



**L-Università
ta' Malta**

**MATSEC
Examinations
Board**



**AM 06 Syllabus
CHEMISTRY**

2028

Contents

Introduction	2
List of Subject Foci	3
List of Learning Outcomes	4
Learning Outcomes and Assessment Criteria.....	5
Scheme of Assessment	57

Introduction

Chemistry is not only a subject of academic study that has engaged scholars worldwide for centuries but also its products and processes have indeed enabled the very existence of societies and allowed their development. It is a central science, which informs several other disciplines including biology, medicine, materials science, agricultural science and others. Chemical warfare and pollution of the environment as well as drug abuse, adulteration of food and similar examples of misuse, or improper use, or disposal of chemicals often cast shadows on this scientific enterprise and generate mistrust of chemicals by the public.

The main purpose of this syllabus is to impart a proper understanding of the core principles and unifying ideas of chemical science, supported and illustrated by a significant body of essential factual knowledge that should help students acquire a meaningful experience of chemistry to serve them adequately in their future studies. The syllabus also provides an opportunity for learners to become properly acquainted with this important science in order for them to be able to better appraise both the opportunities and the challenges, which chemistry and its products present to society.

The syllabus for Chemistry at this level assumes and builds on knowledge of the subject at a level that is equivalent to that covered by the Secondary Education Certificate syllabus.

The syllabus is presented in a number of Learning Outcomes (1 – 16) which include an Assessment Criteria column. This is intended to assist teachers, students and examiners determine the width and the depth of treatment for each subject matter mentioned.

Teachers may of course go beyond the learning outcomes in class but the examination will test the outcomes as described in this syllabus in terms of the Assessment Criteria.

The course is intended to build both a theoretical knowledge base in chemistry that prepares students for further higher studies in the subject as well as to develop basic practical skills. With respect to practical chemistry, students should learn how to perform laboratory work in a manner that respects both their health and the environment; teachers should evaluate the risks associated with practical work with a view to safeguard student safety and environmental integrity.

List of Subject Foci

1. Atomic Structure
2. Electronic Theory and Chemical Bonding
3. States of Matter
4. Quantity of Matter
5. Energetics
6. Phase Equilibria
7. Chemical Equilibrium
8. Ionic Equilibria
9. Redox Equilibria
10. Reaction Kinetics
11. Chemical Periodicity
12. Inorganic Chemistry
13. Organic Chemistry – General Principles
14. Chemistry of Aliphatic and Aromatic Compounds
15. Purification and Structure Determination
16. Quantitative and Qualitative Analysis

List of Learning Outcomes

- ☐ I can explain the structure of an atom and investigate differences between different atoms.
- ☐ I can explain different types of bonding, why they exist and how to investigate them.
- ☐ I can explain the Kinetic theory of gases with respect to ideal and real gases.
- ☐ I can apply the mole concept in calculations.
- ☐ I can show an understanding of enthalpy changes, entropy and Gibbs free energy.
- ☐ I can show an understanding of one-component systems, two-component systems, distillation, and osmosis.
- ☐ I can show an understanding of the concept of chemical equilibrium.
- ☐ I can show an understanding of the concept of homogenous and heterogeneous ionic equilibria.
- ☐ I can show an understanding of what happens during redox reactions, their feasibility and possible applications of these reactions.
- ☐ I can show an understanding of the rate of a chemical reaction and the different factors that influence it.
- ☐ I can show an understanding of the periodic relationships amongst period 3 elements and their compounds.
- ☐ I can show an understanding of the physical and chemical properties of a number of inorganic elements and their compounds.
- ☐ I can show an understanding of the structure and properties of different organic molecules.
- ☐ I can show an understanding of the properties, preparation and reactions of a number of organic compounds.
- ☐ I can show an understanding of various purification techniques and the structure of molecules deduced from spectroscopic data.
- ☐ I can use chemicals and simple apparatus in quantitative and qualitative exercises, observe and record results of experimental work, interpret these observations and deduce correct inferences based on both qualitative and quantitative data.

Learning Outcomes and Assessment Criteria

Subject Focus:	Atomic Structure
Learning Outcome 1	I can explain the structure of an atom and investigate differences between different atoms.

Topic	Sub-Topic	Assessment Criteria
1.1 The fundamental particles		1. Describe the structure of an atom in terms of fundamental particles (protons, electrons and neutrons) using the Bohr model. 2. Recall the charges and relative masses of the fundamental particles.
1.2 The nuclear structure of the atom		1. Define the terms proton (or atomic) number and nucleon (or mass) number. 2. Define the relative isotopic mass. 3. Define the relative atomic mass. 4. Calculate the relative atomic mass of an element from the isotopic masses and relative abundances. 5. Describe the use of isotopes as tracers in mechanistic studies using the example of ^{18}O labelling in esterification.
1.3 Radioactivity	1.3.1 Alpha and beta particles and gamma rays	1. Define alpha and beta particles and gamma rays. 2. Describe the properties of alpha and beta particles and gamma rays. <i>Limited to ionisation power, penetrating properties and deflection in electric fields.</i>
	1.3.2 Half-life	3. Define half-life. 4. Calculate half-life from given data. 5. Calculate the initial or final amount of isotopes after a number of half-lives. <i>Limited to three half-lives.</i> 6. Describe C-14 dating. 7. Apply the concept of half-life in determining the age of a carbon-containing artefact.

Topic	Sub-Topic	Assessment Criteria
	1.3.3 Nuclear equations	8. Write balanced nuclear equations. <i>Limited to bombardment by alpha particles and neutrons.</i> <i>Positron emission will not be examined.</i> <i>Calculations requiring the use of the equation for radioactive decay will not be set.</i>
1.4 Quantised energy levels in atoms		1. Construct the electronic configuration of isolated atoms of elements H to Kr using 1s, 2s, 2p etc. notation and electrons-in-boxes notation and applying the 'building-up' (aufbau) principle. 2. Describe quantised energy levels using evidence from atomic spectra. 3. Explain how transitions between energy levels can give rise to flame colours. <i>Excluding questions on the interpretation of atomic spectra.</i>
1.5 Classification of elements		1. Classify elements into groups periods, and 'blocks', including the first row d-block elements.
1.6 Atomic orbitals	1.6.1 Shapes of orbitals	1. Sketch the shape of s and p orbitals.
	1.6.2 Hybridised orbitals	2. Sketch the shapes of hybridised orbitals: sp, sp² and sp³. 3. Identify the hybridisation of carbon atoms in organic molecules. 4. Relate the hybridisation of atomic orbitals to molecular geometry.
1.7 Ionisation energies and electron affinity and relation to electronic configuration		1. Explain how first ionisation energies vary across the second and third period (both the general trends and irregularities) and down a group in the Periodic Table. 2. Explain why second and subsequent electron affinities are endothermic.

Subject Focus:	Electronic Theory and Chemical Bonding
Learning Outcome 2	I can explain different types of bonding, why they exist and how to investigate them.

Topic	Sub-Topic	Assessment Criteria
2.1 The ionic (electrovalent), covalent and co-ordinate (or dative) bond		1. Describe electronegativity as the ability of an atom to attract the bonding electrons in a covalent bond with reference to Pauling's electronegativity scale. 2. Recall the concept of electron sharing and electron transfer between atoms as complete or partial processes depending on the electronegativity of the atomic species concerned. 3. Explain intermediate bonding in terms of ion and bond polarisation. 4. Define dative bonding exemplified by molecular compounds limited to CO, HNO₃, Al₂Cl₆, NH₃BF₃ molecules; H₃O⁺ and NH₄⁺ ions and complex ions/molecules mentioned in this syllabus.
2.2 Polar covalent bonds and electrical dipoles in molecules		1. Describe polar covalent bonds which may, or may not, give rise to molecules with a permanent dipole. <i>Use of symbols δ^+ and δ^- to represent partial charges in bonds.</i> 2. Explain qualitatively polar covalent bonds in molecules in terms of vector representation of dipole moments, which can lead to molecules having a permanent dipole or no dipole.
2.3 Nature of forces in bonding		1. Draw Lewis structures showing all outer electrons as dots and crosses and/or the corresponding bonds as lines. 2. Describe covalent bonding in terms of overlap of orbitals. 3. Describe hybridisation in carbon molecules to produce σ and π bonds.
2.4 Delocalisation of electrons		1. Interpret evidence for delocalisation from bond lengths and thermochemical data. 2. Describe delocalisation of π bonds in, for example, benzene, buta-1,3-diene, ozone and carbonate, nitrate and alkanoate ions by means of resonance structures (canonical forms) and resonance hybrids (delocalised structures).

Topic	Sub-Topic	Assessment Criteria
2.5 Metallic bonding		1. Explain malleability, ductility and electrical conductivity of metals in terms of delocalised electrons. 2. Relate the strength of metallic bonding to the attraction between metallic nuclei and the delocalised valence electrons. 3. Explain the variation of metallic character down a group and across a period.
2.6 Lattice structures of NaCl and CsCl as typical ionic solids	2.6.1 Unit cells	1. Relate the cationic radius to the coordination number and type of crystal lattice. <i>Candidates are expected to identify the crystal structures from unit cells of NaCl and CsCl but will not be asked to draw the unit cells.</i>
2.7 Shapes of molecules and ions	2.7.1 Valence shell electron pair repulsion (VSEPR) theory	1. Recall the approximate bond angles in molecules and ions having both lone and bonding pairs surrounding the central atom with up to six electron pairs, excluding π electron pairs. 2. Sketch appropriate diagrams of the shapes of molecules/ions mentioned in 1. above. 3. State the names of the shapes mentioned in 1. above.
2.8 Molecular crystals and macromolecular structures		1. Explain physical properties such as melting point, boiling point, sublimation, hardness, brittleness and electrical conductivity in terms of structure, where applicable. <i>Limited to iodine, carbon dioxide, water, carbon (diamond), carbon (graphite), silicon(IV) oxide and other elements listed in Section 11.1.</i>
2.9 Attractive forces	2.9.1 Intermolecular forces	1. Describe permanent dipole-permanent dipole forces, induced dipole-induced dipole forces, permanent dipole-induced dipole forces, ion-permanent dipole forces and hydrogen bonding. 2. Explain the effects of these forces on specific properties of molecular compounds: e.g. boiling points of simple hydrides in Groups 4, 5, 6 and 7 and variation of the boiling points along a homologous series of organic compounds. 3. Explain, in terms of intermolecular forces, the solubility of molecules in polar and non-polar solvents.
	2.9.2 Intermolecular hydrogen bonding	4. Explain the effects of these forces on the formation of carboxylic acid dimers; the open structure of ice leading to its density being lower than that of liquid water; and boiling points of water and alcohols.
	2.9.3 Intramolecular hydrogen bonding	5. Distinguish between intermolecular and intramolecular hydrogen bonding in 4-nitrophenol and 2-nitrophenol isomers in terms of melting point.

Subject Focus:	States of Matter
Learning Outcome 3	I can explain the Kinetic theory of gases with respect to ideal and real gases.

