ZNOTES.ORG

UPDATED TO 2023-2025 SYLLABUS

CAIE IGCSE CHEMISTRY

SUMMARIZED NOTES ON THE THEORY SYLLABUS

1. States of Matter

1.1. Solids, Liquids and Gases

- States of Matter are the different forms in which matter can exist
- The three states of matter are: **Solid, Liquids and Gases**

Properties of Solid, Liquid and Gases

	Solids	Liquids	Gases
Shape	have a definite shape	take the shape of their container	take the shape of their container
Density	high	moderate to high	low
Volume	have a fixed volume	have a fixed volume	don't have a fixed volume - expand to fill the container
Fluidity	non-fluid in nature	fluid in nature	fluid in nature
Particle Space	closely packed together, barely any inter-molecular spaces	closely packed together, with barely any inter-molecular	presence of large, inter-molecular spaces
Particular Movement	particles are immobile; however, they can vibrate about their fixed positions	particles can move past one another	particles are mobile and move randomly
Particular Arrangement	particles are arranged regularly in a lattice	particles are arranged irregularly	particles are arranged irregularly

1.2. Changes of States

Boiling and Evaporation

Boiling and Evaporation are both **endothermic processes**:

- Boiling and Evaporation are processes which involve the conversion of a liquid into a gas.
- Both processes allow **molecules to move further apart from each other**.

Boiling	Evaporation
happens at a set temperature called the boiling point	occurs at any temperature below the boiling point and above the freezing point (liquid)
occurs throughout the liquid	only occurs at the surface
relatively faster process	a slow process

Condensation

Condensation is the process by which a gas converts into a liquid. It happens at the same temperature as the boiling point.

- As temperature decreases, the **energy of particles will decrease**, making it move more slowly.
- Condensation and Freezing are both **energy-given out reactions**.

Freezing, Melting and Sublimation

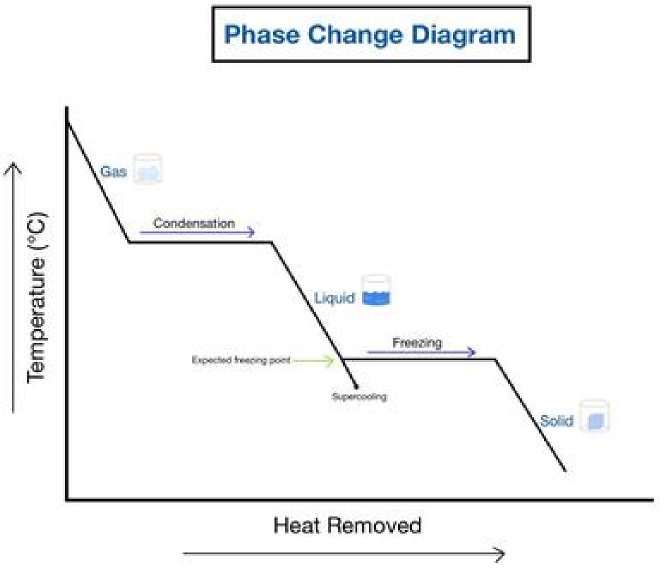
1. Melting is the process in which a solid converts to a liquid. It happens at a set temperature called the melting point.
2. Freezing is the process in which a liquid converts to a solid. It happens at the same temperature as the freezing point.
3. Sublimation occurs when a solid has enough energy to convert into a gas or gas converted into a solid.

1.3. Cooling and Heating Curves

Cooling Curves

1. On cooling, the particles of a gas move slower and slower and the gas contracts
2. The particles are now closer together, and inter-molecular bonds start to form between them once the condensation point is reached
3. The temperature of the substance stops falling; the energy released by bond formation cancels out the energy lost due to cooling
4. Once all the gas has turned into liquid, the temperature starts to fall again, and the liquid begins to contract until the freezing point is reached
5. At the freezing point, inter-molecular bonds between the liquid molecules start to develop to form a solid
6. At the freezing point, the temperature remains constant until all of the liquid has solidified since the energy released due to bond formation cancels out the energy lost due to cooling

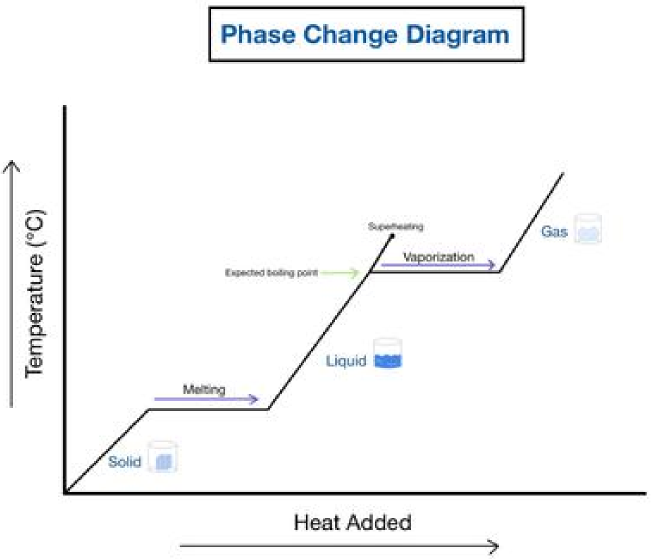
The following curve is obtained if this phenomenon is plotted - the cooling curve:



Heating Curves

1. On heating, the particles of a solid start to vibrate faster and faster in their mean positions, and the solid begins to expand
2. Once the melting point is reached, the inter-molecular bonds between the particles begin to breakdown
3. The temperature of the substance at this point remains constant until all the solid has turned into a liquid because the energy received by the system is cancelled out by the energy used to break intermolecular bonds
4. Once all the solid has turned to liquid, the temperature starts to rise again, and the liquid begins to expand until the boiling point is reached
5. At the boiling point, the intermolecular bonds between the liquid molecules start to break down to form gas
6. At the boiling point, the temperature of the substance remains constant until all of the liquid has vaporised since the energy absorbed by the substance is cancelled by the energy used for the breakdown of intermolecular bonds

The following curve is obtained if this phenomenon is plotted - the heating curve:



1.4. Effects of Temperature and Pressure on Gas

Gases are compressible. By changing the pressure acting on them, their volume may be influenced.

1. An increase in external pressure produces a contraction (decrease) in volume. The gas is said to be compressed.
2. A fall in external pressure produces an expansion (increase) in volume. The gas is said to be decompressed.

The volume of gases may also be influenced by temperature. The temperature of a gas affects its internal pressure and, thereby, its volume.

1. When temperature increases, the gas molecules have increased kinetic energy and hit the walls of their container more often and with greater force. This causes an increase in internal pressure and an increase in volume.
2. When the temperature decreases, the gas molecules have decreased kinetic energy and hit the walls of their container less often and with attenuated force. This causes a decrease in internal pressure and a decrease in volume.

1.5. Diffusion

Diffusion: the net movement of particles from a region of higher concentration to a region of lower concentration as a result of their random movement until equilibrium is reached.

- The rate of diffusion is most rapid in gases > liquids > solids.

1.6. Effect of Relative Molecular Mass in Diffusion

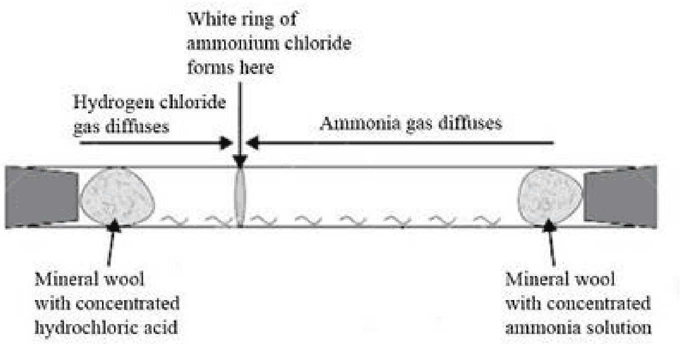
- The rate at which gases diffuse differs and depends on the gas molecules' mass.
- At the same temperature, molecules with a lower mass move faster on average than those with a higher mass.

The HCL and NH_3 Experiment

The experiment involved placing a cotton wool plug soaked in hydrochloric acid (HCl) at one end of a glass tube and a cotton wool plug soaked in ammonia (NH3) at the other.

The tube was left undisturbed for some time, during which the gases diffused towards each other. The diffusion rate was measured by observing the distance travelled by each gas after a fixed time. It was found that ammonia, with a lower molecular mass, diffused faster than hydrochloric acid, which has a higher molecular mass.

Conclusion: Ammonia has a **smaller Mr** than hydrochloric acid, so ammonia has a **greater rate of diffusion**. Therefore, the white ring of ammonium chloride **forms near the hydrochloric acid.**



2. Atoms, Elements and Compounds

Atom: The smallest particle of matter made of protons, neutrons and electrons

2.1. Elements, Compounds and Mixtures

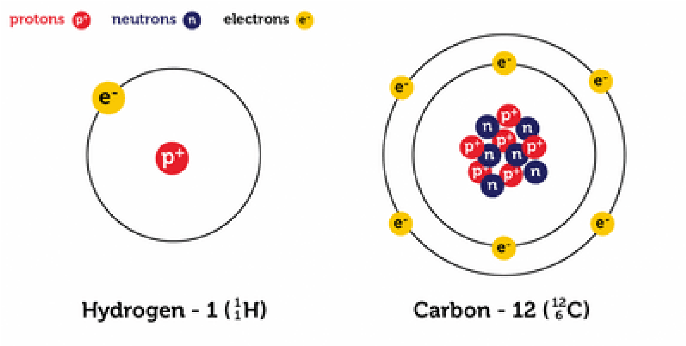
- **Elements**: Group of atoms that share the **SAME number of protons**. It cannot be broken down into simpler substances by chemical methods.
- **Compounds**: Two or more elements **chemically bonded together.**
- **Mixtures**: Two or more elements **not chemically bonded together.**

2.2. Atomic Structure

- An atom comprises three subatomic particles: protons, neutrons and electrons.
- The atom is mostly space with a positively charged nucleus consisting of protons and neutrons in the centre and electrons in the space around the nucleus (held together by the electrostatic force of attraction between them and the positively charged nucleus)
- The characteristics of neutrons, protons and electrons are as follows:

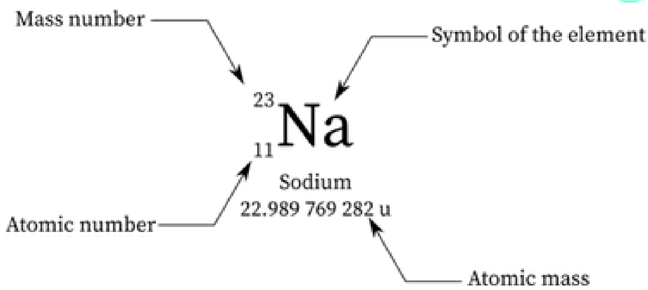
Subatomic particle	Relative Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	1/1840	-1

- Since electrons and protons have opposing and equal charges, **the atom's overall charge is neutral.**
- Neutrons have the purpose of holding the nucleus together. The larger the nucleus gets, the more are the neutrons required to hold the nucleus together



2.3. Proton and Nucleon Number

- The **Proton Number (Atomic Number)** is **the number of protons in the nucleus of an atom**.
- It is unique to each element. It is denoted by the letter "Z". For a neutral atom, the number of protons and electrons are equal; therefore, the proton number (Z) also corresponds to the number of electrons.
- The **Nucleon Number (Mass Number)** is **the total number of protons and neutrons in the nucleus of an atom.**
- The following format is shown below:



2.4. Electronic Configurations of Elements & Ions

Atoms have electrons that orbit around a central nucleus, and these orbits are referred to as electron shells.

- The energy levels of the shells increase as their distance from the nucleus increases.
- The first shell has a max capacity of **2 electrons**, while the subsequent shell can hold up to **8**.

For this syllabus, we only need to know the **general full electronic configuration as (2.8.8)**
(a) Group VIII noble gases have **a full outer shell**
(b) the **number of outer shell electrons** is equal to the **group number in Groups I to VII**
(c) the number of occupied electron shells is **equal to the period number**

2.5. Isotopes

Isotopes: different atoms of the same element that have the same number of protons but different numbers of neutrons.

