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**PEPERIKSAAN
SIJIL TINGGI PERSEKOLAHAN MALAYSIA
(MALAYSIA HIGHER SCHOOL CERTIFICATE)**

CHEMISTRY

Syllabus

Second Edition

This syllabus applies for the 1999 examination and thereafter until further notice. However the form of examination for Chemistry stated in this booklet was first implemented in the 2001 examination as announced through a circular, Pemberitahuan MPM/2(AM)/200. The Teachers/candidates are advised to contact Majlis Peperiksaan Malaysia for the latest information about the syllabus.



MAJLIS PEPERIKSAAN MALAYSIA
(MALAYSIAN EXAMINATION COUNCIL)

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FALSAFAH PENDIDIKAN KEBANGSAAN

Pendidikan di Malaysia ialah suatu usaha yang berterusan ke arah memperkembang potensi individu secara menyeluruh dan bersepadu untuk melahirkan insan yang seimbang dan harmonis dari segi intelek, rohani, emosi, dan jasmani berdasarkan kepercayaan dan kepatuhan kepada Tuhan. Usaha ini bertujuan untuk melahirkan warganegara Malaysia yang berilmu pengetahuan, berketrampilan, berakhlak mulia, bertanggungjawab, dan berkeupayaan mencapai kesejahteraan diri serta memberikan sumbangan terhadap keharmonian dan kemakmuran keluarga, masyarakat, dan negara.

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962 CHEMISTRY

Aims

This syllabus aims to enhance students' knowledge and understanding of chemistry to enable them to further their studies at institutions of higher learning or to assist them to embark on related careers, and also to promote an awareness among them of the role of chemistry in the universe.

Objectives

The objectives of this syllabus are to enable the students to

- (a) know the facts, terminology, and principles of chemistry;
- (b) interpret phenomena by using models, laws, and chemical principles;
- (c) understand, interpret, and use scientific information presented in various forms;
- (d) solve various problems and situations objectively;
- (e) analyse, synthesise, evaluate, and think about information and ideas logically and critically;
- (f) plan and carry out experiments scientifically, and to make deductions;
- (g) use scientific equipment properly and safely;
- (h) develop attitudes and values which are in line with scientific studies and practices, and to stimulate interest in and care for the environment.

Content

1 MATTER (10 periods)

- 1.1 (a) Electrons, protons, and neutrons as fundamental particles which are very important in chemistry; relative charge and relative masses of these particles
- (b) Proton number, nucleon number, and isotopes
Note: Previously proton number and nucleon number were known as atomic number and mass number respectively.
- 1.2 Relative atomic masses, A_r , relative molecular masses, M_r , relative formula masses based on the ^{12}C scale
- 1.3 Determination of the relative masses A_r and M_r by mass spectrometry method
- 1.4 Mole, the Avogadro constant, and applications of the mole concept to gases and solutions
- 1.5 States of matter
 - 1.5.1 Gases
 - (a) The kinetic theory of gases – Boyle's law, Charles' law, and Dalton's law
 - (b) The ideal gas equation $pV = nRT$ and its use
 - (c) Deviations from ideality

1.5.2 Liquids

The kinetic concept of liquid state and simple kinetic molecular model for melting, vaporisation, and vapour pressure of pure liquids

1.5.3 Solids

(a) The structure of lattice, allotrope, and crystal system

(b) The changes in states of matter, phase diagrams of H₂O and CO₂. The processes of vaporisation, boiling, sublimation, freezing, melting, and critical points

Explanatory notes

Candidates should be able to

1. explain the properties of protons, neutrons, and electrons in terms of their relative charges and relative masses
2. predict the behaviour of beams of protons, neutrons, and electrons in both electric and magnetic fields
3. explain the distribution of mass and charges within an atom
4. deduce the numbers of protons, neutrons, and electrons present in both neutral and charged species of a given proton (atomic) number and nucleon (mass) number
5. explain the contribution of protons and neutrons to atomic nuclei in terms of proton number and nucleon number
6. distinguish between isotopes based on the number of neutrons present; state examples of both stable and unstable isotopes
7. define the terms relative atomic, isotopic, molecular, and formula masses based on the ¹²C scale
8. interpret mass spectra in terms of relative abundance of isotopes and molecular fragments
9. calculate relative atomic mass of an element from the relative abundance of its isotopes or its mass spectrum
10. define the term mole in terms of the Avogadro constant
11. calculate the number of moles of reactants, volumes of gases, volumes of solutions, and concentration of solutions leading to stoichiometric deduction
12. explain the pressure and behaviour of ideal gas using the kinetic theory
13. explain qualitatively, in terms of molecular size and intermolecular forces, the conditions necessary for a gas approaching the ideal behaviour
14. define Boyle's law, Charles' law, and Dalton's law
15. calculate the partial pressure of a gas and its composition based on Dalton's law
16. use the $pV = nRT$ equation in calculations; including the determination of the relative molecular mass, M_r
17. explain the limitation of ideality at very high pressures and very low temperatures
18. explain the kinetic concept of the liquid state
19. explain melting, vaporisation, and vapour pressure using simple kinetic molecular theory
20. define the boiling point and freezing point of liquids
21. explain qualitatively the properties of solid in terms of the arrangement of particles in three dimensions and the repeated pattern of unit cells

22. explain the terms lattice, unit cell, and allotrope of carbon (including fullerenes) and sulphur
23. identify the properties of the seven basic crystal structures: cubic, hexagonal, monoclinic, orthorhombic, rhombohedral, tetragonal, and triclinic, with suitable examples
24. sketch the phase diagram for water and carbon dioxide, and explain the anomalous behaviour of water
25. explain phase diagrams as graphical plots of experimentally determined results
26. interpret phase diagrams as curves describing the condition of equilibrium between phases and as regions each representing a single phase
27. predict how a phase may change with changes in temperature and pressure
28. discuss vaporisation, boiling, sublimation, freezing, melting, and critical points of H₂O and CO₂
29. explain the use of dry ice in industry

2 ELECTRONIC STRUCTURE OF ATOMS (5 periods)

- (a) The concept of electronic energy levels treated qualitatively and illustrated by the characteristic line spectra of atomic hydrogen in the Lyman series
- (b) Atomic orbitals. The number and relative energies of the *s*, *p*, and *d* orbitals for the principal quantum numbers 1, 2, and 3: the shape and symmetry of the *s* and *p* orbitals
- (c) The filling of the orbitals according to their energy and the pairing of electrons

Explanatory notes

Candidates should be able to

1. explain the formation of the spectrum of atomic hydrogen
2. calculate the ionisation energy of an atom from the Lyman series converging limit
3. describe the number and relative energies of the *s*, *p*, and *d* orbitals for the principal quantum numbers 1, 2, and 3 including the 4*s* orbitals
4. describe the shape of the *s* and *p* orbitals
5. predict the electronic configuration of atoms and ions given the proton number (and charge)
6. explain and use the Hund's rule and the Pauli Exclusion Principle in the filling of orbitals

3 THE PERIODIC TABLE (10 periods)

- 3.1 Development of the Modern Periodic Table
- 3.2 Building of the Periodic Table based on the proton number and electronic configuration of the elements
Note: IUPAC numbering for groups in the Periodic Table, i.e. Groups 1 to 18, is used.
- 3.3 Elements in Groups 1, 2, 13, 14, 16, 17, and 18 and Periods 1 to 4
- 3.4 Classification of elements into the *s*, *p*, *d*, and *f* blocks
- 3.5 Variation in the physical properties with proton number across Period 2 and Period 3 in terms of atomic radius, melting point, boiling point, enthalpy change of vaporisation, electrical conductivity, ionisation energy, and electronegativity
- 3.6 Variation in the physical properties of the first row *d*-block elements in terms of melting point, density, and successive ionisation energies

