

## AS Level Chemistry - HEAD

3.1 Physical chemistry	3.1.1 Atomic structure	<b>3.1.1.1 Fundamental particles</b> <ul style="list-style-type: none"> <li>Appreciate how knowledge and understanding of atomic structure has evolved over time.</li> <li>Describe modern atomic theory.</li> </ul>		<b>3.1.1.2 Mass number and isotopes</b> <ul style="list-style-type: none"> <li>Determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge.</li> <li>Explain the existence of isotopes.</li> <li>Interpret simple mass spectra of elements to calculate relative atomic mass from isotopic abundance.</li> </ul>		<b>3.1.1.3 Electron configuration</b> <ul style="list-style-type: none"> <li>Define '<i>first ionisation energy</i>'.</li> <li>Write equations for first and successive ionisation energies.</li> <li>Explain how first and successive ionisation energies in Period 3 (Na–Ar) and in Group 2 (Be–Ba) give evidence for electron configuration in sub-shells and in shells.</li> </ul>								
	3.1.2 Amount of substance	<b>3.1.2.1 Relative atomic mass and relative molecular mass</b> In terms of $^{12}\text{C}$ : <ul style="list-style-type: none"> <li>Define '<i>relative atomic mass</i>' (<math>A_r</math>).</li> <li>Define '<i>relative molecular mass</i>' (<math>M_r</math>).</li> </ul>	<b>3.1.2.2 The mole and the Avogadro constant</b> Perform calculations using the Avogadro constant: <ul style="list-style-type: none"> <li>using <math>M_r</math> and amount in moles.</li> <li>using concentration, volume and amount of substance in a solution.</li> </ul>		<b>3.1.2.3 The ideal gas equation</b> Use the ideal gas equation $pV = nRT$ in equations.	<b>3.1.2.4 Empirical and molecular formula</b> <ul style="list-style-type: none"> <li>Calculate empirical formula from composition by mass or percentage by mass data.</li> <li>Calculate molecular formula from the empirical formula and relative molecular mass.</li> </ul>		<b>3.1.2.5 Balanced equations and associated calculations</b> <ul style="list-style-type: none"> <li>Write balanced equations for reactions studied</li> <li>Balance equations for unfamiliar reactions when reactants and products are specified.</li> <li>Use balanced equations to calculate:                             <ul style="list-style-type: none"> <li>- masses</li> <li>- volumes of gases</li> <li>- percentage yields</li> <li>- percentage atom economies</li> <li>- concentrations and volumes for reactions in solutions.</li> </ul> </li> <li>Make up a volumetric solution and carry out a simple acid–base titration (<b>RP</b>).</li> </ul>						
	3.1.3 Bonding	<b>3.1.3.1 Ionic bonding</b> <ul style="list-style-type: none"> <li>Predict the charge on a simple ion using the position of the element in the Periodic Table.</li> <li>Construct formulas for ionic compounds.</li> </ul>		<b>3.1.3.2 Nature of covalent and dative covalent bonds</b> Represent covalent and co-ordinate bonds using lines and arrows.		<b>3.1.3.3 Metallic bonding</b> Describe the relationship between delocalised electrons and positive ions in metallic bonding.		<b>3.1.3.4 Bonding and physical properties</b> <ul style="list-style-type: none"> <li>Describe the four types of crystal structure as either ionic, metallic, giant covalent or molecular.</li> <li>Use diamond, graphite, ice, iodine, magnesium and sodium chloride as examples of the four types of crystal structure.</li> <li>Relate the melting point and conductivity of materials to the type of structure and the bonding present.</li> <li>Explain using diagrams the energy changes associated with changes of state.</li> </ul>		<b>3.1.3.5 Shapes of simple molecules and ions</b> Explain the shapes of, and bond angles in, simple molecules and ions with up to six electron pairs (including lone pairs of electrons) surrounding the central atom.		<b>3.1.3.6 Bond polarity</b> Define ' <i>electronegativity</i> ' explain why some molecules with polar bonds do not have a permanent dipole.		<b>3.1.3.7 Forces between molecules</b> <ul style="list-style-type: none"> <li>Explain the existence of these forces between familiar and unfamiliar molecules: permanent dipole–dipole forces, induced dipole–dipole (van der Waals, dispersion, London) forces, hydrogen bonding.</li> <li>Explain how melting and boiling points are influenced by these intermolecular forces.</li> </ul>

