



**L-Università
ta' Malta**

**MATSEC
Examinations
Board**



IM 06 Chemistry
CHEMISTRY

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Introduction

The syllabus assumes candidates have already followed an introductory course in chemistry and it presents an overview of important concepts in the subject. Most of the concepts studied at intermediate level will not be developed to the same depth as at the Advanced Matriculation level but the syllabus is intended to cover key ideas that allow the student to understand better the nature of chemicals and the forces that drive chemical change.

The influence of chemicals and chemical industry on the quality of life and on the environment should be discussed throughout the course. Examples of everyday applications of chemistry and its products are numerous and they should be invoked by tutors as frequently as possible.

Chemistry is an experimental science and it is recommended that students carry out practical work. It may not be possible for students following this course to participate actively in laboratory work: to make up for this, tutors should be very generous with experimental demonstrations.

The syllabus is presented in a number of Learning Outcomes (1 – 15) 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.

List of Subject Foci

1. Atoms, molecules, ions and the mole concept
2. Stoichiometry
3. States of Matter
4. Chemical bonding
5. Energetics
6. Rates of Reactions
7. Chemical Equilibrium
8. Ionic Theory and Redox Reactions
9. Acids and Bases (Ionic equilibria)
10. Chemical Periodicity
11. Transition Metals
12. Chemistry of the halogens
13. Organic Chemistry: General Principles
14. Pure substances and mixtures
15. Qualitative and Quantitative Analysis

List of Learning Outcomes

- ☐ I can explain the structure of an atom and investigate differences between different atoms.
- ☐ I can apply the mole concept in calculations.
- ☐ I can explain the kinetic theory of gases and solve problems using the ideal gas equation.
- ☐ I can explain different types of bonding, why they exist and how to investigate them.
- ☐ I can define enthalpy changes and calculate enthalpy changes using Hess's law.
- ☐ I can describe and explain the effect of different factors on the rate of a chemical reaction.
- ☐ I can understand the concept of chemical equilibria, predict changes in the position of the equilibrium with a change in conditions and carry out the necessary calculations involving chemical equilibria.
- ☐ I can understand the concepts of oxidation and reduction, and together with the reactivity series I can predict the outcome of certain chemical reactions.
- ☐ I can explain the concept of ionic equilibria and use this concept to calculate a number of quantities associated with these types of equilibria.
- ☐ I can describe periodic relationships among elements in the Periodic Table.
- ☐ I can describe the main characteristics of transition metals.
- ☐ I can describe some of the physical properties, chemical properties and uses of chlorine, bromine and iodine and some of their compounds.
- ☐ I can identify, name, represent and deduce the structure and properties of different organic molecules.
- ☐ I can distinguish between pure substances and mixtures as well as choose the correct technique to separate a mixture of substances.
- ☐ I can show an understanding of qualitative and quantitative exercises.

Learning Outcomes and Assessment Criteria

Subject Focus:	Atoms, molecules, ions and the mole concept
Learning Outcome 1	I can explain the structure of an atom and investigate differences between different atoms.

Topic	Sub-Topic	Assessment Criteria
1.1 Atomic nature of matter		- Recall that an atom is the smallest particle making up an element. - Describe qualitatively the structure of an atom in terms of fundamental particles (protons, electrons and neutrons) using the Bohr model. - Recall the charges and relative masses of the fundamental particles. - Describe matter in terms of atoms, molecules or ions.
1.2 The nuclear structure of the atom		- Define the terms proton (or atomic) number and nucleon (or mass) number. - Define isotopes. - Define the relative isotopic mass. - Define the relative atomic mass. - Calculate the relative atomic mass of an element from the isotopic masses and relative abundances. - Define relative molecular/formula mass.
1.3 Electronic configurations		- Recall the existence of s, p and d orbitals. <i>Candidates will not be expected to draw the shapes of s, p and d orbitals.</i> - Construct the electronic configuration of isolated atoms of elements H to Ar using 1s, 2s, 2p etc. notation and electrons-in-boxes notation and applying the 'building – up' (aufbau) principle.
1.4 Classification of elements		Classify elements into groups periods, and 'blocks', including the first row d-block elements.

