



A-Level Chemistry
Revision notes 2015

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Atomic Structure

The atom is made up of three sub-particles - the proton (+ve. and neutron (no charge. that are located in the nucleus of the atom and the electron (-ve charge. that is found orbiting the nucleus.

The nucleus has a positive charge.

The electrons are arranged in **energy levels** around the nucleus. Two electrons occupy the first, eight the second, eighteen the third.

Within the energy levels are **sub-shells**.

The region the electrons are said to occupy is called an **orbital**.

The four types of orbital are:

1. s-orbitals - spherical can hold 2 electrons.
2. p-orbitals - dumb-bell shaped, they go around in three's (px, py, pz. - so hold altogether a maximum of six electrons.
3. d-orbitals are complicated in shape - they are grouped in five's hence can hold a maximum of 10 electrons.
4. f-orbitals are complicated in shape - they are grouped in seven's - so altogether can hold 14 electrons.

The order that electrons fill is as follows:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p.

First ionisation energy:

The energy required to remove one mole of gaseous atoms to form one mole of gaseous ions.

The value of the first ionisation energy depends upon:

1. The effective nuclear charge
2. The distance between the electron and the nucleus
3. The 'shielding' produced by lower energy levels.

Three types of chemical bonding:

1. **Ionic bonding:** The transfer of electrons from metal atoms to a non-metal atom to form charged ions. The resulting product is held together by electrostatic attractions.
2. **Covalent bonding:** atoms share one or more electrons to form a molecule. A single covalent bond is shared with each atom donating one electron.
3. **Co-ordinate or dative covalent bonding:** a normal covalent bond, each atom donates one electron to the shared pair. In a co-ordinate bond electrons come from the same atom.

The **shape of a molecule** is decided by the **valence shell electron pair repulsion theory:** this states that molecules arrange their electron pairs to minimise repulsions between them.

Distorted shapes arise from the presence of lone pairs of electrons that cause greater repulsion than bonding pairs.

Atoms, Molecules and Stoichiometry

The **relative atomic mass**, A_r or RAM:

The average mass of an element's naturally occurring isotopes relative to the mass of an element of carbon-12.

A **mole** of a substance is the amount of substance that has the same number of particles as there are in exactly 12g of carbon-12.

The particles may be atoms, molecules, ions or even electrons. This number of particles is referred to as **Avogadro's constant**, L and is approximately $6 \times 10^{23} \text{ mol}^{-1}$.

The mass of one mole of a substance is often referred to as **molar mass**.

An instrument called a **mass spectrometer** is used to calculate relative atomic mass. The mass of individual isotopes and their abundance is found in order to calculate RAM.

The **empirical formula** of a compound shows the simplest whole-number ratio of the elements present.

The **molecular formula** shows the total number of atoms of each element present in the molecule.

States of Matter

The **kinetic theory of matter** states that all matter is made up of particles and exists in one of three states, solid, liquid or gas.

The **order** of the particles decreases as you change from solid, liquid to gas - due to decrease in forces between particles.

To change a substance from a solid to finally a gaseous state, energy must be supplied in order to overcome these forces of attractions between particles. As a change of state occurs the temperature of the substance remains constant as the energy supplied is used to overcome these attractive forces.

The **kinetic theory of ideal gases** makes two major assumptions.

1. Gases **do** have a volume.
2. Intermolecular forces of attraction **do** exist.

Real gases deviate from ideal behaviour at low temperatures and high pressure.

There are **three gas laws** that govern the behaviour of gases with regards to changes in temperature, pressure and volume.

1. Boyle's law: For a fixed mass of gas, the pressure is inversely proportional to the volume, if temperature remains constant.

$pV = \text{constant}$

2. Charles' law: For a fixed mass of gas, the volume is proportional to the absolute temperature, if the pressure remains constant.

$V/T = \text{constant}$

3. Pressure law: for a fixed mass of gas, the pressure is proportional to the absolute temperature, if the volume remains constant.

$p/T = \text{constant}$

The **gas constant** depends on the amount of gas, therefore is written as nR , where n = no. of moles. $R = 8.314 \text{ JK}^{-1}\text{mol}^{-1}$.

