

# IB SYLLABUS

## ***1.1 Mole Concept and Avogadro's Constant***

1.1.1. Describe the mole concept and apply it to substances.

The mole concept applies to all kinds of particles: atoms, molecules, ions, formula units etc. The amount of substance is measured in units of moles. The approximate value of Avogadro's constant ( $L$ ),  $6.02 \times 10^{23} \text{ mol}^{-1}$ , should be known.

1.1.2 Calculate the number of particles and the amount of substance (in moles).

Convert between the amount of substance (in moles) and the number of atoms, molecules or formula units.

## ***1.2 Formulas***

1.2.1 Define the term *molar mass* ( $M$ ) and calculate the mass of one mole of a species.

1.2.2 Distinguish between *atomic mass*, *molecular mass* and *formula mass*.

The term *molar mass* (in  $\text{g mol}^{-1}$ ) can be used for all of these.

1.2.3 Define the terms *relative molecular mass* ( $M_r$ ) and *relative atomic mass* ( $A_r$ ).

The terms have no units.

1.2.4 State the relationship between the amount of substance (in moles) and mass, and carry out calculations involving amount of substance, mass and molar mass.

1.2.5 Define the terms empirical formula and molecular formula.

The molecular formula is a multiple of the empirical formula.

1.2.6 Determine the empirical formula and/or the molecular formula of a given compound.

Determine the :

- empirical formula from the percentage composition or from other suitable experimental data
- percentage composition from the formula of a compound
- molecular formula when given both the empirical formula and the molar mass.

## ***1.3 Chemical Equations***

1.3.1 Balance chemical equations when all reactants and products are given.

Distinguish between coefficients and subscripts.

1.3.2 Identify the mole ratios of any two species in a balanced chemical equation.

Use balanced chemical equations to obtain information about the amounts of reactants and products.

1.3.3 Apply the state symbols (s), (l), (g) and (aq).

Encourage the use of state symbols in chemical equations.

## ***1.4 Mass and Gaseous Volume Relationships in Chemical Reactions***

1.4.1 Calculate stoichiometric quantities and use these to determine experimental and theoretical yields.

Mass is conserved in all chemical reactions. Given a chemical equation and the mass or amount (in moles) of one species, calculate the mass or amount of another species.

1.4.2 Determine the limiting reactant and the reactant in excess when quantities of reaction substances are given.

Given a chemical equation and the initial amounts of two or more reactants :

- identify the limiting reactant
- calculate the theoretical yield of a product
- calculate the amount(s) of the reactant(s) in excess remaining after the reaction is complete.

1.4.3 Apply Avogadro's law to calculate reacting volumes of gases.

## 1.5 Solutions

1.5.1 Define the terms *solute*, *solvent*, *solution* and *concentration* ( $\text{g dm}^{-3}$ ) and  $\text{mol dm}^{-3}$ )

Concentration in  $\text{mol dm}^{-3}$  is often represented by square brackets around the substance under consideration, eg  $[\text{CH}_3\text{COOH}]$ .

1.5.2 Carry out calculations involving concentration, amount of solute and volume of solution.

1.5.3 Solve solution stoichiometry problems.

Given the quantity of one species in a chemical reaction in solution (in grams, moles or in terms of concentration), determine the quantity of another species.

## 2.1 The Atom

2.1.1 State the relative mass and relative charge of protons, electrons and neutrons.

The accepted values are :

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

2.1.2 State the position of protons, neutrons and electrons in the atom.

2.1.3 Define the terms *mass number (A)*, *atomic number (Z)* and *isotope*.

2.1.4 State the symbol for an isotope given its mass number and atomic number.

Use notation,  ${}^A_Z X$  eg.  ${}^{12}_6 C$

2.1.5 Explain how the isotopes of an element differ.

Isotopes have the same chemical properties but different physical properties.

Examples such as  ${}^1_1 H$ ,  ${}^2_1 H$ ,  ${}^3_1 H$ ,  ${}^{12}_6 C$ ,  ${}^{14}_6 C$ ,  ${}^{35}_{17} Cl$ ,  ${}^{37}_{17} Cl$  should be considered.

2.1.6 Calculate and explain non-integer atomic masses from the relative abundance of isotopes.

2.1.7 Calculate the number of protons, electrons and neutrons in atoms and ions from the mass number, atomic number and charge.

## 2.2 Electron Arrangement

2.2.1 Describe and explain the difference between a continuous spectrum and a line spectrum.

2.2.2 Explain how the lines in the emission spectrum of hydrogen are related to the energy levels of electrons.

Students should be able to draw an energy-level diagram, show transitions between different energy levels and recognize that the lines in a line spectrum are directly related to these differences. An understanding of convergence is

expected. Series should be considered in the ultraviolet, visible and infrared regions of the spectrum. Calculations, knowledge of quantum numbers and historical references are not required.

2.2.3 Describe the electron arrangement of atoms in terms of main energy levels.

Students should know the maximum number of electrons that can occupy a main energy level (up to  $Z = 18$ ). No knowledge of sublevels s, p, d and f is required. The term *valence electrons* is used to describe the electrons in the highest main energy level.

2.2.4 Determine the electron arrangement up to  $Z = 20$ .

For example, 2.8.7 or 2.8.7 for  $Z = 17$ .

### 3.1 The Periodic Table

3.1.1 Describe the arrangement of elements in the periodic table in order of increasing atomic number.

Names and symbols of the elements are given in the *Chemistry Data Booklet*.

The history of the periodic table is not required.

3.1.2 Distinguish between the terms *group* and *period*.

The numbering system for groups in the periodic table is shown in the data booklet. Students should also be aware of the position of the transition metals in the periodic table.

3.1.3 Deduce the relationship between the electron configuration of elements and their position in the periodic table.

Explanations are only required for the first 20 elements, although general principles can extend to the whole of the periodic table. For example, students should know or be able to predict that K is in group 1 using  $Z = 19$ , but need only know that since Cs is in group 1, it has one electron in its outer shell.

### 3.2 Physical Properties

3.2.1 Describe and explain the periodic trends in atomic radii, ionic radii, ionization energies, electronegativity and melting points for the alkali metals ( $\text{Li} \rightarrow \text{Cs}$ ), halogens ( $\text{F} \rightarrow \text{I}$ ) and period 3 elements ( $\text{Na} \rightarrow \text{Ar}$ ).

Cross reference with topics 2, 4 and 5. Data for all these properties are listed in the data booklet. Explanations for the first four trends should be given in terms of the balance between the attraction of the nucleus for the electrons and the repulsion between electrons. Explanations based on effective nuclear charge are not required.

Ionization energy is defined as the minimum energy required to remove one electron from an isolated gaseous atom.

### 3.3 Chemical Properties

3.3.1 Discuss the similarities in chemical nature of elements in the same group.

The following reactions should be covered :

- alkali metals (Li, Na and K) with water and with halogens ( $\text{Cl}_2$  and  $\text{Br}_2$ ).
- halogens ( $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$ ) with halide ions ( $\text{Cl}^-$ ,  $\text{Br}^-$  and  $\text{I}^-$ )
- halide ions ( $\text{Cl}^-$ ,  $\text{Br}^-$  and  $\text{I}^-$ ) with silver ions.

Reactions of the halogens with alkali and confirmation of the silver halide by reaction with ammonia solution are not required.

3.3.2 Discuss the change in nature, from metallic to non-metallic, of the elements across period 3.

Use the study of the period 3 oxides to illustrate, for example, the change from basic through amphoteric to acidic oxides and their reaction with water. Halides and hydrides are not required.

## 4.1 Ionic Bond

- 4.1.1 Describe the ionic bond as the result of electron transfer leading to attraction between oppositely charged ions.
- 4.1.2 Determine which ions will be formed when metals in groups 1, 3 and 3 lose electrons.
- 4.1.3 Determine which ions will be formed when elements in groups 6 and 7 gain electrons.
- 4.1.4 State that transition metals can form more than one ion.  
Restrict examples to simple ions eg  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ .
- 4.1.5 Predict whether a compound of two elements would be mainly ionic or mainly covalent from the position of the elements in the periodic table, or from their electronegativity values.
- 4.1.6 Deduce the formula and state the name of an ionic compound formed from a group 1, 2 or 3 metal and a group 5, 6 or 7 non-metal.

## 4.2 Covalent Bond

- 4.2.1 Describe the covalent bond as the result of electron sharing.  
The electron pair is attracted by both nuclei leading to a bond which is directional in nature. Both single and multiple bonds should be considered.  
Dative covalent bonds are not required.
- 4.2.2 Draw the electron distribution of single and multiple bonds in molecules.  
Examples should include  $\text{O}_2$ ,  $\text{N}_2$ ,  $\text{CO}_2$ ,  $\text{C}_2\text{H}_4$  (ethene) and  $\text{C}_2\text{H}_2$  (ethyne).
- 4.2.3 State and explain the relationship between the number of bonds, bond length and bond strength.  
The comparison should include bond lengths and bond strengths of :  
- two carbon atoms joined by single, double and triple bonds  
- the carbon atom and the two oxygen atoms in the carboxyl group of a carboxylic acid.
- 4.2.4 Compare the relative electronegativity values of two or more elements based on their positions in the periodic table.  
Precise values of electronegativity are not required.
- 4.2.5 Identify the relative polarity of bonds based on electronegativity values.  
In a covalent bond, electron distribution may not be symmetrical and the electron pair may not be equally shared.
- 4.2.6 Draw and deduce Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom.  
A pair of electrons can be represented by dots, crosses, a combination of dots and crosses or by a line. For example, chlorine can be shown as :  

$$\begin{array}{c} \text{x x} \quad \text{x x} \\ \text{x Cl} \quad \text{x Cl} \quad \text{x} \quad \text{or} \quad \begin{array}{c} \cdot \cdot \quad \cdot \cdot \\ \cdot \cdot \quad \cdot \cdot \\ \cdot \cdot \quad \cdot \cdot \\ \cdot \cdot \quad \cdot \cdot \end{array} \quad \text{or} \quad \begin{array}{c} \text{Cl} \\ \text{---} \\ \text{Cl} \end{array} \quad \text{---} \quad \begin{array}{c} \text{Cl} \\ \text{---} \\ \text{Cl} \end{array} \end{array}$$
- Note : Cl - Cl is not a Lewis structure.
- 4.2.7 Predict the shape and bond angles for molecules with four charge centres on the central atom.  
Use the valence shell electron pair repulsion (VSEPR) theory to predict the shapes and bond angles of molecules and ions having four pairs of electrons (charge centres) around the central atom. Suitable examples are  $\text{NH}_3$ ,  $\text{H}_2\text{O}$  and alkanes (eg  $\text{CH}_4$ ).