- The isotopes of an element have the **same chemical properties** because **they contain the same number of outer shell electrons** and therefore have the **same electronic configuration**.
- **Radioisotopes:** certain isotopes such as Carbon-14 and Tritium have a nucleus so heavy that they are radioactive in nature. Their nucleus is **unstable and breaks up spontaneously.**

2.6. Relative Atomic Masses

- Most elements exist naturally as a mixture of their isotopes. Using the data on the abundance of these naturally occurring isotopes, we can calculate the mass relative atomic mass of the element.
- **NOTE**: for all purposes, the mass numbers of elements have been rounded off to the nearest whole number; however, only Chlorine is used with its actual mass number of 35.5. This is for the sake of simplicity of calculation
- An example for calculating the relative mass and abundance:

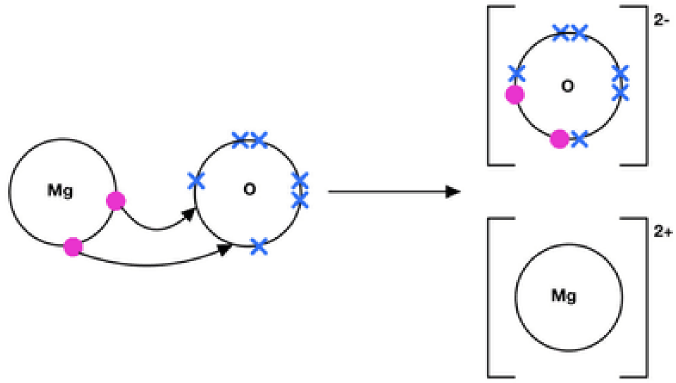
Q. Iridium has two isotopes. These isotopes are Iridium - 191 and Iridium - 193. A natural sample consists of of 37.3% of Iridium - 191. Calculate the relative atomic mass (A_r) of the natural sample of Iridium

A. $\rightarrow$
Step 1. Identify the percentage of Iridium - 193
$\rightarrow$ if the sample consists of 37.3% of Iridium - 191, it must consist of 100% - 37.3% of Iridium - 191
$\rightarrow$ 100 - 37.3 = 62.7%
$\rightarrow$ the sample consists of 62.7% of Iridium - 193
Step 2. Consider a sample of 100 atoms of Iridium. In that sample, 37.3 of atoms should have a mass of 191 and 62.7 atoms should have the mass of 193
THEN
average mass $= \dfrac{(37.3 \times 191) + (62.7 \times 193)}{100} = 192.2$ (4 significant figures)
Answer: 192.4

2.7. Ions and Ionic Bonds

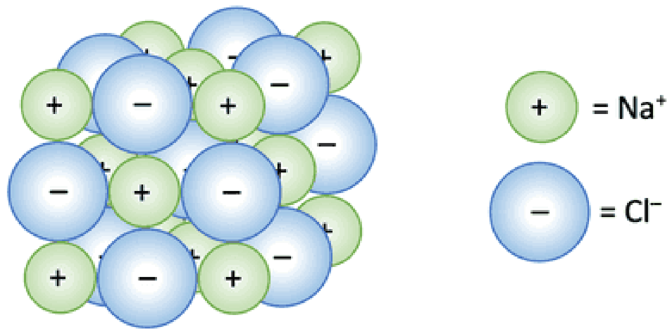
Cations: Positive Ions
Anions: Negative Ions
Ionic Bonds: strong electrostatic attraction between oppositely charged ions (metals + non-metals)

Ionic Compounds Dot and Cross Diagram



Typical ionic compounds are usually **giant lattice structures** with a **regular arrangement of alternating positive and negative ions.**

- Using an X and a dot in your drawings will help differentiate the two ions.



Properties of Ionic Compound

1. High Boiling and Melting Point: A lot of energy is needed to overcome the strong electrostatic forces between oppositely charged ions
2. Good Electrical Conductivity when **molten or aqueous**:
   - As the charges flow, ions can move freely in an aqueous/molten state.
   - Ions are not free to move when solid state, as the **charges cannot flow.**
3. Ionic Compounds are **soluble in water**: The positive and negative ions are **attracted to water.**
4. **Solid in Room Temperature**: Regular arrangement of ions in a lattice with ions of opposite charges next to each other

Other Properties of Ionic Compound

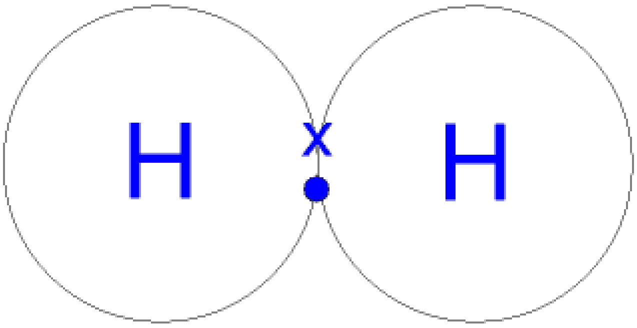
1. Brittle

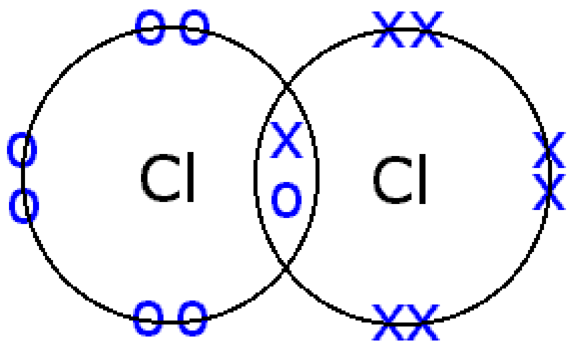
2. Low Volatility

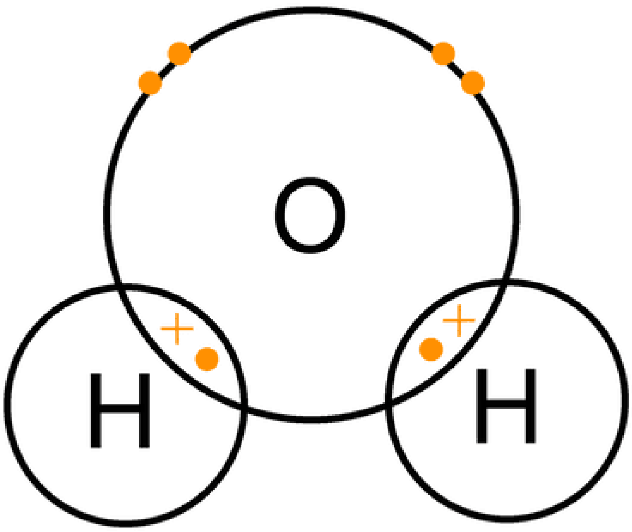
2.8. Simple Molecules and Covalent Bonds

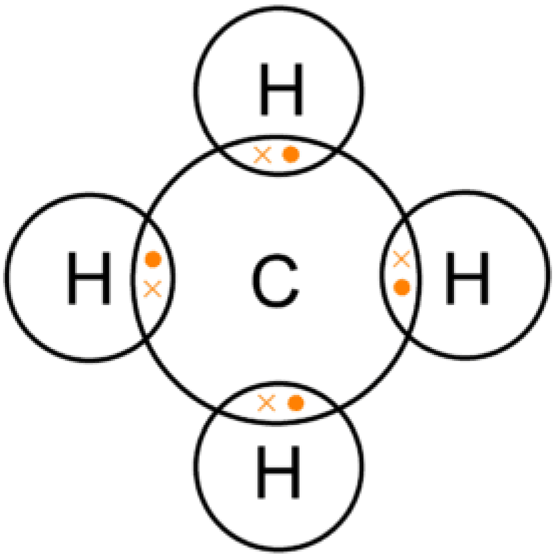
Covalent Bonding: Pairs of electrons shared between two atoms leading to noble gas electronic configuration (2.8.8)

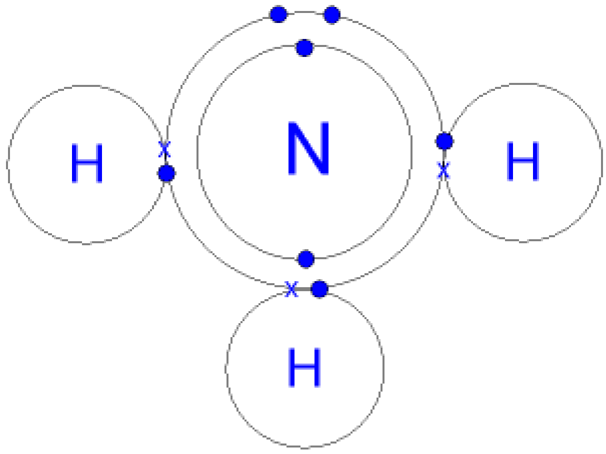
Covalent Bonds Dot and Cross Diagram

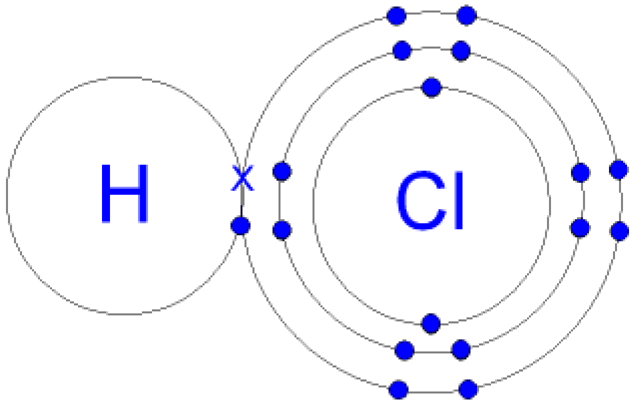












Different Types of Covalent Bonds

1. Single Bonds - e.g., Chlorine
2. Double Bonds - e.g., Carbon Dioxide
3. Triple Bonds - e.g., Nitrogen

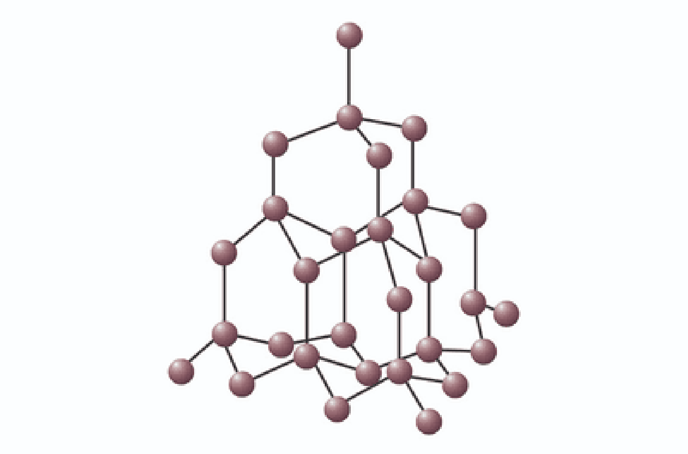
Properties of Covalent Compound

1. The intermolecular forces in covalent compounds are **weak** but have **strong covalent bonds**.
2. Covalent Compounds have low melting and boiling point. They require less energy to break the weak intermolecular forces (same as attractive forces).
3. Poor Electrical Conductivity - **No free electrons or ions present** to carry an electrical current

2.9. Giant Covalent Structures

Giant Covalent (Macromolecular) Structures: solids with very high melting points, where all the atoms are made of pure carbon.

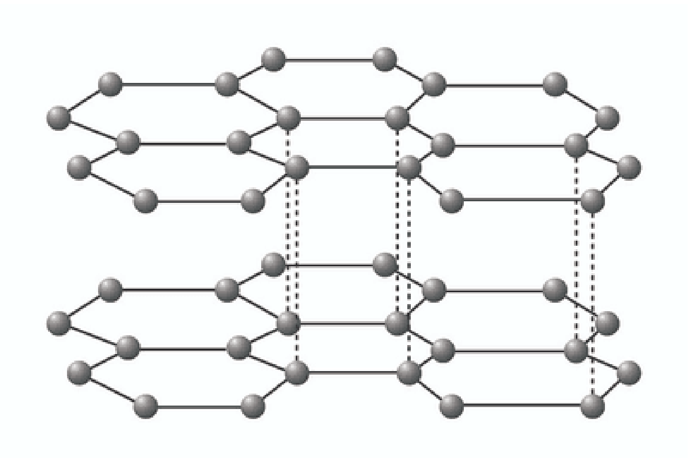
Diamond



Properties

1. Each carbon atom is joined with **four other carbon atoms**
2. **High Melting and Boiling Points** - Strong Covalent Bonds
3. **No Delocalised/Free Moving Electrons**
4. It cannot be scratched easily
5. Transparent colour (Extra information)
6. Cannot conduct electricity **due to no free-moving electrons**
7. **Hard** in structure
8. **Giant Lattice Arrangement**
9. Uses are for **cutting tools**

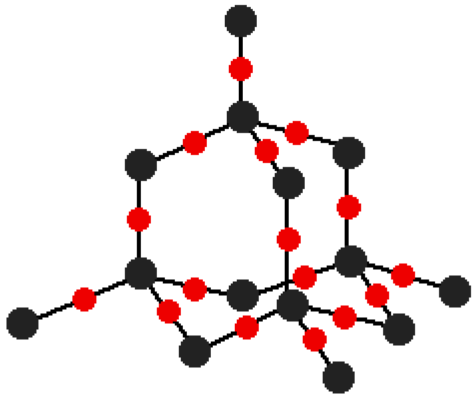
Graphite



Properties

1. Each carbon atom is joined with **three other carbon atoms**
2. **High Melting and Boiling Points** - Strong covalent bonds within the layers (but the layers are attracted to each other by weak intermolecular forces)
3. Contains **Delocalised/Free Moving Electrons**
4. It can be scratched easily
5. Opaque/Black
6. Can conduct electricity **due to free-moving electrons**
7. Soft - **Layers can slide easily**
8. **Layers of hexagonal rings held by weak intermolecular forces**
9. Uses are for **lubricant and electrode** in Electrolysis