Explanatory notes

Candidates should be able to

1. explain the development of the Periodic Table by Newlands, Mendeleev, and Moseley
2. use the aufbau principle and subsequently explain the electronic configuration of atoms with proton numbers 1 to 30 in the Periodic Table
3. identify elements in Groups 1, 2, 13, 14, 15, 16, 17, and 18, and elements in Periods 1 to 4
4. explain the position of elements in the Periodic Table in
 - (a) block *s*, with outer shell configurations s^1 and s^2
 - (b) block *p*, with outer shell configurations from s^2p^1 to s^2p^6
 - (c) block *d*, with outer shell configurations from d^1s^2 to $d^{10}s^2$
5. explain the position of *f*-block elements in the Periodic Table
6. interpret and explain the trend and gradation of atomic radii, melting points, boiling points, enthalpy changes, vapourisation, and electrical conductivities in terms of structure and bonding
7. explain the factors influencing ionisation energies
8. explain the trend in ionisation energy across Period 2 and Period 3 and down a group
9. predict the electronic configuration and position of unknown elements in the Periodic Table from successive values of ionisation energies
10. explain the almost similar physical properties such as density and melting point, in terms of bonding, metallic and ionic radii
11. explain the variation in successive ionisation energies

4 CHEMICAL BONDING (15 periods)

4.1 Electrovalent/ionic bonding

4.2 Covalent bonding

- (a) Covalent bonding in the SO_4^{2-} , CO_3^{3-} , NO_3^- , and CN^- ions
- (b) Hybridisation of the *s* and *p* orbitals for the C, N, and O atoms
- (c) The existence of ionic properties in molecules and covalent properties in ionic compounds
- (d) Co-ordinate covalent bonding
- (e) The repulsion theory between electron pairs and predictions of the shape of molecules and ions

4.3 Metallic bonding

4.4 Intermolecular forces between molecules

- (a) van der Waals forces: permanent dipoles and induced dipoles
- (b) Hydrogen bonding and its effect on physical properties

Explanatory notes

Candidates should be able to

1. explain electrovalent and covalent bonding in terms of 'dot and cross' diagrams

2. explain the Lewis structure of SO_4^{2-} , CO_3^{2-} , NO_3^- , and CN^- ions
3. predict and explain the shape of molecules and ions using the principle of electron pairs repulsion, e.g. linear, trigonal planar, tetrahedral, trigonal bipyramid, octahedral, V-shaped, and pyramid
4. explain the concept of overlapping and hybridisation of the s and p orbitals for the C, N, and O atoms in the CH_4 , C_2H_4 , C_2H_2 , NH_3 , and H_2O molecules
5. explain the differences in the bond angles in the water, ammonia, and methane molecules
6. explain the existence of polar and non-polar bonding in molecules which contain the C–Cl, C–N, C–O, C–Li, C–Si bonds, and explain the covalent properties of ionic compounds such as Al_2O_3 , AlI_3 , and LiI
7. explain the existence of co-ordinate bonding as exemplified by H_3O^+ , NH_4^+ , Al_2Cl_6 , and $[\text{Fe}(\text{CN})_6]^{3-}$
8. explain typical properties associated with electrovalent and covalent bonding
9. explain hydrogen bonding, van der Waals forces, and metallic bonding
10. explain metallic bonding through overlapping of orbitals
11. explain the formation of conduction and valence bands
12. distinguish between conductors, insulators, and semiconductors (Si and Ge) in terms of the location of conduction and valence bands
13. deduce the effect of intermolecular forces between molecules on the physical properties of substances
14. deduce the effect of hydrogen bonding on the physical properties of substances including organic substances
15. deduce the types of bonding present in substances from the given information

5 REACTION KINETICS (11 periods)

- 5.1 Rate of reaction
- 5.2 Collision theory
- 5.3 Rate law
- 5.4 The effect of temperature on rate constants, rates of reaction, and activation energy; Arrhenius's equation and the Boltzmann distribution
- 5.5 The role of catalysts in reactions
- 5.6 The order of reaction and rate constants for zero-, first-, and second-order reactions
- 5.7 The determination of the orders of reaction and the rate constants
- 5.8 Half-life of first-order reactions

Explanatory notes

Candidates should be able to

1. explain and use the terms rate of reaction, rate equation, order of reaction, rate constant, half-life of a first-order reaction, rate determining step, activation energy, and catalyst
2. explain qualitatively, in terms of collision theory, the effects of concentration and temperature on the rate of a reaction

3. calculate a rate constant from initial rates
4. predict an initial rate from rate equations and experimental data
5. suggest an experimental technique for studying the rate of a given reaction
6. calculate $t_{1/2}$ for a first-order reaction
7. use integrated forms of rate equations to determine zero-, first-, and second-order reactions involving a single reactant
8. deduce the order of a reaction by the initial rates method; deduce zero-, first-, and second-order reactions; deduce the order of reaction from concentration-time graphs
9. explain the relationship between the rate constants with the activation energy and temperature using Arrhenius's equation
10. use the Boltzmann distribution to explain the distribution of molecular energy
11. explain the effect of catalysts on the rate of a reaction
12. explain how a reaction, in the presence of a catalysts, follows an alternative path with a lower activation energy
13. explain enzymes as biological catalysts

6 EQUILIBRIA (36 periods)

6.1 Chemical equilibria

- 6.1.1 Reversible reaction, equilibria
- 6.1.2 Mass action law and derivation of equilibrium constants
- 6.1.3 Homogenous and heterogeneous equilibrium constants, K_p and K_c
- 6.1.4 Factors affecting chemical equilibria: Le Chatelier's principle. The following examples are useful.
 - (a) The synthesis of hydrogen iodide
 - (b) The dissociation of dinitrogen tetraoxide
 - (c) The hydrolysis of simple esters
 - (d) The Contact process
 - (e) The synthesis of ammonia
- 6.1.5 Equilibrium constant in terms of partial pressures and concentrations. The qualitative effect of temperature on equilibrium constants

6.2 Ionic Equilibria

- 6.2.1 The Arrhenius, Brønsted-Lowry, and Lewis theories of acids and bases
- 6.2.2 The degree of dissociation of weak acids and bases as the basis of strong/weak electrolytes
- 6.2.3 Dissociation constants K_a , K_b , K_w , pH, pOH, pK_a , pK_b , and pK_w
- 6.2.4 Titration indicators as acids or bases
- 6.2.5 Buffer solutions
- 6.2.6 Heterogeneous equilibria of ions, K_{sp} , and the common ion effect

6.3 Phase Equilibria

6.3.1 Liquid mixtures

- (a) Mixtures of two miscible liquids: Raoult's law applied to vapour pressures of miscible liquids; boiling point/composition curves; fractional distillation; azeotropic mixtures
- (b) Mixture of two immiscible liquids. Steam distillation

6.3.2 Distribution of solute (including gases) between two phases

- (a) Partition coefficient: solvent extraction illustrated by ether extraction. (A suitable example is the distribution of an organic acid in 2-methyl-1-propanol/water)
- (b) Chromatography: a qualitative treatment using the adsorption and/or partition theory

Explanatory notes

Candidates should be able to

1. explain a reversible reaction and dynamic equilibrium in terms of forward and backward reactions
2. calculate the quantities present at equilibrium from given appropriate data
3. deduce expressions for equilibrium constants in terms of concentrations, K_c , and partial pressures, K_p
4. calculate the value of the equilibrium constants in terms of concentrations or partial pressures from appropriate data
5. state Le Chatelier's principle and use it qualitatively from given information
6. discuss the effect of catalysts, or changes in concentration, pressure, or temperature on a system at equilibrium, using Le Chatelier's principle
7. state examples of equilibria which are referred to and studied in general
8. explain the effect of temperature on equilibrium constants from the equation
$$\ln K = -\frac{\Delta H}{RT} + C$$
9. show awareness of the importance of an understanding of chemical equilibrium in chemical industry
10. use Arrhenius, Brønsted-Lowry, and Lewis theories to explain acids and bases
11. identify conjugate acids and bases
12. explain qualitatively the different properties of strong and weak electrolytes
13. explain and use the terms pH, pOH, K_a , pK_a , K_b , and pK_b
14. calculate pH from the H_3O^+ ion concentration for acids (monobasic) and strong and weak bases
15. explain changes in pH during acid-base titrations in terms of strengths of acids and bases
16. select suitable indicators for acid-base titrations
17. explain the significance of the ionic product of water, K_w , and its use in calculations
18. define buffer solutions
19. explain the use of buffer solutions and their importance in biological systems