4.2.8 Identify the shape and bond angles for species with two and three negative charge centres.

Examples should include species with non-bonding as well as bonding electron pairs, eg  $\text{CO}_2$ ,  $\text{SO}_2$ ,  $\text{C}_2\text{H}_2$ ,  $\text{C}_2\text{H}_4$ ,  $\text{CO}_3^{2-}$  and  $\text{NO}_2^-$ .

4.2.9 Predict molecular polarity based on bond polarity and molecular shape.

The polarity of a molecule depends on its shape and on the electronegatives of its atoms, eg  $\text{CO}_2$ ,  $\text{H}_2\text{O}$ .

### **4.3 Intermolecular Forces**

4.3.1 Describe the types of intermolecular force (hydrogen bond, dipole-dipole attraction and van der Waals' forces) and explain how they arise from the structural features of molecules.

All these intermolecular forces are weaker than covalent bonds. For substances of similar molar mass, hydrogen bonds are stronger than dipole-dipole attractions which are stronger than van der Waals' forces. Van der Waals' forces arise from the electrostatic attraction between temporary induced dipoles in both polar and non-polar molecules.

4.3.2 Describe and explain how intermolecular forces affect the boiling points of substances.

The hydrogen bond can be illustrated by comparing physical properties of  $\text{H}_2\text{O}$  and  $\text{H}_2\text{S}$   $\text{NH}_3$  and  $\text{PH}_3$   $\text{C}_3\text{H}_8$ ,  $\text{CH}_3\text{CHO}$  and  $\text{C}_2\text{H}_5\text{OH}$

### **4.4 Metallic Bond**

4.4.1 Describe metallic bond formation and explain the physical properties of metals.

Metallic bonding is explained in terms of a lattice of positive ions surrounded by delocalized valence electrons. The delocalised electrons should be related to the high electrical conductivity, malleability and ductility of metals.

### **4.5 Physical Properties**

4.5.1 Compare and explain the following properties of substances resulting from different types of bonding: melting and boiling points, volatility, conductivity and solubility.

Consider melting points, boiling points and volatility of similar substances, such as  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$ , and substances with different types of bonding and different intermolecular forces. Students should be aware of the effect of impurities on the melting point of a substance.

The solubilities of compounds in non-polar and polar solvents should be compared and explained. Consider also the solubilities of alcohols in water as the length of the carbon chain increases.

4.5.2 Predict the relative values of melting and boiling points, volatility, conductivity and solubility based on the different types of bonding in substances.

## **5.1 States of Matter**

5.1.1 Describe and compare solids, liquids and gases as the three states of matter.

The movement of particles, the attractive forces between particles and the interparticle spacing should be described. A molecular level description of what happens when evaporation, boiling, condensing, melting and freezing occur should be given. Students should understand what is meant by the term *diffusion*.

5.1.2 Describe kinetic energy in terms of the movement of particles whose average energy is proportional to absolute temperature. Students should be able to describe what happens when the temperature is changed.

Kinetic theory should be interpreted in terms of ideal gases consisting of point masses in random motion whose energy is proportional to absolute temperature. Students should be able to describe what happens when the temperature is changed.

5.1.3 Describe the Maxwell-Boltzmann energy distribution curve.

5.1.4 Draw and explain qualitatively Maxwell-Boltzmann energy distribution curves for different temperatures.\

5.1.5 Describe qualitatively the effects of temperature, pressure, and volume changes on a fixed mass of an ideal gas.

5.1.6 State the ideal gas equation,  $PV=nRT$ .

5.1.7 Apply the ideal gas equation in calculations.

Use the relationship between  $P$ ,  $V$ ,  $n$  and  $T$  for gases. Students should be familiar with  $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$  and be able to calculate molar volume.

## **6.1 Exothermic and Endothermic Reactions**

6.1.1 Define the terms *exothermic reaction*, *endothermic reaction* and *standard enthalpy change of reaction* ( $\Delta H^\theta$ )

Standard enthalpy change is heat transferred under standard conditions - pressure 101.3 kPa, temperature 298 K. Only  $\Delta H$  can be measured, not  $H$  for the initial or final state of a system.

6.1.2 State the relationship between temperature change, enthalpy change and whether a reaction is exothermic or endothermic.

Combustion of organic compounds are good examples of exothermic reactions.

6.1.3 Deduce, from an enthalpy level diagram, the relative stabilities of reactants and products and the sign of the enthalpy change for the reaction.

If the final state is more stable (lower on the enthalpy level diagram), this implies that  $H_{final} < H_{initial}$  and  $\Delta H$  must be negative. Energy must be released in going to a more stable state.

6.1.4 Describe and explain the changes which take place at the molecular level in chemical reactions.

Relate bond formation to the release of energy and bond breaking to the absorption of energy.

6.1.5 Suggest suitable experimental procedures for measuring enthalpy changes of reactions in aqueous solution.

Explore different reactions operating at constant pressure (open containers). Use of the bomb calorimeter is not required.

## 6.2 Calculation of Enthalpy Changes

6.2.1 Calculate the heat change when the temperature of a pure substance is altered.

Student should be able to calculate the heat change for a substance given the mass, specific heat and temperature change.

6.2.2 Explain that enthalpy changes of reaction relate to specific quantities of either reactants or products.

Enthalpy changes are measured in joules (J) and are often quoted in  $\text{kJ mol}^{-1}$  or either a reactant or a product.

6.2.3 Analyse experimental data for enthalpy changes of reactions in aqueous solution.

6.2.4 Calculate the enthalpy change for a reaction in aqueous solution using experimental data on temperature changes, quantities of reactants and mass of solution.

Enthalpy change of an acid-base reaction could be investigated.

## 6.3 Hess's Law

6.3.1 Determine the enthalpy change of a reaction which is the sum of two or more reactions with known enthalpy changes.

Use examples of simple two- and three-step processes. Students should be able to construct simple enthalpy cycles, but will not be required to state Hess's law.

## 6.4 Bond Enthalpies

6.4.1 Define the term *average bond enthalpy*.

Bond enthalpies are quoted for the gaseous state and should be recognised as average values obtained from a number of similar compounds. Cross reference with 11.2.6.

6.4.2 Calculate the enthalpy change of a reaction using bond enthalpies.

## 6.5 Entropy

6.5.1 State and explain the factors which increase the disorder (entropy) in a system.

An increase in disorder can result from the mixing of different types of particles, change of state (increased distance between particles), increased movement of particles or increased numbers of particles. An increase in the number of particles in the gaseous state usually has a greater influence than any other possible factor.

6.5.2 Predict whether the entropy change ( $\Delta S$ ) for a given reaction or process would be positive or negative.

From a given equation, identify a single factor which affects the value of  $\Delta S$  and predict the sign of  $\Delta S$ .

## 6.6 Spontaneity

6.6.1 Define standard free energy change of reaction ( $\Delta G^\theta$ ).

6.6.2 State whether a reaction or process will be spontaneous by using the sign of  $\Delta G^\theta$ .

6.6.3 State and predict the effect of a change in temperature on the spontaneity of a reaction, given standard entropy and enthalpy changes.

Use the equation  $\Delta G^\theta = \Delta H^\theta - T \Delta S^\theta$

## 7.1 Rates of Reaction

7.1.1 Define the term *rate of reaction* and describe the measurement of reaction rates.

Rate of reaction can be defined as the decrease in the concentration of reactants per unit time or the increase in the concentration of product per unit time.

7.1.2 Analyze data from rate experiments.

Graphs of changes in concentration, volume or mass against time should be interpreted qualitatively.

## 7.2 Collision Theory

7.2.1 Describe and explain the collision theory.

Students should know that not all collisions lead to a reaction.

7.2.2 Define *activation energy* ( $E_a$ ) and explain that reactions occur when reacting species have  $E > E_a$ .

Molecules must have a minimum energy and appropriate collision geometry in order to react. A simple treatment is all that is required. Cross reference with 5.1.3 and 5.1.4.

7.2.3 Predict and explain, using collision theory, the qualitative effect of particle size, temperature, concentration and catalysts on the rate of reaction.

Increasing the temperature increases the frequency of collisions but, more importantly, the proportion of molecules with  $E > E_a$  increases.

7.2.4 Explain that reactions can occur by more than one step and that one step can determine the rate of reaction.

Few reactions involve just one step although one step in the reaction, the *rate determining step*, determines the reaction rate. Orders of reactions and rate laws are not required.