Silicon (IV) Oxide (SiO_2)



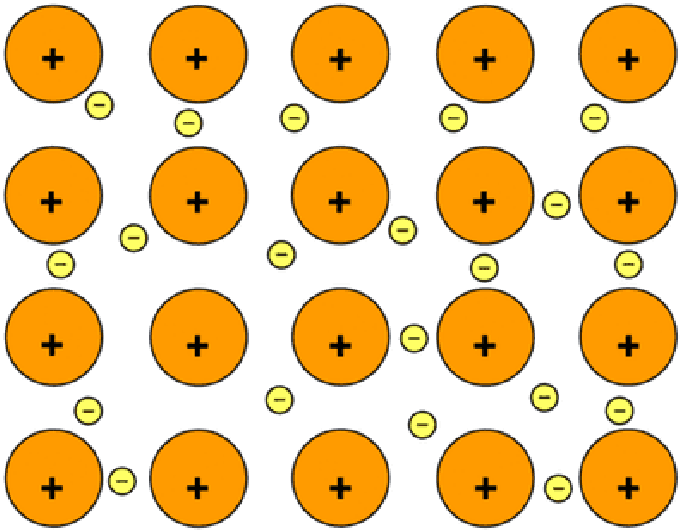
Properties

1. The structure of Silicon (IV) Oxide is **similar/resemblance to that of a diamond.**
2. **Hard Structure**
3. **High melting and boiling point** - More energy to overcome
4. **Rigid Tetrahedral Structure**
5. Does **not conduct electricity**

- **Each Silicon atom** is covalently bonded with **4 Oxygen Atoms**
- **Each Oxygen atom** is covalently bonded with **2 Silicon Atoms**

2.10. Metallic Bonding

Metallic Bonding: the electrostatic attraction between the positive ions in a giant metallic lattice and a "sea" of delocalised electrons.



Properties of Metallic Bonding

1. Metallic Bonds have good electrical conductivity: **Delocalised electrons can move through the structures and carry current.**
2. High Melting and Boiling Point: **More energy to overcome strong forces of attraction between positive metal ions and the sea of delocalised electrons & Vibrate/Transfer Heat**
3. Malleability: Can be **hammered into shapes** as **layers can slide over each other.**
4. Ductility: Can be **drawn into thin wires**

3. Stoichiometry

3.1. Formulae

Charges of Elemental Groups

Group	Charges
I	1
II	2
III	3
IV	+/- 4
V	-3
VI	-2
VII	-1
VIII	0

Charges of Common Ions

Name	Formula	Charges
Ammonium	NH_4^+	+1
Nitrate	NO_3^-	-1
Hydroxide	OH^-	-1
Carbonate	CO_3^{2-}	-2
Sulfate	SO_4^{2-}	-2

3.2. State Symbols & Word Equations

- Balancing equations: A chemical equation is balanced when there are an equal number of atoms and charges on both sides of the equation
- **State symbols:**
  - (s) = solid
  - (l) = liquid
  - (g) = gas
  - (aq) = aqueous solution
- Names of compounds
  - A compound ending with **-ide** only contains two different elements
  - A compound ending with **-ate** contains oxygen

3.3. Relative Masses of Atoms and Molecules

- **Relative Atomic Mass (A_r):** the average mass of the isotopes of an element compared to 1/12th of the mass of an atom of 12C.
- **Relative Molecular Mass (M_r):** sum of relative atomic masses of all the atoms in one molecule of the compound.

3.4. The Mole and the Avogadro Constant

A **mole** of a substance is the amount that contains the same number of units as the number of carbon atoms in 12 grams of carbon-12

- A mole is the A_r or M_r expressed in grams e.g. 1 mole of Carbon-12 is equal to 12 grams.
- It is equal to 6.02×10^{23} particles; this number is called Avogadro's constant.
- 1 mole also occupies a volume of $24dm^3$ at room temperature and pressure

3.5. Number of Moles

$$\text{Amount of Substance} = \frac{\text{mass}}{\text{molar mass}}$$

- Amount of Substance (mol)
- Mass (g)
- Molar Mass (Mr) in (g/mol)

3.6. Number of Moles in Aqueous Solution

$\text{Moles} = \text{concentration}$ x volume

- Moles (mol)
- Concentration (mol/dm^3)
- Volume (24 dm^3)

3.7. Moles in Gases

$$\text{Volume} = \text{No. of Moles} \times 24\text{dm}^3 \text{at r.t.p}$$

- 1 dm^3 = $1000 cm^3$

3.8. Concentration

$$Concentration = \frac{\text{no. of moles}}{\text{volume}}$$

- Moles per dm^3
  - $1 mol/dm^3$
- Grams per dm^3, g/dm^3

Concentration can be measured and converted into g/dm^3 or mol/dm^3, by multiplying the molar mass of the compound.

3.9. Moles Calculation in Acid-Base Titration

$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

20.0 cm^3 of 0.0500 mol/dm^3 HCl reacts with the 25.0 cm^3 of $Ca(OH)_2$.

Determine the concentration of $Ca(OH)_2$ in g/dm^3. Use the following steps.

- Calculate the number of moles in 20.0 cm^3 of 0.0500 mol/dm^3 HCl.

.... mol

We will use the formula:

Concentration x Volume (cm^3)/1000

20 x 0.0500 divide 1000 = 0.001

3.10. Empirical Formulae

- This is the simplest ratio of the atoms in a compound
- For example:
  - Molecular formula of ethane= C_2H_6
  - Empirical formula of ethane = CH_3
- To find out the empirical formula you:
  - Make the percent ratio into the simplest whole-number ratio (NOTE: if given %s, use them as grams)
  - Divide the coefficients of each element symbol by the lowest coefficient

3.11. Molecular Formulae

- It shows the actual number of atoms in one molecule of a substance.

3.12. Percentage Purity & Yield

- $\text{Percentage Purity} = \frac{\text{mass of product (pure)}}{\text{mass of compound (impure)}} \times 100$
- $\text{Percentage Yield} = \frac{\text{actual mass obtained}}{\text{calculated/theoretical mass}} \times 100$

3.13. Percentage Composition by Mass

- (mass of element/molecular mass) x 100

Example Question (0620/42/F/M/23)

Calculate the percentage composition by mass of nitrogen in $(NH_4)_2SO_4$.

percentage of nitrogen = ... % [2]

Step 1: Calculate the molar mass of (NH4)2SO4
Step 2: Find the mass of Nitrogen
Step 3: Use the formula and multiply by 100.
Step 4: That will be the percentage of the question given.

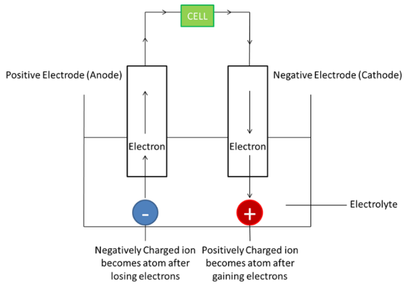
4. Electrochemistry

Electrolysis: the decomposition of an ionic compound, when molten or aqueous solution- by passing of an electric current

- This is possible due to the presence of **mobile electrons/free-moving ions**
- An **electrolyte** is a **molten or aqueous substance that undergoes electrolysis**

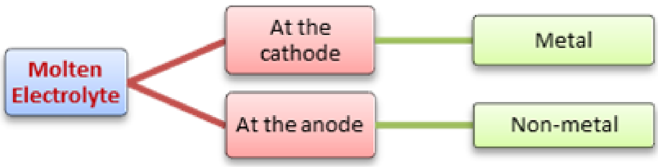
Components of Electrolysis	Definition
Electrodes	Metal or graphite rods that aid the flow of electricity in and out of the electrolyte 1. Anode: Positive electrode 2. Cathode: Negative Electrode (**PANIC**: Positive is Anode, Negative is Cathode)
Anion	Negatively charged ion that moves to anode
Cation	Positively charged ion that moves to the cathode

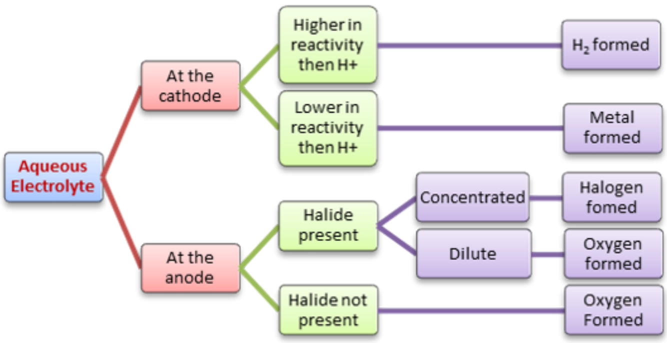
- Note: **Reactive** electrodes participate in the reaction, while **inert** electrodes (Graphite, Carbon) do not react with the cations or anions.



4.2. Reduction and Oxidation

- **Reduction** of positive cations happens at the cathode
- **Oxidation** of negative anions happens at the anode
  For example (Ionic Half Equations)
  - At the anode: $2Cl^- \rightarrow Cl_2 + 2e^-$
  - At the cathode: $2H^+ + 2e^- \rightarrow H_2$





Useful Acronyms

1. **REDCATANOX** (Reduction is Cathode, Anode is Oxidation)
2. **OILRIG** (Oxidation is loss, Reduction is gain)
3. **PANIC** (Positive is Anode, Negative is cathode)
4. **CMAN** (Cathode discharge Metals, Anode Discharge Non-Metals)

4.3. Observations in Electrolysis

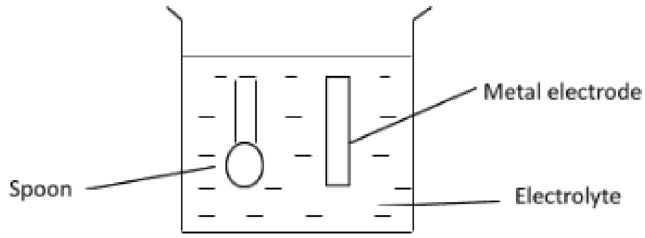
Electrolyte	At Cathode	At Anode
Molten Lead (II) Bromide	Lead	Bromine
Concentrated Hydrochloric Acid	Hydrogen	Chlorine
Concentrated Aqueous Sodium Chloride (Brine)	Hydrogen	Chlorine
Dilute Sulfuric Acid	Hydrogen	Oxygen
Aqueous Copper (II) Sulfate with Graphite Electrodes	Copper	Oxygen \n
Aqueous Copper (II) Sulfate with Copper Electrodes	Copper	Copper \n

- Blue copper (II) sulfate doesn't change as the concentration of Cu2+ ions remains unchanged.
- **Inert** (Unreactive electrodes) are **Platinum, Graphite or Carbon** Electrodes, So they don't **react with the ions during electrolysis.**

4.4. Electroplating

Electroplating: the process of coating the surface of a metal (more reactive) with another metal (less reactive) using electrolysis

- **Components:**
  - Anode: pure metal being used to electroplate the object
  - Cathode: object being electroplated
  - Electrolyte: aqueous solution of the soluble salt of pure metal (same as anode)
- **Used to:**
  - Prevent corrosion
  - Enhance appearance

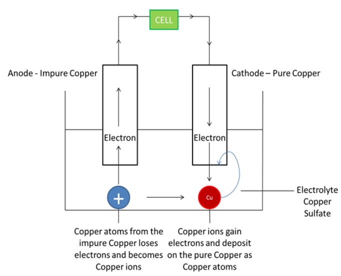


4.5. Refining Metals

- Cathode: a thin strip of pure metal
- Anode: impure metal
- Electrolyte: Aqueous Salt Solution of metal

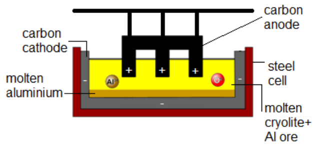
Example:

- **The refining of copper:** Impure copper as the anode and pure copper as the cathode; the aqueous copper (II) sulfate helps the copper ions move from the anode to the cathode. Here the ions gain electrons and become copper atoms, making the pure copper cathode thicker.
  1. Reaction at Anode: $Cu \rightarrow 2e + Cu^{2+}$ (mass decreases)
  2. Reaction at Cathode: $Cu^{2+} + 2e \rightarrow Cu$ (mass increases)



4.6. Extraction of Aluminum

- The main ore of Aluminium Oxide is bauxite
- Aluminum (III) oxide (alumina) is dissolved in molten **cryolite** (Na_3AlF_6) **– this mixture has a lower melting point (industrially preferred)**



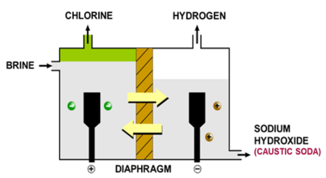
During electrolysis, aluminium ($Al^{3+} + 3e^- \rightarrow Al$) is produced at the carbon cathode, and oxygen ($2O^{2-} + 4e^- \rightarrow O_2$) at the carbon anode.