Topic	Sub-Topic	Assessment Criteria
3.1 Kinetic theory of gases		1. Sketch distribution diagrams showing speed of gas molecules at different temperatures. 2. Relate these diagrams to the concept of activation energy of chemical reactions. <i>The Zartmann experiment will not be examined.</i>
3.2 Ideal gases and the Ideal Gas Law		1. Recall the basic assumptions of the ideal gas model. 2. Explain the Ideal Gas Law. 3. Apply the Ideal Gas Law in the determination of relative molecular mass for gases using the evacuated flask method and volatile liquids using the gas syringe. 4. Solve problems involving the ideal gas equation.
3.3 Real gases and deviation from ideal behaviour		1. Explain deviations of real gases from ideality. 2. Relate the use of the Van der Waals equation to non-ideal behaviour. <i>The Van der Waals equation will be provided.</i> <i>No numerical questions will be set to test knowledge of Van der Waals equation.</i>
3.4 Dalton's Law of Partial Pressures		1. State Dalton's law of partial pressures. 2. Solve problems involving partial pressures using Dalton's law.
3.5 Vapour pressure and saturated vapour pressure		1. Describe the difference between vapour pressure and saturated vapour pressure. 2. Explain the relationship between saturated vapour pressure and temperature. 3. Explain vaporisation and fusion in terms of the kinetic molecular model.

Subject Focus:	Quantity of Matter
Learning Outcome 4	I can apply the mole concept in calculations.

Topic	Sub-Topic	Assessment Criteria
4.1 Moles of substance	4.1.1 The mole	1. Define the mole. 2. Describe matter in terms of atoms, molecules or ions. 3. Calculate in terms of molar quantities. 4. Convert between different units of volume – m^3 ; dm^3 or L; cm^3 or mL. 5. Calculate molar concentration in units of mol dm^{-3} or mol L^{-1} . 6. Calculate mass concentration in units of g dm^{-3} or g L^{-1} .
	4.1.2 Avogadro constant	7. Recall the value of the Avogadro constant. 8. Calculate the number of particles in a certain amount of substance using the Avogadro constant.
4.2 Empirical and molecular formulae		1. Determine the empirical formula of a compound from given data. 2. Determine the molecular formula of a compound from the empirical formula and molar mass.
4.3 Chemical equations	4.3.1 Balancing equations	1. Construct both full and ionic chemical equations. 2. Apply chemical equations for calculations involving reacting substances in terms of amounts and concentrations measured in moles and mol dm^{-3} respectively.
	4.3.2 Limiting reagent	3. Deduce the limiting reagent in a chemical reaction from given data. 4. Apply the concept of limiting reagent in calculations.
	4.3.3 Percentage yield	5. Calculate the percentage yield of a product in a chemical reaction. 6. State why chemical reactions do not always produce a 100% yield. 7. Calculate the percentage purity of a substance from given data.

Topic	Sub-Topic	Assessment Criteria
4.4 Titrimetric analysis	4.4.1 Titrations	1. Define primary standard and standard solution. 2. Recall the use of anhydrous sodium carbonate and hydrated ethanedioic acid as primary standards. 3. Describe the preparation of a standard solution. 4. Describe the experimental details of titrimetric analysis. <i>Limited to acid-base and redox titrations.</i>
	4.4.2 Back titrations	5. Carry out a titrimetric analysis (refer to Topic 16 for more details). 6. Describe the experimental details of a back titration. <i>Limited to acid-base and redox systems.</i> 7. Carry out a back titration involving acid-base or redox systems (refer to Topic 16 for more details).

Subject Focus:	Energetics
Learning Outcome 5	I can show an understanding of enthalpy changes, entropy and Gibbs free energy.

Topic	Sub-Topic	Assessment Criteria
5.1 Changes in energy	5.1.1 Exothermic and endothermic processes	1. Recognise the Joule (J) as the unit of energy. 2. Describe endothermic and exothermic energy changes for phase changes and chemical changes.
	5.1.2 Energy level (enthalpy) diagrams	3. Sketch energy level (enthalpy) diagrams for exothermic and endothermic processes.
5.2 Standard enthalpy changes		1. Define the following enthalpy changes: standard enthalpy change of reaction, formation, atomisation, ionisation, combustion, solution (or dissolution), neutralisation, solvation including hydration, electron affinity, bond enthalpy terms, bond dissociation enthalpies and lattice enthalpies. 2. Relate the enthalpy of solution to the lattice enthalpy and enthalpies of hydration. 3. Distinguish between the first and second electron affinity for di-negative ions. 4. Explain the difference between bond enthalpy terms and bond dissociation enthalpies. 5. Estimate the enthalpy change of a reaction using bond enthalpies. 6. Explain why enthalpy values obtained from bond enthalpy calculations are approximate.
5.3 Hess's law and the Born-Haber cycle.	5.3.1 Hess's law and its use in simple calculations	1. State Hess's Law. 2. Calculate enthalpy changes indirectly by the construction of simple enthalpy cycles or energy level diagrams from given thermochemical data.
	5.3.2 Born-Haber cycle	3. Construct Born-Haber cycles. 4. Calculate enthalpy changes indirectly by the construction of Born-Haber cycles from given thermochemical data.
5.4 Theoretical and experimental values of enthalpy changes		1. Explain the deviation between theoretical and experimental lattice enthalpies of silver halides. 2. Interpret the difference in terms of a degree of covalency in the bonding of this formally ionic structure.

Topic	Sub-Topic	Assessment Criteria
Topic	Sub-Topic	Assessment Criteria
		3. Explain the difference between expected and experimental values of enthalpies of formation or combustion of benzene and similar simple hydrocarbons as evidence of delocalised structures.
5.5 Calorimetry	5.5.1 Determination of enthalpy change of neutralization, solution and reaction	1. Plot temperature-time curves from the data obtained from calorimetry experiments. 2. Interpret temperature-time curves to determine temperature change. 3. Describe the relationship $\Delta H = -mc\Delta T$. 4. Calculate enthalpy changes of neutralization, solution and reaction from temperature changes and/or temperature-time curves using $\Delta H = -mc\Delta T$ 5. Calculate the enthalpy change of an acid-base reaction using $\Delta H = -mc\Delta T$.
	5.5.2 Simple thermometric acid-base titrations	6. Describe a simple thermometric acid-base titration. 7. Plot temperature-volume curves from the data obtained from simple thermometric acid-base titrations. 8. Determine the end-point from the temperature-volume curve.
5.6 Entropy and Gibbs free energy	5.6.1 The system and its surroundings	1. Distinguish between the concept of a system and its surroundings. 2. State that ΔH alone is not able to fully predict spontaneous change.
	5.6.2 Entropy	3. Describe entropy as a measure of disorder of a system. 4. Predict whether ΔS is positive, negative or nearly zero in simple changes exemplified by state change, combustion, dissolution, and dimerisation. 5. Calculate ΔS from absolute entropy values.
	5.6.3 Gibbs free energy	6. Explain that processes are spontaneous when the free energy change is negative. 7. Describe the relationship between free energy, enthalpy and entropy, namely, $\Delta G = \Delta H - T\Delta S$. 8. Calculate the temperature at which processes start or cease to be spontaneous by use of the equation $\Delta G = \Delta H - T\Delta S$.
	5.6.4 Kinetic versus thermodynamic stability	9. Explain the difference between the kinetic and thermodynamic stability of chemical processes.

Subject Focus:	Phase Equilibria
Learning Outcome 6	I can show an understanding of one-component systems, two-component systems, distillation, and osmosis.

Topic	Sub-Topic	Assessment Criteria
6.1 Phase changes for one-component systems		1. Draw pressure-temperature diagrams for the one-component systems: water and carbon dioxide. 2. Identify the boiling, melting, triple points and critical temperature in pressure-temperature diagrams. 3. Explain sublimation and supercritical fluids by using pressure-temperature diagrams. 4. Discuss the use of supercritical carbon dioxide as a 'green' solvents which can replace toxic organic solvents, e.g. in the food industry.
6.2 Two-component systems	6.2.1 Mixtures of two miscible liquids and Raoult's Law	1. Define Raoult's Law. 2. State examples of binary mixtures which are approximately ideal and which show both positive and negative deviations from Raoult's Law. 3. Calculate the vapour pressure of a component above an ideal binary mixtures using Raoult's law. 4. Calculate the composition of the vapour above an ideal binary mixture using Dalton's law. 5. Explain the enthalpy changes on mixing components of ideal and non-ideal mixtures. 6. Sketch pressure-composition and temperature-composition diagrams for approximately ideal mixtures and mixtures which show both positive and negative deviations from Raoult's Law. 7. Explain fractional distillation for ideal and non-ideal mixtures and the formation of azeotropic mixtures.
6.3 Immiscible liquids and steam distillation		1. Explain why a mixture of two immiscible liquids boils at a temperature below the boiling point of both liquids. 2. Describe the advantage of steam distillation over simple distillation.

Topic	Sub-Topic	Assessment Criteria
		3. Deduce an equation that relates the mass ratio of the substances in the vapour phase to the vapour pressures of the liquids and their relative molecular masses. 4. Calculate the mass of a compound obtained by steam distillation. 5. Calculate the relative molecular mass of an organic compound obtained from steam distillation.
6.4 Osmosis and osmotic pressure	6.4.1 Osmosis	1. Define the term colligative property of solutions. 2. Define osmosis as an example of a colligative property of solutions. <i>Questions on the determination of partial degree of dissociation/association will not be set.</i>
	6.4.2 Osmotic pressure	3. Define osmotic pressure, Π . 4. Calculate the variables in the equation $\Pi V = nRT$ where n is the amount, in moles, of particles of solute.
	6.4.3 Reverse osmosis	5. Explain reverse osmosis. 6. Relate reverse osmosis to local desalination of seawater. 7. Compare reverse osmosis to distillation in terms of energy considerations.

Subject Focus:	Chemical Equilibrium
Learning Outcome 7	I can show an understanding of the concept of chemical equilibrium.

Topic	Sub-Topic	Assessment Criteria
7.1 Concept of dynamic equilibrium	7.1.1 Characteristics of an equilibrium mixture	1. Identify the main characteristics of the equilibrium state, as constancy of concentrations/pressures resulting from the equality of the rates of the forward and reverse reactions.
	7.1.2 Equilibrium constants	2. Derive equilibrium constant in terms of concentrations and of pressures (K_c and K_p) for homogenous and heterogenous equilibria including units as appropriate. 3. Deduce the extent of reaction from the magnitude of the equilibrium constant. 4. Describe equilibrium constant expressions involving condensed (liquid or solid) phases in heterogenous equilibria. 5. Calculate K_c and K_p from provided data. 6. Calculate equilibrium partial pressures from K_p and other provided data. 7. Calculate equilibrium concentrations from K_c and other provided data.
7.2 Degree of dissociation		1. Derive an expression for the degree of dissociation of covalent molecular substances converting into smaller molecules. 2. Relate the degree of dissociation to K_c and K_p . <i>Problems requiring solving quadratic equations using the formula will not be set.</i>
7.3 Experimental methods of investigating chemical equilibrium		1. Explain how the equilibrium reaction between a carboxylic acid and an alcohol can be investigated experimentally.
7.4 Le Chatelier's Principle		1. Explain the qualitative effect on the equilibrium position and on the value of the equilibrium constant for changes in concentration, temperature and (where relevant) pressure on: i. wholly gaseous equilibria; ii. solid-gas equilibria; iii. equilibria in the liquid phase (e.g. esterification) or in solution.

Topic	Sub-Topic	Assessment Criteria
	7.4.1 Haber Process	2. Explain the effect of the introduction of an inert gas on an equilibrium mixture at constant volume. 3. Explain the effects of the presence of catalysts on equilibria. 4. Describe the principle using the Haber process for the manufacture of ammonia as an example. <i>Details of the manufacturing plant will not be tested.</i> 5. Explain why and how compromise conditions are required in the Haber process.
7.5 Distribution of a non-volatile solute between two immiscible solvents		1. Derive expressions for the partition constant (K_d) for a non-volatile solute between two immiscible solvents. 2. Explain single and multiple solvent extraction of a non-volatile solute. 3. Solve numerical questions on the partition constant related to solvent extraction, which will be limited to single and double extractions. <i>Discussion of the association and dissociation effects on partition equilibrium is not required.</i>

Subject Focus:	Ionic Equilibria
Learning Outcome 8	I can show an understanding of the concept of homogenous and heterogenous ionic equilibria.

