

Subject: CAPE Chemistry

UNIT 1: CHEMICAL PRINCIPLES AND APPLICATIONS

Grade: 12

Term: 1-3

Module topics: Fundamentals in chemistry, kinetics and equilibria, Chemistry of the elements

MODULE 1: FUNDAMENTALS IN CHEMISTRY (TERM 2)

GENERAL OBJECTIVES

On completion of this Module, students should:

1. Understand that theories in chemistry are subject to change.
2. Understand the theory of atoms as a useful construct that explains the structure and behaviour of matter, and the impact of nuclear chemistry on society.
3. Understand the development of the periodic table for the classification of elements.
4. Appreciate that the forces of attraction between particles influence the properties and behaviour of matter.
5. Understand the mole concept.
6. Understand redox reactions.
7. Understand the kinetic theory.
8. Understand concepts associated with energy changes.
9. Develop the ability to perform calculations involving energy changes.

Specific Objectives

1. Atomic Structure and the Periodic Table

Students should be able to:

- 1.1. Discuss the process of theoretical change with respect to Dalton's atomic theory.
- 1.2. Describe the structure of the atom.
- 1.3. Define the following terms: (a) mass number; (b) isotopes; and, (c) relative atomic and isotopic masses based on the $^{12}_6\text{C}$ scale.
- 1.4. Explain the phenomenon of radioactivity.

- 1.5. Cite the use of radioisotopes;
- 1.6. Calculate the relative atomic mass of an element, given isotopic masses and abundances;
- 1.7. Explain how data from emission spectra provide evidence for discrete energy levels within the atom; Bohr model, simple treatment of the emission spectrum of hydrogen; Lyman series, Balmer series; ΔE or $de = hv$.
- 1.8. Describe the atomic orbitals; Principal quantum numbers, s, p and d orbitals; relative energies of 4s and 3d orbitals.
- 1.9. Describe the shapes of the s and p orbitals;
- 1.10. Determine the electronic configurations of atoms and ions in terms of s, p and d orbitals; Consider elements from atomic numbers 1 to 30.
- 1.11. State the factors which influence the first ionisation energy of elements; Include atomic radii, nuclear charge, shielding.
- 1.12. Explain how ionisation energy data provide evidence for sub-shells; and, Use Period 3 as an example.
- 1.13. Derive the electronic configuration of an element from data on successive ionisation energies.

2. Forces of Attraction

Students should be able to:

- 2.1. State the various forces of attraction between particles; Ionic bonds, covalent bonds, hydrogen bonds, metallic bonds, Van der Waals forces. (Permanent permanent dipole; induced-induced dipole or temporary/instantaneous induced dipole).
- 2.2. State the relationship between forces of attraction and states of matter.
- 2.3. Relate physical properties of matter to differences in strength of forces of attraction; Variation in melting points, boiling points and solubilities.
- 2.4. Explain the formation of the following: (a) ionic bonds; (b) covalent bonds; and, (c) metallic bonds. Covalent bonds should be discussed in terms of orbital overlap which results in the formation of sigma (σ) and pi (π) bonds. Metallic bonding is to be treated as a lattice of positive ions surrounded by mobile electrons. Electronegativity and polarity of bonds should be included.
- 2.5. Describe co-ordinate (dative covalent) bonding; Use 'dot-cross' diagrams; refer to simple systems (for example, BF_3/NH_3).

2.6. Describe the origin of intermolecular forces; Refer to hydrogen bonding; Van der Waals forces, permanent dipole.

2.7. Predict the shapes of, and bond angles in simple molecules and ions; Application of the VSEPR theory to include the following systems: trigonal (for example, BF_3), linear (for example, BeCl_2), tetrahedral (for example, NH_4^+ , CH_4), pyramidal (for example, H_3O^+ , CH_3 , and NH_3), non-linear (for example, H_2O), octahedral (for example, SF_6). Construct molecular models and measure bond angles.

2.8. Explain the shapes and bond angles of simple organic compounds; Ethane, ethene and benzene; apply the concept of hybridisation and resonance. Include sp^2 and sp^3 hybridisation.

2.9. Predict the shapes and bond angles of molecules similar to ethane; and, Simple substituted derivatives, for example, dichloroethane.

