagram or a number
eg. oxygen (2,6)

Sub-atomic Particles



Proton: Positive subatomic particle in the nucleus. Relative mass 1, charge +1

Neutron: Neutral subatomic particle in the nucleus. Relative mass 1, Charge 0 (no charge)

Electron: Negative subatomic particle orbiting the nucleus. Very small relative mass. Charge -1. Can be represented by dots or crosses

Atomic radius: 0.1 nm

Nucleus: The centre of an atom in which most of the mass of the atom is concentrated

5.1.1.5 Size and mass of atoms

- Atoms are **very small**, having a radius of about 0.1 nm (1 x 10-10 m).
- Atomic mass number: The sum (total) of the protons and neutrons in the nucleus of an atom of an element.
- Atomic (Proton) number: The number of protons in an atom of an element. Balanced by number of electrons in an atom of that element. (so atoms have no overall charge).

Name of particle	Relative mass
Proton	1
Neutron	1
Electron	Very small

12
C
carbon
6

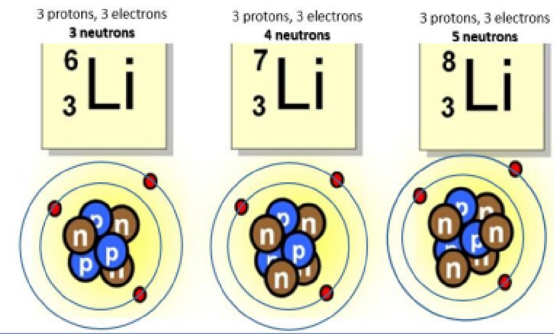
5.1.2.2 Development of the periodic table

- Early versions organized by atomic mass
- Didn't take account of isotopes
- Many elements missing
- Mendeleev ordered elements by atomic (proton) number
- Left gaps for undiscovered elements. Later discoveries proved him right.

5.1.1.6 Relative atomic mass

Isotopes are atoms of the same element with **different numbers of neutrons** in the nucleus.

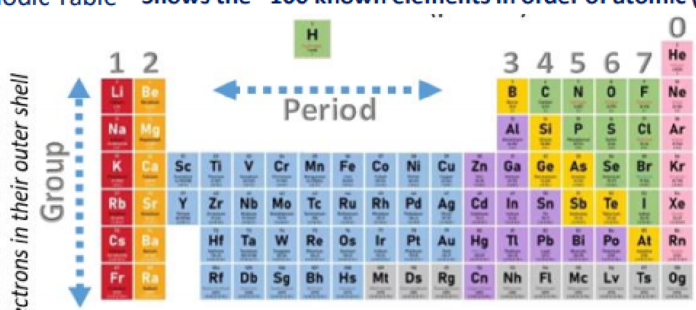
Relative atomic mass: Average value that takes account of the **abundance** of the different isotopes of that element.



5.1.2 Periodic Table

Shows the ~100 known elements in order of atomic (proton) number

Elements in groups have similar properties and have the same number of electrons in their outer shell

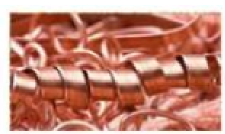


Group 0: Noble Gases Have full outer shells so stable and unreactive. Boiling points increase going down the group.

Group 1: Alkali Metals 1 electron on the outer shell. Reactivity INCREASES going down the group. Vigorous reactions with oxygen, chlorine and water.

Group 7: Halogens 7 electrons on the outer shell. Non-metals. Exist as diatomic (2 atoms) molecules. Reactivity decreases going down the group.

5.1.2.3 Metals & Non-metals



vs

Elements that react to form positive ions are **metals**.

Elements that do NOT form positive ions are **non-metals.**

Metals	Non-metals
Good conductors of heat and electricity	Bad conductors of heat and electricity
Malleable: can be beaten into thin sheets, hammered into shape	Brittle: breaks easily if solid
Ductile: can be stretched into wires	Non-ductile: snap easily
Shiny (lustre)	Dull

Knowledge Organiser – Separate Chemistry only

5.1.3 Properties of Transition metals

- A block of elements found between Groups 2 and 3 in the middle of the Periodic Table.

E.g. copper, nickel, chromium, manganese, cobalt and iron.

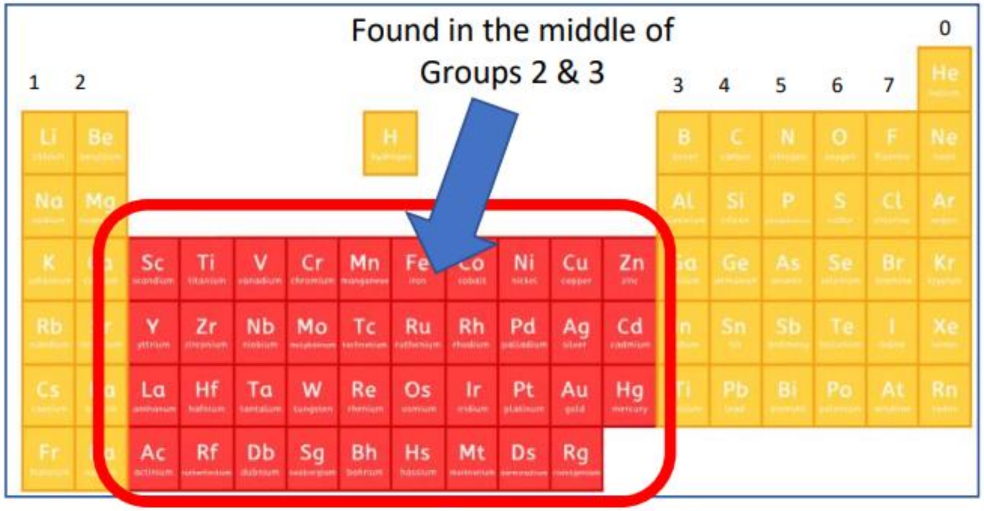
- Strong, shiny, conductors of heat/elec.
- Good catalysts (speed up chemical reactions without getting used up themselves)

E.g. iron is a catalyst in the Haber Process making ammonia for fertilisers.

- Can form more than ion

E.g. copper can become Cu^{+} or Cu^{2+} & iron can become Fe^{2+} or Fe^{3+}

- The ions and compounds formed from transition metals are often colourful



Exam Question: List three properties of transition metals that are different from the metals in group 1 (the alkali metals) (3)

Answer: The three main differences are:

- high melting point (group 1 metals have low melting points)
- hard (group 1 metals are soft)
- high density (group 1 metals have lower densities)

Uses of transition metals

Transition metals have a wide range of uses. Their properties are very similar but not identical. It is important to choose the right transition metal for the required purpose.

Transition Metal	Use	Reason
Gold	Jewellery	Does not react with air or water at room temperature Can be bent and hammered into shape (malleable)
	Electrical conductors	Good conductor of electricity
Silver	Jewellery	Does not react with air or water at room temperature
	Printed circuit boards and electrical contacts	Good conductor of electricity
Copper	Electrical wires	Good conductor of electricity Can be shaped into wires (ductile)
	Printed circuit boards	Good conductor of electricity
	Water pipes	Does not react with water at room temperature Can be hammered or bent into shape
Iron	Building materials (eg bridges, building, ships, cars)	Strong, sheets are easily shaped, and cheap compared to most other metals.
	Catalyst (eg in Haber process to produce ammonia)	Strong, sheets are easily shaped, and cheap compared to most other metals.
Chromium	Coat other metals eg coat iron on bikes	Stays shiny when polished, and resistant to corrosion.
	Catalyst in chemical industry	Increases the rate of certain reactions but can be recovered, unchanged, at the end

Knowledge Organiser – 5.1 Atomic structure & the periodic table

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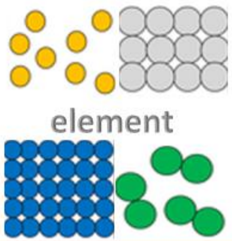
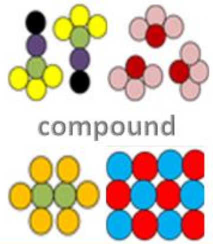


Made from only one type of atom e.g. Gold is made of only gold atoms



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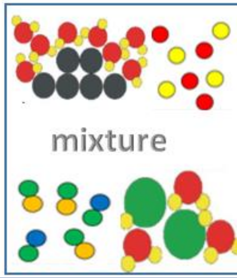

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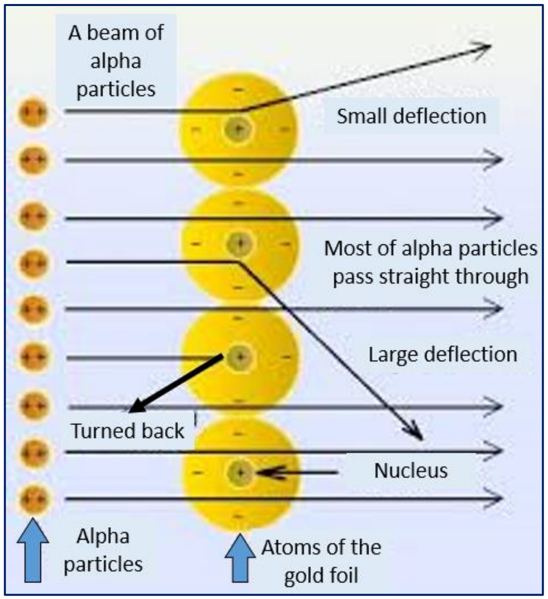

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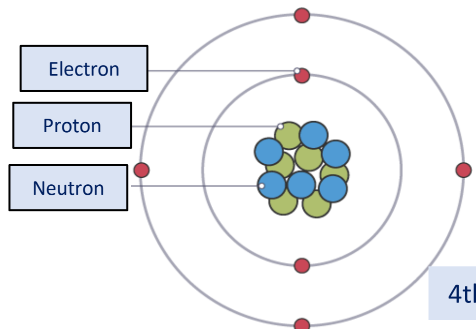
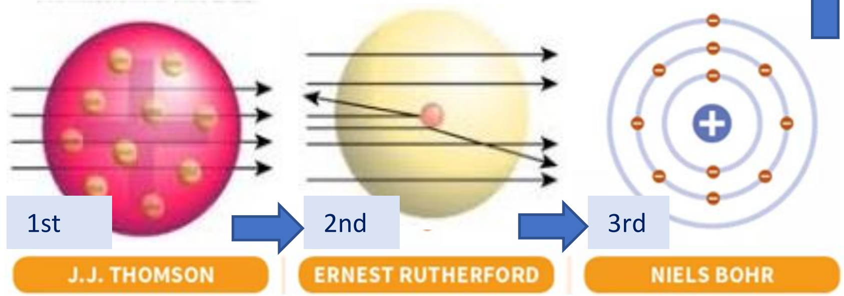


4th



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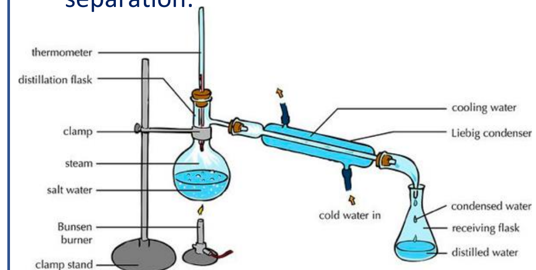
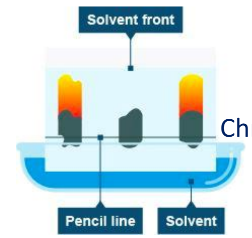


Simple distillation



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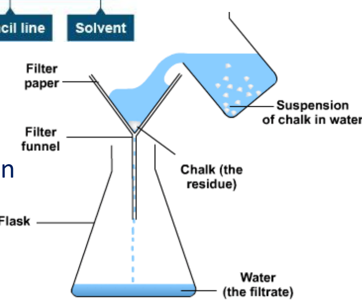
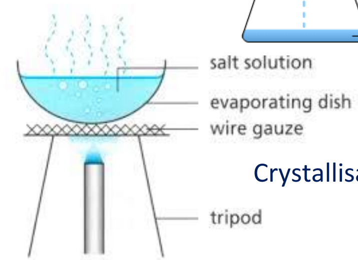


Filtration



Crystallisation

Knowledge Organiser – 5.1 Atomic structure & the periodic table

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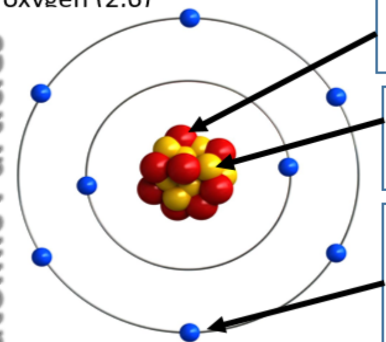
Next shell 8, next shell 8. (2,8,8)

Electronic structure an be shown as a diagram or a number eg. oxygen (2,6)

Sub-atomic Particles



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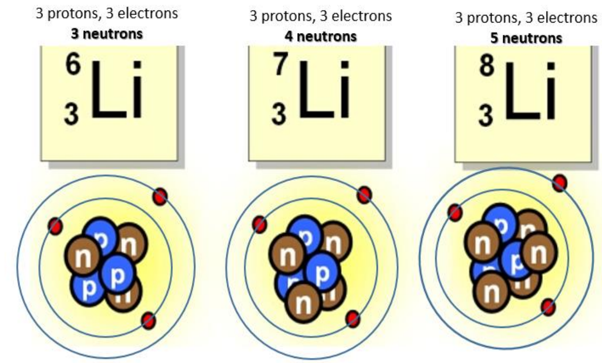
5.1.1.6 Relative atomic mass



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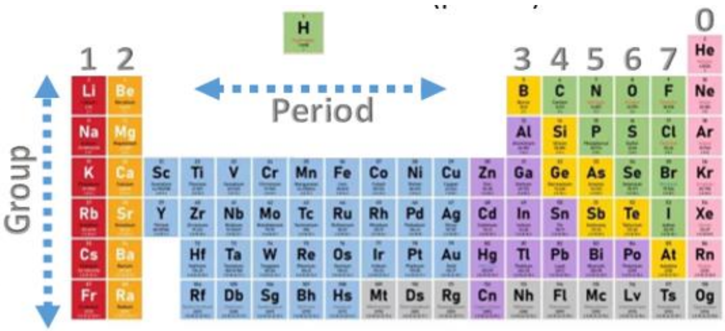


Group 0: Noble Gases
Have full outer shells so stable and unreactive. Boiling points increase going down the group.

Group 1: Alkali Metals
1 electron on the outer shell. Reactivity INCREASES going down the group. Vigorous reactions with oxygen, chlorine and water.

Group 7: Halogens
7 electrons on the outer shell. Non-metals. Exist as diatomic (2 atoms) molecules. Reactivity decreases going down the group.

5.1.2.3 Metals & Non-metals

vs

Elements that react to form positive ions are **metals**.

Elements that do NOT form positive ions are **non-metals**.

Metals	Non-metals
Good conductors of heat and electricity	Bad conductors of heat and electricity
Malleable: can be beaten into thin sheets, hammered into shape	Brittle: breaks easily if solid
Ductile: can be stretched into wires	Non-ductile: snap easily
Shiny (lustre)	Dull

Knowledge Organiser – 5.2 Structure & bonding

5.2.1.1 Chemical bonds

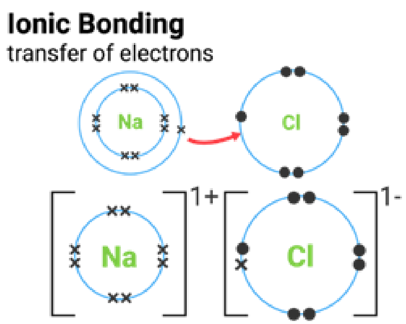
There are three types of strong chemical bonds: ionic, covalent and metallic.

Ionic bonding: particles are oppositely charged ions. Ionic bonding occurs in compounds formed from metals combined with non-metals

Covalent bonding the particles are atoms which share pairs of electrons. Covalent bonding occurs in most non-metallic elements and in compounds of non-metals

Metallic bonding the particles are atoms which share delocalised electrons. Metallic bonding occurs in metallic elements and alloys.

5.2.1.2 Ionic Bonding

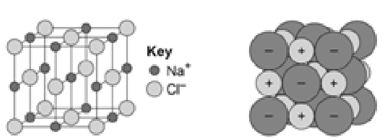


- Between a metal atom and a non-metal atom
- Metals lose electrons to form positive ions
- Non-metals gain electrons & form negative ions
- Electrons **transferred** (**ions** formed)
- Strong **electrostatic** forces
- Giant **lattice** structures
- High melting/boiling points
- If **molten** or in **solution** ions will conduct electricity

5.2.1.3 Ionic compounds

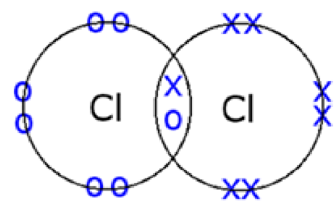
- An ionic compound is a **giant structure of ions.**
- Ionic compounds are held together by **strong electrostatic forces of attraction** between oppositely charged ions.
- These **forces act in all directions in the lattice** and this is called **ionic bonding.**

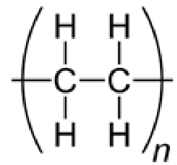
Eg: structure of sodium chloride

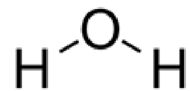


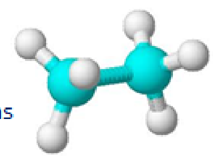
5.2.1.4 Covalent Bonding

Sharing electrons









- Between two non-metal atoms
- Electrons are **shared**
- A **covalent** bond is one pair of shared electrons
- Covalent bonds are ALWAYS STRONG

5.2.2.1. The three states of matter

- **Freezing** take place at the **melting point**
- **Boiling** and **condensing** take place at the **boiling point.**
- **Particle theory** can help to explain melting, boiling, freezing and condensing.
- The **amount of energy needed to change state** from solid to liquid and from liquid to gas **depends on the strength of the forces** between the particles of the substance.
- **The stronger the forces between the particles the higher the melting point and boiling point of the substance.**

5.2.2.3 Properties of Ionic compounds

- have **high melting points** and **high boiling points because** of the **large amounts of energy needed to break the many strong bonds.**
- When melted or dissolved in water, ionic compounds **conduct electricity because the ions are free to move and so charge can flow.**

5.2.2.4 Properties of Small molecules

- Usually **gases or liquids** with **low melting point & low boiling point.**
- Weak **intermolecular** forces (because they are small molecules) which are overcome when substance melts or boils.
- e.g. gases, water
- Do not conduct electricity as no overall electric charge.

5.2.2.5 Polymers

- **Long** molecules with atoms linked by **strong covalent bonds.**
- Solid at room temperature as **relatively strong intermolecular forces.**
- **Repeating** units
  e.g. plastics

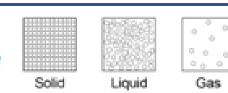
5.2.2.6 Giant covalent structures

Giant lattices

- High melting point and boiling point
- Strong covalent bonds which must be overcome to melt or boil.
- e.g. silicon dioxide, diamond, graphite

5.2.2.2. State symbols

In chemical equations, the three states of matter are shown as (s), (l) and (g).

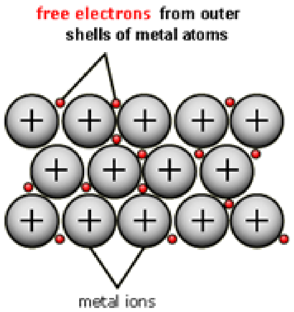


(aq) for aqueous solutions eg salt water or acid solutions.

Knowledge Organiser – 5.2 Structure & bonding

5.2.1.5 Metallic Bonding

- Bonding between atoms of a metal
- **Delocalised** electrons (negative) & metal ions (positive)
- Shared delocalised electrons form strong metallic bonds
- Delocalised electrons **conduct** heat and electricity



- Pure metals are soft: layers of atoms can slide over each other

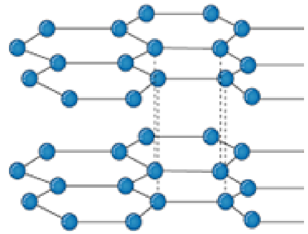
5.2.2.7 Properties of metals and alloys

- Metals have **giant structures of atoms** with **strong metallic bonding.** Therefore most metals have **high melting and boiling points.**
- In **pure metals, atoms are arranged in layers**, which allows metals to be **bent and shaped.** (malleable)
- **Pure metals** are too soft for many uses and so **are mixed with other metals to make alloys which are harder.**
- In **alloys**, different atoms **disrupt** the layers
- Alloys are **harder** than pure metals

5.2.2.8 Metals as conductors

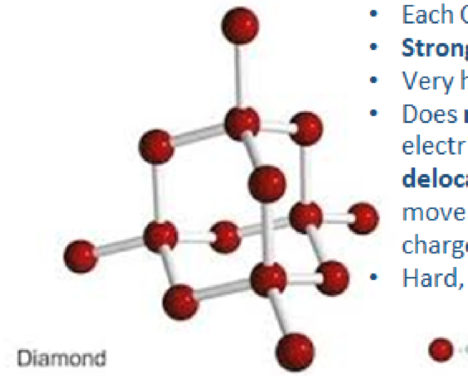
- Metals are **good conductors of electricity** because the **delocalised electrons** in the metal **carry electrical charge** through the metal.
- Metals are **good conductors of thermal energy** because **energy is transferred by the delocalised electrons.**

5.2.3 Structure & bonding of carbon



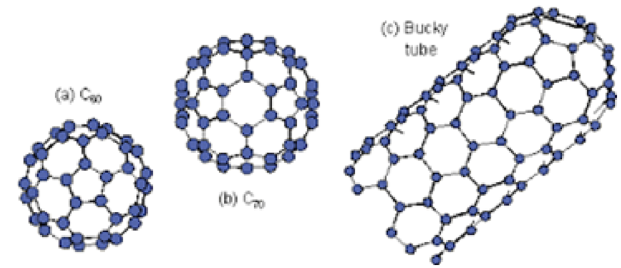
Graphite

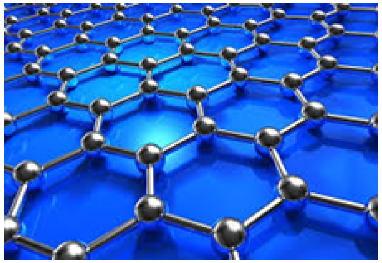
- Giant lattice (in layers)
- Each C forms **3 bonds**
- **Layers** of hexagonal rings with no bonds between layers
- **Weak forces** between layers to slide off easily (used as **lubricant**, graphite pencil)
- Giving **1 delocalised** electron
- Good conductor as has the delocalised electron can **move and carry the charge.**



Diamond: Giant lattice

- Each C forms **4 bonds**
- **Strong** covalent bonds
- Very high mpt/bpt
- Does **not** conduct electricity as has **no delocalised electrons** to move and carry the charge.
- Hard, used in **drill bits**





Graphene

- A **single** layer of graphite
- Useful in **electronics** and **composites**
- 3 covalent bonds leaving free electrons to conduct electricity

Fullerenes

- Hollow shapes
- Hexagonal rings, but may also contain rings of 5 or 7 Cs
- Buckminsterfullerene (C_{60}) spherical.
- Carbon nanotubes are **cylindrical**. Very useful for **nanotechnology, electronics**

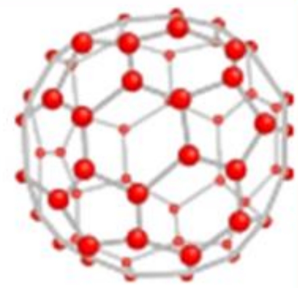
Note: carbon is a non-metal so the bonds between carbon atoms must be **COVALENT.**

Knowledge Organiser – Structure & Bonding Separate Chemistry

Nanoscience

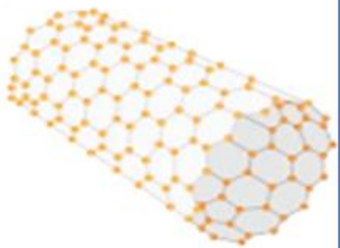
- Structures are 1 – 100nm in size
- A few hundred atoms big
- Smaller than fine particles ($PM_{2.5}$)
- Coarse particles (PM_{10}) are also called "dust"
- As the side of a cube DECREASES by a factor of 10, the surface area to volume ratio INCREASES by a factor of 10
- Nanoparticles have a high SA to V ratio so they have different properties than the same material in bulk.
- Smaller quantities are needed as they're more effective.