The **ideal gas equation** can now be written as:

$pV = nRT$

Units used must be SI.

Alternative uses of this equation are:

$$P_1V_1/T_1 = P_2V_2/T_2$$

where 1 represents the gas conditions before any change, 2 represents gas conditions after a change.

To calculate **molecular mass**: M_r ,

$$M_r = mRT/pV$$

There are three types of **intermolecular forces**:

1. van der Waal's:

Caused by non-polar molecules having temporary dipoles (due to movement of electrons) that cause an imbalance of electrons in neighbouring molecules. Hence, creating **electrostatic attractions**.

Example: methane CH_4

2. Permanent dipole:

A polar molecule contains permanent dipoles, due to the molecule being unsymmetrical in terms of shape or type of atom present. The size of this force is determined by the electronegativities of the atoms present.

Solids whose particles are held by permanent dipoles have greater boiling points than those held by van der Waal's due to their permanent nature.

Example: HCl

3. Hydrogen bond:

A strong electrostatic attraction between the poorly shielded proton of the hydrogen atom bonded to a small highly electronegative atom, such as N, O or F and a lone pair of electrons on a neighbouring molecule.

Example: water - H_2O .

Solids

Can be classified as one of five types:

1. Metallic:

Atoms held together by electrostatic forces between pseudo cations and delocalised electrons. Have high melting points and are good conductors of heat and electricity.

2. Giant ionic:

Ions held in a giant lattice due to electrostatic attraction between cations and anions. Soluble in water, good conductors when dissolved or in molten state, brittle.

3. Giant covalent:

In general each atom (C or Si) can be imagined situated in the centre of a tetrahedron strongly bonded to four other atoms. Covalent linking of these atoms occurs throughout the lattice.

Diamond has C atoms with the above arrangement, leading to its properties of poor conductor of electricity and heat, hard, very high melting point.

Graphite also has this giant linkage of covalent bonds between carbon atoms, however only three bonds are made by each atom, leaving a delocalised electron on each atom.

The carbon atoms are arranged in flat parallel layers. Between layers are weak van der Waal forces. Graphite is hard, a good conductor of heat and electricity with a very high melting point.

4. Simple molecular:

Weak van der Waal forces hold molecules in lattice (e.g. iodine) they have low melting points, are non-conductors of electricity and are insoluble in polar solvents such as water.

5. Hydrogen bonded:

High melting point in comparison to similar compounds due to presence of strong intermolecular forces. Ice has a less dense solid than liquid due to solid structure having much more free space between molecules.

Chemical Energetics

Chemical reactions can either release energy to the surroundings, **exothermic** or energy can be transferred from the surroundings, **endothermic**.

The **law of conservation of energy** states that energy cannot be destroyed or created.

Enthalpy change is the term used to describe the energy exchange that takes place with the surroundings at constant pressure. H is the symbol.

$$\Delta H = \Delta H \text{ products} - \Delta H \text{ reactants}$$

The units for enthalpy change are Kilojoules per mole.

An exothermic reaction has a **negative** enthalpy value whereas an endothermic enthalpy has a **positive** value.

Standard conditions are used to measure enthalpies of different reactions so that they may be fairly compared.

The **standard enthalpy change of reaction**: ΔH_r^\ominus is the enthalpy change when amounts of reactants shown in the equation for the reaction, react under standard conditions to give products in their standard state.

The **standard enthalpy of formation**: ΔH_f^\ominus is the enthalpy change when one mole of a compound is formed from its elements under standard conditions.

The **standard enthalpy of combustion**: ΔH_c^\ominus is the enthalpy change when one mole of an element or compound reacts completely with oxygen under standard conditions.

For a chemical reaction to occur bonds must break before new bonds can be made. When bonds break energy is absorbed (endothermic). When bonds form energy is released (exothermic).

Bond energy:

The amount of energy required to break a covalent bond; it indicates the strength of a bond. Value always quoted as bond energy per mole, E. The value is positive due to it being an **endothermic** reaction.