2.10. Describe qualitatively the lattice structure of crystalline solids and their relation to physical properties. Simple molecular (for example, I_2), hydrogen bonded (for example, ice), giant molecular (for example, SiO_2), ionic (for example, NaCl), metallic (for example, Cu), giant atomic (for example, graphite and diamond) structures.

3. The Mole Concept

Students should be able to:

3.1. Apply Avogadro's law; Perform calculations involving molar volumes.

3.2. Define the mole.

3.3. Define the term 'molar mass'.

3.4. Write balanced molecular and ionic equations.

3.5. Perform calculations based on the mole concept.

3.6. Apply the mole concept to molecular and ionic equations.

3.7. Calculate empirical and molecular formulae; Combustion data; absolute masses or relative abundances of elements.

3.8. Perform titrimetric analyses; and, Conduct acid/base titrations and redox titrations. (dichromate (VI)), hydrogen peroxide, iodide thiosulfate, manganate (VII); mean (consecutive accurate values within 0.10 cm^3 of each other), significant figures.

3.9. Use results from titrimetric analyses to calculate: (a) mole ratios; (b) molar concentration; and, (c) mass concentration.

4. Redox Reactions

Students should be able to:

- 4.1. Explain redox reactions in terms of electron transfer and changes in oxidation state (number).
- 4.2. Construct relevant half equations for redox reactions; Redox equations should be constructed under both acidic and basic conditions.
- 4.3. Deduce balanced equations for redox reactions from relevant half equations.
- 4.4. Order elements in terms of oxidising or reducing ability. Perform simple displacement reactions to order elements in terms of oxidising or reducing ability; addition of zinc to copper (II) sulfate solution; addition of chlorine water to bromide or iodide solutions.

5. Kinetic Theory

Students should be able to:

- 5.1 state the basic assumptions of the kinetic theory with reference to an ideal gas.
- 5.2. Explain the differences between real and ideal gases; Qualitative treatment only – the conditions which are necessary for a gas to approach ideal behaviour, the limitations of ideality at very high pressures and very low temperatures. Include graphical representations.
- 5.3. State Boyle's law and Charles' law; Include graphical representations.
- 5.4. Perform calculations using: (a) Boyle's law; (b) Charles' law; and, (c) the ideal gas equation ($pV = nRT$); and, Calculations involving the use of Van der Waals equation of state are not required. Include calculations of relative molar mass.
- 5.5. Explain the following: (a) the liquid state; (b) melting; and, (c) vaporisation.

6. Energetics

Students should be able to:

- 6.1. State that chemical reactions take place through energy changes (usually in the form of heat) associated with the breaking and making of bonds; Note that bond making is an exothermic process, that is: ΔH - ve while bond breaking is an endothermic process, that is: ΔH + ve.
- 6.2. State that energy changes occur in chemical reactions associated with the making and breaking of bonds;
- 6.3. Explain the differences between exothermic and endothermic reactions using energy profile diagrams.
- 6.4. Explain the term 'bond energy'; Calculations involving bond energy data.
- 6.5. Explain how bond energy data may be used to show the relationship between strength of covalent bonds and reactivity of covalent molecules; Lack of reactivity of nitrogen. Consider factors which affect bond energy.

6.6. Apply concepts associated with enthalpy changes; Include enthalpy change of formation, combustion, neutralisation, reaction, hydration, solution, atomisation, ionisation energy, electron affinity and lattice energy.

6.7. Explain the effect of ionic charge and radius on the magnitude of lattice energy; No calculation needed.

6.8. State Hess's law of constant heat summation; and, Use standard conditions.

6.9. Calculate enthalpy changes from appropriate experimental data.

MODULE 2: Kinetics and Equilibria (TERM 3)

GENERAL OBJECTIVES

On completion of this Module, students should:

1. Understand the concepts associated with reaction rates;
2. Understand the concepts associated with chemical equilibrium;
3. Appreciate that equilibrium concepts can be applied to chemical systems;
4. Appreciate that principles of kinetics and equilibria can be applied to industrial and biological processes.