Topic	Sub-Topic	Assessment Criteria
8.1 Acids and bases	8.1.1 Arrhenius, Brønsted-Lowry and Lewis definitions	1. Define Arrhenius, Brønsted-Lowry and Lewis acids and bases. 2. Explain conjugate pairs.
	8.1.2 Strong and weak acids and bases	3. Explain the proticity or basicity of acids and acidity of bases (e.g. sulfuric acid as a 'diprotic' or 'dibasic' acid and carbonate as a 'diacid' base). 4. Compare strong and weak acids; and strong and weak bases. 5. Explain acid and base strength in terms of their dissociation constants and pK values.
	8.1.3 K_a and K_b	6. Derive an expression for the dissociation constants K_a and K_b and include their units. 7. Explain with the aid of equations the amphoteric nature of water, hydrogencarbonate ion and hydrogensulfate(VI) ion. 8. Explain the stepwise dissociation of dibasic acids with an associated K_a value for each step. 9. Calculate K_a and K_b from provided data. 10. Calculate equilibrium concentrations of $[H^+]$ ions from K_a and other provided data. 11. Calculate equilibrium concentrations of $[OH^-]$ ions from K_b and other provided data.
	8.1.4 K_w	12. Derive an expression for the ionic product of water, K_w . 13. Convert K_a , K_b and K_w to pK values and vice versa. 14. Relate K_w to K_a and K_b .
	8.1.5 pH of aqueous solutions	15. Calculate the pH and pOH of aqueous solutions. 16. Calculate the $[H_3O^+]$ and $[OH^-]$ from the pH and vice versa.
	8.1.6 Hydrolysis of salts	17. Explain qualitatively the hydrolysis of salts including a description of the approximate pH of a solution of the salt. <i>Limited to: salts containing trivalent cations; salts of a strong acid and a weak base and vice versa.</i>

Topic	Sub-Topic	Assessment Criteria
Topic	Sub-Topic	Assessment Criteria
8.2 Indicators typified by phenolphthalein and methyl orange	8.2.1 Theory of indicators	1. Explain the use of indicators in acid-base titrations using their K_a values. 2. Assign appropriate indicators for titrations. 3. Recall the pH range of phenolphthalein and methyl orange.
	8.2.2 pH curves	4. Draw pH curves. 5. Interpret pH curves. <i>Limited to titrations between a monobasic acid and a monoacid base.</i>
8.3 Buffer solutions		1. Explain a buffer system. 2. Explain how a buffer system works to keep the pH fairly stable. 3. Explain applications of buffers. <i>Limited to blood and calibration of a pH meter.</i> 4. Calculate the pH of buffer solutions. 5. Calculate the composition of buffer solutions.
8.4 Role of solvent in equilibria involving ionisation of molecular solutes	8.4.1 Degree of ionisation	1. Calculate the degree of ionisation of weak electrolytes, which are acidic or basic. 2. Describe qualitatively the conductivity of solutions of strong and weak electrolytes. 3. Describe the connection between conductivity of a solution and the presence of dissolved ions.
8.5 Heterogeneous ionic equilibria	8.5.1 Solubility product K_{sp}	1. Derive solubility product expressions for sparingly soluble salts. 2. Relate K_{sp} to the molar solubility. 3. Solve mathematical equations involving K_{sp} and molar solubility of ionic sparingly soluble solutes of the type 1:1 (e.g. AgCl or $BaSO_4$).
	8.5.2 Factors influencing the solubility product equilibrium	4. Describe the effect on the solubility product equilibrium by addition (common ion effect) or removal of an ion involved in the equilibrium by acid-base or complex ion formation.

Subject Focus:	Redox Equilibria
Learning Outcome 9	I can show an understanding of what happens during redox reactions, their feasibility and possible applications of these reactions.

Topic	Sub-Topic	Assessment Criteria
9.1 Redox reactions		1. Deduce the oxidation number (state) of elements in different chemical species. 2. Explain the connection between oxidation number and bonding. <i>For example carbon in CO, CO₂ and hydrocarbons and oxygen in H₂O and H₂O₂.</i> 3. Construct balanced redox equations using half equations. 4. Explain disproportionation and comproportionation reactions. 5. State common oxidising agents and reducing agents covered in other parts of the syllabus.
9.2 Electrodes, galvanic cells and standard electrode potentials.		1. Explain electrode potentials. 2. Describe qualitatively how an electrode potential is set up in a half-cell. 3. Describe galvanic cells. 4. Describe the hydrogen electrode as a typical standard electrode.
9.3 Cell representation		1. Illustrate a galvanic cell using a schematic diagram. 2. Represent the cell diagram in standard notation such as: i. Zn(s) Zn²⁺(aq) Cu²⁺(aq) Cu(s) ii. Pt 2Br⁻(aq), Br₂(aq) Co³⁺(aq), Co²⁺(aq) Pt 3. Determine the emf of a galvanic cell. 4. Determine the anode and the cathode of a galvanic cell.
	9.3.1 The equation $\Delta G = -z E^{\circ}_{\text{rxn}} F$	5. Relate the free energy change of a reaction to its E° value using the equation $\Delta G = -z E^{\circ}_{\text{rxn}} F$ 6. Determine whether redox reactions are energetically feasible by using the electrode potentials of the reaction.

Topic	Sub-Topic	Assessment Criteria
9.4 Application of electrode potentials to predict redox change and reactivity	9.4.1 Electrochemical Series	1. Explain the electrochemical series. 2. Predict the reactivity of elements and molecules/ions depending on their position in the electrochemical series.
	9.4.2 Corrosion	3. Explain corrosion as an electrochemical process. 4. Explain the use of sacrificial protection to prevent rusting. 5. Predict redox change under non-standard conditions using the Nernst Equation. <i>The Nernst equation will be provided where its use is required.</i>
	9.4.3 Fuel cells	6. Describe the theory of fuel cells using the hydrogen – air cell as an example. <i>Operational details and construction of fuel cells will not be tested.</i>

Subject Focus:	Reaction Kinetics
Learning Outcome 10	I can show an understanding of the rate of a chemical reaction and the different factors that influence it.

Topic	Sub-Topic	Assessment Criteria
10.1 The experimental investigation of reaction rate for simple reactions		1. Describe suitable methods by which the rate of a chemical reaction can be determined. <i>The methods that will be tested are:</i> - gas volume changes by gas syringe; - colorimetry; - conductivity measurements; - changes of pH for (slow) reactions involving the consumption or formation of hydroxonium ions; - titrimetry for (slow) reactions. 2. Describe the relative merits and demerits of each method. 3. Interpret the given data obtained from each method. <i>This may include calculations and plotting of graphs.</i>
10.2 The rate equation	10.2.1 Initial rates method	1. Define the term order of reaction. 2. Define the term rate constant. 3. Deduce the order of reaction with respect to a reactant from the data provided. <i>This will be limited to zero, first and second order.</i> <i>Fractional and pseudo orders will not be tested.</i> 4. Calculate the overall order of a chemical reaction. 5. Derive the rate equation of a chemical reaction. <i>Integrated rate equations will not be tested.</i> 6. Deduce the unit of the rate constant of a chemical reaction.

Topic	Sub-Topic	Assessment Criteria
	10.2.2 Half-life for first order reactions	7. Calculate the variables in the rate equation. 8. Present kinetic data graphically, which includes plots for concentration against time, rate against concentration and log rate against log concentration. 9. Interpret graphical representations of kinetic data. 10. State that the half-life of a first order reaction is constant. 11. Determine the half-life of a first order reaction from given data and graphs. 12. Sketch the decay curve. 13. Plot the decay curve from given data. 14. Determine the half-life of a first order reaction from a plotted decay curve.
10.3 Factors affecting the rate of a chemical reaction	10.3.1 The collision theory	1. Describe the criteria that lead to an effective collision between reactant particles. 2. Explain the effect of the following factors on the rate of a chemical reaction: pressure; concentration; surface area; temperature and catalysts.
	10.3.2 The Arrhenius equation	3. Define the term activation energy. 4. Describe the significance of the pre-exponential factor in the Arrhenius equation. 5. Relate the temperature to the rate of a chemical reaction as represented by the Arrhenius equation. 6. Relate the activation energy to the rate of a chemical reaction as represented by the Arrhenius equation. <i>The equation will be provided when required and numerical calculations will not be set.</i>
	10.3.3 Maxwell-Boltzmann distribution of energies	7. Sketch on the same pair of axes, a graph showing the distribution of energies at temperature T and at T+10K. 8. Interpret the relationship between temperature and rate of reaction as expressed by the graphs. 9. Relate the areas under the graph with the number of particles that may or may not be involved in an effective collision. 10. Relate the area under the graph with the number of particles that may be involved in a successful collision without and with a catalyst.

Topic	Sub-Topic	Assessment Criteria
10.4 The relationship between the rate law and the mechanism of a chemical reaction	10.4.1 Multistep reactions and the rate-determining step	1. Describe examples of mechanisms with a slow step followed by a faster step/s. <i>Limited to S_N1 for halogenoalkanes and to the reaction between CO and NO_2.</i> 2. Describe mechanisms with a fast reversible first step followed by a slow step. <i>Limited to gaseous reactions of NO with O_2.</i> 3. Identify the rate-determining step of a chemical reaction of any other given mechanism from its rate equation. 4. Determine the order of a chemical reaction from its mechanism.
	10.4.2 Molecularity	5. Define the molecularity of a reaction. 6. Determine the molecularity of the rate-determining step. 7. Explain the difference between the order and the molecularity of a chemical reaction.
	10.4.3 Third order reactions	8. Describe the improbability of three-particle collisions.
	10.5 Photochemical reactions and free radical mechanisms	1. Identify the initiation, propagation and termination steps of free radical mechanisms. 2. Describe the free radical mechanism of the monohalogenation of alkanes and of the polymerisation of ethene.
10.6 Catalysis	10.6.1 Homogeneous and heterogeneous catalysis	1. Define homogeneous and heterogeneous catalysts. 2. Describe the action of a homogeneous catalyst as described by the intermediate compound formation theory. 3. Illustrate the action of a heterogeneous catalyst as described by adsorption theory. 4. Describe the role of catalysts in speeding up a reaction by providing an alternative pathway that is characterised by a lower activation energy. 5. Explain the lack of influence of catalysts on the equilibrium composition of a reversible reaction. 6. Sketch labelled enthalpy level diagrams (energy profiles) for reactions with and without catalysts.

Topic	Sub-Topic	Assessment Criteria
	10.6.2 Catalysts in industrial processes	7. Describe the role of the catalytic converter in changing unburnt hydrocarbons, CO and nitrogen oxides in car exhaust into less harmful products. <i>Details of different types of converters will not be tested.</i>
	10.6.3 Catalyst poisoning	8. Define the term catalyst poisoning. 9. Describe catalyst poisoning in the catalytic converter and in the use of the Lindlar catalyst.
	10.6.4 Autocatalysis	10. Define the term autocatalysis. 11. Describe autocatalysis in the chemical reaction between acidified manganate(VII) and ethanedioate ions. 12. Sketch a graph of concentration against time for an autocatalytic reaction. 13. Interpret a plot of concentration against time for an autocatalytic reaction.