## 8.1 Dynamic equilibrium

8.1.1 Outline the characteristics of a system in a state of equilibrium.

Many chemical reactions are reversible and never go to completion. Equilibrium can be approached from both directions. For a system in equilibrium the rate of the forward reaction equals the rate of the reverse reaction and the concentrations of all reactants and products remains constant. The system is closed and macroscopic properties remain constant.

Use phase equilibrium as an example of dynamic equilibrium involving physical changes.

## 8.2 The position of equilibrium

8.2.1 State the equilibrium constant expression ( $K_c$ ) for a homogeneous reaction.

Consider equilibria involving one phase, gases or species in aqueous solution. The equilibrium constant is specific to a given system and varies with temperature. No calculation is required.

8.2.2 Deduce the extent of reaction from the magnitude of the equilibrium constant.

When  $K_c > 1$ , products exceed reactants at equilibrium. When  $K_c < 1$ , reaction goes almost to completion.



When  $K_c < 1$ , reactants exceed products at equilibrium. When  $K_c > 1$ , reaction hardly proceeds.

8.2.3 Describe the qualitative effects of changes of temperature, pressure and concentration on the position of equilibrium and value of the equilibrium constant.

Use Le Chatelier's principle to predict the effects of these changes on the position of equilibrium. The value of the equilibrium constant ( $K_c$ ) is only affected by temperature. The position of equilibrium may change without the value of the  $K_c$  changing.

8.2.4 State and explain the effect of a catalyst on an equilibrium reaction.

8.2.4 Describe and explain the application of equilibrium and kinetics concepts in the Haber process and the Contact process.

## ***9.1 Properties of Acids and Bases***

9.1.1 Outline the characteristic properties of acids and bases in aqueous solution.

The properties that must be considered are : effects on indicators and reactions of acids with bases, metals and carbonates. Bases which are not hydroxides, such as ammonia, soluble carbonates and hydrogencarbonates, should be included. Alkalis are bases that dissolve in water.

## ***9.2 Strong and Weak Acids and Bases***

Note : Bronsted-Lowry definitions of acids and bases are not required for this sub-topic.

9.2.1 Describe and explain the differences between strong and weak acids and bases in terms of the extent of dissociation, reaction with water and conductivity.

The term *ionization* can be used instead of *dissociation*. Solutions of equal concentration can be compared by pH and/or conductivity.

9.2.2 State whether a given acid or base is strong or weak.

Specified strong acids are hydrochloric acid, nitric acid and sulfuric acid. Specified weak acids are ethanoic acid and carbonic acid (aqueous carbon dioxide). Specified strong bases are all group 1 hydroxides and barium hydroxide. Specified weak bases are ammonia and ethylamine.

9.2.3 Describe and explain data from experiments to distinguish between strong and weak acids and bases, and to determine the relative acidities and basicities of substances.

## ***9.3 The pH Scale***

9.3.1 Distinguish between aqueous solutions that are acidic, neutral or basic using the pH scale.

9.3.2 Identify which of two or more aqueous solutions is more acidic or basic, using pH values.

Measures pH using a pH meter or pH paper. Students should know that pH paper contains a mixture of indicators. The theory of pH meters is not required.

9.3.3 State that each change of one pH unit represents a tenfold change in the hydrogen ion concentration  $[H^+(aq)]$ .

Relate integral values of pH to  $[H^+(aq)]$  expressed as powers of ten.

Calculation of pH from  $[H^+(aq)]$  is not required.

9.3.4 Deduce changes in  $[H^+(aq)]$  when the pH of a solution changes by more than one pH unit.

## **9.4 Buffer Solutions**

9.4.1 Describe a buffer solution in terms of its composition and behaviour.

A buffer resists change in pH when a small amount of a strong acid or base is added. Suitable examples include

ammonium chloride/ammonia solution and ethanoic acid/sodium ethanoate.

Blood is an example of a buffer solution.

9.4.2 Describe ways of preparing buffer solutions.

## **9.5 Acid-base Titrations**

9.5.1 Draw and explain a graph showing pH against volume of titrant for titrations involving strong acids and bases.

## **10.1 Oxidation and Reduction**

10.1.1 Define *oxidation* and *reduction* in terms of electron loss and gain.

Introduce the concept of half-equation

10.1.2 Calculate the oxidation number of an element in a compound.

Oxidation numbers should be shown by a sign (+ or -), eg. +7 for Mn in  $KMnO_4$

10.1.3 State and explain the relationship between oxidation numbers and the names of compounds.

Oxidation numbers in the names of compounds are represented by Roman numerals, eg. Iron (II) oxide, Iron(III)oxide

10.1.4 Identify whether an element is oxidised or reduced and identify simple redox reactions using oxidation numbers.

Appropriate reactions to illustrate this can be found in topics 3 and 11.

Possible examples: iron (II) and (III), manganese(II) and (VII), chromium(III) and (VI), copper(I) and (II) oxides of sulfur and oxyacids, halogens and halide ions.

Define the terms oxidising agent and reducing agent.

## **10.2 Reactivity**

10.2.1 Deduce a reactivity series based upon the chemical behaviour of a group of oxidising and reducing agents

Displacement reactions of metals and halogens(see 3.3.1) provide a good experimental illustration of reactivity. Standard electrode potentials or reduction potentials are not required.

10.2.2 Deduce the feasibility of a redox reaction from a given reactivity series.

10.2.3 Describe and explain how a redox reaction is used to produce electricity in a voltaic cell.

Students should be able to draw a diagram of a simple half-cell and show how two half-cells can be connected by a salt bridge to form a whole cell. Suitable examples of half-cells are Mg, Zn, Fe and Cu in solutions of their ions.

## 10.3 Electrolysis

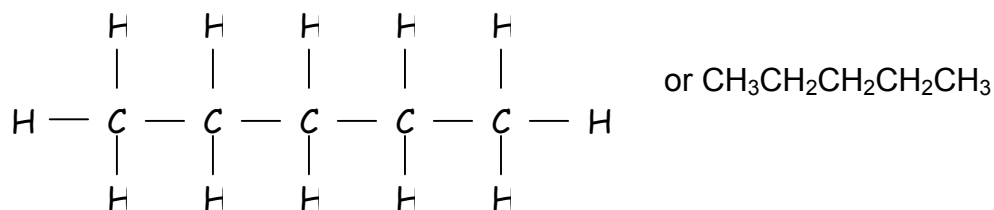
- 10.3.1 Draw a diagram identifying the essential components of an electrolytic cell.  
An electrolytic cell converts electrical energy to chemical energy. The diagram should include the source of the electric current and conductors, positive and negative electrodes and electrolyte.
- 10.3.2 Describe how current is conducted in an electrolytic cell.  
Students should describe how reactions occur.
- 10.3.3 Deduce the products for the electrolysis of a molten salt.  
Equations showing the formation of products at each electrode should be given.
- 10.3.3 Distinguish between the use of a spontaneous redox reaction to produce electricity in a voltaic cell and the use of electricity to carry out a non-spontaneous redox reaction in an electrolytic cell.  
Some teachers may wish to describe reactions at the electrodes in a cell in terms of reduction at the cathode and oxidation at the anode, but this is not required.
- 10.3.4 Describe and explain the use of electrolysis in electroplating.  
Restrict this to copper plating.

## 11.1 Homologous Series

- 11.1.1 Describe the features of a homologous series.  
Features include a general formula and neighbouring members differing by  $\text{CH}_2$ , with similar chemical properties and with a gradation in physical properties.
- 11.1.2 Predict and explain the trends in boiling points of members of a homologous series.  
In a homologous series there is a gradual increase in boiling points as the number of carbon atom increases. Cross reference with 4.3.

## 11.2 Hydrocarbons

- 11.2.1 Draw structural formulas for the isomers of the non-cyclic alkanes up to  $\text{C}_6$ .  
Structural formulas should indicate clearly the bonding between atoms. For example, for pentane :



- 11.2.2 State the names of alkanes up to  $\text{C}_6$ .  
Name these using IUPAC rules. Consider both straight and branch-chained alkanes.
- 11.2.3 Explain the relative inertness of alkanes.  
Refer to bond enthalpies. See 6.4.
- 11.2.4 Draw structural formulas and state the names for straight-chain alkenes ( $\text{C}_n\text{H}_{2n}$ , where  $n$  is between 2 and 5).

Geometric (*cis-trans*) isomers are not required.

11.2.5 Describe complete and incomplete combustion of hydrocarbons.

The formation of CO and C during incomplete combustion should be related to environmental impacts and oxidation-reduction.

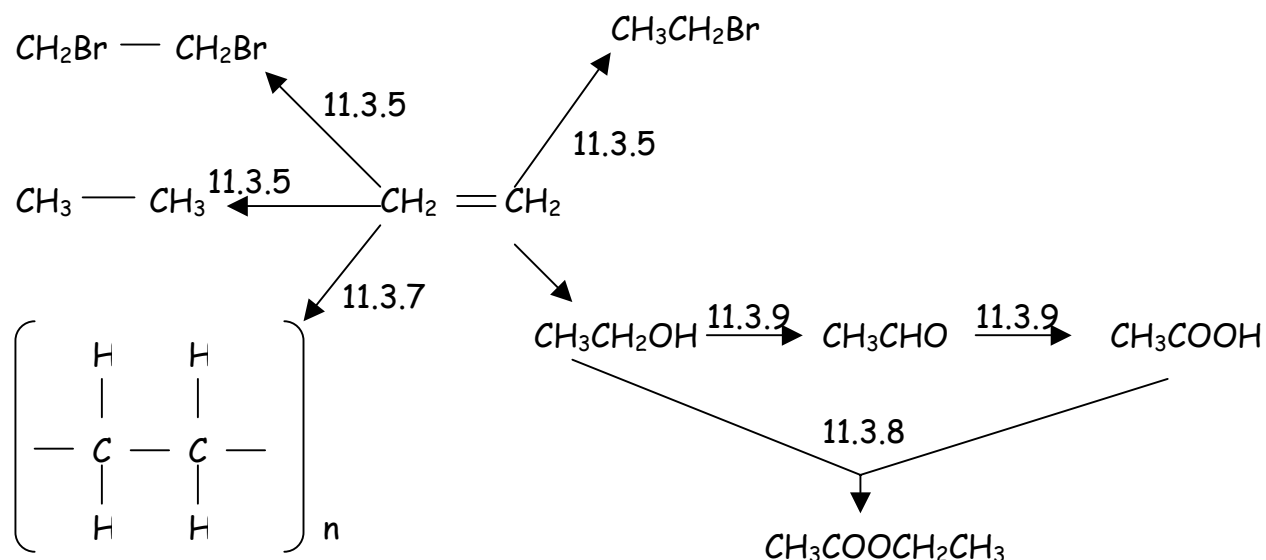
11.2.6 State that the combustion of hydrocarbons is an exothermic process.