SPECIFIC OBJECTIVES

1. Rates of Reaction

Students should be able to:

- 1.1. Explain the concepts associated with reaction rates; Include a study of rate constant, order of reaction, half-life, rate-determining step, activation energy, collision theory, (simple treatment only), and catalysis. Include enzymes in industrial and biological processes.
- 1.2. Design suitable experiments for studying the factors which affect rates of reactions; Include effects of concentration, temperature and catalysts. Conduct suitable experiments for studying the factors which affect rates of reactions; express results in the form of tables and graphs.
- 1.3 Construct rate equations of the form: $\text{Rate} = k [\text{A}]^n [\text{B}]^m$ limited to simple cases involving zero, first and second order reactions.
- 1.4 Rate equations may be derived or deduced from experimental data supplied.

- 1.5 Deduce the order of reaction from appropriate data; Include deductions of possible reaction mechanisms.
- 1.6 Interpret concentration against time and concentration against rate for zero and first order reactions; Qualitative and quantitative treatments required.
- 1.7 Perform calculations from rate data; Calculate initial rates and rate constants.
- 1.8 Perform simple calculations using half-life data; and, Limited to first order reactions.
- 1.9 Explain the effect of temperature and catalysts on the rate of the reaction using Boltzmann distribution of energies (and of collision frequency). Include the use of Boltzmann distribution curves.

2. Principles of Chemical Equilibrium

Students should be able to:

- 2.1 explain the concept of dynamic equilibrium; Consider examples of static and dynamic equilibrium. Refer to physical and chemical processes.
- 2.2. State the characteristics of a system in dynamic equilibrium.
- 2.3. Define the terms K_c and K_p ; Write equilibrium constant expressions in terms of K_c and K_p .
Conduct a simple experiment to determine the value of K_c for a reaction.
- 2.4. Perform calculations involving equilibrium constants in terms of concentration, (K_c) and partial pressure, (K_p).
- 2.5. State Le Chatelier's principle.
- 2.6. Apply Le Chatelier's principle to explain the effects of changes in temperature, concentration and pressure on a system in equilibrium; and, Include reference to the characteristics of a system in dynamic equilibrium.
- 2.7. Interpret how changes in concentration, pressure, temperature or the presence of a catalyst may affect the value of the equilibrium constant.

3. Acid/Base Equilibria

Students should be able to:

3.1. Explain the differences in behaviour of strong and weak acids and bases, using Bronsted-Lowry theory.

3.2. Define the terms K_a , pH , pK_a , and pK_b , K_w and pK_w .

3.3. Perform calculations involving pH , pOH , K_a , pK_a K_w and pK_w , K_b and pK_b ; Quadratic equations are not required.

3.4. Describe the changes in pH during acid/base titrations; Include a study of titration curves.

3.5. Explain what is meant by the pH range of indicator.

3.6. State the basis for the selection of acid/base indicator for use in titrations. Include phenolphthalein and methyl orange. Titration curves. Perform experiments to show that the effectiveness of different indicators is related to the pH changes which occur during titration.

4. Buffers and pH

Students should be able to:

4.1. Define the term 'buffer solution'.

4.2. Explain how buffer solutions control pH .

4.3. Calculate the pH of buffer solutions from appropriate data; and, perform simple experiments to determine the pH of buffer solutions.

4.4. Discuss the importance of buffers in biological systems and in industrial processes. Include reference to blood buffer systems such as hydrogencarbonate, phosphate and amino- acid systems, enzyme catalysed reactions and the food processing industry.

5. Solubility Product

Students should be able to:

5.1. Define the term solubility product, K_{sp} . Write equilibrium constant expression for K_{sp} .

- 5.2. Explain the principles underlying solubility product and the common ion effect;
- 5.3. Perform calculations involving solubility product.
- 5.4. Relate the solubility product principle to the selective precipitation of substances. Include reference to qualitative analysis and kidney stone formation.

6. Redox Equilibria

Students should be able to:

- 6.1. Define the terms standard electrode potential and standard cell potential;
- 6.2. Describe the standard hydrogen electrode; include labelled diagram of standard hydrogen electrode.
- 6.3. Describe methods used to measure the standard electrode potentials of: (a) metals or nonmetals in contact with their ions in aqueous solutions; and, (b) ions of the same element in different oxidation states;
- 6.4. Calculate standard cell potentials from standard electrode potentials of two half cells;
- 6.5. Use standard electrode potentials of cells: (a) to determine the direction of electron flow; and, (b) to determine the feasibility of a reaction; include cell diagram or notation of the type $\text{zn(s)}|\text{zn}^{2+}(\text{aq})||\text{cu}^{2+}(\text{aq})|\text{cu(s)}$.
- 6.6. Predict how the value of an electrode potential varies with concentration; and, no treatment of the nernst equation is required. Apply le chatelier's principle.
- 6.7. Apply the principles of redox processes to energy storage devices.

