

Chemistry Syllabus, Grade 12

General Objectives of Grade 12 Chemistry

To develop understanding and acquire knowledge of:

- types, properties and processes of formation of solutions;
- different concepts of acids and bases, their properties and qualitative and quantitative expressions;
- energy changes in physical and chemical processes and laws governing them;
- fundamental concepts related to the interconversion of chemical and electrical energy;
- qualitative and quantitative aspects of electrochemistry and its applications;
- occurrence, abundance and metallurgical processes of some useful elements;
- analyzing the risks and benefits of the development and application of synthetic polymers and suggesting possible methods of control.

To develop skills and abilities of:

- preparing solutions of specific concentration and solving quantitative problems involving solutions;
- solving equilibrium problems involving acids and bases;
- predicting whether a solution of a specific salt will be acidic, alkaline or neutral;
- calculating energy changes in chemical reactions;
- determining spontaneity of chemical reactions;
- calculating problems related to electrolysis;
- constructing model galvanic cell;
- designing and conducting simple experiments relevant to their level.

To develop habits and attitudes of:

- appreciating the roll of some chemical processes such as buffer systems in balancing natural phenomena in our lives;
- practicing the chemical and physical methods of preventing corrosion;
- realising that elements are cycling in nature;
- being confident and independent;
- relating a scientific knowledge to everyday applications.

Unit 1: Solutions (30 periods)

Unit out comes: Students should be able to:

- Know the types of solutions;
- Understand the solution formation process, the rate of solution , the heat of solution & solubility;
- Describe the dependence of solubility on temperature & pressure of solution;
- Know how to solve problems involving concentration of solutions & express the result in various units;
- Describe, using the concept of equilibrium, the behavior of ionic solutes in solutions that are unsaturated, saturated and supersaturated;
- Prepare solutions of required concentration by dissolving a solute or diluting a concentrated solution;
- Understand the relationship between the colligative properties of solution & know how to solve problems involving colligative properties of solution;
- Describe scientific enquiry skills along this unit: observing, classifying, comparing & contrasting, communicating, measuring, asking questions, drawing conclusion, applying concept and problem solving.

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| <p><i>Students will be able to:</i></p> <ul style="list-style-type: none"> • Define the terms mixture, homogenous and heterogeneous mixtures, solute, solvent, solution. • distinguish between homogenous and heterogeneous mixtures • describe suspension and colloids | <p>1. Solutions 1.1 Homogeneous and heterogeneous mixtures <i>(1 period)</i></p> <ul style="list-style-type: none"> • Some basic definitions - mixture, homogeneous and heterogeneous mixtures, solute, solvent, solution • Suspension and colloids | <p>Students should be aware by now that understanding and using proper terminologies from the very beginning helps master basic concepts in science.</p> <p>Student should know that mixtures are divided in to three categories: suspensions, colloids, and solutions. A suspension is a dispersion of fine solid particles in a liquid or gas, removable by filtration. It could also separate into its components under the influence of gravity. A mixture of sand and water is an example.</p> <p>Students should know that in a solution, the particles of the solute are of the size of individual atoms, molecules or ions. Because of their small size they mingle with the particles of the solvent and never separate, no matter how long the solution is allowed to stand. Between these two extremes are mixtures called colloids. In a colloid, the particle size is much larger than in a solution, but not so large that they separate on standing.</p> <p>Ask students to dissolve sand, Gentian Violet and bile in water in three different flasks. Set them aside and see if they become any clearer on standing for some time. You might need to decant them after 30 minutes. Do you see any change upon further standing? Pass light through each mixture and see which one scatters light without separating on standing. What do you call this type of mixture?</p> <p>Students should come to the conclusion that both suspension and colloid are examples of heterogeneous mixtures whereas solution is a homogeneous mixture.</p> <p>Students should know that a solution has at least two pure components: solute and solvent. There could be more than one solute component in a solution but only one of the components is usually taken as the solvent. A substance dissolved in another substance is called solute and the substance in which the solute or solutes dissolve is the solvent</p> |