Subject Focus:	Chemical Periodicity
Learning Outcome 11	I can show an understanding of the periodic relationships amongst period 3 elements and their compounds.

Topic	Sub-Topic	Assessment Criteria
11.1 Periodic relationships amongst the elements Li to Ar		- Recall that periodic classification of elements is done in terms of electronic structure. - Describe how properties of period 2 and period 3 elements exhibit periodic behaviour as illustrated by their trends in melting and boiling points, electrical conductivity and ionisation energy.
11.2 Periodic relationships amongst the oxides and hydroxides of period 3 elements		- Recall the formulae of the oxides of period 3 elements. - Describe the physical state of the oxides of period 3 elements. - Relate the physical state of the oxides of period 3 elements to their structure and bonding. - State the acid-base character of the oxides and hydroxides of metals and the oxides of non-metals. <i>This is limited to the following: Na_2O, MgO, Al_2O_3, NaOH, Mg(OH)_2, Al(OH)_3, SiO_2, P_4O_{10}, SO_2, SO_3 and Cl_2O.</i>
11.3 Periodic relationships amongst the chlorides of the elements Na to P		- Recall the formulae of the chlorides of the elements Na to P. - Describe the physical state of the chlorides of the elements Na to P. - Relate the physical state of the chlorides of the elements Na to P to their structure and bonding. - Describe the chemical reactions of the chlorides of the elements Na to P with water.
11.4 Periodic relationships amongst the hydrides of period 3 elements		- Recall the formulae of the simple hydrides of period 3 elements. - Describe the physical state of the hydrides of period 3 elements. - Relate the physical state of the hydrides of period 3 elements to their structure and bonding. <i>Excluding aluminium hydride.</i>

Subject Focus:	Inorganic Chemistry
Learning Outcome 12	I can show an understanding of the physical and chemical properties of a number of inorganic elements and their compounds. N.B. In exam questions involving preparations, only reagents, conditions and the relevant equations will be required.

Topic	Sub-Topic	Assessment Criteria
12.1 S-block chemistry <i>Note: Group 1 elements will be limited to lithium, sodium and potassium. Group 2 elements include beryllium, magnesium, calcium, strontium and barium.</i>		- Describe the physical properties (melting points and density) of the elements. - Describe the oxidation states and electronic configuration in terms of orbitals for s-block elements. - Describe the flame test, including a reference to the flame colour for: Li, Na, K, Ca, Sr and Ba. - Describe the reaction of the elements with water and with oxygen. - Explain the reasons for diagonal relationships for: Li and Mg; Be and Al; including examples to illustrate the relation. - Draw structures, including any covalent bonds and any delocalization features, of the following ions: oxides, peroxides, superoxides, hydroxides, nitrides, carbonates, hydrogencarbonates, chlorides, nitrates(V), nitrates(III), sulfates(VI) and sulfates(IV). <i>Delocalisation will be limited to carbonate, nitrate(V) and sulfate(VI) ions.</i> - Describe the effect of heat on nitrates(V), hydrogencarbonates and carbonates.
	12.1.1 Trends in thermal stability	8. Explain trends in thermal stability for nitrates(V) and carbonates.
	12.1.2 Trends in solubility	9. Explain trends in solubility on the basis of hydration and lattice enthalpies for hydroxides and sulfates(VI).
12.2 Aspects of the chemistry of aluminium		- Describe the amphoteric nature of the metal and its oxide and hydroxide. - Explain the passivation of the metal. - Draw the structure of the hydrated ion. - Explain the acidity of the hydrated ion. - Explain the structure of the dimeric form of gaseous Al_2Cl_6 and its equilibrium with the monomeric form. - Recall alums as double salts of general formula $\text{M}^{\text{I}}\text{M}^{\text{III}}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ where M^{III} could be an ion other than aluminium. <i>Preparation of alums is not required.</i>

Topic	Sub-Topic	Assessment Criteria
Topic	Sub-Topic	Assessment Criteria
12.3 The chemistry of carbon, silicon, tin and lead	12.3.1 Allotropy of carbon	1. Draw the structure of graphite and diamond. 2. Describe the structure of the allotropes of carbon (graphite, diamond and fullerenes typified by C_{60}). 3. Compare the properties of the macromolecular allotropes (graphite and diamond) to the simple molecular fullerenes (typified by C_{60}).
	12.3.2 Main oxidation states and inert pair effect	4. Predict the relative stability of the two main oxidation states on descending the group. 5. Explain the redox character of simple compounds of Sn(II) and Pb(IV) as exemplified by $SnCl_2$ and PbO_2 .
	12.3.3 Principal oxides	6. Describe the bonding in CO , CO_2 , SiO_2 , PbO and PbO_2 and the acid/base character of these oxides.
	12.3.4 Halides	7. Describe the bonding in di- and tetrachlorides. 8. Describe the action of water for the tetra- and dichlorides. <i>Lack of hydrolysis of CCl_4 should be explained in terms of steric hindrance.</i>
	12.3.5 Toxicity of lead	9. Explain the use of tetraethyllead in petrol engines and why it has been phased out.
12.4 The chemistry of fluorine, chlorine, bromine and iodine	12.4.1 Simple tests for these elements	1. Explain the bleaching property of chlorine and bromine. 2. Describe the colour of solutions of bromine and iodine in organic solvents and the starch test for iodine.
	12.4.2 Trends in physical and chemical properties	3. Explain the trends in boiling points and bond dissociation enthalpies for the series X_2 and HX ($X = F, Cl, Br$ and I). 4. Explain the trends in the first ionisation energy and the first electron affinity of the halogens.
	12.4.3 Reactions of ionic halides	5. Describe reactions of ionic chlorides, bromides and iodides with silver ions, phosphoric(V) and sulfuric(VI) acids. <i>The reaction with concentrated sulfuric(VI) acid should be explained in relation to the reducing properties of X^- and HX.</i> 6. Explain why the solubility of AgX ($X = Cl, Br, I$) is influenced in the presence of ammonia.

Topic	Sub-Topic	Assessment Criteria
		7. Describe the formation of the corresponding silver(I) complexes.
Topic	Sub-Topic	Assessment Criteria
	12.4.4 Acidity of HX and anomalous behaviour of HF	8. Explain the trend in acid strength for HX (X = F, Cl, Br, I) in aqueous media through bond dissociation enthalpy. 9. Explain the hydrogen bonding in HF and in the ion HF_2^- .
	12.4.5 Oxidising properties of the halogens	10. Calculate the oxidation states of the halogens Cl, Br and I in their compounds. 11. Explain the trend in the oxidising strength of the halogens X_2 (X = Cl, Br, I). 12. Write equations for the displacement reactions for X_2 and disproportionation reactions of chlorine. 13. Write equations describing the disproportionation reactions of chlorate(I) (OCl^-).
	12.4.6 Oxo anions	14. Describe the structures of chlorate(V) and chlorate(VII) ions. 15. Describe the effect of heat on KClO_3 and the oxidation of iodide by chlorate(V). 16. Explain the use of chlorate(V) and chlorate(VII) as oxidising agents in pyrotechnics. 17. Recall that chlorate(V) and chlorate(VII) are toxic.
	12.4.7 Iodine in titrimetry	18. Describe the reaction of iodine (prepared by the oxidation of iodide with Cu(II) or iodate(V)) with thiosulfate ions and its use in redox titrimetry.
	12.4.8 Triiodide ion	19. Describe the solubility of iodine in potassium iodide solution.
12.5 The chemistry of transition metals	12.5.1 d-block elements and transition metals	1. Recall the 3d block elements. 2. Write the electronic configurations of the metals Sc to Zn (given their atomic number). 3. Discuss the difference between d-block elements and transition metals. <i>A transition metal is a d-block element that forms at least one species in which the d-orbital is incompletely filled (Sc to Cu).</i> <i>Candidates should be able to write the electronic configuration of the d-block elements in their common ions and oxidation states.</i>

Topic	Sub-Topic	Assessment Criteria
	12.5.2 Characteristic properties of transition metals	4. Illustrate the metallic character, variable oxidation states (explained on the basis of the electron configuration of these elements), catalytic properties, and formation of coloured complex ions. <i>The colour is explained through electronic transitions of d-electrons between the ground state and an excited state when light is absorbed. If some wavelengths of light are absorbed, then the other wavelengths will be transmitted or reflected.</i>
	12.5.3 Transition metal complexes	5. Describe the bonding in transition metal complexes in terms of dative covalency between ligands and the central metal atom/ion. <i>Candidates should be able to classify ligands as monodentate, bidentate and hexadentate from the formula or structure of the ligand or complex and identify that the total number of bonds from the ligands to the central transition metal atom/ion is known as the co-ordination number.</i> 6. Describe the structure of ethane-1,2-diamine, ethanedioate and aminoethanoate ion as bidentate ligands. 7. Recall that EDTA^{4-} is a hexadentate ligand. <i>The structure of EDTA^{4-} will not be examined.</i>
	12.5.4 Stability of complex ions and ligand exchange	8. Explain why polydentate ligands easily displace monodentate ligands to form more stable chelate complexes.
	12.5.5 Nomenclature of complex ions	9. Name complexes that contain the following monodentate and bidentate ligands: water, ammonia, hydroxide, halides, cyanide, carbon monoxide, nitrogen(II) oxide, thiocyanate, ethane-1,2-diamine, ethanedioate and aminoethanoate ion. <i>The naming of complexes is to be carried out using the publication Chemical Nomenclature for Use in Matriculation Examinations (2021) as a guide.</i>
	12.5.6 Shapes of di-, tetra-, and hexa-coordinated systems	10. Describe the structure of complex ions for the: - linear complexes $[\text{Cu}(\text{NH}_3)_2]^+$, $[\text{CuCl}_2]^-$ and $[\text{Ag}(\text{NH}_3)_2]^+$; - tetrahedral complexes $[\text{CoCl}_4]^{2-}$, $[\text{FeCl}_4]^-$ and $[\text{CuCl}_4]^{2-}$; - square planar complexes $[\text{Cu}(\text{CN})_4]^{2-}$;

- octahedral complexes of transition metals with monodentate and bidentate ligands.
Candidates are expected to include drawings of the structures as part of the description.