Topic	Sub-Topic	Assessment Criteria
1.5 Ionisation energies and relation to electronic configuration		1. Define first ionisation energy. 2. 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.

Subject Focus:	Stoichiometry
Learning Outcome 2	I can apply the mole concept in calculations.

Topic	Sub-Topic	Assessment Criteria
2.1 The mole		1. Define the mole. 2. Calculate the number of particles in a certain amount of substance using the Avogadro constant. 3. Calculate the relative molecular/formula mass of a substance from given data.
2.2 Empirical and molecular formulae		1. Determine the empirical formula of a compound from masses or percentage by mass of the elements. 2. Determine the molecular formula of a compound from the empirical formula and relative molecular/formula mass.
2.3 Chemical Equations	2.3.1 Balanced equations	1. Construct both full and ionic balanced chemical equations to represent relative numbers of particles in chemical reactions. 2. Represent the states of substances by using symbols (s), (l), (g) and (aq) for solid, liquid, gas and aqueous solutions respectively. 3. Apply chemical equations for calculations involving reacting substances in terms of amounts and concentrations measured in moles and mol dm^{-3} respectively.

Subject Focus:	States of Matter
Learning Outcome 3	I can explain the kinetic theory of gases and solve problems using the ideal gas equation.

Topic	Sub-Topic	Assessment Criteria
3.1 Molar volume of gases and Avogadro's Law.		1. Recall Avogadro's Law. 2. Calculate the ratios of reactants and products in chemical equations using gas volumes.
3.2 Kinetic theory of gases		1. Sketch distribution diagrams showing energy 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 assessed.</i>
3.3 Ideal gases and the ideal gas law		1. Recall the basic assumptions of the ideal gas model. <i>This should be limited to the size of particles and intermolecular forces.</i> 2. Explain the ideal gas equation. 3. Solve problems involving the ideal gas equation.
3.4 Vapour pressure and saturated vapour pressure		1. Explain vapour pressure as evidence of the presence of a vapour in contact with the evaporating liquid/solid. 2. Explain saturated vapour pressure.

Subject Focus:	Chemical bonding
Learning Outcome 4	I can explain different types of bonding, why they exist and how to investigate them.

Topic	Sub-Topic	Assessment Criteria
4.1 Ionic (Electrovalent) Bonding		1. Describe Ionic bonding as an electrostatic attraction between oppositely charged ions. 2. Describe the physical properties of NaCl in relation to its lattice structure. <i>Physical properties limited to melting point, conductivity, physical state and solubility.</i> <i>Structure of the unit cell is not examinable.</i>
4.2 Covalent bonding		1. Draw Lewis structures showing all outer electrons as dots and crosses and/or the corresponding bonds as lines. 2. Recall the concept of electrostatic attraction in covalent bonding. 3. Recall that covalent bonding is present in both simple and macromolecules. 4. Describe the structure and bonding in macromolecules. <i>Macromolecules limited to silicon(IV) oxide, graphite and diamond.</i>
	4.2.1 Dative bonding	5. Define dative bonding. <i>Limited to CO molecule, H_3O^+, NH_4^+ ions and complex ions (refer to 11.3).</i> 6. Apply the octet rule in predicting formulae. <i>To include exceptions to octet rule for the electron deficient molecules and 'octet expansion'. Limited to BF_3 and SF_6.</i>
4.3 Polar covalent bonds and polarity in covalent molecules		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 unequal electron sharing depending on the electronegativity of the atomic species concerned. 3. Explain intermediate bonding in terms of bond polarisation.

