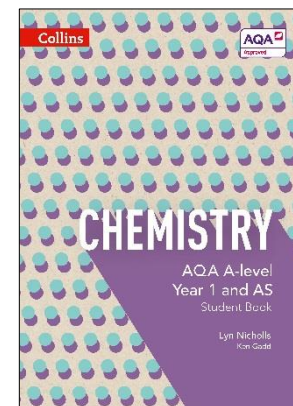


**Collins**

# AQA A-level Chemistry Year 1 and AS Scheme of Work



## Scheme of Work

### AQA A-level Chemistry Year 1 and AS

This course covers the requirements of the first year of the AQA AS and A-level Chemistry specification. These schemes of work are designed to accompany the use of Collins' AQA A-level Chemistry Year 1 and AS Student Book, and references to sections in that book are given for each lesson.

We have assumed that 120 one-hour lessons are taught during the year. Each lesson is matched to the Specification Content. Learning outcomes for each lesson are listed, as are the key Mathematical Skills, Practical Skills, and Apparatus and Techniques Skills that the lesson provides opportunities to practise. It is suggested in which lessons the six Required Practicals may be carried out, to help you plan for these and the sourcing of necessary equipment.

The schemes suggested are of course flexible, and editable, to correspond with your timetabling and to enable you to plan your own route through the course.

#### KEY

The codes in the 'skills covered' column refer to the skills in the AQA specification.

MS - Mathematical Skills

PS - Practical Skills

AT- Apparatus and Techniques Skills

## Scheme of Work

### AQA A-level Chemistry Year 1 and AS (120 hours)

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 1 – Atomic structure (6 hours)					
1 Sub-atomic particles Review of theories and evidence relating to the structure of the atom, from early ideas up to the Bohr model.	<ul style="list-style-type: none"> <li>Describe the main theories and models of the atom up to the Bohr model</li> <li>Explain how ideas about the nature of protons, neutrons and electrons of the atoms have changed over time</li> <li>Link and explain the evidence that led to the development of new theories and models of the atom</li> </ul>	3.1.1.1 Fundamental particles Appreciate that knowledge and understanding of atomic structure has evolved over time.  Protons, neutrons and electrons: relative charge and relative mass.  An atom consists of a nucleus containing protons and neutrons surrounded by electrons.		1.1	
2 Atoms and isotopes Use ideas about mass number and atomic number to explain the existence and significance of isotopes.  Provide examples of importance isotopes.	<ul style="list-style-type: none"> <li>Determine the number of fundamental particles in atoms and ions using <math>A</math>, <math>Z</math> and charge</li> <li>Explain why some isotopes are significant</li> <li>Make predictions about the physical and chemical properties of isotopes</li> </ul>	3.1.1.2 Mass number and isotopes The principles of a simple time of flight (TOF) mass spectrometer, limited to electron spray ionisation, acceleration, ion drift, ion detection, data analysis.  The mass spectrometer gives accurate information about relative isotopic mass and also about the relative abundance of isotopes.	MS 1.1: Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures.  MS 1.2: Students calculate weighted means, e.g. calculation of an atomic mass based on supplied isotopic abundances.	1.1 1.2 1.3 1.4	
3 Mass spectrometry	<ul style="list-style-type: none"> <li>Describe how a time of flight mass spectrometer works</li> </ul>			1.4	

Outline of how a time of flight (TOF) mass spectrometer works and interpret data from the mass spectrum of an element.	<ul style="list-style-type: none"> <li>Predict how different particles might behave in the time of flight mass spectrometer</li> <li>Interpret the mass spectrum for different elements and deduce information about the mass of different isotopes</li> </ul>				
4 Interpreting mass spectra  Interpret data from the mass spectra of elements and molecules and explain their formation.	<ul style="list-style-type: none"> <li>Identify different elements using mass spectra</li> <li>Identify molecular ions from their m/z values from a mass spectra</li> <li>Explain how different molecular ions from a mass spectra are formed</li> </ul>	3.1.1.2 Mass number and isotopes  Mass spectrometry can be used to identify elements.  Mass spectrometry can be used to determine relative molecular mass.  Students should be able to: interpret simple mass spectra of elements; calculate relative atomic mass from isotopic abundance, limited to mononuclear ions.	MS 3.1: Students interpret and analyse spectra.	1.5 1.6	
5 Emission spectra  Introduction to the idea of emission spectra and how they provide evidence for energy shells.  Describe the electronic configuration of elements in terms of s, p and d orbitals	<ul style="list-style-type: none"> <li>Describe the shapes of s and p electron orbitals</li> <li>Write the electronic configuration of atoms and ions up to Z=36 in terms of level and sub levels (orbitals), s, p and d</li> <li>Explain the link between emission spectra and energy levels</li> </ul>	3.1.1.3 Electronic configuration  Electron configurations of atoms and ions up to Z = 36 in terms of levels and sub-levels (orbitals) s, p and d.		1.7	
6 Evidence form ionisation energies  Describe what ionisation energy is.	<ul style="list-style-type: none"> <li>Define first ionisation energy</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</li> </ul>	3.1.1.3 Electronic configuration  Ionisation energies.  Students should be able to: define first ionisation energy; write equations for first and successive		1.8 1.9	

<p>Use trends in first ionisation energy across periods to evidence relating to the existences of energy shells and sub-shells</p>	<p>Group 2 (Be–Ba) give evidence for electron configuration in sub-levels and in levels</p>	<p>ionisation energies; 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.</p>			
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One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 2 – Amount of Substance (16 hours)					
1 Relative Masses  Recap of ideas from chapter 1 relating to atomic and molecular mass and using mass spectra. Use formulae to calculate relative formula or molecular mass	<ul style="list-style-type: none"> <li>Define relative atomic and molecular mass</li> <li>Use mass spectra to calculate atomic and molecular masses</li> <li>Calculate the relative formula and molecular masses of different compounds from their formulae.</li> </ul>	<p>3.1.2.1 Relative atomic mass and relative molecular mass</p> <p>Relative atomic mass and relative molecular mass in terms of <math>^{12}\text{C}</math>.</p> <p>The term relative formula mass will be used for ionic compounds.</p> <p>Students should be able to: define relative atomic mass (<math>A_r</math>); define relative molecular mass (<math>M_r</math>).</p>		2.1	
2 The mole  Recap of the concept of the mole. Practical activity to deduce the value of Avogadro's constant	<ul style="list-style-type: none"> <li>State what Avogadro's constant is and use it to find the numbers of particles in different numbers of moles</li> <li>Describe how Avogadro's constant was derived</li> <li>Explain why the mole is important to chemists</li> </ul>	<p>3.1.2.2 The mole and the Avogadro constant</p> <p>The Avogadro constant as the number of particles in a mole.</p> <p>The mole as applied to electrons, atoms, molecules, ions, formulae and equations.</p> <p>The concentration of a substance in solution measured in <math>\text{mol dm}^{-3}</math>.</p> <p>Students should be able to carry out calculations: using the Avogadro constant; using mass of substance, relative molecular mass and amount in moles; using concentration, volume and amount of substance in a solution.</p>	<p>MS 0.1: Students carry out calculations using numbers in standard and ordinary form, e.g. using the Avogadro constant.</p> <p>MS 0.4: Students carry out calculations using the Avogadro constant.</p> <p>MS 1.1: Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated results can only be reported to the limits of the least accurate measurement.</p>	2.2	
3 Mole Calculations  Recap of calculations involving moles, masses and concentration.	<ul style="list-style-type: none"> <li>Calculate the mass given the number of moles for a range of different compounds</li> <li>Calculate the numbers of moles given the mass in a range of different compounds</li> <li>Calculate concentrations of liquids from masses and moles.</li> </ul>			2.2	
4 Solutions				2.7	

<p>Describe how to make up a standard solution. Make up different solutions of known concentration. Become familiar with a range of calculations involving solutions in terms of masses, moles and concentration.</p>	<ul style="list-style-type: none"> <li>Describe how to make a standard solution</li> <li>Explain how to identify the concentration of an unknown solutions by comparing the colour against solutions of known concentration</li> <li>Carry out different types of calculations involving solutions.</li> </ul>	<p>Students will not be expected to recall the value of the Avogadro constant.</p>			
<p>5 Ideal gas equation</p> <p>Introduction to the idea of an ideal gas Relate the ideal gas equation to experiments by Charles and Boyle.</p>	<ul style="list-style-type: none"> <li>Describe the differences between an ideal gas and a real gas</li> <li>Explain the effects of temperature and pressure on a gas</li> <li>Carry out calculations using the ideal gas equation</li> </ul>	<p>3.1.2.3 The ideal gas equation</p> <p>The ideal gas equation <math>pV = nRT</math> with the variables in SI units.</p> <p>Students should be able to use the equation in calculations.</p> <p>Students will not be expected to recall the value of the gas constant, <math>R</math>.</p>	<p>AT a, b and k</p> <p>PS 3.2: Students could be asked to find the <math>M_r</math> of a volatile liquid.</p> <p>MS 0.0: Students understand that the correct units need to be in <math>pV = nRT</math></p> <p>MS 2.2, 2.3 and 2.4: Students carry out calculations with the ideal gas equation, including rearranging the ideal gas equation to find unknown quantities.</p>	<p>2.3</p>	
<p>6 Relative molecular mass of a volatile liquid</p> <p>Use an experimental method to make measurements to calculate the <math>M_r</math> of hexane.</p>	<ul style="list-style-type: none"> <li>Describe an experiment to determine the molecular mass of a volatile liquid</li> <li>Make precise measurements and use sensitive equipment like gas syringes with care.</li> <li>Use experimental data and the ideal gas equation to calculate the <math>M_r</math> of a volatile liquid</li> </ul>			<p>2.3</p>	
<p>7 Calculating empirical formula</p>		<p>3.1.2.4 Empirical and molecular formula</p>	<p>AT a and k</p>	<p>2.4</p>	

Recap of the definition of empirical formula Calculation of empirical formula using an experimental methods	<ul style="list-style-type: none"> <li>Explain the difference between empirical formula and molecular formula</li> <li>Describe how to find the empirical mass of a compound by direct reaction and by continuous variation.</li> <li>Use empirical data to calculate the empirical formula.</li> </ul>	<p>Empirical formula is the simplest whole number ratio of atoms of each element in a compound.</p> <p>Molecular formula is the actual number of atoms of each element in a compound.</p> <p>The relationship between empirical formula and molecular formula.</p>	PS 2.3 and 3.3: Students could be asked to find the empirical formula of a metal oxide.		
8 Determine the formulae of compounds  Introduction to techniques to calculate molecular formula from empirical formula, molecular mass, data from CHNS analysers and high level mass spectrometry	<ul style="list-style-type: none"> <li>Calculate molecular formula from the empirical formula and relative molecular mass</li> <li>Describe how CHNS elemental analysers and high level resolution mass spectrometry are used to find the molecular mass</li> <li>Calculate molecular formula from data from CHNS and high level mass spectrometry.</li> </ul>	Students should be able to: empirical formula from data giving composition by mass or data giving percentage by mass; calculate molecular formula from the empirical formula and relative molecular mass.		2.4	
9 Balanced chemical equations  Recap of features of balanced symbol equations. How to write more complicated balanced symbol and ionic equations.	<ul style="list-style-type: none"> <li>Describe how to use titrations to determine the stoichiometry of neutralisation reactions</li> <li>Write balanced equations for a range of different reactions given the formula of reactants and products</li> <li>Write balanced symbol and ionic equations from word equations</li> </ul>	<p>3.1.2.5 Balanced equations and associated calculations</p> <p>Equations (full and ionic). Percentage atom economy is:</p> $\frac{\text{molecular mass of desired product}}{\text{sum of molecular masses of all reactants}} \times 100$ <p>Economic, ethical and environmental advantages for society and for industry of developing chemical processes with a high atom economy.</p>	AT a, d, e, f and k  PS 4.1: Students could be asked to find: the concentration of ethanoic acid in vinegar; the mass of calcium carbonate in an indigestion tablet; the $M_r$ of $\text{MgCO}_3$ ; the $M_r$ of succinic acid; the mass of aspirin in an aspirin tablet; the yield for the conversion of magnesium to magnesium oxide; the $M_r$ of a hydrated salt (e.g. magnesium	2.5	
10 Yields from reactions	<ul style="list-style-type: none"> <li>Use empirical data to calculate percentage yield</li> </ul>	Students should be able to use balanced equations to calculate:		2.5 2.6	