Topic	Sub-Topic	Assessment Criteria
	12.5.7 Isomerism in complex ions	11. Describe structural and cis/trans isomerism in octahedral complexes for hydrates of CrCl_3 and optical isomerism in octahedral complexes with three bidentate ligands. <i>Candidates are expected to include drawings of the structures as part of the description.</i>
	12.5.8 Coloured complex ions	12. Describe the dependency of the colour of complex ions on: - the transition metal ion. Exemplified by the different colours of the hexaaquacomplexes of Fe(II), Co(II) and Cu(II) ions; - the oxidation state of the transition metal ion. Exemplified by the different colours of Fe(II), Fe(III) and Cr(II), Cr(III) hexaaquacomplexes; - the type of ligands bonded to the transition metal ion. Exemplified by aqua, ammine and chloro complexes of Cu(II), $[\text{Fe}(\text{H}_2\text{O})_5(\text{NO})]^{2+}$ and $[\text{Fe}(\text{H}_2\text{O})_5(\text{SCN})]^{2+}$. 13. Describe the absence of colour of complexes of Cu(I) and Ag(I) due to their d^{10} configuration.
	12.5.9 Ligand exchange reactions	14. Describe ligand exchange reactions for similar sized ligands such as the replacement of water ligands with ammonia ligands in Cu(II) and Co(II) ions. 15. Describe incomplete ligand exchange reactions, limited to the formation of $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$, $[\text{Fe}(\text{H}_2\text{O})_5(\text{NO})]^{2+}$ and $[\text{Fe}(\text{H}_2\text{O})_5(\text{SCN})]^{2+}$. 16. Describe the exchange of water ligands with the larger chloride ligands in Cu(II) and Fe(III) ions which brings about a change in coordination number. 17. Write equilibrium reactions for the stepwise stability constants of ligand exchange reactions and vice-versa.
	12.5.10 Stability of chelating ligands	18. Infer the overall stability constant for a ligand exchange reaction from the stepwise stability constants. 19. Predict ligand exchange from values of stability constants. 20. Explain that polydentate ligands replace monodentate ligands due to the chelate effect. <i>Such effect has to be explained in terms of the balance between entropy and enthalpy changes in these reactions.</i>

Topic	Sub-Topic	Assessment Criteria
	12.5.11 Transition metal compounds	21. Describe haem as an iron(II) complex with a polydentate ligand. 22. Explain the ability of Fe(II) in blood haemoglobin to form a co-ordinate bond with oxygen and carbon monoxide and hence the toxicity of the latter. 23. Explain the non-existence of iron(III) iodide and iron(III) carbonate. 24. Explain the acidity of hexaaquairon(III) ion and compare with that of the corresponding iron(II) species. 25. Describe how $[\text{Fe}(\text{SCN})(\text{H}_2\text{O})_5]^{2+}$ is involved in qualitative analysis for Fe(III). 26. Describe how Cu_2O is formed in the Fehling's test for aldehydes and how Cu^{2+} can be converted into CuCl and CuI. 27. Describe the instability of $\text{Cu}^+(\text{aq})$ ion with respect to disproportionation and its stability with ammonia and chloride ligands. 28. Explain the use of acidified manganate(VII) as a strong oxidising agent in titrimetry, inorganic and organic chemistry. 29. Describe the reduction of manganate(VII) in alkaline solution to form the dark brown solid, MnO_2. 30. Describe the chromate(VI)-dichromate(VI) equilibrium and the effect of pH change on this equilibrium. 31. Recall that Cr(VI) compounds are carcinogenic.
12.6 The chemistry of hydrogen	12.6.1 Manufacture of hydrogen	1. Describe the manufacture of the element from natural gas by steam reformation and by electrolysis of acidified water.
	12.6.2 Laboratory preparation	2. Explain methods of preparation of hydrogen on a small scale using reactions involving metal and acid or aluminium and alkali.
	12.6.3 Properties as reducing agent	3. Describe reactions of hydrogen acting as a reducing agent with non-metals (oxygen, nitrogen and sulfur) and metal oxides as well as with organic molecules, e.g. unsaturated hydrocarbons.

12.6.4 Ionic, covalent and interstitial hydrides

- Describe the type of bonding in salt-like hydrides (containing H^- ion) and covalent hydrides.
- Identify the connection between interstitial hydrides and storage of hydrogen and its potential in the hydrogen economy.

Topic	Sub-Topic	Assessment Criteria
	12.6.5 Complex hydrides	- Describe the structure and bonding in $LiAlH_4$ (lithium tetrahydridoaluminate). - Describe the use of $LiAlH_4$ as a reducing agent in organic chemistry for the $C=O$ and the $COOH$ groups but which does not reduce the $C=C$ double bond. - Describe the use of $NaBH_4$ (sodium tetrahydridoborate) as a reducing agent for the $C=O$ group.
	12.6.6 Water: its physical properties and solvent action	Explain the physical properties of water in terms of hydrogen bonding and its solvent properties in terms of its polarity.
12.7 The chemistry of nitrogen	12.7.1 Structure and bonding of N_2	- Explain the lack of reactivity of nitrogen as a consequence of the strength of the triple bond. - Describe the reaction of atmospheric nitrogen with atmospheric oxygen at high temperatures to form NO. - Describe the industrial reaction of nitrogen with hydrogen. <i>Refer to Sub-Topic 7.4.1.</i> - Describe the reactions of nitrogen with Li and group 2 elements Mg to Ba.
	12.7.2 Ammonia and its reactions as base, ligand and reducing agent	Describe the properties of ammonia as a weak Lowry-Brønsted base, a good monodentate ligand (with Cu^{2+}, Zn^{2+}, Ni^{2+} in qualitative analysis) and a reducing agent.
	12.7.3 Ammonium salts	Explain the thermal instability of ammonium nitrate(III) and ammonium nitrate(V) as resulting from the redox reaction between the reducing cation and the oxidising anion.
	12.7.4 Oxides of nitrogen N_2O , NO , NO_2 , N_2O_4	- Draw electronic structures and bonding in these oxides showing the presence of odd electron systems where relevant. - Illustrate the molecular shapes of these oxides. - Describe the disproportionation of NO_2 when dissolved in water or alkali.
	12.7.5 Nitric(III) (nitrous) acid and salts	Describe the in situ preparation of nitric(III) acid and nitrates(III) of s-block elements.

	12.7.6 Nitric(V) (nitric) acid and salts	11. Describe the chemistry for the preparation of nitric(V) acid by reaction of a non-volatile acid on nitrates(V). 12. Describe reactions of nitric(V) acid acting as an acid, an oxidising agent and a nitrating agent in organic chemistry.
Topic	Sub-Topic	Assessment Criteria
		13. Describe the brown ring test for nitrates(V). <i>Description to include explanation of reactions taking place.</i> 14. Describe the use of Devarda's alloy or aluminium in alkali to test for nitrates(V). <i>Description to include explanation of reactions taking place.</i>
12.8 The chemistry of oxygen	12.8.1 Allotropes of oxygen: O ₂ and O ₃ (trioxygen or ozone)	1. Describe that oxygen is a main component of the atmosphere. 2. Describe a test for oxygen in the lab. 3. Explain how trioxygen forms in the stratosphere from photochemical reactions and its action as a screen of ultraviolet radiation. 4. Recall that trioxygen forms in the lower atmosphere as an air pollutant formed from other gaseous pollutants. <i>Chemistry of formation will not be tested.</i>
	12.8.2 Preparation of oxygen	5. Describe the preparation of oxygen gas from lead(IV) oxide and hydrogen peroxide.
	12.8.3 Reaction of oxygen with metals and non-metals	6. Describe the preparation of compounds of oxygen by direct union between oxygen and various elements. <i>For oxides stated in other parts of the syllabus.</i> 7. Classify oxides as acidic, basic, amphoteric or neutral. <i>For oxides stated in other parts of the syllabus.</i> 8. Distinguish between normal and mixed oxides (typified by iron(II,III) oxide and lead(II,IV) oxide) in terms of the species present in these compounds.
	12.8.4 Hydrogen peroxide and peroxides	9. Describe the chemistry and conditions required for the preparation of hydrogen peroxide from barium peroxide.

10. Describe the main properties of hydrogen peroxide namely, as oxidising agent, as a reducing agent with oxidants stronger than itself (limited to manganate(VII) and dichromate(VI)), and as a thermally unstable compound.

Topic	Sub-Topic	Assessment Criteria
	12.8.5 Superoxide ion	11. Explain the relatively weak nature of the peroxide bond and the ease of its homolysis to form radicals. 12. Illustrate the structure and bonding of the peroxide ion. 13. Describe the reaction of the peroxide ion with water, acid and with carbon dioxide. 14. Illustrate the structure and bonding of the superoxide ion. 15. Describe the reaction of potassium superoxide with water and its use in rebreathers.
12.9 The chemistry of sulfur	12.9.1 Allotropes of sulfur (S_8)	1. Describe the effect of temperature on the stability of rhombic and monoclinic sulfur. <i>Knowledge of crystal shapes is not required.</i>
	12.9.2 Hydrogen sulfide	2. Recall the rotten-egg odour of hydrogen sulfide and its toxicity. 3. Describe how hydrogen sulfide can be prepared from sulfides. 4. Explain the properties of hydrogen sulfide as a weak diprotic acid and a reducing agent.
	12.9.3 Sulfur(IV) oxide, sulfuric(IV) acid, sulfates(IV)	5. Describe how sulfur(IV) oxide, sulfuric(IV) and sulfates(IV) can be made and interconverted in the laboratory. 6. Explain the reducing character of sulfur(IV) oxide, sulfuric(IV) and sulfates(IV).
	12.9.4 Sulfur(VI) oxide and sulfuric(VI) acid	7. Describe how sulfur(VI) oxide, sulfuric(VI) and sulfates(VI) can be made and interconverted in the laboratory. 8. Describe reactions of concentrated sulfuric(VI) acid as a nonvolatile proton donor capable of displacing more volatile acids from their salts, an oxidising agent, a strong dehydrating agent and a sulfonating agent in organic chemistry.
	12.9.5 Structure of oxoanions of sulfur	9. Draw the electronic structure and shape of the ions: sulfate(IV), sulfate(VI) and thiosulfate.

12.9.6 Thiosulfate

10. Explain the fact that the two sulfur atoms in thiosulfate are present in different structural environments and the implication of this on the mean oxidation state of the sulfur atoms.

The individual oxidation states of the sulfur atoms will not be required.

11. Describe the reaction of the thiosulfate ion with acids, iodine and chlorine.

Topic**Sub-Topic****Assessment Criteria**

12.9.7 Sulfur oxides as environmental pollutants

12. Explain why sulfur oxides are often present in the atmosphere as contaminants, their origin in fossil fuels and their role in the formation of acid rain and its environmental impact; the importance of using low-sulfur fuels as a means of reducing atmospheric sulfur oxides.

Techniques of desulfurisation of fuels are not required.

12.10 Qualitative analysis of inorganic compounds and ions

Note: Limited to reactions of elements and compounds mentioned in Learning Outcome 12.

1. Deduce inferences from given observations of inorganic reactions.

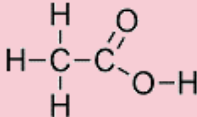
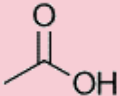
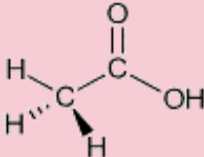
2. Identify elements/compounds/ions from inferences.

3. Describe inorganic reactions from given observations.

Subject Focus: Organic Chemistry – General Principles

Learning Outcome 13

I can show an understanding of the structure and properties of different organic molecules.