MODULE 3: CHEMISTRY OF THE ELEMENTS (TERM 1)

GENERAL OBJECTIVES

On completion of this Module, students should:

1. Use fundamental concepts to rationalise the physical and chemical properties of elements and their compounds;
2. Appreciate that the properties of elements are related to their compounds and their uses; and,
3. Understand the principles underlying the identification of anions and cations

Specific Objectives

1. Period 3: Sodium to Argon

Students should be able to:

1.1. Explain the variations in physical properties of the elements in terms of structure and bonding; Include reference to melting point and electrical conductivity. Atomic and ionic radii, electronegativity and density.

1.2 describe the reactions of the elements with oxygen, chlorine and water; No treatment of peroxides or superoxides required.

1.3 explain the variation in oxidation number of the oxides and chlorides.

1.4 describe the reactions of the oxides and chlorides with water; Include equations. Conduct experiments to investigate the reactions of the oxides and chlorides with water; include relevant equations.

1.5 explain the trend in the acid/base behaviour of the oxides and hydroxides; Include equations. Conduct experiments to investigate the acid/base behavior of the oxides and hydroxides; include relevant equations.

1.6. Predict the types of chemical bonding present in the chlorides and oxides; and, Refer to differences in electronegativities and ionic radii of the elements.

1.7 discuss the uses of some of the compounds of aluminium and phosphorous. Limited to the use of aluminium hydroxide in antacid medication, white phosphorous used in flares and military applications, red phosphorous used at the side of match boxes and argon used in fluorescent and incandescent lighting.

2. Group II Elements

Students should be able to:

2.1. Explain the variations in properties of the elements in terms of structure and bonding; Include reference to atomic and ionic radii and ionisation energies.

2.2. Describe the reactions of the elements with oxygen, water, and dilute acids; Include equations.

2.3. Explain the variation in the solubility of the sulfates; Qualitative treatment only is required. Simple explanations in terms of lattice and hydration energies.

2.4. Explain the variation in the thermal decomposition of the carbonates and nitrates; and, Include equations.

2.5. Discuss the uses of some of the compounds of magnesium and calcium. Limited to the use of magnesium oxide, calcium oxide, calcium hydroxide and calcium carbonate.

3. Group IV Elements

Students should be able to:

- 3.1. Explain the variations in physical properties of the elements in terms of structure and bonding; Include reference to variations in metallic character and electrical conductivity.
- 3.2. Describe the bonding of the tetrachlorides;
- 3.3. Explain the reactions of the tetrachlorides with water; Include equations.
- 3.4. Discuss the trends in: 1.11 (a) bonding; (b) acid/base character; and, (c) thermal stability of the oxides of oxidation states II and IV; Make reference to E^θ values of the elements. Include equations.
- 3.5. Discuss the relative stabilities of the oxides and aqueous cations of the elements in their higher and lower oxidation states;

4. Group VII Elements

Students should be able to:

- 4.1 explain the variations in physical properties of the elements in terms of structure and bonding; Volatility, density, colour, and state.
- 4.2. Explain the relative reactivities of the elements as oxidising agents; Include reactions with sodium thiosulfate and refer to E^θ values. Use solutions of the elements with bleach, bromine water, and iodine solution.
- 4.3. Describe the reactions of the elements with hydrogen; Include equations.
- 4.4. Explain the relative thermal stabilities of the hydrides; Include bond energies in explanations.
- 4.5. Describe the reactions of the halide ions with: 1.14 (a) aqueous solution of AgNO_3 followed by aqueous ammonia; and, (b) concentrated sulfuric acid; and, Perform experiments of halide ions with aqueous AgNO_3 followed by aqueous ammonia.
- 4.6. Describe the reactions of chlorine with cold and hot aqueous solution of sodium hydroxide. Include changes in oxidation number and the process of disproportionation.