20. calculate the pH of buffer solutions from given appropriate data
21. explain and use the term solubility product, K_{sp}
22. calculate K_{sp} from given concentrations, and *vice versa*
23. explain the common ion effect including buffer solutions
24. predict the possibility of precipitation from given data of concentration of solutions
25. state and apply Raoult's law
26. explain qualitatively the effect of a non-volatile solute on the vapour pressure of a solvent, and hence on its melting point and boiling point
27. describe qualitatively the relationship between boiling point, enthalpy change of vaporisation, and intermolecular forces
28. interpret the boiling point-composition curves for mixtures of two miscible liquids in terms of 'ideal' behaviour or positive or negative deviations from Raoult's law
29. explain and use the term azeotropic mixture
30. explain the limitations on the separation of two components forming an azeotropic mixture
31. explain the principles involved in fractional distillation of ideal liquid mixtures
32. explain qualitatively the advantages and disadvantages of fractional distillation under reduced pressure
33. explain the steam distillation of two immiscible liquids
34. explain the use of these methods of distillation in industries and the separation of liquid mixture components in the laboratory
35. calculate composition, vapour pressure, and relative molecular mass
36. define the term partition coefficient
37. explain how solvent extraction works in the extraction of ether
38. calculate the partition coefficient for systems in which the solute is in the same molecular state in the two solutions
39. describe briefly and qualitatively paper, column, thin layer, and gas/liquid chromatography in terms of adsorption and/or partition
40. describe the use of chromatography as an analytical tool
41. describe the extensive use of these methods and gas chromatography in industry and medicine

7 **ELECTROCHEMISTRY** (13 periods)

- 7.1 Oxidation and reduction reactions in the Daniell cell
- 7.2 Half-reaction and half-cell. Redox potential
- 7.3 Redox reactions and electromotive force (e.m.f) – of cells and the standard hydrogen electrode
- 7.4 Electrode potentials and the formation of the electrochemical series
- 7.5 Cell potentials from the combination of various electrode potentials; spontaneous and non-spontaneous electrode reactions
- 7.6 Nernst equation and its use
- 7.7 The use and principles of electrochemistry in the prevention of corrosion

- 7.8 The mechanism of electrolysis as opposed to an electrochemical cell
- 7.9 Faraday's first and second laws and their use
- 7.10 Extraction and manufacturing of aluminium; recycling of aluminium
- 7.11 Manufacturing of chlorine by the electrolysis of brine either by mercury cathode cell or diaphragm cell

Explanatory notes

Candidates should be able to

1. explain the functions and cell diagram of the Daniell cell
2. construct redox equations using half-equations
3. explain the standard hydrogen electrode
4. calculate the (e.m.f.) of a cell using the E° values and write the redox equations
5. explain the methods used to determine standard electrode redox potentials
6. predict the stability of aqueous ions from E° value
7. determine the direction of electron flow in a simple cell of given electrode potentials
8. predict the feasibility of a reaction from E°_{cell} value
9. calculate E°_{cell} from the concentration value of solutions using the Nernst equation
10. explain the electrochemistry principle in the prevention of rusting and in dental filling
11. describe the importance of the development of better improved batteries for electric cars in terms of smaller size, lower mass, and higher voltage
12. calculate the quantity of product liberated during electrolysis
13. predict the product liberated during electrolysis based on electrochemistry
14. define and explain the relationship between the Faraday constant, the Avogadro constant, and the charge on the electron
15. calculate the number of coulomb used, the mass of material, and/or gas volume liberated during electrolysis
16. explain the uses of electrolysis in the manufacturing of aluminium, treatment of effluent (Ni, Cr, and Cd), electroplating of plastics, and anodisation
17. explain the advantages of recycling aluminium compared with extracting aluminium
18. explain the manufacturing of chlorine by the electrolysis of brine

8 THERMOCHEMISTRY AND CHEMICAL ENERGETICS (7 periods)

- 8.1 Enthalpy changes, ΔH , of reactions, formation, combustion, neutralisation, hydration, fusion, atomisation, solution
- 8.2 Hess' law
- 8.3 Lattice energies for simple ionic crystals. A qualitative appreciation of the effects of ionic charge and ionic radius on the magnitude of lattice energy
- 8.4 The Born-Haber cycle for the formation of simple ionic crystals and their aqueous solution
- 8.5 The solubility of solids in liquids

Explanatory notes

Candidates should be able to

1. explain that most chemical reactions are accompanied by energy changes, principally in the form of heat energy; the energy changes can be exothermic or endothermic
2. calculate the heat energy change from experimental measurements using the relationship: energy change = $mc\Delta T$
3. define the terms: enthalpy change of formation, combustion, hydration, solution, neutralisation, atomisation, ionisation energy
4. explain the terms enthalpy change of reaction and standard conditions
5. calculate enthalpy changes from experimental results
6. state Hess' law and its use to find enthalpy changes that cannot be determined directly, e.g. an enthalpy change of formation from enthalpy changes of combustion
7. construct energy level diagrams relating the enthalpy to reaction path and activation energy
8. define lattice energy for simple ionic crystals in terms of the change from gaseous ions to solid lattice
9. explain qualitatively the effects of ionic charge and ionic radius on the numerical magnitude of lattice energy values
10. construct energy cycles for the formation of simple ionic crystals and their aqueous solutions
11. calculate enthalpy changes from energy cycles
12. explain qualitatively the influence on solubility of the relationship between enthalpy change of solution, lattice energy of solid, and enthalpy change of hydration, or other solvent/solute interaction

9 PERIOD 3 AND GROUP 2: SELECTED PROPERTIES (4 periods)

- 9.1 Reaction of Period 3 elements, sodium to argon, with oxygen and water
- 9.2 Acid/base properties of oxides and hydrolysis of oxides (reaction with water)
- 9.3 Trends in the properties of the nitrates, carbonates, hydroxides, and sulphates of Group 2 elements

Explanatory notes

Candidates should be able to

1. describe the reactions of Period 3 elements with oxygen and water
2. interpret the ability of elements to act as oxidising and reducing agents
3. describe and explain the acid/base properties of the oxides of Period 3 elements
4. describe the properties and the reactions with water of the oxides of Period 3 elements
5. describe the classification of the oxides of Period 3 elements as basic, amphoteric, or acidic based on their reactions with water, acid, and alkali
6. describe the thermal decomposition of the nitrates, carbonates, and hydroxides of Group 2 elements
7. interpret qualitatively and explain the thermal decomposition of the nitrates, carbonates, and hydroxides in terms of the charge density and polarity of large anions

- interpret and explain qualitatively the variation in solubility of sulphate, in terms of the relative magnitudes of the enthalpy change of hydration for the relevant ions, and the corresponding lattice energy
- interpret, and make predictions from, the trends in physical and chemical properties of Group 2 compounds

10 GROUP 13: ALUMINIUM (4 periods)

- Extraction (refer to 7.10)
- Bonding and properties of the oxides and chlorides related to charge density of the Al^{3+} ions
- Acidic character of aqueous aluminium salts and their reactions with concentrated aqueous alkalis
- Resistance of aluminium to corrosion
- Uses of aluminium and its compounds in industry

Explanatory notes

Candidates should be able to

- explain the electrolytic method of extraction of aluminium from pure bauxite
- explain the relationship between charge density of the Al^{3+} ions and their bonding properties, and their oxides and chlorides
- explain aluminium's affinity for oxygen in the Thermite process
- explain the acidic character of aqueous aluminium salts and their reactions with concentrated aqueous alkalis
- explain why the aluminium metal resists corrosion
- describe some important uses of the aluminium metal in relation to its resistance to corrosion, good electrical and thermal conductivity, and low density
- describe the uses of compounds containing aluminium in modern industry, and examples of such compounds like alloys, alums, zeolites, and ceramics