See 6.3 and 6.4.

### 11.3 Other Functional Groups

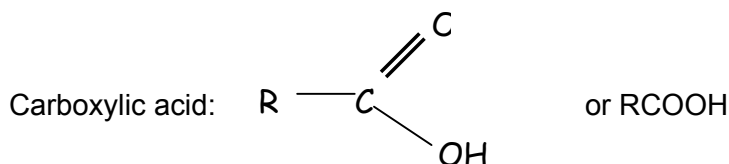
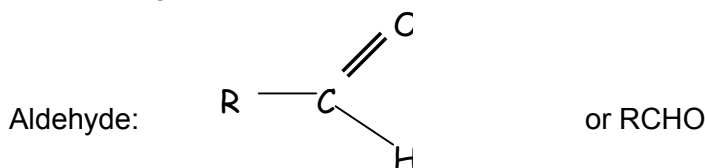
Along with alkanes and alkenes, compounds containing one or more functional groups have been chosen to introduce students to :

- interrelationships involving significant functional groups
- important reaction types such as addition, substitution, oxidation, condensation, esterification and polymerization.
- This is expressed in the following scheme :



11.3.1 Draw and state the names of compounds containing up to five carbon atoms with one of the following functional groups : aldehyde, ketone, carboxylic acid, alcohol, amide, amine, ester and halogenoalkane.

Functional groups in full and condensed forms are required, eg :



11.3.2 Explain that functional groups can exist as isomers.

Examples include :

ethanoic acid ( $\text{CH}_3\text{COOH}$ ) and methyl methanoate ( $\text{HCOOCH}_3$ ) propanal ( $\text{CH}_3\text{CH}_2\text{CHO}$ ) and propanone ( $\text{CH}_3\text{COCH}_3$ ).

11.3.3 Outline the existence of optical isomers.

Restrict this to the fact that, if a carbon atom has four different substituents, the molecule exists in two enantiomeric forms that rotate the plane of polarized light in opposite directions. Students should be able to identify a chiral (asymmetric centre).

## ***12.1 The Mass Spectrometer***

12.1.1 State the principles of a mass spectrometer and outline the main stages in its operation.

A simple diagram of a single beam mass spectrometer is required. The following stages of operation should be considered : vaporization, ionization, acceleration, deflection and detection.

12.1.2 Describe how the mass spectrometer may be used to determine relative isotopic, atomic and molecular masses using the  $^{12}\text{C}$  scale.

Students should be able to calculate the relative atomic mass from the abundance of the isotopes (see 2.1.6). Interpretation of fragmentation patterns is not required.

## ***12.2 Electron Configuration of Atoms***

12.2.1 State and explain how evidence from first and successive ionization energies accounts for the existence of the main energy levels and sub-levels.

Interpretation of graphs of first ionization and successive ionization energies versus atomic number provides evidence for the existence of the main energy levels and sub-levels.

12.2.2 State how orbitals are labelled.

Limit this to  $n < 5$ .

12.2.3 State the relative energies of s, p, d and f orbitals.

12.2.4 State the number of orbitals at each energy level.

12.2.5 Draw the shape of an s orbital and the shapes of the  $p_x$ ,  $p_y$  and  $p_z$  orbitals.

12.2.6 State the Aufbau principle.

Reference should be made to Hund's rule.

12.2.7 Apply the Aufbau principle to electron configurations.

Apply the Aufbau principle for an atom up to  $Z = 54$ , eg for  $Z = 23$  the electronic configuration is  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$  or  $[\text{Ar}] 4s^2 3d^3$  or  $[\text{Ar}] 3d^3 4s^2$ . Exceptions to this rule are not expected.

12.2.8 Relate the electron configuration of an atom to its position in the periodic table.

Students should be able to label the s, p, d and f blocks of the periodic table.

## ***13.1 Periodic Trends Na → Ar (the third period)***

13.1.1 Explain the physical properties of the chlorides and oxides of the elements in the third period (Na → Ar) in terms of their bonding and structure.

Refer to the following oxides and chlorides :

- Oxides :  $\text{Na}_2\text{O}$ ,  $\text{MgO}$ ,  $\text{Al}_2\text{O}_3$ ,  $\text{SiO}_2$ ,  $\text{P}_4\text{O}_6$  and  $\text{P}_4\text{O}_{10}$ ,  $\text{SO}_2$  and  $\text{SO}_3$ ,  $\text{Cl}_2\text{O}$  and  $\text{Cl}_2\text{O}_2$ .
- Chlorides :  $\text{NaCl}$ ,  $\text{MgCl}_2$ ,  $\text{Al}_2\text{Cl}_6$ ,  $\text{SiCl}_4$ ,  $\text{PCl}_3$  and  $\text{PCl}_5$  and  $\text{Cl}_2$  (sulfur chloride is not required).

Limit the explanation to the physical states of the compounds under standard conditions and electrical conductivity in the molten state only.

13.1.2 Describe the chemical trends for the chlorides and oxides referred to in

13.1.1. Include relevant equations.

Limit this to acid-base properties of the oxides and the reactions of the chlorides and oxides with water.

## **13.2 d-block Elements (first row)**

13.2.1 List the characteristic properties of transition elements.

Restrict this to variable oxidation states, complex ion formation, coloured compounds and catalytic properties.

13.2.2 Identify which elements are considered to be typical of the d-block elements. Sc and Zn are not typical.

13.2.3 Describe the existence of variable oxidation states in d-block elements.

The 4s and 3d sub-levels are close in energy. Students should know that all d-block elements can show an oxidation state of +2. In addition, they should be familiar with the oxidation states of the following : Cr (+3,+6), Mn(+4, +7), Fe (+3) and Cu (+1).

13.2.4 Define the term *ligand*.

13.2.5 Describe how complexes of d-block elements are formed.

Suitable examples are :  $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ ,  $[\text{Fe}(\text{CN})_6]^{3-}$ ,  $[\text{Cu}(\text{NH}_3)_4]^{2+}$ ,  $[\text{Ag}(\text{NH}_3)_2]^+$ . Only monodentate ligands are required.

13.2.6 Explain why some complexes of d-block elements are coloured.

Students need only know that in complexes the d orbitals are split into two sets at different energy levels and the electronic transitions that take place between them are responsible for their colours.

13.2.7 Outline the catalytic behaviour of d-block elements and their compounds.

Limit this to :

- $\text{MnO}_2$  in the decomposition of hydrogen peroxide
- $\text{V}_2\text{O}_5$  in the Contact process
- Fe in the Haber process
- Ni in the conversion of alkenes to alkanes.

The mechanisms of action are not required.

## **14.1 Shapes of Molecules and Ions**

14.1.1 State and predict the shape and bond angles using the VSEPR theory for 5- and 6-negative charge centers.

The shape of the molecules / ions and bond angles if all pairs of electrons are shared, and the shape of the molecules/ions if one or more lone pairs surround the central atom, should be considered. Examples such as  $\text{PCl}_5$ ,  $\text{SF}_6$  and  $\text{XeF}_4$  can be used.

## **14.2 Hybridization**

14.2.1 Describe  $\sigma$  and  $\pi$  bonds.

Treatment should be restricted to :

- bonds - electron distribution has axial symmetry around the axis joining the two nuclei
- $\pi$  bonds resulting from the combination of parallel p orbitals
- double bonds formed by a  $\sigma$  and a  $\pi$  bond
- triple bonds formed by a  $\sigma$  and two  $\pi$  bonds.

14.2.2 State and explain the meaning of the term *hybridisation*.

Hybridization should be explained in terms of the mixing of atomic orbitals to form new orbitals for bonding. Students should consider  $sp$ ,  $sp^2$  and  $sp^3$  hybridisation, and the shapes and orientation of these orbitals.

14.2.3 Discuss the relationships between Lewis structures, molecular shapes and types of hybridisation ( $sp$ ,  $sp^2$  and  $sp^3$ ).

Using examples from inorganic as well as organic chemistry, students should write the Lewis structure, deduce the shape of the molecule and recognise the type of hybridisation.

### **14.3 Delocalisation of Electrons**

14.3.1 State what is meant by the delocalisation of  $\pi$  electrons and explain how this can account for the structures of some substances.

Examples such as  $\text{NO}_3^-$ ,  $\text{NO}_2^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{O}_3$ ,  $\text{RCOO}^-$  and benzene can be used. (These could also be dealt with through the resonance approach).