To measure enthalpy changes we use the following calculation:

$$\text{Energy transfer} = mc \Delta T \text{ (joules)}$$

c = specific heat capacity.

m = mass.

$$\Delta T = \text{temperature change.}$$

Hess' law states that the total enthalpy change for a chemical reaction is independent of the route by which the reaction takes place.

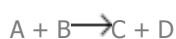
Reaction Kinetics

Chemical reactions occur at a definite rate which is determined by the reaction conditions.

Example of calculation:

Rate = change of volume/time

Consider the reaction:



The rate equation can be expressed as:

$$\text{Rate} = \frac{-d[A]}{dt} = \frac{-d[B]}{dt} = \frac{-d[C]}{dt} = \frac{d[D]}{dt}$$

To find out how [A] or [B] affect the rate we have to perform a series of experiments in which one concentration is varied whilst the other remains constant.

Usually it is found that:

$$\frac{-d[A]}{dt} \text{ is proportional to } [A]^x \text{ and } [B]^y$$

Or

$$\frac{-d[A]}{dt} = k [A]^x [B]^y$$

k = **Rate constant**

x and y are **orders** with respect to A and B.

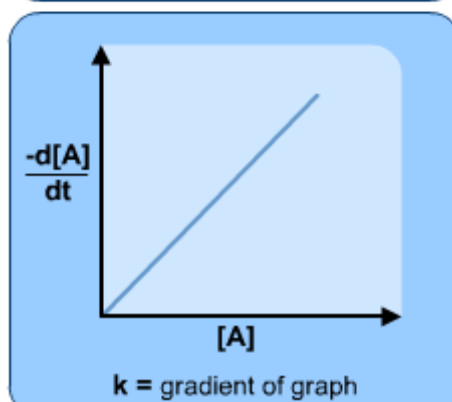
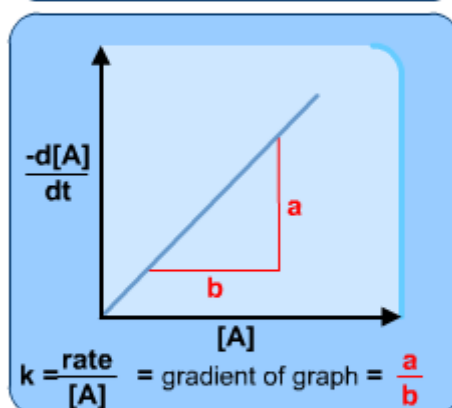
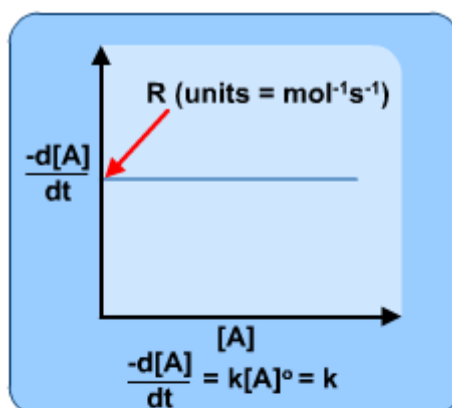
The **order** of a reaction with respect to a given reactant is defined as 'power of its concentration in the rate equation'.

The overall order is the sum of the powers of the concentrations of the reactants that appear in the rate equation.

To find the order of a reaction with respect to one of the reactants A,

1. Plot [A] against time
2. Calculate the rate at 5 or 6 different times by drawing tangents to the curve at these times and finding the gradients.
3. Plot the rate against [A]. If this is a straight line then the reaction is first order in A. If not a straight line then, plot rate against $[A]^2$. A straight line shows the reaction is second order in A.

Examples of these graphs are shown below:



The value of k, the rate constant is found by taking the gradient of the graph.

The **half-life** of a reactant is the time taken for the initial concentration to fall by half.

The effect of temperature on rate of reaction is summarised by two theories:

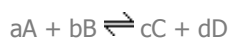
1. **The collision theory** states that molecules must collide with sufficient energy (activation energy) if a reaction is to take place. As temperature increases more molecules gain this activation energy, hence more collisions occur per second, rate increases.