Topic	Sub-Topic	Assessment Criteria
		4. Describe polar covalent bonds which may, or may not, give rise to molecules with a permanent dipole. <i>Limited to H₂O, CO₂ and CCl₄.</i> <i>Use of symbols $\delta+$ and $\delta-$ to represent partial charges in bonds.</i>
4.4 Shapes of molecules and simple ions	4.4.1 Valence shell electron pair repulsion (VSEPR) theory.	1. Sketch appropriate diagrams of the shapes of molecules/ions in BeF₂, BF₃, CH₄, SF₆, NH₃, H₂O, NH₄⁺, H₃O⁺. 2. State the names of the shapes of molecules/ions mentioned in 4.4.1. 3. Recall the approximate bond angles in BF₃, CH₄, NH₃, H₂O.
4.5 Physical properties of simple 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) and silicon(IV) oxide.</i>
4.6 Intermolecular forces		1. Define permanent dipole-permanent dipole, induced dipole-induced dipole attractions and hydrogen bonds. 2. Explain the effects of these forces on specific properties of molecular compounds: e.g. boiling points of simple hydrides in Groups 4 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.
	4.6.1 Open crystal structure of ice	4. Compare the density of ice to that of liquid water in terms of the open crystal structure of ice.
4.7 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.

Subject Focus:	Energetics
Learning Outcome 5	I can define enthalpy changes and calculate enthalpy changes using Hess's law.

Topic	Sub-Topic	Assessment Criteria
5.1 Changes in energy	5.1.1 Exothermic and endothermic processes	1. State that fuels and food contain energy. 2. Recognise the Joule (J) as the unit of energy. 3. Describe endothermic and exothermic energy changes for phase changes and chemical changes. 4. Recognise the ΔH notation used for enthalpy change.
	5.1.2 Energy level (enthalpy) diagrams	5. Sketch energy level (enthalpy) diagrams for exothermic and endothermic processes.
5.2 Standard enthalpy changes		1. Define the following standard enthalpy changes of: formation, atomisation, combustion, solution (or dissolution) and neutralisation. <i>Experimental determination of these enthalpy changes will not be assessed.</i> 2. Construct thermochemical equations for the enthalpy changes mentioned in 1. above. 3. Explain enthalpy of reaction in terms of bond breaking and bond making processes. 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. <i>Excluding aromatic compounds.</i>
5.3 Hess's law	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. <i>Limited to calculation of enthalpy changes of formation and combustion.</i>

Subject Focus:	Rates of Reactions
Learning Outcome 6	I can describe and explain the effect of different factors on the rate of a chemical reaction.

Topic	Sub-Topic	Assessment Criteria
6.1 Reaction rate		1. Describe the reaction rate as the increase of product concentration or decrease of reactant concentration with time.
6.2 Factors affecting the rate of a chemical reaction	6.2.1 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. <i>Treatment of the collision theory should be qualitative.</i>
	6.2.2 Activation energy	3. Define the term activation energy.
	6.2.3 Catalysts	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. State the effect of a catalyst on the rate of reaction.
	6.2.4 Maxwell-Boltzmann distribution of energies	6. Sketch on the same pair of axes, a graph showing the distribution of energies at two different temperatures. 7. Recall that the rate of reaction nearly doubles every 10K rise in temperature. 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.

Subject Focus:	Chemical Equilibrium
Learning Outcome 7	I can understand the concept of chemical equilibria, predict changes in the position of the equilibrium with a change in conditions and carry out the necessary calculations involving chemical equilibria.

Topic	Sub-Topic	Assessment Criteria
7.1 Concept of dynamic equilibrium	7.1.1 Reversible reactions	1. Interpret the $\rightleftharpoons$ notation as showing a reversible reaction.
	7.1.2 Dynamic equilibrium	2. State that chemical equilibria are dynamic. 3. Describe the dynamic equilibrium state as a chemical reaction in which the rates of the forward and reverse reactions are equal.
7.2 The Equilibrium Law	7.2.1 Equilibrium constant K_c	1. Derive equilibrium constant in terms of concentrations (K_c) for equilibria including units as appropriate. <i>Only homogenous equilibria will be assessed.</i>
		2. Calculate K_c from the data provided. <i>Calculations of equilibrium concentrations from K_c will not be assessed.</i> 3. Deduce the extent of reaction from the magnitude of the equilibrium constant.
7.3 Le Chatelier's Principle	7.3.1 Effect of change in conditions on equilibrium position	1. Explain the qualitative effect on the equilibrium position for changes in concentration, temperature and (where relevant) pressure on equilibria. <i>The effect of the addition of an inert gas at constant volume on equilibrium will not be assessed.</i>
	7.3.2 Haber process	2. Explain the effects of the presence of catalysts on equilibria. 3. Describe the Haber process for the manufacture of ammonia as an example of an industrial application of a reversible process. <i>Details of the manufacturing plant will not be assessed.</i> 4. Explain why compromise conditions are required in the Haber process. 5. Describe the use of ammonia obtained from the Haber process in the production of artificial fertilisers.