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



Buckminsterfullerene has the formula C_{60}

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



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

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

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

Knowledge Organiser – 5.2 Structure & bonding

5.2.1.1 Chemical bonds

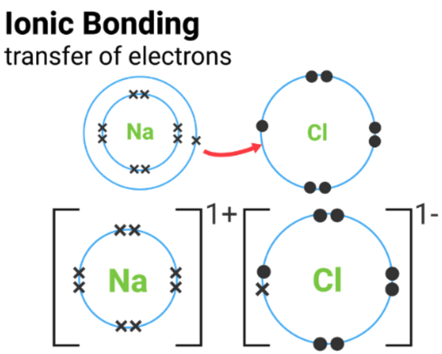
There are three types of strong chemical bonds: ionic, covalent and metallic.

Ionic bonding: particles are oppositely charged ions. Ionic bonding occurs in compounds formed from metals combined with non-metals

Covalent bonding the particles are atoms which share pairs of electrons. Covalent bonding occurs in most non-metallic elements and in compounds of non-metals

Metallic bonding the particles are atoms which share delocalised electrons. Metallic bonding occurs in metallic elements and alloys.

5.2.1.2 Ionic Bonding

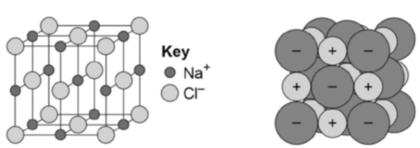


- Between a metal atom and a non-metal atom
- Metals lose electrons to form positive ions
- Non-metals gain electrons & form negative ions
- Electrons **transferred** (**ions** formed)
- Strong **electrostatic** forces
- Giant **lattice** structures
- High melting/boiling points
- If **molten** or in **solution** ions will conduct electricity

5.2.1.3 Ionic compounds

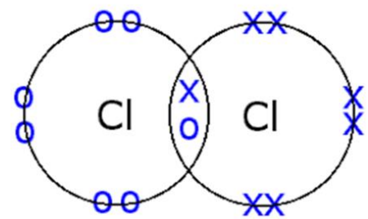
- An ionic compound is a **giant structure of ions**.
- Ionic compounds are held together by **strong electrostatic forces of attraction** between oppositely charged ions.
- These **forces act in all directions in the lattice** and this is called **ionic bonding**.

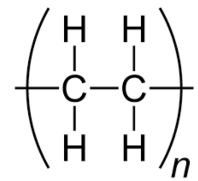
Eg: structure of sodium chloride

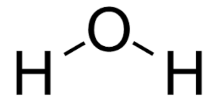


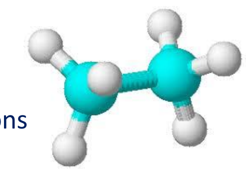
5.2.1.4 Covalent Bonding

Sharing electrons









- Between two non-metal atoms
- Electrons are **shared**
- A **covalent** bond is one pair of shared electrons
- Covalent bonds are ALWAYS STRONG

5.2.2.1. The three states of matter

- **Freezing** take place at the **melting point**
- **Boiling** and **condensing** take place at the **boiling point**.
- **Particle theory** can help to explain melting, boiling, freezing and condensing.
- The **amount of energy needed to change state** from solid to liquid and from liquid to gas **depends on the strength of the forces** between the particles of the substance.
- **The stronger the forces between the particles the higher the melting point and boiling point of the substance**.

5.2.2.3 Properties of Ionic compounds

- have **high melting points** and **high boiling points because** of the **large amounts of energy needed to break the many strong bonds.**
- When melted or dissolved in water, ionic compounds **conduct electricity because the ions are free to move and so charge can flow**.

5.2.2.4 Properties of Small molecules

- Usually **gases or liquids** with **low melting point & low boiling point.**
- Weak **intermolecular** forces (because they are small molecules) which are overcome when substance melts or boils.
- e.g. gases, water
- Do not conduct electricity as no overall electric charge.

5.2.2.5 Polymers

- **Long** molecules with atoms linked by **strong covalent bonds.**
- Solid at room temperature as **relatively strong intermolecular forces.**
- **Repeating** units

e.g. plastics

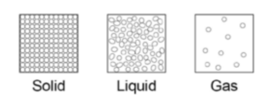
5.2.2.6 Giant covalent structures

Giant lattices

- High melting point and boiling point
- Strong covalent bonds which must be overcome to melt or boil.
- e.g. silicon dioxide, diamond, graphite

5.2.2.2. State symbols

In chemical equations, the three states of matter are shown as (s), (l) and (g).

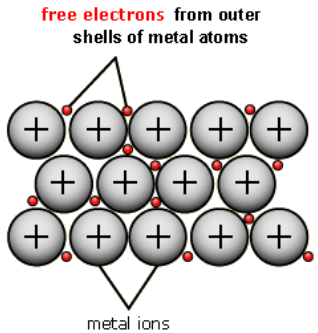


(aq) for aqueous solutions eg salt water or acid solutions.

Knowledge Organiser – 5.2 Structure & bonding

5.2.1.5 Metallic Bonding

- Bonding between atoms of a metal
- **Delocalised** electrons (negative) & metal ions (positive)
- Shared delocalised electrons form strong metallic bonds
- Delocalised electrons **conduct** heat and electricity



- Pure metals are soft: layers of atoms can slide over each other

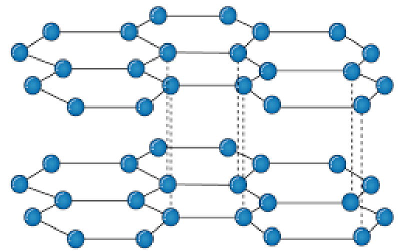
5.2.2.7 Properties of metals and alloys

- Metals have **giant structures of atoms** with **strong metallic bonding**. Therefore most metals have **high melting and boiling points.**
- In **pure metals, atoms are arranged in layers**, which allows metals to be **bent and shaped**. (malleable)
- **Pure metals** are too soft for many uses and so **are mixed with other metals to make alloys which are harder**.
- In **alloys**, different atoms **disrupt** the layers
- Alloys are **harder** than pure metals

5.2.2.8 Metals as conductors

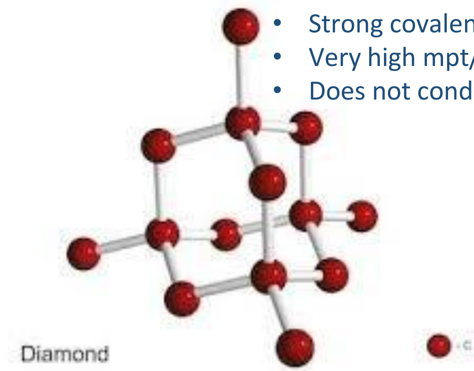
- Metals are **good conductors of electricity** because the **delocalised electrons** in the metal **carry electrical charge** through the metal.
- Metals are **good conductors of thermal energy** because **energy is transferred by the delocalised electrons**.

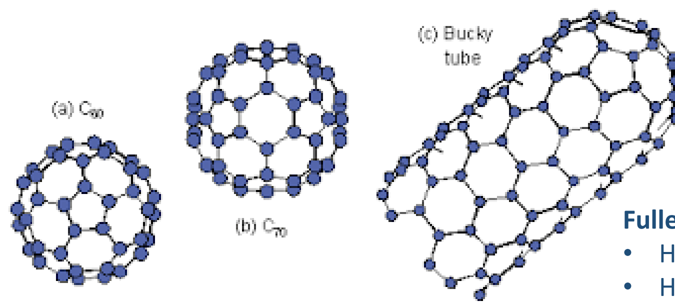
5.2.3 Structure & bonding of carbon



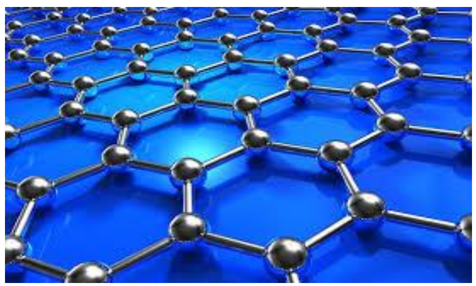
Graphite
- Giant lattice (in layers)
- Each C forms 3 bonds
- Layers of hexagonal rings with no bonds between layers
- Giving 1 delocalised electron
- Good conductor

Diamond: Giant lattice
- Each C forms 4 bonds
- Strong covalent bonds
- Very high mpt/bpt
- Does not conduct electricity





Fullerenes
- Hollow shapes
- Hexagonal rings, but may also contain rings of 5 or 7 Cs
- Buckminsterfullerene (C_{60}) spherical.
- Carbon nanotubes are cylindrical. Very useful for nanotechnology, electronics



Graphene
- A single layer of graphite
- Useful in electronics and composites

Note: carbon is a non-metal so the bonds between carbon atoms must be COVALENT.

Knowledge Organiser – 5.3 Quantitative Chemistry

5.3.1.1 Conservation of mass and balanced chemical equations

Reacting masses

In all chemical reactions the **total mass of reactants used is equal to the total mass of the products** made: Reactants ⟶ Products

5.3.1.2 Relative Formula Mass (M_r)

Relative atomic mass

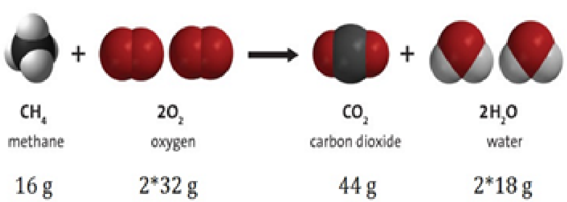
Different atoms have different masses.
Atoms have such a small mass it is more convenient to know their masses compared to each other.
Carbon is taken as the standard atom and has a relative atomic mass (A_r) of 12.

Relative formula mass

To find the relative formula mass (M_r) of a compound, you just add together the A_r values for all the atoms in its formula.

Example 1:
Find the M_r of carbon monoxide (CO).
The A_r of carbon is 12 and the A_r of oxygen is 16
So the M_r of carbon monoxide is 12 + 16 = 28.

Example 2:
Find the M_r of carbon dioxide (CO_2)
The A_r of carbon is 12 and the A_r of oxygen is 16, but there are 2 atoms of oxygen in the formula.
So the M_r of Carbon dioxide is 12 + 16 + 16 = 44



Mass of reactants (80g) Mass of products (80g)

5.3.1.3 Mass change when a reactant or product is a gas

Apparent loss of gain in mass when a **gas** is a product or reactant and is **gained** or **released** to the atmosphere in an **non-enclosed system.**

5.3.1.4 Chemical Measurements

Measurements have uncertainty.
You need to be able to look at the range of measurements about the mean (average) as a measure of uncertainty.

5.3.2.5 Limiting reactants (HT only)

- In a chemical reaction involving two reactants, it is common to use an <u>excess</u> of one of the reactants to <u>**ensure**</u> that all of the other reactant is <u>**used**</u>.
- The **reactant that is completely used up** is called the **limiting reactant** because it limits the amount of products.

5.3.2 Amounts of substances in relation to masses of pure substances (HT only)

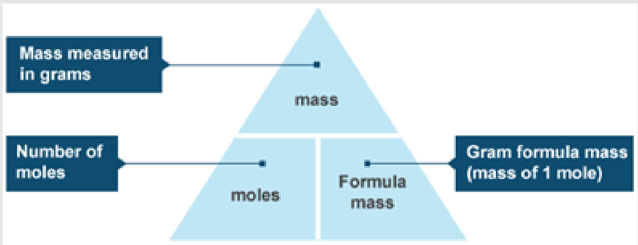
The Mole:

- The unit for amount of substance is called the **mole**, shown as **mol**. One mole of atoms, *ions* or *molecules* is around 6×10^{23} (6 followed by 23 zeroes). This is called the Avogadro constant.
- This is the same number as the number of carbon atoms in 12 g of carbon.

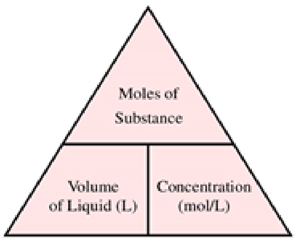
This equation shows how **molar mass, number of moles** and **mass** are related:

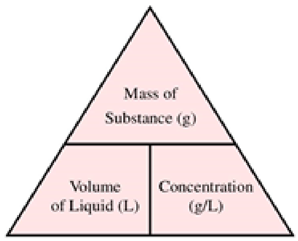
number of moles = mass ÷ molar mass

This can be rearranged to find the mass if the number of moles and molar mass are known, or to find the molar mass if the mass and number of moles are known.



Finding the number of moles
Example
What is the number of moles of carbon dioxide molecules in 22 g of CO_2?
A_r of C = 12, A_r of O = 16
The relative formula mass M_r of carbon dioxide = 12 + 16 + 16 = 44
This means that the molar mass of carbon dioxide = 44 g/mol
number of moles = 22 ÷ 44 = 0.5 mol





Knowledge Organiser – 5.3 Quantitative Chemistry

Spec	Question	Answer
5.3.1.1	What is the law of conservation of mass?	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.
5.3.1.1	What does the conservation of mass mean in terms of chemical reactions?	This means that chemical reactions can be represented by symbol equations which are balanced in terms of the numbers of atoms of each element involved on both sides of the equation.
5.3.1.2	What is the relative formula mass (M_r) of a compound?	The relative formula mass (M_r) of a compound is the sum of the relative atomic masses of the atoms in the numbers shown in the formula
5.3.1.2	What happens to the sum of the relative formula masses of the reactants & products?	The sum of the relative formula masses of the reactants in the quantities shown equals the sum of the relative formula masses of the products in the quantities shown.
5.3.1.3	How can we explain a change in mass?	This can usually be explained because a reactant or product is a gas and its mass has not been taken into account.
5.3.1.3	Give 2 examples of reactions where there appears to be a change in mass	• when a metal reacts with oxygen the mass of the oxide produced is greater than the mass of the metal • thermal decompositions of metal carbonates carbon dioxide is produced and escapes into the atmosphere leaving the metal oxide as the only solid product.
5.3.1.4	When there is uncertainty about a result, what 2 things should you do?	• represent the distribution of results and make estimations of uncertainty • use the range of a set of measurements about the mean as a measure of uncertainty.
5.3.2.1	What are chemical amounts measured in and what is its unit?	Chemical amounts are measured in moles. The symbol for the unit mole is mol.

Spec	Question	Answer
5.3.2.1 HT	What is the mass of one mole equal to?	The mass of one mole of a substance in grams is numerically equal to its relative formula mass. One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance.
5.3.2.1 HT	What is Avogadro's number, including its value?	The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant. The value of the Avogadro constant is 6.02×10^{23} per mole
5.3.2.2 HT	How many moles of reactants and products in: $Mg + 2HCl\ MgCl_2 + H_2$	one mole of magnesium reacts with two moles of hydrochloric acid to produce one mole of magnesium chloride and one mole of hydrogen gas.
5.3.2.3 HT	How are the balancing numbers in a symbol equation calculated?	The balancing numbers in a symbol equation can be calculated from the masses of reactants and products by converting the masses in grams to amounts in moles and converting the numbers of moles to simple whole number ratios.
5.3.2.4 HT	What is a limiting reactant and how does the limiting reactant affect the amount of products produced?	The reactant that is completely used up is called the limiting reactant because it limits the amount of products. The effect of a limiting quantity of a reactant on the amount of products it is possible to obtain in terms of amounts in moles or masses in grams.
5.3.2.5 HT	How is the concentration of a solution measured?	The concentration of a solution can be measured in mass per given volume of solution, e.g. grams per dm^3 (g/dm^3).
5.3.3.1 HT	Why is it not always possible to obtain the calculated amount of product?	• the reaction may not go to completion because it is reversible • some of the product may be lost when it is separated • some of the reactants may react in ways different to the expected reaction.
5.3.3.1 HT	How do you calculate percentage yield?	$\% \text{ Yield} = \frac{\text{Mass of product actually made}}{\text{Maximum theoretical mass of product}} \times 100$
5.3.3.2 HT	How is percentage atom economy calculated?	The percentage atom economy of a reaction is calculated using the balanced equation for the reaction as follows: $= \frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula masses of all reactants from equation}} \times 100$
5.3.4 HT	What information do you need to calculate the concentration of a soln?	If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated.
5.3.5 HT	What is the volume of one mole of any gas at room temp and pressure?	. The volume of one mole of any gas at room temperature and pressure (20ºC and 1 atmosphere pressure) is 24 dm^3

Knowledge Organiser – 5.3 Quantitative Chemistry (Separate Only)

4.3.3.1 Percentage Yield

The law of conservation of mass states that no atoms are gained or lost in a chemical reactions, BUT, it is not always possible to obtain calculated amounts of a product because:

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

4.3.3.1 Percentage Yield is calculated using the equation:

$$\% \text{ yield} = \frac{\text{mass of product actually made}}{\text{maximum theoretical mass of product}} \times 100$$

4.3.3.1 Example of % yield

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

Answer:

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

% yield of $CuSO_4$ crystals = **90%**

4.3.3.2 Reaction Pathways (HT)

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

- *Percentage yield*
- *Atom economy*
- *Rate of reaction*
- *Position of the equilibrium*
- *Usefulness/ toxicity of any by-product*

The **raw materials** needed for a particular reactions may affect its chosen pathway. E.g. crude oil is a **finite** resource, however plant sugars are **renewable** and can be replenished as long as other plants are replanted. Pathways involving renewable resources are **favourable**.

4.3.3.2 Atom Economy

The % atom economy can be calculated using the equation:

$$\frac{M_r \text{ of } \textbf{desired product} \text{ from equation}}{\text{Sum of } M_r \text{ of } \textbf{ALL reactants} \text{ from equation}} \times 100$$

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

4.3.3.2 Atom Economy e.g:

When glucose (M_r = **180**) is fermented, ethanol (M_r =**46**) is made.

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

Calculate the atom economy for this reaction if ethanol is what the manufacturer requires.

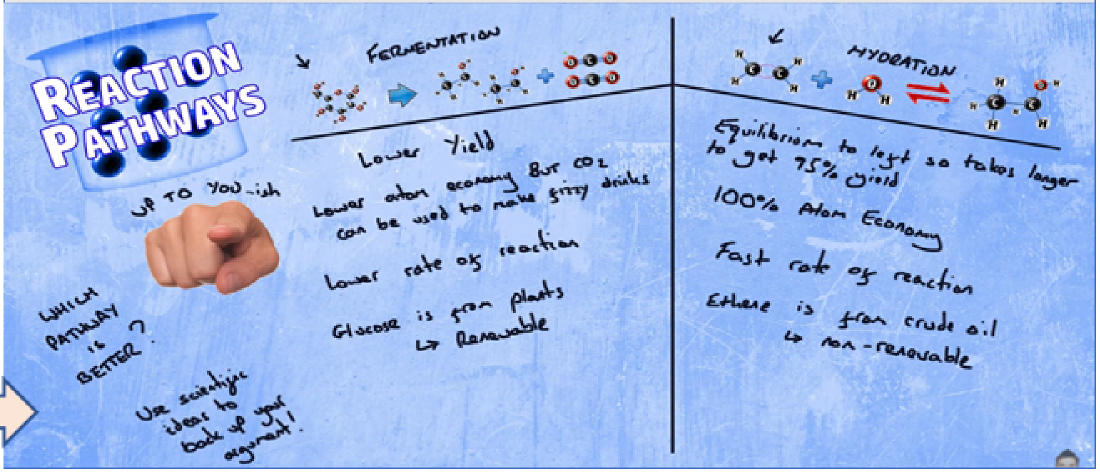
Answer:

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

Atom economy = **51.1%**

Example in Context: **Ethanol** can be made through **fermentation** of glucose or the **hydration** of ethene. Use the data in the table and your own knowledge to ***evaluate*** the 2 methods and ***conclude*** which is the best method for ethanol production? (6)

Method	% yield	Atom economy (%)	Rate of reaction
Fermentation	15	51.1	low
hydration	95	100	high



Knowledge Organiser – 5.3 Quantitative Chemistry (Separate Only)

4.3.5 Volume of gases (HT)

The volume of one mole of **ANY** gas at room temperature and pressure (20^0C and 1 atmospheric pressure) is **$24 dm^3$!!**

Conversion: $24 dm^3$ → **x 1000** → $24\ 000 cm^3$

To calculate a known volume of gas:

Volume = amount in moles x molar volume

E.g. Determine the volume of 0.55 mol of carbon monoxide (CO) at room temp and pressure.

Answer:

Volume = amount in mol x molar volume

= 0.55 mol x 24 dm^3

= <u>13.2 dm^3</u>

4.3.5 Rearranging the equation:

Number of moles of gas = $\frac{\text{volume of gas } (dm^3)}{24\ (dm^3)}$

OR = $\frac{\text{volume of gas } (cm^3)}{24\ 000\ cm^3}$

1) How many **moles** of gas are found in 48 dm^3 of $CO_{2(g)}$?

Moles = $48 dm^3$ / $24 dm^3$ = 2 moles

2) Calculate the **volume** of gas (in cm^3) in 1.5 moles of N_2O_4

Volume = 1.5 mol x 24 000 cm^3 = **$36000 cm^3$**

E.g. 1) Determine the amount of hydrogen gas that occupies $198 cm^3$ at room temp and pressure.

Number of moles of gas = $\frac{\text{volume of gas } (dm^3)}{\text{molar volume } (dm^3)}$

Amount in mol = $\frac{198\ cm^3}{24\ 000 cm^3}$ = **<u>0.0083 mol</u>**

Calculating a <u>volume</u> from a mass (HT)

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

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

a) Determine the molar amount of sodium (A_r = 23)

Amount in mol = $\frac{\text{mass (g)}}{\text{atomic mass}}$

= 3.5g / 23

= **<u>0.15 mol</u>** of sodium

b) Determine the molar amount of hydrogen.