## AS Level Chemistry - HEAD

<b>3.1.4 Energetics</b>	<p><b>3.1.4.1 Enthalpy change</b></p> <ul style="list-style-type: none"> <li>Define standard enthalpy of combustion (<math>\Delta_c H^\ominus</math>).</li> <li>Define standard enthalpy of formation (<math>\Delta_f H^\ominus</math>).</li> </ul>	<p><b>3.1.4.2 Calorimetry</b></p> <ul style="list-style-type: none"> <li>Recall that the heat change, <math>q</math>, in a reaction is given by the equation <math>q = mc\Delta T</math> where <math>m</math> is the mass of the substance that has a temperature change <math>\Delta T</math> and a specific heat capacity <math>c</math>.</li> <li>Use this equation to calculate the molar enthalpy change for a reaction and in related calculations.</li> <li>Measurement of an enthalpy change. <b>(RP)</b></li> </ul>	<p><b>3.1.4.3 Applications of Hess's law</b></p> <p>Use Hess's law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation.</p>	<p><b>3.1.4.4 Bond enthalpies</b></p> <ul style="list-style-type: none"> <li>Define the term '<i>mean bond enthalpy</i>'</li> <li>Use mean bond enthalpies to calculate an approximate value of <math>\Delta_H</math> for reactions in the gaseous phase.</li> <li>Explain why values from mean bond enthalpy calculations may vary.</li> </ul>	
<b>3.1.5 Kinetics</b>	<p><b>3.1.5.1 Collision theory</b></p> <p>In terms of chemical reactions:</p> <ul style="list-style-type: none"> <li>define the term '<i>activation energy</i>'.</li> <li>explain why most collisions do not lead to a reaction.</li> </ul>	<p><b>3.1.5.2 Maxwell–Boltzmann distribution</b></p> <p>Draw and interpret distribution curves for different temperatures using Maxwell–Boltzmann distributions of molecular energies in gases.</p>	<p><b>3.1.5.3 Effect of temperature on reaction rate</b></p> <ul style="list-style-type: none"> <li>Use the Maxwell–Boltzmann distribution to explain why a small temperature increase can lead to a large increase in rate.</li> <li>Investigate how the rate of a reaction changes with temperature <b>(RP)</b>.</li> </ul>	<p><b>3.1.5.4 Effect of concentration and pressure</b></p> <ul style="list-style-type: none"> <li>Explain the qualitative effect of changes in concentration on collision frequency.</li> <li>Explain the qualitative effect of a change in the pressure of a gas on collision frequency.</li> </ul>	<p><b>3.1.5.5 Catalysts</b></p> <p>Use a Maxwell–Boltzmann distribution to help explain how a catalyst increases the rate of a reaction involving a gas.</p>
<b>3.1.6 Chemical equilibria</b>	<p><b>3.1.6.1 Chemical equilibria and Le Chatelier's principle</b></p> <ul style="list-style-type: none"> <li>Use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium.</li> <li>Explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used.</li> </ul>		<p><b>3.1.6.2 Equilibrium constant <math>K_c</math> for homogeneous systems</b></p> <ul style="list-style-type: none"> <li>Construct an expression for <math>K_c</math> for a homogeneous system in equilibrium.</li> <li>Calculate a value for <math>K_c</math> from the equilibrium concentrations for a homogeneous system at constant temperature.</li> <li>Perform calculations involving <math>K_c</math>.</li> <li>Predict the qualitative effects of changes of temperature on the value of <math>K_c</math>.</li> </ul>		