11 GROUP 14: C, Si, Ge, Sn, Pb (9 periods)

- General study of the chemical trends for Group 14 elements and their oxides and chlorides
- Bonding, molecular form, sublimation, thermal stability, and hydrolysis of tetrachlorides, including mixed chlorides of carbon (freon) (refer to 17.7)
- Bonding, acid/base nature, and thermal stability of the oxides of oxidation states II and IV
- Relative stability for higher and lower states of oxidation for elements from carbon to lead in their oxides, chlorides, and aqueous cations
- Structure of the covalent crystal of carbon, e.g., diamond and graphite
 - Use of carbon in industry

- 11.6 (a) Use of silicon and silicate in industry
(b) Structure of silicate
(c) Quartz glass and the effect of adding metallic and non-metallic oxides
- 11.7 Use of tin alloys

Explanatory notes

Candidates should be able to

1. explain the changing trends in chemical properties of Group 14 elements and their oxides and chlorides
2. describe and explain the bonding in and molecular shape of the tetrachloride
3. explain the formation of freon and the ill-effects of its use
4. explain the volatility, thermal stability, and hydrolysis of the chloride in terms of structure and bonding
5. describe and explain the bonding, acid-base nature, and the stability of the oxides of oxidation states II and IV
6. describe and explain the relative stability of higher and lower oxidation states of the elements in their oxides and aqueous cations
7. explain the existence of different coordination numbers
8. explain the descending property of catenation in the group
9. explain the relationship between the structure of covalent crystal for graphite and diamond and their uses
10. explain the uses of carbon as a component in composite material (such as bullet-proof jackets, fibre, and amorphous carbon)
11. explain the use of silicone as a semiconductor and silicone as a fluid, elastomer, and resin
12. explain the importance of silicate as a basic material for cement, glass, and ceramics
13. explain silicate, SiO_4^{4-} , as a primary structural unit and connect these units to form chains (pyroxenes and amphiboles), sheets (mica), and framework structure (quartz)
14. outline the structure of kaolinite and montmorillonite
15. explain the composition of soda glass, borosilicate glass, quartz glass, and lead glass
16. explain the effects of adding metallic and non-metallic oxides to glass
17. describe the use of tin alloys in the production of solder and pewter

12 GROUP 15: NITROGEN AND ITS COMPOUNDS (9 periods)

12.1 Nitrogen

- (a) The inert nature of nitrogen
- (b) Principle of nitrogen extraction from air

12.2 Ammonia

- (a) Its formation from ammonium salts (refer to practical chemistry)
- (b) Its properties as a base
- (c) Its uses particularly in the manufacture of nitric acid and fertilisers

12.3 Oxides of nitrogen (NO, NO₂, N₂O₄)

- (a) Formation of the oxides of nitrogen in internal combustion engines and from lightning
- (b) Use of catalytic converters to reduce air pollution

Explanatory notes

Candidates should be able to

1. explain the bonding in the nitrogen molecules
2. explain the inertness of nitrogen in terms of high bonding energy as well as the non-polar nature of nitrogen molecules
3. show knowledge of the principles of extraction of nitrogen from air
4. explain the formation of ammonia from ammonium salts
5. explain the properties of ammonia as a base
6. explain the importance of ammonia particularly in the manufacture of nitric acid and fertilisers
7. explain the structure of the oxides of nitrogen (NO, NO₂, N₂O₄)
8. explain how the oxides of nitrogen are produced in internal combustion engines and from lightning
9. explain the function of catalytic converters to reduce air pollution by oxidising CO to CO₂ and reducing NO_x to N₂
10. explain the effects of the oxides of nitrogen and sulphur on air pollution

13 GROUP 17: Cl, Br, I (6 periods)

- 13.1 Variations in the volatility and colour intensity of the elements
- 13.2 Relative reactivity of the elements as oxidising agents
- 13.3 Reactions of the elements with hydrogen, and the relative stability of the hydrides (with reference to the reactivity of fluorine comparatively)
- 13.4 Reactions of halide ions
- 13.5 Reactions of chlorine with aqueous sodium hydroxide
- 13.6 Important uses of the halogens and halogen compounds

Explanatory notes

Candidates should be able to

1. state that the colour intensity of the halogens increase on going down the group in the Periodic Table
2. explain how the volatility of the halogens decrease on going down the group in the Periodic Table
3. explain the relative reactivity of these elements as oxidising agents
4. explain the order of reactivity between the halogens and hydrogen as well as relative stability of the halides
5. explain and write equations of reactions between halide ions with aqueous silver ions followed by aqueous ammonia (refer to practical chemistry) and with concentrated sulphuric acid

6. explain the reactions of chlorine with cold and hot aqueous sodium hydroxide
7. describe the important uses of the halogens and halogen compounds as antiseptic, bleaching agent, in purifying water, and in black-and-white photography (explained as an example of redox reaction)

14 AN INTRODUCTION TO THE CHEMISTRY OF *d*-BLOCK ELEMENTS (12 periods)

- 14.1 Chemical properties of *d*-block elements (Ti, Mn, Cr, and Fe)
 - (a) Variable oxidation states and colour changes (refer to 7)
Note: No discussion of the zero and negative oxidation states is required.
 - (b) Formation of complex ions by exchange of ligands
 - (c) Catalytic properties
 - (i) Heterogeneous catalysis
 - (ii) Homogeneous catalysis
- 14.2 Nomenclature of complex compounds
- 14.3 Bonding in complex ions
- 14.4 Stereoisomerism of complex ions and compounds (refer to 15.3)
- 14.5 Use of *d*-block elements and their compounds, e.g. chromium, manganese, cobalt, and titanium(IV) oxide

Explanatory notes

Candidates should be able to

1. explain variable oxidation states in terms of the energies of *3d* and *4s* orbitals
2. explain the colour changes of transition metal ions in terms of a half-filled *3d* orbitals and state the colours of the complex aqueous ions of elements
3. state the principal oxidation numbers of these elements in their common cations, oxides, and oxo ions
4. explain qualitatively the relative stabilities of these oxidation states
5. explain the use of standard redox potentials in predicting the relative stabilities of aqueous ions
6. explain the terms complex ion and ligand
7. explain the formation of complex ions by exchange of ligands, e.g. water, ammonia, cyanide ions, thiocyanide ions, ethanedioate ions, ethylenediaminetetraethanoate, and halide ions; examples: $[\text{Fe}(\text{CN})_6]^{4-}$, $[\text{Fe}(\text{CN})_6]^{3-}$, $[\text{Fe}(\text{H}_2\text{O})_5(\text{SCN})]^{2+}$
8. explain the catalytic activity of transition elements in terms of variable oxidation states (homogenous catalysis) or adsorption by coordinate bond information (heterogeneous catalysis)
9. explain heterogeneous catalysis, e.g. Ni and Pt in the hydrogenation of alkenes and Fe/Fe₂O₃ in the Haber process
10. explain the availability of the *d*-orbitals resulting in a different mechanism of lower activation energy
11. explain homogeneous catalysis in terms of the reduction of Fe³⁺ by iodide ions and the subsequent reoxidation by S₂O₈²⁻ ions
12. show understanding of the naming of complex compounds

13. discuss coordinate bond formation between ligands and the central metal ions, and state the types of ligands, i.e. monodentate, bidentate, and hexadentate
14. explain the phenomenon of stereoisomerism: (a) geometric isomers, e.g. *cis* and *trans* – $[\text{Cr}(\text{NH}_3)_4\text{Cl}_2]^+$, *mer* and *fac* – $[\text{Cr}(\text{NH}_3)_3\text{Cl}_3]$; (b) optical isomers, e.g. *d* and *l* – $[\text{Cr}(\text{C}_2\text{O}_4)_3]^{3-}$
15. explain the use of chromium in the hardening process of steel and in stainless steel, cobalt and manganese in alloys, and TiO_2 as a pigment, e.g. in plastics and paint