Recap percentage yield and atom economy. Carry out a number of calculations of percentage yield from empirical and given data.	<ul style="list-style-type: none"> <li>Calculate the theoretical and percentage yield from balanced symbol equations.</li> <li>Calculate percentage atom economy from balanced symbol equations</li> </ul>
11 Thermal decomposition Carry out the thermal decomposition of calcium carbonate, hydrated copper (II) sulfate and calculate percentage yields and percentage atom economies.	<ul style="list-style-type: none"> <li>Identify and describe what is meant by a thermal decomposition reactions</li> <li>Calculate percentage yields from thermal decomposition reactions</li> <li>Calculate the formula of hydrates salts from thermal decomposition reactions.</li> </ul>
12 Titrations  Learn how to prepare a standard solution and carry out a titration accurately. Use a titration to calculate concentration.	<ul style="list-style-type: none"> <li>Use equipment with care and precision to prepare a standard solution</li> <li>Carry out a titration to produce accurate and reliable data, calculating the percentage measurement error</li> <li>Use a titration to calculate the concentration of a solution</li> </ul>
13 Ethanoic acid in vinegar Use a titration to determine the percentage composition of	<ul style="list-style-type: none"> <li>Describe how to determine the concentration of ethanoic acid in vinegar</li> <li>Carry out a titration to produce accurate and reliable data,</li> </ul>

masses; volumes of gases; percentage yields; percentage atom economies; concentrations and volumes for reactions in solutions.

sulfate) by heating to constant mass.

AT a and k: Students could be asked to find the percentage conversion of a Group 2 carbonate to its oxide by heat.

AT d, e, f and k: Students could be asked to determine the number of moles of water of crystallisation in a hydrated salt by titration.

MS 0.2: Students construct and/or balance equations using ratios. Students calculate percentage yields and atom economies of reactions.

2.5	
2.7	Required Practical 1: Make up a volumetric solution and carry out a simple acid-base titration.
2.7	

ethanoic acid in vinegar	<p>calculating the percentage measurement error</p> <ul style="list-style-type: none"> <li>Use data from the titration to calculate the percentage concentration of ethanoic acid in vinegar</li> </ul>		
<p>14 Relative formula of a hydrated salt</p> <p>Use a titration between hydrated sodium carbonate and hydrochloric acid to determine the relative formula of the hydrated salt</p>	<ul style="list-style-type: none"> <li>Describe how use a titration to determine the formula of hydrated sodium carbonate</li> <li>Carry out the titration to produce accurate and reliable data</li> <li>Use results to calculate the formula of the hydrated sodium carbonate</li> </ul>		2.7
<p>15 Relative formula mass of succinic acid</p> <p>Carry out a titration between succinic acid and sodium hydroxide to determine the molecular formula of succinic acid.</p>	<ul style="list-style-type: none"> <li>Describe how use a titration to calculate the relative formula of succinic acid</li> <li>Carry out the titration to produce accurate and reliable data</li> <li>Use reliable results and the balanced equation to determine the molecular formula of succinic acid</li> </ul>		2.7
<p>16 Analysing medicines</p> <p>Carry out titrations to determine the amount of calcium carbonate in indigestion tablets and/ or aspirin in aspirin tablets.</p>	<ul style="list-style-type: none"> <li>Describe how use a titration to determine the amount of calcium carbonate in indigestion tablets or aspirin in aspirin tablets</li> <li>Carry out the titrations to produce accurate and reliable data</li> <li>Use reliable results and balanced equations to determine the amount of active ingredient in the medicines</li> </ul>		2.7



One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 3 – Bonding (9 hours)					
<p>1 Ionic bonds</p> <p>Recap of what is meant by an ionic bond and use dot and cross diagrams of electronic structures and formulae to represent ionic compounds</p>	<ul style="list-style-type: none"> <li>explain what is meant by “ionic bonding”</li> <li>draw appropriate dot and cross diagrams to show electronic structures of ions in an ionic compound</li> <li>determine the formula of simple and complex ionic compounds</li> </ul>	<p>3.1.3.1 Ionic bonding</p> <p>Ionic bonding involves electrostatic attraction between oppositely charged ions in a lattice. The formulas of compound ions e.g. sulfate, hydroxide, nitrate, carbonate and ammonium.</p> <p>Students should be able to: predict the charge on a simple ion using the position of the element in the Periodic Table; construct formulas for ionic compounds.</p>		3.1	
<p>2 Covalent Bonds</p> <p>Use appropriate conventions to display the bonds in covalent compounds. Introduce the idea of bond polarity and dative covalent/ co-ordinate bonds</p>	<ul style="list-style-type: none"> <li>Draw appropriate dot and cross diagrams to represent covalent bonds compound, including dative covalent bonds.</li> <li>use the idea of electronegativity to explain why electron density in covalent bonds may not be symmetrical</li> <li>Explain a molecule may contain polar bonds, but the molecules itself may or may not be polar.</li> </ul>	<p>3.1.3.2 Nature of covalent and dative covalent bonds</p> <p>A single covalent bond contains a shared pair of electrons.</p> <p>Multiple bonds contain multiple pairs of electrons.</p> <p>A co-ordinate (dative covalent) bond contains a shared pair of electrons with both electrons supplied by one atom.</p> <p>Students should be able to represent: a covalent bond using a line; a co-ordinate bond using an arrow.</p>		3.2	
<p>3 Metals</p> <p>Recap on metallic bonding. Compare properties of substances which are bonded by ionic,</p>	<ul style="list-style-type: none"> <li>identify whether elements or compounds have ionic, covalent or metallic bonding</li> <li>explain the bonding and structure of metals</li> </ul>	<p>3.1.3.3 Metallic bonding</p> <p>Metallic bonding involves attraction between delocalised electrons and positive ions arranged in a lattice.</p>		3.3	

covalent or metallic bonding.	<ul style="list-style-type: none"> <li>use ideas about bonding to explain the properties of metals</li> </ul>				
<p>4 Crystal structure</p> <p>Describe different crystal structures. Carry out a practical to grow a range of different crystal structures.</p>	<ul style="list-style-type: none"> <li>make observations of the growth of a range of different crystals including silver, potassium nitrate and benzoic acid</li> <li>describe the four types of crystal structure: ionic; metallic; macromolecular (giant covalent); molecular</li> <li>draw diagrams to show the crystal structures of diamond, graphite, ice, iodine, magnesium and sodium chloride</li> </ul>	<p>3.4 Bonding and physical properties</p> <p>The four types of crystal structure: ionic; metallic; macromolecular (giant covalent); molecular.</p> <p>The structures of the following crystals as examples of these four types of crystal structure: diamond; graphite; ice; iodine; magnesium; sodium chloride</p> <p>Students should be able to: relate the melting point and conductivity of materials to the type of structure and the bonding present; explain the energy changes associated with changes of state; draw diagrams to represent these structures involving specified numbers of particles.</p>	<p>AT a, b, h and k</p> <p>PS 1.1: Students could be asked to find the type of structure of unknowns by experiment (e.g. to test solubility, conductivity and ease of melting).</p>	3.4	
<p>5 Bonding and physical structures</p> <p>Use a series of experimental tests to determine the solubility, conductivity and ease of melting to determine the crystal structure in a range of unknown substances.</p>	<ul style="list-style-type: none"> <li>Recall examples of substances with ionic, metallic, macromolecular and molecular bonding.</li> <li>relate the melting point and conductivity of materials to the type of structure and the bonding present</li> <li>explain the energy transfers that happen with changes of state</li> </ul>			3.4	
<p>6 Shapes of molecules and ions</p> <p>Introduction to the idea of shapes and bond</p>	<ul style="list-style-type: none"> <li>describe different shapes of simple molecules</li> <li>explain why simple molecules and ions have specific shapes</li> </ul>	3.1.3.5 Shapes of simple molecules and ions	MS 0.3 and 4.1: Students could be given familiar and unfamiliar examples of species and asked to	3.5	

<p>angles of simple molecules, using the VSEPR principle.</p>	<p>using the idea of electron pair repulsion</p> <ul style="list-style-type: none"> <li>work out the shapes and bond angles of simple molecules and ions using the principle of valence shell electron pair repulsion (VSEPR)</li> </ul>	<p>Bonding pairs and lone (non-bonding) pairs of electrons as charge clouds that repel each other.</p> <p>Pairs of electrons in the outer shell of atoms arrange themselves as far apart as possible to minimise repulsion.</p> <p>Lone pair-lone pair repulsion is greater than lone pair-bond pair repulsion, which is greater than bond pair-bond pair repulsion.</p> <p>The effect of electron pair repulsion on bond angles.</p> <p>Students should be able to explain the shapes of, and bond angles in, simple molecules and ions with up to 6 electron pairs (including lone pairs of electrons) surrounding the central atom.</p>	<p>deduce the shape according to valence shell electron pair repulsion (VSEPR) principles.</p> <p>Bond angles for NH<sub>3</sub>, CH<sub>4</sub> and H<sub>2</sub>O should be known and other bond angles could be estimated.</p>		
<p>7 Polar bonds and intermolecular forces</p> <p>Explain the different types of intermolecular force.</p> <p>Use the angle of deflection of different liquid streams with an electrically-charged plastic ruler to investigate bond polarity.</p>	<ul style="list-style-type: none"> <li>describe the different types of intermolecular bonds (permanent dipole–dipole forces, induced dipole–dipole (van der Waals’) forces and hydrogen bonding)</li> <li>explain the type of intermolecular bonds between molecules of familiar and unfamiliar compounds using ideas about electronegativity</li> <li>explain how melting and boiling points are influenced by these intermolecular forces</li> </ul>	<p>3.1.3.6 Bond polarity</p> <p>Electronegativity as the power of an atom to attract the pair of electrons in a covalent bond.</p> <p>The electron distribution in a covalent bond between elements with different electronegativities will be unsymmetrical. This produces a polar covalent bond and may cause a molecule to have a permanent dipole.</p> <p>Students should be able to: use partial charges to show that a bond is polar; explain why some molecules with polar bonds do not have a permanent dipole.</p>		<p>3.6 3.7</p>	