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| <ul style="list-style-type: none"> define heat of solution, solvation energy and hydration energy. apply the concept of heat of solution to the solution of ammonium nitrate crystal. apply the concept of heat of solution to the solution of sodium hydroxide crystal; explain how heat of solution is influenced by the inter particle interaction forces. <ul style="list-style-type: none"> describe the distinctions among unsaturated, saturated and supersaturated solutions saturated solutions of sodium sulphate prepare supersaturated solution of sodiumthiosulphate <ul style="list-style-type: none"> explain the equilibrium nature of saturated solution | <ul style="list-style-type: none"> The rate of dissolution Energy changes in the solution process <ul style="list-style-type: none"> Heats of solution Heats of hydration of ionic solid in water <p>1.4 Solubility as an equilibrium process (4 periods)</p> <ul style="list-style-type: none"> Saturated, unsaturated and supersaturated solution | <p>The heat of solution is the energy absorbed or liberated when a solution is formed. For liquid solutions formed from a solvent A and solute B, an ideal solution results when the A-B attractions are the same as the A-A and B-B attractions. For an ideal solution, $\Delta H_{\text{soln}} = 0$. When the A-B attractions are greater than the A-A and B-B attractions, $\Delta H_{\text{soln}} < 0$ and the solution process is exothermic. When the A-B attractions are weaker, $\Delta H_{\text{soln}} > 0$ and the solution process is endothermic.</p> <p>For solutions of solids in liquids the lattice energy (the energy required to separate the solute particles from a crystal) and hydration energy (or solvation energy- the energy released when the solute particle is placed into the solvent cage) must be considered.</p> <p>Students have to appreciate because solutions are mixtures but not pure compounds they are not governed by the law of constant composition. Hence one can make solutions of different concentrations from the same solute and solvent. They can be unsaturated, saturated or even supersaturated.</p> <p>Students should be aware that a water-soluble substance dissolves in water until no more can dissolve at the given temperature. At the start the solution is unsaturated; it can dissolve more solute. Eventually the solution becomes saturated and can not dissolve any more solute. This behavior is also observed in solutions involving solvents other than water.</p> <ul style="list-style-type: none"> An unsaturated solution is a solution which can dissolve more solute at a given temperature. A saturated solution is a solution which cannot dissolve any more solute at a given temperature <p>Students should prepare unsaturated and saturated solution of sodium sulfate Students should also prepare supersaturated solution of sodiumthiosulphate. It can be prepared by adding sodiumthiosulphate to hot water and letting it to cool slowly. Then add a small seed crystal to cause crystallization to occur.</p> <p>Students should know that a saturated solution is an equilibrium system in which the rate of dissolution is equal to the rate of crystallization. The solute is in equilibrium with its ions or molecules in solution. At a given temperature the concentration of the solute and its ions or molecules in a saturated solution are constant.</p> |