According to the equation, the molar ratio of sodium : hydrogen is **2 : 1**

Therefore, 0.15 mol of Na produces **0.075 mol** of hydrogen gas.

c) Determine the volume of hydrogen gas.

volume = amount in mol x molar volume

= 0.075 mol x $24 dm^3$

<u>= 1.8 dm^3</u>

Molar ratios: $H_2 + Cl_2 \rightarrow 2HCl$

When hydrogen reacts with chlorine, hydrogen chloride is produced. In terms of the molar ratio, **$10 cm^3$ of hydrogen would react with 10 cm^3** chlorine:- a molar ratio of **1:1**.

The molar ratio between hydrogen and the hydrogen chloride is, however, **1 : 2**.

So $10 cm^3$ of hydrogen would make $20 cm^3$ of hydrogen chloride.

Calculating a <u>mass</u> from a volume (HT)

Sodium reacts with chlorine to produce sodium chloride.

$$2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2NaCl_{(s)}$$

Determine the mass of sodium chloride (M_r **58.5**) That can be produced from **$685 cm^3$** of chlorine.

Amount of chlorine in $685 cm^3$ is:

Number of moles of gas = $\frac{\text{volume of gas } (cm^3)}{24\ 000\ cm^3}$

mol of chlorine = $\frac{685\ cm^3}{24\ 000\ cm^3}$

= <u>0.029 mol.</u>

Mole ratio of chlorine : sodium is **1 : 2**

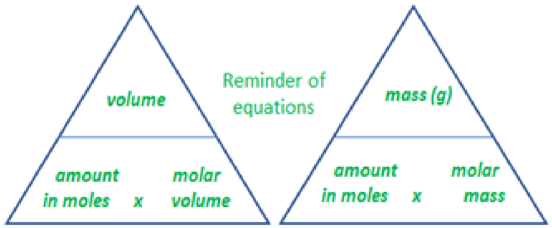
Therefore 0.029 mol of chlorine would produce (0.029 mol x 2) = **0.058 mol** of sodium chloride.

Mass of NaCl = moles of NaCl x Mr of NaCl

= 0.058 mol x 58.5

<u>= 3.393g</u>

<u>Answer</u>: <u>3.393g</u> of sodium chloride will be made from 686cm3 of chlorine gas.



Knowledge Organiser – 5.3 Quantitative Chemistry

5.3.1.1 Conservation of mass and balanced chemical equations

Reacting masses

In all chemical reactions the **total mass of reactants used is equal to the total mass of the product**s made: Reactants ⟶ Products

5.3.1.2 Relative Formula Mass (M_r)

Relative atomic mass

Different atoms have different masses.

Atoms have such a small mass it is more convenient to know their masses compared to each other.

Carbon is taken as the standard atom and has a relative atomic mass (A_r) of 12.

Relative formula mass

To find the relative formula mass (M_r) of a compound, you just add together the A_r values for all the atoms in its formula.

Example 1:

Find the M_r of carbon monoxide (CO).

The A_r of carbon is 12 and the A_r of oxygen is 16

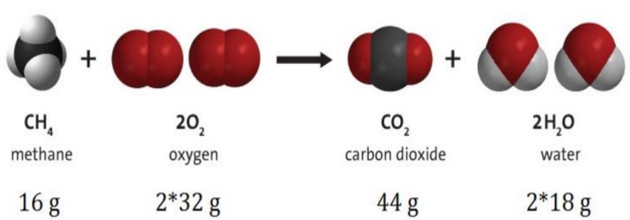
So the M_r of carbon monoxide is 12 + 16 = 28.

Example 2:

Find the M_r of carbon dioxide (CO2)

The A_r of carbon is 12 and the A_r of oxygen is 16, but there are 2 atoms of oxygen in the formula.

So the M_r of Carbon dioxide is 12 + 16 + 16 = 44



Mass of reactants (80g) Mass of products (80g)

5.3.1.3 Mass change when a reactant or product is a gas

Apparent loss of gain in mass when a **gas** is a product or reactant and is **gained** or **released** to the atmosphere in an **non-enclosed system**.

5.3.1.4 Chemical Measurements

Measurements have uncertainty.

You need to be able to look at the range of measurements about the mean (average) as a measure of uncertainty.

5.3.2.5 Limiting reactants (HT only)

- In a chemical reaction involving two reactants, it is common to use an excess of one of the reactants to ensure that all of the other reactant is used.
- The **reactant that is completely used up** is called the **limiting reactant** because it limits the amount of products.

5.3.2 Amounts of substances in relation to masses of pure substances (HT only)

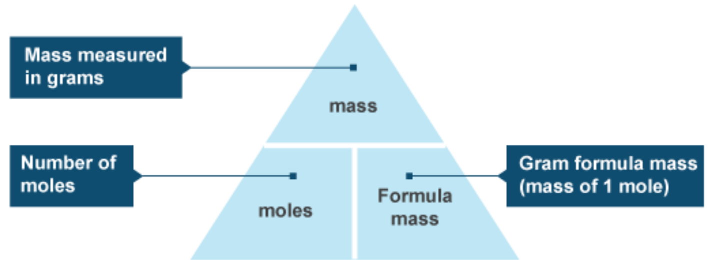
The Mole:

- The unit for amount of substance is called the **mole**, shown as **mol**. One mole of atoms, *ions* or *molecules* is around 6×10^{23} (6 followed by 23 zeroes). This is called Avogadro constant.
- This is the same number as the number of carbon atoms in 12 g of carbon.

This equation shows how **molar mass**, **number of moles** and **mass** are related:

number of moles = mass ÷ molar mass

This can be rearranged to find the mass if the number of moles and molar mass are known, or to find the molar mass if the mass and number of moles are known.



Finding the number of moles

Example

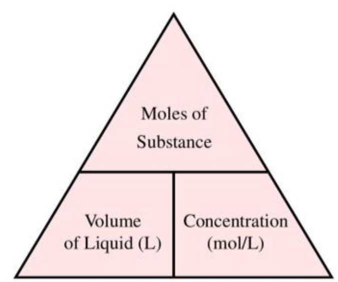
What is the number of moles of carbon dioxide molecules in 22 g of CO2?

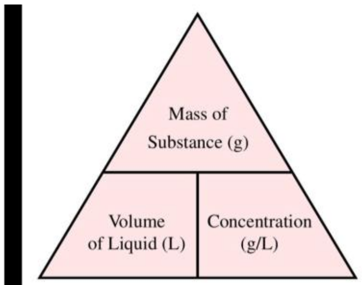
A_r of C = 12, A_r of O = 16

The relative formula mass M_r of carbon dioxide = 12 + 16 + 16 = 44

This means that the molar mass of carbon dioxide = 44 g/mol

number of moles = 22 ÷ 44 = 0.5 mol





Knowledge Organiser – 5.3 Quantitative Chemistry

Spec	Question	Answer
5.3.1.1	What is the law of conservation of mass?	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.
5.3.1.1	What does the conservation of mass mean in terms of chemical reactions?	This means that chemical reactions can be represented by symbol equations which are balanced in terms of the numbers of atoms of each element involved on both sides of the equation.
5.3.1.2	What is the relative formula mass (Mr) of a compound?	The relative formula mass (Mr) of a compound is the sum of the relative atomic masses of the atoms in the numbers shown in the formula
5.3.1.2	What happens to the sum of the relative formula masses of the reactants & products?	The sum of the relative formula masses of the reactants in the quantities shown equals the sum of the relative formula masses of the products in the quantities shown.
5.3.1.3	How can we explain a change in mass?	This can usually be explained because a reactant or product is a gas and its mass has not been taken into account.
5.3.1.3	Give 2 examples of reactions where there appears to be a change in mass	• when a metal reacts with oxygen the mass of the oxide produced is greater than the mass of the metal • thermal decompositions of metal carbonates carbon dioxide is produced and escapes into the atmosphere leaving the metal oxide as the only solid product.
5.3.1.4	When there is uncertainty about a result, what 2 things should you do?	• represent the distribution of results and make estimations of uncertainty • use the range of a set of measurements about the mean as a measure of uncertainty.
5.3.2.1	What are chemical amounts measured in and what is its unit?	Chemical amounts are measured in moles. The symbol for the unit mole is mol.

Spec	Question	Answer
5.3.2.1 HT	What is the mass of one mole equal to?	The mass of one mole of a substance in grams is numerically equal to its relative formula mass. One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance.
5.3.2.1 HT	What is Avogadros number, including its value?	The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant. The value of the Avogadro constant is 6.02×10^{23} per mole
5.3.2.2 HT	How many moles of reactants and products in: Mg + 2HCl MgCl2 + H2	one mole of magnesium reacts with two moles of hydrochloric acid to produce one mole of magnesium chloride and one mole of hydrogen gas.
5.3.2.3 HT	How are the balancing numbers in a symbol equation calculated?	The balancing numbers in a symbol equation can be calculated from the masses of reactants and products by converting the masses in grams to amounts in moles and converting the numbers of moles to simple whole number ratios.
5.3.2.4 HT	What is a limiting reactant and how does the limiting reactant affect the amount of products produced?	The reactant that is completely used up is called the limiting reactant because it limits the amount of products. The effect of a limiting quantity of a reactant on the amount of products it is possible to obtain in terms of amounts in moles or masses in grams.
5.3.2.5 HT	How is the concentration of a solution measured?	The concentration of a solution can be measured in mass per given volume of solution, eg grams per dm3 (g/dm3).
5.3.3.1 HT	Why is it not always possible to obtain the calculated amount of product?	• the reaction may not go to completion because it is reversible • some of the product may be lost when it is separated • some of the reactants may react in ways different to the expected reaction.
5.3.3.1 HT	How do you calculate percentage yield?	$\% \text{ Yield} = \dfrac{\text{Mass of product actually made}}{\text{Maximum theoretical mass of product}} \times 100$
5.3.3.2 HT	How is percentage atom economy calculated?	The percentage atom economy of a reaction is calculated using the balanced equation for the reaction as follows: $= \dfrac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula masses of all reactants from equation}} \times 100$
5.3.4 HT	What information do you need to calculate the concentration of a soln?	If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated.
5.3.5 HT	What is the volume of one mole of any gas at room temp and pressure?	. The volume of one mole of any gas at room temperature and pressure (20oC and 1 atmosphere pressure) is 24 dm3

Knowledge Organiser – 5.4 Chemical Changes

5.4.1 Reactivity of metals

Metals react with oxygen to produce **metal oxides.**
The reactions are **oxidation reactions because the metals gain oxygen.**

5.4.1.2 The reactivity series

- 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.
- A more reactive metal can displace a less reactive metal from a compound.

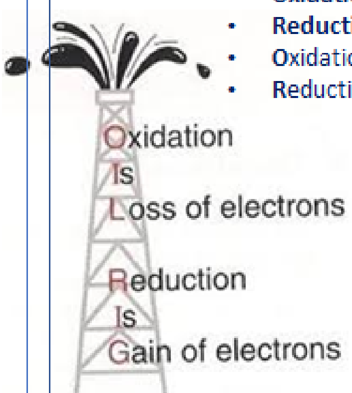
potassium	**most reactive**	**K**
sodium		**Na**
calcium		**Ca**
magnesium		**Mg**
aluminium		**Al**
carbon		**C**
zinc		**Zn**
iron		**Fe**
tin		**Sn**
lead		**Pb**
hydrogen		**H**
copper		**Cu**
silver		**Ag**
gold		**Au**
platinum	**least reactive**	**Pt**

5.4.1.3 Extraction of metals and reduction

- **Unreactive** metals are found as pure elements (e.g. gold) but most are compounds.
- Those below carbon can be **extracted** from oxides using carbon.
- Those above carbon need to be extracted using **electrolysis**.

5.4.1.4 Oxidation and reduction in terms of electrons (HT only)

- **Oxidation** involves gain of oxygen
- **Reduction** involves loss of oxygen
- Oxidation Is Loss of electrons
- Reduction Is Gain of electrons



5.4.2.1 Reactions of acids with metals

Acids react with some metals to produce salts and hydrogen.

acid + metal → salt + hydrogen
acid + alkali → salt + water
acid + carbonate → salt + water + carbon dioxide

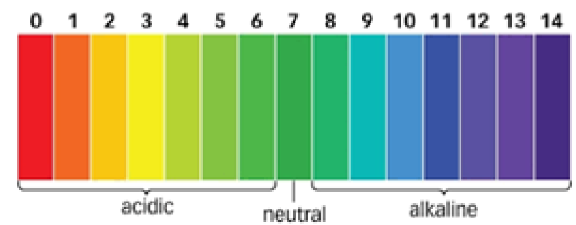
HCl - hydrochloric acid produces chlorides
HNO_3 - nitric acid produces nitrates
H_2SO_4 - sulfuric acid produces sulfates

(HT only)

- explain in terms of gain or loss of electrons, these are redox reactions
- identify which species are oxidised and which are reduced in given chemical equations.

5.4.2.2 Neutralisation of acids and salt production

- Acids are neutralised by alkalis (eg soluble metal hydroxides) and bases (eg insoluble metal hydroxides and metal oxides)



5.4.2.4. The pH scale and neutralisation

- **Acids produce hydrogen ions (H+)** in aqueous solutions.
- Aqueous **solutions of alkalis contain hydroxide ions (OH–).**
- The **pH scale**, from 0 to 14, is a measure of the **acidity** or **alkalinity** of a solution, and can be measured using **universal indicator** or a **pH probe**.
- A solution with pH 7 is neutral.
- In **neutralisation reactions** between an acid and an alkali, **hydrogen ions react with hydroxide ions to produce water.**

HIGHER TIER

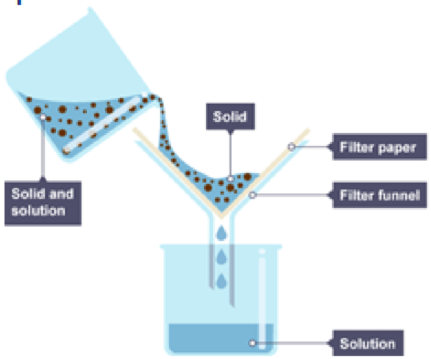
Strong acids (HCl, HNO_3, H_2SO_4) fully **ionise**
Weak acids (ethanoic, citric, carbonic) **partially** ionise

Knowledge Organiser – 5.4 Chemical Changes

5.4.2.3 Soluble salts

- Soluble salts can be made from **acids** by reacting them with solid insoluble substances, such as **metals, metal oxides, hydroxides or carbonates.**
- The solid is added **in excess** to the acid until no more reacts and the **excess solid is filtered** off to produce a **solution** of the **salt.**
- Salt solutions can be **crystallised** to produce solid salts. **Pat crystals dry** with a paper towel.

RPA Preparation of a soluble salt



1. Add excess solid to acid
2. React
3. Filter off unreacted solid

4. Warm to evaporate
5. Allow to crystallise and dry

5.4.3. Electrolysis

5.4.3.1 The process of electrolysis

- Ionic compounds can be electrolysed when liquid or molten, as the ions are then **free to move and carry the charge**
- An electric current is passed through the **electrolyte**
- **Positive** ions move to the negative electrode (**cathode**)
- **Negative** ions move to the positive electrode (**anode**)
- Aluminium is extracted by electrolysis from a mixture of **aluminium oxide** and **cryolite**

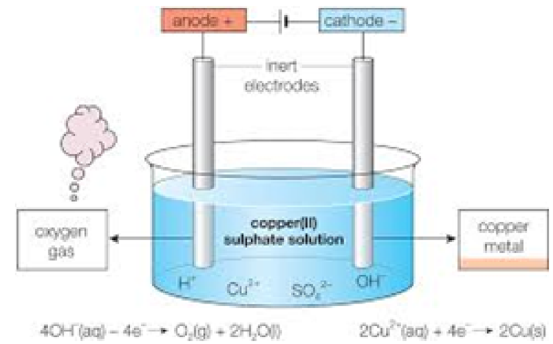
5.4.3.2 Electrolysis of molten ionic compounds

- When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes
- the metal (lead) is produced at the cathode
- the non-metal (bromine) is produced at the anode.

5.4.3.3 Using electrolysis to extract metals

- **Metals** can be **extracted** from molten compounds using **electrolysis.**
- Electrolysis is used **if the metal is too reactive to be extracted by reduction** with carbon or if the metal reacts with carbon.
- **Large amounts of energy** are used in the extraction process to melt the compounds and to produce the electrical current.
- **Aluminium** is manufactured by the **electrolysis** of a **molten mixture of aluminium oxide and cryolite using carbon** as the **positive** electrode (**anode**).

RPA Electrolysis of aqueous solution



5.4.3.4. Electrolysis of aqueous solutions

- The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.
- At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen.
- At the positive electrode (anode), oxygen is produced unless the solution contains halide ions when the halogen is produced.
- This happens because in the aqueous solution water molecules break down producing hydrogen ions and hydroxide ions that are discharged.

5.4.3.5 Reactions at electrodes as half equations (HT)

During **electrolysis**, at the **cathode** (negative electrode), **positively charged ions gain electrons.** Ie **reductions reactions**
At the **anode** (positive electrode), **negatively charged ions lose electrons.** I.e. **oxidations.**
Reactions at electrodes can be represented by half equations,

E.g. $2H^+ + 2e^- \rightarrow H_2$

and

$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ or

$4OH^- - 4e^- \rightarrow O_2 + 2H_2O$

Knowledge Organiser – 5.4 Chemical Changes (Separate Chemistry)

Titration Method - RPA

A Volumetric pipette is used to measure out a fixed volume of solution

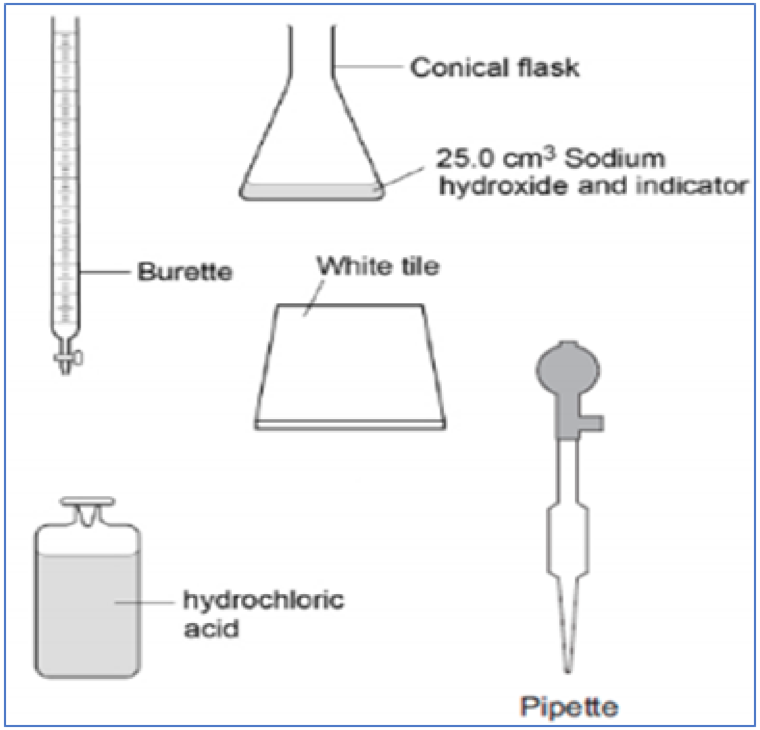
A burette is used to measure the volume of the solution added

- Wash a **volumetric pipette** with **distilled water** followed by some of the **alkali**
- Using the **pipette**, measure **25cm^3 alkali** (eg **NaOH**) into a conical flask
- Add a few drops of indicator (eg **phenolphthalein**) to the solution in the conical flask and swirl. The mixture should turn **pink**.
- Place a **white tile** under the flask
- Rinse a **burette** with **distilled water** followed by some of the **acid**, allowing some of the acid to pass **through the tap** (filling the **jet**)
- Fill the burette up to the mark using the **acid** (eg **hydrochloric acid**) & place **over the conical flask**
- Record **initial reading** on the **burette**
- Open tap to **slowly** release acid into the conical flask whilst **swirling**
- Keep on repeating this until the indicator permanently changes colour (**end point**) (eg phenolphthalein goes **colourless**)
- Record **final volume** reading on the **burette** by reading the bottom of the **meniscus**.
- Work out the volume of acid (**titre**) that was run into the flask
- Repeat the whole process at **least three times** until you get **concordant** titres (within **0.1cm^3** of each other)
- Calculate the **mean** titre
- Use results to calculate **concentration of the alkali in mol/dm^3**

The pipette allows the same volume of acid to be added each time, helping to make the results repeatable.

Accuracy

- Make sure the burette is vertical.
- Take the readings from the bottom of the meniscus.
- Near to the end-point, rinse the inside of the flask with distilled water.
- Add the acid drop by drop.



Knowledge Organiser – 5.4 Chemical Changes (Separate Chemistry)

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

Titration calculations RPA – Past paper question 1:
A student titrated hydrochloric acid with **25cm^3** of **0.10 mol/dm^3** sodium hydroxide solution. The equation for the titration is:

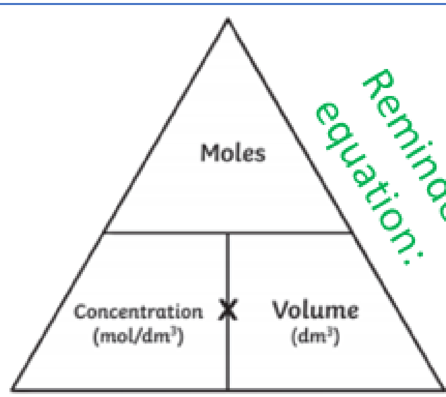
$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

	Titre 1	Titre 2	Titre 3	Titre 4	Titre 5
Volume HCl cm^3	13.60	12.10	11.10	12.15	12.15

Use concordant results in the table to calculate:
a) The **mean** titre
b) **Concentration** of the hydrochloric acid solution

Answer:
a) Concordant results are those within 0.10 cm^3 of each other.

Mean titre = $\frac{12.10 + 12.15 + 12.15}{3}$ = **12.13**



Reminder of equation:

Draw a table like this and fill it in with the information you know from the question. Change cm^3 to dm^3.

	ACID	ALKALI
No. of moles		
Conc. (mol/dm^3)		
volume (dm^3)		

Answer

b)	ACID	ALKALI
No. of moles	Step 2	Step 1
Conc. (mol/dm^3)	Step 3	0.1
volume (dm^3)	Mean= 0.01213	0.025

Step 1- work out moles of alkali:
Moles = conc x volume = 0.1 x 0.025 = 0.0025 mol

Step 2 – work out moles of other reactant:
Mole ratio of 1NaOH to 1HCl so same number = 0.0025mol

Step 3- work out con of other reactant:
Conc = $\frac{\text{moles}}{\text{volume}}$ = $\frac{0.0025 \text{ mol}}{0.01213 \text{ dm}^3}$ = **0.206 mol/ dm^3**

b)	ACID	ALKALI
No. of moles	0.0025 mol	0.0025 mol
Conc. (mol/dm^3)	**0.206**	0.1
volume (dm^3)	Mean= 0.01213	0.025

Knowledge Organiser – 5.4 Chemical Changes (Separate Chemistry)

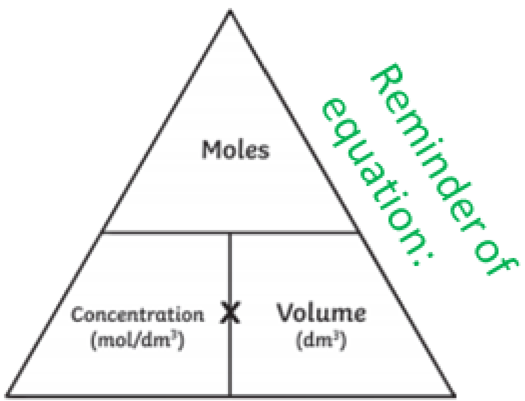
Titration calculations RPA – Past paper question 2:
A student titrated **20cm³** of **1.0 mol/dm³** sulfuric acid with **25cm³** sodium hydroxide solution. The equation for the titration is:

$$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$$

What was the concentration of **sodium hydroxide**?