Topic	Sub-Topic	Assessment Criteria
13.1 Functional groups and homologous series	13.1.1 Types of formulae	1. Deduce possible structural formulae from molecular formulae. 2. Apply structural formulae as follows: i. condensed formulae such as CH_3COOH; ii. display formulae where individual bonds are shown explicitly, such as:  iii. skeletal formulae, such as:  <i>Unless specified:</i> - structural formulae of organic compounds can be given as described in 13.1.1.2. - for compounds containing cyclic structures, the cyclic part can be represented using the skeletal formula and the aliphatic part can be represented by any formula of the above. 3. Apply wedge and dashed/hatched line convention to represent three-dimensional structure, such as: 

Topic	Sub-Topic	Assessment Criteria
13.2 Catenation and spatial distribution in carbon compounds	13.1.2 Functional groups	4. Identify functional groups in the following classes: alkenes, alkynes, arenes, alcohols, ethers, phenols, halogenoalkanes and halogenoarenes, aldehydes, ketones, carboxylic acids and derivatives, amines and diazonium compounds.
	13.1.3 Nomenclature of organic compounds	5. Deduce systematic names of organic compounds from their structure and vice versa (refer to publication Chemical Nomenclature for Use in Matriculation Examinations (2021)).
	13.2.1 Catenation	1. Explain why the element carbon is unique in exhibiting catenation.
13.3 Isomerism in organic compounds	13.2.2 Spatial distribution of the bonds in simple carbon compounds	2. Explain the shape of the bonds and bond angles in alkanes (tetrahedral), alkenes and carbonyl groups (planar) and alkynes (linear) in terms of hybridised sp^3 , sp^2 and sp orbitals.
	13.3.1 Structural isomerism	1. Describe different types of structural isomerism. <i>Limited to: chain, positional, functional and keto-enol tautomerism.</i>
	13.3.2 Stereoisomerism	2. Describe stereoisomerism namely cis-trans (geometrical) (restricted to C=C and C=N systems), optical isomerism (restricted to enantiomerism) and conformational isomerism (restricted to cyclohexane). <i>The E/Z notation for geometrical isomers will not be examined.</i>
	13.3.3 Optical isomerism	3. Describe the property of optical activity in terms of dextrorotatory (+) and laevorotatory (-) compounds and their effect on plane polarised light. 4. Describe chiral molecules as those that do not possess a plane of symmetry. <i>The use of the descriptors D and L and R and S is not required.</i> <i>Descriptions will be limited to molecules having one chiral centre.</i>
	13.3.4 Drawing of isomers	5. Identify one or more chiral centres in a given organic molecule. 6. Explain the term racemate/racemic mixture of optical isomers in terms of external compensation. <i>Resolution of racemic mixture will not be tested.</i> 7. Draw structural and stereoisomers listed in 13.3.1 and 13.3.2. <i>For optical isomers, the wedge and dashed/hatched line convention should be used.</i>

Topic	Sub-Topic	Assessment Criteria
	13.3.5 Properties of isomers	8. Explain how size, shape and polarity of isomers can influence physical properties. 9. Compare the volatility of linear and branched isomers. 10. Describe the effect of intramolecular versus intermolecular hydrogen bonding on boiling point as exemplified by the nitrophenols.
13.4 Delocalisation of electrons in organic molecules		1. Identify the presence of delocalisation in conjugated systems. 2. Explain how delocalisation lowers the energy of the system and is thus a favourable structural feature. 3. Explain the concept of delocalisation in benzene.
13.5 Redox equations in organic chemistry		1. Construct oxidation and reduction equations using [O] and [H] respectively unless specifically asked to give half equations.

Subject Focus:	Chemistry of Aliphatic and Aromatic Compounds
Learning Outcome 14	I can show an understanding of the properties, preparation and reactions of a number of organic compounds. N.B. In exam questions involving preparations, only reagents, conditions and the relevant equations will be required. Where details of a laboratory preparation are required, they would be explicitly asked for. Chemical equations are not required in reaction schemes of organic compounds unless otherwise stated.

Topic	Sub-Topic	Assessment Criteria
14.1 Petroleum hydrocarbons	14.1.1 Sources of petroleum hydrocarbons	1. Recall that alkanes, alkenes and aromatic hydrocarbons can be obtained from petroleum by fractional distillation. <i>Details of the industrial process, the different fractions and their uses are not required.</i>
	14.1.2 Combustion of petroleum hydrocarbons	2. Explain how the combustion reactions of alkanes and other hydrocarbons lead to their use as fuels in industry, in the home and in transport. 3. Explain how the combustion of hydrocarbons leads to the formation of carbon dioxide, an important greenhouse gas.
	14.1.3 Environmental considerations	4. Explain the formation of pollutants: carbon monoxide, oxides of nitrogen and unburnt hydrocarbons arising from the internal combustion engine. 5. Explain how catalysts minimise the effect of pollutants mentioned in 4. above.
14.2 Alkanes, alkenes and alkynes	14.2.1 Alkanes and cycloalkanes	1. Recall the general unreactivity of alkanes.
	14.2.2 Substitution reaction	2. Describe free radical substitution by chlorine and bromine using appropriate chemical equations.
	14.2.3 Ring strain	3. Describe how ring strain in molecules, exemplified by cyclopropane, leads to enhanced reactivity relative to corresponding non-cyclic alkanes.
	14.2.4 Alkenes as unsaturated hydrocarbons	4. Describe alkenes as unsaturated hydrocarbons containing a double bond made up of a σ and a π bond.
	14.2.5 Preparation	5. Explain the double bond as an electron rich centre susceptible to electrophilic attack. 6. Describe the preparation of alkenes from alcohols and halogenoalkanes through elimination reactions.
	14.2.6 Addition reactions	7. Describe the reaction of alkenes with hydrogen, halogens, hydrogen halides, and water in the presence of sulfuric(VI) acid.

Topic	Sub-Topic	Assessment Criteria
Topic	Sub-Topic	Assessment Criteria
	14.2.7 Markovnikov's rule	8. Deduce the structure of the major product in the addition of halogen halides and water with an unsymmetrical alkene. 9. Explain the formation of the major product in terms of relative carbocation stability.
	14.2.8 Ozonolysis	10. Describe ozonolysis as the addition reaction of ozone to a double bond to produce an unstable ozonide. 11. Describe reductive ozonolysis as a reaction producing carbonyl compounds as products of the reduction of the ozonide. 12. Deduce the location of the double bond and the structure of the alkene from the products of the reductive ozonolysis reaction.
	14.2.9 Oxidation of the double bond	13. Describe the reaction with cold, dilute, alkaline manganate(VII) ions to form the diol.
	14.2.10 Addition polymerisation	14. Deduce the repeating unit of a polymer obtained from a given monomer. 15. Draw the repeating unit for poly(ethene), poly(propene), poly(chloroethene), poly(phenylethene). 16. Recall the preparation of low density and high density poly(ethene) from thermal and catalytic polymerisation. <i>Details of conditions are not required.</i> 17. Relate the uses of low density and high density poly(ethene) to their structure.
	14.2.11 Poly(alkene)s as waste products	18. Explain the lack of reactivity and non-biodegradability of poly(alkene)s and hence the accumulation in the environment.
	14.2.12 Alkynes	19. Describe alkynes as unsaturated hydrocarbons containing a triple bond made up of a σ and two π bonds.
	14.2.13 Preparation	20. Describe the preparation of alkynes by dehydrohalogenation reactions. 21. Describe the preparation of ethyne from calcium carbide. <i>Details of the preparation of calcium carbide are not required.</i>

Topic	Sub-Topic	Assessment Criteria
	14.2.14 Addition reactions	22. Describe the reactions with hydrogen, halogens and hydrogen halides, to form partial or fully saturated products and with water to form carbonyl compounds via tautomerisation of the enol. <i>Including the importance of chloroethene as a feedstock for poly(chloroethene) (PVC) manufacture.</i>
	14.2.15 Combustion	23. Describe the exothermic reaction of ethyne with oxygen and its use in oxy-acetylene flame.
	14.2.16 Acidity of terminal alkynes	24. Describe the reaction of terminal alkynes with ammoniacal silver nitrate and copper(I) chloride.
14.3 Arenes: benzene and alkylbenzenes	14.3.1 Structure and bonding	1. Describe the structure and bonding in benzene.
	14.3.2 Electrophilic substitution	2. Describe the nitration, sulfonation, alkylation, alkanoylation and the halogenation of the aromatic ring.
		3. Describe the concept of activating and deactivating groups and their directing influence on further substitution of the aromatic ring. <i>Such groups are to be exemplified by: -CH₃, -OH, -NH₂, -X, -NO₂, -CHO, -COOH.</i>
	14.3.3 Addition reactions	4. Apply canonical formulae, including Kekulé structures for aromatic molecules, for the description of delocalisation in benzene, phenol/phenoxide, phenylamine, nitrobenzene and chlorobenzene. <i>Mechanistic details are limited to OH/O⁻, amino, nitro and chloro groups bonded to benzene.</i>
	14.3.4 Side chain reactions of methylbenzene i. halogenation ii. oxidation	5. Describe the addition of hydrogen and chlorine to benzene under appropriate conditions. 6. Describe the free radical substitution of the methyl side chain with chlorine or bromine. 7. Describe the conversion of the -CH ₃ side chain to -CHO and -COOH using chromium(VI) dichloride dioxide and manganate(VII) respectively.
14.4 Alcohols, ethers, phenols	14.4.1 Alcohol classification	1. Identify a given alcohol group as primary, secondary or tertiary.
	14.4.2 Alcohol physical properties	2. Explain the relatively high boiling point and miscibility with water of lower members in terms of hydrogen bonding in comparison to ethers.

Topic	Sub-Topic	Assessment Criteria
	14.4.3 Alcohol preparation	3. Describe the preparation of alcohols from halogenoalkanes, alkenes, aldehydes, ketones, carboxylic acids and Grignard reagents.
	14.4.4 Industrial preparation of ethanol	4. Describe the preparation of ethanol by the fermentation of sugars to form ethanol in drinks and biofuels and by the steam hydration of ethene.
	14.4.5 Alcohol reactions	5. Describe oxidation reactions and the haloform reaction. 6. Recall Lucas test. 7. Explain how tertiary alcohols can be distinguished from the other two classes using Lucas Test. <i>The test has dubious validity in distinguishing secondary from primary alcohols and for this reason it will not be employed in the practical examination for this purpose.</i> 8. Describe the preparation of alkoxides, halogenoalkanes, esters, alkenes and ethers from alcohols.
	14.4.6 Aromatic alcohols	9. Describe aromatic alcohols, typified by phenylmethanol as different from phenols.
	14.4.7 Polyhydric alcohols	10. Describe the properties of polyhydric alcohols; limited to ethane-1,2-diol and propane-1,2,3-triol.
	14.4.8 Ethers preparation	11. Describe the preparation of symmetrical ethers from alcohols and symmetrical/unsymmetrical ethers from alkoxides and halogenoalkanes.
	14.4.9 Ether reaction	12. Describe the reaction of HI with ethers.
	14.4.10 Ether uses	13. Recall the use of ethoxyethane as a solvent and in extraction procedures despite its inflammable nature.
	14.4.11 Phenol preparation	14. Describe the preparation of phenol from sulfonic acids and diazonium salts.
	14.4.12 Phenol acidity	15. Explain the weakly acidic nature of phenol by its failure to liberate carbon dioxide with sodium carbonate and sodium hydrogencarbonate.
	14.4.13 Phenol reactions	16. Describe the preparation of esters and ethers from phenols. 17. Describe mono-substitution reactions (including coupling reactions with diazonium ions to form azo dyes) in the ring as well as the tribromination and trinitration. 18. Describe the reduction to benzene.

19. Recall that compounds with hydroxyl attached to the benzene ring give a characteristic violet colouration with neutral iron(III) chloride which also serves as a diagnostic test.