Molten Cryolite

- High temperatures are needed to melt is expensive. Therefore, molten cryolite **lowers the melting point/operating temperature and increases conductivity.**
- Due to the high temperature, the oxygen **reacts with the carbon** in the graphite anode to form CO_2, so the anode must **be periodically replaced.**

4.7. Electrolysis of Brine

- Brine is concentrated aqueous NaCl solution
- Ions present: Na^+, H^+, Cl^- and OH^-

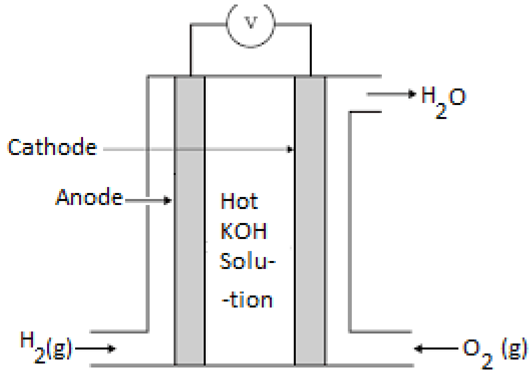


At the anode	At the cathode
Made of titanium	Made of steel
Cl^- ions; Chlorine gas	Hydrogen cations reduced to H_2 molecules

Unreacted ions (Na^+, H^+ and OH^-) move through porous membrane due to difference in liquid pressure
Left Na^+ and OH^- which form aqueous sodium hydroxide

4.8. Hydrogen-Oxygen Fuel Cells

Hydrogen-Oxygen Fuel Cells: Uses **hydrogen and oxygen as the main reactants** to produce electricity; the only product released is **water**.



Chemical Equation: $2H_2 + O_2 \rightarrow 2H_2O$

Advantages of motor vehicles	Disadvantages of motor vehicles
Renewable source	Large fuel tank required
Lesser flammability from petrol	Currently expensive
Emission Free (No carbon pollutions)	Lesser Hydrogen Filling stations
Non-toxic	

5. Chemical Energetics

5.1. Exothermic & Endothermic Reactions

Chemical reactions involve energy transfer between the system (the chemical reaction) and its surroundings.

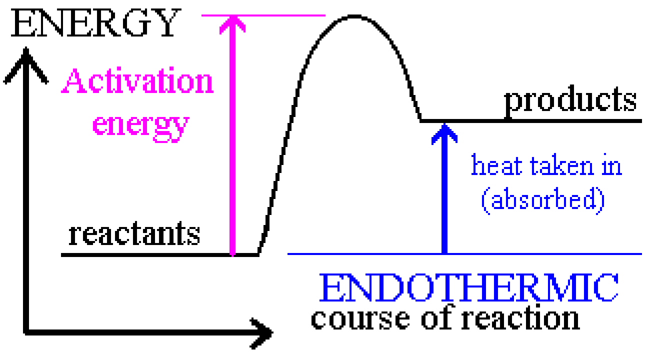
Exothermic Reaction	Endothermic Reaction
Heat energy is released into the surroundings	Heat energy is absorbed from the surroundings
Bond making reactions	Bond breaking reactions
Surrounding temperature increases	Surrounding temperature decreases

5.2. Energy Level Diagrams

- Energy level diagrams represent chemical reactions that include the relative energies of the reactant and product.
- A reaction's energy change (ΔH) is represented by the difference in height between the reactant and its product.
- The activation energy (Ea) is the minimum energy required for the reaction to take place.

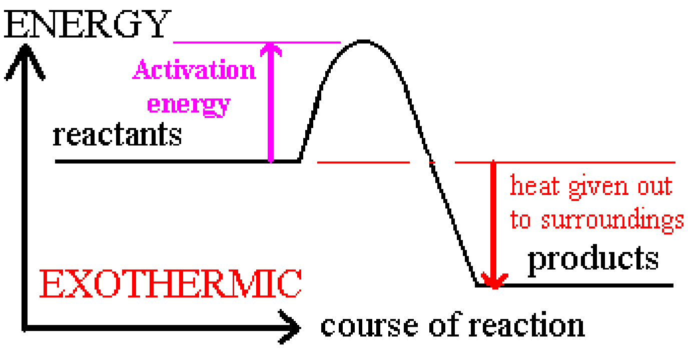
Endothermic Energy Level Diagram

- ~~The system gains energy; higher activation energy is required: Energy is taken in.~~



Exothermic Energy Level Diagram

- The system loses energy; lower activation energy is required: Energy is given out.



5.3. Bond Energy

- This is the energy required or released when a bond is formed or broken respectively. The unit measure of this energy is kJ/mol.
- The energy change in a reaction is calculated using the following formula:

$$\Delta H = Bond\ Breaking - Bond\ Forming$$

- If the overall heat energy value is negative, the reaction is exothermic
- If the overall heat energy value is positive, the reaction is endothermic

Example

$H_2 + Cl_2 \longrightarrow 2HCl$

Bond	Bond energy (kj/mol)
H - H	436
Cl - Cl	243
H - Cl	432

Bond breaking $\longrightarrow$ 436 + 243 = 679 kj/mol
Bond forming $\longrightarrow$ 2 (432) = 864 kj/mol
Thus,
Δ H $\longrightarrow$ 679 - 864 = -185 kj/mol
The reaction is exothermic.

6. Chemical Reactions

Physical Change	Chemical Change
The reaction is easily reversible	The reaction is harder to reverse
The product has no new chemical properties	Chemical product has different properties
Ex. dissolving a solute in a solvent	Energy change [exothermic/endothermic]

6.2. Rates of Reaction

Collision Theory
Successful collisions have enough activation energy to break pre-existing bonds and form new bonds at the moment of impact.
Rates of reaction

- The measure of the speed of the collision
- Calculated by the concentration of reactant used up or product produced per unit of time
- Unit = $(mol/dm^3)/s$

6.3. Concentration

Increasing the concentration of reactants increases the rate of reaction

- Higher-concentration reactants contain more particles per unit volume, increasing the successful collision and reaction rates.
- When the concentration changes in the rate of reaction graph, the ~~collision energy will remain the same~~, but the collision rate will increase (**activation energy does not change**).

Less particles, less frequent and successful collision

More particles, more frequent and successful collision

6.4. Temperature

Increasing temperature increases the rate of reaction

- Increased temperatures lead to increased average kinetic energy of particles. Particle movement produces energy greater than/equal to activation energy; an increased successful collision rate leads to an increased reaction rate.

Particles have less energy, less frequent and successful collision

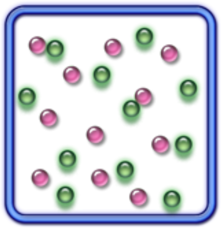
Particles have high energy, more frequent and successful collision

6.5. Surface Area of Solids

Decreasing the particle size increases the rate of reaction

- Decreasing particle size increases surface area; more reactant particles are exposed to collide, so the successful collision rate increases. This results in an increased rate of reaction

Less particles, less frequent and successful collision



More particles, more frequent and successful collision

Explosive Combustion

- Fine particles are combustible in the air due to a larger surface area
- The rate of reaction is high, making them explosive
- Examples: methane gas in coal mines and flour milling

6.6. Pressure

Increasing the pressure in a gaseous system increases the rate of reaction

- The distance between particles is reduced under pressure
- There are more particles per unit volume; the successful collision rate increases, resulting in an increased reaction rate.

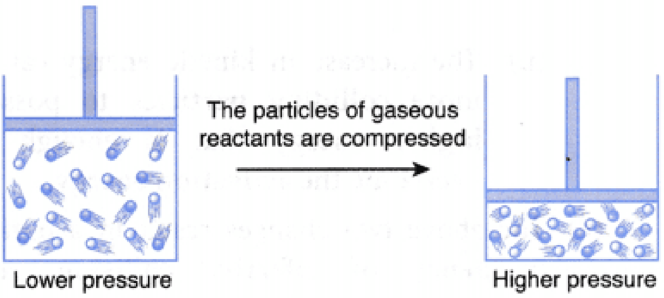
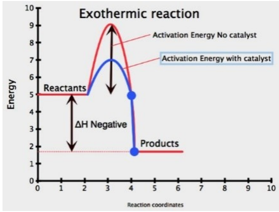


Figure *Increase in pressure of gaseous reactants*

6.7. Catalyst

A catalyst is a substance that increases the reaction rate **by lowering the activation energy and is left unchanged at the end of the reaction.**

- More particles will have an energy greater than or equal to the activation energy, therefore successful collision rate increases resulting in an increased rate of reaction
- For gaseous reactants, if the catalyst is solid metal, the catalyst provides a surface for the reaction.
- The larger the surface area of the metal catalyst, the larger the area for the reaction to take place; therefore higher the rate of reaction

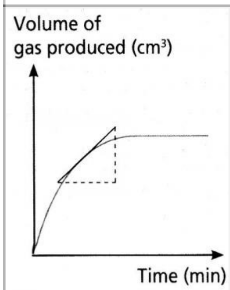
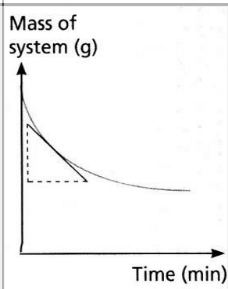
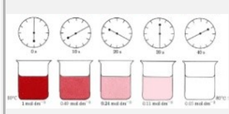


Enzymes Affecting Rate of Reaction

- Enzymes are biological catalysts that speed up reactions but remain chemically unchanged.
- Enzymes function best at optimum temperature and pH level; otherwise, they may denature and completely stop functioning.

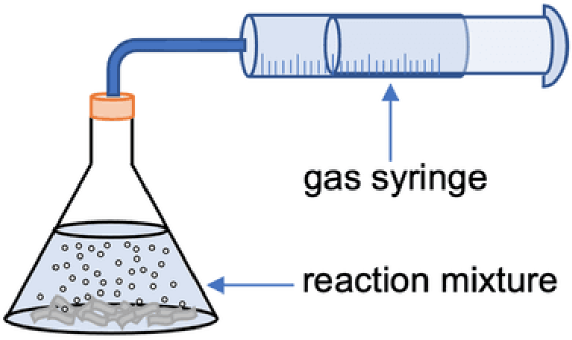
6.8. Rates of Reaction Graphs

Interpreting graphs:
A graph with a steeper gradient at the beginning and reaching a horizontal gradient faster depicts a high rate of reaction.

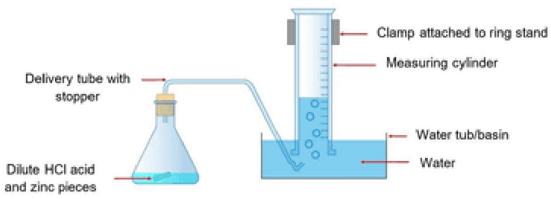
Gas Evolved	Mass Loss	Colour Change
		
If a gas evolves, measure the volume of gas produced per unit of time using a gas syringe	If a gas evolves, measure the loss in mass per unit time by placing it on a balance and then putting cotton wool on top to allow gas to pass but not enter	If a colour change, measure the time taken to turn cloudy

6.9. Evaluating Practical Methods

1. **Change in Mass method**
   - Not suitable for experiments where hydrogen gas is given off (because density decreases, too small mass)
2. Easier to set up the **gas syringe** and obtain reliable measurements



3. **An inverted Measuring cylinder** to collect gas over water is harder to set up.



6.10. Reversible Reactions

The symbol of reversible reactions are: ⇌

- The products can then react with each other or decompose to form the reactant molecules.
- Two equations you should know: \n $CuSO_4 \cdot 5H_2O\ (\text{blue}) \rightleftharpoons CuSO_4\ (\text{white}) + H_2O$ $CoCl_2 \cdot 6H_2O\ (\text{pink}) \rightleftharpoons CoCl_2\ (\text{blue}) + H_2O$
- (anhydrous by heating; hydrated form by adding water)

Reversible Reaction in a closed system when equilibrium:

- Rate of forward reaction = rate of reverse reaction
- Concentrations of all reactants and products remain constant and are no longer changing

Equilibrium

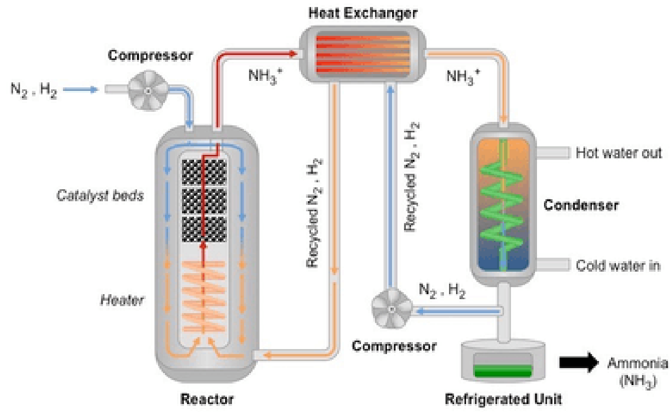
- **Le Châtelier's Principle:** if conditions of equilibrium are changed, the position of the equilibrium moves to oppose change
- **Temperature:** Temperature lowered; equilibrium moves in exothermic direction. Temperature increases; equilibrium moves in the endothermic direction.
- **Pressure:** Pressure raised; equilibrium moves to the side with the fewest gas molecules. Pressure lowered; equilibrium moves to the side with most gas molecules.
- **Concentration:** Decreasing reactant concentration or increasing product concentration; equilibrium moves to the reactant side. Increasing reactant concentration or decreasing product concentration; equilibrium moves to the product side.