Draw a table like this and fill it in with the information you know from the question. Change cm³ to dm³.

	ACID	ALKALI
No. of moles		
Conc. (mol/dm³)	1.0	
volume (dm³)	**20cm³**	**25cm³**



Answer

	ACID	ALKALI
No. of moles	Step 1	Step 2
Conc. (mol/dm³)	1.0	Step 3
volume (dm³)	0.020	0.025

Step 1- work out moles of acid:
Moles = conc x volume = 1.0 x 0.20 = 0.02 mol

Step 2 – work out moles of other reactant:
Mole ratio of 1acid : 2 alkali so = 0.04 mol

Step 3- work out con of other reactant:

$$\text{Conc} = \frac{\text{moles}}{\text{volume}} = \frac{0.04\ \text{mol}}{0.025\text{dm}^3} = \mathbf{1.6\ mol/\ dm^3}$$

	ACID	ALKALI
No. of moles	0.02 mol	0.04 mol
Conc. (mol/dm³)	1.0	**1.6**
volume (dm³)	0.020	0.025

Knowledge Organiser – 5.4 Chemical Changes

5.4.1 Reactivity of metals

Metals react with oxygen to produce **metal oxides**.
The reactions are **oxidation reactions because the metals gain oxygen**.

5.4.1.2 The reactivity series

- 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.
- A more reactive metal can displace a less reactive metal from a compound.

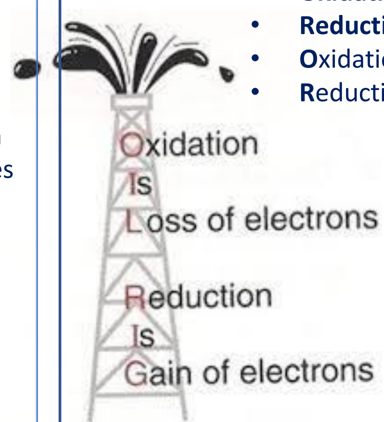
potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

5.4.1.3 Extraction of metals and reduction

- **Unreactive** metals are found as pure elements (eg gold) but most are compounds.
- Those below carbon can be **extracted** from oxides using carbon.
- Those above carbon need to be extracted using **electrolysis**.

5.4.1.4 Oxidation and reduction in terms of electrons (HT only)

- **Oxidation** involves gain of oxygen
- **Reduction** involves loss of oxygen
- **O**xidation **I**s **L**oss of electrons
- **R**eduction **I**s **G**ain of electrons



5.4.2.1 Reactions of acids with metals

Acids react with some metals to produce salts and hydrogen.

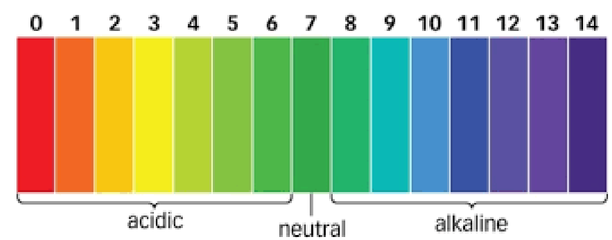
acid + metal → salt + hydrogen
acid + alkali → salt + water
acid + carbonate → salt + water + carbon dioxide

HCl - hydrochloric acid produces chlorides
HNO_3 - nitric acid produces nitrates
H_2SO_4 - sulfuric acid produces sulfates

(HT only)
- explain in terms of gain or loss of electrons, these are redox reactions
- identify which species are oxidised and which are reduced in given chemical equations.

5.4.2.2 Neutralisation of acids and salt production

- Acids are neutralised by alkalis (eg soluble metal hydroxides) and bases (eg insoluble metal hydroxides and metal oxides)



5.4.2.4. The pH scale and neutralisation

- **Acids produce hydrogen ions (H+)** in aqueous solutions.
- Aqueous **solutions of alkalis contain hydroxide ions (OH–).**
- The **pH scale**, from 0 to 14, is a measure of the **acidity** or **alkalinity** of a solution, and can be measured using **universal indicator** or a **pH probe**.
- A solution with pH 7 is neutral.
- In **neutralisation reactions** between an acid and an alkali, **hydrogen ions react with hydroxide ions to produce water.**

HIGHER TIER

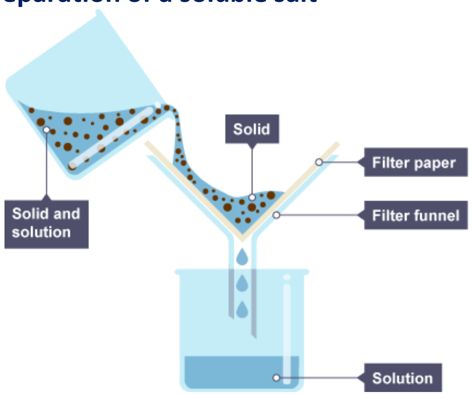
Strong acids (HCl, HNO_3, H_2SO_4) fully **ionise**
Weak acids (ethanoic, citric, carbonic) partially ionise

Knowledge Organiser – 5.4 Chemical Changes

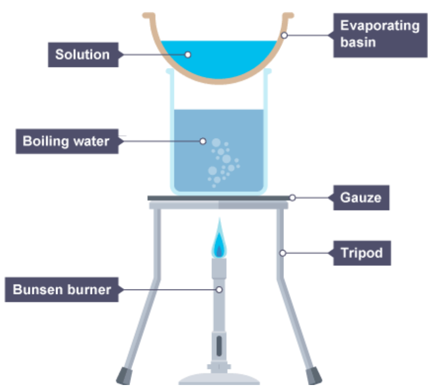
5.4.2.3 Soluble salts

- Soluble salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates.
- The solid is added to the acid until no more reacts and the excess solid is filtered off to produce a solution of the salt.
- Salt solutions can be crystallised to produce solid salts.

RPA Preparation of a soluble salt



1. Add excess solid to acid
2. React
3. Filter off unreacted solid



4. Warm over water bath then leave to evaporate
5. Allow to crystallise. Dry the pure crystals

5.4.3. Electrolysis

5.4.3.1 The process of electrolysis

- Ionic compounds can be electrolysed when liquid or molten, as the ions are then free to move
- An electric current is passed through the electrolyte
- Positive ions move to the negative electrode (cathode)
- Negative ions move to the positive electrode (anode)
- Aluminium is extracted by electrolysis from a mixture of aluminium oxide and cryolite

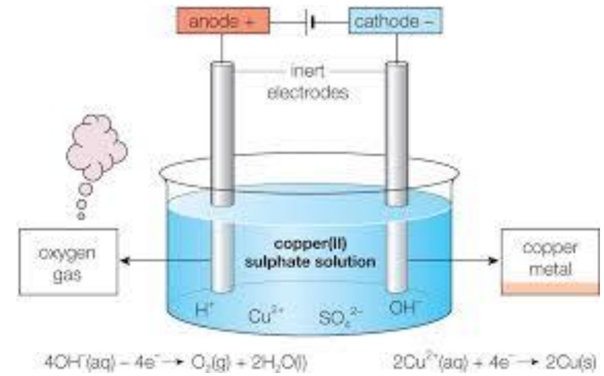
5.4.3.2 Electrolysis of molten ionic compounds

- When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes
- the metal (lead) is produced at the cathode
- the non-metal (bromine) is produced at the anode.

5.4.3.3 Using electrolysis to extract metals

- **Metals** can be **extracted** from molten compounds using **electrolysis**.
- Electrolysis is used **if the metal is too reactive to be extracted by reduction** with carbon or if the metal reacts with carbon.
- **Large amounts of energy** are used in the extraction process to melt the compounds and to produce the electrical current.
- **Aluminium** is manufactured by the **electrolysis** of a **molten mixture of aluminium oxide and cryolite using carbon** as the **positive** electrode (**anode**).

RPA Electrolysis of aqueous solution



5.4.3.4. Electrolysis of aqueous solutions

- The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.
- At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen.
- At the positive electrode (anode), oxygen is produced unless the solution contains halide ions when the halogen is produced.
- This happens because in the aqueous solution water molecules break down producing hydrogen ions and hydroxide ions that are discharged.

5.4.3.5 Representation of reactions at electrodes as half equations (HT only)

During **electrolysis**, at the **cathode** (negative electrode), **positively charged ions gain electrons**. Ie **reductions reactions** At the **anode** (positive electrode), **negatively charged ions lose electrons**. Ie **oxidations**. Reactions at electrodes can be represented by half equations, for example:

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

and

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$ or

$$4OH^- - 4e- \rightarrow O_2 + 2H_2O$$

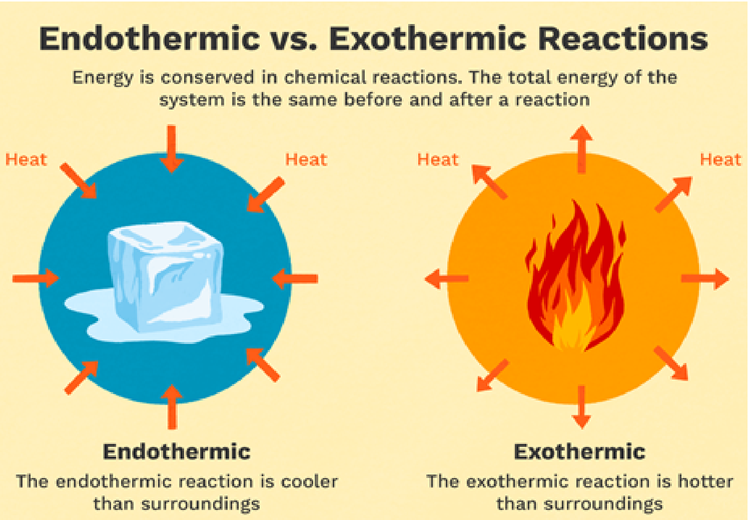
Knowledge Organiser – 5.5 Energy Changes (separate chemistry)

5.5.1 Exothermic and endothermic reactions

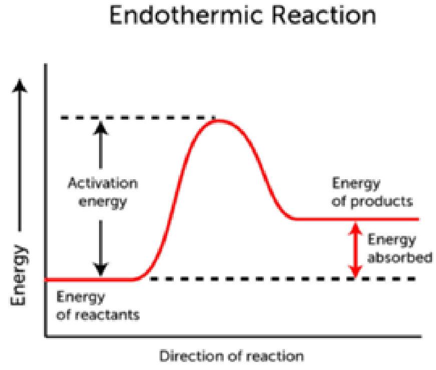
5.5.1.1 Energy transfer during Exothermic and endothermic reactions

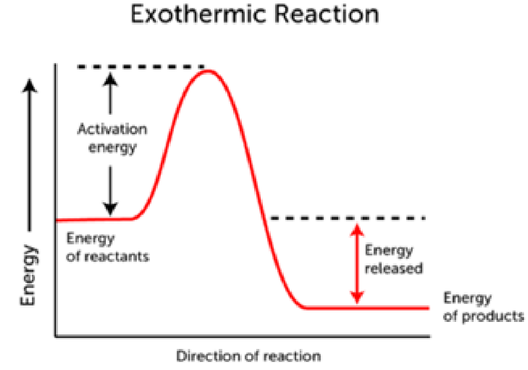
- **Energy is conserved** in chemical reactions.
- The amount of energy in the universe at the end of a chemical reaction is the same as before the reaction takes place.
- If a **reaction transfers energy to the surroundings** the **product** molecules must have **less energy than the reactants, by the amount transferred.**

- **Exothermic reactions give out energy to the surroundings. Exo = exit.**
  - Examples: combustion, neutralisation, hand warmers.
- **Endothermic reactions take in energy from the surroundings. Endo = enter.**
  - Examples: thermal decomposition, reaction of citric acid and sodium hydrogencarbonate and sports injury packs.



5.5.1.2 Reaction Profiles



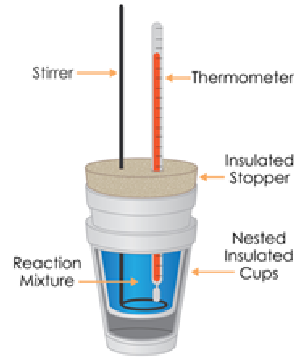


- In an **endothermic** reaction profile the **products finish higher in energy** than the reactants.
- In an **exothermic reaction** profile the **products finish lower in energy** than the reactants.

- **Activation energy** is the **minimum** energy required for a reaction to happen when particles **collide.**

- The **overall energy change** is the difference between the relative energy of the reactants and the products.

RPA Investigate the variables that affect temperature in reacting solutions



The variables you could change are:

- Type of reactant (metal, carbonate, alkali)
- Type of acid used.
- Concentration of acid.
- Size of reactant pieces (if solid).
- Concentration of alkali.

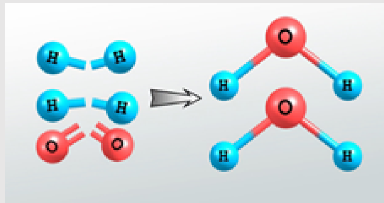
If one of these variables is changing, then all others stay the same.

Knowledge Organiser – 5.5 Energy Changes (separate chemistry)

5.5.1.3 Energy change of reactions (HT only)

During a chemical reaction:

- **Energy must be put in** to **break bonds** in the **reactants.**
- **Energy is given** out when **bonds in the products are formed.**



- If **overall energy change is negative** = **exothermic** reaction.
- If **overall energy change is positive** = **endothermic** reaction.

In **exothermic** reactions, the **energy released from forming new bonds** is **greater** than the energy needed to break existing bonds. In **endothermic** reactions, the **energy needed to break existing bonds** is **greater** than the energy released from forming new bonds.

The **difference between the sum of the energy needed to break bonds** in the reactants and the **sum of the energy released when bonds in the products are formed** is the **overall energy change of the reaction.**

Example:

Bond	Average bond energy (kJ mol⁻¹)
H — H	436
O — H	463
O ═ O	498

Bonds broken:

- 2 x **H-H** = 2 x 436 = 872 kJ/mol
- **O=O** = 498 kJ/mol
- Total = 872 + 498 = 1370 kJ/mol

Bonds formed:

- 4 x **H-O** = 4 x 463 = 1852 kJ/mol
- Total = 1852 kJ/mol

- **Exothermic** energy change is **negative**
- **Endothermic** energy change is **positive.**

Total energy change = reactants - products:

1370 kJ/mol – 1852 kJ/mol = - 482 kJ/mol

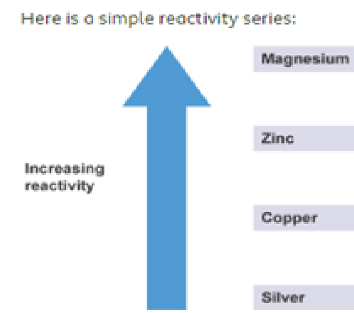
4.5.2.1 Cells and Batteries

Chemical cells use chemical reactions to transfer energy by **electricity.**
The **voltage** of a cell depends upon a number of factors, including what the **electrodes** are made from, and the substance used as the **electrolyte.**

- In non-rechargeable cells, eg alkaline cells, a voltage is produced until one of the reactants is used up. When this happens, we say the battery 'goes flat'.
- In rechargeable cells and batteries, like the one used to power your mobile phone, the chemical reactions can be reversed when an external circuit is supplied.

- A simple cell can be made by connecting two different metals in contact with an electrolyte.
- A number of cells can be connected in series to make a battery, which has a higher voltage than a single cell.
- If we connect different combinations of metals to make a cell, we find that the voltage changes.

What affects the voltage of a cell?



4.5.2.2 Fuel cells

Fuel cells work in a different way than chemical cells. Fuel cells produce a **voltage** continuously, as long as they are supplied with:

✓ a constant supply of a suitable **fuel**
✓ oxygen, eg from the air

The fuel is **oxidised** electrochemically, rather than being burned, so the reaction takes place at a lower temperature than if it was to be burned. Energy is released as electrical energy, not **thermal energy** (heat).

Hydrogen-oxygen fuel cells:
Hydrogen-oxygen fuel cells are an alternative to rechargeable cells and batteries. In a hydrogen-oxygen fuel cell, hydrogen and oxygen are used to produce a voltage. Water is the only product. The overall reaction in a hydrogen-oxygen fuel cell is:

hydrogen + oxygen → water

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

4.5.2.2 Electrode half equations (HT)

At the negative electrode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

At the positive electrode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

When you add these two half equations together, you get the following overall equation:

$$2H_2 + 4OH^- + O_2 + 2H_2O + 4e^- \rightarrow 4H_2O + 4e^- + 4OH^-$$

The hydroxide ions, electrons and two H_2O molecules will now cancel because they are on both sides, leaving the overall equation: $2H_2 + O_2 \rightarrow 2H_2O$

Knowledge Organiser – 5.5 Energy Changes

5.5.1 Exothermic and endothermic reactions

5.5.1.1 Energy transfer during Exothermic and endothermic reactions

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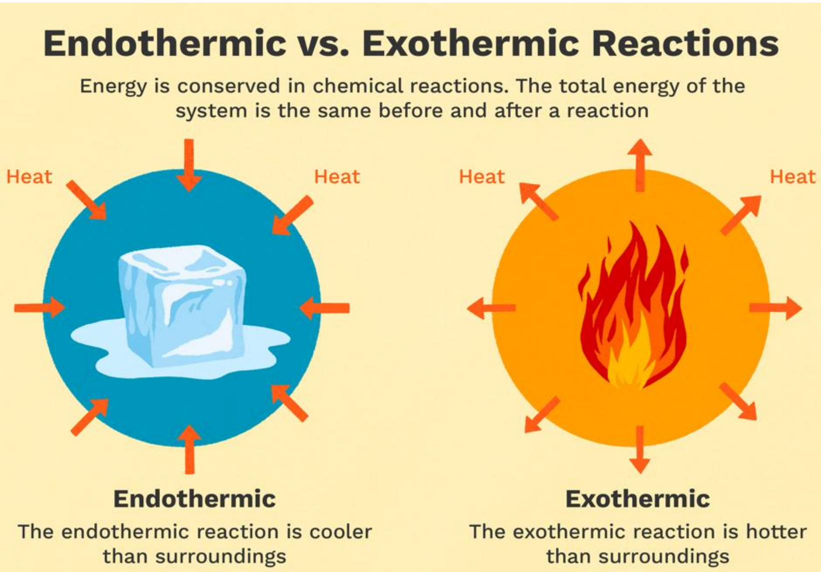
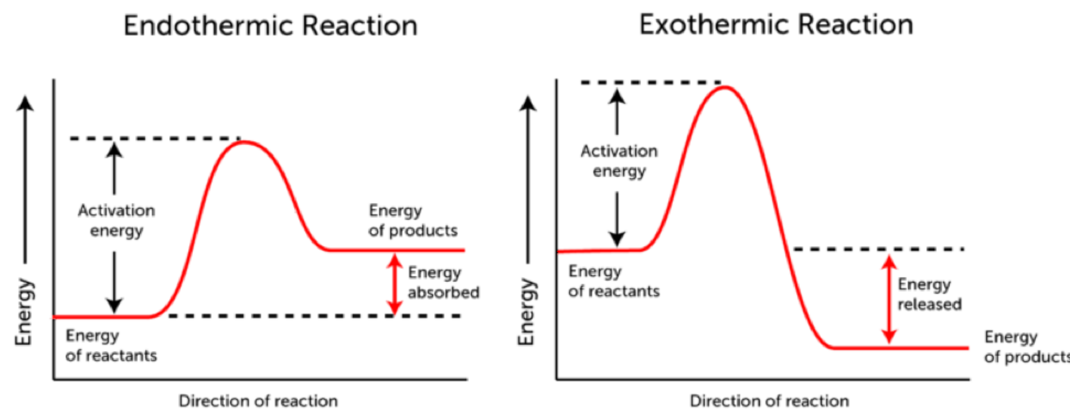


5.5.1.2 Reaction Profiles



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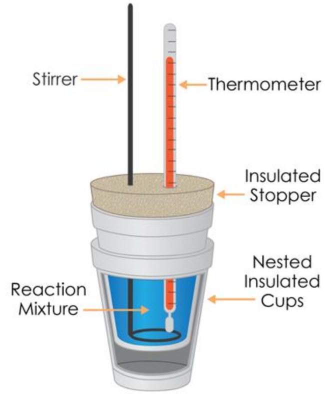

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Knowledge Organiser – 5.5 Energy Changes

5.5.1.3 Energy change of reactions (HT only)

During a chemical reaction:

- **Energy must be put in** to **break bonds** in the **reactants**.
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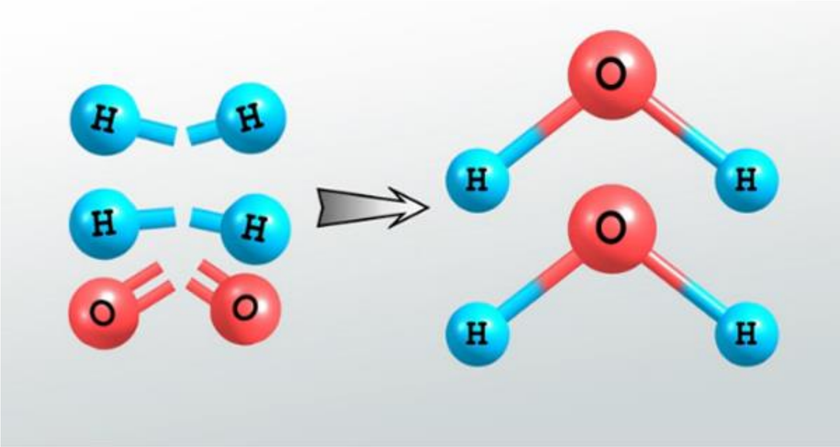

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- Total = 1852 kJ/mol

Total energy change = reactants - products:

1370 kJ/mol – 1852 kJ/mol = - 482 kJ/mol

Knowledge Organiser – 5.6 The Rate & Extent of Chemical Change

5.6.1.1 Calculating rates of reactions

The rate of a reaction is a measure of how quickly a **reactant** is used up, or a **product** is formed.

Rate is calculated by:

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

OR...