<p>8 Hydrogen bonding</p> <p>Use ideas about melting, boiling point and density to explain the formation of hydrogen bonds</p>	<p>Students should be able to</p> <ul style="list-style-type: none"> <li>• give examples of molecules which exhibit hydrogen bonding</li> <li>• explain the formation of hydrogen bonds</li> <li>• explain why hydrogen bonds are the strongest of the intermolecular bonds</li> </ul>	<p>3.1.3.7 Forces between molecules</p> <p>Forces between molecules: permanent dipole–dipole forces; induced dipole–dipole (van der Waals’, dispersion and London) forces; hydrogen bonding.</p> <p>The melting and boiling points of molecular substances are influenced by the strength of these intermolecular forces. The importance of hydrogen bonding in the low density of ice and the anomalous boiling points of compounds.</p>	<p>AT d and k</p> <p>PS 1.2: Students could try to deflect jets of various liquids from burettes to investigate the presence of different types and relative size of intermolecular forces.</p>	<p>3.7</p>	
<p>9 Revision of ideas about bonding</p> <p>Groups work together to produce a large display about bonding. The features of each type of bonding with suitable examples, without external resources if possible. 3-D models should be made from everyday materials to accompany the display.</p>	<ul style="list-style-type: none"> <li>• give examples of substances which exhibit each type of bonding discussed</li> <li>• Explain in detail how each type of bonding works and which types of elements are involved</li> <li>• Compare and contrast the properties of substance which exhibit different types of bonding.</li> </ul>	<p>Students should be able to: explain the existence of these forces between familiar and unfamiliar molecules; explain how melting and boiling points are influenced by these intermolecular forces.</p>		<p>3.1 to 3.7</p>	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 4 –The Periodic Table (3 hours)					
1 Classifying elements  Recap ideas about ionisation energy and energy shells. Use data about first ionisation to make predictions about the position of elements on the Periodic Table	<ul style="list-style-type: none"> <li>describe how the Periodic Table is organised</li> <li>classify an element as s, p, d or f block according to its position in the Periodic Table</li> <li>Use data on first ionisation energies and proton number to predict the position of elements on the periodic table</li> </ul>	3.2.1.1 Classification  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		4.1	
2 Trends in physical properties  Use ideas about nuclear charge, distance from the nucleus, electron shielding and paired or unpaired electrons to explain the trends in atomic radius and first ionisation energies in Period 3. Explain how atomic radius changes down group 2.	<ul style="list-style-type: none"> <li>Explain what is meant by the atomic radius</li> <li>explain the trends in atomic radius, first ionisation energy and electrical conductivity of the elements Na–Ar</li> <li>explain the melting point of the elements in terms of their structure and bonding</li> </ul>	3.2.1.2 Physical properties of Period 3 elements  The trends in atomic radius, first ionisation energy and melting point of the elements Na–Ar.  The reasons for these trends in terms of the structure of and bonding in the elements.		4.2	
3 Reactions of Period 3 elements  Carry out investigations/ demonstrations of the reactions of group 3	<ul style="list-style-type: none"> <li>describe the reactions of Na and Mg with water</li> </ul>			4.3	



elements with water, oxygen and chlorine	<ul style="list-style-type: none"><li>• describe the reactions of Period 3 elements with oxygen and with chlorine</li><li>• Make predictions about how electronegativity is linked to the bonding in oxides and chlorides of group 3 elements</li></ul>				
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One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 5 – Introduction to organic chemistry (9 hours)					
1 Organic and inorganic  Introduce the idea of organic compounds Explain why C has its own branch of chemistry using ideas about valency and mean bond enthalpy.	<ul style="list-style-type: none"> <li>Identify some organic compounds</li> <li>Explain the key differences between organic compounds and inorganic compounds</li> <li>explain why carbon is able to form a large number of different compounds</li> </ul>	3.3.1.1 Nomenclature  Organic compounds can be represented by: empirical formula; molecular formula; general formula; structural formula; displayed formula; skeletal formula.  The characteristics of a homologous series, a series of compounds containing the same functional group.		5.1 5.2	
2 Types of formulae  Introduce the rules and conventions for drawing different types of formulae and explain when they are appropriate to use.	<ul style="list-style-type: none"> <li>use empirical formulae, molecular formulae, general formulae, structural formulae, displayed formulae and skeletal formulae to represent organic compounds</li> <li>Explain the differences between the different types of formulae</li> <li>Convert from one type of formulae to another</li> </ul>	IUPAC rules for nomenclature.  Students should be able to: draw structural, displayed and skeletal formulas for given organic compounds; apply IUPAC rules for nomenclature to name organic compounds limited to chains and rings with up to six carbon atoms each;		5.3	
3 Molecular models and drawing structures  Use 3-D models to represent simple organic molecules and use these to draw 3-D representations of these molecules.	<ul style="list-style-type: none"> <li>use molecular model kits to build models of organic molecules</li> <li>recognise the three dimensional aspect of molecules</li> <li>use appropriate notation to represent the three dimensional aspect of molecules</li> </ul>	up to six carbon atoms each; 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.		5.2	
4 Functional groups				5.4	

<p>Describe the features of a homologous series using data on boiling points. Introduce the following functional groups:- alkanes, alkenes, halogenoalkanes, alcohols, carboxylic acids, aldehydes, ketones and esters</p>	<ul style="list-style-type: none"> <li>• describe what is meant by a homologous series and a functional group</li> <li>• recognise and name the functional groups for alkanes, alkenes, halogenoalkanes, alcohols, carboxylic acids, aldehydes, ketones and esters</li> <li>• Identify functional groups from displayed and molecular formulae.</li> </ul>				
<p>5 Identifying functional groups by chemical reactions</p> <p>Use practical procedures to carry out a number of tests to identify alkenes, halogenoalkanes, alcohols, aldehydes and carboxylic acids.</p>	<ul style="list-style-type: none"> <li>• Follow procedures safely</li> <li>• Describe the tests for alkenes, halogenoalkanes, alcohols, aldehydes and carboxylic acids</li> <li>• Use information from chemical tests to make inferences about functional group(s) present</li> <li>•</li> </ul>				
<p>6 Naming organic compounds</p> <p>Introduce IUPAC rules for nomenclature for naming alkanes, alkenes, halogenoalkanes, alcohols, carboxylic acids, aldehydes and ketones</p>	<ul style="list-style-type: none"> <li>• Explain the importance of having a common international system for naming organic compounds</li> <li>• apply IUPAC rules for nomenclature to name organic compounds limited to chains and rings with up to six carbon atoms</li> <li>• apply IUPAC rules for nomenclature to draw the structure of organic</li> </ul>			5.5	

	compounds limited to chains and rings with up to six carbon atoms				
7 Introduction to reaction mechanisms  Introduce the stages of the free radical mechanism. Explain the rules for using curly arrows in reaction mechanisms.	<ul style="list-style-type: none"> <li>• explain free radical mechanisms</li> <li>• write balanced equations for the steps in a free radical mechanism</li> <li>• use curly arrows to show the movement of electrons during a chemical reaction</li> </ul>			5.6	
8 Isomers  Introduce what an isomer is by showing different examples. Explain the key differences between structural and stereoisomers.	<ul style="list-style-type: none"> <li>• explain the differences between structural and stereoisomers</li> <li>• Draw the structures of chain, position, functional group and E and Z isomers</li> <li>• Apply CIP priority rules to E and Z isomers</li> </ul>			5.7	
9 Melting points and boiling points of isomers  Practical investigation of the melting points of propan-1-ol and propan-2-ol and use of data books to analyse the melting and boiling points of different isomers. These are related to the structure of the isomers to explain the differences.	<ul style="list-style-type: none"> <li>• carry out a practical investigation safely</li> <li>• measure the boiling point of an organic compound</li> <li>• explain why isomers have different boiling and melting points</li> </ul>			5.7	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 6 – The alkanes (8 hours)					
<p>1 Fractional distillation of crude oil</p> <p>Practical investigation of fractional distillation to recap the process. Revision of different fractions and their uses</p>	<ul style="list-style-type: none"> <li>Explain why alkanes are saturated hydrocarbons</li> <li>Describe how alkanes in petroleum can be separated by fractional distillation</li> <li>Explain the uses of different fractions</li> </ul>	<p>3.3.2.1 Fractional distillation of crude oil</p> <p>Alkanes are saturated hydrocarbons. Petroleum is a mixture consisting mainly of alkane hydrocarbons that can be separated by fractional distillation.</p>	<p>AT a, d and k</p> <p>PS 1.2: Fractional distillation of a crude oil substitute.</p>	6.1	
<p>2 Cracking processes</p> <p>Recap what happens during cracking. Explain the differences between thermal and catalytic cracking.</p>	<ul style="list-style-type: none"> <li>explain the economic reasons for cracking</li> <li>explain the differences between thermal cracking and catalytic cracking</li> <li>use chemical equations to describe cracking</li> </ul>	<p>3.3.2.2 Modification of alkanes by cracking</p> <p>Cracking involves the breaking of C–C bonds in alkanes.</p> <p>Thermal cracking takes place at high pressure and high temperature and produces a high percentage of alkenes (mechanism not required).</p>		6.2	
<p>3 Catalytic cracking</p> <p>Carry out an investigation of the catalytic cracking of paraffin.</p>	<ul style="list-style-type: none"> <li>follow and experimental procedure safely to crack liquid paraffin</li> <li>describe the differences between batch and continuous processes</li> <li>explain why zeolites are used as the catalysts</li> <li></li> </ul>	<p>Catalytic cracking takes place at a slight pressure, high temperature and in the presence of a zeolite catalyst and is used mainly to produce motor fuels and aromatic hydrocarbons (mechanism not required).</p> <p>Students should be able to explain the economic reasons for the cracking of alkanes.</p>		6.2	

<p>4 Combustion of alkanes</p> <p>Use ideas about energy diagrams and enthalpy of combustion to explain why alkanes are useful for complete combustion. Recap the differences between complete and incomplete combustion</p>	<ul style="list-style-type: none"> <li>• explain why alkanes are used as fuels</li> <li>• describe the differences between complete and incomplete combustion</li> <li>• write balanced chemical equations for complete and incomplete combustion reactions</li> </ul>	<p>3.3.2.3 Combustion of alkanes</p> <p>Alkanes are used as fuels.</p> <p>Combustion can be complete or incomplete.</p> <p>The internal combustion engine produces a number of pollutants including NO<sub>x</sub>, CO, carbon and unburned hydrocarbons.</p> <p>These gaseous pollutants from internal combustion engines can be removed using catalytic converters.</p>		6.3	
<p>5 Measuring energy when alkanes burn</p> <p>Use calorimetry to measure the energy changes during the combustion of n hexane, heptane and octane and calculate the enthalpy change for each.</p>	<ul style="list-style-type: none"> <li>• describe how to carry out a practical investigation to measure the energy released when alkanes burn</li> <li>• calculate the enthalpy change from experimental data</li> <li>• calculate the measurement uncertainties and identify the systematic errors</li> </ul>	<p>Combustion of hydrocarbons containing sulfur leads to sulfur dioxide that causes air pollution.</p> <p>Students should be able to explain why sulfur dioxide can be removed from flue gases using calcium oxide or calcium carbonate.</p>		6.3	
<p>6 Problems with alkane combustion</p> <p>Explore the different products formed by the internal combustion engine, the problems they cause and ways to limit their impact.</p>	<ul style="list-style-type: none"> <li>• Name the pollutants produced by the internal combustion engine and describe the problems they cause.</li> <li>• explain how catalytic converters remove the pollutants from the internal combustion engine</li> <li>• explain why sulfur dioxide can be removed from flue gases using calcium oxide or calcium carbonate</li> </ul>			6.4	
<p>7 Chlorination of alkanes</p>		<p>3.3.2.4 Chlorination of alkanes</p>		6.5	