15 THE CHEMISTRY OF CARBON (10 periods)

- 15.1 Bonding of the carbon atoms: the shapes of the ethane, ethene, ethyne, and benzene molecules
- 15.2 General, empirical, molecular, and structural formulae for organic chemistry
- 15.3 Isomerism – structural, geometric, and optical
- 15.4 Classification based on functional groups (general formula)
- 15.5 Nomenclature and structural formulae for each functional/radical group (refer to their trivial names)
- 15.6 Nucleophile and electrophile
- 15.7 Structure and its effect on
 - (a) physical properties; e.g. boiling point, melting point, and solubility in water
 - (b) acidity and basicity

The effect of the structure and delocalisation of electrons on the relative acid or base strength, i.e. proton donors or acceptors, in ethylamine, phenylamine, ethanol, phenol, and chlorine-substituted ethanoic acids

Explanatory notes

Candidates should be able to

1. explain the concept of hybridisation in the bonding of carbon atoms with reference specially to carbon atoms which have a valency of four and the types of hybridisation such as the following:
sp – linear, *sp*² – triangular, *sp*³ – tetrahedral
2. describe the formation of σ and π bonds as exemplified by diagrams of the overlapping of orbitals in CH_4 , C_2H_4 , C_2H_2 , and C_6H_6 molecules
3. explain the concept of delocalisation of π electrons in benzene rings (aromaticity)
4. explain the meaning of general, empirical, molecular, and structural formulae of organic compounds
5. calculate empirical formulae and derive molecular formulae
6. interpret structural isomerism with reference to the ability of carbon atoms to link together with each other in a straight line and/or in branches
7. explain geometric/*cis-trans* isomerism in alkenes in terms of restricted rotation due to π bond/ $\text{C}=\text{C}$ bonds
8. explain the meaning of a chiral centre and how such a centre gives rise to optical isomerism
9. identify chiral centres and/or *cis-trans* isomerism in a molecule of given structural formula
10. determine the possible isomers for an organic compound of known molecular formula

11. describe the classification of organic compounds by functional groups and the nomenclature of classes of organic compounds according to the IUPAC system of the following classes of compounds:
 - (a) alkanes, alkenes
 - (b) haloalkanes
 - (c) alcohols (including primary, secondary and tertiary) and phenols
 - (d) aldehydes and ketones
 - (e) carboxylic acids and esters
 - (f) primary amines, amides, and amino acids
12. define the terms nucleophile and electrophile
13. describe the relationship between the size of molecules in the homologous series and the melting and boiling points
14. explain the attractive forces between molecules (van der Waals forces and hydrogen bonding)
15. explain the meaning of Lewis acids and bases in terms of charge/electron density
16. explain why many organic compounds containing oxygen/nitrogen which have lone pair electrons (as Lewis bases) form bonds with electron acceptors (as Lewis acids)
17. explain how nucleophiles such as OH^- , NH_3 , H_2O , Br^- , I^- and carbonion have Lewis base properties, whereas electrophiles such as H^+ , NO_2^+ , Br_2 , AlCl_3 , ZnCl_2 , FeBr_3 , BF_3 , and carbonium ions have Lewis acid properties
18. explain induction effect which can determine the properties and reactions of functional groups
19. explain how most functional groups such as $-\text{NO}_2$, $-\text{CN}$, $-\text{COOH}$, $-\text{COOR}$, $>\text{C}=\text{O}$, SO_3H , $-\text{X}$ (halogen), $-\text{OH}$, $-\text{OR}$, $-\text{NH}_2$, C_6H_5 are electron acceptors whereas functional groups such as $-\text{CH}_3$, $-\text{R}$ (alkyl or aryl) are electron donors
20. explain how the concept of induction can account for the differences in acidity between CH_3COOH , ClCH_2COOH , Cl_2CHCOOH , and Cl_3CCOOH ; between $\text{ClCH}_2\text{CH}_2\text{CH}_2\text{COOH}$ and $\text{CH}_3\text{CH}_2\text{CHClCOOH}$
21. use the concept of delocalisation of electrons to explain the differences in acidity between ethanol and phenol, as well as the differences in basicity between CH_3NH_2 and $\text{C}_6\text{H}_5\text{NH}_2$

16 HYDROCARBONS (14 periods)

16.1 Alkanes, exemplified by ethane

- (a) Free radical substitution, e.g. the effect of chlorination of hydrocarbons in water on the environment (refer to 13.5)
- (b) Free radical reactions illustrated by the reaction of methane with chlorine
- (c) Crude oil as a source of energy and chemicals; cracking reactions
- (d) The unreactivity of alkanes towards polar reagents
- (e) The effect of hydrocarbons on the environment

16.2 Alkenes, exemplified by ethene

- (a) Electrophilic addition
- (b) Addition by electrophilic reagents, illustrated by the reactions of bromine/ethene or hydrogen bromide/ethene

- (c) Simple tests for alkenes
 - (d) The importance of ethene in industry
- 16.3 Arenes, exemplified by benzene and methylbenzene (toluene)
- (a) Electrophilic substitution reactions
 - (b) Influence of substitution groups to the benzene ring – on further substitution. The effect of induction as exemplified by $-\text{OH}$, $-\text{Cl}$, $-\text{CH}_3$, $-\text{NO}_2$, $-\text{COCH}_3$, $-\text{NH}_2$
 - (c) Oxidation of the side-chain to produce RCOOH . Halogenation in the side-chain or aromatic nucleus depending on reaction conditions
 - (d) Effects of arene compounds on health

Explanatory notes

Candidates should be able to

1. explain alkanes as saturated aliphatic hydrocarbons
2. explain the construction of the alkane series (straight and branched) and IUPAC nomenclature of alkanes for C_1 to C_5
3. name alkyl groups derived from alkanes and identify primary, secondary, tertiary, and quaternary carbons
4. explain the homolytic cleavage of bonds which produces free radicals that determine the mechanism of a reaction
5. explain the halogenation of alkanes and its mechanism as well as the oxidation of alkane with limited and excess oxygen, and the use of alkanes as fuels
6. explain the mechanism of free radical substitution
7. explain the use of crude oil as source of aliphatic hydrocarbons
8. explain how cracking reactions can be used to obtain alkanes and alkenes of lower M_r from larger hydrocarbon molecules
9. show awareness of the environmental consequences of CO , CO_2 , and unburnt hydrocarbons arising from the internal combustion engine
10. define alkenes as unsaturated aliphatic hydrocarbons with one or more double bonds
11. name alkenes according to the IUPAC nomenclature and the common names for C_1 to C_5
12. explain the mechanism of electrophilic addition in alkenes
13. explain the chemistry of alkenes as exemplified by the following reactions of ethene:
 - (a) addition of hydrogen, steam, hydrogen halides, halogens (decolourisation) of bromine water and concentrated sulphuric acid
 - (b) oxidation – decolourisation of the manganate(VII) ions
 - (c) polymerisation
14. explain the use of bromination reaction and decolourisation of MnO_4^- ions as simple tests for alkenes and unsaturated compounds
15. explain briefly the preparation of chloroethane, epoxyethane, ethane-1,2-diol, and polyethene
16. explain the nomenclature of aromatic compounds derived from benzene and the use of *o* (*ortho*), *m* (*meta*), and *p* (*para*) or the numbering of substitution groups to the benzene ring

17. describe the mechanism of electrophilic substitution in arenes illustrated by the nitration of benzene
18. explain the chemistry of arenes as exemplified by substitution reactions of halogen, HNO_3 , $\text{CH}_3\text{I}/\text{AlCl}_3$, CH_3COCl , and SO_3 with benzene and methylbenzene (toluene)
19. explain why the benzene ring is more stable against oxidants like KMnO_4 and $\text{K}_2\text{Cr}_2\text{O}_7$ compared with aliphatic alkenes
20. explain how a reaction between an alkylbenzene and an oxidant such as acidified KMnO_4 will cause the carbon atoms which are joined directly to the benzene ring to become a carboxyl group, and the remaining alkyl chains will turn into water and CO_2
21. determine the products of halogenation of methylbenzene (toluene) with Lewis acid catalysts such as AlCl_3 and FeCl_3 and in the presence of light only
22. predict the reaction products when the substitution group in benzene is an electron accepting or donating group
23. explain the uses of arene compounds as solvents
24. explain that arene compounds are carcinogenic