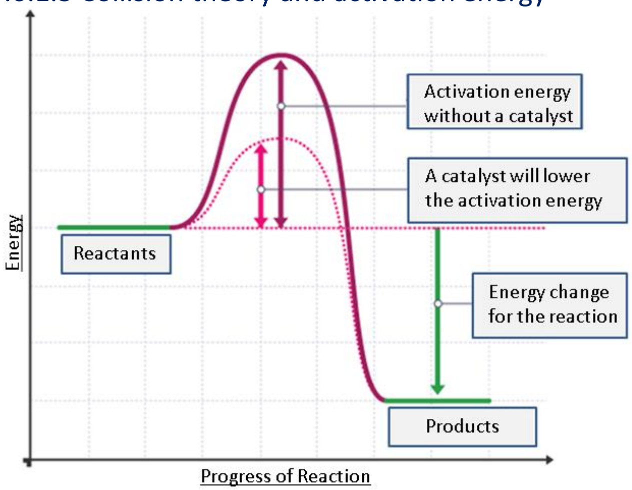
$$\textit{mean rate of reaction} = \frac{\textit{quantity of product formed}}{\textit{time taken}}$$

Eg If 4 moles of a **product** were made during ten seconds, the average rate of reaction would be 4 ÷ 10 = **0.4 mol/s**

Or, if 40g of a **reactant** was used up during 10 seconds, the average rate of reaction would be 40g ÷ 10 = **4 g/s**

Or, if 50 cm^3 of **product** was made during 25 seconds, the average rate of reaction would be 50cm^3 ÷ 25 = **2cm^3/s**

5.6.1.3 Collision theory and activation energy



A catalyst provides an alternative **reaction pathway**

→ that has a lower **activation energy**

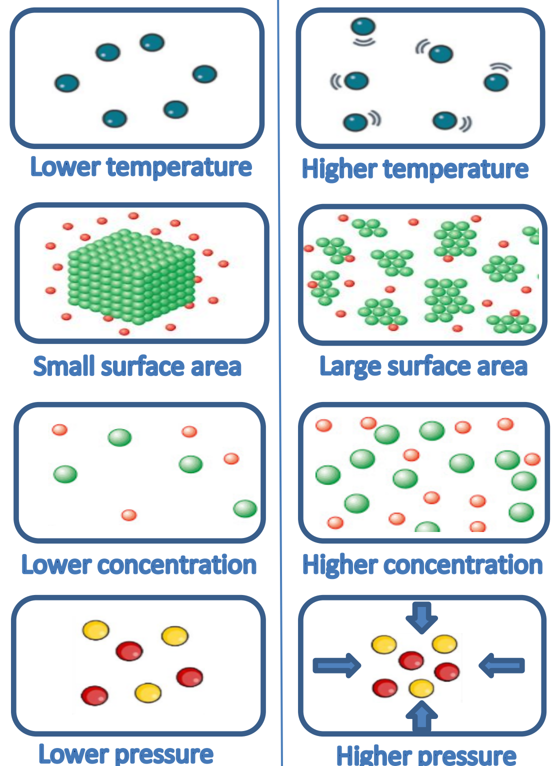
→ does not change the frequency of collisions

→ does increase the frequency of successful collisions

→ more particles have energy greater than the activation energy

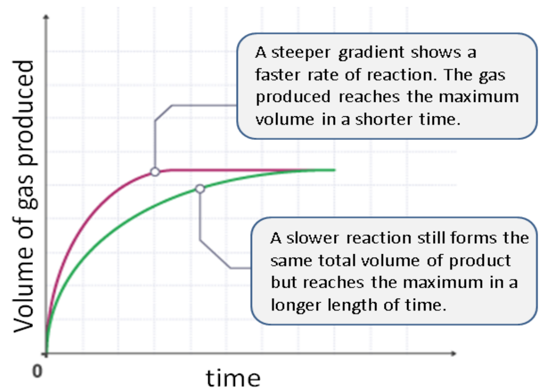
→ therefore more successful collisions

5.6.1.2 Factors affecting rates of reactions



Less chance of reactant particles colliding

More chance of reactant particles colliding



5.6.1.4 Catalysts

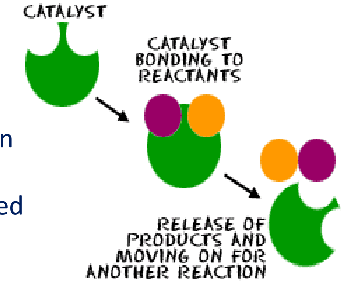
A **catalyst** is a substance that:

- **speeds up** the rate of a reaction
- **does not alter the products** of the reaction
- **not chemically changed or used up** at end of reaction
- Is **only needed in small masses**

❖ Not all reactions have suitable catalysts.

Enzymes act as catalysts in biological systems..

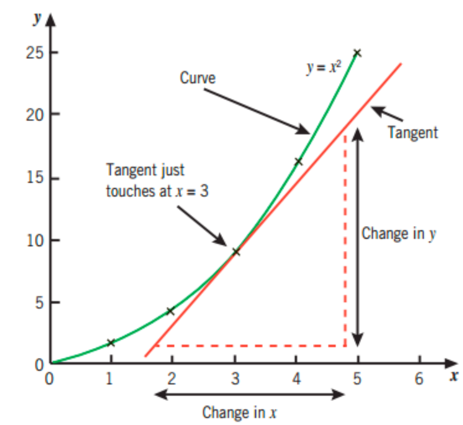
Catalysts are not included in the chemical equation for reaction as they are not used or produced in it.



5.6.1.2 Calculating gradient of a graph

For **straight** line: Draw a **triangle** against the line, divide the vertical measurement by the horizontal to get the rate of reaction.

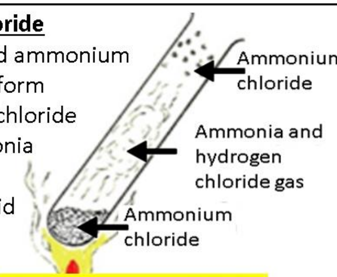
For a **curve**, draw a **tangent** against the line.



5.6.2.1 Reversible Reactions A

Heating ammonium chloride

When heated, white solid ammonium chloride decomposes to form ammonia and hydrogen chloride gas. When cooled, ammonia and hydrogen chloride react to form a white solid of ammonium chloride:

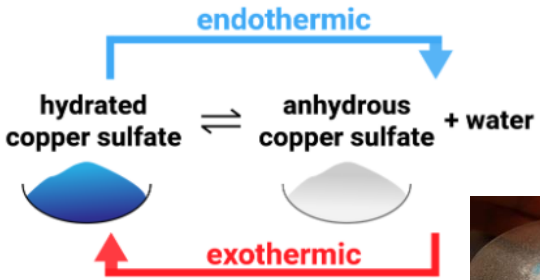


$$NH_4Cl \rightleftharpoons NH_3 + HCl$$

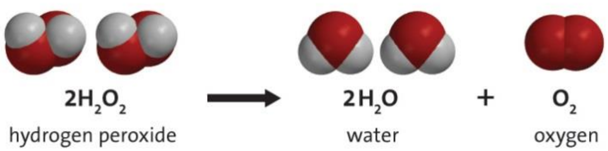
Knowledge Organiser – 5.6 The Rate & Extent of Chemical Change

5.6.2.2 Energy changes and reversible reactions

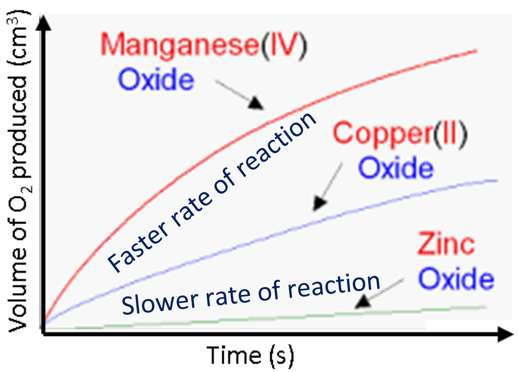
The **reaction** between **anhydrous** copper(II) sulphate and water is used as a test for water.
The white solid turns blue in the presence of water.



Testing different catalysts



Manufacturers can try different catalysts for reactions to find the one that forms the products the quickest, therefore making more profit. The results of a trial of 3 metals salts that act as catalysts for the decomposition of H_2O_2 are below.



5.6.2.3 Equilibrium

If a reversible reaction takes place in a closed system (where no reactants can enter and no products can escape), the forwards and backwards reaction reach a state of **equilibrium**, they occur at exactly the same rate.

Example of a reversible reaction:
Nitrogen gas is reacted with hydrogen gas to make ammonia gas. The forward reaction (making ammonia) is **exothermic**.

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

The backwards reaction (making nitrogen and hydrogen gas is **endothermic**).

The equilibrium position is:

- to the left if the concentrations of N_2 and H_2 are greater than the concentration of NH_3
- to the right if the concentration of NH_3 is greater than the concentrations of N_2 and H_2

5.6.2.4 The effect of changing conditions on equilibrium (HT Only)

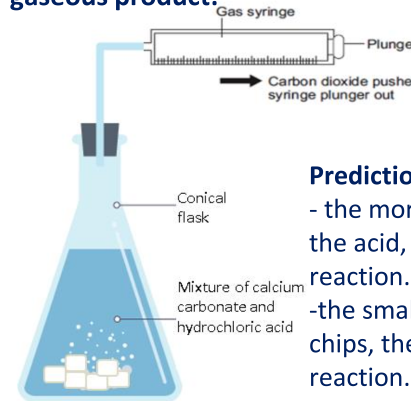
Le Chatelier's principle
The equilibrium position can be changed by changing the reaction conditions through:

- changing the **pressure**
- changing the **concentration**
- changing the **temperature**

The system will respond to counteract the change.

- Industry uses this principle regularly to increase the amount of product they make (for the best profits!).

RPA Finding the rate of reaction by volume of a gaseous product:

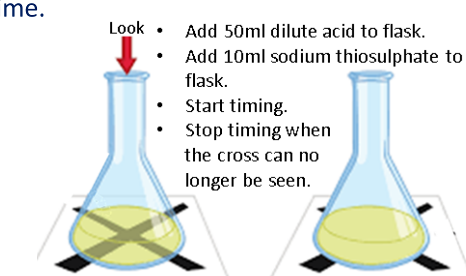


Prediction
- the more concentrated the acid, the faster the reaction.
-the smaller the marble chips, the faster the reaction.

The reaction is complete when no more gas is being produced.

Hazard	Possible harm	Possible precaution
Hydrochloric acid	Causes skin and eye irritation	Wear eye protection
Fizzing in the reaction mixture	Acidic spray or foam may damage skin and eyes	Use a large conical flask so there is plenty of space inside; do not look over the top when adding the calcium carbonate

RPA: Finding the rate of reaction by turbidity.
Sodium thiosulfate solution reacts with dilute hydrochloric acid. The **sulfur** produced forms a cloudy yellow-white **precipitate** during the reaction. The time taken for this to achieve a given cloudiness provides a way to measure the reaction time.



- Add 50ml dilute acid to flask.
- Add 10ml sodium thiosulphate to flask.
- Start timing.
- Stop timing when the cross can no longer be seen.

Prediction – the reaction is faster if the reactants are hotter or more concentrated.

Knowledge Organiser – 5.6 The Rate & Extent of Chemical Change

More details for Higher Tier - If a system is at equilibrium and a change is made to any of the conditions (temperature, concentration, pressure), then that system will respond to counteract the change. This is called **Le Chatelier's Principle.**

Changing the concentration, temperature and pressure of a reaction system can make a big change to where the equilibrium lies, and industry uses this principle regularly to increase the amount of product they make (for the best profits!).

5.6.2.5 Effect of changing concentration on equilibrium (HT Only)

Changing concentration

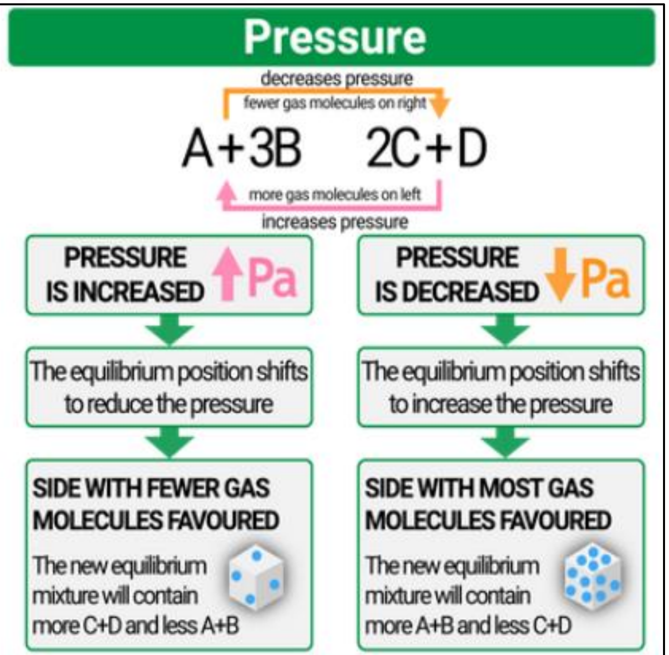
if you add more reactant, the equilibrium will shift to the right to reduce the concentration of reactant (and make more product)

if you remove some of the product, the equilibrium will shift to the right to increase the concentration of the product

5.6.2.5 Effect of changing pressure on equilibrium (HT Only)

Changing pressure

If you increase the pressure then the equilibrium will shift to reduce it (by favouring which ever side of the reaction has the fewest molecules of gas)



Catalysts

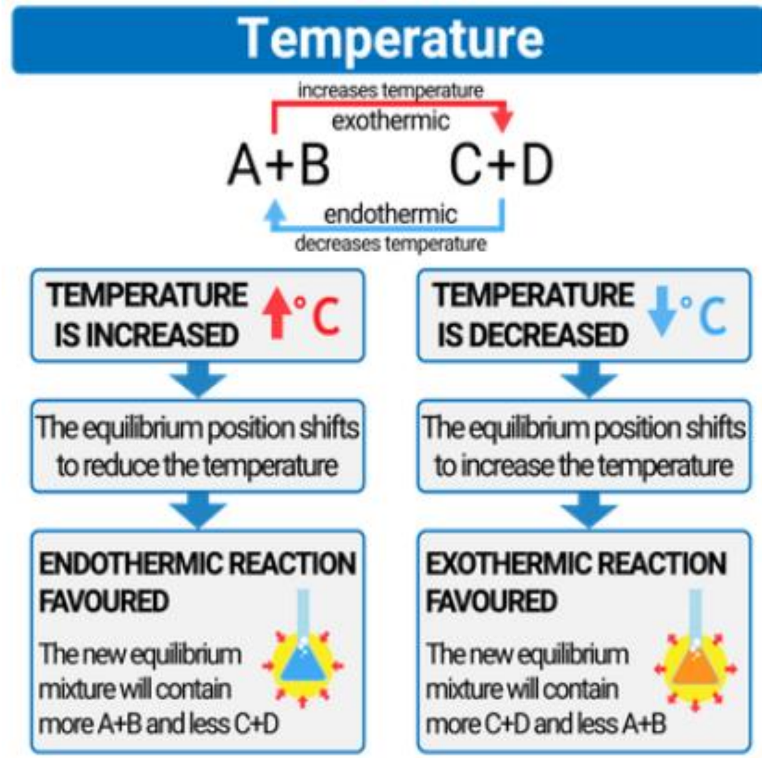
Increase the rate of both the forward and backwards reactions equally (so does not change the equilibrium position)

5.6.2.5 Effect of changing temperature on equilibrium (HT Only)

Changing temperature

If the temperature is increased then the equilibrium position will shift to reduce the temperature (so will favour the endothermic reaction)

If the temperature is decreased then the equilibrium position will shift to increase the temperature (so will favour the exothermic reaction)



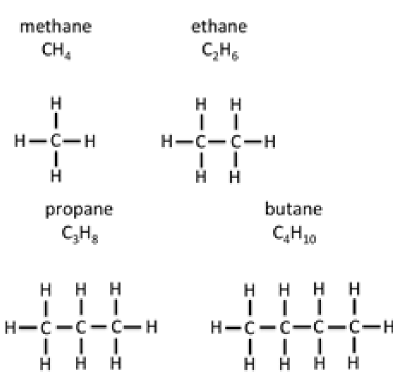
Knowledge Organiser – 5.7 Organic Chemistry

5.7..1.1 Crude oil & hydrocarbons

- **Hydrocarbons** are **compounds** that contain hydrogen and carbon **atoms** only.
- **Crude oil** is a **finite** resource that is found in the Earth's crust. It is the remains of **organisms** that lived and died millions of years ago - mainly plankton which was buried in mud.
- Crude oil is a complex **mixture** of hydrocarbons. The carbon atoms in these molecules are joined together in chains and rings.
- Crude oil is an important source of **fuels** such as petrol, diesel, kerosene, heavy fuel oil and liquefied petroleum gases, & **feedstock** for the petrochemical industry

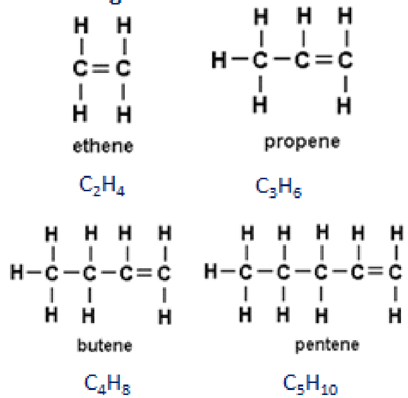
5.7.1.2 Alkanes C_nH_{2n+2}

- Saturated hydrocarbons
- contain H and C atoms
- Single covalent bonds between carbon atoms
- Majority of compounds in crude oil are alkanes



5.7.1.2 Alkenes CnH2n

- Unsaturated hydrocarbons
- Contain H and C
- A double covalent bond between some carbon atoms so they are more reactive.
- Formed from alkanes during **cracking.**



5.7.1.2 Petrochemical industry

Petrochemical: a substance made from crude oil using chemical reactions

Solvent: The liquid in which the solute dissolves to form a solution.

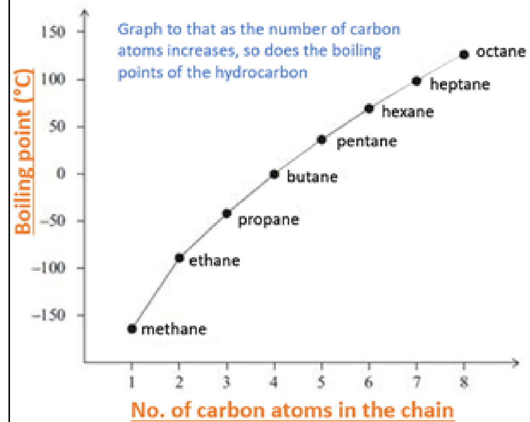
Lubricant: A lubricant is anything which reduces the friction between two surfaces.

Detergent: A mixture of chemicals which have cleaning properties when dissolved in water, and are able to dissolve grease.

Feedstock: A raw material used to provide reactants for industrial reactions.

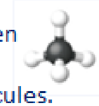
Polymer: A large molecule formed from many identical smaller molecules known as monomers.

5.7.1.3 Properties of Hydrocarbons: Alkenes

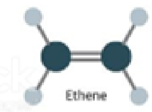


5.7.1.4 Cracking & Alkenes

Hydrocarbons can be broken down (cracked) to produce smaller, more useful molecules.



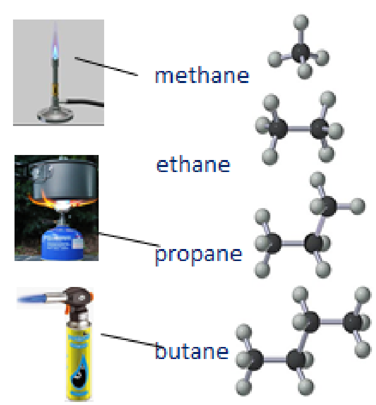
Cracking products include alkanes and an alkene.



There is a high demand for small molecules and so some of the products of cracking are useful as fuels.

5.7.1.3 Properties of Hydrocarbons: Alkanes

- If a substance is **more** viscous it is thick and sticky. Solids are most viscous.
- If a substance is **less** viscous, it is easier to pour. Gases are least viscous.



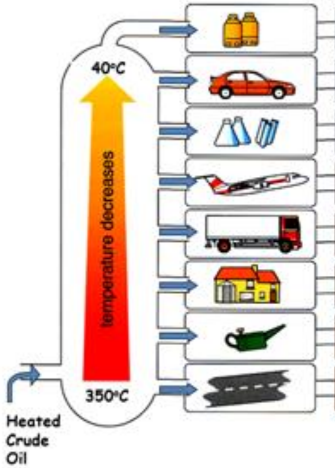
The more carbons there are in the chain, the more **viscous** the hydrocarbon. It becomes thicker. Its **viscosity** increases.

The fewer carbons there are in the chain, the more **volatile** the hydrocarbon is, therefore the easier it is to **ignite**. Its **flammability** increases.

- **Flammability** is measure of how **easy the hydrocarbon is to ignite and burn.**
- Small hydrocarbons such as methane are more **volatile.** This means they **evaporate easily.**
- This means that they **turn into a gas at a lower temperature** and can **ignite easily**

Knowledge Organiser – 5.7 Organic Chemistry

5.7.1.2 Fractional Distillation



Fraction	Boiling pt.	C atoms in chain
Liquid petroleum gas	< 25°C	3
Petrol (gasoline)	25 - 60°C	8
Naphtha	60 - 180°C	10
Paraffin	180 - 220°C	12
Diesel	220 - 250°C	20
Fuel oil	250 - 300°C	40
Lubricating oil	300 - 350°C	80
Bitumen	> 350°C	120

Fractional distillation works because the **different liquids have different boiling points.** When the mixture is heated:

- **vapours rise** through a column which is hot at the bottom, and cooler at the top
- vapours **condense** when they reach a part of the column that is below the temperature of their boiling point
- each liquid is led away from the column to be collected

5.7.1.4 Steam Cracking

C_2H_6 → C_2H_4 + H_2

Uses a temperature of approx. 550°C and no catalyst

5.7.1.4 Thermal Decomposition

Example

$C_{10}H_{22} \xrightarrow[\text{catalyst}]{800^\circ C} C_2H_4 + C_3H_6 + C_5H_{12}$

Decane → Ethene + Propene + Pentane

Used to make plastics (polymers) Polyethene / polypropene

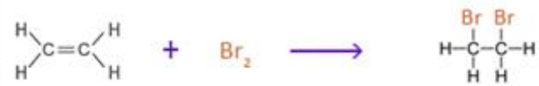
Used in petrol

As a large molecule is broken into smaller ones using heat, this is another example of a thermal decomposition reaction.

5.7.1.4 Testing for double bonds

This is the way to test for a double C=C bond in a molecule.

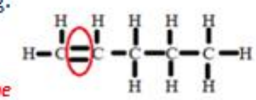
$C_2H_4(g)$ $Br_2(aq)$ $C_2H_4Br_2(aq)$



Ethene gas is added to **bromine water**.
The bromine water is **yellow/ orange** at the start.
When it reacts with ethene is turns **colourless**.

The double **covalent** bond in the ethene opens up to join with the bromine atoms from the bromine water.

This test works with **any** hydrocarbon that has a **double bond** e.g.