Subject Focus:	Ionic Theory and Redox Reactions
Learning Outcome 8	I can understand the concepts of oxidation and reduction, and together with the reactivity series, I can predict the outcome of certain chemical reactions.

Topic	Sub-Topic	Assessment Criteria
8.1 Oxidation and reduction	8.1.1 Redox reactions	1. Recognise redox reactions. 2. Describe redox reactions in terms of loss or gain of electrons, oxygen or hydrogen. 3. Explain disproportionation reactions. 4. Construct balanced redox equations using half equations. 5. Explain the concept of spectator ions. 6. Deduce the oxidation number of elements in different chemical species. <i>(Limited to binary compounds between metals and non-metals only; excluding peroxides and superoxides).</i> 7. Define oxidising and reducing agents. 8. Identify common oxidising agents and reducing agents covered elsewhere in this syllabus. Note: Practicals are strongly recommended when covering this section.

Subject Focus	Acids and Bases (Ionic equilibria)
Learning Outcome 9	I can explain the concept of ionic equilibria and use this concept to calculate a number of quantities associated with these types of equilibria.

Topic	Sub-Topic	Assessment Criteria
9.1 Acids and bases	9.1.1 Strong and weak acids and bases	1. Define Arrhenius acids and bases. 2. Explain the proticity or basicity of acids and acidity of bases (e.g. sulfuric(VI) acid as a 'diprotic' or 'dibasic' acid and carbonate as a 'diacid' base). 3. Define strong and weak acids, and strong and weak bases, according to the extent of their dissociation.
	9.1.2 Brønsted-Lowry theory	4. Define the Brønsted-Lowry theory for acids and bases. 5. Explain conjugate pairs.
	9.1.3 K_a and K_b	6. Derive an expression for the dissociation constants K_a and K_b and include their units. 7. Explain acid and base strength in terms of their dissociation constants values. 8. Define amphoteric substances.
	9.1.4 K_w	9. Explain with the aid of equations the amphoteric nature of water. 10. Derive an expression for the ionic product of water, K_w . <i>Numerical calculations involving K_a and K_b will not be assessed.</i>
9.2 pH, indicators and salts	9.2.1 pH of aqueous solutions	1. Define pH. 2. Relate pH to the pH scale. 3. Calculate the pH and pOH of aqueous solutions. 4. Calculate the $[H_3O^+]$ and $[OH^-]$ from the pH and vice versa.

Topic	Sub-Topic	Assessment Criteria
		5. State that the pH meter and universal indicators can be used to measure pH. <i>Students are expected to be familiar with the logarithmic function as required for calculation of pH.</i> <i>Theory of indicators will not be assessed.</i> <i>Students are not expected to recall the colour changes and pH ranges of any indicator.</i> Hands on practicals are highly recommended in this topic.
9.3 Buffer solutions		1. Describe an acidic buffer system. 2. Explain how an acidic buffer system works to keep the pH stable. 3. Explain how a buffer system helps maintain the pH of human blood stable. <i>Problems involving numerical calculations of buffer solutions will not be set.</i>

Subject Focus:	Chemical Periodicity
Learning Outcome 10	I can describe periodic relationships among elements in the Periodic Table.