17 HALOALKANES (6 periods)

- 17.1 (a) Nucleophilic substitution reactions
(b) Reactivity of chlorobenzene compared with chloroalkanes
- 17.2 Nucleophilic substitution at the carbon atom in haloalkanes
- 17.3 Relative reactivity of primary, secondary, and tertiary alkyl halides
- 17.4 Elimination reactions
- 17.5 Haloalkanes as the starting material in the preparation of organometals
- 17.6 Uses of haloalkanes
- 17.7 Effects of haloalkanes on the environment

Explanatory notes

Candidates should be able to

1. explain the chemistry of haloalkanes as exemplified by
 - (a) nucleophilic substitution reactions represented by the hydrolysis of bromoethane, the formation of nitriles, and the formation of primary amines by reaction with NH_3
 - (b) haloalkane elimination reactions which produce alkenes
2. explain the mechanism of nucleophilic substitution in haloalkanes
3. explain the difference in the reactivity of chlorobenzene compared with chloroalkanes, with particular reference to hydrolysis reactions
4. explain the relative reactivity of primary, secondary, and tertiary alkyl halides
5. explain the use of haloalkanes in the synthesis of organometal compounds such as the following:
 - (a) Reactions with lithium and magnesium to give organolithium and organomagnesium (Grignard reagents) compounds, and their respective uses in reactions with carbonyl compounds, silicon halides (SiCl_4), and tin halides (SnCl_4)

- (b) Reactions of chloroethane with Na/Pb alloy to give tetraethyllead(IV)
- 6. explain the uses of fluoroalkanes and chlorofluoroalkanes as inert substances for aerosol propellants, detergents, coolants, fire-extinguishers, and insecticides (DDT)
- 7. explain the bad effects of DDT and chlorofluoroalkanes on the environment, especially in the depletion of the ozone layer
- 8. explain the formation of ozone in the atmosphere and in factories
- 9. explain the role of ozone in the atmosphere, i.e. the ozone layer as an absorber of ultraviolet light

18 HYDROXY COMPOUNDS (10 periods)

- 18.1 Classification of hydroxy compounds into primary, secondary, and tertiary alcohols
- 18.2 Reactions of hydroxy compounds
 - (a) Reactions of alcohols as exemplified by ethanol
 - (b) Use of triiodomethane test
 - (c) Comparison of the acidity of phenol, alcohol, and water (refer to 15.7)
 - (d) Reactions of phenol
- 18.3 Preparation and manufacture of hydroxy compounds
 - (a) Alcohol – in the laboratory and industry
 - (b) Phenol – in industry
- 18.4 Use of hydroxy compounds
 - (a) Alcohol
 - (b) Phenol in industry

Explanatory notes

Candidates should be able to

1. describe the classification of hydroxy compounds into primary, secondary, and tertiary alcohols, and explain the nomenclature of alcohols
2. explain that most alcohol reactions can be divided into two groups, i.e.
 - (a) the RO–H bond is broken and H is replaced by other groups because alcohol and phenol have the characteristics of weak acid (formation of hydroxides and phenoxides)
 - (b) the R–OH is broken and –OH is replaced by other groups caused by nucleophilic substitution reactions
3. explain the chemistry of alcohol compounds as exemplified by the following reactions.
 - (a) The formation of halogenoalkanes
 - (b) The formation of an alkoxide with sodium
 - (c) Oxidation to carbonyl compounds and carboxylic acids
 - (d) Dehydration to alkenes and ether
 - (e) Esterification
 - (f) Acylation

4. explain the difference in reactivity of primary, secondary, and tertiary alcohols as exemplified by
 - (a) the reaction rate of such alcohols to give haloalkanes
 - (b) the reaction products of $\text{KMnO}_4/\text{K}_2\text{Cr}_2\text{O}_7$ oxidation in the presence of sulphuric acid
5. explain the reactions of the alcohol with the structure

$$\begin{array}{c} \text{H} \\ | \\ \text{CH}_3 - \text{C} - \\ | \\ \text{OH} \end{array}$$
 with alkaline aqueous iodine to give triiodomethane
6. discuss the relative acidities of water, phenol, and ethanol with particular reference to the inductive effect and the effect of delocalised electrons
7. explain the reactions of phenol with sodium hydroxide, sodium, and acid chlorides
8. explain the reactions of phenol with bromine water and aqueous iron(III) chloride
9. describe the preparation and manufacture of alcohol from
 - (a) the hydration of ethane (in the laboratory and industry)
 - (b) natural gas and fermentation process (ethanol)
10. describe the process of manufacturing phenol involving the use of cumene
11. explain the uses of alcohols as antiseptic, solvent, and fuel
12. explain the use of phenol in the manufacture of cyclohexanol and hence nylon-6,6

19 CARBONYL COMPOUNDS (6 periods)

- 19.1 (a) Aldehydes, exemplified by ethanal (acetaldehyde) and benzaldehyde (phenylmethanal)
 - (b) Ketones, exemplified by propanone and phenylethanone
- 19.2 Preparation of carbonyl compounds
- 19.3 Chemical reactions
 - (a) Aldehydes
 - oxidation to carboxylic acids (with KMnO_4)
 - oxidation with Fehling's solution and Tollen's reagent
 - (b) Ketones
 - reaction with alkaline iodine as a triiodomethane test for methylketone
 - (c) Aldehyde and Ketone
 - reduction to alcohols
 - condensation with 2,4-dinitrophenylhydrazine
 - nucleophilic addition with hydrogen cyanide
- 19.4 Reducing sugar and carbohydrates as examples of natural compounds which have the functional groups $-\text{OH}$ and $>\text{C}=\text{O}$, e.g. glucose and cellulose, with cellulose as a polymer for glucose

Explanatory notes

Candidates should be able to

1. write the general formula for aldehydes and ketones, and name their compounds according to the IUPAC nomenclature and their common names
2. write the reaction equations for the preparation of aldehydes and ketones
3. determine the properties of an unknown carbonyl compound – aldehyde or ketone – on the results of simple tests (e.g. Fehling's solution, Tollens' reagent, and ease of oxidation)
4. explain the reaction of ketone compounds with the structure $\text{CH}_3-\text{C}=\text{O}$ with alkaline aqueous iodine to give triiodomethane
$$\begin{array}{c} \text{CH}_3-\text{C}=\text{O} \\ | \end{array}$$
5. explain the reduction reactions of aldehydes or ketones to primary or secondary alcohols through catalytic hydrogenation reaction and with LiAlH_4
6. explain the use of 2,4-dinitrophenylhydrazine reagent as a simple test to detect the presence of $>\text{C}=\text{O}$ groups
7. explain the mechanism of the nucleophilic addition reactions of hydrogen cyanide with aldehydes or ketones
8. explain that natural compounds such as glucose, sucrose, and other carbohydrates have the $>\text{C}=\text{O}$ group
9. explain the characteristics of glucose as a reducing sugar

20 CARBOXYLIC ACIDS (4 periods)

Carboxylic acids as exemplified by ethanoic acid and benzoic acid

- (a) Preparation of carboxylic acids by the oxidation of primary alcohols and hydrolysis of nitriles
- (b) Chemical reactions of carboxylic acids
- (c) Uses of carboxylic acids in industry