Topic	Sub-Topic	Assessment Criteria
14.5 Halogenoalkanes and halogenoarenes	14.5.1 Halogenoalkanes classification	1. Identify a given halogenoalkane group as primary, secondary or tertiary.
	14.5.2 Halogenoalkanes preparation	2. Describe the preparation of monohalogenoalkanes from alcohols and alkenes. <i>Limited to chloro-, bromo- and iodoalkanes</i>
		3. Describe the preparation dihalogenoalkanes from alkenes and alkynes. <i>Limited to chloro-, bromo- and iodoalkanes</i>
	14.5.3 Halogenoalkanes reactions	4. Describe the conversion of halogenoalkanes: i. by substitution into alcohols (by aqueous OH^-), ethers, amines, nitriles, esters and alkylarenes; ii. by elimination into alkenes and alkynes (using alcoholic OH^-); iii. into Grignard reagents; limited to monohalogenoalkanes; <i>For (iii), description should include that dry conditions are required to prevent hydrolysis of Grignard reagents.</i> iv. into alkanes by reaction with sodium under dry conditions.
	14.5.4 Halogenoalkanes reactivity	5. Explain the reactivity of monohalogenoalkanes in terms of their structure (primary, secondary, tertiary halogenoalkanes) and the halogen atom (Cl, Br, I). 6. Explain the lack of reactivity of fluoroalkanes in terms of the C–F bond strength.
	14.5.5 Halogenoarenes preparation	7. Describe the preparation by direct halogenation (where appropriate) and via diazonium compounds. <i>Limited to chloro-, bromo- and iodoarenes.</i>
	14.5.6 Halogenoarenes reactivity	8. Explain the unreactivity of the substituent halogen atom with respect to nucleophilic substitution reactions in terms of delocalisation of electrons. <i>Limited to chlorobenzene.</i>

14.5.7 Halogenoarenes reaction with magnesium

9. Describe the preparation of Grignard reagents.

Topic	Sub-Topic	Assessment Criteria
	14.5.8 Uses of organohalogen compounds other than in synthetic chemistry	10. Recall that the use of organohalogen compounds as solvents is being phased out due to their toxicity and other environmental impacts. 11. Explain the impact on the ozone layer caused by the use of CFCs and other halogeno compounds.
14.6 Aldehydes and ketones	14.6.1 Preparation	1. Describe the preparation of aldehydes and ketones from corresponding alcohols. 2. Describe the preparation of aromatic aldehydes by oxidation of a methyl group on a benzene ring using chromium(VI) dichloride dioxide. 3. Describe the preparation of aromatic ketones by alkanoylation.
	14.6.2 Addition reactions of the carbonyl group	4. Describe the addition reactions with H^- (as in the hydridometallates), ROH, HCN and NaHSO_3 (including the use of this reagent in the purification of aldehydes).
	14.6.3 Condensation reactions	5. Describe the reactions with hydrazine, phenylhydrazine, and 2,4-dinitrophenylhydrazine. 6. Describe the use of 2,4-dinitrophenylhydrazine as a test for the carbonyl group. 7. Explain the use of 2,4-dinitrophenylhydrazones for identification purposes.
	14.6.4 Other reactions	8. Describe the reaction with PCl_5, the haloform reaction, the aldol reaction (limited to the formation of the β-hydroxyaldehyde only), the Cannizzaro reaction, the reduction with hydrogen. 9. Distinguish between the type of carbonyl compound (aldehyde or ketone) based on oxidation employing weak oxidant systems; limited to Fehling's reagent and Tollen's reagent.
14.7 Carboxylic acids	14.7.1 Preparation	1. Describe the preparation of carboxylic acids: i. from alcohols and aldehydes by oxidation; ii. from carboxylic acid derivatives by hydrolysis. 2. Describe the preparation of aromatic carboxylic acids from alkylbenzenes by the oxidation of the alkyl side chain.

14.7.2 Weak acidity

3. Explain the weak protic acidity of carboxylic acids in terms of their structure and relative stability of the carboxylate anion.

Topic	Sub-Topic	Assessment Criteria
	14.7.3 α -substitution and acidity	4. Describe the increase or decrease in acidity of chlorine or alkyl substituted carboxylic acids in terms of inductive effect. 5. Compare the acidity of carboxylic acids with that of alcohols and phenols. 6. Account for the difference in acidity of carboxylic acids, alcohols and phenols in terms of structure. <i>The effect of substituents on the acidity of phenols will not be examined.</i>
	14.7.4 Reduction	7. Describe the conversion of $-\text{COOH}$ to $-\text{CH}_2\text{OH}$ using LiAlH_4 .
	14.7.5 Exceptional behaviour of methanoic acid	8. Describe the anomalous behaviour of methanoic acid and its salts as typified by dehydration, reaction with PCl_5 , acidified manganate(VII) and Tollens' reagent.
	14.7.6 Dicarboxylic acids	9. Recall the use of ethanedioic acid and its salts as standard reducing agents in titrimetry. 10. Explain the formation of acid anhydrides from dicarboxylic acids that allow it on the basis of structure.
	14.7.7 Carboxylate preparation	11. Describe methods of preparation of carboxylate salts.
	14.7.8 Carboxylate alkalinity	12. Explain the alkalinity through hydrolysis reactions.
	14.7.9 Soaps	13. Recall that sodium or potassium carboxylates having long alkyl chains can act as soaps.
	14.7.10 Decarboxylation	14. Describe the use of soda lime to bring about decarboxylation of carboxylates. <i>Represent this in an equation with NaOH.</i>
14.8 Acid derivatives	14.8.1 Preparation: esters, acid chlorides, acid anhydrides, acid amides, nitriles	1. Describe the preparation of esters by the reaction of: i. carboxylic acids with alcohols; ii. acid chlorides with alcohols or phenols; iii. acid anhydrides with alcohols or phenols; iv. a carboxylate with a halogenoalkane.

2. Describe the preparation of acid chlorides by the reaction of carboxylic acids with PCl_5 or SOCl_2 (sulfur(IV) dichloride oxide).
3. Describe the preparation of acid anhydrides from acid chlorides and a carboxylate.

Topic	Sub-Topic	Assessment Criteria
		4. Describe the preparation of acid amides, by: i. thermal decomposition of ammonium carboxylates; ii. reaction of esters with ammonia; iii. reaction acid chlorides with ammonia; iv. reaction of acid anhydrides with ammonia. 5. Describe the preparation of nitriles by: i. dehydration of acid amides using phosphorus(V) oxide; ii. the reaction of halogenoalkanes with an alcoholic solution of KCN. <i>Recognise the reaction in 5. ii. as a step-up technique in organic synthesis.</i> 6. Describe the preparation of aromatic nitriles from diazonium salts. 7. Compare the hydrolysis of acid chlorides and halogenoalkanes. 8. Describe the acid and alkaline hydrolysis of amides, esters and nitriles. 9. Describe the hydrolysis of acid chlorides and acid anhydrides. 10. Describe the reactions of acid chlorides and acid anhydrides with alcohols. 11. Describe the Hofmann degradation of amides. <i>Recognise the reaction in 11. as a step down technique in organic synthesis.</i> <i>An equation will not be required.</i> 12. Describe the reduction of acid amides and nitriles to amines. 13. Describe the reduction of esters to alcohols. <i>The reduction of acid chlorides and acid anhydrides will not be tested.</i> 14. Describe the reaction of acid amides with nitric(III) acid.
14.8.2 Reactions: hydrolysis, with alcohols, Hofmann degradation, reduction, HNO_2		

14.9 Amines and diazonium compounds	14.9.1 Amines structure	1. Identify a given amine as primary, secondary or tertiary. 2. Identify quaternary ammonium compounds.
	14.9.2 Preparation	3. Describe the preparation of primary amines by Hofmann degradation, reduction of amides and reduction of nitriles.

Topic	Sub-Topic	Assessment Criteria
		4. Describe the preparation of aromatic amines by reduction of nitroarenes by tin and concentrated hydrochloric acid. 5. Describe the preparation of a quaternary ammonium halide from the reaction of a tertiary amine with a halogenoalkane.
	14.9.3 Basicity	6. Explain the relative basic strength of ammonia, primary and secondary aliphatic amines and benzenamine in an aqueous solution in terms of their structure. 7. Explain the difference in basic strength between primary amines and acid amides.
	14.9.4 Reactions	8. Describe the preparation of alkylammonium salts from corresponding amines. 9. Explain the alkylation reaction. 10. Explain that benzenamine is an activated ring as exemplified by its reaction with aqueous bromine. 11. Describe the reaction of nitric(III) acid with primary aliphatic amines to produce nitrogen as a product of decomposition of unstable aliphatic diazonium compounds and its use as a test for primary aliphatic amines.
	14.9.5 Diazonium salts: structure	12. Draw structures of diazonium salts.
	14.9.6 Preparation	13. Describe the preparation of aryl diazonium salts by reaction of a primary aromatic amine with dilute hydrochloric acid and sodium nitrate(III) at a temperature below 5 °C. <i>Knowledge of the mechanism of the preparation of the diazonium ion is not required.</i>
	14.9.7 Reactions: i. substitution with loss of diazo group	14. Explain the relative stability of the diazonium ion on the basis of delocalization of charge through the mesomeric effect. 15. Describe the conversion of the diazonium salt to phenols, halogenoarenes, arenes, and nitriles

ii. diazo group is retained

16. Describe the coupling reactions of diazonium salts with phenols to form azo dyes.

Topic	Sub-Topic	Assessment Criteria
14.10 Amino acids	14.10.1 Properties	1. Recall that amino acids are typical difunctional molecules, which are important in biological systems. 2. Describe the reactions of amino acids as pertaining to the carboxyl and amino groups. 3. Explain the formation of dipolar ions (zwitterions) in amino acids and its influence on physical properties such as solubility in water and organic solvents and melting points. 4. Describe the effect of varying pH on the ionisation of an amino acid in aqueous solution with reference to its isoelectric point. <i>Note: Technique of electrophoresis will not be examined.</i> 5. Explain the existence of optical activity of some amino acids. 6. Explain the preparation of transition metal complexes using amino acid salts as bidentate ligands. 7. Describe a practical method of separation followed by identification of amino acids in a mixture using TLC, internal standards and R_f values.
	14.10.2 Preparation	8. Describe the synthesis of amino acids from aldehydes via hydroxynitriles.
14.11 Condensation polymers		1. Define condensation polymerisation. 2. Describe the formation of polypeptides as natural polyamides by condensation polymerisation between given amino acids. 3. Identify the peptide link in a given polypeptide. 4. Describe the relationship between amino acids and polypeptides. 5. Outline the formation of synthetic polyesters from difunctional molecules.

Limited to the formation of PET, poly(ethene terphthalate) and PLA, poly(lactic acid), from its monomer/s.

Note: the systematic name of PLA, poly(2-hydroxypropanoate), will not be assessed.

Experimental conditions are not required.

6. Recall the presence of the ester linkage in biological molecules such as triglycerides.

Topic	Sub-Topic	Assessment Criteria
		7. Outline the formation of polyamides from difunctional molecules. <i>Limited to the formation of nylon-66 and kevlar from the appropriate diamine and diacid/diacid chloride.</i> <i>Experimental conditions are not required.</i> 8. Recall the similarity between the amide and peptide linkage in polyamides and polypeptides/natural proteins respectively. 9. Draw the repeating unit for PET, PLA, nylon-66 and kevlar. 10. Explain the properties and major commercial uses of PET, PLA, nylon-66 and kevlar. 11. Explain the environmental advantage of the ease of hydrolysis of both the ester and amide (peptide) link. 12. Compare the hydrolytic properties of condensation polymers in contrast with the lack of reactivity of poly(alkene)s;
14.12 Mechanistic aspects	14.12.1 Ionic mechanisms	1. Describe the following mechanistic terms: nucleophile, electrophile, free radical. 2. Explain the following concepts: inductive and mesomeric effects. 3. Describe bimolecular nucleophilic substitution (S_N2) and unimolecular nucleophilic substitution (S_N1) reactions of halogenoalkanes. 4. Explain why primary halogenoalkanes react by a bimolecular mechanism while tertiary halogenoalkanes react by a unimolecular mechanism. 5. Describe electrophilic substitution of arenes. <i>Limited to nitration, alkylation, alkanoylation and chlorination.</i>

6. Describe the nucleophilic addition to the carbonyl group typified by the reactions of HCN and hydrazines with aldehydes and ketones.