Topic	Sub-Topic	Assessment Criteria
10.1 Periodic relationships		1. Recall that the periodic classification of elements is done in terms of electronic structure. 2. Describe the variation in properties as illustrated by their trends in melting and boiling points, valency, atomic radius, first ionisation energy and metallic character. <i>Limited to Period 2 and Period 3 elements.</i>

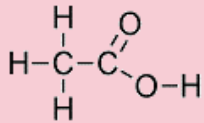
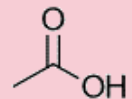
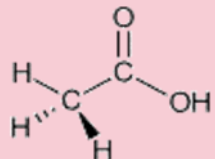
Subject Focus:	Transition Metals
Learning Outcome 11	I can describe the main characteristics of transition metals.

Topic	Sub-Topic	Assessment Criteria
11.1 D-block elements and transition metals		1. Recall 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>
11.2 Characteristic properties		1. State the colour of coordination compounds containing the following ions: $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$, $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$, $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$, $[\text{Cu}(\text{NH}_3)_4]^{2+}$, $[\text{CuCl}_2]^-$, $[\text{Mn}(\text{H}_2\text{O})_6]^{2+}$. 2. Describe the presence of variable oxidation states in transition metal ions/compounds. <i>Limited to: Mn(II), Mn(IV), Mn(VII), Fe(II), Fe(III), Cu(I) and Cu(II).</i> 3. Recall the catalytic properties of transition metal ions and compounds. <i>Limited to the following reactions: MnO_2 in the decomposition of hydrogen peroxide; Fe in the Haber Process; Raney Ni in hydrogenation of unsaturated hydrocarbons.</i>
11.3 Coordination compounds		1. Define a ligand. <i>Limited to monodentate.</i> 2. Recall the following examples of ligands: H_2O, NH_3, Cl^-, CO. 3. Describe metal-ligand bonding in terms of dative covalent (coordinate) bonding. 4. Define coordination number of a complex. 5. Draw linear and octahedral complex ions. 6. Name the following complex ions: $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$, $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$, $[\text{Cu}(\text{NH}_3)_2]^+$, $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$, $[\text{Cu}(\text{NH}_3)_4]^{2+}$, $[\text{Mn}(\text{H}_2\text{O})_6]^{2+}$.

Subject Focus:	Chemistry of the halogens
Learning Outcome 12	I can describe some of the physical properties, chemical properties and uses of chlorine, bromine and iodine and some of their compounds.

Topic	Sub-Topic	Assessment Criteria
12.1 Trends in physical and chemical properties		1. Explain the trend in boiling points. 2. Explain the trend in oxidising power of the elements. <i>Limited to: chlorine, bromine and iodine.</i>
12.2 The use of iodine in titrimetry		1. Write an equation for the reaction of iodine with thiosulfate ions. 2. Describe the use of this reaction in titrimetry.
12.3 Uses of chlorine and the chlorate(I) ion		1. Describe some of the uses of chlorine. 2. Write the formula of the chlorate(I) ion. 3. State the use of chlorate(I) as a bleaching agent and as a disinfectant.

Subject Focus:	Organic Chemistry: general principles
Learning Outcome 13	I can identify, name, represent and deduce the structure and properties of different organic molecules.

Topic	Sub-Topic	Assessment Criteria
13.1 Functional groups and homologous series.	13.1.1 Structural 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 structural formulae of organic compounds can be given as described in 13.1.1.2.</i> 3. Apply wedge and dashed/hatched line convention to represent three-dimensional structure, such as: 
	13.1.2 Functional groups	4. Identify functional groups in the following classes: alkenes, alkynes, alcohols, ethers, halogenoalkanes, aldehydes, ketones, carboxylic acids and amines.
	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)). <i>Limited to the compounds mentioned in this syllabus.</i>