Explanatory notes

Candidates should be able to

1. name aliphatic and aromatic acids according to IUPAC nomenclature, and their common names for C_1 to C_6
2. write the reaction equations for the formation of carboxylic acids from alcohols, aldehydes, and nitriles
3. explain the acidic characteristics of carboxylic acids compared with HCl , HNO_3 , or H_2SO_4 and their reaction with bases to form salts
4. explain the substitution of the $-\text{OH}$ group by the nucleophiles $-\text{OR}$ and $-\text{Cl}$ to form esters and acyl chlorides respectively
5. explain the reduction of carboxylic acids to primary alcohols
6. explain the oxidation and dehydration of methanoic and ethanedioic acids
7. explain the various uses of carboxylic acids in foods, perfumes, polymers, and other industries

21 CARBOXYLIC ACID DERIVATIVES (6 periods)

21.1 Acyl chlorides

- (a) Relative reactivity – hydrolysis of acyl chlorides and alkyl chlorides
- (b) Chemical reactions with alcohols, phenols, and primary amines

21.2 Esters as exemplified by ethyl ethanoate and phenyl benzoate

- (a) Preparation of esters from carboxylic acids and acyl chlorides
- (b) Chemical reactions of esters
- (c) Fats and oils as natural esters: general structure of glycerol ester derivatives of given long-chain carboxylic acids
- (d) Uses of esters

21.3 Amides as exemplified by ethanamide

- (a) Preparation of amides
- (b) Hydrolysis of amides

Explanatory notes

Candidates should be able to

1. compare the ease of hydrolysis of acyl chlorides with that of alkyl chlorides
2. explain the formation of esters by the reactions of acyl chlorides with alcohols and phenols
3. explain the formation of amides by the reaction of acyl chlorides with primary amines
4. explain the acid and base hydrolysis of esters: saponification
5. explain the reduction of esters to primary alcohols
6. explain the importance of esters as fats and oils
7. explain the use of esters as taste enhancers, flavourings, preservatives, solvents, and in the production of polyesters
8. explain the nomenclature of amides according to the IUPAC system
9. explain the preparation of primary and secondary amides from reactions of acid chlorides with amines
10. explain the hydrolysis of amides in acidic or basic conditions

22 AMINES (4 periods)

Primary amines as exemplified by ethylamine and phenylamine

- (a) Preparation of amines (refer to 17.1)
- (b) Chemical reactions of amines with mineral acids, nitrous acid, and with aqueous bromine (refer to practical chemistry)

Explanatory notes

Candidates should be able to

1. name amines according to the IUPAC nomenclature and their common names

2. explain the preparation of ethylamine by the reduction of nitriles, phenylamine by the reduction of nitrobenzene, and amine by the nucleophilic substitution reaction between ammonia and alkyl halides
3. explain the formation of salts when amines react with mineral acids
4. differentiate between primary aliphatic amines and primary aryl (aromatic) amines by their respective reactions with nitrous acid and aqueous bromine
5. explain the structures of diazonium ion and diazonium salt
6. explain the formation of dyes by the coupling reaction of the diazonium salt (benzenediazonium chloride) and phenol

23 AMINO ACIDS AND PROTEINS (4 periods)

- 23.1 Amino acids as exemplified by aminoethanoic acid (glycine) and 2-aminopropanoic acid (alanine)
- (a) General formula and structure of amino acids
 - (b) Acid/base properties
 - (c) Formation of zwitterions
 - (d) Formation of the peptide linkage
- 23.2 Protein – structure and hydrolysis of proteins
- 23.3 Importance of amino acids and proteins

Explanatory notes

Candidates should be able to

1. explain the structure and general formula of amino acid with one amino group
2. identify the chiral centre in the amino acid molecules
3. explain the acid/base properties of amino acids
4. explain the formation of zwitterions
5. explain the peptide linkage as amide linkage formed by the condensation between amino acids as exemplified by glycylalanine and alanilglycine
6. explain that the protein structure is based on the peptide linkage
7. explain the hydrolysis of proteins
8. explain the roles of amino acids and proteins in the biological system

24 POLYMERS (9 periods)

- 24.1 Introduction – artificial/synthetic polymers, natural polymers, and copolymers
- 24.2 Preparation
- (a) Condensation polymerisation
 - (b) Addition polymerisation, including free radicals and ionic (cationic and anionic) reactions – use of Ziegler-Natta catalyst
- 24.3 Classification of polymers
- 24.4 Use of polymers in daily life

24.5 Natural rubber (2-methylbuta-1,3-diene) polymer or polyisoprene

- (a) Microstructure: *cis* and *trans* forms
- (b) Production of bulk rubber from latex
- (c) Vulcanisation
- (d) Uses of latex and bulk rubber

Explanatory notes

Candidates should be able to

1. explain the meaning of monomer, polymer, repeating unit, and copolymer
2. identify monomers in a polymer
3. explain polymerisation
4. explain condensation polymerisation and addition polymerisation
5. explain addition polymerisation which involves free radicals and ionic reactions mechanism
6. explain the use and role of the Ziegler-Natta catalyst in the addition polymerisation process
7. suggest polymers which can be prepared from a monomer or a pair of monomers
8. identify the manufacturing process for the preparation of polyethane/polyethylene/PE, polypropene/propylene/PP, polyphenylethene/polystyrene/PS, and SBR synthetic rubber, i.e. phenylethene-butadiene-1,3-diene/styrene-butadiene copolymer
9. explain the classification of polymers as thermosetting, thermoplastic, and elastomer depending on their thermal properties
10. explain the uses of polyethylene, polypropylene, polystyrene, and synthetic rubber (styrene-butadiene copolymer) in the making of domestic products
11. explain the isoprene monomer in natural rubber
12. explain the existence of two isomers in poly(2-methylbuta-1,3-diene)
 - (a) the elastic *cis* form (from the *Hevea brasiliensis* trees)
 - (b) the inelastic *trans* form (from the gutta-percha trees)
13. explain the process of producing bulk rubber from latex
14. explain the formation of cross-linkages via sulphur to change the physical properties of natural rubber
15. explain the use of latex in the making of gloves, threads, etc, and the use of bulk rubber in the making of tyres, belts, etc

Practical Syllabus

School-based Assessment of Practical (Paper 3)

School-based assessment of practical work will only be carried out during the school term of form six for candidates from government and private schools which have been approved by the Malaysian Examinations Council to carry out the school-based assessment. *Individual private candidates, candidates from private schools which have no permission to carry out the school-based assessment of practical work, candidates who repeat upper six (in government or private schools), and candidates who do not attend classes of lower six and upper six for two consecutive years (in government or private schools) are not allowed to take this paper.*

The Malaysian Examinations Council will specify 15 compulsory experiments to be carried out by candidates and to be assessed by subject teachers in schools. The experiments will cover three aspects, i.e. volumetric analysis, determination of physical quantities, and techniques. Details of the topic, aim, theory, apparatus, and method of each of the experiments will be compiled and distributed to all schools.

Candidates will be assessed based on the following.

- (a) The use and organisation of techniques, apparatus, and materials
- (b) Observations, measurements, and recording
- (c) The interpretation of experimental observations and data
- (d) The designing and planning of investigations
- (e) Scientific and critical attitudes

Written Practical Test (Paper 4)

Individual private candidates, candidates from private schools which have no permission to carry out the school-based assessment of practical work, candidates who repeat upper six (in government or private schools), and candidates who do **not** attend classes of lower six and upper six for two **consecutive** years (in government or private schools) are required to take this paper.

The written practical test will be made up of three structured questions to test the candidates' understanding of practical procedures in the laboratory. The Malaysian Examinations Council will not be strictly bound by the syllabus in setting questions. Where appropriate, candidates will be given sufficient information to enable them to answer the questions. Only knowledge of theory within the syllabus and knowledge of usual laboratory practical procedures will be expected.

Questions to be set will test candidates' ability to

- (a) record readings from diagrams of apparatus;
- (b) describe, explain, comment on, or suggest experimental arrangements, techniques, and procedures;
- (c) complete tables of data and/or plot graphs;
- (d) interpret, draw conclusions from, and evaluate observations and experimental (including graphical) data;
- (e) describe tests for gases, ions, oxidising and reducing agents, and/or make deductions from such tests;
- (f) perform simple calculations based on experiments.