<p>Explain the stages of the free radical mechanism during homolytic fission of a chloroalkane.</p>	<ul style="list-style-type: none"> <li>• write equations to show the free radical substitution of methane with chlorine</li> <li>• identify the initiation, propagation and termination steps in the free radical mechanism</li> <li>• Explain why ultraviolet light can be used to break the C-Cl bond using <math>E = h\nu</math></li> </ul>	<p>The reaction of methane with chlorine.</p> <p>Students should be able to explain this reaction as a free-radical substitution mechanism involving initiation, propagation and termination steps.</p>			
<p>8 Debating the uses of alkanes as fuels</p> <p>Work in groups to prepare a debate sources to explain whether or alkanes should remain our primary source fuel, assuming that supplies do not run out. Evidence from chemical reactions and other sources must be used. Research can be carried out to find further evidence. Groups should prepare arguments for both sides of the debate.</p>	<ul style="list-style-type: none"> <li>• Explain the importance of alkanes as fuels and how they are derived</li> <li>• Explain all the problems surrounding their use and the impact these have on humans and other organisms</li> <li>• Use evidence well in conjunction with reasoned arguments</li> </ul>			6.1 – 6.5	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 7 – Energetics (9 hours)					
<p>1 Enthalpy change</p> <p>Recap energy level diagrams and introduce the idea of enthalpy. Define standard enthalpies of combustion and formation and use them to calculate relative molecular mass.</p>	<ul style="list-style-type: none"> <li>describe reactions as endothermic or exothermic</li> <li>explain that enthalpy change (<math>\Delta H</math>) is the energy transferred by heating or cooling, measured under conditions of constant pressure</li> <li>define standard enthalpy of combustion (<math>\Delta_c H^\ominus</math>) and standard enthalpy of formation (<math>\Delta_f H^\ominus</math>)</li> </ul>	<p>3.1.4.1 Enthalpy change</p> <p>Reactions can be endothermic or exothermic.</p> <p>Enthalpy change (<math>\Delta H</math>) is the heat energy change measured under conditions of constant pressure.</p> <p>Standard enthalpy changes refer to standard conditions, i.e. 100 kPa and a stated temperature (e.g. <math>\Delta H_{298}^\ominus</math>).</p> <p>Students should be able to define: standard enthalpy of combustion (<math>\Delta_c H^\ominus</math>); standard enthalpy of formation (<math>\Delta_f H^\ominus</math>).</p>		<p>7.1</p> <p>7.2</p> <p>7.3</p> <p>7.4</p>	
<p>2 Calorimetry</p> <p>Recap ideas of finding the energy of burning fuels. Use calorimetry to investigate and evaluate the enthalpy of combustion of methanol.</p>	<ul style="list-style-type: none"> <li>Describe how calorimetry can be used to measure the <math>\Delta_c H^\ominus</math> of methanol</li> <li>use <math>q = mc\Delta T</math> to calculate the molar enthalpy change for a reaction using data collected by calorimetry</li> <li>Evaluate calorimetric methods and make suggestions for improvement</li> </ul>	<p>3.1.4.2 Calorimetry</p> <p>The heat change, <math>q</math>, in a reaction is given by the equation <math>q = mc\Delta T</math>.</p> <p>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>.</p>	<p>MS 0.0 and 1.1: Students understand that the correct units need to be used in <math>q = mc\Delta T</math>.</p> <p>Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures.</p>	<p>7.5</p>	<p>Required practical 2: Measurement of an enthalpy change.</p>
<p>3 Enthalpy of solution</p> <p>Carry out experiments to measure the enthalpy of</p>	<ul style="list-style-type: none"> <li>describe dissolution as an endothermic or exothermic process</li> </ul>	<p>Students should be able to: use this equation to calculate the molar enthalpy change for a</p>		<p>7.5</p>	<p>Required practical activity 2: Measurement of an enthalpy change.</p>



<p>solution of a range of different solids</p>	<ul style="list-style-type: none"> <li>explain how to and carry out experiments to determine enthalpy change of solution</li> <li>use <math>q = mc\Delta T</math> to calculate the enthalpy change of solution</li> </ul>	<p>reaction; use this equation in related calculations.</p> <p>Students will not be expected to recall the value of the specific heat capacity, <math>c</math>, of a substance.</p>	<p>Students understand that calculated results can only be reported to the limits of the least accurate measurement.</p>		<p>Student Book 7.5 page 152</p>
<p>4 Enthalpy of neutralisation</p> <p>Calculate expected temperature changes from given neutralisation reactions and compare predictions to empirical values. Compare temperature changes obtained in different reaction vessels.</p>	<ul style="list-style-type: none"> <li>explain the term enthalpy of neutralisation</li> <li>Explain how to measure the enthalpy of neutralisation</li> <li>Predict temperature changes from the equation <math>q = mc\Delta T</math></li> </ul>		<p>AT a and k</p> <p>PS 2.4, 3.1, 3.2, 3.3 and 4.1: Students could be asked to find <math>\Delta H</math> for a reaction by calorimetry. Examples of reactions could include: dissolution of potassium chloride; dissolution of sodium carbonate; neutralising NaOH with HCl; displacement reaction between <math>\text{CuSO}_4 + \text{Zn}</math>; combustion of alcohols.</p>	<p>7.5</p>	
<p>5 Enthalpy of combustion</p> <p>Introduce the bomb calorimeter and compare values for enthalpy of combustion obtained using simple calorimetry with a bomb calorimeter. Explain the differences in findings.</p>	<ul style="list-style-type: none"> <li>define calculate standard enthalpies of combustion using the equation <math>q = mc\Delta T</math></li> <li>describe a bomb calorimeter</li> <li>explain how a bomb calorimeter is different to simpler calorimeters that are used to determine enthalpies of combustion</li> </ul>			<p>7.5</p>	
<p>6 Hess's law</p> <p>Introduce Hess's law and recap the definitions of enthalpy of formation,</p>	<ul style="list-style-type: none"> <li>Explain Hess's law</li> <li>Identify enthalpies of combustion and formation Hess's cycles</li> </ul>	<p>3.1.4.3 Applications of Hess's Law</p> <p>Hess's Law.</p>	<p>MS 2.4: Students carry out Hess's law calculations.</p>	<p>7.6</p>	

combustion and reaction. Use Hess's law to calculate values of any one of these enthalpies	<ul style="list-style-type: none"> <li>use Hess's Law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation</li> <li></li> </ul>	Students should be able to use Hess's Law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation.	AT a and k PS 2.4, 3.2 and 4.1: Students could be asked to find $\Delta H$ for a reaction using Hess's law and calorimetry, then present data in appropriate ways. Examples of reactions could include: thermal decomposition of $\text{NaHCO}_3$ ; hydration of $\text{MgSO}_4$ ; hydration of $\text{CuSO}_4$		
7 Hydration of a metal salt  Determine the enthalpy of solution for anhydrous and hydrated magnesium sulfate experimentally and use these values and Hess's Law to calculate the enthalpy of reaction to form the hydrated salt.	<ul style="list-style-type: none"> <li>collect experimental data to calculate enthalpy changes</li> <li>use Hess's law to calculate enthalpy changes for reactions that cannot be determined directly</li> <li>Explain how Hess's law can be used to calculate enthalpy changes for these reactions</li> </ul>			7.6	
8 A Thermal decomposition reaction  Work out how to measure the enthalpy change for the thermal decomposition of sodium hydrogen carbonate using enthalpy of neutralisation. Plan and carry out the investigation.	<ul style="list-style-type: none"> <li>collect experimental data to calculate enthalpy changes</li> <li>work out how Hess's law to calculate enthalpy changes for the decomposition of sodium hydrogen carbonate</li> <li>explain and plan an investigation to use Hess's law to calculate enthalpy changes for this reaction</li> </ul>			7.6	
9 Bond enthalpies  Recap energy level diagrams and ideas of	<ul style="list-style-type: none"> <li>define the term mean bond enthalpy</li> </ul>	3.1.4.4 Bond enthalpies  Mean bond enthalpy.	MS 1.2: Students understand that bond enthalpies are mean values across a range of	7.7	

<p>bond making and bond breaking. Use data on bond enthalpies to calculate enthalpies of reaction, combustion and formation.</p>	<ul style="list-style-type: none"> <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 differ from those determined using Hess's Law</li> </ul>	<p>Students should be able to: define the term mean bond enthalpy; use mean bond enthalpies to calculate an approximate value of <math>\Delta H</math> for reactions in the gaseous phase; explain why values from mean bond enthalpy calculations differ from those determined using Hess's Law.</p>	<p>compounds containing that bond.</p>		
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One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 8 – Kinetics (10 hours)					
1 Rate of reaction  Recap of what is meant by the term rate of reaction and how it is calculated. Explain how different factors affect the rate.	<ul style="list-style-type: none"> <li>Recall the definition of rate of reaction</li> <li>Describe how different factors affect the rate of reaction</li> <li>Explain how reactions take place in terms of collisions, and how various factors increase the rate.</li> </ul>	3.1.5.1 Collision theory  Reactions can only occur when collisions take place between particles having sufficient energy. This energy is called the activation energy.  Students should be able to: define the term activation energy; explain why most collisions do not lead to a reaction.		8.1	
2 Collision theory  Introduces the idea of orientation and activation. Uses these ideas to explain why most collision do not lead to a reaction. Introduces the idea of one and two-step reactions	<ul style="list-style-type: none"> <li>describe how moving particles store energy (translational, vibrational and rotational)</li> <li>define the term activation energy</li> <li>explain why most collisions do not lead to a reaction</li> </ul>			8.2 8.3	
3 Distribution of stored energy  Introduces the idea that frequency of collision is proportional to square root of the temperature. Introduces the Maxwell-Boltzmann distribution of energy.	<ul style="list-style-type: none"> <li>Explain the shape of different energy –level diagrams, including two-step reactions.</li> <li>Draw the Maxwell-Boltzmann distribution of stored energy in gas molecules</li> <li>Explain the shape of the curve</li> <li></li> <li></li> </ul>	3.1.5.2 Maxwell–Boltzmann distribution  Maxwell–Boltzmann distribution of molecular energies in gases.  Students should be able to draw and interpret distribution curves for different temperatures.		8.3 8.4 8.5	