2. **The transition state theory** explains the nature of an 'energy barrier' by the existence of an intermediate 'activated complex' or transition state formed during the reaction. This is a high energy species in which old bonds are partially broken and new bonds partially made.

Catalysts provide an alternative pathway that has a lower activation energy than the original one. The catalyst is involved in the reaction but is reformed at the end.

Chemical Equilibria

A reversible reaction is one where there is a forward and backwards reaction occurring:



The double arrow signifies a reversible reaction.

If in the above reactions the concentrations of the reactants and products does not change, although the reaction is still in progress, then the forward rate must equal the backward rate.

A situation known as **dynamic equilibrium is achieved**.

Any dynamic equilibrium can be described in terms of its equilibrium constant, K_c .

The **equilibrium constant** is the product of the molar concentrations of the products raised to the power of the coefficient in the stoichiometric equation, divided by the product of molar concentrations of the reactants, each raised to the power of its coefficient in the stoichiometric equation.

For the above reaction,

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

[] represents concentration of species.

Le Chatelier's principle states that the position of the equilibrium of a system changes to minimise the effect of any imposed change in conditions.

Changing **concentration**, does not change the numerical value of the equilibrium, but does change its position. In general equilibrium shifts to the right if concentration of reactant is increased.

Changing **pressure**, for reactions involving gases may cause a change in the position of the equilibrium, but the constant remains the same. For a reaction where there is an increase in the number of moles from reactants to products, increasing the pressure moves the equilibrium to the left.

Changing **temperature** does shift the equilibrium. For an exothermic reaction, the increase in temperature causes the equilibrium to shift to the left, since the opposing endothermic reaction will tend to reduce the temperature.

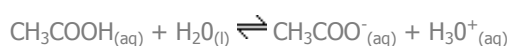
A catalyst has no effect on the position of the equilibrium, but does increase the rate of the forward and backward reactions.

Ionic Equilibria

The **Bronsted-Lowry theory** states that an acid is a substance which donates protons, and a base is a substance that accepts protons.

The equilibrium law can be applied to aqueous solutions of acids.

For example, the following equilibrium is established in an aqueous solution of ethanoic acid:



The equilibrium constant is given by:

$$K_a = \left(\frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \right)$$

Or

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

K_a is the **Acid dissociation constant** which is a measure of the strength of an acid.

$\text{p}K_a = -\lg K_a$ - for most acids this gives the range of values between 0-14. Strong acids have low $\text{p}K_a$ values.

K_w is the **ionic product of water**, it is based on the equilibrium that occurs due to water's ability to self-ionise.

$$K_w = [\text{H}^+][\text{OH}^-]$$

At 25°C it has a value of 1.0×10^{-14}

$$\text{p}K_w = 14.$$

To calculate pH:

$$\text{pH} = -\lg [\text{H}^+]$$

A buffer is a solution in which the pH of which does not change significantly when small amounts of acid or base is added to it.

Solubility product is a measure of a compound's solubility: K_{sp} .

Electrochemistry

Redox reactions involve **oxidation and reduction** reactions occurring simultaneously.

Oxidation is the loss of electrons.

Reduction is the gain of electrons.

Displacement reactions of metals and their ions in solution are an example of redox but also of **dynamic equilibria**. Metal atoms can lose electrons to become ions (oxidation) or metal ions can gain electrons to become atoms (reduction).

An **electrochemical cell** converts chemical energy into electrical energy. Since metals can be oxidised or reduced depending on their chemical environment.

Electrodes are charged (usually metal or graphite) rods.

Electrolytes are solutions that contain ions.

Salt bridge completes the circuit and prevents the build-up of charge in either half-cell, by allowing the passage of ions.

In order to find the **Standard electrode potential** of a half cell we use a **standard hydrogen electrode** (potential zero) and measure the EMF between the two cells.

$$\Delta E^{\ominus} = \Delta E^{\ominus} \text{ right-hand half-cell} - \Delta E^{\ominus} \text{ left-hand half-cell.}$$

The **Standard electrode potential** of a metal is the potential acquired when the metal is immersed in a 1 mol dm⁻³ solution of its ions at a temp. of 25°C - **symbol** - E

We can tabulate the order of oxidising/reducing ability of a system - this we call the **electrochemical series**. The most +ve E value is at the top - the greatest oxidising agent. The most -ve E value at the bottom - the greatest reducing agent.