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| <ul style="list-style-type: none"> • define solubility • describe the factors that affect solubility of substances • investigate the effect of temperature on solubility of sodium sulphate • conduct an experiment to determine solubility of table salt and sugar • state Henry's law • use Henry's law to calculate concentration of gaseous solute in a solution • define concentration of a solution • define mass percentage, ppm and ppb of a solute in a solution • calculate the mass percentage, ppm and ppb of a solute in a solution from a given information • define mole fraction | <ul style="list-style-type: none"> • Effect of temperature on solubility • Effect of pressure on solubility of gases - Henry's Law 1.5 Ways of expressing concentrations of solution (7 periods) • Mass percentage, ppm and ppb of solute • Mole fraction • Molarity • Normality • Molality • Conversion of concentration units | <p>Students should know that the solubility of a substance at a given temperature is the maximum quantity of substance, in moles or grams, which dissolves in a solvent to form 1 liter of saturated solution at that temperature. The solubility of substances changes with temperature.</p> <p>Ask students if they can draw parallel reasoning on factors that affect rates of dissolution with the overall factors that affect rates of chemical reactions. You may help them if they have forgotten what they learned in previous lessons.</p> <p>Among the many factors that determine the solubility, temperature should be given due consideration.</p> <p>A rise in temperature increases solubility if the dissolving of additional solute is endothermic. A fairly good rule of thumb is that the solubility of most solids and liquids in a solid solvent increases with increasing temperature. The solubilities of gases, however, almost always decrease with increasing temperature.</p> <p>Students should appreciate that the solubilities of almost all solid solutes increase with temperature whereas those of gases invariably decrease.</p> <p>Pressure has virtually no effect on the solubility of solids or liquids in liquid solvents. The solubilities of gases, however, are very markedly affected by pressure changes. Before you proceed ask students to use Le Chatelier's principle to predict this.</p> <p>This is summarized in the form of Henry's law. Henry's law relates the concentration of a dissolved gas, C_g, to its partial pressure, P_g, over the solution.</p> <p style="text-align: center;">$C_g = k_g P_g$, where k_g is Henry's law constant.</p> <p>Treat Henry's law only qualitatively.</p> <p>Students should be aware that the quantity of solute in a solution is a very important property of a solution. The quantity of solute per unit quantity of solution is called its concentration. Knowledge of it can be very useful in everyday life. For example, human beings estimate the quantity of solute in beverages, sugar, salt and other food solutions by tasting. Some farmers estimate the acid concentration in soils by tasting it to check whether the soil is suitable for planting crops. However, tasting is not an accurate and safe method of estimating the concentration of solute in solution. It doesn't allow the exact quantity of solute in solution to be known since different people have different tastes, apart from taste buds being unable to measure exact quantities. For example, many medicines in solution are prepared to specified concentrations of solute in solution to give correct dosage, otherwise they could be poisonous. Different forms of concentration arise due to the measurement of the quantities of solute and solvent in different units. Students should know the definitions of each concentration units.</p> |

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| <ul style="list-style-type: none"> • calculate mole fraction of a solute and a solvent in a solution • define molarity • calculate molarity of a solution from a given information • define the terms equivalent mass, number of equivalents and normality • calculate normality of a solution from a given information • define molality • calculate molality of a solution from a given information • Inter convert various concentration expressions • prepare molar solutions of different substances • prepare normal solutions of different substances • prepare molal solutions of different substances • explain dilution process • calculate the volume or concentration changes during dilution of solution • prepare a dilute solution from concentrated solution | <p>1.6 Preparation of solutions (2 periods)</p> <ul style="list-style-type: none"> • Diluting solutions | <p>It is easy to confuse some of the units with similar names so students have to be cautious when using them. Students should know that there are many ways in which the concentration of a solution can be described. Students could describe the concentration of a solution using appropriate units such as :</p> <ul style="list-style-type: none"> • Mass percentage, ppm and ppb of solute • Mole fraction • Molarity • Normality and • Molality <p>Let the Students calculate mass percentage, ppm, ppb, molarity, molality and normality of a solution from given information.</p> <p>Students should learn how to convert among the different units.</p> <p>Students should be aware that a common task in school, medical, industrial and other chemical laboratories is the preparation of solutions of known concentration. For aqueous solutions, distilled, demineralized or deionized water is used. Other solvents can also be used depending on the solution specified. Solutions are usually prepared from liquids or solids, and occasionally gases as solutes. Solutions in school science laboratories are prepared by teachers or laboratory technicians. Students should be able to calculate the quantities needed to prepare solutions. Students should know that a solution whose concentration is accurately known is called a standard solution. Students should prepare 1M, 1N and 1m solution of available chemical in a laboratory.</p> <p>Students should be aware that stock solutions are prepared very concentrated. When we are to use them they usually need to be weakened or diluted. Students have to use the following formula to dilute solutions.</p> |