<p>4 Determining rates of reaction</p> <p>Introduces the idea of zero, first and second order with respect to different reactants. Investigation to calculate the order of a reaction from concentration-time graphs</p>	<ul style="list-style-type: none"> <li>Recognise zero, first and second order reactions from concentration-time graphs</li> <li>interpret graphs of reactant concentration or product concentration against time</li> <li>calculate the rate of reaction by drawing tangents to the curve graphs</li> </ul>	<p>3.1.5.3 Effect of temperature on reaction rate</p> <p>Meaning of the term rate of reaction.</p> <p>The qualitative effect of temperature changes on the rate of reaction. Students should be able to use the Maxwell–Boltzmann distribution to explain why a small temperature increase can lead to a large increase in rate.</p>	<p>AT a, b, k and l</p> <p>PS 2.4 and 3.1: Students could investigate the effect of temperature on the rate of reaction of sodium thiosulfate and hydrochloric acid by an initial rate method.</p>	8.6	
<p>5 Initial rate method</p> <p>Introduces the initial rate method and how it can be used to determine how temperature affects the rate of reaction between sodium thiosulfate and hydrochloric acid</p>	<ul style="list-style-type: none"> <li>describe how to find the initial rate of reaction</li> <li>explain how to use the initial rate method to compare factors that affect rate</li> <li>compare different ways of measuring rate of reaction</li> </ul>			8.6	
<p>6 Concentration, pressure and reaction rate</p> <p>Questions and graphs are used to determine the effect of concentration and pressure on the rate of reaction and how this effect can be measured.</p>	<ul style="list-style-type: none"> <li>Describe how concentration and pressure affect the rate of reaction</li> <li>Make predictions about how concentration and pressure affect the rate of a reaction on a graph</li> <li>Use collision theory to explain the effect of concentration and pressure on the rate of reaction</li> </ul>			8.6	
<p>7 Continuous monitoring method</p>		3.1.5.4 Effect of concentration and pressure	AT a, e, k and i	8.6	Required practical activity 3:

Data for different reactions are given from which graphs are to be plotted. Rate of reactions are calculated from these continuous monitoring methods.	<ul style="list-style-type: none"> <li>Describe what is meant by continuous monitoring</li> <li>Plot and range of graphs including concentration of product formed and concentration of reactant lost for different reactions</li> <li>Explain how continuous monitoring may be used to investigate the rate of reaction</li> </ul>	The qualitative effect of changes in concentration on collision frequency. The qualitative effect of a change in the pressure of a gas on collision frequency.  Students should be able to explain how a change in concentration or a change in pressure influences the rate of a reaction.	Students could investigate the effect of changing the concentration of acid on the rate of a reaction of calcium carbonate and hydrochloric acid by a continuous monitoring method.		Investigation of how the rate of a reaction changes with temperature.
8 Investigating concentration and reaction rate  Investigate the reaction between calcium carbonate and hydrochloric acid at different concentrations to calculate the rate of reaction.	<ul style="list-style-type: none"> <li>Describe how to carry out an investigation using continuous monitoring to find the rate of reaction, using different concentration of acid</li> <li>Collect accurate and reliable data, plot graphs of the data and use these to calculate the rate of reaction</li> <li>Peer review results and account for anomalies</li> </ul>			8.6	
9 Investigating catalysis  Investigate the effect of different substances to catalyse the reactions of the decomposition of hydrogen peroxide and the reaction between sodium thiosulfate and iron (III) nitrate.	<ul style="list-style-type: none"> <li>Explain the term catalyst</li> <li>Recognise that catalysts are specific to particular reactions</li> <li>Describe the effect of different catalysts on the rate of reaction</li> </ul>	3.1.5.5 Catalysts  A catalyst is a substance that increases the rate of a chemical reaction without being changed in chemical composition or amount.  Catalysts work by providing an alternative reaction route of lower activation energy.  Students should be able to use a Maxwell–Boltzmann distribution to help explain how a catalyst increases the rate of a reaction involving a gas.		8.8	
10 Catalysts and reaction rate  Introduce the idea that some catalysts do take	<ul style="list-style-type: none"> <li>Recognise that some catalysts actually take part in a chemical</li> </ul>			8.8	

<p>part in a reactions and investigate an example with the reaction between potassium sodium tartrate and hydrogen peroxide with cobalt (II) chloride as a catalyst. Relate catalysis to the Maxwell Boltzmann distribution.</p>	<p>reaction to form an intermediate step</p> <ul style="list-style-type: none"> <li>• Make careful observations from investigations</li> <li>• use a Maxwell–Boltzmann distribution to help explain how a catalyst increases the rate of a reaction involving a gas</li> </ul>				
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One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 9 – Equilibria (7 hours)					
<p>1 Dynamic Equilibrium</p> <p>Recap ideas about reversible reactions. Use the effect of heat on iodine crystals, the dissociation of ammonium chloride, carbon dioxide in water and chromate (VI) and dichromate (VI) and to investigate dynamic equilibrium.</p>	<ul style="list-style-type: none"> <li>Give examples of different reversible changes</li> <li>Explain how reversible reactions and dynamic equilibrium are linked</li> <li>Apply ideas of dynamic equilibrium to practical investigations</li> </ul>	<p>3.1.6.1 Chemical equilibria and Le Chatelier's principle</p> <p>In a reversible reaction at equilibrium: forward and reverse reactions proceed at equal rates; the concentrations of reactants and products remain constant.</p> <p>Le Chatelier's principle can be used to predict the effects of changes in temperature, pressure and concentration on the position of equilibrium in homogeneous reactions.</p> <p>A catalyst does not affect the position of equilibrium.</p> <p>Students should be able to: use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium; explain why for a</p>	<p>PS 1.1: Students could carry out test-tube equilibrium shifts to show the effect of concentration and temperature (e.g. <math>\text{Cu}(\text{H}_2\text{O})_6^{2+}</math> with concentrated HCl).</p> <p>MS 0.3: Students estimate the effect of changing experimental parameters on a measurable value, e.g. how the value of <math>K_c</math> would change with temperature, given different specified conditions.</p> <p>MS 1.1: Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated results can only be reported to</p>	9.1	
<p>2 Investigating dynamic equilibrium</p> <p>Investigate what happens at equilibrium through the changes between iodine in cyclohexane and potassium iodide</p>	<ul style="list-style-type: none"> <li>Use observations from reactions in equilibrium to make inferences</li> <li>explain that in a reversible reaction at equilibrium forward and reverse reactions proceed at equal rates</li> <li>explain that in a reversible reaction at equilibrium the concentrations of reactants and products remain constant</li> </ul>	<p>Students should be able to: use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium; explain why for a</p>	<p>MS 1.1: Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated results can only be reported to</p>	9.1	
<p>3 The equilibrium constant <math>K_c</math></p> <p>Introduce what is meant by a homogenous and heterogeneous system. Introduce the expression</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> </ul>	<p>Students should be able to: use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium; explain why for a</p>	<p>MS 1.1: Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated results can only be reported to</p>	9.2	



for $K_c$ and use it to perform calculations		reversible reaction used in an industrial process, a compromise temperature and pressure may be used.	the limits of the least accurate measurement.		
4 Determining $K_c$ for the reaction of ethanol and ethanoic acid  Use the reaction between ethanol and ethanoic acid to determine the value of $K_c$	<ul style="list-style-type: none"> <li>describe how to carry out an investigation to determine <math>K_c</math></li> <li>Determine the value of <math>K_c</math> from data obtained from reliable and accurate observations.</li> <li>compare the value to accepted values and account for differences</li> </ul>	3.1.6.2 Equilibrium constant $K_c$ for homogeneous systems  The equilibrium constant $K_c$ is deduced from the equation for a reversible reaction. The concentration, in $\text{mol dm}^{-3}$ , of a species X involved in the expression for $K_c$ is represented by [X]  The value of the equilibrium constant is not affected either by changes in concentration or addition of a catalyst.  Students should be able to: construct an expression for $K_c$ for a homogeneous system in equilibrium; calculate a value for $K_c$ from the equilibrium concentrations for a homogeneous system at constant temperature; perform calculations involving $K_c$ ; predict the	MS 2.2 and 2.3: Students calculate: the concentration of a reagent at equilibrium; the value of an equilibrium constant $K_c$  PS 2.3: Students could determine the equilibrium constant, $K_c$ , for the reaction of ethanol with ethanoic acid in the presence of a strong acid catalyst to ethyl ethanoate.	9.2	
5 Factors affecting equilibrium  Introduce Le Châtelier's principle and how it is used to make predictions about the equilibrium position with different changes. Use the demonstration of the equilibrium between $\text{ICl (l)}$ and $\text{Cl}_2 \text{(g)}$ and $\text{ICl}_3 \text{(s)}$ to demonstrate the effect of temperature on the position of equilibrium.	<ul style="list-style-type: none"> <li>describe Le Châtelier's principle</li> <li>predict how changes in temperature, pressure and concentration affect the position of equilibrium, using this principle</li> <li>Make quantitative predictions about the effect of concentration, temperature and pressure on the equilibrium position, using this principle.</li> <li></li> </ul>			9.3	
6 Investigating equilibria in solutions of metal ions  Use reactions with complex ions to investigate the effects of	<ul style="list-style-type: none"> <li>Describe some examples of complex ions</li> <li>Make and record careful observations relating to reactions involving complex ions</li> </ul>			9.3	

<p>concentration and temperature on the position of equilibrium. Apply Le Châtelier's principle to predict and explain the changes.</p>	<ul style="list-style-type: none"> <li>• Use Le Châtelier's principle to predict what effect changes to concentration and temperature have on given reactions</li> </ul>	<p>qualitative effects of changes of temperature on the value of <math>K_c</math></p>		
<p>7 Equilibria and the chemical industry</p> <p>Recap ideas about the Haber Process and introduce the manufacture of ethyl butanoate and nitric acid as industrial reactions in dynamic equilibrium. Analyse these reactions to determine the optimum conditions for product yield and compare to the actual reaction conditions</p>	<ul style="list-style-type: none"> <li>• Describe industrial reactions which are in dynamic equilibrium</li> <li>• Use Le Châtelier's principle to predict what the optimum conditions for the reactions are likely to be</li> <li>• Explain why compromises in temperature and pressure need to be made.</li> </ul>		<p>9.3</p>	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 10 – Redox reactions (4 hours)					
<p>1 Oxidation and reduction</p> <p>Recap ideas about reduction and oxidation, using the electrolysis of sodium chloride solution. Use the reaction between sodium and chlorine to investigate redox reactions, writing half and full equations.</p>	<ul style="list-style-type: none"> <li>Describe oxidation and reduction processes in electrolysis and other reactions in terms of electron transfer</li> <li>Explain what happens in a redox reaction like the reaction between sodium and chlorine</li> <li>Use simple half equations to write full equations</li> </ul>	<p>3.1.7 Oxidation, reduction and redox equations</p> <p>Oxidation is the process of electron loss and oxidising agents are electron acceptors.</p> <p>Reduction is the process of electron gain and reducing agents are electron donors.</p> <p>The rules for assigning oxidation states.</p> <p>Students should be able to: work out the oxidation state of an element in a compound or ion from the formula;</p>		10.1	
<p>2 Oxidation states</p> <p>Introduce the term oxidation state and how to determine what they are for different elements in a compound. Use reactions of manganese to investigate the different oxidation states it exists in.</p>	<ul style="list-style-type: none"> <li>explain the term oxidation state</li> <li>work out the oxidation state of an element in a compound or ion from the formula</li> <li>make accurate observations to predict the oxidation state of manganese in different reactions</li> </ul>	<p>write half-equations identifying the oxidation and reduction processes in redox reactions combine half-equations to give an overall redox equation.</p>		10.2	
<p>3 Redox reactions</p> <p>Recap</p>	<ul style="list-style-type: none"> <li>identify the oxidation states of an element as a reactant and product in a reaction</li> <li>Identify which “atoms” are oxidised and reduced during chemical changes</li> </ul>			10.3	