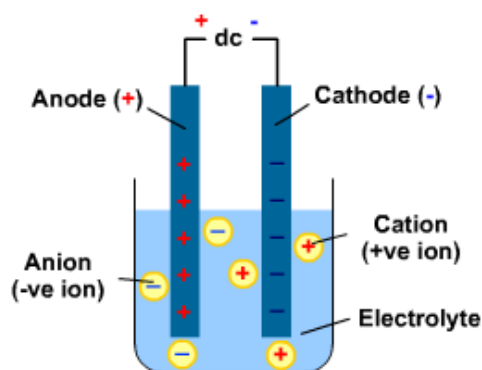
To make a prediction about a reaction, remember that the system which is lower in the series will lose electrons and the one higher in the series will gain electrons.

Different types of cells/batteries are used in everyday life, e.g. lead-acid battery and the dry cell - all carrying out redox reactions in order to convert chemical energy to electrical energy.

Rusting is an example of an electrochemical process.

Electrolysis is the decomposition of a compound using electricity.

An electrolysis cell is illustrated below:



The cell consists of two electrodes (cathode -ve and anode +ve) dipped into a compound in a molten state or in solution - the **electrolyte**.

When a d.c. current is passed the compound splits up due to the anions (-ve) being attracted to the anode and the cations (+ve being attracted to the cathode).

Reduction occurs at the cathode, **oxidation** at the anode.

Predicting what products occur at the electrodes is confused if compound is in aqueous solution due to the presence of OH⁻ (anion) and H⁺(cation) these compete with the other ions present. The electrochemical series is used to predict which cation is most readily reduced, and which anion is most readily oxidised - the other ions remain in solution.

In the **Chloro-alkali industry** chlorine gas, hydrogen gas and sodium hydroxide are produced from the electrolysis of **brine**.

The quantity of electricity passed is proportional to the amount of substance discharged at the electrode.

Quantity of electricity (charge) = current x time

One mole of electrons has a charge of 96500C. This is known as the **Faraday constant (F)**.

$$F = L \times e$$

L = Avagadro's no.

e = charge of an electron

The number of moles of electrons required to discharge 1 mole of ions is equal to the charge on the ion.

Group II and Group IV

Group II elements are less reactive than the Group I element in the same period. In compounds, the Group II elements always have an oxidation number of **+2**.

The ease with which an atom loses electrons increases with increasing atomic number in Group II.

The ease of **thermal decomposition** of Group II nitrates and carbonates decreases down the Group. Thermal decomposition is determined by the ability of the cation present to polarise the large carbonate or nitrate ions.

The **solubility** of Group II sulphates decreases down the Group. This trend is explained by the decrease in magnitude of the enthalpy change of hydration of M^{2+} .

The **metallic character** of the elements in Group IV increases with atomic number and the **non-metallic** character decreases with atomic number.

The stability of the **+4** oxidation state in Group IV **decreases** with increasing atomic number and the stability of the **+2** oxidation state **increases** with increasing atomic number.

All the **tetrachlorides** of Group IV except carbon are readily hydrolysed by cold water to form the corresponding oxide and HCl.

The **thermal stability** of the Group IV **tetrachlorides** decreases with increasing atomic number.

The **thermal stability** of the group IV **dioxides decreases** with increasing atomic number of the Group IV element.

Group VII

Reactivity decreases down the group with increasing atomic number.

Fluorine has the greatest oxidising power of all the halogens. Iodine has the greatest reducing power.

Within a compound, the halogen is most commonly in an **oxidation state of -1**.

Halogens usually make ionic bonds with metals. However they will form covalent bonds with themselves to form diatomic molecules or with a hydrogen atom.

The **hydrogen halides** are colourless acidic gases that dissolve in water to form acid solutions.

Thermal stability of the hydrogen halides decreases with increasing atomic number of the halogen.