### **14.4 Structures of Allotropes of Carbon**

14.4.1 Describe and explain the structures and properties of diamond, graphite and fullerene.

Students should recognize the type of hybridisation present in each allotrope and the delocalisation of electrons in graphite and  $\text{C}_{60}$  fullerene.

### **15.1 Standard Enthalpy Changes of Reaction**

15.1.1 Define and use the terms standard state and standard enthalpy change of formation ( $\Delta H_f^\ominus$ ).

15.1.2 Calculate the enthalpy change of a reaction using standard enthalpy changes of formation.

### **15.2 Lattice Enthalpy**

15.2.1 Define the term lattice enthalpy.

The sign of  $\Delta H_{\text{lattice}}$  indicates whether the lattice is being formed or broken.

15.2.2 Compare the effect of both the relative sizes and the charges of ions on the lattice enthalpies of different ionic compounds.

The relative value of the theoretical lattice enthalpy increases with higher ionic charge and smaller ionic radius due to increased attractive forces.

15.2.3 Construct a Born-Haber cycle and use it to calculate an enthalpy change.

15.2.4 Analyze theoretical and experimental lattice enthalpy values.

A significant difference between the two values indicates covalent character.

### **15.3 Spontaneity of a Reaction**

15.3.1 Calculate the standard entropy change for a reaction ( $\Delta S^\ominus$ ) using values of absolute entropies.

15.3.2 Calculate  $\Delta G^\ominus$  for a reaction using the equation  $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$  or by using values of the standard free energy change of formation,  $\Delta G_f^\ominus$ .

### **16.1 Rate Expression**

16.1.1 Define the terms rate constant and order of reaction.

16.1.2 Derive the rate expression for a reaction from data.

$$\text{Rate} = k[\text{A}]^m[\text{B}]^n$$

where  $k$  = rate constant,  $[\text{A}]$  = concentration of A in  $\text{mol dm}^{-3}$  etc.

$m$  and  $n$  = integers,  $m + n$  = overall order of the reaction.

16.1.3 Draw and analyze graphical representation for zero-, first- and second- order reactions.

16.1.4 Define the term *half-life* and calculate the half-life for first-order reactions only.

The half-life should be calculated from graphs and by using the integrated form of the rate equation. The integrated rate equation for second-order reactions is not required.

## **16.2 Reaction Mechanism**

16.2.1 Define the terms rate-determining step, molecularity and activated complex.

16.2.2 Describe the relationship between mechanism, order, rate-determining step and activated complex.

Limit examples to one- or two-step reactions where the mechanism is known. Students should understand what an activated complex (transition state) is and how the order of a reaction relates to the mechanism.

## **16.3 Activation Energy**

16.3.1 Describe qualitatively the relationship between the rate constant (k) and temperature (T).

16.3.2 Describe how the Arrhenius equation can be used to determine the activation energy and the Arrhenius constant (A).

Arrhenius equation :

$$k = Ae^{\left(\frac{-E_a}{RT}\right)}$$

A relates to the geometric requirements of the collisions (see 7.2). Direct substitution using simultaneous equations and a graphical method can be used. The logarithmic form of the Arrhenius equation is :

$$\ln k = -\frac{E_a}{RT} + \ln A$$

Both methods should be explained, but actual calculations are not needed.

16.3.3 Draw and explain enthalpy level diagrams for reactions with and without catalysts.

16.3.4 Distinguish between homogeneous catalysts and heterogeneous catalysts.

Homogeneous catalyst - reactants and catalyst are in the same phase.

Heterogeneous catalyst - reactants and catalyst are in different phases.

16.3.5 Outline the use of homogeneous and heterogeneous catalysts.

Examples include hydrogenation using metals (see 13.2.7) and acid catalyzed formation of esters.

## **17.1 Phase Equilibrium**

17.1.1 State and explain the equilibrium established between a liquid and its own vapor.

Liquid-vapor equilibrium is a dynamic equilibrium established when the rate of condensation equals the rate of vaporization. The vapor pressure is independent of the volume of the container, liquid or vapor.

17.1.2 State and explain the qualitative relationship between vapor pressure and temperature.

Students should be able to show the relationship graphically and explain it in terms of kinetic theory.

17.1.3 State and explain the relationship between enthalpy of vaporization, boiling point and intermolecular forces.



Students should be able to predict the relative strength of intermolecular forces of different liquids when given the physical properties, or vice versa. Cross reference with 4.3.

## 17.2 The Equilibrium Law

- 17.2.1 Solve homogeneous equilibrium problems using the expression for  $K_c$ . Calculate  $K_c$  given all equilibrium concentrations. Given  $K_c$  and other appropriate concentrations, find an equilibrium concentration.  $K_p$  and  $K_{sp}$  are not required, nor is use of the quadratic expression.

## 18.1 Bronsted-Lowry Acids and Bases

- 18.1.1 Define acids and bases according to the Bronsted-Lowry theory.  
18.1.2 Identify whether or not a compound could act as a Bronsted-Lowry acid or base.  
18.1.3 Identify the conjugate acid-base pairs in a given acid-base reaction.  
18.1.4 Determine the structure for the conjugate acid (or base) of any Bronsted-Lowry base (or acid).

The members of a conjugate acid-base pair always differ by a single proton ( $H^+$ ). Structures of conjugate acid-base pairs should always make clear the approximate location of the proton transferred, eg.  $CH_3COOH/CH_3COO^-$  rather than  $C_2H_4O_2/C_2H_3O_2^-$ .

## 18.2 Lewis Theory

- 18.2.1 Define and apply the terms *Lewis acid* and *Lewis base*.  
A Lewis acid-base reaction involves the formation of a new covalent bond in which both electrons are provided by one species. Such bonds are called dative covalent bonds. The formation of complexes (see 13.2.4 and 13.2.5) is usually a Lewis acid-base reaction.

## 18.3 Calculations Involving Acids and Bases

Note : A proton in water can be written as  $H^+(aq)$  or  $H_3O^+(aq)$ ; the former is adopted here.

- 18.3.1 State the expression for the ionic product constant of water ( $K_w$ ).  
 $K_w = [H^+(aq)][OH^-(aq)] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$  at 298 K but this varies with temperature.  
18.3.2 Deduce  $[H^+(aq)]$  and  $[OH^-(aq)]$  for water at different temperatures given  $K_w$  values.  
18.3.3 Define pH, pOH and  $pK_w$ .  
18.3.4 Calculate  $[H^+(aq)]$ ,  $[OH^-(aq)]$ , pH and pOH from specified concentrations.  
The values of  $[H^+(aq)]$  or  $[OH^-(aq)]$  are directly related to the concentration of the acid or base.  
18.3.5 State the equation for the reaction of any weak acid or weak base with water, and hence derive the ionization constant expression.  
In general  $HA(aq) \leftrightarrow H^+(aq) + A^-(aq)$   
 $B(aq) + H_2O(l) \leftrightarrow BH^+(aq) + OH^-(aq)$  (base hydrolysis)  
Then  $K_a = \frac{[H_A(aq)][A^-(aq)]}{[HA(aq)]}$  and  $K_b = \frac{[BH^-(aq)][OH^-(aq)]}{[B(aq)]}$   
Examples used should involve the transfer of only one proton.  
18.3.6 Derive the expression  $K_a \times K_b = K_w$  and use it to solve problems for any weak acid and its conjugate base and for any weak base and its conjugate acid.

- 18.3.7 State and explain the relationship between  $K_a$  and  $pK_a$  and between  $K_b$  and  $pK_b$ .
- 18.3.8 Determine the relative strengths of acids or their conjugate bases from  $K_a$  or  $pK_a$  values.
- 18.3.9 Apply  $K_a$  or  $pK_a$  in calculations.  
Calculations can be performed using various forms of the acid ionisation constant expression (see 18.3.5). Students should state when approximations are used in equilibrium calculations. Use of the quadratic expression is not required.
- 18.3.10 Calculate the pH of a specified buffer system.  
Calculations will involve the transfer of only one proton. Cross reference with 9.4.

## **18.4 Salt Hydrolysis**

- 18.4.1 State and explain whether salts form acidic, alkaline or neutral aqueous solutions.  
Examples should include salts formed from the four possible combinations of strong and weak acids and bases. The effect of the charge density of the cations in groups 1, 2, 3 and d-block elements should also be considered, eg.  $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$  or  $[\text{Fe}(\text{OH})(\text{H}_2\text{O})_5]^{2+} + \text{H}^+$ .

## **18.5 Acid -base Titrations**

- 18.5.1 Draw and explain the general shapes of graphs of pH against volume of titrant for titrations involving monoprotic acids and bases.  
All combinations should be covered : strong acid + strong base, strong acid + weak base, weak acid + strong base and weak acid + weak base.

## **18.6 Indicators**

- 18.6.1 Describe qualitatively how an acid-base indicator works.  
Use  $\text{HIn}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{In}^-(\text{aq})$  or similar colour A colour B
- 18.6.2 State and explain how the pH range of an acid-base indicator relates to its  $pK_a$  value.
- 18.6.3 Determine an appropriate indicator for a titration, given the equivalence point of the titration and  $K_a$  (or  $pK_a$ ) values for possible indicators.

## **19.1 Redox Equations**

- 19.1.1 Balance redox equations in acid solution.  
Half-equations and oxidation numbers may be used.  $\text{H}^+(\text{aq})$  and  $\text{H}_2\text{O}$  should be used where necessary to balance half-equations.

## **19.2 Standard Electrode Potentials**

**Note:** In 19.2.1 to 19.2.4 half-equations can be used to introduce redox couples, including  $\text{H}^+/\text{H}_2$  and a selection of common couples from the electrochemical series. The Daniell cell provides a good illustration of the principles under consideration here.