Pentene

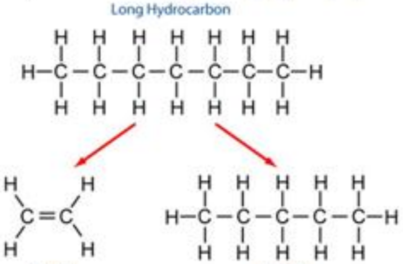
5.7.1.4 Equations for Cracking

Hexane can be cracked to form butane & ethene:

Word Equation: hexane → butane + ethane

Symbol Equation: $C_6H_{14} \rightarrow C_4H_{10} + C_2H_4$

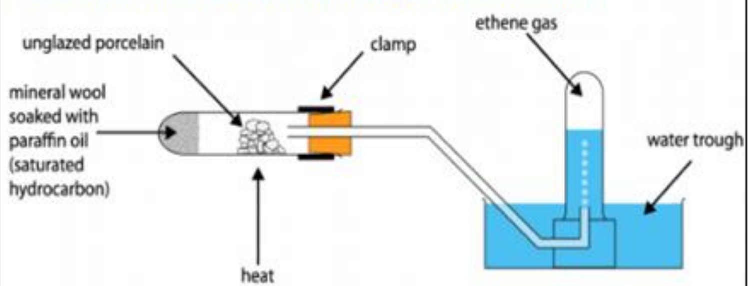
Long Hydrocarbon



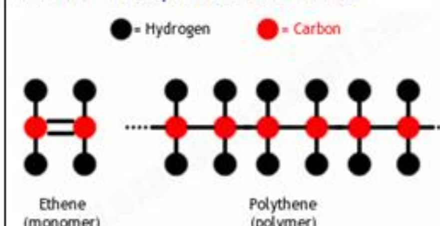
Ethene Alkane

5.7.1.4 Catalytic Cracking

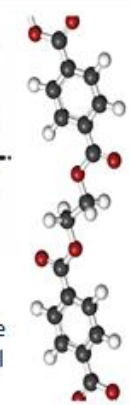
Uses a temperature of approx. 550 °C and a **catalyst** known as a zeolite which contains aluminium oxide and silicon oxide.



5.7.1.4 How plastics are made



Polymers, whether artificial (such as the plastic shown) or natural, are made of repeating chains of smaller chemical units (**monomers**).

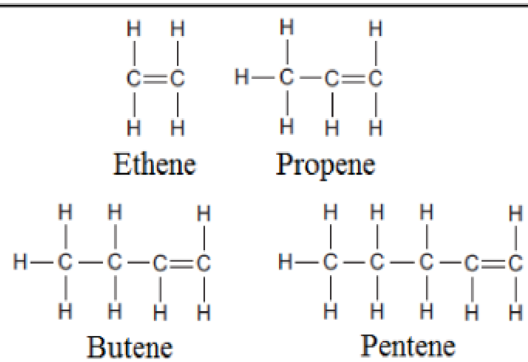


Knowledge Organiser – 5.7 Organic Chemistry (Separate Chemistry)

Key terms	
Functional group	An atom or group of atoms that give organic compounds their characteristic reactions.
Homologous series	Family of organic compounds with the same functional group.
Double bond	A covalent bond made by the sharing of two pairs of electrons.
Unsaturated hydrocarbon	A hydrocarbon whose molecule contains at least one carbon-carbon double bond.
Alkene	A **hydrocarbon** containing at least one **double bond**. They follow the formula **C_nH_{2n}**. Used to make **polymers**.
Bromine water	A chemical that is **brown/orange** in colour. If added to an **alkene** it reacts and changes to **colourless**. Alkanes do not produce a change in colour.
Addition	two molecules add together to form a single product with 100% atom economy.
Oxidising agent	A substance that has the ability to oxidise other substances. Its symbol is [O]

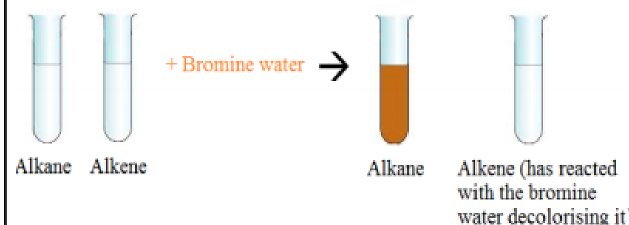
Structure of Alkenes

Alkenes are unsaturated hydrocarbons. The general formula of the alkenes containing one double bond is **C_nH_{2n}**



Reactions of the alkenes

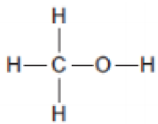
It is the **C=C double bond** that makes the **alkenes far more reactive than the alkanes**. Alkenes will react with hydrogen, water (steam) and the halogens, by addition of atoms across the C=C double bond so that the double bond becomes a single carbon-carbon bond.

Combustion	Alkenes will burn in oxygen to produce carbon dioxide and water. $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ **Alkenes release less energy** per mole in combustion **than alkanes** hence the **alkanes tend to be used as fuels**, whereas the alkenes are not.
Reaction with halogens	Ethene reacts with bromine to form dibromoethane in an **addition** reaction. $CH_2{=}CH_2 + Br_2 \rightarrow CH_2BrCH_2Br$ When you test ethene with **orange bromine water** it turns the bromine water from orange to colourless.  The alkenes also react in a similar way with the other halogens, chlorine and iodine.
Reaction with hydrogen	Alkenes **reacts with hydrogen** in the presence of a **nickel catalyst** at a temperature of about 150°C to **produce an alkane**. $C_2H_4 + H_2 \rightarrow C_2H_6$ This reaction is used to add hydrogen across double bonds in unsaturated oils making margarine.
Reaction with water (steam)	Ethene **reacts with steam** in the presence of a **catalyst** to make ethanol. $C_2H_4 + H_2O \rightleftharpoons C_2H_5OH$ The reaction also requires heat and high pressure. The reaction is **reversible** so unreacted steam and ethane are recycled over the catalyst.

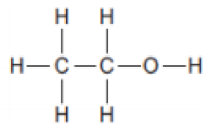
Knowledge Organiser – 5.7 Organic Chemistry (Separate Chemistry)

Structure of Alcohols

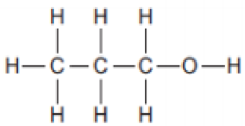
Alcohols contain the -OH functional group.



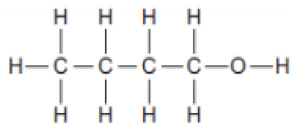
Methanol



Ethanol



Propanol



Butanol

Reactions of the alcohols

Combustion	Alcohols are **flammable** and will burn in oxygen with a **clean blue flame** to produce **carbon dioxide** and **water**. $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
With sodium	React with sodium metal to produce a solution **of sodium alkoxide and hydrogen gas.** $2C_2H_5OH + 2Na \rightarrow 2C_2H_5ONa + H_2$ If sodium ethoxide, or any other sodium alkoxide is dissolved in water, **effervescence (bubbles)** are observed and you get a **strongly alkaline solution**.
Oxidation	Combustion is one way to oxidise an alcohol, however you can also oxide an alcohol using an **oxidizing agent** such as **potassium dichromate**. An alcohol is oxidized to a **carboxylic acid** when boiled with **acidified** potassium dichromate. $C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$
With water	Alcohols dissolve many of the same substances as water. They also dissolves some organic compounds that water cannot, making them **excellent solvents**. The first four alcohols dissolve well with water making a neutral solution.

Uses of alcohols

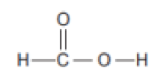
Alcohols are used as solvents in products such as perfumes, aftershaves and mouthwashes. Ethanol is the main alcohol in alcoholic drinks. Ethanol is also used in spirit burners and as a fuel, for e.g. as a biofuel in cars.

Manufacture of ethanol

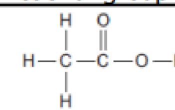
Fermentation	Ethanol is made by **fermenting sugars** from plant material with **yeast**. The reaction is typically carried out between **20-30°C**. $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ All equipment must be **sterile** at the start. It also has to be carried out under **anaerobic (without air)** conditions, otherwise the ethanol would react with oxygen and turn into vinegar. Ethanol made by fermentation is termed a **biofuel.**
From ethene	Ethanol can also be made from reacting ethene (obtained from cracking of crude oil) and steam in the presence of a catalyst. This method uses up crude oil, a non renewable resource.

Structure of Carboxylic acids

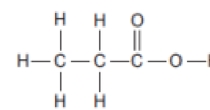
Carboxylic acids contain the -COOH functional group.



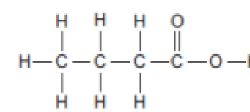
Methanoic acid



Ethanoic acid



Propanoic acid



Butanoic acid

Reactions of Carboxylic acids

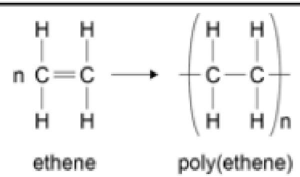
With metal carbonates	Forms a salt, water and carbon dioxide $2CH_3COOH + Na_2CO_3 \rightleftharpoons 2CH_3COONa + H_2O + CO_2$ Effervescence (bubbles) observed as $CO_{2(g)}$ forms
In water (HT)	Aqueous solutions of carboxylic acids are **weak acids** & only **partially ionise** (have higher pH than strong acids of same concentration). $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
With alcohols	**Esters** are formed. **A sulfuric acid catalyst** is required. $CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$ In this reaction, the ester **ethyl ethanoate** forms. Esters are **sweet/fruity smelling** & used in perfumes & food flavourings.

Knowledge Organiser – 5.7 Organic Chemistry (Separate Chemistry)

Section 1: Key terms	
Polymer	**Very large covalently bonded** molecules with **many repeating units** (poly means many).
Monomer	**Small reactive molecules** which join together to make a polymer (mono means one).
Plastics	Made of very large covalently bonded molecules called polymers
Addition polymerisation	The reaction between **alkene monomers** to form a polymer
Condensation polymerisation	Usually involves a **small molecule released** in the reaction (like **water or HCl**), as the polymer forms.
Monosaccharide	Simple **carbohydrates** made from **one sugar unit** e.g. glucose.
Polysaccharide	A polymer made from monosaccharide monomers e.g. starch or cellulose).
Protein	Polymers of **amino acids**
DNA	**D**eoxyribo**n**ucleic **a**cid is made up from monomers called **nucleotides**
Nucleotides	**Monomers** used to make DNA. There are four different types that can react to form DNA polymers.

Section 2: Addition poly[redacted]

One of the most important ways that chemicals from crude oil are used is to make polymers. Alkenes can be used to make polymers such as poly(ethene) and poly(propene) by addition polymerisation.



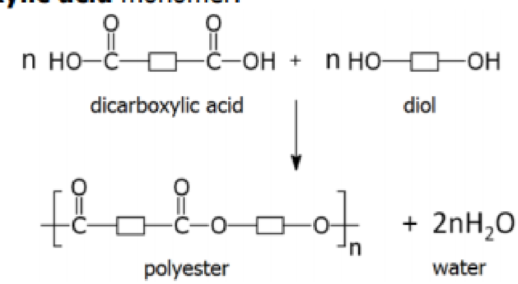
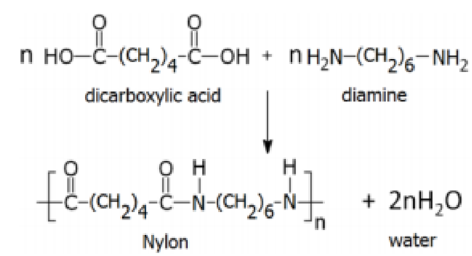
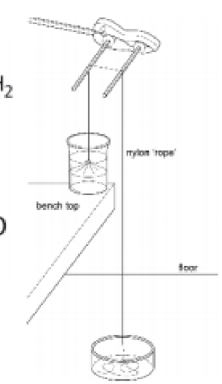
In addition polymers the repeating unit has the **same atoms as the monomer** because when the C=C bond **"opens up"** in polymerisation, **no other molecule** is formed in the reaction.

Uses	**Polyethene** is very useful as it is strong, transparent and easily shaped. Used to make drinks bottles, washing up bowls, dustbins and cling film. **Polypropene** forms a very strong tough plastic. Used to make carpets, milk crates and ropes.

Section 3: Condensation [redacted]ation (HT)

As well as addition polymerisation (which requires monomers with a C=C), chemists can also make polymers from another type of reaction called **condensation polymerisation.**

Condensation polymerisation involves monomers with **two functional groups**. When these types of monomers join together, they usually lose small molecules such as water or HCl, and so the reactions are called condensation reactions. Two products are usually formed.

Examples	Polyester (used to make clothing) and nylon (used to make rope and stockings).
Forming a polyester	Requires an **diol (dialcohol)** monomer and a **dicarboxylic acid** monomer. 
Forming nylon	Requires a **diamine** monomer and a **dicarboxylic** acid monomer.   Nylon thread can be made using the apparatus shown in the diagram

Knowledge Organiser – 5.7 Organic Chemistry (Separate Chemistry)

Natural polymers

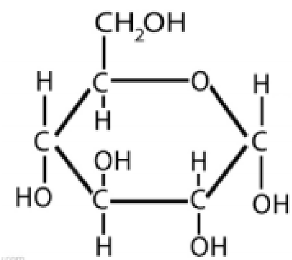
Naturally occurring polymers are found in all living things (e.g. polymers that make up starch, cellulose, proteins and DNA. They are formed during **condensation polymerisation** reactions.

Making polysaccharides from sugars

Simple carbohydrates (monosaccharides) are compounds containing carbon, hydrogen and oxygen e.g. glucose $C_6H_{12}O_6$

Monosaccharides can bond together to make polymers (polysaccharides). **Starch and cellulose are polysaccharides** made from **glucose** monomers. Plants use the starch they make from glucose as energy stores.

Glucose monomers → Starch polymer + water

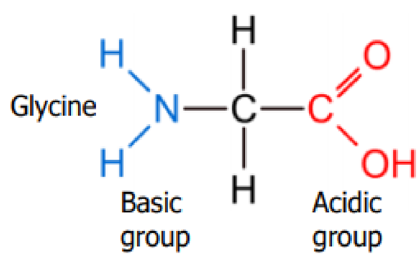


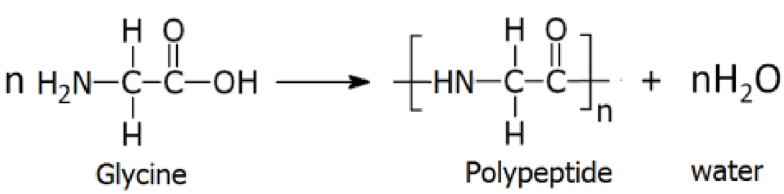
Glucose

Making polypeptides and proteins from amino acids. (HT)

The monomers of proteins are called **amino acids**. Amino acids have **two functional groups**, one basic (the amine group – NH_2) and one acidic (carboxylic acid group -COOH). The simplest amino acid is glycine.

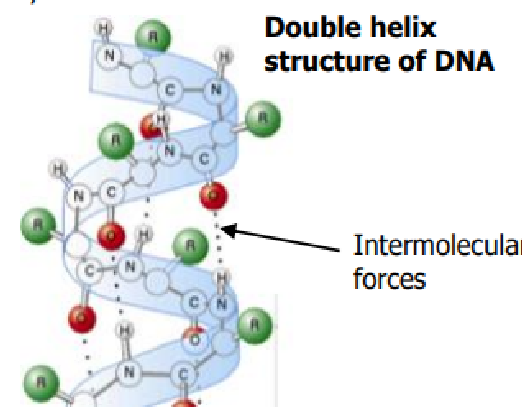
Many more glycine monomers can link together form a polypeptide molecule. There are about 20 amino acids that join together in a variety of sequences that make up more than 1000 proteins in your body.





DNA

DNA (**d**eoxyribo**n**ucleic **a**cid) is a natural polymer **essential for life** because it enables living things to develop and function. It is made up from monomers called **nucleotides**. DNA's structure contains a **genetic code** that determines the different **amino acid sequences** of every protein in living organisms and viruses.

Nucleotide	Based on the sugar **deoxyribose**, bonded to a phosphate group and a base.
How is DNA made	By the condensation polymerisation of repeating units of nucleotide monomers. DNA is a **polynucleotide**.
Structure of DNA	Most DNA molecules are **two polymer chains**, made from **four different nucleotide** monomers, in the form of a **double helix**. The two polymer strands run in opposite directions to each other and are held in place by the **intermolecular forces** down the length of each polymer strand.  There are **four different nucleotide** monomers that can react to form DNA polymers.

Knowledge Organiser – 5.7 Organic Chemistry

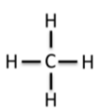
4.3.1.1 Crude oil & hydrocarbons

- **Hydrocarbons** are **compounds** that contain hydrogen and carbon **atoms** only.
- **Crude oil** is a **finite** resource that is found in the Earth's crust. It is the remains of **organisms** that lived and died millions of years ago - mainly plankton which was buried in mud.
- Crude oil is a complex **mixture** of hydrocarbons. The carbon atoms in these molecules are joined together in chains and rings.
- Crude oil is an important source of **fuels** such as petrol, diesel, kerosene, heavy fuel oil and liquefied petroleum gases, & **feedstock** for the petrochemical industry

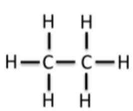
4.3.1.1 Alkanes C_nH_{2n+2}

- Saturated hydrocarbons
- contain H and C atoms
- Single covalent bonds between carbon atoms
- Majority of compounds in crude oil are alkanes

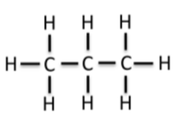
methane
CH_4



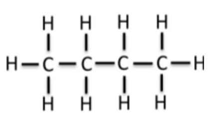
ethane
C_2H_6



propane
C_3H_8

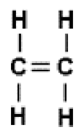


butane
C_4H_{10}

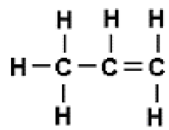


4.3.1.1 Alkenes C_nH_{2n}

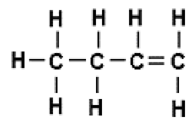
- Unsaturated hydrocarbons
- Contain H and C
- A double covalent bond between some carbon atoms so they are more reactive.
- Formed from alkanes during **cracking.**



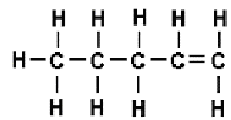
ethene
C_2H_4



propene
C_3H_6



butene
C_4H_8



pentene
C_5H_{10}

5.7.1.2 Petrochemical industry

Petrochemical: a substance made from crude oil using chemical reactions

Solvent: The liquid in which the solute dissolves to form a solution.

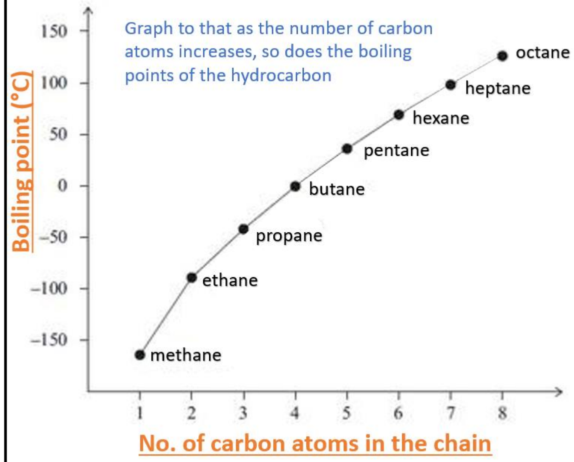
Lubricant: A lubricant is anything which reduces the friction between two surfaces.

Detergent: A mixture of chemicals which have cleaning properties when dissolved in water, and are able to dissolve grease.

Feedstock: A raw material used to provide reactants for industrial reactions.

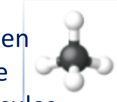
Polymer: A large molecule formed from many identical smaller molecules known as monomers.

5.7.1.3 Properties of Hydrocarbons: Alkenes

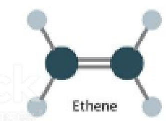


5.7.1.4 Cracking & Alkenes

Hydrocarbons can be broken down (cracked) to produce smaller, more useful molecules.



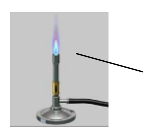
Cracking products include alkanes and an alkene.



There is a high demand for small molecules and so some of the products of cracking are useful as fuels.

5.7.1.3 Properties of Hydrocarbons: Alkanes

- If a substance is **more** viscous it is thick and sticky. Solids are most viscous.
- If a substance is **less** viscous, it is easier to pour. Gases are least viscous.

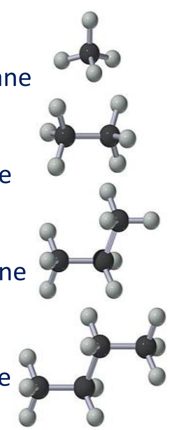


methane

ethane

propane

butane



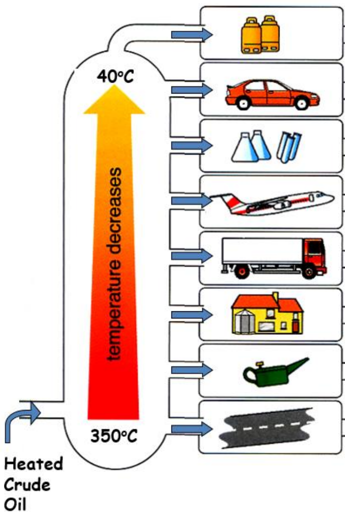
The more carbons there are in the chain, the more **viscous** the hydrocarbon. It becomes thicker. Its **viscosity** increases.

The fewer carbons there are in the chain, the more **volatile** the hydrocarbon is, therefore the easier it is to **ignite**. Its **flammability** increases.

- **Flammability** is measure of how **easy the hydrocarbon is to ignite and burn**.
- Small hydrocarbons such as methane are more **volatile**. This means they **evaporate easily**.
- This means that they **turn into a gas at a lower temperature** and can **ignite easily**

Knowledge Organiser – 5.7 Organic Chemistry

5.7.1.2 Fractional Distillation

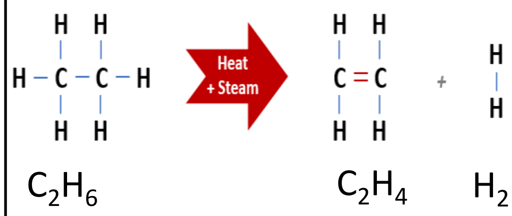


Fraction	Boiling pt.	C atoms in chain
Liquid petroleum gas	< 25°C	3
Petrol (gasoline)	25 - 60°C	8
Naphtha	60 - 180°C	10
Paraffin	180 - 220°C	12
Diesel	220 - 250°C	20
Fuel oil	250 - 300°C	40
Lubricating oil	300 - 350°C	80
Bitumen	> 350°C	120

Fractional distillation works because the **different liquids have different boiling points.** When the mixture is heated:

- **vapours rise** through a column which is hot at the bottom, and cooler at the top
- vapours **condense** when they reach a part of the column that is below the temperature of their boiling point
- each liquid is led away from the column to be collected

5.7.1.4 Steam Cracking



Uses a temperature of approx. 550°C and no catalyst

5.7.1.4 Thermal Decomposition

Example

$C_{10}H_{22} \xrightarrow[\text{catalyst}]{800^\circ C} C_2H_4 + C_3H_6 + C_5H_{12}$

Decane → Ethene + Propene + Pentane

Used to make plastics (polymers) Polyethene / polypropene

Used in petrol

As a large molecule is broken into smaller ones using heat, this is another example of a thermal decomposition reaction.

5.7.1.4 Testing for double bonds

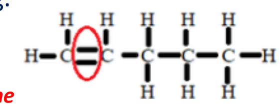
This is the way to test for a double C=C bond in a molecule.