	<ul style="list-style-type: none"> <li>• Explain which are the reducing and oxidizing agents in given reactions</li> </ul>			
<p>4 Redox equations</p> <p>Use metal displacement reactions to further investigate redox reactions. Practice writing full redox reactions and balancing them.</p>	<ul style="list-style-type: none"> <li>• write complex half-equations identifying the oxidation and reduction processes in redox reactions</li> <li>• combine half-equations to give an overall redox equation</li> <li>• Balance redox equations by balancing electrons, balancing charge with <math>H^+</math> and balancing with <math>H_2O</math>.</li> </ul>		10.3	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 11 – Group 2, the alkaline earth metals (5 hours)					
<p>1 Physical properties</p> <p>Recap ideas of atomic radius, nuclear charge, ionisation energy and electron shielding. Use data to explore the differences between the physical properties of group 1 and group 2 me</p>	<ul style="list-style-type: none"> <li>Compare how the physical properties of group 1 and group 2 metals change down and between the groups</li> <li>explain the trends in atomic radius between group 1 and 2 metals and first ionisation energy using ideas about atomic radius, nuclear charge and electron shielding</li> <li>explain the differences in melting point of the group 1 and 2 elements in terms of their structure and bonding</li> </ul>	<p>3.2.2 Group 2, the alkaline earth metals</p> <p>The elements in Group 2 are called the alkaline earth metals. The trends in the solubilities of the hydroxides and the sulfates of these elements are linked to their use. Barium sulfate, magnesium hydroxide and magnesium sulfate have applications in medicines whilst calcium hydroxide is used in agriculture to change soil pH, which is essential for good crop production and maintaining the food supply.</p>	<p>AT c and k</p> <p>PS 2.2: Students could test the reactions of Mg–Ba with water and Mg with steam and record their results.</p> <p>AT d and k</p> <p>PS 2.2: Students could test the solubility of Group 2</p>	11.1	
<p>2 Chemical properties</p> <p>Recap familiar reactions of group 2 compounds including reactions involving hard water, reactions of limestone. Compare the reactivity of Ca and Mg with water and oxygen, making predictions about how the reactivity changes down the group</p>	<ul style="list-style-type: none"> <li>describe some common reactions of group 2 elements and compounds</li> <li>describe the reactions the reactions of Mg–Ba with oxygen and water</li> <li>explain the trends in chemical reactivity down the group</li> </ul>	<p>The trends in atomic radius, first ionisation energy and melting point of the elements Mg–Ba</p> <p>Students should be able to: explain the trends in atomic radius and first ionisation energy; explain the melting point of the elements in terms of their structure and bonding.</p>	<p>hydroxides by mixing solutions of soluble Group 2 salts with sodium hydroxide and record their results.</p> <p>AT d and k</p> <p>PS 2.2: Students could test the solubility of Group 2</p>	11.3	
<p>3 Group 2 metal salts</p> <p>Use reactions between group 2 nitrates and sodium hydroxide, sodium carbonate and sodium sulfate to</p>	<ul style="list-style-type: none"> <li>make careful observations to draw inferences and account for anomalous results</li> </ul>	<p>The reactions of the elements Mg–Ba with water. The use of magnesium in the extraction of titanium from <math>\text{TiCl}_4</math>. The relative solubilities of the hydroxides of the elements Mg–Ba in water. <math>\text{Mg}(\text{OH})_2</math> is sparingly soluble.</p>	<p>hydroxides by mixing solutions of soluble Group 2</p>	11.3	

determine the trend in solubility of the group 2 hydroxides, carbonates and sulfates.	<ul style="list-style-type: none"> <li>explain the relative solubilities of the sulfates of the elements Mg–Ba in water</li> <li>explain why BaCl<sub>2</sub> solution is used to test for sulfate ions and why it is acidified</li> <li></li> </ul>	<p>The use of Mg(OH)<sub>2</sub> in medicine and of Ca(OH)<sub>2</sub> in agriculture.</p> <p>The use of CaO or CaCO<sub>3</sub> to remove SO<sub>2</sub> from flue gases.</p> <p>The relative solubilities of the sulfates of the elements Mg–Ba in water.</p> <p>BaSO<sub>4</sub> is insoluble.</p> <p>The use of acidified BaCl<sub>2</sub> solution to test for sulfate ions.</p> <p>The use of BaSO<sub>4</sub> in medicine</p> <p>explain why BaCl<sub>2</sub> solution is used to test for sulfate ions and why it is acidified.</p>	<p>2 salts with sodium hydroxide and record their results.</p> <p>Students could test the solubility of Group 2 sulfates by mixing solutions of soluble Group 2 salts with sulfuric acid and record their results.</p> <p>Students could test for sulfate ions using acidified barium chloride and record their results.</p> <p>Research opportunity: Students could investigate the use of BaSO<sub>4</sub> in medicine.</p>		
<p>4 Uses of Group 2 and their compounds</p> <p>Use research and questions to identify and explain the main uses of group 2 elements and compounds in commercial and industrial applications</p>	<ul style="list-style-type: none"> <li>Describe some key uses of group 2 element and compounds</li> <li>explain the use of magnesium in the extraction of titanium</li> <li>explain why some applications use compounds of Group 2 elements</li> <li></li> </ul>			11.2 11.4	
<p>5 Revision of Group 2 and their compounds</p> <p>Summarise the features of group 2 elements and compounds in a ten minute radio programme. Explain all the main trends in physical properties and chemical reactions, and describe their main uses.</p>	<ul style="list-style-type: none"> <li>Describe some uses of Group 2 elements and compounds</li> <li>Describe some reactions of group 2 elements and compounds</li> <li>Explain the trends in all physical and chemical properties without any resources</li> </ul>			11.1 to 11.4	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 12 – Group 7 (17), the Halogens (6 hours)					
<p>1 physical properties of halogens</p> <p>Recap of the appearance and properties of halogens and types of bonding they exhibit. Use data to describe the trends in melting and boiling points and electronegativity and explain trends using ideas structure, bonding, nuclear charge, electron shielding and atomic radius.</p>	<ul style="list-style-type: none"> <li>describe the appearance and properties of the halogens at room temperature</li> <li>explain the trends in the boiling point of the elements in terms of their structure and bonding</li> <li>explain the trend in electronegativity of the halogens and the impact of this on reactivity</li> </ul>	<p>3.2.3.1 Trends in properties</p> <p>The trends in electronegativity and boiling point of the halogens.</p> <p>Students should be able to: explain the trend in electronegativity; explain the trend in the boiling point of the elements in terms of their structure and bonding.</p> <p>The trend in oxidising ability of the halogens down the group, including the displacement reactions of halide ions in aqueous solution.</p> <p>The trend in reducing ability of the halide ions, including the reactions of solid sodium halides with concentrated sulfuric acid.</p> <p>The use of acidified silver nitrate solution to identify and distinguish between halide ions.</p> <p>The trend in solubility of the silver halides in ammonia.</p> <p>Students should be able to explain why: silver nitrate solution is used to identify halide ions; the silver nitrate solution is acidified; ammonia solution is added.</p>	<p>AT d and k</p> <p>PS 2.2: Students could carry out test-tube reactions of solutions of the halogens (Cl<sub>2</sub>, Br<sub>2</sub>, I<sub>2</sub>) with solutions containing their halide ions (e.g. KCl, KBr, KI).</p> <p>Students could record observations from reactions of NaCl, NaBr and NaI with concentrated sulfuric acid.</p> <p>Students could carry out tests for halide ions using acidified silver nitrate, including the use of ammonia to distinguish the silver halides formed.</p>	<p>12.1</p> <p>12.2</p> <p>12.3</p>	
<p>2 Redox reactions of halogens</p> <p>Use displacement reactions to investigate the trend in oxidising ability down the halogen group. Demonstrate reactions between sulphuric acid and sodium halides to investigate reducing ability.</p>	<ul style="list-style-type: none"> <li>explain the trend in oxidising ability of the halogens down the group</li> <li>carry out and explain test tube reactions of halogen solutions with solutions of their halide ions</li> <li>explain the trends in reducing ability of the halide ions, including the reactions of solid sodium halides with concentrated sulfuric acid</li> </ul>				
<p>3 Identifying halides</p> <p>Use experiments with acidified silver nitrate</p>	<ul style="list-style-type: none"> <li>carry out tests for halide ions using acidified silver nitrate solution and</li> </ul>				

<p>solution and concentrated ammonia solution to identify and distinguish between halide ions; investigate the solubility of halides in dilute ammonia and excess concentrated ammonia solution.</p>	<p>ammonia solution to distinguish between the halides formed</p> <ul style="list-style-type: none"> <li>• make observations of how halide ions behave in dilute and excess concentrated ammonia solution</li> <li>• Explain why silver nitrate is used, why it is acidified and why ammonia solution is used.</li> </ul>				
<p>4 Required practical 4 Carry out simple tests to identify cations – group 2 M<sup>2+</sup>; ammonium NH<sub>4</sub><sup>+</sup> and anions group 7 (17), X<sup>-</sup> halide ions; hydroxide OH<sup>-</sup>; carbonate CO<sub>3</sub><sup>2-</sup>; sulfate SO<sub>4</sub><sup>2-</sup>.</p>	<ul style="list-style-type: none"> <li>• work safely in a laboratory</li> <li>• use laboratory apparatus for a variety of experimental techniques, including qualitative tests for ions</li> <li>• draw conclusions from observations made</li> </ul>	<p>3.2.3.2 Uses of chlorine and chlorate(I)</p> <p>The reaction of chlorine with water to form chloride ions and chlorate(I) ions.</p> <p>The reaction of chlorine with water to form chloride ions and oxygen.</p> <p>Appreciate that society assesses the advantages and disadvantages when deciding if chemicals should be added to water supplies.</p> <p>The use of chlorine in water treatment.</p> <p>Appreciate the benefits to health of water treatment by chlorine outweigh its toxic effects.</p>	<p>Research opportunity: Students could investigate the treatment of drinking water with chlorine. Students could investigate the addition of sodium fluoride to water supplies.</p>	<p>12.3</p>	<p>Required practical activity 4: Carry out simple test-tube reactions to identify:</p> <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></li> </ul>
<p>5 Chlorine and water supplies</p> <p>Introduce the term disproportionation and apply it to the reaction between chlorine and water. Use data to explore the use of chlorine in water supplies</p>	<ul style="list-style-type: none"> <li>• describe how chlorine reacts with water to form chloride and chlorate ions and chlorine gas</li> <li>• Use ideas about oxidation and reduction to explain how chlorine reacts with water</li> <li>• explain the benefits and disadvantages of adding chlorine to water supplies</li> </ul>	<p>The reaction of chlorine with cold, dilute, aqueous NaOH and uses of the solution formed.</p>		<p>12.4</p>	