- 19.2.1 Describe the standard hydrogen electrode.  
Laboratory work using the standard hydrogen electrode is not required.

19.2.2 Define the term *standard electrode potential* and explain the measurement of standard electrode potentials to produce the electrochemical series.

19.2.3 Define the term *cell potential* and calculate cell potentials using standard electrode potentials.

19.2.4 Predict whether a reaction will be spontaneous using standard electrode potential ( $E^\ominus$ ) values.

Students should be able to predict the direction of electron flow in an external circuit and the reaction taking place in a cell.

Relate positive  $E^\ominus$  values for spontaneous reactions to negative  $\Delta G^\ominus$  values (see 6.6).

### 19.3 Electrolysis

19.3.1 List and explain the factors affecting the products formed in the electrolysis of aqueous solutions.

Factors to be considered are position in the electrochemical series, nature of the electrode and concentration. Suitable examples for electrolysis include water, aqueous sodium chloride and aqueous copper (II) sulfate.

19.3.2 List the factors affecting the amount of product formed during electrolysis.

Factors are charge on the ion, current and duration of electrolysis.

19.3.3 Determine the relative amounts of the products formed during the electrolysis of aqueous solutions.

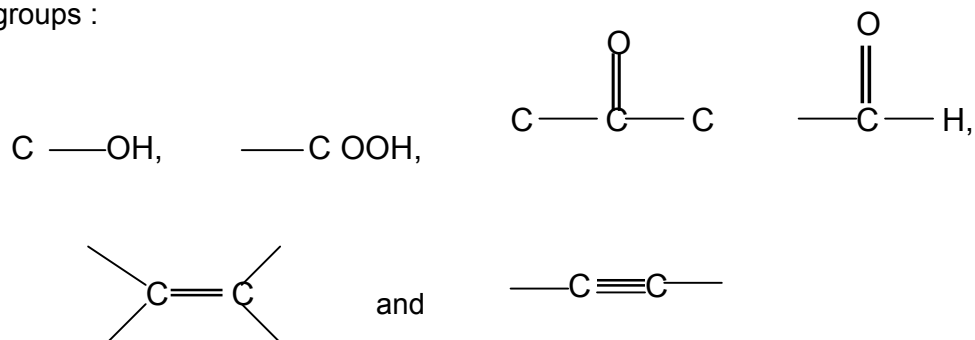
### 20.1 Determination of Structure

20.1.1 State that the structure of a compound can be determined using information from a variety of spectroscopic and chemical techniques.

Students should realize that information from only one technique is usually insufficient to determine or confirm a structure.

20.1.2 Describe and explain how information from an infrared spectrum can be used to identify functional groups in a compound.

Restrict this to using infrared spectra to show the presence of the functional groups :



and to match the fingerprint region to a known spectrum.

20.1.3 Describe and explain how information from a mass spectrum can be used to determine the structure of a compound.

Restrict this to using mass spectra to determine the relative molecular mass of a compound and to identify simple fragments, for example :

- $(M_r - 15)^+$  loss of  $\text{CH}_3$
- $(M_r - 29)^+$  loss of  $\text{C}_2\text{H}_5$  or  $\text{CHO}$

- $(M_r - 31)^-$  loss of  $\text{CH}_3\text{O}$
- $(M_r - 45)^+$  loss of  $\text{COOH}$

20.1.4 Describe and explain how information from a  $^1\text{H}$  NMR spectrum can be used to determine the structure of a compound.

Restrict this to using NMR spectra to determine the number of different environment in which hydrogen is found the number of hydrogen atoms in each environment. Splitting patterns are not required.

## 20.2 Hydrocarbons

20.2.1 State and explain the low reactivity of alkanes in terms of the inertness of C-H and C-C bonds.

20.2.2 State that alkanes can react with halogens and distinguish between *homolytic* and *heterolytic fission*.

Students should be able to define and recognize a *free radical*. Mechanisms are not required.

20.2.3 Describe and explain the structure of benzene using chemical and physical evidence.

Consider the special stability of the ring system (heat of combustion or hydrogenation of  $\text{C}_6\text{H}_6$  in comparison to that of cyclohexene, cyclohexadiene and cyclohexatriene), as well as benzene's tendency to undergo substitution rather than addition reactions.

## 20.3 Nucleophilic Substitution Reaction

20.3.1 Distinguish between primary, secondary and tertiary halogenoalkanes.

20.3.2 Describe and explain the  $\text{S}_{\text{N}}1$  and  $\text{S}_{\text{N}}2$  mechanisms in nucleophilic substitution.

Students must be able to draw a stepwise mechanism. Examples of nucleophiles should include  $\text{CN}^-$ ,  $\text{OH}^-$  and  $\text{NH}_3$  for each reaction type.

20.3.3 Describe and explain the molecularity for the  $\text{S}_{\text{N}}1$  and  $\text{S}_{\text{N}}2$  mechanisms.

The predominant mechanism for tertiary halogenoalkanes is  $\text{S}_{\text{N}}1$  and for primary halogenoalkanes it is  $\text{S}_{\text{N}}2$ . Both mechanisms occur for secondary halogenoalkanes.

20.3.4 Describe how the rate of nucleophilic substitution in halogenoalkanes depends on both the identity of the halogen and whether the halogenoalkane is primary, secondary or tertiary.

## 20.4 Alcohols

20.4.1 Describe the dehydration reaction of alcohols to form alkenes.

20.4.2 Determine the products formed by the oxidation of primary, secondary and tertiary alcohols using acidified potassium dichromate (VI) solution.

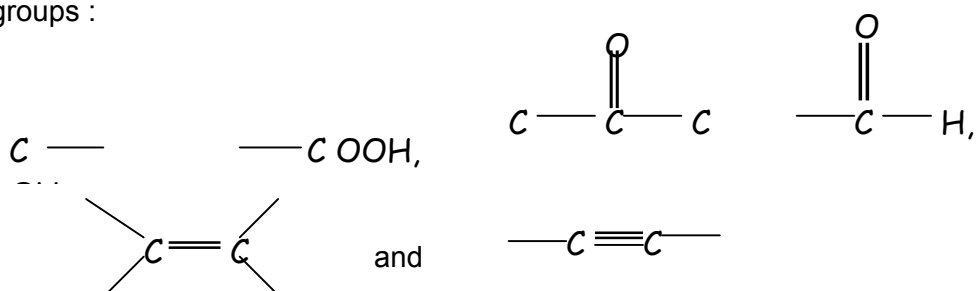
A.1 Determination of Structure

A.1.1 State that the structure of a compound can be determined using information from a variety of spectroscopic and chemical techniques.

Students should realize that information from only one technique is usually insufficient to determine or confirm a structure.

A.1.2 Describe and explain how information from an infrared spectrum can be used to identify functional groups in a compound.

Restrict this to using infrared spectra to show the presence of the functional groups :



and to match the fingerprint region to a known spectrum.

A.1.3 Describe and explain how information from a mass spectrum can be used to determine the structure of a compound.

Restrict this to using mass spectra to determine the relative molecular mass of a compound and to identify simple fragments, for example :

- $(M_r - 15)^-$  loss of  $\text{CH}_3$
- $(M_r - 29)^-$  loss of  $\text{C}_2\text{H}_5$  or  $\text{CHO}$
- $(M_r - 31)^-$  loss of  $\text{CH}_3\text{O}$
- $(M_r - 45)^+$  loss of  $\text{COOH}$

A.1.4 Describe and explain how information from a  $^1\text{H}$  NMR spectrum can be used to determine the structure of a compound.

Restrict this to using NMR spectra to determine the number of different environment in which hydrogen is found the number of hydrogen atoms in each environment. Splitting patterns are not required.

A.1.5 Describe and explain the structure of benzene using chemical and physical evidence.

Consider the special stability of the ring system (heat of combustion or hydrogenation of  $\text{C}_6\text{H}_6$  in comparison to that of cyclohexene, cyclohexadiene and cyclohexatriene), as well as benzene's tendency to undergo substitution rather than addition reactions.

## ***A.2 Rate Expression***

A.2.1 Define the terms rate constant and order of reaction.

A.2.2 Derive the rate expression for a reaction from data.

$$\text{Rate} = k[\text{A}]^m[\text{B}]^n$$

where  $k$  = rate constant,  $[\text{A}]$  = concentration of A in  $\text{mol dm}^{-3}$  etc.

$m$  and  $n$  = integers,  $m + n$  = overall order of the reaction.

A.2.3 Draw and analyse graphical representation for zero-, first- and second- order reactions.

A.2.4 Define the term *half-life* and calculate the half-life for first-order reactions only.

The half-life should be calculated from graphs and by using the integrated form of the rate equation. The integrated rate equation for second-order reactions is not required.

## ***A.3 Reaction Mechanism***

A.3.1 Define the terms rate-determining step, molecularity and activated complex.

A.3.2 Describe the relationship between mechanism, order, rate-determining step and activated complex.

Limit examples to one- or two-step reactions where the mechanism is known. Students should understand what an activated complex (transition state) is and how the order of a reaction relates to the mechanism.

## ***A.4 Nucleophilic Substitution Reaction***

A.4.1 Distinguish between primary, secondary and tertiary halogenoalkanes.

A.4.2 Describe and explain the S<sub>N</sub>1 and S<sub>N</sub>2 mechanisms in nucleophilic substitution. Students must be able to draw a stepwise mechanism. Examples of nucleophiles should include -CN, -OH and NH<sub>3</sub> for each reaction type.

A.4.3 Describe and explain the molecularity for the S<sub>N</sub>1 and S<sub>N</sub>2 mechanisms. The predominant mechanism for tertiary halogenoalkanes is S<sub>N</sub>1 and for primary halogenoalkanes it is S<sub>N</sub>2. Both mechanisms occur for secondary halogenoalkanes.