Knowledge of the mechanism of the elimination of water from the addition product is not required.

7. Describe the electrophilic addition to carbon-carbon double bond as typified by reaction of halogens and hydrogen halides with alkenes.

Topic			Sub-Topic	Assessment Criteria
			14.12.2 Markovnikov's Rule	8. Explain the basis of Markovnikov's Rule in terms of the stability of carbocations.
			14.12.3 Homolytic mechanisms	9. Describe the homolytic reactions involved in the halogenation of alkanes, cycloalkanes and methylbenzene. <i>Limited to monohalogenation.</i> 10. Describe the homolytic reactions involved in the polymerisation of alkenes.
			14.13 Organic laboratory preparations	1. Describe in detail the laboratory preparation and isolation of pure samples of the following organic substances: i. ethanal from ethanol; ii. an ester from the reaction of a carboxylic acid and an alcohol.
			14.14 Organic reaction pathways and qualitative analysis	1. Devise a synthetic pathway involving organic reactions. 2. Identify intermediates, including reagents and conditions, in a given synthetic pathway. <i>Synthetic pathways will be limited to 4 steps unless an intermediate is given.</i>

Subject Focus:	Purification and Structure Determination
Learning Outcome 15	I can show an understanding of various purification techniques and the structure of molecules deduced from spectroscopic data.

Topic	Sub-Topic	Assessment Criteria
15.1 Purification and testing for purity of organic compounds	15.1.1 Physical purification	1. Describe the following techniques used in the purification of organic compounds: filtration, solvent extraction, recrystallisation, drying (for gases, liquids and solids), simple, fractional and steam distillation.
	15.1.2 Chemical purification	2. Describe the use of sodium hydrogensulfate(IV) in the purification of aldehydes.
	15.1.3 Derivatives for identification	3. Describe the process of derivatisation for identification purposes in aldehydes and ketones.
	15.1.4 Thin layer chromatography (TLC)	4. Explain the principles of chromatography as used in TLC. 5. Describe the procedure that is used when carrying out TLC.
	15.1.5 Tests for purity	6. Recall that the presence of impurities causes a melting point to deviate from the standard value. 7. Apply the principle of melting points as a test for the purity of a compound. 8. Describe how TLC can be used to check the purity of a substance. 9. Determine the presence of a substance in a mixture through the use of standards and by applying the concept of R_f values obtained from chromatograms. 10. Recall that boiling points obtained during distillation are a measure of purity.
15.2 The identification of structural features of organic compounds from physical and chemical properties		1. Deduce the presence of specific structural features from physical properties and results of reactions with reagents specified in this syllabus.
15.3 The determination of structural formulae from spectroscopy	15.3.1 Mass Spectrometry	1. Interpret the mass spectrum of an element to obtain the abundance and mass of each isotope and the relative atomic mass of the element. 2. Interpret simple mass spectra involving only singly-charged ions to obtain information about the structure of a molecular substance and the molecular mass of a substance. <i>Interpretation of organic mass spectra will be limited to simple fragmentation, where further fragmentation of a resulting fragment will not be examined.</i> <i>No knowledge of peak formation via rearrangements is expected.</i>

Topic	Sub-Topic	Assessment Criteria
Topic	Sub-Topic	Assessment Criteria
	15.3.2 Infrared (IR) Spectroscopy	3. Explain the peak height ratios for the mass spectra of Cl_2 and Br_2. <i>Students will be provided with the relative abundances and atomic masses of chlorine and bromine.</i> 4. Interpret infrared spectra to obtain information about parts of the structure of organic molecules using given tables of wavenumbers. <i>Infrared absorption data will be provided when required.</i> <i>Identification will be restricted to simple functional groups mentioned in other parts of the syllabus and based on stretching absorptions.</i> <i>Identification using the fingerprint region will not be examined.</i>
	15.3.3 Proton Nuclear Magnetic Resonance (^1H NMR) Spectroscopy	5. Interpret high-resolution ^1H NMR spectra and chemical shift data provided to obtain information about the structure or parts of the structure of organic molecules. <i>Carbon-13 NMR spectroscopy will not be examined.</i> <i>Chemical shift data will be provided when required.</i> <i>Identification should be restricted to functional groups mentioned in other parts of this syllabus.</i> 6. Explain the use of tetramethylsilane (TMS) as a standard in ^1H NMR spectroscopy. 7. Calculate the relative abundances of the chemically different protons in a molecule from given integration data of ^1H NMR spectra. 8. Apply the n+1 rule to determine the splitting patterns of adjacent, chemically different protons in a molecule. <i>Limited to doublet, triplet and quartet formation in aliphatic compounds.</i> <i>Candidates will be expected to identify aromatic protons from chemical shift data but will not be expected to analyse their splitting patterns.</i> 9. Identify whether a signal in a ^1H NMR spectrum is due to a proton in O–H and N–H by proton exchange using D_2O. <i>Details of instrumentation for this section are not required.</i> <i>Emphasis should be on interpretation of data to obtain information rather than spectroscopic theory.</i>

Topic	Sub-Topic	Assessment Criteria
15.4 Combined techniques		1. Deduce the structure of organic compounds using a combination of data from elemental and chemical analysis, mass spectrometry, IR spectroscopy and ^1H NMR spectroscopy.

Subject Focus: Quantitative and Qualitative Analysis

Learning Outcome 16	I can use chemicals and simple apparatus in quantitative and qualitative exercises, observe and record results of experimental work, interpret these observations and deduce correct inferences based on both qualitative and quantitative data. N.B. Candidates are instructed to use lab coats, disposable gloves and eye protection (safety goggles) during the practical sessions and to avoid wearing clothes or articles that increase the risk of accident in the laboratory.
----------------------------	--

Topic	Sub-Topic	Assessment Criteria
16.1 Quantitative exercises		1. Weigh a known mass of solid. 2. Prepare a solution of known concentration using a volumetric flask. 3. Prepare a diluted solution from a given solution using a volumetric flask. 4. Record burette and thermometer readings. 5. Calculate titre values and average titre value. 6. Obtain a set of concordant, accurate and precise titre values. 7. Plot temperature against time graphs. 8. Calculate temperature changes from a plotted graph.
	16.1.1 Acid-Base Titrations	9. Carry out titrations (including back titrations) involving solutions of acids and bases. <i>Indicator colour changes will be specified except for colour changes involving phenolphthalein and methyl orange.</i>
	16.1.2 Redox Titrations	10. Carry out redox titrations involving potassium manganate(VII) as an oxidising agent. 11. Carry out iodometric titrations involving the formation of iodine from the reaction of iodide with copper(II) or iodate(V) as oxidising agents followed by titration with thiosulfate. <i>Complexometric and argentometric titrations will not be assessed.</i>
	16.1.3 Calorimetry	12. Carry out calorimetry experiments to determine the enthalpy of neutralisation or the enthalpy change of the displacement reactions of metals.

Topic	Sub-Topic	Assessment Criteria
<i>Thermometric titrations will not be assessed.</i>		

Topic	Sub-Topic	Assessment Criteria
16.2 Qualitative exercises	16.2.1 Inorganic	1. Carry out test-tube reactions by following the instructions given to identify the following ions: ammonium, potassium, sodium, silver, calcium, magnesium, strontium, barium, aluminium, zinc, tin(II), chromium(III), manganese(II), manganese(IV), iron(II), iron(III), copper(I), copper(II), oxide, hydroxide, carbonate, hydrogencarbonate, nitrate(V), nitrate(III), carboxylate, sulfate(IV), sulfate(VI), thiosulfate, chloride, bromide, iodide, phosphate, manganate(VII). <i>Instructions should indicate the presence of binary salt mixtures or double salts.</i> <i>Instructions should include volumes and masses that can be measured with the provided equipment.</i> 2. Record observations on colours of solutions and precipitates, effervescence, colours of gases and pH. 3. Infer the need to test for gases from observations. <i>Gases are limited to: ammonia, carbon dioxide, oxygen, sulfur(IV) oxide, hydrogen halides (HX, X = Cl, Br, I) and nitrogen(IV) oxide.</i> 4. Deduce inferences from recorded observations.
	16.2.2 Organic	5. Carry out test-tube reactions by following the instructions given to identify the following: Double and triple bonds (alkenes, alkynes), hydroxyl (alcohols and phenols), carbonyl (aldehydes and ketones), carboxyl, carboxylate salts, amides, and amino groups. <i>Both aliphatic and aromatic compounds may be assessed.</i> <i>A combination of functional groups may be assessed, e.g. amino acids, hydroxybenzoic acid, oxopentanoic acid etc ...</i> 6. Record observations on colours of solutions and precipitates, effervescence, colours of gases and pH.

7. Infer the need to test for gases from observations.

Gases are limited to: ammonia, carbon dioxide, nitrogen(IV) oxide and nitrogen (by elimination).

8. Deduce inferences from recorded observations.

Scheme of Assessment

The examination consists of two written papers, Paper I and Paper II and a practical Paper III. Each of the papers is of three hours duration.

Paper I

- a written paper which carries 40% of the total score;
- consists of six to ten compulsory structured questions;
- candidates will write their answers on the examination paper in the spaces provided.

Paper II

- a written paper which carries 40% of the total score;
- consists of two sections each containing four extended response questions carrying equal marks - five questions are to be answered from this paper, two questions from each section and any one other.
- candidates will write their answers on a separate booklet provided by MATSEC.

Paper III

- a practical examination which carries 20% of the total score;
- an open book examination – candidates may use any printed material which assists them in their work.

Candidates may use an electronic calculator in all parts of the examination.

The following data is provided in the examination papers, when applicable:

- Molar Gas Constant, $R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$
- Molar volume at STP = 22.40 L
- Faraday's Constant = 96500 C mol⁻¹
- Ionic product of water, $K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$
- 1 atm = 101325 Pa
- SHC of water, $c_w = 4.20 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$
- Density of water, $\rho_w = 1.00 \text{ g cm}^{-3}$

Familiarity with Periodic Table

Candidates are expected to be familiar with the structure of the Periodic Table and the group affiliation of each element with atomic number from 1 to 30. However, a copy of the Periodic Table will be provided for Paper II of the examination although not for Paper I. Relative atomic masses will always be given where necessary.

Mathematical skills

In order to understand certain concepts required by the syllabus, candidates will need to be able to:

- Recognise and use expressions in decimal and standard form; use ratios, fractions and percentages and find arithmetic means; make estimates of results; use an appropriate number of significant figures; use calculators to find and use x^y , $\sqrt{x}$, $\frac{1}{x}$, $\log_{10} x$.
- Change the subject of an equation; substitute numerical values into algebraic equations using units for physical quantities; use logarithms in relation to quantities, which range over several orders of magnitude.
- Appreciate angles and shapes in regular two and three-dimensional structures and to represent three-dimensional forms in two dimensions; understand the symmetry of two-dimensional and three-dimensional shapes.
- Plot two variables from given data; understand that $y = mx + c$ represents a linear relationship and be able to determine the slope and intercept of a line; draw and use the slope of a tangent to a curve as a measure of rate of change.

Language skills

Candidates are reminded of the importance of the use of good English in answering questions in the examination. Orderly and detailed presentation will be rewarded in answers to extended-response questions and calculations.