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| <ul style="list-style-type: none"> • use stoichiometrically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substances • explain the relationship between reacting ions, spectator ions, precipitation and solubility • write net ionic equations • define colligative properties • list the important properties of solvents that are affected by the formation of a solution • state Raoult's law • explain the change in the colligative properties of solvent when a certain solute is added • describe Vant Hoff's factor • calculate vapor pressure, the boiling point, and freezing point of a solvent | <p>1.7 Solution Stoichiometry (3 periods)</p> <ul style="list-style-type: none"> • Mole - mass • Mole - volume • Mole - number of particles <p>1.8 Describing reactions in solution (1 period)</p> <ul style="list-style-type: none"> • Molecular equation • Ionic equation <p>1.9 Colligative properties of solutions (5 periods)</p> <ul style="list-style-type: none"> • Vapor pressure lowering • Boiling point elevation • Freezing point depression • Osmotic pressure | <p>$C_i V_i = C_f V_f$ where C_i is initial concentration, C_f is final concentration, V_i is initial volume and V_f is final volume. Ask students to prepare a 100ml, 1M sucrose solution and then take a certain volume of it to prepare 100ml, 0.1 M solution by dilution.</p> <p>Students should know that the principles of stoichiometry apply to reactants in solutions as well. Like in any other reaction, students should write balanced chemical equations that accurately describe the chemical reaction. Students should be able to appreciate that it is the particles dissolved in the solvent (assuming the solvent does not participate in the reaction) that take part in the reaction. If a solution is diluted by simply adding to it, the number of particle reactants remains the same. Students have to be aware of this when doing solution stoichiometry. Students could calculate amounts of reactants and products in a reaction of pure and dissolved substance.</p> <p>Students should know that if all the ions in an ionic equation for a reaction do not cancel, then there is a net ionic equation and a reaction will occur. However, if all the ions do cancel, then there is no net ionic equation and no net reaction. When writing ionic equations, formulas for weak electrolytes are written in molecular form. Keep the discussion qualitative so that students do not get bored by the arithmetic involved. You should of course be prepared to give additional work to able students.</p> <p>Students should appreciate that solutions have different physical properties from those of the pure solvents they are made of. Interestingly, these differences in properties depend not on the nature of the solute dissolved in the solvent but on the number of discrete particles per unit volume of the solution. Hence they are called colligative, because the word colligative means collective. The colligative properties of solutions are</p> <ol style="list-style-type: none"> 1. vapor pressure lowering 2. boiling point elevation 3. freezing point depression and 4. osmotic pressure <p>Even though most of these effects are small, they have many practical applications, including some that are vital to biological systems. Ask students if equal masses of different solutes in the same amount of solvent will show the same vapor pressure, same boiling point, same freezing point and same osmotic pressure. You may tell students, in the temperate latitudes where winters are unbearably called, antifreeze</p> |

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| after a certain amount of solute is added <ul style="list-style-type: none">• define Osmosis and Osmotic pressure of a solution• calculate the osmotic pressure of a solution• compare and contrast change in colligative properties of electrolytic and non electrolytic solution | | such as ethylene glycol are added to radiators of cars. Ask for explanation. People often talk about bodies of cars imported used being damaged. Provoke discussion with colligative properties of solutions in mind. Introduce the formulas for all four colligative properties of solutions and make sure students have at least qualitative understanding of them. Make especial note of the similarity in form between the ideal gas equation and the van't Hoff's equation for osmotic pressure. Give some examples to show how osmotic pressure is used in biochemical laboratories to determine molecular masses of huge molecules like proteins and nucleic acids. |