In aqueous solution, the halogen with the lower atomic number can displace the halide ion with the higher atomic number.

Acidified silver nitrate can be used to distinguish between aqueous halide ions. The silver halides formed can be distinguished by the colour of the precipitate, and by their differing solubilities in aqueous ammonia.

Transition Metals

Definition: An element that has its 'd' orbitals partly filled, in some of its compounds.

They are hard, lustrous, weakly electropositive metals with high melting and boiling points.

The high melting points are due to the 3d and 4s electrons being available for delocalisation, hence the large **electrostatic attraction**.

Density increases across period as nuclear charge increases.

The 3d orbital has a higher energy than the 4s - hence electrons fill the 4s before the 3d.

On ionisation, the 4s empties before the 3d due to repulsion between 3d and 4s electrons. Therefore, the 4s are pushed to a higher energy level than 3d.

They have variable oxidation states due to the ability to lose electrons from 4s and 3d relatively easily.

The increase in stability of +2 oxidation state is explained by the greater difficulty of removing a third electron as the nuclear charge increases. However, Fe^{3+} is more stable than Fe^{2+} due to stability of $3d^5$ arrangement.

They form **coloured compounds**. Compounds that are coloured have electrons promoted from a ground state to an excited state. This is made possible due to the splitting of the d-orbitals into respective higher and lower energy levels. When the excited electron 'drops' back to a ground state a photon of light energy is released.

They form **complexes** - a complex is formed when a central metal atom or ion is surrounded by species that donate lone pairs of electrons.

A **ligand** is a species that donates a lone pair of electrons to a central metal atom or ion. The ligands make coordinate (or dative) bonds with the central atom/ion.

Ligand exchange may occur if the complex ion produced is more stable.

The addition of aqueous solutions of sodium hydroxide or ammonia may cause the production of coloured hydroxide precipitates. Some of these precipitates dissolve in excess alkali, while others dissolve in ammonia.

Periodicity

A group in the Periodic Table contains elements with the **same** outer shell electronic configuration but with **differing** proton numbers.

Periodic variations may be observed across Periods in physical properties such as ionisation energies, atomic radii, boiling points and electrical conductivities.

Across a Period the structures of the elements change from giant metallic, through giant molecular to simple molecular. Group 0 consists of individual atoms.

Chemically, the elements change from reactive metals, through less reactive metals and less reactive non-metals to reactive non-metals. Group 0 are highly unreactive.

There are periodic variations in the formulae and properties of oxides and chlorides of elements, from **ionic compounds in Group I** to **molecular covalent compounds in Group VII**.

Across **Period 3**, the trend in the reactions of the elements and their oxides and chlorides with water shows **decreasing pH** in the solutions formed.

Across Period 3 there is a general trend from the formation of a strongly alkaline solution by the ionic oxide in Group I, to the formation of increasingly acidic solutions by the molecular oxides of Group V to VII.

Between these two extremes lie the **basic oxide of Group II**, the **insoluble amphoteric oxide of Group III** and the **insoluble but acidic oxide of Group IV**.

All the chlorides of Period 3 elements dissolve readily in water. Across the Period there is a general trend of increasing speed of reaction and increasing acidity of the solution.

General Principles

Carbon atoms bond covalently by sharing electrons.

In general, organic compounds divide into **aromatic compounds**, which contain **benzene rings**, and **aliphatic compounds**, which do not. Hydrocarbons may be classified as **alkanes, alkenes or arenes**.

Each of these groups produces a **homologous series**, of compounds with **increasing numbers of atoms**.

There are many examples of different organic compounds with the same molecular formula in which the atoms are joined together in different arrangements. Such compounds are called **isomers**.

Chemists use a variety of types of formulae to represent organic molecules and help distinguish between isomers. Such formulae include structural and displayed formulae. **Structural isomers** have the same molecular formula but have different structural formulae.

Many hydrocarbons exist as structural isomers. **Geometric or cis-trans isomers** have the same molecular but have different displayed formulae. Alkenes may exist as geometric isomers.

Organic reagents can be classified as either:

nucleophiles- attack centres of low electron density.

electrophiles- attack centres of high electron density.