Topic	Sub-Topic	Assessment Criteria
Topic	Sub-Topic	Assessment Criteria
13.2 Redox equations in organic chemistry		1. Construct oxidation and reduction equations using [O] and [H] respectively unless specifically asked to give half equations.
13.3 Catenation and spatial distribution in carbon compounds	13.3.1 Catenation	1. Explain why the element carbon is unique in exhibiting catenation in terms of the strength of the C-C bond.
	13.3.2 Spatial distribution of the bonds in simple carbon compounds	2. Recall the shape of the bonds and bond angles in alkanes (tetrahedral), alkenes (trigonal planar).
13.4 Isomerism in organic compounds	13.4.1 Structural isomerism	1. Describe different types of structural isomerism. <i>Limited to chain, positional and functional isomerism.</i>
	13.4.2 Stereoisomerism	2. Describe stereoisomerism namely cis-trans (geometrical) of the C=C bond, and optical isomerism (restricted to enantiomerism in 2-hydroxypropanoic acid and 2-aminopropanoic acid). <i>Geometrical isomerism is limited to compounds with the same groups/atoms across the double bond.</i>
	13.4.3 Optical isomerism	3. Describe the property of optical activity in terms of dextrorotatory (+) and laevorotatory (-) compounds and their effect on plane polarized light. 4. Describe 2-hydroxypropanoic acid and 2-aminopropanoic acid as chiral molecules, which do not possess a plane of symmetry. <i>The use of the descriptors D and L and R and S is not required.</i>
	13.4.4 Drawing and identification of isomers	5. Explain the biological importance of optical isomers. 6. Relate the biological importance of optical isomers to compounds with specific reference to Ibuprofen and Thalidomide. 7. Draw structural and stereoisomers listed in 13.4.1.1 and 13.4.2.2. 8. Identify structural and stereoisomers listed in 13.4.1.1 and 13.4.2.2. <i>For optical isomers, the wedge and dashed/hatched line convention should be used.</i>

Topic	Sub-Topic	Assessment Criteria
13.5 Alkanes, alkenes and alkynes	13.5.1 Alkanes	1. Recall the general unreactivity of alkanes.

Topic	Sub-Topic	Assessment Criteria
	13.5.2 Substitution reaction	2. Describe substitution by chlorine and bromine using appropriate overall chemical equations. <i>Description should include reference to multiple substitutions but chemical equations should be limited to mono substitutions.</i>
	13.5.3 Alkenes as unsaturated hydrocarbons	3. Describe the use of alkanes as fuels. 4. Describe alkenes as unsaturated hydrocarbons containing a double bond made up of a σ and a π bond. 5. Identify the double bond as an electron rich centre susceptible to electrophilic attack.
	13.5.4 Preparation	6. Outline the preparation of alkenes from alcohols. <i>Limited to dehydration with excess concentrated sulfuric(VI) acid at 180°C; including chemical equation.</i>
	13.5.5 Electrophilic addition reactions	7. Describe the reaction of alkenes with hydrogen, halogens, hydrogen halides, and water in the presence of sulfuric(VI) acid.
	13.5.6 Markovnikov's rule	8. Deduce the structure of the major product in the addition of hydrogen halides with propene.
	13.5.7 Alkynes	9. Recall that alkynes are unsaturated hydrocarbons containing a triple bond made up of a σ and two π bonds.
	13.5.8 Combustion	10. Describe the exothermic reaction of ethyne with oxygen and its use in oxy-acetylene flame.
13.6 Alcohols and ethers	13.6.1 Alcohol classification	1. Identify a given alcohol group as primary, secondary or tertiary.
	13.6.2 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.

13.6.3 Laboratory preparation of alcohols	3. Outline the laboratory preparation of alcohols from alkenes.
13.6.4 Industrial preparation of ethanol	4. Outline the formation of ethanol by the fermentation of sugars to form ethanol in drinks and biofuels.

Topic	Sub-Topic	Assessment Criteria
	13.6.5 Alcohol reactions	5. Outline the oxidation reactions of primary alcohols to aldehydes and carboxylic acids, and secondary alcohols to ketones. <i>Limited to the use of acidified dichromate(VI) as oxidising agent.</i> 6. Outline the formation of chloroalkanes (via reaction with PCl_5), esters, alkenes and ethers from alcohols. 7. Compare the reaction of alcohols with sodium with that of water with sodium.
	13.6.6 Polyhydric alcohols	8. Outline the physical properties of polyhydric alcohols; limited to ethane-1,2-diol and propane-1,2,3-triol. 9. Relate the properties of the polyhydric alcohols mentioned in 8. above to their uses.
	13.6.7 Ethers preparation	10. Outline the preparation of ethoxyethane from ethanol.
	13.6.8 Ether unreactivity	11. Recall the relative unreactivity of ethers.
	13.6.9 Ether uses	12. Recall the use of ethoxyethane as a solvent and in extraction procedures despite its inflammable nature.