  Oxidising Agents are **Electron Acceptors** and Reducing Agents are **Electron Donor.**

6.11. Haber Process

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

Materials (Reactants) in Haber Process
Nitrogen: From the **air**
Hydrogen: From the reaction between **Methane with Steam**



Conditions

Temperature: **450°C**
Pressure: **200 atm (20,000 kPa)**
Catalyst: **Iron catalyst**
Why are these conditions chosen in the Haber Process?

1. **Pressure is not too high** - (Although increased pressure is more yield, however, more pressure will require more expensive equipment, and it can explode).
2. **Optimum temperature** - (If lower temperature, more ammonia yield, however too low, the rate is uneconomical.

6.12. Contact Process

$2SO_2 + O_2 \rightleftharpoons 2SO_3$ (g) [Exothermic]

Materials (Reactants) in Contact Process
Sulfur Dioxide: **Burning sulfur or Roasting Sulfide ores**
Oxygen: From the air

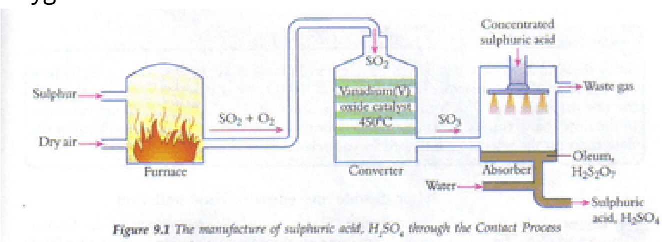


Figure 9.1 The manufacture of sulphuric acid, H_2SO_4 through the Contact Process

Conditions

Temperature: **450°C**
Pressure: **2 atm (200 kPa)**
Catalyst: **Vanadium (V) oxide catalyst**
Why are these conditions chosen in the Contact Process?

1. **Pressure is high** - (Increase pressure, more yield)
2. **Optimum temperature** - (If lower temperature, more sulfuric acid yield, however too low, the rate is uneconomical.

6.13. Redox

Redox: A simultaneous oxidation and reduction reaction

Oxidation	Reduction
Loss of electrons	Gain of electrons
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen

NOTE: **Roman Numerals** next to an element's name are the **oxidation number of an element in a compound**, e.g. Iron (II) and Iron (III). The reaction product formula depends on this.
Reducing agents are oxidised, and oxidising agents are reduced.

Identifying Redox Reactions

1. **Oxidation State:**

The oxidation state highlights electron movement in a reaction
eg. $CuO + Mg \rightarrow MgO + Cu$
$Cu^{2+} + Mg \rightarrow Mg^{2+} + Cu$ [oxide is a spectator and is removed as it doesn't change its oxidation state]
Here Copper (II) is reduced while magnesium is oxidized; the reaction is redox

2. **Indicators:**
   - **Acidified Aqueous Potassium manganate (VII)** is a deep purple oxidising agent; when added to a reducing agent changes from purple to colourless.
   - **Aqueous Potassium iodide** is a reducing; when added to an oxidising agent changes colourless to yellow-brown.

Redox Reactions by Changes in Oxidation Number

Rules	Example
The oxidation numbers of the element in their uncombined state are zero	In H_2, the oxidation number of H is "0."
The oxidation number of a monatomic ion is the same as the charge of the ion.	In Zn^{2+}, the oxidation number is "+2."
The sum of the oxidation numbers in a compound is zero	In H_2O, the oxidation number of H is +1, and O is -2, (2 x (+1) + (1+(-2)) = 0
The sum of the oxidation numbers in an ion is equal to the charge of the ion	In MnO^{4-} the oxidation number of Mn is +7, therefore (1 x (-7)) + (4 x (-2)) = -1

7. Acids, Bases and Salts

7.1. Properties of Acids

Acid: a substance that produces hydrogen ions (H^+) when dissolved in water.
Definition of Acids: **Proton donors**.
Indicators

- Have a pH between 1 (strong) and 6 (weak)
- Turns blue **litmus** red
- Turns **methyl orange** indicator red
- Colourless in **Thymolphthalein**

Weak and Strong Acids

1. **Strong acids: completely dissociated in aqueous solution** producing lots of H^+ ions

e.g: $HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$

2. **Weak acids: partially dissociated in aqueous solution** producing few H^+ ions

e.g: $CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COO^-\ (aq)$
Chemical properties

- Acid + metal → salt + hydrogen gas
- Acid + base → salt + water
- Acid + metal carbonate → salt + carbon dioxide + water

7.2. Properties of Bases

Bases: substances which neutralize acids to form salt and water only.
Definition of Bases: They are **proton acceptors** (form OH^- ions)
Indicators

- Have a pH between 8 (weak) and 14 (strong)
- Turns red **litmus** blue
- Turns **methyl orange** indicator yellow
- Turns Blue in **Thymolphthalein**

Weak and Strong Alkalis (Soluble Bases)

1. **Strong alkalis:** completely dissociates in aqueous solution producing lots of OH^- ions

e.g.: NaOH (aq) → Na^+ (aq) + OH^- (aq)

2. **Weak alkalis** partially ionize in water producing OH^- ions

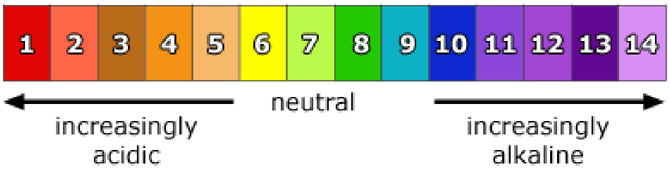
e.g: NH_4OH (aq) ⇌ NH_4^+ (aq) + OH^- (aq)

Chemical properties

- Base + acid → salt + water (+ CO_2 when base is a metal carbonate)
- Base + ammonium salt → salt + ammonia gas + water

7.3. Neutral

- Neutral substances are pH 7.
- **pH scale**
  pH is the concentration of H^+ ions per dm^3 of solution
  Universal indicator solution is used to determine the pH of a substance by matching the color change to the pH color chart.



7.4. Types of Oxides

- Metal oxides are **basic** in nature e.g. Copper oxide and Calcium oxide
- Non-metal oxides are **acidic** in nature e.g. sulfur dioxide and carbon dioxide
- Aluminium, and zinc form **amphoteric oxides** e.g. zinc oxide
- Oxides that react with neither acids nor bases are **neutral** e.g. nitrous monoxide and carbon monoxide

Colours of Transition Metal Compound

Metal Compounds	Colour
Copper (II) Sulfate	Blue
Copper (II) Oxide	Black
Copper (II) Carbonate	Green
Manganese (IV) Oxide	Black
Iron (II) Salts	Pale Green
Iron (III) Salts	Brown or Orange

Preparation of Soluble & Insoluble Salts

7.5. Soluble Salts

Method A: Soluble Salts from Excess Insoluble Bases (metal, metal oxide, carbonates)

1. Warm acid (increases the speed of reaction)
2. Add an excess reactant + stir
3. Filter mixture
4. Transfer to evaporating basin
5. Heat using a Bunsen burner
6. Leave to cool until crystallisation point
7. Wash crystals with distilled water
8. Dry crystals on filter paper

Method B: Titration

1. Place a known volume of alkali using a volumetric pipette into a conical flask.
2. Add indicator (e.g. thymolphthalein)
3. Titration: add acid using a burette until the endpoint has reached
4. Record the volume of acid added
5. Repeat without indicator
6. Transfer to evaporating basin
7. Heat with Bunsen burner
8. Leave to cool until crystallisation point
9. Wash crystals with distilled water
10. Dry crystals on filter paper

Insoluble Salt

Precipitation: Insoluble Solid forms between two aqueous solutions.

1. Mix two soluble salts
2. Filter to remove the precipitate
3. Wash the precipitate with distilled water
4. Leave to dry

7.6. Preparation of Salts

- A salt is a compound formed when a metal replaces all the hydrogen atoms of an acid.
- Naming salts involves two parts; the name of the metal and the acid ending
  eg. calcium + hydrochloric acid = calcium chloride
- The Water of Crystallisation is the **water molecules present in hydrated crystals.**

Type of Salt	Acid used
Sulfate	Sulfuric acid
Nitrate	Nitric acid
Chloride	Hydrochloric acid
Ethanoate	Ethanoic acid

- Salts can either be soluble or insoluble

Soluble Salts	Insoluble Salts
All sodium, potassium and Ammonium salts	None
All nitrates	None
Chlorides	Except for silver and lead
Sulfates	Except for barium, lead and calcium
Potassium, Sodium and Ammonium Carbonates	All other carbonates
Sodium, Potassium and Ammonium Hydroxides (partially calcium hydroxide)	Nearly all hydroxides

Identification of Ions and Gases

7.7. Test for Aqueous Cations

Cations	With aqueous NaOH	With aqueous Ammonia
Aluminum (Al^{3+})	**White** soluble precipitate turns colourless in excess	**White** precipitate, insoluble in excess
Ammonium (NH_4^+)	**Ammonia gas** produced on warming	
Calcium (Ca^{2+})	**White** precipitate, insoluble in excess	Faint or no precipitate
Copper (Cu^{2+})	**Light Blue precipitate,** insoluble in excess	**Light Blue** precipitate, soluble in excess to give a **dark blue** solution
Iron(II) (Fe^{2+})	**Green** precipitate, insoluble in excess	**Green** precipitate, insoluble in excess
Iron(III) (Fe^{3+})	**Red-brown** precipitate, insoluble in excess	**Red-brown** precipitate, insoluble in excess
Zinc (Zn^{2+})	**White** precipitate, soluble and turns colourless in excess	**White** precipitate, soluble and turns colourless in excess
Chromium (Cr^{3+})	**Green** precipitate, soluble in excess	**Green** precipitate, insoluble in excess

7.8. Test for Anions

Sulfate ions (**SO_4^{2-}**):
Add dilute nitric acid, then add aq. barium nitrate
White precipitate formed

Sulfite ions (**SO_3^{2-}**):
Add acidified aqueous potassium manganate (VII) and heat
Colour changes from purple to colourless

Halide ions:
Add nitric acid, then aqueous silver nitrate

Chloride (Cl^-)	**White precipitate**
Bromide (Br^-)	Cream precipitate
Iodide (I^-)	Yellow precipitate

Nitrate ions (**NO_3^-**):
Add aqueous sodium hydroxide, then add warm aluminum foil
Pungent gas produced turns damp red litmus blue

Carbonate ions (**CO_3^{2-}**):
Add dilute hydrochloric acid
If bubbles/ gas produced turn limewater cloudy, carbonate ions present

7.9. Test for Gases

Gas	Test and Test Result
Ammonia (NH_3)	Damp red litmus paper turns blue
Carbon Dioxide (CO_2)	Turns limewater milky
Chlorine (Cl_2)	Bleaches red/blue litmus paper
Hydrogen (H_2)	Place lighted splint, squeaky pop
Oxygen (O_2)	Place glowing splint, splint relights
Sulfur Dioxide (SO_2)	Turns Acidified Aqueous Potassium Manganate (VII) from purple to colourless

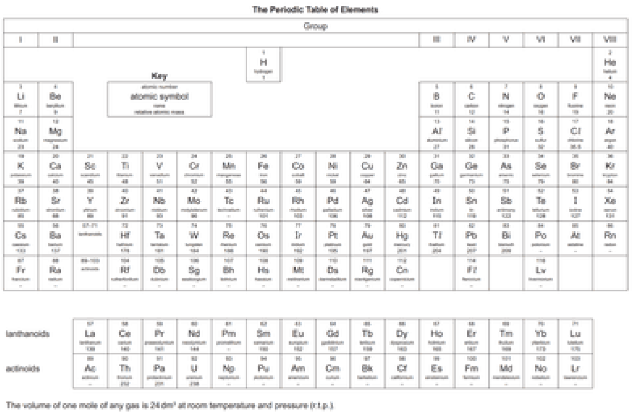
7.10. Flame Tests

Metal Ion	Flame Colour
Lithium	Red
Sodium	Yellow
Calcium	Orange-red
Potassium	Lilac
Barium	Light-Green
Copper (II)	Blue-Green

8. The Periodic Table

- The Periodic table is a method of classifying elements.
- Elements are arranged in order of increasing atomic number (each proceeding element has one more proton)
- Made up of rows called periods and columns called **groups**, the position of an element helps determine its electronic configuration.