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define the terms mixture, homogenous and heterogeneous mixtures, solute, solvent, solution, rate of dissolution, heat of solution, solvation energy, hydration energy, solubility, concentration of a solution, mass percentage, ppm and ppb of a solute in a solution, mole fraction, molarity, molality, equivalent mass, number of equivalents and normality; distinguish between homogenous and heterogeneous mixtures; explain the different types of solutions and give examples for each type of the solutions; explain how the "like dissolves like" rule depends on inter-particle forces of interactions and predict relative solubilities; discuss the factors that affect the rate of dissolution; apply the concept of heat of solution to the solution of ammonium nitrate and sodium hydroxide crystals; explain how heat of solution is influenced by the inter particle interaction forces; describe the distinctions among unsaturated, saturated and supersaturated solutions; explain the equilibrium nature of saturated solution; describe the factors that affect solubility of substances; conduct an experiment to determine solubility of table salt and sugar; state Henry's law and use it to calculate concentration of gaseous solute in a solution; calculate the mass percentage, ppm and ppb of a solute in a solution, mole fraction of a solute and a solvent in a solution, molarity of a solution, normality of a solution, and

molality of a solution from a given information; prepare molar solutions, normal solutions and molal solutions of different substances; inter convert concentration units; explain dilution process and calculate the volume or concentration changes during dilution of solution; prepare a dilute solution from concentrated solution; use stoichiometrically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substances; explain the relationship between reacting ions, spectator ions, precipitation and solubility; write net ionic equations; define colligative properties, Osmosis and Osmotic pressure of a solution; list the important properties of solvents that are affected by the formation of a solution; state Raoult's law; explain the change in the colligative properties of solvent when a certain solute is added; calculate vapor pressure boiling point, and freezing point of a solvent after a certain amount of solute is added, and the osmotic pressure of a solution; compare and contrast change in colligative properties of electrolytic and non electrolytic solution.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 2: Acid-base Equilibria (26 periods)

Unit outcomes: Students will be able to:

- Understand the Arrhenius, Bronsted-Lowery & Lewis concepts of acids & bases;
- Understand the dissociation of water, weak monoprotic & polyprotic acids, & weak bases;
- Know how to solve equilibrium problems involving concentration of reactants & products, k_a , k_b , P^H & P^{OH} ;
- Understand the common ion effect, buffer solution, hydrolysis of salts , acid-base indicators and acid-base titrations;
- Explain how buffering action affects our daily lives using examples;
- Determine the equivalents of acid or base that are required to neutralize specific amount of acid or base;
- Predict, in qualitative terms, whether a solution of a specific salt will be acidic, basic or neutral;
- Know how to solve problems involving concentration and P^H of acid-base titration;
- Describe scientific enquiry skills along this unit: classifying, communicating, asking questions, applying concepts and making generalization.

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| <p><i>Students will be able to:</i></p> <ul style="list-style-type: none"> • define acid by the Arrhenius concept • explain why proton exists bounded to water molecule , as H_3O^+, in all acid-base reactions; • give examples of Arrhenius acids • define base by the Arrhenius concept • give examples of Arrhenius bases; • define acid by the Bronsted-Lowry concept; • give examples of Bronsted-Lowry acids • define base by the Bronsted-Lowry concept • give examples of Bronsted-Lowry bases | <p>2. Acid-base Equilibria 2.1 Acid-base concepts <i>(5 periods)</i></p> <ul style="list-style-type: none"> • Arrhenius concept of acids and bases • Bronsted-Lowry concept of acids and bases <ul style="list-style-type: none"> - Conjugate acid-base pairs - Auto ionization of substances - Amphiprotic species | <p>It has to be clear from the outset that this unit deals quantitatively with the equilibria involving the self ionization of water and the dissociation of weak (not strong) acids and bases. These are important in any aqueous system, particularly biological ones where many important molecules behave as weak acids or bases.</p> <ul style="list-style-type: none"> • Since many of the numerical problems in this unit require the application of some simple algebra, ample examples with detailed step-by-step procedures need to be given and students need to be warned not to rush and skip steps in the reasoning. • Definitions of acids and bases according to Arrhenius, Bronsted-Lowry and Lewis should be discussed in groups by the students. Students should be able to see that they are not in contradiction to one another and hence each one of them can still be used; but that one is more comprehensive than the others. They should be able to see this for themselves during the discussion. <p>Students should be aware that by Bronsted-Lowry definition, a great variety of chemical properties and chemical reactions can be correlated, including reactions that take place in solvents other than water or in solvent at all.</p> <p>Students should know some terms like conjugate acid-base pair, auto ionization and amphiprotic species under Bronsted-Lowry concepts.</p> <p>Students should be able to give examples of Arrhenius, Bronsted-Lowry and Lewis acids and identify conjugated acid – base pairs.</p> |