A.4.4 Describe how the rate of nucleophilic substitution in halogenoalkanes depends on both the identity of the halogen and whether the halogenoalkane is primary, secondary or tertiary.

## ***A.5 Calculations Involving Acids and Bases***

Note: A proton in water can be written as H<sup>+</sup>(aq) or H<sub>3</sub>O<sup>+</sup>(aq); the former is adopted here.

A.5.1 State the expression for the ionic product constant of water (K<sub>w</sub>).

$K_w = [H^+(aq)][OH^-(aq)] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$  at 298 K but this varies with temperature.

A.5.2 Deduce [H<sup>+</sup>(aq)] and [OH<sup>-</sup>(aq)] for water at different temperatures given K<sub>w</sub> values.

A.5.3 Define pH, pOH and pK<sub>w</sub>.

A.5.4 Calculate [H<sup>+</sup>(aq)], [OH<sup>-</sup>(aq)], pH and pOH from specified concentrations.

The values of [H<sup>+</sup>(aq)] or [OH<sup>-</sup>(aq)] are directly related to the concentration of the acid or base.

A.5.5 State the equation for the reaction of any weak acid or weak base with water, and hence derive the ionization constant expression.

In general  $HA(aq) \leftrightarrow H^+(aq) + A^-(aq)$

$B(aq) + H_2O(l) \leftrightarrow BH^+(aq) + OH^-(aq)$  (base hydrolysis)

Then  $K_a = \frac{[H_A(aq)][A^-(aq)]}{[HA(aq)]}$  and  $K_b = \frac{[BH^-(aq)][OH^-(aq)]}{[B(aq)]}$

Examples used should involve the transfer of only one proton.

A.5.6 State and explain the relationship between K<sub>a</sub> and pK<sub>a</sub> and between K<sub>b</sub> and pK<sub>b</sub>.

A.5.7 Determine the relative strengths of acids or their conjugate bases from  $K_a$  or  $pK_a$  values.

A.5.8 Apply  $K_a$  or  $pK_a$  in calculations.

Calculations can be performed using various forms of the acid ionisation constant expression (see 18.3.5). Students should state when approximations are used in equilibrium calculations. Use of the quadratic expression is not required.

A.5.9 Calculate the pH of a specified buffer system.

Calculations will involve the transfer of only one proton. Cross reference with 9.4.

## ***B.1 Pharmaceutical Products***

B.1.1 List the effects of drugs and medicines

Generally a drug or medicine is any chemical which does one or more of the following:

- alters incoming sensory sensations
- alters mood or emotions
- alters physiological state, including consciousness, activity level or coordination.

Stress the importance of the body's natural healing processes and the placebo effect.

B.1.2 Outline the stages involved in research, development and testing of new pharmaceutical products.

Refer to the Thalidomide case as an example of what can go wrong. The use of combinatorial chemistry is not required here, but is covered in B.8.4.

B.1.3 Describe the different methods of administering drugs.

The four main methods are oral, rectal, inhalation, and parenteral (by injection). Injection may be intravenous, intramuscular, or subcutaneous.

B.1.4 Discuss terms *lethal dosage* ( $LD_{50}$ ) *tolerance*, and *side effects*.

$LD_{50}$  is the lethal dose required for 50% of the population.

A person who develops tolerance requires a larger dose of a drug in order to achieve the effect originally obtained by a smaller dose. Stress that the difference between the main effect and the side effects is relative. For example, morphine is often used as a pain killer with intestinal constipation being a side effect. For a person with diarrhoea the constipation induced becomes the main effect, with the pain relief being the side effect. The risk:benefit ratios should be considered.

## ***B.2 Antacids***

B.2.1 State and explain how excess acidity in the stomach can be reduced by the use of different bases.

Examples should include aluminum and magnesium compounds and sodium hydrogen carbonate. Students should be able to write balanced equations for neutralization reactions and know that antacids are often combined with alginates (which produce a neutralizing layer preventing the acid in the stomach from rising into the esophagus and causing heartburn), and with anti-foaming agents (such as dimethicone).

## ***B.3 Analgesics***

B.3.1 Describe and explain the different ways that analgesics prevent pain.

Mild analgesics function by intercepting the pain stimulus at the source, often by interfering with the production of substances (eg prostaglandins) that cause pain, swelling or fever. Strong analgesics work by temporarily bonding to receptor sites in the brain, preventing the transmission of pain impulses without depressing the central nervous system.

B.3.2 describe the use of derivatives of salicylic acid as mild analgesics and compare the advantages and disadvantages of using aspirin and paracetamol (acetaminophen)

Aspirin has been found to be useful in preventing the recurrence of heart attacks. The disadvantages of aspirin include ulceration and stomach bleeding, allergic reactions and Reye's syndrome in children (a potentially fatal liver and brain disorder). Paracetamol is very safe in the correct dose but can, rarely, cause blood disorders and kidney damage. Overdosage can lead to serious liver damage, brain damage and even death.

B.3.3 Compare the structures of morphine, codeine and the semi-synthetic opiate, heroin.

Stress the simple modification to the structure of morphine which results in the semi-synthetic drug, heroin.

B.3.4 Discuss the advantages and disadvantages of using morphine and its derivative as strong analgesics.

Include the social as well as physiological effects of both short - and long-term use.

## ***B.4 Depressants***

B.4.1 Describe the effects of depressants.

At low doses a depressant may exert little or no effect. At moderate doses the compound may induce sedation (soothing, reduction of anxiety). At higher doses it may induce sleep and at extremely high doses it may cause death. Depressants are often prescribed as anti-depressants because they relieve depression.

B.4.2 Discuss the social and physiological effects of the use and abuse of ethanol. Include effects on the family, cost to society and the short- and long-term health effects.

**\*B.4.3 Describe and explain the techniques used for the detection of ethanol in the breath and in the blood or urine.**

**Include potassium(VI)dichromate in the breathalyzer, analysis of the blood or urine by chromatography and absorption of infra-red radiation in the intoximeter.**

B.4.4 Describe the synergistic effects of ethanol with other drugs.

Examples include increased risk of stomach bleeding with aspirin, and increased risk of heavy sedation with any drug that has a sedative effect on the central nervous system.

B.4.5 List other commonly used depressants and describe their structure.

Limit this to a brief mention of the use of diazepam (Valium®), nitrazepam (Mogadon®) and fluoxetine hydrochloride (Prozac®)

## ***B.5 Stimulants***

B.5.1 List the physiological effects of stimulants.

B.5.2 Compare amphetamines and adrenaline.

Amphetamines and adrenaline are chemically similar in that both derive from the phenylethylamine structure. Amphetamines mimic the effects of adrenaline and are known as sympathomimetic drugs.



- B.5.3 Discuss the short- and long-term effects of nicotine consumption.
- Short-term effects: increased heart rate and blood pressure and reduction of urine output, as well as stimulating effects.
  - Long-term effects: increased risk of heart disease, coronary thrombosis and peptic ulcers. Discuss also the addictive properties of nicotine and the further risks of associated with smoking tobacco.
- B.5.4 Describe the effects of caffeine and compare its structure with that of nicotine. Caffeine is a respiratory stimulant. When consumed in large amounts it can cause anxiety, irritability and sleeplessness. It is a weak diuretic. Both caffeine and nicotine contain tertiary amine group.

## ***B.6 Antibacterials***

- B.6.1 Outline the historical development of penicillins.  
Include the discovery by Fleming and the development by Florey and Chain.
- B.6.2 Compare *broad-spectrum* and *narrow-spectrum* antibiotics.
- B.6.3 Explain how penicillins work and discuss the effects of modifying the side chain.  
Penicillins work by interfering with the chemicals that bacteria need to form normal cell walls. Modifying the side chain results in penicillins which are more resistant to the penicillinase enzyme.
- B.6.4 Discuss and explain the effect overprescription of penicillins has, and the use of penicillins in animal feedstock.

## ***B.7 Antivirals***

- B.7.1 State how viruses are different from bacteria.
- B.7.2 Describe the different ways in which antiviral drugs work.  
Antiviral drugs may work by altering the cell's genetic material so that the virus cannot use it to multiply. Alternatively they may prevent the viruses from multiplying by blocking enzyme activity within the host cell.
- B.7.3 Discuss the difficulties associated with solving AIDS problems.  
Specific proteins on the HIV virus bind to receptor protein on certain white blood cells (T cells). Because of the ability of the HIV viruses to mutate and because their metabolism is linked closely with that of the cell, effective treatment with antiviral drugs is very difficult, as is vaccine development.

## **Higher level**

### ***B.8 Stereochemistry in Drug Action and Design***

- B.8.1 Describe the importance of geometrical isomerism in drug action.  
Students should be aware that *cis*- and *trans*-isomerism can occur in inorganic complexes and that the two different isomers can have different pharmacological effects. The anti-cancer drug cisplatin is a good example.
- B.8.2 Discuss the importance of chirality in drug action.  
The two enantiomers in a racemic mixture of a drug may have very different effects, eg. Thalidomide. One enantiomer of Thalidomide alleviates morning sickness in pregnant women, whilst the other enantiomer causes deformities in the limbs of the fetus.
- B.8.3 Describe the use of chiral auxiliaries to form the desired enantiomer.  
A chiral auxiliary is used to convert a non-chiral molecule into just the desired enantiomer, thus avoiding the need to separate enantiomers from a racemic mixture. It works by attaching itself to the non-chiral molecules to create the stereochemical conditions necessary to force the reaction to follow a certain

path. Once the new molecule has been formed, the auxiliary can be taken off (recycled) to leave the desired enantiomer. An example is the synthesis of Taxol, an anti-cancer drug.