$C_2H_4(g)$ $Br_2(aq)$ $C_2H_4Br_2(aq)$

Ethene gas is added to **bromine water**.
The bromine water is **yellow/ orange** at the start.
When it reacts with ethene is turns **colourless**.

The double **covalent** bond in the ethene opens up to join with the bromine atoms from the bromine water.

This test works with **any** hydrocarbon that has a **double bond** e.g.



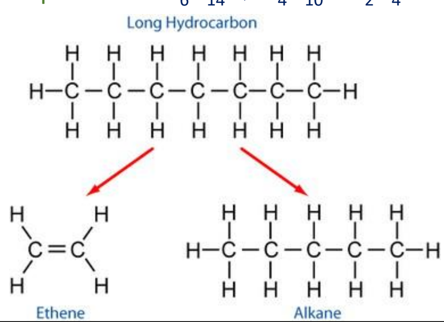
Pentene

5.7.1.4 Equations for Cracking

Hexane can be cracked to form butane & ethene:

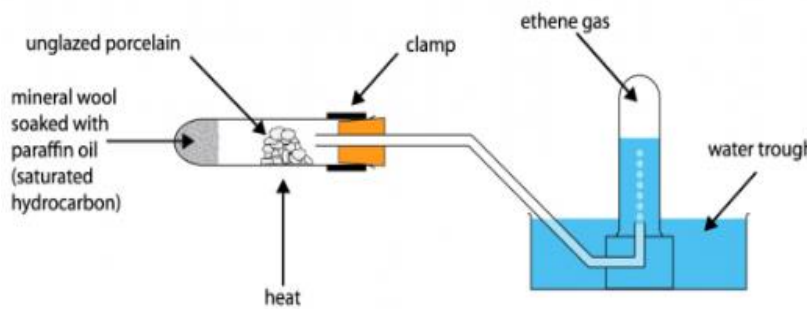
Word Equation: hexane → butane + ethane

Symbol Equation: $C_6H_{14} \rightarrow C_4H_{10} + C_2H_4$

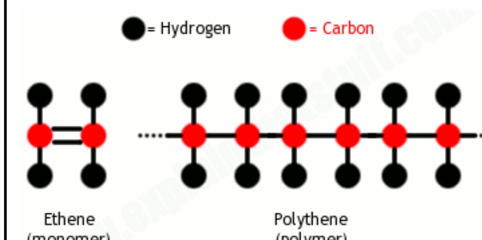


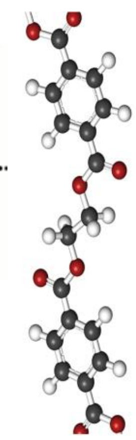
5.7.1.4 Catalytic Cracking

Uses a temperature of approx. 550 °C and a **catalyst** known as a zeolite which contains aluminium oxide and silicon oxide.



5.7.1.4 How plastics are made





Polymers, whether artificial (such as the plastic shown) or natural, are made of repeating chains of smaller chemical units (**monomers**).

Knowledge Organiser – 5.8 Chemical analysis

5.8.1.1 Pure substances

- In chemistry a **pure** substance contains only one type of element or one type of compound.
  - Example: pure water only contains H_2O.
- **Pure** substances **melt and boil** at specific **temperatures** which can be used to **identify** a substance and test if it is pure.
- In everyday life, a **pure** substance is something that has had nothing else added to it and is in its **natural** state.
  - Example: 'pure' orange juice is not chemically pure but doesn't have any chemicals added that aren't from oranges..

5.8.1.2 Formulations

A **formulation** is a mixture that has been **designed** as useful **product.**

Many products are **complex** mixtures in which each chemical has a particular purpose.

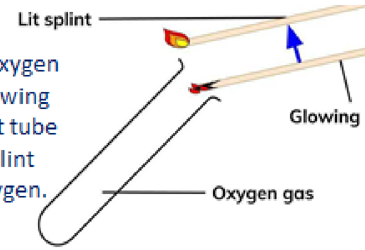
All the ingredients must be mixed in the right quantities so the product has the correct properties.

For example: Fuels, medicine, paint

5.8.2 Identification of common gases

Oxygen

The test for oxygen is to put a glowing splint in a test tube of gas. The splint relights in oxygen.

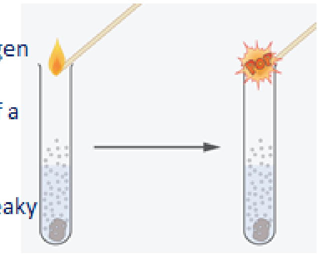


Chlorine

The test for chlorine is to put damp litmus paper in the gas. The chlorine will bleach the paper white.

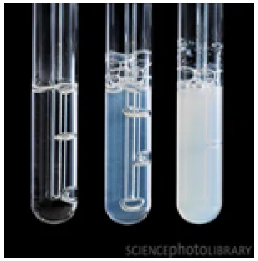
Hydrogen

The test for hydrogen is to put a burning splint at the end of a test tube of gas. Hydrogen burns rapidly with a squeaky pop.



Carbon dioxide

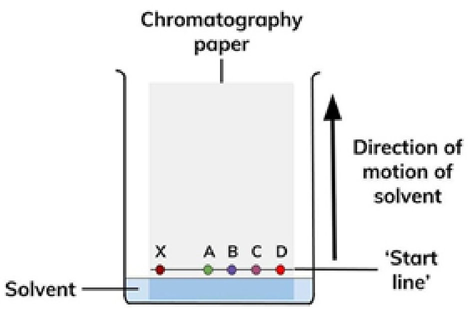
The test for carbon dioxide is to bubble the gas through limewater (calcium hydroxide). The limewater will turn milky (cloudy).

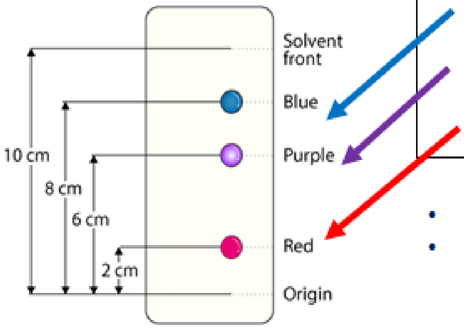


5.8.1.3 Chromatography

- Chromatography **separates** mixtures and can **identify** substances.
- Two phases are used:
  1. The **mobile** phase (moves) is the **solvent.**
  2. The **stationary** phase (doesn't move) is the paper.
- Separation depends on the distribution between phases. **The more soluble a substance is, the more time it spends in the mobile phase.**

RPA : Chromatography





R_f of blue spot $= \frac{8}{10} = 0.8$

R_f of purple spot $= \frac{6}{10} = 0.6$

R_f of red spot $= \frac{2}{10} = 0.2$

- R_f must between 0 and 1.
- In an exam you may be asked to measure these so have a ruler and measure to the nearest mm!

- The **R_f value** is a ratio of how far the substance has travelled with regard to the solvent.

$$Rf = \frac{\text{distance travelled by the component}}{\text{distance travelled by the solvent}}$$

- Each **pure substance has a unique R_f value in each solvent which can identify it.**

- The start line must be drawn in pencil because ink will run.
- The solvent must start below the line otherwise your substances will leech into the solvent.

Chromatography can identify what substances are in a mixture.

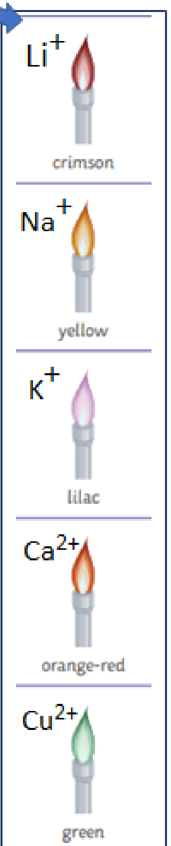
Knowledge Organiser – 5.8 Chemical analysis (Separate Chemistry)

4.8.3.1 Flame Tests

Metal ions when heated produce a **variety** of flame colours. These colours are used to identify the **metal ions** present in a sample.

1. Dip **wire loop** into sample.
2. Place loop into **roaring** Bunsen flame
3. Observe **colour** of flame and cross reference

Mixtures sometimes make it difficult to isolate just **one** colour

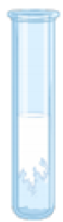
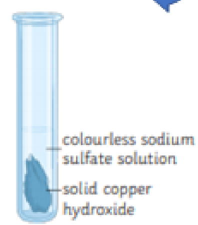
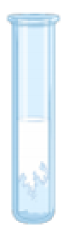
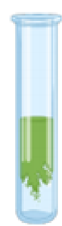
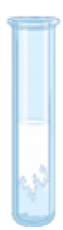


4.8.3.2 Metal Hydroxides

To identify **metal ions, sodium hydroxide** solution should be added to the sample. This causes a **precipitate** to form:

- Calcium or magnesium → **white** (doesn't dissolve)
- Aluminium → **white** (**dissolves** in excess $NaOH_{(aq)}$)
- Iron (II)→ **green**
- Iron (III) → **brown**
- Copper (II) → **blue**

Magnesium sulfate + sodium hydroxide → magnesium hydroxide + sodium sulfate

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

4.8.3.3 Metal Carbonates

- To identify carbonate ions, place a small volume of **limewater** into **test tube.**
- Add the sample and a few drops of **hydrochloric acid** (source of H+ ions) using a **pipette.**
- Seal the test tube with a **bung and delivery tube.**
- Place end of the delivery tube **into** the limewater.
- If the limewater turns **milky-white**, the gas produced is **CO_2** and the sample is a carbonate.

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

4.8.3.4 Halides

- Using a **pipette**, add a few drops of **nitric acid** solution to the sample, then a few drops of **silver nitrate solution.**
- Observe the colour of the precipitate.

Cl- → **Chloride** → **white** precipitate
Br- → **Bromide** → **cream** precipitate
I- → **Iodide** → **yellow** precipitate

4.8.3.5 Sulfates

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

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

4.8.3.2 Half Equations

These can show what happens in **precipitation** reactions but do **not** show spectator ions (those not involved)

<u>Copper (II)</u>

$Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$ Cu^{2+} means atom **lost 2 electrons,** therefore *must* **gain 2 electrons**; 1 from **each** OH^- ion.

<u>Iron (III)</u>

$Fe^{3+} + 3OH^- \rightarrow Fe(OH)_3$ Fe^{3+} means Fe atom **lost 3 electrons,** therefore *must* **gain 3 electrons**; 1 from **each** OH^- ion.

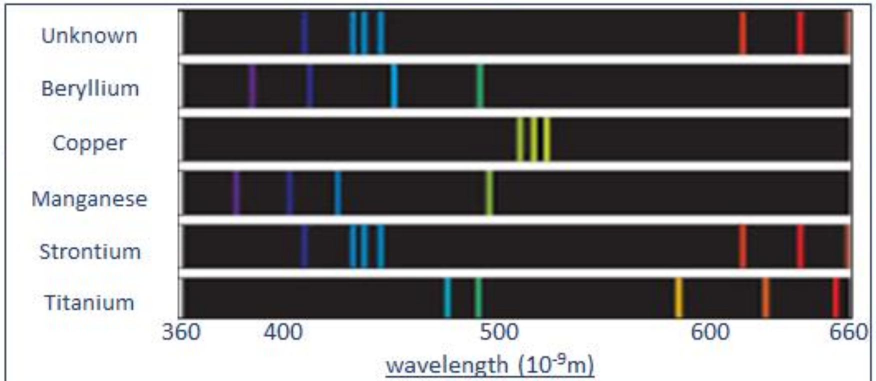
Knowledge Organiser – 5.8 Chemical analysis (Separate Chemistry)

4.8.3.6 Instrumental Methods

The **benefits** are:
- Rapid results
- Accurate
- Sensitive

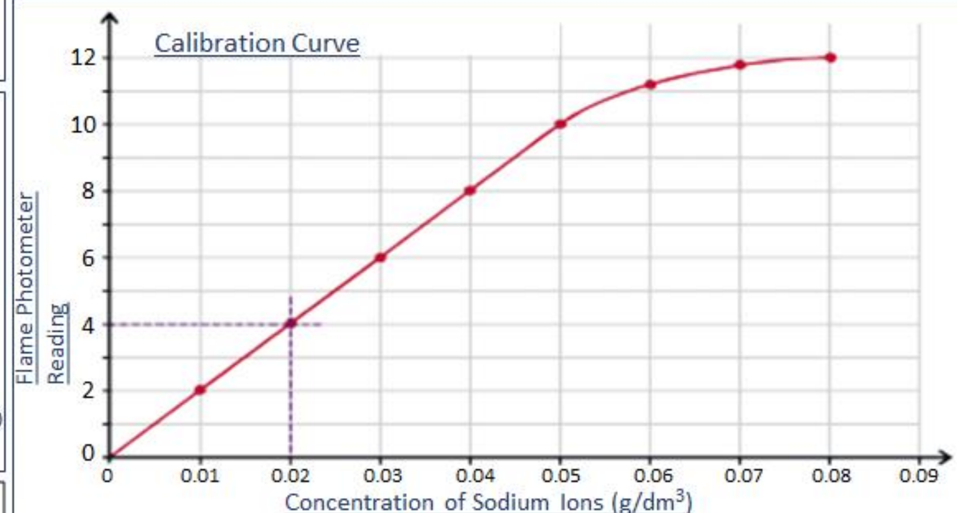
The **drawbacks** are:
- Expensive equipment
- Special training needed

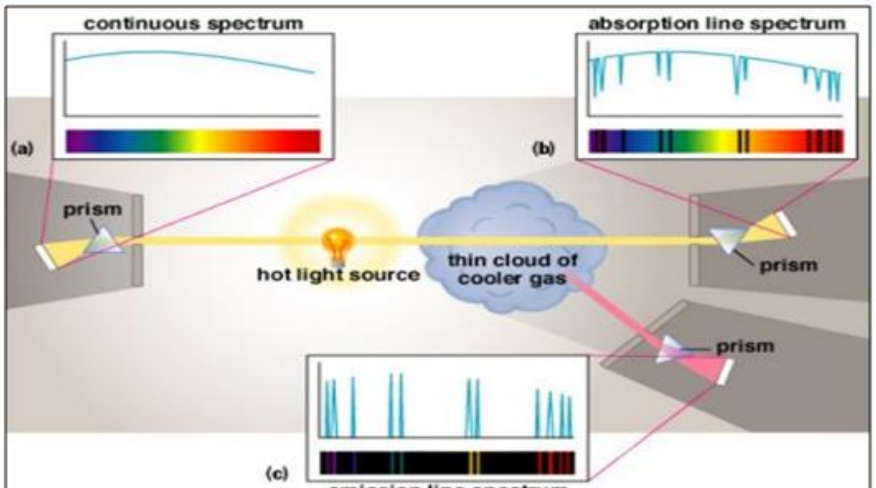
4.8.3.7 Flame Emission Spectroscopy This is an instrumental method of analysis. It is a technique used to identify metal ions in solution.



4.8.3.7 Calibration Curves

The readings for different concentrations of **metal ions** in solution are taken and can then be used to plot a calibration curve





4.8.3.7 Emission Line Spectrum (pl. spectra)

A big advantage of this technique over flame testing is that it can analyse mixtures of metal ions in solution….

1. Sample is **heated** in a flame
2. Electrons in the metal ions are **excited** by the thermal energy. As a result, they move to a **higher energy level**.
3. When the electrons **fall back** into a lower energy level, they **release energy** in the form of light.
4. The emitted **wavelengths** of light are analysed instrumentally, by a **spectroscope**.
5. The output is a line spectrum that is **compared** with reference line spectra of **known** metal ions.
6. Each metal ions produces a **unique emission spectrum** so can be **identified** and **concentration** measured.

Knowledge Organiser – 5.8 Chemical analysis

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  - ➢ Example: pure water only contains H_2O.
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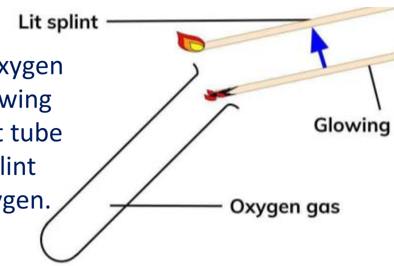

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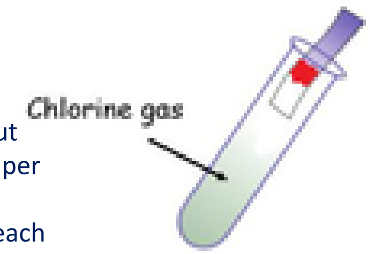

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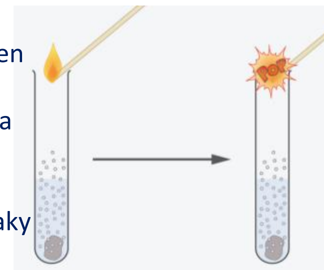
The test for hydrogen is to put a burning splint at the end of a test tube of gas. Hydrogen burns rapidly with a squeaky pop.



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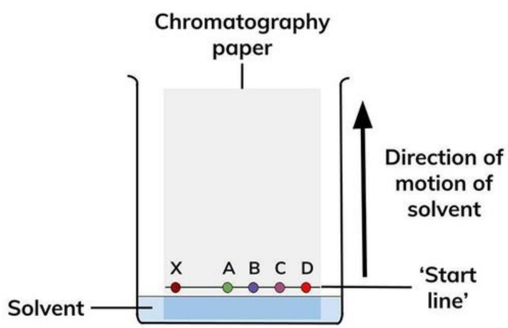

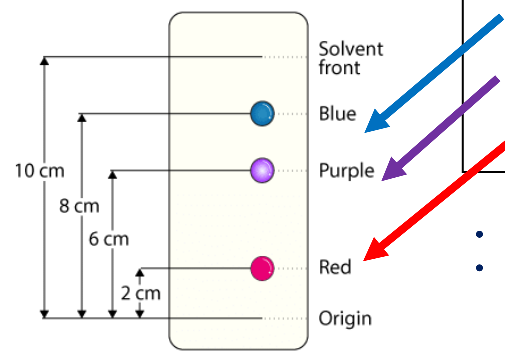

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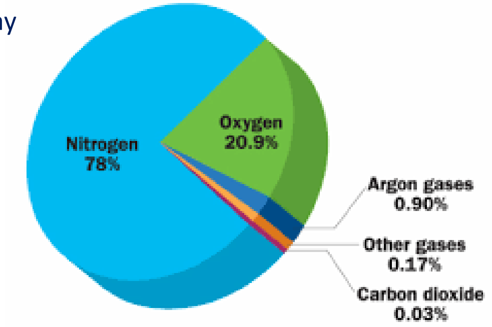
Knowledge Organiser – 5.9 Chemistry of the atmosphere

5.9.1.1 Proportion of different gases in the atmosphere

The **atmosphere** today has remained in the same proportions for 200 million years.



5.9.1.3 How oxygen increased

Algae and plants produced the oxygen now present in the atmosphere through **photosynthesis**.

This started 2.7 billion years ago and oxygen has steadily increased allowing more plants and animals to evolve.

5.9.1.1 How carbon dioxide decreased

There are **4** main reasons carbon dioxide **decreased**:

1. **Plants and algae** used CO_2 in **photosynthesis** to produce **oxygen**.

$$6CO_2 + 6H_2O \xrightarrow{\text{Sunlight}} C_6H_{12}O_6 + 6O_2$$

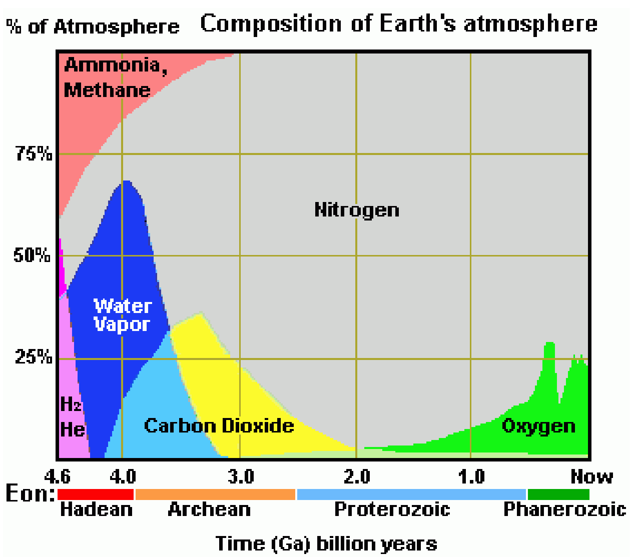
Carbon Dioxide Water Glucose (Energy) Oxygen

2. CO_2 was used up in the **formation of fossil fuels** (however, we are letting it back out!)
3. CO_2 is trapped when **sedimentary rocks formed**.
4. CO_2 **dissolved** in the newly formed **oceans.**

5.9.1.2 The Earth's early atmosphere

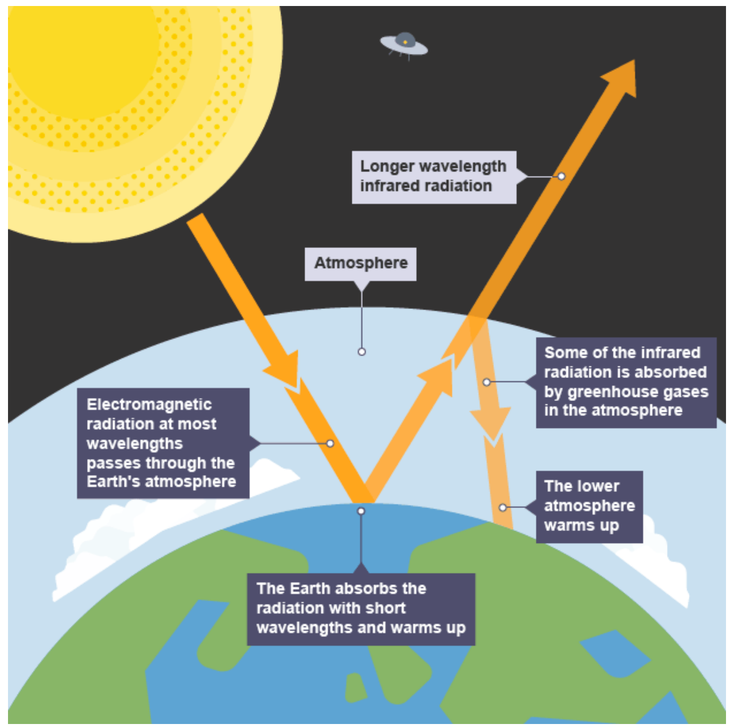
There are several theories about what was in the **early atmosphere** because we are looking 4.6 billion years ago. The main theory suggests:

- Initially a large amount of **volcanic** activity released gases to form the early atmosphere
- These gases were a mixture of mainly **carbon dioxide** and **water** with some **nitrogen**, **methane** and **ammonia**.
- Water vapour then **condensed** to form oceans.



5.9.2.1 Greenhouse gases

Greenhouse gases keep temperatures on Earth high enough to support life. The 3 main gases are **CO_2, methane and water vapour**.



5.9.2.2 Human activities which contribute to an increase in greenhouse gases in the atmosphere

Humans are increasing the levels of CO_2 and methane.

- **Carbon dioxide** can be increased through **burning fossil fuels** and **deforestation**.
- **Methane** levels can be increased by **mass livestock farming** and **rice farming**.