## AS Level Chemistry - HEAD

3.2 Inorganic Chemistry	3.1.7 Redox	<b>3.1.7 Oxidation, reduction and redox equations</b> <ul style="list-style-type: none"> <li>Describe the processes of oxidation and reduction in terms of electrons.</li> <li>Recall the rules for assigning oxidation states</li> <li>Work out the oxidation state of an element in a compound or ion from the formula</li> <li>Write half-equations identifying the oxidation and reduction processes in redox reactions</li> <li>Combine half-equations to give an overall redox equation.</li> </ul>	
	C.2.1 Periodicity	<b>3.2.1.1 Classification</b> Recall that an element is classified as s, p, d or f block according to its position in the Periodic Table, which is determined by its proton number.	<b>3.2.1.2 Physical properties of Period 3 elements</b> <ul style="list-style-type: none"> <li>Explain the trends in atomic radius and first ionisation energy.</li> <li>Explain the melting point of the elements in terms of their structure and bonding.</li> </ul>
	3.2.2 Group 2, alkaline earth metals	<b>3.2.2 Group 2, the alkaline earth metals</b> <ul style="list-style-type: none"> <li>Explain the trends in atomic radius and first ionisation energy.</li> <li>Explain the melting point of the elements in terms of their structure and bonding.</li> <li>Explain why BaCl<sub>2</sub> solution is used to test for sulfate ions and why it is acidified.</li> <li>Describe the reactions of the elements Mg–Ba with water.</li> <li>Describe the use of magnesium in the extraction of titanium from TiCl<sub>4</sub>.</li> <li>Describe the relative solubilities of the hydroxides of the elements Mg–Ba in water.</li> <li>Give examples of the use of Mg(OH)<sub>2</sub> in medicine and of Ca(OH)<sub>2</sub> in agriculture.</li> <li>Explain the use of CaO or CaCO<sub>3</sub> to remove SO<sub>2</sub> from flue gases.</li> <li>Describe the relative solubilities of the sulfates of the elements Mg–Ba in water.</li> <li>Describe the use of acidified BaCl<sub>2</sub> solution to test for sulfate ions.</li> <li>Give examples of the use of BaSO<sub>4</sub> in medicine.</li> </ul>	
	3.2.3 Group 7(17), the halogens	<b>3.2.3.1 Trends in halogen properties</b> <ul style="list-style-type: none"> <li>Explain the trend in electronegativity.</li> <li>Explain the trend in the boiling point in terms of structure and bonding.</li> <li>Explain the trend in oxidising ability of the halogens down the group, including displacement reactions of halide ions in aqueous solution.</li> <li>Explain the trend in reducing ability of the halide ions, including the reactions of solid sodium halides with concentrated sulfuric acid.</li> </ul>	<b>3.2.3.2 Uses of chlorine and chlorate(I)</b> <ul style="list-style-type: none"> <li>Describe the reaction of chlorine with water.</li> <li>Appreciate that society assesses the advantages and disadvantages when deciding if chemicals should be added to water supplies.</li> <li>Explain the use of chlorine in water treatment.</li> <li>Appreciate that the benefits to health of water treatment by chlorine outweigh its toxic effects.</li> <li>Describe the reaction of chlorine with cold, dilute, aqueous NaOH and uses of the solution formed. Carry out simple test-tube reactions to identify:                             <ul style="list-style-type: none"> <li>- cations – Group 2, NH<sub>4</sub><sup>+</sup></li> <li>- anions – Group 7 (halide ions), OH<sup>-</sup>, CO<sub>3</sub><sup>2-</sup>, SO<sub>4</sub><sup>2-</sup> (<b>RP</b>)</li> </ul> </li> </ul>

## AS Level Chemistry - HEAD

3.3 Organic Chemistry	3.3.1 Introduction to Organic Chemistry	<b>3.3.1.1 Nomenclature</b> <ul style="list-style-type: none"> <li>Draw structural, displayed and skeletal formulas for given organic compounds.</li> <li>Apply IUPAC rules for nomenclature to name organic compounds limited to chains and rings with up to six carbon atoms each.</li> <li>Apply IUPAC rules for nomenclature to draw the structure of an organic compound from the IUPAC name limited to chains and rings with up to six carbon atoms each.</li> </ul>		<b>3.3.1.2 Reaction mechanisms</b> <ul style="list-style-type: none"> <li>Construct balanced equations for the steps in a free-radical mechanism.</li> <li>Outline mechanisms by drawing the structures of the species involved and curly arrows to represent the movement of electron pairs.</li> </ul>		<b>3.3.1.3 Isomerism define the term structural isomer</b> <ul style="list-style-type: none"> <li>Draw the structures of chain, position and functional group isomers.</li> <li>Define the term stereoisomer.</li> <li>Draw the structural formulas of E and Z isomers.</li> <li>Apply the CIP priority rules to E and Z isomers.</li> </ul>
	3.3.2 Alkanes	<b>3.3.2.1 Fractional distillation of crude oil</b> <ul style="list-style-type: none"> <li>Define the term alkane.</li> <li>Explain the process of fractional distillation as used to separate petroleum.</li> </ul>	<b>3.3.2.2 Modification of alkanes by cracking</b> <ul style="list-style-type: none"> <li>Describe and explain the conditions involved in the cracking of alkanes.</li> <li>Explain the economic reasons for cracking alkanes.</li> </ul>	<b>3.3.2.3 Combustion of alkanes</b> <ul style="list-style-type: none"> <li>Describe the complete and incomplete combustion of alkanes.</li> <li>Explain the formation and consequences of the following pollutants: NO<sub>x</sub>, CO, carbon and unburned hydrocarbons.</li> <li>Explain why sulfur dioxide can be removed from flue gases using calcium oxide or calcium carbonate.</li> </ul>		<b>3.3.2.4 Chlorination of alkanes</b> Explain the reaction of methane with chlorine as a free-radical substitution mechanism involving initiation, propagation and termination steps.
	3.3.3 Halogenoalkanes	<b>3.3.3.1 Nucleophilic substitution</b> Outline the nucleophilic substitution mechanisms of the reactions of halogenoalkanes with the nucleophiles OH <sup>-</sup> , CN <sup>-</sup> and NH <sub>3</sub> .	<b>3.3.3.2 Elimination</b> <ul style="list-style-type: none"> <li>Describe the concurrent substitution and elimination reactions of a halogenoalkane (e.g. 2-bromopropane with potassium hydroxide).</li> <li>Explain the role of the reagent as both nucleophile and base.</li> <li>Outline the mechanisms of these reactions.</li> </ul>		<b>3.3.3.3 Ozone depletion</b> <ul style="list-style-type: none"> <li>Use equations, such as the following, to explain how chlorine atoms catalyse decomposition of ozone:  <math>Cl^{\bullet} + O_3 \rightarrow ClO^{\bullet} + O_2</math> and <math>ClO^{\bullet} + O_3 \rightarrow 2O_2 + Cl^{\bullet}</math> </li> <li>Appreciate that results of research by different groups in the scientific community provided evidence for legislation to ban the use of CFCs as solvents and refrigerants. Chemists have now developed alternative chlorine-free compounds.</li> </ul>	