13.7 Aldehydes and ketones

1. Outline the preparation of aldehydes by the oxidation of primary alcohols.
Limited to the use of acidified dichromate(VI) as oxidising agent.
2. Outline the oxidation of aldehydes to carboxylic acids.
Limited to the use of acidified dichromate(VI) as oxidising agent.
3. Describe Fehling's test as a means of distinguishing between aldehydes and ketones.
4. Outline the reduction of aldehydes and ketones using LiAlH_4 in ether to give primary and secondary alcohols respectively.
5. Outline the addition of HCN to form hydroxynitriles as a typical reaction of carbonyl compounds.
6. Outline the hydrolysis of hydroxynitriles to hydroxycarboxylic acids.
7. Recall the existence of the carbonyl group in biological molecules such as sugars.
8. Recall the use of carbonyl compounds as organic solvents as exemplified by butanone.

Topic	Sub-Topic	Assessment Criteria
13.8 Carboxylic acids		1. Outline the preparation of carboxylic acids by the oxidation of primary alcohols and aldehydes. 2. Describe the weak acidic nature of carboxylic acids in terms of the ability to partially dissociate in water and generate H^+ ions. 3. Outline the formation of salts from carboxylic acids. 4. Describe how salt formation increases the solubility of the organic compound. 5. Recall that sodium or potassium carboxylates having long alkyl chains can be obtained by alkaline hydrolysis of triglycerides and can act as soaps. 6. Outline the formation of esters by the reaction of carboxylic acids with alcohols. 7. Recall that the carboxylic acid group is present in biological molecules such as amino acids.
13.9 Amines		1. Describe the basic nature of the amino group as the ability to accept H^+ ions. 2. Outline the formation of salts from primary amines. 3. Recall that primary amines and ammonia form soluble ammonium salts.

13.10 Amino acids

4. Recall that the amino group is present in biological molecules such as amino acids.

1. Recall that amino acids are difunctional molecules that are important in biological systems.
2. Explain the formation of dipolar ions (zwitterions) in amino acids and its influence on physical properties such as solubility in water and melting points.
3. Describe the effect of varying pH on the ionisation of an amino acid in aqueous solution with reference to its isoelectric point.

Note: Technique of electrophoresis will not be assessed.

4. Explain the existence of optical activity of some amino acids as exemplified by 2-aminopropanoic acid.
5. Recall that amino acids can be separated by thin layer chromatography.
6. Recall that polypeptides are formed by the condensation polymerisation between amino acids.
7. Identify the peptide link in a given polypeptide.

Topic	Sub-Topic	Assessment Criteria
13.11 Polymers	13.11.1 Addition polymers	1. Define addition polymerisation. 2. Deduce the repeating unit of a polymer obtained from a given alkene monomer. 3. Draw the repeating unit for poly(ethene) and poly(chloroethene). 4. Describe the uses of poly(ethene) and poly(chloroethene). 5. Explain the lack of reactivity and non-biodegradability of poly(alkene)s and hence the accumulation in the environment.
	13.11.2 Condensation polymers	6. Define condensation polymerisation. 7. Outline the formation of polyesters from difunctional molecules. <i>Limited to the formation of PLA, poly(lactic acid), from its monomer.</i> <i>Note: the systematic name of PLA, poly(2-hydroxypropanoate), will not be assessed.</i> <i>Experimental conditions are not required.</i>

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

9. Outline the formation of polyamides from difunctional molecules.

Limited to the formation of nylon-66 from the appropriate diamine and diacid.

Experimental conditions are not required.

10. Recall the similarity between the amide and peptide linkage in polyamides and polypeptides/natural proteins respectively.