- **Period number**: number of electron shells
- **Group number**: number of outer shell electrons
- Elements in the same group have similar chemical properties.



8.2. Periodic Trends

1. The table moves from metals on the left to non-metals on the right.
2. Down a group of metals, elements become more reactive.
3. With non-metals going down a group, reactivity decreases.

8.3. Group I Properties

- Group I metals: Lithium, sodium and potassium

Chemical Properties	Physical Properties
Readily react with oxygen and water; stored in oil	Good conductors of heat and electricity
React violently with chlorine	Soft and easy to cut
Burst into flames when heated with oxygen[red flame for lithium; yellow flame for sodium; lilac flame for potassium]	Shiny when freshly cut
Produce soluble white compounds.	Melting points decrease down the group.
React with water to form alkaline metal and hydrogen gas	Increase density down the group

- Predicting the properties of other Group I alkali metals: Rubidium, Caesium and Francium [reactivity increases down the group]

Element	Reaction with Water
Lithium	Floats and gives off hydrogen gas (effervescence)
Sodium	Vigorous Reaction and moves very quickly
Potassium	Explosive Reaction, lilac flame

8.4. Group VII Properties

Properties	Patterns
States and Colours, at RTP: Fluorine- Yellow-green gas Chlorine- Pale Yellow-Green gas Bromine- Red-Brown liquid Iodine- Grey-Black solid	Down the group, size, mass and density increase
Poisonous	Down the group, the colour darkens
Diatomic; form halide ions in a displacement reaction	Reactivity decreases down the group, because it has to gain an electron, so the closer the electron is to the positive nucleus, the more easily it will be gained, so atoms with fewer shells will react more easily.
Do not conduct electricity	melting point increases down the group

8.5. Transition Elements

Physical Properties

- High melting & boiling points
- Malleable and ductile
- Good conductors of heat & electricity
- High density

Chemical Properties

- Act as catalysts
- Form coloured compounds
- Variable Oxidation Numbers (Iron (II) or Iron (III))

8.6. Noble Gases

Properties	Uses
Density increases down the group	Helium-filling balloons
Monoatomic and colourless	Argon – Lamps
M.P. and B.P. increases the group	Neon – advertising signs
Don't conduct electricity	
Inert & stable due to full outer shell electrons	

9. Metals

9.1. Properties of Metals

Physical Properties

1. Good conductors of heat (when molten) and electricity
2. High melting and boiling points
3. Malleable and Ductile
4. High densities
5. Solids at room temperature (except mercury in liquid)

Chemical Properties

1. Metals + Acids → Salt + Hydrogen
2. Metal + Oxygen → Metal Oxide
3. Metal + Water → Metal Hydroxide + Hydrogen
4. Metal + Steam → Metal Oxide + Hydrogen

Non-Metals

Properties of Non-Metals

Physical Properties

1. Poor thermal conductor
2. Poor conductor of electricity (Except graphite)
3. Brittle and not malleable
4. Solids and gases at room temperature (bromine is liquid)

9.2. Uses of Metals

1. **~~Aluminium~~**
    - Manufacture of Aircraft/Cars (Low density)
    - Food Containers (Resistant to corrosion)
    - Overhead electrical cable (Good conductor of electricity/ductile)
2. **Zinc**
    - Galvanizes Iron = coats it to stop it from rusting (protective coating)
    - Alloys – brass/bronze
    - Batteries
    - Sacrificial Protection
3. **Copper**
    - Electrical Wiring (Good conductor of electricity/Ductile)

9.3. Alloys and their Properties

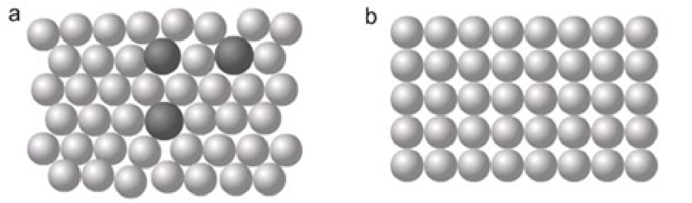
Alloy: a mixture of two or more metals or a metal and non-metal

- Alloys are useful because they are harder and stronger/useful than pure metal; they can also resist rusting with air and water.

Alloy	Made from	Special Properties	Uses
Brass	Copper and zinc	Stronger and more resistant to corrosion/rusting	Electrical fittings, car radiators

Alloy	Made from	Special Properties	Uses
Stainless steel	Iron, carbon, chromium and nickel	resistant to corrosion/rusting	Kitchen sinks, cutlery, surgical instruments

- Due to the irregularity in atom sizes and structure, metal alloys, which stop layers from sliding over each other, are stronger.
- This is what the structure of an alloy (a) looks like compared to a pure metal (b).



9.4. Reactivity Series

NOTE: Aluminum - despite its high placement in the reactivity series, it is seemingly unreactive due to its protective aluminium oxide layer.

General Reactivity Series (descending order of reactivity)	The Chemical Reactivity of Metals
K - Potassium Na - Sodium Ca – Calcium Mg – Magnesium Al – Aluminum **C – Carbon** Zn – Zinc Fe – Iron Pb – Lead **H – Hydrogen** Cu – Copper Ag – Silver Au – Gold	**Water:** 1. metal + cold water → metal hydroxide + hydrogen 2. metal + steam → metal oxide + hydrogen
Everything above hydrogen can displace hydrogen from its acid, and hydrogen cannot reduce its oxides.	**Oxygen:** **metal +** oxygen → metal oxide
Metals above carbon, their oxides cannot be reduced by carbon	**Dilute acids:** In a metal and acid reaction, the hydrogen atom in the acid is replaced by the metal atom to form a product of salt and hydrogen

Copper, Silver and Gold cannot react with Dilute Hydrochloric acid because its **too unreactive**

9.5. Displacement Reactions

In these reactions, metals compete for oxygen or anions

- Oxidisation is the loss of electrons
- Reduction is the gain of electrons
- The more reactive metal will displace the less reactive metal from oxygen or an anion.
- If the more reactive metal has oxygen or an anion, no reaction occurs
- The bigger the difference in reactivity between the two metals, the faster the reaction

Example: $Mg + O_2 \rightarrow 2MgO$
Here magnesium is oxidised while the oxygen atom is reduced

9.6. Corrosion of Metals

Rusting: the corrosion of iron and steel to form rust (**hydrated iron (III) oxide**) via oxidation

The conditions for rusting to occur:

1. Water
2. Air containing oxygen

Equation for Rusting of Iron

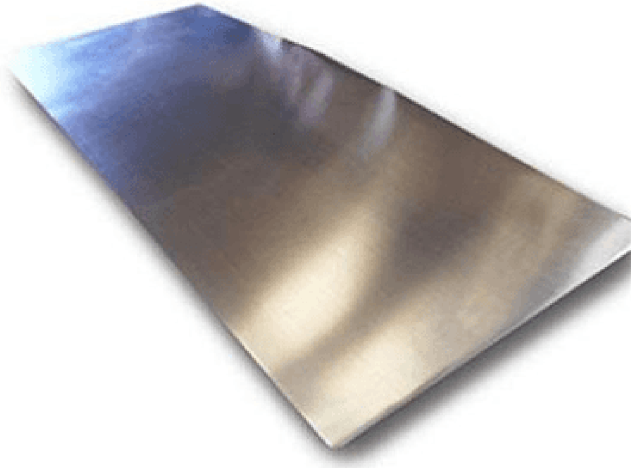
Iron + Water + Oxygen → Hydrated Iron (III) Oxide
$4Fe\ (s) + 2H_2O\ (l) + 3O_2\ (g) \rightarrow 2Fe_2O_3{\cdot}H_2O\ (s)$

Two Ways of Preventing Rust

1. **Barrier Method**
   - Coating a material to prevent the iron/steel from being in contact with water and oxygen (Painting, Greasing & Plastic Coating)
2. **Sacrificial Method**
   - When more reactive metals corrode to less reactive metals, they lose electrons in preference to iron.

Galvanisation of Iron

It is the **protection of iron and steel objects** by coating them with a **layer of zinc.**



The barrier method is due to the zinc layer preventing exposure to air and water and sacrificial protection due to zinc being more reactive than iron, corroding in preference to iron.

9.7. Extraction of Metals

- Process of separating a particular metal from its compound; metal ore
- Ore is more difficult to decompose from gold to potassium (expensive)

Metal	Extraction Method
K - Potassium Na - Sodium Ca – Calcium Mg – Magnesium Al – Aluminum	Reduction via electrolysis
Carbon	
Zn – Zinc Fe – Iron Pb – Lead	Reducing by Carbon
Hydrogen	
Cu – Copper Ag – Silver Au – Gold	Occur naturally

9.8. Extraction of Iron

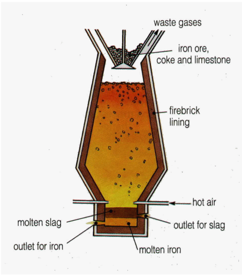
- Ore haematite (**Fe_2O_3**) + Impurities
- Burning of coke (Carbon) to **provide heat and produce carbon dioxide**

Ore is crushed and mixed with carbon and limestone (**$CaCO_3$**) and transferred into the blast furnace

$C + O_2 \rightarrow CO_2$ (exothermic)

$C + O_2 \rightarrow CO_2$ (exothermic)
$CaCO_3 \rightarrow CaO + CO_2$ (thermal decomposition)
$CO_2 + C \rightarrow 2CO$

2. Carbon monoxide reduces Iron(III) oxide from the ore to iron
   $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
3. The Calcium oxide reacts with impurities like silica to form slag (a waste product)
   $CaO + SiO_2 \rightarrow CaSiO_3$
4. Uses of slag include making roads and cement



The remaining waste gases that leave the blast furnace are **Carbon Dioxide, Carbon Monoxide, and Nitrogen.**

10. Chemistry of the Environment

10.1. Chemical Tests for Water

Test	Type of test	Positive result
Anhydrous Cobalt (II) Chloride	Chemical	Turns from blue to pink
Anhydrous Copper (II) Sulfate	Chemical	Turns from white to blue
Test Melting and Boiling Point	Physical	M.P at 0°C and B.P at 100°C

- **Distilled Water is used** in practical chemistry rather than tap Water because it has fewer chemical impurities.

10.2. Natural Sources of Water

Water is an important source in the natural world. However, with the changing world and massive urbanisation. **Our water may contain substances such as:**

1. Dissolved Oxygen (**this is important for aquatic life**)
2. Metal Compounds (Provide **essential minerals for life - however, some are toxic**)
3. Plastics (**harm aquatic life**)
4. Sewage (**contains harmful microbes which cause diseases**)
5. Harmful microbes
6. Nitrates fromfertiliserss
7. Phosphate from fertilisers and detergents **(leads to deoxygenation of water and damage to aquatic life (Eutrophication)**

10.3. Treatment of Domestic Water Supply

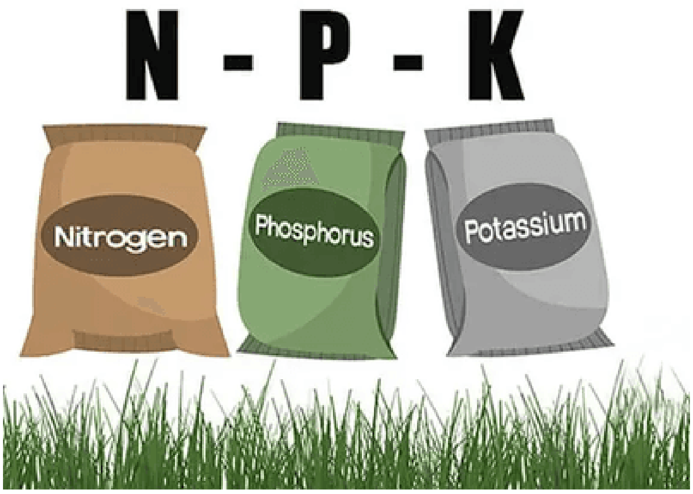
1. Water is pumped into screens to **remove solid, insoluble impurities.**
2. A **sedimentation** process making small clay pieces stick together and are then removed.
3. The water then undergoes **filtration** through layers of sand and gravel to **remove larger, insoluble debris**.
4. **Carbon** is also added into filtered water to **remove taste and odour**.
5. The **chlorination process** adds chlorine gas bubbled into the water to **kill bacteria and other microbes**; the acidic effect on the water is reversed by adding an alkali, sodium hydroxide.