B.8.4 Explain the use of combinatorial chemistry to synthesis new drugs.

Combinatorial chemistry is used to synthesize a large number of different compounds and screen them for biological activity, resulting in a 'combinatorial library' (for example the 'mix and split' process whereby polypeptides can be made by every combination of amino acids, using polystyrene resin beads). Stress the importance of solid phase chemistry.

## ***B.9 Anesthetics***

B.9.1 Compare local and general anesthetics in terms of their mode of action.

B.9.2 Compare the structures and effects of cocaine, procaine and lidocaine.

B.9.3 Discuss the advantages and disadvantages of nitrous oxide, ethoxyethane, trichloromethane, cyclopropane and halothane.

Nitrous oxide is not very potent, trichloromethane leads to liver damage, ethoxyethane and cyclopropane are highly flammable. Halothane(2-bromo-2-chloro-1,1,1-trifluoroethane) is widely used but is potentially harmful to the ozone layer.

B.9.4 Calculate the partial pressures of component gases in an anesthetic mixture.

Knowledge of how to use Dalton's law of partial pressures is required. Students are not expected to state the law.

## ***B.10 Mind-altering drugs***

B.10.1 Describe the effects of lysergic acid diethylamide(LSD), mescaline, psilocybin and tetrahydrocannabinol(THC).

B.10.2 Discuss the structural similarities and differences between LSD, mescaline, and psilocybin.

Stress the similarity of all three drugs and compare them to the indole ring.

B.10.3 Discuss the arguments for and against legalization of cannabis.

Arguments for legalization include the ability of cannabis to offer relief for certain diseases. Arguments against legalization include the possible harmful effects and the possibility of cannabis users moving on to harder drugs.

## ***D.1 Primary Air Pollution***

D.1.1 Describe the sources of carbon monoxide, oxides of nitrogen and sulfur, particulates and hydrocarbons in the atmosphere.

Include both natural and man-made sources. Balanced equations should be used where possible.

D.1.2 Outline the effects of primary air pollution on health.

Students should be familiar with at least one harmful effect of each of the substances in D.1.1.

D.1.3 Discuss methods for the reduction of primary air pollution.

Limit this to the following methods :

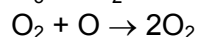
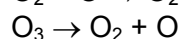
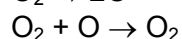
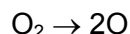
- CO - catalytic converters
- NO<sub>x</sub> - catalytic converters, lean burn engines, recirculation of exhaust gases
- SO<sub>x</sub> - alkaline scrubbing, removal of sulfur-containing compounds from coal and oil, limestone-based fluidized beds
- Particulates - electrostatic precipitation

- Hydrocarbons - catalytic converters.

## ***D.2 Ozone Depletion***

D.2.1 Describe the formation and depletion of ozone by natural processes.

Refer to the following equations.



D.2.2 List the pollutants, and their sources, that cause the lowering of ozone concentration.

Consider chlorofluorocarbons (CFCs) and nitrogen oxides.

D.2.3 State the environmental effects of ozone depletion.

Include the increased incidence of skin cancer and eye cataracts, and the suppression of plant growth.

D.2.4 Discuss the alternatives to CFCs in terms of their properties.

Alternatives include hydrocarbons, fluorocarbons and hydrofluorocarbons (HFCs). Include toxicity, flammability, the relative weakness of the C-Cl bond and the ability to absorb infrared radiation.

## ***D.3 Greenhouse Effect and Global Warming***

D.3.1 Describe the *greenhouse* effect.

Greenhouse gases allow the passage of incoming solar radiation but absorb the heat radiation from the Earth, maintaining a mean global temperature.

The greenhouse effect is a normal and necessary condition for life on Earth.

D.3.2 List the main greenhouse gases and their sources, and discuss their relative effects.

The greenhouse gases to be considered are CH<sub>4</sub>, H<sub>2</sub>O, CO<sub>2</sub> and N<sub>2</sub>O which have natural and man-made origins. Their effects depend on their abundance and their ability to absorb heat radiation.

D.3.3 Discuss the influence of increasing amounts of greenhouse gases on global warming.

Effects include climate change, thermal expansion of the oceans and melting of the polar ice caps.

D.3.4 Outline the influence of particulates on the Earth's surface temperature.

Particulates can lower the temperature by reflecting sunlight.

## ***D.4 Acid Rain***

D.4.1 State what is meant by acid rain and outline its origins.

Rain is naturally acidic because of dissolved CO<sub>2</sub>; acid rain has a pH of less than 5.6. Acid rain is caused by oxides of sulfur and nitrogen. Students should know the equations for the burning of sulfur and nitrogen and for the formation of H<sub>2</sub>SO<sub>3</sub> and H<sub>2</sub>SO<sub>4</sub>.

D.4.2 Discuss the environmental effects of acid rain and possible methods to counteract them.

## ***D.5 Water Suitable for Drinking***

D.5.1 Discuss the demand for fresh water and reasons for the inadequacy of its supply.

Only a small fraction of the Earth's water supply is fresh water. Of this fresh water, over 80% is in the form of ice caps and glaciers. Water is mainly used for agriculture and industry.

D.5.2 Compare the advantages and disadvantages of treating drinking water with chlorine and ozone.

Include cost, retention time and formation of chlorinated organic compounds.

D.5.3 Discuss ways to obtain fresh water from sea water using distillation, reverse osmosis and ion exchange.

D.5.4 Discuss ways to reduce the amount of water used and to recycle water.

## ***D.6 Dissolved Oxygen in Water***

D.6.1 Outline the importance of dissolved oxygen in water.

D.6.2 Outline biological oxygen demand (BOD) as a measure of oxygen - demanding wastes in water.

Refer to the amount of oxygen needed to decompose waste matter over a definite period of time. No distinction between biological and biochemical oxygen demand will be made.

D.6.3 Distinguish between aerobic and anaerobic decomposition of organic material in water.

D.6.4 Describe the influence of sewage, detergents and fertilisers on the growth of aquatic plants, and the effect of their subsequent decomposition on oxygen concentration (eutrophication).

The additional nitrogen and phosphorus compounds encourage growth of aquatic plants often in the form of 'algal blooms' or, in coastal areas, 'red tides'.

D.6.5 Discuss the effect of heat on dissolved oxygen and metabolism in water.

## ***D.7 Waster Water Treatment***

D.7.1 Outline the primary and secondary stages of sewage treatment and state what is removed during each stage.

For primary treatment filtration, flocculation and sedimentation should be covered. For secondary treatment mention the use of oxygen and bacteria (eg. the activated sludge process).

D.7.2 Discuss the increasing use of tertiary treatment.

Include removal of heavy metals and phosphates by chemical precipitation and nitrates by chemical or biological processes.

## **Higher level**

### ***D.8 Smog***

D.8.1 Compare reducing and photochemical smog.

D.8.2 Describe the catalytic effect of particulates and nitrogen oxides on the oxidation of sulfur dioxide.

- Particulates and  $\text{SO}_2$  - heterolytic catalysis to form  $\text{SO}_3$
- $\text{NO}_x + \text{SO}_2$  - free radical catalysis to form  $\text{SO}_3$

D.8.3 Outline the formation of secondary pollutants in photochemical smog.

Treatment should be restricted to the formation of radicals from the reaction of nitrogen oxides with sunlight and the reaction of these radicals with

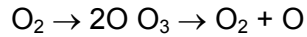
hydrocarbons, leading to the formation of aldehydes and peroxyacylnitrates (PANs).

D.8.4 Discuss the formation of thermal inversions and their effects on air quality.

### ***D.9 Ozone Depletion***

D.9.1 Explain the dependence of O<sub>2</sub> and O<sub>3</sub> dissociation on the wavelength of light.

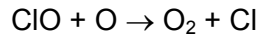
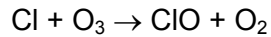
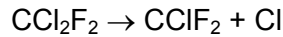
$\lambda = 242 \text{ nm}$   $\lambda = 330 \text{ nm}$



The energy needed should be related to the bonding in O<sub>2</sub> and O<sub>3</sub>.

D.9.2 Describe the steps in the catalysis of O<sub>3</sub> depletion by CFCs and NO<sub>x</sub>.

For example :



NO<sub>x</sub> similar pathway

D.9.3 Outline the reasons for greater ozone depletion in polar regions.

Consider the seasonal variation in temperature in the upper atmosphere.

Refer to surface catalysis on ice particles.

D.9.4 Describe the properties required for sun-screening compounds.

Such compounds should contain conjugated double bonds, eg.

paraaminobenzoic acid (PABA), so that absorption of ultraviolet light is possible.

### ***D.10 Toxic Substances in Water***

D.10.1 Discuss the different approaches to expressing toxicity.

Include the advantages and disadvantages of LD<sub>50</sub> (lethal dose in 50% of the population) and maximum daily tolerance.

D.10.2 State the principal toxic types of chemicals that may be found in polluted water.

Include heavy metals, pesticides, dioxins and polychlorinated biphenyls (PCBs).

D.10.3 Outline the sources, health and environmental effects of cadmium, mercury and lead compounds.

- Cadmium - metal plating, some rechargeable batteries, pigments
- Mercury - seed dressing to prevent mould, batteries
  - Lead - some kinds of paint, as tetraethyl lead in gasoline

D.10.4 Describe the sources and possible health effects of nitrates in drinking water.

Include the formation of carcinogenic nitrosamines and a possible link to the formation of nitrites leading to oxygen depletion in the body.