Questions to be set in the written practical test will test candidates' knowledge and understanding in three aspects as follows.

(a) Volumetric analysis

Candidates should understand experiments involving the procedures and calculations of volumetric analysis for acid-base and redox titrations. Candidates should also know the calculations for purity determination and stoichiometry.

(b) Determination of physical quantities

Candidates should understand experiments involving the measurements of some quantities in the following topics: thermochemistry, reaction kinetics, equilibrium, solubility, and electrochemistry.

(c) **Techniques**

Candidates should know the following techniques: qualitative analysis of ions and functional groups, synthesis, and separation processes.

Although systematic analysis and a knowledge of traditional methods of separations will **not** be required, it will be assumed that candidates will be familiar with the simple reactions of the following ions: NH_4^+ , Mg^{2+} , Al^{3+} , Ca^{2+} , Cr^{3+} , Mn^{2+} , Fe^{2+} , Fe^{3+} , Ni^{2+} , Cu^{2+} , Zn^{2+} , Ba^{2+} , Pb^{2+} , CO_3^{2-} , NO_3^- , NO_2^- , S^{2-} , SO_4^{2-} , SO_3^{2-} , $\text{S}_2\text{O}_3^{2-}$, Cl^- , Br^- , I^- , MnO_4^- , CH_3CO_2^- , $\text{C}_2\text{O}_4^{2-}$.

A knowledge of simple organic reactions, e.g. test-tube reactions indicating the presence of unsaturation and functional groups will be required.

The substances to be asked in questions may contain ions not included in the above list; in such cases, candidates will **not** be expected to identify the ions but to draw conclusions of a general nature.

Form of Examination

Candidates are required to enter for Papers 1, 2, and either Paper 3 or Paper 4.

<i>Paper</i>	<i>Format of paper</i>	<i>Marks</i>	<i>Duration</i>
Paper 1	50 compulsory multiple-choice questions are to be answered.	50 (to be scaled to 60)	1¾ hours
Paper 2	<i>Section A:</i> 4 compulsory short structured questions are to be answered. <i>Section B:</i> 4 questions are to be answered out of 6 essay questions.	40 60 (15 per question) Total: 100 (to be scaled to 120)	2½ hours
Paper 3	School-based Assessment of Practical: 15 compulsory experiments are to be carried out.	20	School term
Paper 4	Written Practical Test: 3 compulsory structured questions are to be answered.	30 (to be scaled to 20)	1 hour

Summary of Important Quantities

The following list is intended as a guide to the important quantities which might be encountered in teaching and used in question papers. The list is not exhaustive.

<i>QUANTITY</i>	<i>COMMON SYMBOL</i>	<i>SI UNIT</i>
mass	m	kg, g
length	l	m
time	t	s
electric current	I	A
thermodynamic temperature	T	K
amount of substance	n	mol
<i>Other quantities</i>		
temperature	θ, t	°C, Kelvin
volume	V, v	m ³
density	ρ	kg m ⁻³ , g cm ⁻³
pressure	p	Pa
frequency	ν, f	Hz
wavelength	λ	m, mm, nm
speed of electromagnetic waves	c	m s ⁻¹
Planck constant	h	J s
electric potential difference	V	V
electrode potential (standard) redox	$(E^\ominus)E$	V
electromotive force	E	V
molar gas constant	R	J K ⁻¹ mol ⁻¹
heat capacity	C	J K ⁻¹
specific heat capacity	c	J K ⁻¹ kg ⁻¹
radioactive decay constant	λ	s ⁻¹
half-life	$T_{1/2}, t_{1/2}$	s
atomic mass	m_a	kg
relative atomic/isotopic mass	A_r	–
molecular mass	m	kg
relative molecular mass	M_r	–
molar mass	M	kg mol ⁻¹
nucleon number	A	–
proton number	Z	–
neutron number	N	–
number of molecules	N	–
number of molecules per unit volume	n	m ⁻³
Avogadro constant	L	mol ⁻¹
Faraday constant	F	C mol ⁻¹
enthalpy change of reaction	ΔH	J, kJ
molar enthalpy change of reaction	ΔH^\ominus	J mol ⁻¹ , kJ mol ⁻¹
ionisation energy	I	kJ mol ⁻¹
lattice energy	–	kJ mol ⁻¹
bonding energy	–	kJ mol ⁻¹
electron affinity	–	kJ mol ⁻¹
rate constant	k	as appropriate
equilibrium constant	K, K_p, K_c	as appropriate
acid (and base) dissociation constant	$K_a, (K_b)$	as appropriate

QUANTITY

COMMON SYMBOL

SI UNIT

order of reaction	n, m	
mole fraction	x	
concentration, molarity	c	$\text{mol dm}^{-3}, \text{M}$
mass density	c	kg dm^{-3}
partition coefficient	K	
degree of dissociation	α	
ionic product	K	as appropriate
solubility product	K_{sp}	as appropriate
ionic product of water	K_{w}	$\text{mol}^2 \text{dm}^{-6}$
pH	pH	

Note: A Data Booklet is available for use in paper 2. This booklet contains a copy of the Periodic Table.

The Periodic Table

Group																								
1 (I)	2 (II)	3	4	5	6	7	8	9	10	11	12	13 (III)	14 (IV)	15 (V)	16 (VI)	17 (VII)	18 (VIII)							
1.0 H 1		<table border="1" style="margin: auto;"> <tr> <td>a</td> <td>a = Relative atomic mass</td> </tr> <tr> <td>X</td> <td>X = Atomic symbol</td> </tr> <tr> <td>b</td> <td>b = Proton number</td> </tr> </table>										a	a = Relative atomic mass	X	X = Atomic symbol	b	b = Proton number							4.0 He 2
a	a = Relative atomic mass																							
X	X = Atomic symbol																							
b	b = Proton number																							
6.9 Li 3	9.0 Be 4											10.8 B 5	12.0 C 6	14.0 N 7	16.0 O 8	19.0 F 9	20.2 Ne 10							
23.0 Na 11	24.3 Mg 12											27.0 Al 13	28.1 Si 14	31.0 P 15	32.1 S 16	35.5 Cl 17	40.0 Ar 18							
39.1 K 19	40.1 Ca 20	45.0 Sc 21	47.9 Ti 22	50.9 V 23	52.0 Cr 24	54.9 Mn 25	55.8 Fe 26	58.9 Co 27	58.7 Ni 28	63.5 Cu 29	65.4 Zn 30	69.7 Ga 31	72.6 Ge 32	74.9 As 33	79.0 Se 34	79.9 Br 35	83.8 Kr 36							
85.5 Rb 37	87.6 Sr 38	88.9 Y 39	91.2 Zr 40	92.9 Nb 41	95.9 Mo 42	[98] Tc 43	101 Ru 44	103 Rh 45	106 Pd 46	108 Ag 47	112 Cd 48	115 In 49	119 Sn 50	122 Sb 51	128 Te 52	127 I 53	131 Xe 54							
133 Cs 55	137 Ba 56	La to Lu	178 Hf 72	181 Ta 73	184 W 74	186 Re 75	190 Os 76	192 Ir 77	195 Pt 78	197 Au 79	201 Hg 80	204 Tl 81	207 Pb 82	209 Bi 83	[209] Po 84	[210] At 85	[222] Rn 86							
[223] Fr 87	[226] Ra 88	Ac to Lr																						

139 La 57	140 Ce 58	141 Pr 59	144 Nd 60	[145] Pm 61	150 Sm 62	152 Eu 63	157 Gd 64	159 Tb 65	163 Dy 66	165 Ho 67	167 Er 68	169 Tm 69	173 Yb 70	175 Lu 71
[227] Ac 89	232 Th 90	231 Pa 91	238 U 92	[237] Np 93	[244] Pu 94	[243] Am 95	[247] Cm 96	[247] Bk 97	[251] Cf 98	[252] Es 99	[257] Fm 100	[258] Md 101	[259] No 102	[262] Lr 103

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