5. First Row Transition Elements

Students should be able to:

- 5.1 define the term transition element; D-block elements forming one or more stable ions with incomplete d-orbitals.
- 5.2. Describe the characteristics of transition elements; Include variation in oxidation number, complex formation, coloured compounds, catalytic activity, magnetic properties.

5.3. Discuss qualitatively the properties of transition elements when compared to those of calcium as a typical s-block element; 1.1 Melting point, density, atomic radius, ionic radius, first ionisation energy, and conductivity.

5.4. Determine the electronic configuration of the first row transition elements and of their ions; Mention changes in oxidation number.

5.5. Explain the relatively small changes in atomic radii, ionic radii, and ionisation energies of the elements across the period;

5.6. Explain the formation of coloured ions by transition elements; d-orbital separation of energy in octahedral complexes.

5.7. Describe the variation in oxidation states of vanadium; Refer to E^θ values. Perform experiments to include the use of an acidified solution of ammonium vanadate (V) and granulated zinc.

5.8. Predict the shapes of complexes of transition elements; Octahedral, tetrahedral and square planar.

5.9. Discuss the use of: $\text{Fe}^{3+}(\text{aq})/\text{Fe}^{2+}(\text{aq})$, $\text{MnO}_4^-(\text{aq})/\text{Mn}^{2+}(\text{aq})$, and $\text{Cr}_2\text{O}_7^{2-}(\text{aq})/\text{Cr}^{3+}(\text{aq})$ as redox systems.

5.10. Explain the principle of ligand exchange. Stability constants and the CO/O₂ haemoglobin and $\text{NH}_3(\text{aq})/\text{Cu}^{2+}(\text{aq})$ systems. Perform experiments to demonstrate ligand exchange. Include reactions involving $\text{Co}^{2+}(\text{aq})$, $\text{Cu}^{2+}(\text{aq})$.

6. Identification of Cations and Anions

Students should be able to:

6.1. Identify cations: K^+ , Na^+ , Ca^{2+} , Ba^{2+} , Cu^{2+} by their flame tests; Refer to atomic emission spectra.

6.2. Identify cations $\text{Mg}^{2+}(\text{aq})$, $\text{Al}^{3+}(\text{aq})$, $\text{Ca}^{2+}(\text{aq})$, $\text{Cr}^{3+}(\text{aq})$, $\text{Mn}^{2+}(\text{aq})$, $\text{Fe}^{2+}(\text{aq})$, $\text{Fe}^{3+}(\text{aq})$, $\text{Cu}^{2+}(\text{aq})$, $\text{Zn}^{2+}(\text{aq})$, $\text{Ba}^{2+}(\text{aq})$, $\text{Pb}^{2+}(\text{aq})$, $\text{NH}_4^+(\text{aq})$; Include the reactions with $\text{OH}^-(\text{aq})$, $\text{CO}_3^{2-}(\text{aq})$ and $\text{NH}_3(\text{aq})$ and confirmatory tests. Perform experiments of the identified cations with hydroxide and aqueous ammonia. Where possible perform confirmatory tests of the identified cations.

6.3. Explain the principles upon which the reactions in Specific Objective 6.2 are based; Refer to equilibrium concepts.

6.4. Write ionic equations for the reactions in Specific Objective 6.2; Include state symbols.

6.5. Identify anions: CO_3^{2-} , NO_3^- , SO_4^{2-} , $\text{SO}_3^{2-}(\text{aq})$, Cl^- , Br^- , I^- , CrO_4^{2-} ; and, Include the reactions with $\text{HCl}(\text{aq})$, conc H_2SO_4 , $\text{Pb}^{2+}(\text{aq})$, $\text{Ag}^+(\text{aq})$, followed by $\text{NH}_3(\text{aq})$, $\text{Ca}(\text{OH})_2(\text{aq})$, $\text{Ba}^{2+}(\text{aq})$, followed by dilute acid. For NO_3^- use copper turnings and conc. H_2SO_4 or add aluminium (powder) or zinc (powder) in the alkaline solution and confirmatory

tests for gases where applicable. Perform experiments to identify the anions CO_3^{2-} , NO_3^- , SO_4^{2-} , SO_3^{2-} (aq), Cl^- , Br^- , I^- , CrO_4^{2-} .

6.6. Write ionic equations for the reactions in Specific Objective 6.5.