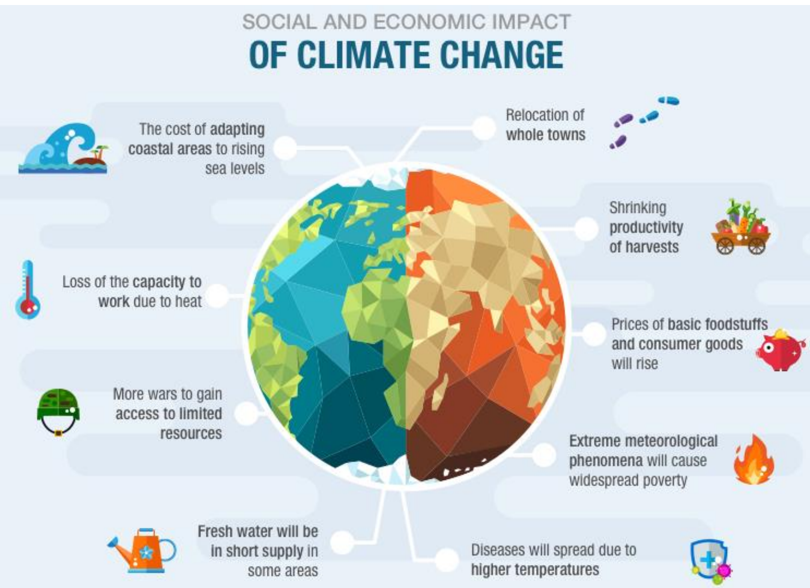
To know the fact from fiction about global warming we must look at **peer-reviewed articles**.
Simplified models can often be misrepresented in the media or maybe **biased**.

Knowledge Organiser – 5.9 Chemistry of the Atmosphere

5.9.2.3 Global Climate Change

An increase in average global temperatures is a major cause of **climate change**. Climate change could cause:

- **Melting ice** which could cause rising sea levels and flooding.
- **Ocean acidification** which could damage sea life.
- **More extreme weather** which could lead to housing damage, wildfires, droughts, flooding etc.
- **Changing climate temperatures** could mean different crop seasons and migrating species.



5.9.2.4 The carbon footprint and its reduction

The **carbon footprint** is the total amount of $\mathbf{CO_2}$ and other greenhouse gases **emitted** over the **full life cycle** of a product, service or event. To reduce a carbon footprint you need to reduce the amount of CO_2 and methane you produce.

Find out yours → https://footprint.wwf.org.uk/#/

5.9.3.1 Atmospheric pollutants and their sources

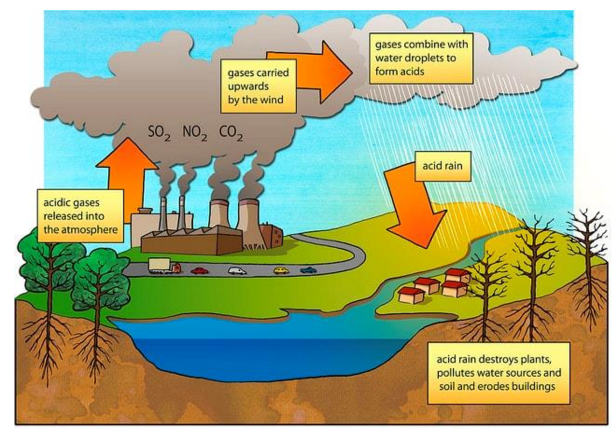
Most **pollutants** come from **combusting fuels**.

Products released can be **CO_2, H_2O, carbon monoxide, sulphur dioxide, and nitrogen oxides**.

Fuels may also produce solid **particulates** of carbon (soot).

5.9.3.2 Properties and effects of atmospheric pollutants

Product	Carbon monoxide	Sulphur dioxide and nitrogen oxides	Particulates
Problem?	Colourless, odourless, toxic gas.	Cause acid rain and respiratory problems.	Cause global dimming and health problems.



Knowledge Organiser – 5.10 Using resources

5.10.1.1 Using the Earth's resources

Humans use the Earth's resources to provide **warmth, shelter, food and transport**. Natural resources provide **food, timber, clothing, and fuels**.
Finite (non-renewable) resources are processed to provide **energy and materials**. These include **fossil fuels** and **metals**.
Renewable resources can provide **energy** and **materials** as well. These include **solar power** and **sustainable crop growth**.

5.10.1.3 Waste water

Urban life and **industry** produces a lot of **waste water**. This water must be treated (**cleaned**) before going back into the environment.

- **Sewage waste** requires removal of **organic matter** and **harmful microbes**. This treatment includes:
  - **Screening** and **grit removal**.
  - **Sedimentation** to produce sewage **sludge** and **effluent**.
  - **Anaerobic** digestion of sewage **sludge**.
  - **Aerobic** biological treatment of **effluent**.
- **Industrial** waste require removal of **organic matter** and **harmful chemicals**.

5.10.1.2 Potable water

Water is **essential** for life. Humans need water with **low levels** of dissolved **salts** and **microbes**. **Safe** water is called **potable water** (**not** the same as **pure** water. In the UK **rain** collects in **lakes and rivers**. Potable water is produced by:

- Choosing the right body of **fresh** water.
- Passing the water through **filter beds**.
- **Sterilising** using **chlorine**, **ozone**, or **UV** light.

If there isn't much fresh water, **sea water** can be **desalinated** by **distillation** or **reverse osmosis** but this takes a lot of **energy**.

Potable water RPA

This practical involves **testing** a sample of water then **purifying** the sample.
Step 1: To analyse the water, you use **universal indicator** to test the **pH** of 10 cm^3 of the water samples. You then evaporate **10 cm^3** of the samples and record the **mass** of **solids** that were **dissolved** in the water.
Step 2: Distilling the water. Set up the distillation equipment and **gently boil** the water. Collect the distilled water in a **cooled tube**. Test the **purity** of this water by finding the **boiling point.**

5.10.2.1 Life cycle assessments

Life cycle assessments (LCAs) are done to assess the environmental impact of products in each of the following stages:

- **Extracting** and **processing** raw materials.
- **Manufacturing** and **packaging**.
- **Use** and **operation** during its lifetime.
- **Disposal** at the end of its useful life, including **transport** and **distribution** at each stage.

It's easy to put numbers on the use of **water, resources, energy** sources, and some **waste** production. However, it is less easy to giving numbers to **pollutant effects**, so it is down to **judgement**.
This means that LCA is **not** a purely objective process, there is a **little guesswork**.

LCAs can be used to **evaluate** a product but **companies** may **leave parts out** to give a **misleading** representation for **advertising** (**biased**)

5.10.2.2 How to reduce resource use

By using less limited materials, the impact on the environment can be reduced. There are three ways to lessen environmental impact:

- ✓ Reduce use of limited resources. Metals, glass, building materials, clay ceramics and most plastics come from limited raw materials. The fuels for the processes come from limited resources. Mining from the Earth causes environmental impact.
- ✓ Reuse. Some products, such as glass bottles, can be reused. Glass bottles can be crushed and melted to make different glass products. Other products cannot be reused so are recycled.
- ✓ Recycle. Metals, plastics, glass and other materials can be melted and recasted into new products meaning no new materials are needed. Some materials need a lot of separation.

Knowledge Organiser – 5.10 Using resources

Comparative LCAs
Used to evaluate which of two alternative products will have a lower negative impact on the environment. For example, we can compare plastic carrier bags and paper carrier bags:

Life cycle stage	Plastic carrier bags	Paper carrier bags
Raw materials	Crude oil is a finite resource; fractional distillation, cracking and polymerisation all require a lot of energy.	Can be made from recycled paper, or from trees. Making paper from trees requires more energy than recycling paper, but much less than making plastics.
Manufacture	Cheaper to make large quantities of bags from plastic.	More expensive to make bags from paper because the handles must be glued on.
Use	Lower impact on the environment because plastic bags are usually stronger so they can be reused many times.	Relatively short lifetime; can only be reused a limited number of times.
Disposal	Can sometimes be collected and recycled; if disposed of as litter, they do not biodegrade; in landfill, may take decades or centuries to degrade.	Can be recycled easily; if disposed of in landfill, they biodegrade quickly.

Biological methods of metal extraction (HT)
The Earth's supply of metal **ores** is limited. Eg, high-grade copper ores are becoming harder to find and mine. There are some alternative methods to extract **metals** from low-grade copper ores that use living **organisms**.
These have advantages and disadvantages compared to the usual **extraction** methods.

Phytomining
Plants absorb **mineral ions** through their **roots**. **Phytoextraction** makes use of this:
- plants are grown on a **low-grade ore**
- the plants **absorb metal ions** through their roots and **concentrate** these ions in their **cells**
- the plants are **harvested** and **burnt**
- the **ash** left behind contains **metal compounds**

Phytoextraction is slow but it:
- reduces the need to obtain **new ore** by **mining**
- **conserves** limited supplies of **high-grade ores**
- reduces the amount of **rock waste** that must be disposed of after **traditional** mining
- Can be used to "clean up" industrial wastelands

Bioleaching
Certain **bacteria** can break down **low-grade ores** to produce an **acidic solution** containing **copper ions**. The solution is called a **leachate** and the process is called **bioleaching**.

Bioleaching does **not need high temperatures** but it produces **toxic substances**, including **sulfuric acid**, which damage the **environment**.

Processing the metal compounds
Iron is **more reactive** than copper. It can **displace copper** from the **leachate**. For example:

iron + copper sulfate → iron(II) sulfate + copper

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

Since **iron is cheaper** than copper, the use of **scrap iron** is a **cost-effective** way to produce copper from the **leachate**.

Alternatively, the copper compounds can be **dissolved** and the solution **electrolysed** to produce **copper metal.**

Knowledge Organiser – 5.10 Using resources (Separate Chemistry)

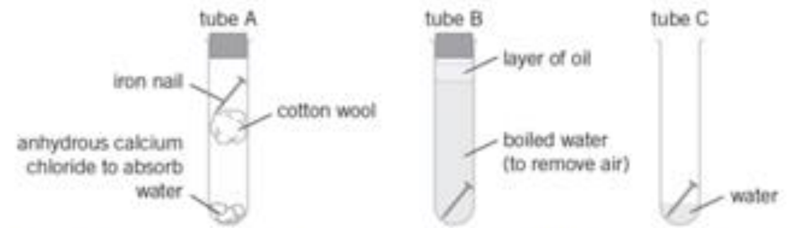
Key Terms

Key Terms	
Corrosion	**Breakdown of materials** due to chemical reactions. It is a form of **erosion**.
Rusting	The **corrosion** of **iron**.
Rust	Rust is hydrated Iron(III)oxide.
Sacrificial protection	An effective way to **prevent rusting** whereby a metal **more reactive than iron** is attached to or coated on an object.
Galvanised	Iron or steel objects that have been **protected from rusting** by a thin layer of **zinc metal** at their surface.
Oxidation	**Loss** of electrons.
Reduction	**Gain** of electrons.
Reducing agent	Tend to get **oxidised themselves** (and hence reduce other species).
Alloy	A mixture of two or more elements, at least one of which is a metal. For e.g. Steel is an alloy of Iron and carbon.
Bronze	Alloy of copper and tin.
Brass	Alloy of copper and zinc.
Steels	Alloys of iron containing specific amounts of carbon and/or other metals.
Hydrated	A substance that contains water in its crystals.
Polymers	A substance made from very large molecules, polymers are made up of many repeating units.
Thermosoftening polymers	Soften and melt when they are heated. Can be remoulded.
Thermosetting polymers	Do not melt when they are heated. Cannot be remoulded.
Composites	Two materials combined to make a material with useful properties.
Ceramics	Materials made by heating clay to high temperatures making hard materials which are excellent insulators.

Rusting

For iron to rust, both **air** and **oxygen** are needed. Providing a barrier between iron either air (oxygen) and water protects the iron from rusting.

Iron + oxygen + water → hydrated iron(III)oxide



Tube A tests to see if air alone makes iron rust. Tube B tests to see if water alone will make iron rust. Tube 3 tests to see if air and water will make iron rust. **Rusting is only observed** in **tube 3** illustrating that both **air and water** are **needed** for iron to rust.

Sacrificial protection provides **protection against rusting**. The iron needs to be attached to a more **reactive metal** (**galvanising** it) for e.g. Zinc, magnesium or aluminium. The zinc is a **stronger reducing agent** than iron, so it has a stronger **tendency to form positive ions** by giving away electrons. As the zinc atoms lose electrons they become **oxidised**. Therefore any water or oxygen reacts with the zinc instead of the iron (protecting the iron from oxidation).

Useful alloys

Alloys are **harder** than **pure** metals because the **regular layers** are **distorted** by **differently sized atoms** and hence **cannot slide**.

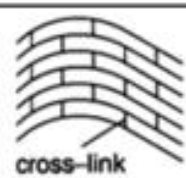
Pure iron is too **soft** for it be useful in its pure form. Steel is an alloy of iron which contains **carefully controlled quantities of carbon** so that it's hardness is controlled.

Steels	**Properties**	**Uses**
High carbon steel	Very hard but brittle	Cutting tools (chisels)
Low carbon steel	Softer but easily shaped	Bodies of cars
Stainless steel	Chromium-nickel steels resistant to corrosion	Cooking utensils, cutlery
Nickel steel alloys	Resistant to stretching	Bridges, bicycle chains

Knowledge Organiser – 5.10 Using resources (Separate Chemistry)

The properties of polymers

The properties of polymers depends on what monomers they are made from the conditions under which they are made.

Thermosoftening polymers	Soften or melt easily when heated because their **intermolecular forces** between the chains are **weak**.	
Thermosetting polymers	Contain **crosslinks** (strong covalent bonds) between chains so they do not soften or melt easily.	
High density polyethene	Made using very high pressures and trace of oxygen. Polymer chains are randomly **branched, can't pack closely** together resulting in a **low density.**	
Low density polyethene	Made using a catalyst at 50°C and a slightly raised pressure. Made of **straight chain** molecules which are **closely packed**, stronger and more **dense**.	**Straight chain** 

Glass, ceramic and composites

Glass	The most common form of glass is Soda Glass which is made by heating a mixture of sand (SiO_2), limestone ($CaCO_3$) and sodium carbonate (soda) at 1500°C. As it cools down the glass turns into a solid. Different types of glass exist depending on amounts of each of the reactants; borosilicate glass involves an extra compound- B_2O_3. • Atoms arranged irregularly • Transparent, brittle, high melting point, keeps its shape (not flexible)
Ceramics	Wet clay is moulded into a desired shape, then heated in a furnace to 1000°C • Used in bricks, tiles, crockery, bathroom furniture • Atoms are held together in a giant covalent lattice, generally in a regular pattern • Hard but brittle, electrical insulators
Composites	Materials made from two or more different materials, with one material acting as a binder for the other material, **reinforcing it.** Usually fibres or fragments of one material are held in a 'matrix' (network of atoms) by the other. • Glass-ceramic composites are very hard and tough (not brittle) • Fibreglass (polymer-ceramic) is a low density, tough, flexible material- e.g. used in kayaks • Plywood, carbon fibres and cement are other examples

The Haber process

The Haber process is used to manufacture ammonia, which can be used to produce nitrogen-based fertilisers. The **raw materials** are **nitrogen** (from the air) and **hydrogen** (from natural gas, mainly **methane**).

The nitrogen and hydrogen are purified then passed over an **iron catalyst** at a **high temperature** of 450°C and a **high pressure** (200 atmospheres) to make **ammonia** NH_3.

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

The reaction is **reversible** so ammonia can break down again into nitrogen and hydrogen. The ammonia is removed by cooling the gases so that the ammonia liquefies. It can then be separated from the unreacted nitrogen and hydrogen gas.

The unreacted nitrogen and hydrogen gases are recycled back into the reaction mixture so that they can react again on the surface of the iron catalyst.

Knowledge Organiser – 5.10 Using resources (Separate Chemistry)

The Haber process key terrms	
Reversible reaction	A reaction in which the **products can also form the reactants**. Its symbol is ⇌ Shown as: A + B ⇌ C + D
Exothermic	A reaction that **transfers energy to** the **surroundings**
Endothermic	A reaction that **takes in energy from** the **surroundings**
Equilibrium (HT)	Equilibrium is reached when the **forward and backwards reactions** occur at **exactly the same rate**. The **amounts of reactants and products present remain constant**. Requires a **sealed container**.
Le Chatelier's Principle (HT)	When a **change in conditions** is introduced to a system at equilibrium, the **position of equilibrium shifts** so as **to cancel out the change.**

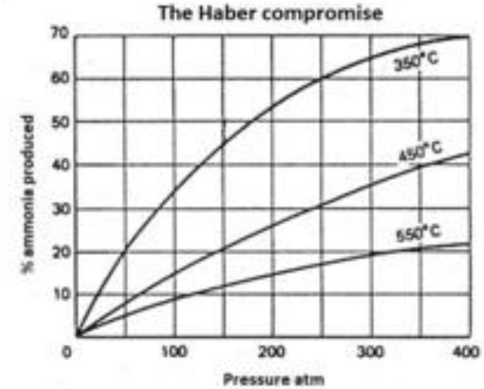
Changing conditions in the Haber Process

Equation for the Haber process: $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$
ΔH is negative (exothermic in forwards direction).

Changing temperature	The Haber process is an **exothermic** process (ΔH is negative). If the temperature is **decreased**, the equilibrium moves to the **exothermic side** and **more NH_3** is made.
Changing the pressure	**Increasing the pressure** results in the equilibrium moving to the right hand side as there are **less gas molecules**.
Catalyst	The iron catalyst speeds up the rate of the forwards and backwards reaction **equally**, hence it **doesn't affect the yield** of ammonia but does result in ammonia being produced **quicker**.

The Haber compromise (HT)

Lowering the temperature **slows** down the rate of reaction, taking **longer** for ammonia to be produced.
Increasing the pressure means stronger, more **expensive equipment** is needed. This **increases** the **cost** of producing ammonia.
Hence a **compromise** is reached achieving an **acceptable yield** in a **reasonable timeframe** while keeping **costs down**.
A pressure of **200 atmospheres** and a temperature of **450°C**.



Fertilisers

Compounds of nitrogen, phosphorus and potassium are used as fertilisers to improve agricultural productivity.

NPK fertilisers	NPK fertilisers contain compounds of all three elements. **Nitrogen** for cell growth and making proteins in plants **Phosphorus** needed to make DNA **Potassium** needed to make enzymes involved in respiration and photosynthesis.
Synthesis	Fertilisers are made by reacting an **acid** and **base** together e.g. Ammonia + nitric acid → ammonium nitrate Ammonia + phosphoric acid → ammonium phosphate Ammonia + sulphuric acid → ammonium sulfate
Obtaining raw materials	Phosphates are obtained from **phosphate rocks**. Phosphate rocks all contains the phosphate ion PO_4^{3-}. The rocks are **insoluble so cant be used directly as fertilisers,** but react with **acids** to make the **soluble phosphate** compounds. Potassium chloride and potassium sulfate are obtained by **mining** and are **soluble** so **can be directly used** as fertilisers. Nitric acid is required to make nitrate fertilisers (ammonia from the Haber process is oxidised to make nitric acid).
Phosphate rock fertilisers	Phosphate rock + nitric acid → phosphoric acid + calcium nitrate Phosphate rock + sulphuric acid → calcium phosphate + calcium sulfate Phosphate rock + phosphoric acid → calcium phosphate

Knowledge Organiser – 5.10 Using resources

5.10.1.1 Using the Earth's resources

Humans use the Earth's resources to provide **warmth, shelter, food and transport**. Natural resources provide **food, timber, clothing, and fuels**.
Finite (non-renewable) resources are processed to provide **energy and materials**. These include **fossil fuels** and **metals**.
Renewable resources can provide **energy** and **materials** as well. These include **solar power** and **sustainable crop growth**.

5.10.1.3 Waste water

Urban life and **industry** produces a lot of **waste water**. This water must be treated (**cleaned**) before going back into the environment.

- **Sewage waste** requires removal of **organic matter** and **harmful microbes**. This treatment includes:
  - **Screening** and **grit removal**.
  - **Sedimentation** to produce sewage **sludge** and **effluent**.
  - **Anaerobic** digestion of sewage **sludge**.
  - **Aerobic** biological treatment of **effluent**.
- **Industrial** waste require removal of **organic matter** and **harmful chemicals**.

5.10.1.2 Potable water

Water is **essential** for life. Humans need water with **low levels** of dissolved **salts** and **microbes. Safe** water is called **potable water** (**not** the same as **pure** water. In the UK **rain** collects in **lakes and rivers**. Potable water is produced by:

- Choosing the right body of **fresh** water.
- Passing the water through **filter beds**.
- **Sterilising** using **chlorine**, **ozone**, or **UV** light.

If there isn't much fresh water, **sea water** can be **desalinated** by **distillation** or **reverse osmosis** but this takes a lot of **energy**.

Potable water RPA

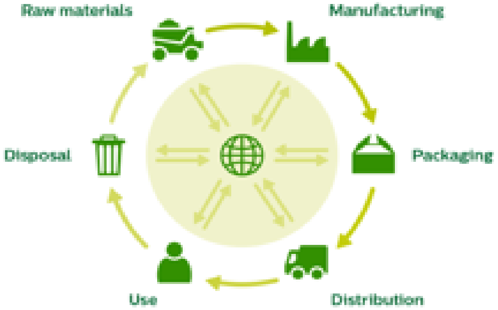
This practical involves **testing** a sample of water then **purifying** the sample.
Step 1: To analyse the water, you use **universal indicator** to test the **pH** of 10 cm^3 of the water samples. You then evaporate **10 cm^3** of the samples and record the **mass** of **solids** that were **dissolved** in the water.
Step 2: Distilling the water. Set up the distillation equipment and **gently boil** the water. Collect the distilled water in a **cooled tube**. Test the **purity** of this water by finding the **boiling point.**

5.10.2.1 Life cycle assessments

Life cycle assessments (LCAs) are done to assess the environmental impact of products in each of the following stages:

- **Extracting** and **processing** raw materials.
- **Manufacturing** and **packaging**.
- **Use** and **operation** during its lifetime.
- **Disposal** at the end of its useful life, including **transport** and **distribution** at each stage.

It's easy to put numbers on the use of **water, resources, energy** sources, and some **waste** production. However, it is less easy to giving numbers to **pollutant effects**, so it is down to **judgement**.
This means that LCA is **not** a purely objective process, there is a **little guesswork**.



LCAs can be used to **evaluate** a product but **companies** may **leave parts out** to give a **misleading** representation for **advertising** (**biased**)

5.10.2.2 How to reduce resource use

By using less limited materials, the impact on the environment can be reduced. There are three ways to lessen environmental impact:

✓ **Reduce** use of limited resources. Metals, glass, building materials, clay ceramics and most plastics come from limited raw materials. The fuels for the processes come from limited resources. Mining from the Earth causes environmental impact.

✓ **Reuse.** Some products, such as glass bottles, can be reused. Glass bottles can be crushed and melted to make different glass products. Other products cannot be reused so are recycled.

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Knowledge Organiser – 5.10 Using resources

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Used to evaluate which of two alternative products will have a lower negative impact on the environment. For example, we can compare plastic carrier bags and paper carrier bags:

Life cycle stage	Plastic carrier bags	Paper carrier bags
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Plants absorb **mineral ions** through their **roots**. **Phytoextraction** makes use of this:
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