The properties of an organic molecule are predominantly determined by the properties of the **functional group** in that compound. Functional groups are atoms or combinations of atoms such as **-OH, -COOH**.

Aliphatic Compounds

Alkanes - C_nH_{2n+2}

Trends in melting point, boiling point and volatility can be explained by the weak intermolecular forces called induced dipole - induced dipole forces (**van der Waal's forces**).

As the chain length of alkanes increases, so does the number of electrons, which increases the strength of the van der Waal forces.

Branching reduces the surface area over which induced dipole - induced dipole forces can act.

Alkanes are **non-polar molecules**, so do not react with polar reagents.

Free radicals are produced by the homolytic fission of covalent bonds, for which light can provide energy.

Alkanes will react with free radicals - **free radical substitution**.

Alkenes

Alkenes are **more reactive** than alkanes because of the presence of the **electron rich double bond**.

Alkenes react by electrophilic addition:

With halogens to give dihalogenoalkanes.

Hydrogen to give alkanes.

Hydrogen halides to give halogenoalkanes.

Water to give alcohols.

Alkenes can be oxidised by cold dilute potassium manganate (VII) to give diols.

Alkenes can be oxidised by hot acidified concentrated potassium manganate (VII) to give carbonyl compounds, the reaction cleaving the carbon-carbon bond.

Halogenoalkanes

Halogenoalkanes can be prepared by the reaction of alcohols with concentrated hydrobromic acid, concentrated hydrochloric acid and concentrated hydroiodic acid or by the reaction of the appropriate phosphorus halides with an alcohol.

Halogenoalkanes react by **nucleophilic substitution**

With cyanide ion to form nitriles.

Ammonia to give amines.

Aqueous hydroxide ion to give alcohols.

Halogenoalkanes can eliminate hydrogen halides to form alkenes when treated with hot ethanolic alkali.

Alcohols

Alcohols contain the hydroxyl, OH, functional group attached directly to a carbon atom.

Short chain alcohols are soluble in water but long chain are not.

Primary alcohols can be oxidised to give carboxylic acids using acidified potassium dichromate (VI).

Secondary alcohols can be oxidised to ketones but tertiary alcohols cannot be oxidised.

Alcohols can be dehydrated to give alkenes.

Carboxylic acids and their derivatives

Can be prepared by the acid- or base catalysed hydrolysis of esters, amides or nitriles.

Carboxylic acids react with alcohols in the presence of an acid catalyst, such as concentrated sulphuric acid, to give esters.

Acyl chlorides are very useful in converting carboxylic acids to esters and amides.

Esters, acyl chlorides and amides are all derivatives of carboxylic acids

Aromatic and Plastics

Aromatic compounds

The **benzene ring** is a regular hexagon of carbon atoms joined by six pi bonds. Six pi electrons are delocalised in an orbital above and below the ring.

Arenes react by **electrophilic substitution** where an electrophile is substituted for a hydrogen atom, which is lost as a proton.

An arene is chlorinated or brominated by the reaction between chlorine and bromine and the arene in the presence of iron (III) chloride or aluminium chloride.

A nitrating mixture consists of concentrated sulphuric acid and concentrated nitric acid, and produces NO_2^+ that can react with arenes to form nitroarenes.

Phenol and phenylamine have highly activated benzene rings and form trisubstituted substitution products at positions 2, 4 and 6 with electrophiles.

Methylbenzene reacts faster than benzene towards electrophilic reagents and forms substitution products at position 2, 4 and 6.

Plastics

Alkenes are **monomers** that form **addition polymerisation**. Addition occurs due to the cleaving of the carbon double bond.

If ethene is the monomer the polymer made is polyethene.

In **condensation polymerisation** the addition of two monomers releases a small molecule, usually water.

In the formation of polyester the two polymers used are ethane 1,2-diol and benzene 1,4 dicarboxylic acid. When these add, an ester link is formed and water is released.

Polypeptides and proteins consist of long chains of amino acids.

Polypeptides are formed by a condensation polymerisation reaction between two amino acids.



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