Uses of Water

Home	Industry
Drinking, cooking and washing	Water jet cutting and water blasting
In car radiators, for gardens and plants	As a solvent in refining ores
	Generating hydroelectricity

10.4. Fertilisers

Fertilisers: Substances added to the soil and taken up by plants to increase crop yield.
Substances contain inside fertilisers are **Ammonium Salts and Nitrates**.



N.P.K - Nitrogen, Phosphorus, and Potassium Fertilisers are found inside fertilisers, essential to **improve plant growth.**

Functions of Elements

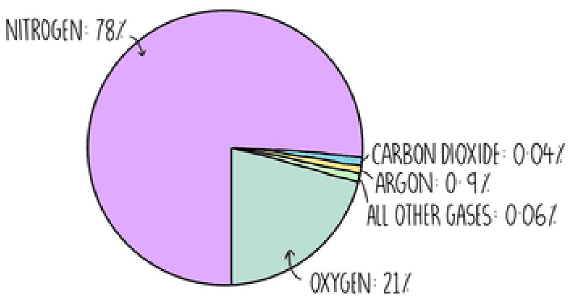
1. Nitrogen - Makes chlorophyll and protein. Promotes healthy leaves
2. Phosphorus - Promotes healthy roots
3. Potassium - Promotes growth and healthy fruits and flowers

Reaction with any alkali substance (except ammonia) displaces ammonia from its compound, for example:
Calcium hydroxide + Ammonium chloride → Calcium chloride + Ammonia + Water

10.5. Air Quality and Climate

The pie chart below presents the components present in clean air:

- Primary: **Nitrogen (78%), Oxygen (21%)**
- Secondary: **Noble gases (mainly Argon) and Carbon Dioxide (1%)**

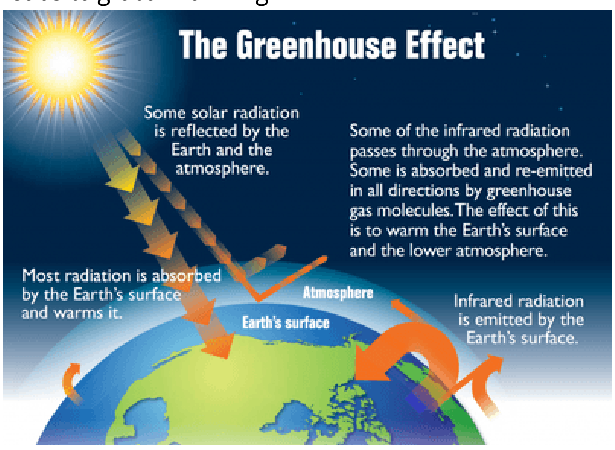


10.6. Pollutants in Air

Pollutant	Source	Negative impact
Carbon monoxide (CO)	Incomplete combustion of carbon-containing fuels (ex. Internal combustion engines)	Binds with haemoglobin, constricting oxygen supply in cells; leads to fatigue/ death
Carbon Dioxide (CO_2)	Complete Combustion of Carbon Containing Fuels	Increased global warming leads to climate change.
Methane (CH_4)	Decomposition of vegetation and waste gases from digestion in animals	Increased global warming leads to climate change.
Sulfur Dioxide (SO_2)	Combustion of fossil fuels which contain sulfur compounds	Causes acid rain.
Nitrogen Oxides (NO_2)	High temperatures that trigger a reaction between N_2 and O_2 (from air)	Causes respiratory problems and photochemical smog; contributes to acid rain
Lead Compounds	Combustion of leaded fuels	Damages brain and nerve cells in young children

Greenhouse Gases Impact on Global Warming

1. **Short wavelength radiation** from Sun reaches the Earth's surface
2. Some thermal energy is **absorbed and heats oceans/lands**
3. Earth radiates some thermal energy **as longer wavelength radiation**
4. Greenhouse gases absorb some of the infrared radiation and **re-emit in all directions**
5. Some infrared radiation comes back to Earth's surface, and this reduces the heat loss to space and leads to global warming



To reduce the effect of climate change:

- Planting trees
- Reduction in livestock farming
- Less use of fossil fuels
- Increased use of hydrogen and renewable energy (e.g. wind, solar)

To reduce the effect of acid rain occurring:

- Reduced emissions of sulfur dioxide by using low-sulfur fuels
- **Flue-gas desulfurisation** with calcium oxide (removing sulfur dioxide from plastic combustion)
- Use of catalytic converters in vehicles

10.7. Photosynthesis

Photosynthesis: the reaction between carbon dioxide and water to produce glucose and oxygen in the presence of chlorophyll and using energy from light.
Word Equation: **Carbon Dioxide + Water → Glucose + Oxygen**
Balanced Chemical Equation:
$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

10.8. Catalytic Converters

1. Present in car exhausts; contains transition metal catalysts of platinum and rhodium
2. Aids redox reactions to neutralize toxic pollutants formed as a result of incomplete fuel combustion: (a) Carbon Monoxide (b) Nitrogen Oxides (c) Unburned hydrocarbons
3. Reaction equations:
   (a) $2CO + O_2 \rightarrow 2CO_2$
   (b) $2NO + 2CO \rightarrow N_2 + 2CO_2$
   (c) $C_8H_{18} + 12\frac{1}{2}O_2 \rightarrow 8CO_2 + 9H_2O$

11. Organic Chemistry

11.1. Organic Formulae, Functional Group and Terminology

Suffix	Compound type
-ane	Alkane
-ene	Alkene
-anol	Alcohol
-anoic acid	Carboxylic acid
-yl/-anoate	Ester

Homologous Series: a group of organic compounds that have similar chemical properties due to being part of the:

- Same Functional Group
- Same General Formula
- Differing from One member to the next by a CH_2 unit
- Displaying a trend in physical properties
- Sharing similar chemical properties

Essential Terminology

1. Functional Group: an **atom or group of atoms** that **determine the chemical properties** of a homologous series
2. Structural Isomers: compounds with the **same molecular formula** but **different structural formula.**
3. Saturated Compounds: molecules in which all carbon-carbon bonds are **single bonds.**
4. Unsaturated Compounds: molecules in which **one or more carbon-carbon bonds** are not single.

11.2. Fuels

Common fossil fuels include: coal, natural gas [main constituent: methane] and petroleum
Petroleum: a mixture of hydrocarbons (**Carbon and Hydrogen ONLY**) which can be separated into useful fractions by fractional distillation
Going down the fractions

- Increasing chain length
- Increasing temperature
- Lower volatility
- Increase Boiling Points
- Increase Viscosity (harder to flow)

Petroleum Fraction	Uses
Refinery gas	heating and cooking
Gasoline fraction	fuels for cars
Naphtha fraction	chemical feedstock
Kerosene/Paraffin fraction	jet fuel
Diesel oil/Gas oil fraction	fuel in diesel engines
Fuel oil fraction	fuel in ships and home heating systems
Lubricating fraction	lubricants, waxes and polishes
Bitumen	making roads

Complete & Incomplete Combustion

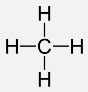
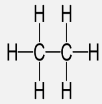
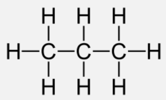
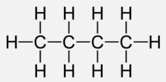
Complete Combustion
A complete combustion reaction with any organic compounds will produce **Carbon Dioxide and Water** as its product. The general equation is:
$C_XH_Y + O_2 \rightarrow CO_2 + H_2O$

Incomplete Combustion
An incomplete combustion reaction with any organic compounds will produce **Carbon Monoxide and Water** as its product. The general equation is:
$C_XH_Y + O_2 \rightarrow CO + H_2O$

11.3. Alkanes

Alkanes: saturated hydrocarbons [single carbon bonds] that are generally unreactive; however, they do undergo combustion reactions

General formula = C_nH_{2n+2}

Methane: CH_4 (n=1)	Ethane: C_2H_6 (n=2)
	
Propane: C_3H_8 (n=3)	**Butane: C_4H_{10} (n=4)**
	

Substitution Reaction

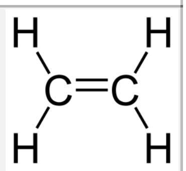
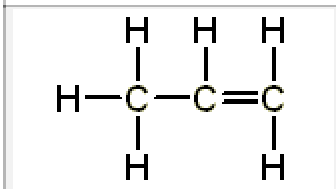
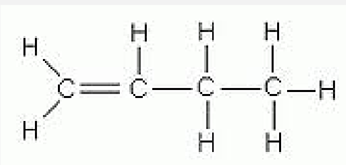
Alkanes go through substitutional reactions, where the atom of another element under UV light replaces the hydrogen atom. (Photochemical)

- The presence of UV light is to **provide activation energy**

$CH_4 + Cl_2 \rightarrow$ (light) $\rightarrow HCl + CH_3Cl$ / CH_2Cl_2 / $CHCl_3$ / CCl_4
Compounds = chloromethane/di/tri/tetrachloromethane

11.4. Alkenes

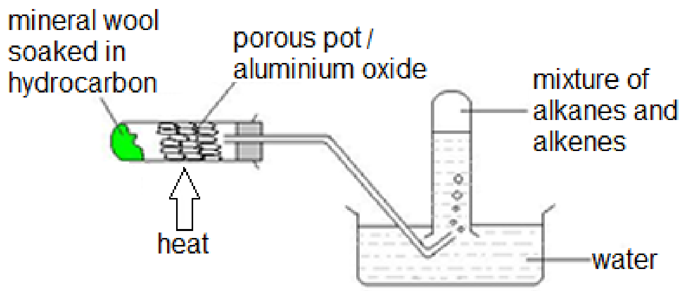
Alkenes: unsaturated hydrocarbons [at least one double bond between 2 carbon atoms]
Have isomers: same molecular formula but different structural formula (placement of double bond shifts), e.g. but-1-ene and but-2-ene
General formula = C_nH_{2n}
Functional group: C=C bond

(n=1)	Ethene: C_2H_4 (n=2)
N/A	
Propene: C_3H_6 (n=3)	**But-1-ene: C_4H_8 (n=4)**
	

Catalytic Cracking (Alkane)

- Thermal decomposition reaction, in which an alkene (and sometimes hydrogen) are produced from an alkane.
- Hydrocarbon is heated, and vapours are passed over a catalyst (alumina or silica)
- **Cracking** always produces a short-chain compound with a C=C bond **(matches the supply of fractions with demand and produces alkene for feedstock.)**

e.g. Cracking of ethane will give ethene and hydrogen



- Butane → Ethane + Ethene ; $C_4H_{10} \rightarrow C_2H_6 + C_2H_4$

How to distinguish between saturated and unsaturated hydrocarbons?
Using Bromine water (orange-brown):

- Saturated: remains orange (unreactive)
- Unsaturated: turns colourless/decolourises

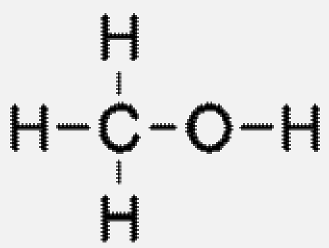
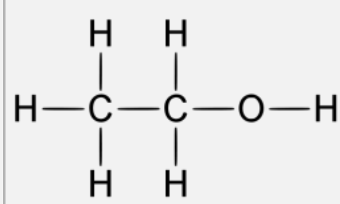
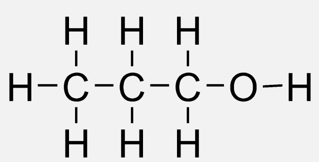
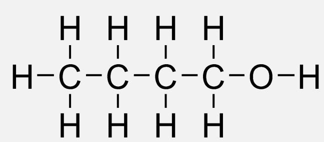
11.5. Alkenes' Addition Reactions

Alkene addition reactions **only form one product.**

1. **With Bromine:** (the test for saturation - orange brown-colourless)
   e.g. ethene (g) + bromine (aq) → 1,2-dibromoethane (l)
2. **With steam (hydration):** forms alcohols with heat (300°C), pressure (**60atm/6000kPa**) and an acid catalyst (phosphoric acid)
   e.g. ethene (g) + steam (g) ⇌ ethanol (l)
3. **With hydrogen (hydrogenation):** a double bond breaks down to form an alkane with a heat of 200 degrees and a catalyst (nickel)
   e.g. ethene (g) + hydrogen (g) → ethane (g)

11.6. Alcohols

General formula = $C_nH_{2n+1}OH$
Functional group: OH

Methanol: CH_3OH(n=1)	Ethanol: C_2H_5OH(n=2)
	