## AS Level Chemistry - HEAD

<b>3.3.4 Alkenes</b>	<b>3.3.4.1 Structure, bonding and reactivity</b> <ul style="list-style-type: none"> <li>Describe the bonding in alkenes.</li> <li>Explain how alkenes react and bond.</li> </ul>	<b>3.3.4.2 Addition reactions of alkenes</b> <ul style="list-style-type: none"> <li>Outline the mechanisms for the electrophilic addition reactions of alkenes with HBr, H<sub>2</sub>SO<sub>4</sub> and Br<sub>2</sub> the use of bromine to test for unsaturation.</li> <li>Explain the formation of major and minor products by reference to the relative stabilities of primary, secondary and tertiary carbocation intermediates</li> </ul>	<b>3.3.4.3 Addition polymers</b> <ul style="list-style-type: none"> <li>Draw the repeating unit from a monomer structure.</li> <li>Draw the repeating unit from a section of the polymer chain.</li> <li>Draw the structure of the monomer from a section of the polymer.</li> <li>Explain why addition polymers are unreactive.</li> <li>Explain the nature of intermolecular forces between molecules of polyalkenes.</li> </ul>
<b>3.3.5 Alcohols</b>	<b>3.3.5.1 Alcohol production</b> <ul style="list-style-type: none"> <li>Explain the meaning of '<i>biofuel</i>'.</li> <li>Justify the conditions used in the production of ethanol by fermentation of glucose.</li> <li>Write equations to support the statement that ethanol produced by fermentation is a carbon neutral fuel and give reasons why this statement is not valid.</li> <li>Outline the mechanism for the formation of an alcohol by the reaction of an alkene with steam in the presence of an acid catalyst.</li> <li>Discuss the environmental (including ethical) issues linked to decision making about biofuel use.</li> </ul>	<b>3.3.5.2 Oxidation of alcohols</b> <ul style="list-style-type: none"> <li>Classify alcohols as primary, secondary or tertiary.</li> <li>Describe the different oxidation reactions of alcohols.</li> <li>Write equations for these oxidation reactions.</li> <li>Explain how the method used to oxidise a primary alcohol determines whether an aldehyde or carboxylic acid is obtained.</li> <li>Use chemical tests to distinguish between aldehydes and ketones including Fehling's solution and Tollens' reagent.</li> </ul>	<b>3.3.5.3 Elimination</b> Outline the mechanism for the elimination of water from alcohols.
<b>3.3.6 Organic analysis</b>	<b>3.3.6.1 Identification of functional groups by test-tube reactions</b> <ul style="list-style-type: none"> <li>Identify the functional groups using reactions in the specification.</li> <li>Perform Tests for alcohol, aldehyde, alkene and carboxylic acid. <b>(RP)</b></li> </ul>	<b>3.3.6.2 Mass spectrometry</b> Use precise atomic masses and the precise molecular mass to determine the molecular formula of a compound.	<b>3.3.6.3 Infrared spectroscopy</b> <ul style="list-style-type: none"> <li>Recall that bonds in a molecule absorb infrared radiation at characteristic wavenumbers.</li> <li>Explain the link between absorption of infrared radiation by bonds in CO<sub>2</sub>, methane and water vapour and global warming.</li> <li>Use infrared spectra and the Chemistry Data Sheet or Booklet to identify particular bonds, and therefore functional groups, and also to identify impurities.</li> </ul>