<p>6 Chlorine in bleach</p> <p>Use a demonstration to show the bleaching action of sodium chlorate (I) food dyes. Use a class practice to investigate the concentration of household bleach.</p>	<ul style="list-style-type: none"> <li>• describe the chemical reaction between chlorine and cold dilute sodium hydroxide solution</li> <li>• carry out an experiment and process the results to estimate the concentration of sodium chlorate(I) in different commercial bleaches</li> <li>• Calculate the concentration of chlorine using data from the investigation</li> </ul>		12.4	
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One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 13 – Halogenoalkanes (7 hours)					
<p>1 Describing halogenoalkanes and their reactions</p> <p>Recap the structure and naming of halogenoalkanes. Identify polarity in the C-X bond and recap ideas about the use of curly arrows in reaction mechanisms.</p>	<ul style="list-style-type: none"> <li>write chemical, structural and displayed formulae for halogenoalkanes</li> <li>explain why halogenoalkanes are polar and the impact of this on reaction mechanisms</li> <li>use curly arrows to describe reaction mechanisms</li> </ul>	<p>3.3.3.1 Nucleophilic substitution</p> <p>Halogenoalkanes contain polar bonds.</p> <p>Halogenoalkanes undergo substitution reactions with the nucleophiles <math>\text{OH}^-</math>, <math>\text{CN}^-</math> and <math>\text{NH}_3</math>.</p> <p>Students should be able to: outline the nucleophilic substitution mechanisms of these reactions; explain why the carbon-halogen bond enthalpy influences the rate of reaction.</p>	<p>AT a, b and k</p> <p>PS 4.1: Students could follow instructions when carrying out test-tube hydrolysis of halogenoalkanes to show their relative rates of reaction.</p> <p>AT d, g and k: Students could prepare a chloroalkane, purifying the product using a separating funnel and distillation.</p>	13.1	
<p>2 Nucleophilic substitution reactions of halogenoalkanes</p> <p>Introduce the term nucleophile with examples. Explain the reaction mechanism for nucleophilic substitution for a range of common nucleophiles, using curly arrow.</p>	<ul style="list-style-type: none"> <li>explain the term nucleophile</li> <li>explain why nucleophiles attack the C-X bond</li> <li>outline the nucleophilic substitution mechanism of halogenoalkanes with the nucleophiles <math>\text{OH}^-</math>, <math>\text{NH}_3</math> and <math>\text{CN}^-</math></li> </ul>			13.1 13.2	
<p>3 Hydrolysis of halogenoalkanes</p> <p>Compare the rate of hydrolysis of 1-bromobutane and 1-</p>	<ul style="list-style-type: none"> <li>explain the term hydrolysis</li> <li>carry out test tube hydrolysis reactions of halogenoalkanes and describe methods to compare the rate of reaction</li> </ul>			13.2	

chlorobutane using a practical investigation	<ul style="list-style-type: none"> <li>explain why the carbon-halogen bond enthalpy influences the rate of reaction</li> <li>construct</li> </ul>				
<p>4 Mechanisms for the hydrolysis of primary and tertiary halogenoalkanes</p> <p>Draw examples of primary, secondary and tertiary halogenoalkanes and explain the differences between them. Explain the reaction mechanisms of primary and tertiary halogenoalkanes with water using curly arrows</p>	<ul style="list-style-type: none"> <li>identify and draw displayed formulas for primary, secondary and tertiary halogenoalkanes</li> <li>explain why there are differences in the reaction mechanisms of primary, secondary and tertiary halogenoalkanes</li> <li>explain the mechanisms for the hydrolysis of primary and tertiary halogenoalkanes</li> </ul>			13.2	
<p>5 Preparation of a halogenoalkane</p> <p>Follow instructions, prepare a sample of 2-chloro-2-methylpropane</p>	<ul style="list-style-type: none"> <li>follow instructions and work safely in the lab</li> <li>explain how to prepare a sample of a halogenoalkane</li> <li>explain how to purify the product using a separating funnel and distillation</li> </ul>			13.2	
<p>6 Elimination reactions of halogenoalkanes</p> <p>Explain the difference between substitution and elimination reactions and why elimination reactions occur instead of</p>	<ul style="list-style-type: none"> <li>outline the mechanism of elimination reactions of halogenoalkanes</li> <li>explain how hydroxide ions act as a base in elimination reactions</li> <li>predict the products of competing substitution and elimination reactions between hydroxide ions and halogenoalkanes</li> </ul>	<p>3.3.3.2 Elimination</p> <p>The concurrent substitution and elimination reactions of a halogenoalkane (e.g. 2-bromopropane with potassium hydroxide).</p> <p>Students should be able to: explain the role of the reagent as both</p>		13.2	

<p>substitution reactions. Explain the mechanism of the reaction using curly arrows.</p>		<p>nucleophile and base; outline the mechanisms of these reactions.</p>			
<p>7 Ozone depletion</p> <p>Describe the role of ozone and evidence for its depletion. Use reactions to explain how it is being depleted. Explain some of the measures used to restore its concentrations</p>	<ul style="list-style-type: none"> <li>describe the role of naturally formed ozone in the upper atmosphere</li> <li>use equations to explain how chlorine atoms catalyse the decomposition of ozone</li> <li>explain the role of the scientific community in providing evidence for legislation to ban the use of CFCs as solvents and refrigerants and find alternatives</li> </ul>	<p>3.3.3.3 Ozone depletion</p> <p>Ozone, formed naturally in the upper atmosphere, is beneficial because it absorbs ultraviolet radiation.</p> <p>Chlorine atoms are formed in the upper atmosphere when ultraviolet radiation causes C–Cl bonds in chlorofluorocarbons (CFCs) to break.</p> <p>Chlorine atoms catalyse the decomposition of ozone and contribute to the formation of a hole in the ozone layer.</p> <p>Appreciate that results of research by chemists provided evidence for legislation to ban the use of CFCs as solvents and refrigerants. Chemists have now developed alternative chlorine-free compounds.</p> <p>Students should be able to: use equations, such as the following, to explain how chlorine atoms catalyse decomposition of ozone:</p> $\text{Cl}\cdot + \text{O}_3 \rightarrow \text{ClO}\cdot + \text{O}_2 \text{ and } \text{ClO}\cdot + \text{O}_3 \rightarrow 2\text{O}_2 + \text{Cl}\cdot$	<p>Research opportunity: Students could investigate the role of chemists in the introduction of legislation to ban the use of CFCs and in finding replacements.</p>	<p>13.3</p>	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 14 – Alkenes (7 hours)					
1 Alkenes Recap the main features of alkenes, including their ability to form geometric and stereoisomers. Explain the differences between sigma and pi bonds.	<ul style="list-style-type: none"> <li>Describe the essential features of alkenes as a homologous series</li> <li>Recall the types of isomers present in alkenes</li> <li>Explain the nature of the bonding in alkenes including pi and sigma bonds</li> </ul>	<p>3.3.4.1 Structure, bonding and reactivity</p> <p>Alkenes are unsaturated hydrocarbons.</p> <p>Bonding in alkenes involves a double covalent bond, a centre of high electron density.</p>		14.1	
2 Electrophilic addition reaction of alkenes Introduce the term electrophile. Using display formulae, explore the electrophilic addition reactions of different electrophiles with alkenes.	<ul style="list-style-type: none"> <li>name the features of an electrophile and give some examples</li> <li>explain why alkenes are susceptible to attack by electrophiles</li> <li>describe the mechanisms of the electrophilic addition reactions of alkenes with HBr, H<sub>2</sub>SO<sub>4</sub> and Br<sub>2</sub> using curly arrows</li> </ul>	<p>3.3.4.2 Addition reactions of alkenes</p> <p>Electrophilic addition reactions of alkenes with HBr, H<sub>2</sub>SO<sub>4</sub> and Br<sub>2</sub>.</p> <p>The use of bromine to test for unsaturation.</p> <p>The formation of major and minor products in addition reactions of unsymmetrical alkenes.</p> <p>Students should be able to: outline the mechanisms for these reactions; explain the formation of major and minor products by reference to the relative stabilities of primary, secondary and tertiary carbocation intermediates.</p>	AT d and k PS 4.1: Students could test organic compounds for unsaturation using bromine water and record their observations.	14.2	
3 Hydration of alkenes Carry out a practical to investigate the electrophilic addition reaction between hex-1-ene and sulphuric acid with water.	<ul style="list-style-type: none"> <li>Use equipment safely and follow a procedure</li> <li>Describe the procedure to produce hexan-2-ol from hex-1-ene</li> <li>Explain the reaction mechanism for the addition reaction, explaining why hexan-2-ol was made instead of hexan-1-ol</li> </ul>			14.2	

<p>4 The addition of reactions of unsymmetrical alkenes</p> <p>Use molecular models to explore the possible products from asymmetrical alkenes. Explain the stability of primary, secondary and tertiary carbocations, and use these ideas to predict the products formed from addition reactions of unsymmetrical alkenes</p>	<ul style="list-style-type: none"> <li>• identify symmetrical and unsymmetrical alkenes</li> <li>• recognise primary, secondary and tertiary carbocations</li> <li>• predict the major and minor products in addition reactions of unsymmetrical alkenes</li> </ul>			14.2	
<p>5 Determining the degree of unsaturation in oils</p>	<ul style="list-style-type: none"> <li>• describe the test for unsaturation</li> <li>• describe a procedure to test the unsaturation of different oils using bromine water</li> <li>• explain the reaction mechanism of the reaction between bromine water and less familiar alkenes using curly arrows</li> </ul>			14.2	
<p>6 Polymers</p> <p>Recap the terms monomer, polymer, polymerisation and repeat unit, and the use of the repeating unit to show the structure of the polymer from the monomer. Research different common addition polymers and</p>	<ul style="list-style-type: none"> <li>• describe the term polymer and use the IUPAC rules for naming polymers</li> <li>• explain how polymers are formed from alkenes and substituted alkenes, drawing the repeating unit of a polymer from the monomer and from the polymer chain</li> <li>• Understand the nature of the intermolecular forces in polyalkenes.</li> </ul>	<p>3.3.4.3 Addition polymers</p> <p>Addition polymers are formed from alkenes and substituted alkenes.</p> <p>The repeating unit of addition polymers.</p> <p>IUPAC rules for naming addition polymers.</p> <p>Addition polymers are unreactive.</p>	<p>AT k</p> <p>PS 1.2</p> <p>Making poly(phenylethene) from phenylethene.</p>	14.3	

<p>explain their properties and nature of their intermolecular forces</p>		<p>Appreciate that knowledge and understanding of the production and properties of polymers has developed over time.</p>			
<p>7 Making polymers</p> <p>Demonstrate the preparation of poly(phenylethene) and describe the process. Explain the chemical changes taking place and the role of pieces of equipment</p>	<ul style="list-style-type: none"> <li>• Identify the monomer for making poly(phenylethene) and write the equation for the polymerisation reaction.</li> <li>• describe how to make poly(phenylethene) also known as polystyrene in the laboratory</li> <li>• Compare the laboratory preparation with industrial methods</li> </ul>	<p>Typical uses of poly(chloroethene), commonly known as PVC, and how its properties can be modified using a plasticiser.</p> <p>Students should be able to: draw the repeating unit from a monomer structure; draw the repeating unit from a section of the polymer chain; draw the structure of the monomer from a section of the polymer; explain why addition polymers are unreactive; explain the nature of intermolecular forces between molecules of polyalkenes.</p>		<p>14.3</p>	