11. Draw the repeating unit for PLA and nylon-66.

12. Explain the properties and major commercial uses of PLA and nylon-66.

13. Explain the environmental advantage of the ease of hydrolysis of both the ester and amide (peptide) link.

14. Compare the hydrolytic properties of condensation polymers in contrast with the lack of reactivity of poly(alkene)s;

Subject Focus:	Pure substances and mixtures
Learning Outcome 14	I can distinguish between pure substances and mixtures as well as choose the correct technique to separate a mixture of substances.

Topic	Sub-Topic	Assessment Criteria
14.1 Pure substances and mixtures		1. Distinguish between pure substances and mixtures. <i>In terms of boiling and melting points.</i> <i>Knowledge related to cooling curves is not expected.</i> <i>Mixtures are limited to: solid/liquid, liquid/liquid, and solid/solid excluding alloys.</i> 2. Distinguish between solvent, solute, and solution.
14.2 Separation of mixtures		1. Describe the following techniques used in the purification of compounds: filtration, distillation, sublimation, crystallisation, thin layer chromatography and solvent extraction. <i>Solvent extraction will be limited to the qualitative treatment of the partitioning of a solute between two immiscible solvents treated as an equilibrium process.</i>

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

2. Choose the correct technique/s from 1. above to separate/purify a mixture of substances.

Subject Focus:	Qualitative and Quantitative Analysis
Learning Outcome 15	I can show an understanding of qualitative and quantitative exercises.

Topic	Sub-Topic	Assessment Criteria
15.1 Qualitative analysis		1. Identify ions from given reactions and/or observations. <i>Limited to: sodium, potassium, magnesium, calcium, barium, ammonium, copper(II), iron(II), iron(III), manganese(II), zinc, aluminium; chloride, bromide, iodide, sulfate(VI), sulfate(IV), thiosulfate, carbonate, nitrate(V), nitrate(III), hydroxide.</i> <i>Tests for chloride, bromide and iodide will be limited to: reactions with silver nitrate(V).</i> <i>Tests for nitrate(V) ions will be limited to reduction by Devarda's alloy.</i> 2. Describe tests to distinguish between two different ions. <i>Limited to the ions listed in 1. above.</i> <i>Description to include equations for precipitation reactions only.</i>
15.2 Quantitative analysis	15.2.1 Balanced equations	1. Construct both full and ionic balanced 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. <i>Calculations involving volumetric analysis are limited to acid-base titrations and redox titrations involving manganate(VII) – iron(II) and iodine - thiosulfate.</i> <i>Calculations involving back titrations are limited to the addition of known excess of acid or alkali to a reactant solution and determining the excess by titration.</i>
	15.2.2 Limiting reagent	3. Recall the use of phenolphthalein and methyl orange as indicators. 4. Deduce the limiting reagent in a chemical reaction from given data. 5. Apply the concept of limiting reagent in calculations.
	15.2.3. Percentage yield	6. Calculate the percentage yield of a product in a chemical reaction. 7. State why chemical reactions do not always produce a 100% yield. 8. Calculate the percentage purity of a substance from given data.

Scheme of Assessment

The Examination consists of one written three-hour paper.

The paper will be divided into three sections as follows:

- Section A
 - consists of eight to ten compulsory questions of the fill-in type requiring short answers;
 - carries 30 marks.
- Section B
 - consist of four to six compulsory structured questions;
 - carries 30 marks.
- Section C
 - candidates to choose two out of four long questions;
 - carries 40 marks.

The minimum mathematical requirements of the syllabus are similar to those for the SEC examination in Chemistry and additional mathematical skills will be indicated in the relevant section of the syllabus. Questions will be set in SI units except that pressures may be expressed in atmospheres (atm.). The Periodic Table, complete with atomic numbers and relative atomic masses, will be provided.

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

- Molar Gas Constant, $R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$
- Molar volume at STP = 22.40 L
- Avogadro's Constant = $6.02 \times 10^{23} \text{ mol}^{-1}$
- Ionic product of water, $K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$
- 1 atm = 101325 Pa
- $0^\circ\text{C} = 273 \text{ K}$