Propanol: C_3H_7OH (n=3)	**Butanol: C_4H_9OH (n=4)**
	

Methods of Production

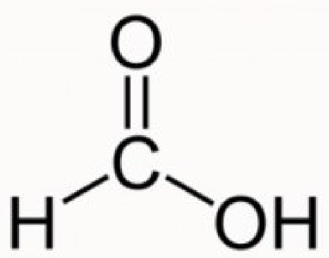
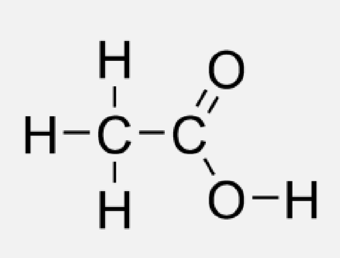
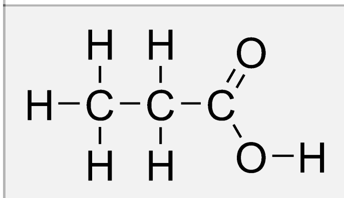
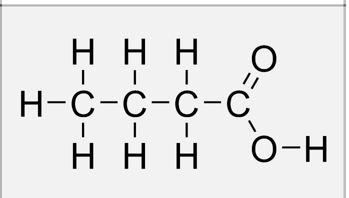
Fermentation of Aqueous Glucose (for Ethanol)	Catalytic Addition of Steam to Ethene
Yeast is added to dissolved glucose. \n Products: ethanol, carbon dioxide and the Temperature between 25-35 °C for optimal enzyme activity.	The ethene reacts with steam (reversibly) to form ethanol in the following conditions: 300°C, 60 atm (6000 kPa) \n Catalyst - phosphoric acid [while low temp. Gives better yield, high temp. is used for a faster rate of reaction]
The slow reaction produces a dilute solution that requires processing. Can only be produced in batches	The fast reaction produces pure ethanol. Continuous production (no batches)
Produces greenhouse gas (CO_2)	No greenhouse pollutants
Uses renewable resources	Uses non-renewable resources (crude oil)

Uses of Ethanol:

- Solvent in glues, printing inks & perfumes
- Fuel

11.7. Carboxylic Acids

General formula: $C_nH_{2n+1}COOH$
Functional group: COOH

Methanoic Acid: CH_2O_2 (n=1)	Ethanoic Acid: CH_3COOH (n=2)
	
Propanoic Acid: $C_3H_6O_2$ (n=3)	**Butanoic acid: $C_4H_8O_2$ (n=4)**
	

Ethanoic acid:

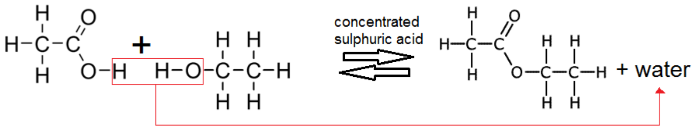
- Weak acid with high pH and low dissociation
- Formed by:
  - Bacterial Oxidation of Ethanol (Vinegar Production)
  - With acidified aqueous potassium manganate (VII) (Good Oxidising Agent)

Carboxylic acids react with alcohols (with an acid catalyst/Heat) to give esters, in a **condensation reaction**, for example:

- Ethanoic acid + ethanol ⇌ ethyl ethanoate + water (alcohol = -yl & carboxylic acid = -oate)
- Carboxylic Acids also have **different structural isomers.**

11.8. Polymers

- Large molecules are built up from small units known as monomers.
- Monomers are **a single unit of a substance.**



- Polymers can have different linkages depending on the type of polymerisation and monomer

Example:

Small units (monomers)	Linkages
Polyester (PET)	Ester
Polyamide & Protein	Amide

Plastics (PET)

Plastics (**polyethylene terephthalate)** are made from polymers. However, they are not biodegradable. (NOTE: PET can be hydrolysed back to monomers and re-polymerised)

Environment Challenges caused by Plastics

1. Disposal in Landfill sites
2. Accumulation in Oceans
3. Formations of toxic gases from burning

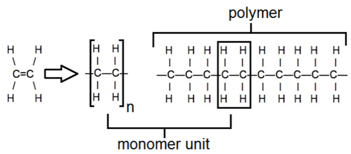
11.9. Properties of Polymers

- It can be moulded under heat and pressure due to its low density
- Low conductivity (retain heat)
- Resistant to corrosion
- Non-biodegradable and lead to plastic waste

11.10. Addition and Condensation Polymerisation

Addition Polymerisation

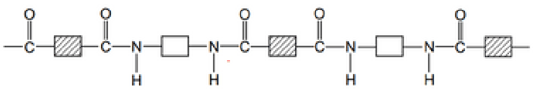
- This only occurs in monomers that contain double carbon (**C=C)** bonds
- Polymers produced using alkene monomers
- Forms only a polymer molecule
- Poly(ethene): is a polymer produced from ethene by addition polymerisation



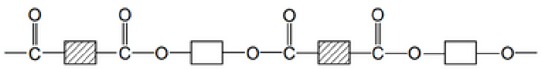
Double bond splits, and the polymer is formed. (Polymers have no double bonds)

Condensation Polymerisation

- When two different monomers are linked together with the removal of a smaller molecule, usually **water** (forms one H_20 molecules per linkage).
  1. **Nylon** (polyamide) is made from a **dicarboxylic acid monomer and a diamine monomer** (a compound with an NH_2 functional group). Forms **amide linkage.**

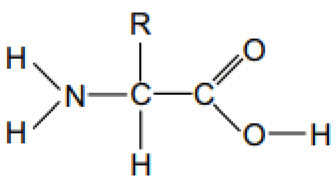
     

  2. **PET** (polyester) is made from a dicarboxylic acid monomer and diols (alcohol with an -OH functional group). Forms **ester linkage.**

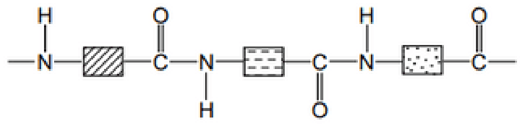
11.11. Protein & Natural Polyamides

Proteins act as the natural polyamides formed from amino acids monomers.
General Structure of Amino Acids:



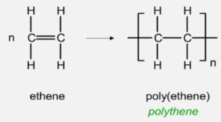
Proteins:

- Proteins contain the same linkages (amide links) as nylon but with different units from amino acid monomers. Their main structure you should know is:



12. Experimental Techniques and Chemical Analysis

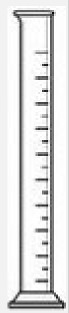
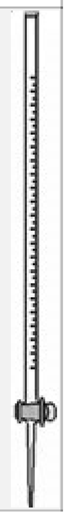
12.1. Measurement

	Use	Structure
Poly(ethene)	Plastic bags and gloves, clingfilm (low density), mugs, bowls, chairs, dustbins (high density)	

Variable	Unit(s)	Apparatus
Time	min/sec	Stopwatch
Temperature	ºC	Thermometer
Mass	grams	Balance

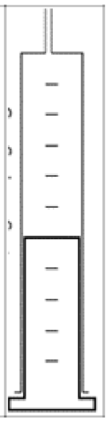
Measuring Volume in Liquids

Approximate measure to most accurate measure.

Name	Measuring Cylinder	Volumetric Pipettes (fixed volumes)	Burettes (variable volume)
Image			

Measuring Volume in Gases

A gas syringe is used to measure the volume of gases gradually

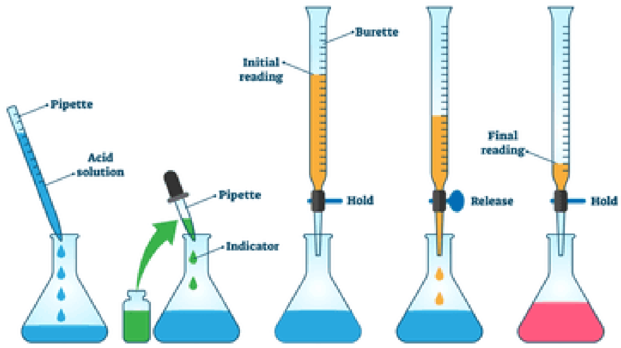


12.2. Mixture of Substances

- Mixture: A mixture is a substance made from two or more other substances not chemically bonded together
- Solution: A mixture of one or more solutes dissolved in a solvent
- Solute: A substance that dissolved in a solvent
- Solvent: A substance that dissolves a solute
- Saturated Solution: A solution containing the maximum concentration which cannot be dissolved further at a given temperature.

12.3. Acid-Base Titration

A method of quantitative chemical analysis where an acid is added slowly to a base until it has been neutralised.



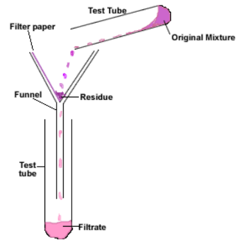
In Acid-Base Titrations, we must add a few drops of indicator (passes the endpoint) to check whether there are colour changes in the chemical reaction.

Indicators	Colour in Acid	Colour in Neutral	Colour in Base
Thymolphthalein	Colourless	Colourless	Blue
Methyl Orange	Red	Orange	Yellow

12.4. Filtration

Used to separate an insoluble solid from a liquid

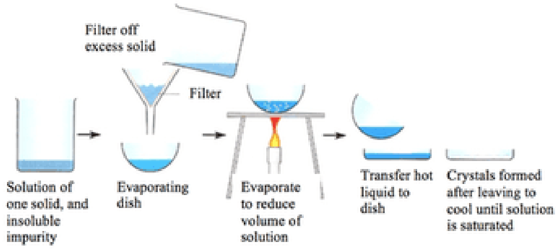
- The mixture goes through a funnel with filter paper into a flask.
- Insoluble residue remains in the funnel
- Filtrate flows through the funnel and gets collected in the flask or test tube or a beaker



12.5. Crystallisation

Used to separate dissolved solid from a solution

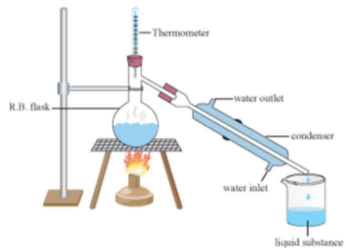
- The solution is heated to increase concentration (solvent evaporates)
- A drop of solution is placed on a slide to check for crystal formation
- The solution is left till it reaches the crystallisation point.
- Crystals are filtered from the solution, washed with distilled water, and dried between the filter paper.



12.6. Simple Distillation

Used to separate a solvent from a solution

- The impure liquid is heated in a round bottom flask
- When it boils, the steam rises into the attached condenser
- Condenser cools the steam to a pure liquid, and it drops into the beaker
- A thermometer is placed to ensure the **highest boiling point is not exceeded.**

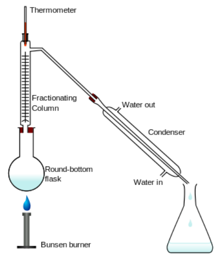


12.7. Fractional Distillation

Used to separate miscible liquids

- Mixture is heated
- Substances, due to their different boiling points, rise in different fractions.
- A mixture of gases condenses on the beads in the fractional column.

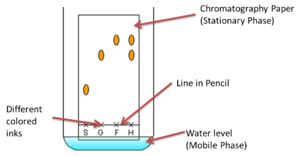
- The beads are heated to the boiling point of the lowest substance so that the substance being removed cannot condense on the beads.
- The other substances continue to condense and will drip back into the flask.
- The beaker can be changed after every fraction



12.8. Chromatography

Used to separate substances in a solvent with different solubilities.

1. Drop the substance onto the start line (pencil) drawn on chromatography paper.
2. Paper is placed in a beaker with solvent; the paper must touch the surface of the solvent (**water or ethanol**) while the line must be above the liquid.
3. The solvent travels up the paper.
4. Different solubilities lead to different travel rates (**high solubility -> high travel rate**)
   - The stationary phase is the material on which separation takes place
   - The mobile phase is the mixture you want to separate, dissolved in a solvent.



- Interpreting simple chromatograms:
  Chromatograms are the visual outputs on the chromatography paper
  - **Number of rings/dots = number of substances**
  - If two dots travel the same distance up the paper, they are the **same substance.**
  - The pure substance only gives **one spot.**
- Retention Value:
  Used to identify a substance, calculated by the formula:

$$\text{Rf Value} = \frac{\text{Distance moved by substance}}{\text{Distance moved by solvent}}$$

Locating Agents

- Used to make colourless chromatograms visible
  - Dry paper in the oven
  - Spray it with locating an agent
  - Heat it for 10 minutes in the oven

12.9. Separation Techniques

- Can be done by dissolving one in an appropriate solvent
- Then filter one and extract the other from the solution by evaporation
- If one solid is magnetic, it can use a magnet, e.g. sand and iron fillings

Component 1	Component 2	Technique	Example
Solid	Solid	Use a solvent	Sand and Salt
Insoluble Solid	Liquid	Use Filtration - Separating the solid and the residue	Copper (II) Oxide and Water
Soluble Solid	Liquid	Crystallization - Liquid is evaporated to leave the solid	Copper (II) sulfate from water
Liquid	Liquid (Miscible)	Distillation (2 Liquids) or Fractional Distillation (More than two liquids)	Ethanol & Water (Simple), Crude Oil (Fractional Distillation)

12.10. Purification

Purity in Substances

- Assessing purity

Pure substances	Have a definite, sharp m.p./b.p.
Impure substances	Have a lower m.p and a higher b.p

This assessment of substance purity is important, especially in food consumption, as its intake can be dangerous.

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