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 15 – Alcohols (8 hours)					
<p>1 Structure and physical properties alcohols</p> <p>Recap the structure of alcohols and introduce primary, secondary and tertiary alcohols. Use data to analyse and compare the boiling points of alkanes, alkenes and alcohols, explain this in terms of hydrogen bonding</p>	<ul style="list-style-type: none"> <li>use IUPAC rules to name alcohols and identify primary, secondary and tertiary alcohols</li> <li>interpret data on boiling points of the homologous series studied so far</li> <li>explain how hydrogen bonds form in alcohols and the impact this has on their properties</li> </ul>	<p>3.3.5.1 Ethanol production</p> <p>Alcohols are produced industrially by hydration of alkenes in the presence of an acid catalyst.</p> <p>Ethanol is produced industrially by fermentation of glucose. The conditions for this process.</p> <p>Ethanol produced industrially by fermentation is separated by fractional distillation and can then be used as a biofuel.</p>	<p>AT a, d and k</p> <p>PS 1.2: Students could produce ethanol by fermentation, followed by purification by fractional distillation.</p>	15.1	
<p>2 Industrial manufacture of ethanol</p> <p>Introduce the commercial production of ethanol from the hydration of ethene. Discuss the reasons for the reaction conditions and outline the mechanism for the reaction</p>	<ul style="list-style-type: none"> <li>describe how alcohols are produced industrially by the hydration of ethene</li> <li>use le Chatelier's principle to explain the conditions used for its manufacture</li> <li>outline the mechanism for the reaction between ethene and steam</li> </ul>	<p>Students should be able to: explain the meaning of the term biofuel; justify the conditions used in the production of ethanol by fermentation of glucose; 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; outline the mechanism for the formation of an alcohol by the reaction of an alkene with steam in the presence of an acid catalyst; discuss the environmental (including ethical) issues linked to decision making about biofuel use.</p>		15.2	
<p>3 Ethanol by fermentation</p> <p>Carry out a practical to produce ethanol by fermentation. Discuss the features and uses of</p>	<ul style="list-style-type: none"> <li>explain the term biofuel and the ethical issues associated to their use</li> <li>explain the conditions and processes required for the</li> </ul>			15.2	



the apparatus and reagents used. Discuss the use of ethanol as a biofuel and the ethical implications arising from its use.	<p>production of ethanol by fermentation</p> <ul style="list-style-type: none"> <li>use equations to support the statement that ethanol produced by fermentation is a carbon neutral fuel and explain why this statement is not valid</li> </ul>				
<p>4 Oxidation of alcohols</p> <p>Use a class practical with acidified potassium dichromate to carry out oxidation reactions with primary, secondary and tertiary alcohols. Explain what is happening in each case and why.</p>	<ul style="list-style-type: none"> <li>describe some oxidizing agents</li> <li>describe how primary alcohols are oxidised to aldehydes, which can be further oxidised to carboxylic acids; secondary alcohols are oxidised to ketones; tertiary alcohols do not oxidise easily</li> <li>write chemical equations for the oxidation reactions using [O] as the oxidant</li> </ul>	<p>3.3.5.2 Oxidation of alcohols</p> <p>Alcohols are classified as primary, secondary and tertiary.</p> <p>Primary alcohols can be oxidised to aldehydes which can be further oxidised to carboxylic acids.</p> <p>Secondary alcohols can be oxidised to ketones.</p> <p>Tertiary alcohols are not easily oxidized. Acidified potassium dichromate(VI) is a suitable oxidising agent.</p>	AT b, d and k	15.3	<p>Students could carry out the preparation of an aldehyde by the oxidation of a primary alcohol.</p> <p>Students could carry out the preparation of a carboxylic acid by the oxidation of a primary alcohol.</p>
<p>5 Preparation of an aldehyde by the oxidation of a primary alcohol</p> <p>Carry out a class practical to prepare an aldehyde from a primary alcohol. Explain how the chemical tests can distinguish between aldehydes and ketones</p>	<ul style="list-style-type: none"> <li>describe how to prepare an aldehyde from a primary alcohol</li> <li>explain how the method used to oxidise a primary alcohol determines whether an aldehyde or a carboxylic acid is produced</li> <li>use chemical tests to distinguish between an aldehyde and a ketone</li> </ul>	<p>Students should be able to: write equations for these oxidation reactions (equations showing [O] as oxidant are acceptable); explain how the method used to oxidise a primary alcohol determines whether an aldehyde or carboxylic acid is obtained; use chemical tests to distinguish between aldehydes and ketones including Fehling's solution and Tollen's reagent.</p>		15.3	
<p>6 Preparation of a carboxylic acid by the oxidation of a primary alcohol</p>	<ul style="list-style-type: none"> <li>carry out the preparation of a carboxylic acid from a primary alcohol</li> </ul>			15.3	

Use a class practical to investigate the preparation of a carboxylic acid from the oxidation of a primary alcohol. Explain how wine growers prevent oxidation of alcohol	<ul style="list-style-type: none"> <li>Explain how the method used to oxidise the primary alcohol determines whether an aldehyde or a carboxylic acid is obtained.</li> <li>Know how to use Fehling's solution and Tollen's reagent to distinguish between an aldehyde and a ketone</li> </ul>				
7 Elimination reactions  Use molecular models to demonstrate elimination reactions. Use curly arrow to explain the mechanism of an elimination reaction in detail.	<p>Describe how alkenes can be formed from alcohols in acid catalysed reactions</p> <ul style="list-style-type: none"> <li>Explain how alkenes produced from alcohols can be used to produce addition polymers without the use of monomers derived from crude oil</li> <li>Outline the mechanism for the elimination of water from alcohols</li> </ul>	3.3.5.3 Elimination  Alkenes can be formed from alcohols by acid-catalysed elimination reactions (mechanism not required).  Alkenes produced by this method can be used to produce polymers without using monomers derived from crude oil.	AT b, d, g and k  PS 4.1: Students could carry out the preparation of cyclohexene from cyclohexanol, including purification using a separating funnel and by distillation.	15.4	
8 Preparation and purification of cyclohexene from cyclohexanol  Class carry out required practical to prepare and purify cyclohexene from cyclohexanol Calculate the atom economy and percentage yield for the reaction.	<ul style="list-style-type: none"> <li>Describe how to carry out the preparation of cyclohexene from cyclohexanol</li> <li>Explain how to purify the product using a separating funnel and distillation</li> <li>Calculate a percentage yield</li> </ul>			15.3 15.4	Required practical 5: Distillation of a product from a reaction.

One hour lessons	Learning Outcomes	Specification Content	Skills Covered	Student Book Section	Required Practicals
CHAPTER 16 – Organic analysis (6 hours)					
<p>1 identifying functional groups</p> <p>Carry out a variety of different tests to identify functional groups studied in this specification</p>	<ul style="list-style-type: none"> <li>recall the reactions of functional groups in this specification</li> <li>describe and carry out test tube reactions to identify alcohol, aldehydes, alkene, halogenoalkane and carboxylic acid on known samples, either by using given methods, or by researching their own methods</li> <li>record their observations in a table and make suitable deductions</li> </ul>	<p>3.3.6.1 Identification of functional groups by test-tube reactions</p> <p>The reactions of functional groups listed in the specification.</p> <p>Students should be able to identify the functional groups using reactions in the specification.</p>	<p>AT b, d and k</p> <p>PS 2.2, 2.3 and 4.1:</p> <p>Students could carry out test-tube reactions in the specification to distinguish alcohols, aldehydes, alkenes and carboxylic acids.</p>	16.1	
<p>2 Required practical activity</p> <p>Students carry out their own tests to determine the identity of unknown functional groups as part of required practical 6.</p>	<ul style="list-style-type: none"> <li>use suitable chemical tests to identify the functional groups present in organic compounds</li> <li>Decide which tests to carry out to determine the identity of the following functional groups:- alcohol, aldehydes, alkene and carboxylic acid</li> <li>explain the chemistry of each test</li> </ul>				Required practical 6: Tests for alcohol, aldehyde, alkene and carboxylic acid.
<p>3 Mass spectrometry</p> <p>Recap of the time of flight spectrometer. Interpret a range of mass spectra. Compare the differences between</p>	<ul style="list-style-type: none"> <li>explain how mass spectrometry can be used to determine the molecular formula of a compound</li> <li>use precise atomic masses to calculate precise molecular masses</li> </ul>	<p>3.3.6.2 Mass spectrometry</p> <p>Mass spectrometry can be used to determine the molecular formula of a compound.</p> <p>Students should be able to use precise atomic masses and the precise</p>		16.2	

low and high resolution mass spectrometers.	<ul style="list-style-type: none"> <li>use precise molecular masses to determine the molecular formula of a compound</li> </ul>	molecular mass to determine the molecular formula of a compound.			
4 Infrared spectroscopy Introduce infrared radiation and explain its effect on the bonds in organic molecules. Explain how the spectrometer works	<ul style="list-style-type: none"> <li>explain the link between wavelength and wavenumber</li> <li>know that bonds in a molecule absorb infrared radiation at characteristic wavenumbers</li> <li>explain the principles of infrared spectrometry</li> </ul>	3.3.6.3 Infrared spectroscopy  Bonds in a molecule absorb infrared radiation at characteristic wavenumbers.  'Fingerprinting' allows identification of a molecule by comparison of spectra.	Students should be able to use data in the Chemistry Data Booklet to suggest possible structures for molecules.	16.3	
5 Interpreting infrared spectra  Interpret a range of different infrared spectra and identify the functional groups present. Use spectra to explain why carbon dioxide, methane and water are greenhouse gases.	<ul style="list-style-type: none"> <li>use infrared spectra and the Chemistry Data sheet to identify particular bonds and hence functional groups as well as impurities</li> <li>know that fingerprinting allows the identification of a molecule by comparing spectra</li> <li>explain the link between the absorption of infrared radiation by the bonds in carbon dioxide, water and methane and global warming</li> </ul>	Students should be able to use infrared spectra and the Chemistry Data Booklet to identify particular bonds, and therefore functional groups, and also to identify impurities.  The link between absorption of infrared radiation by bonds in CO <sub>2</sub> , methane and water vapour and global warming.		16.3	
6 Revision of organic chemistry  Groups work together to produce a set of detailed revision cards or posters or PowerPoints. They summarise all the reaction mechanisms studied in the specification. These are	<ul style="list-style-type: none"> <li>provide displayed formulae of all functional groups and name all organic compounds given</li> <li>Explain how reactions connect different functional groups together and describe reaction conditions for all reactions.</li> <li>Explain all reaction mechanisms studied in full using curly arrows correctly.</li> </ul>			All chapters relating to organic chemistry	

collated together to produce a reaction pathway connecting all different functional groups, showing the reactions between them and the reaction conditions for all reactions.					
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