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| <ul style="list-style-type: none"> • explain what conjugate acids and conjugate bases are • identify the acid-base conjugate pairs from the given reaction; • write an equation for self-ionization of water and ammonia • explain what is meant by amphiprotic species • give examples of reactions of amphiprotic species; • define acid by the Lewis concept • give examples of Lewis acids • define base by the Lewis concept • give examples of Lewis bases • describe the ionization of water <ul style="list-style-type: none"> • drive the expression of ion product for water, K_w • explain the effect of temperature on K_w • explain why water is a weak electrolyte | <ul style="list-style-type: none"> • Lewis concept of acids and bases <p>2.2 Ionic Equilibria of weak acids and bases (9 periods)</p> <ul style="list-style-type: none"> • Ionization of water <ul style="list-style-type: none"> - Ion product for water, K_w | <p>Students should know that the most general theory of acids and bases is that of Lewis. This theory encompasses the Bronsted-Lowry theory, and it goes one step further by ascribing acid-base behavior to many reactions that do not involve the transfer of a proton.</p> <p>Students should appreciate one direction in which the science of chemistry matured, as narrow definitions of acids and bases progressively widened to encompass different species, physical states, solvent systems and reaction types.</p> <p>Students should know that equilibria involving the ionization of weak acids and weak bases in aqueous solution are very important in many fields of science.</p> <ul style="list-style-type: none"> • Students have to know that in any solution in which water is the solvent, the equilibrium condition for the ionization of water $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ holds at $25^\circ C$. <p>They should also be able to calculate pH/pOH given $[H^+]$ or $[OH^-]$ and $[H^+]$ or $[OH^-]$ given pH/pOH .</p> <p>Students should understand that the strengths of acids and bases can be described in many ways. Some of the ways are:</p> <ul style="list-style-type: none"> • H^+ and OH^- ions concentration • P^H and P^{OH} • Percent dissociation • K_a and K_b |

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| <ul style="list-style-type: none"> • use K_w to calculate $[H_3O^+]$ or $[OH^-]$ in aqueous solution • define P^H • define P^{OH} • explain the relationship between P^H and P^{OH} • calculate P^H from $[H^+]$ and $[H^+]$ from P^H • calculate P^{OH} from $[OH^-]$ and $[OH^-]$ from P^{OH} • write an expression for the percent ionization of weak acids or weak bases • calculate the percent dissociation of weak acids and bases • write the expression for the acid dissociation constant, K_a; • calculate K_a for an acid from the concentration of a given solution and its P^H; • Calculate $[H^+]$ and P^H of an acidic solution from given values of K_a and initial concentration of the solution. • Write the expression for the base dissociation constant, K_b • Calculate K_b for a base from the concentration of the basic solution and its P^{OH}. • Calculate the $[OH^-]$ and P^{OH} of a basic solution | <ul style="list-style-type: none"> • Measures of the strength of acids and bases in aqueous solution; <ul style="list-style-type: none"> - H^+ ion concentration, $[H^+]$, and OH^- ion concentration, $[OH^-]$ - P^H and P^{OH} - percent ionization - Ionization (dissociation) constant <ul style="list-style-type: none"> * Acid ionization constant, K_a * Base ionization constant, K_b | <p>Here K_a and K_b are treated more.</p> <p>Students should know that distilled or deionized water is an extremely weak electrolyte. It conducts electricity very weakly and any conduction by it can only be detected by very sensitive galvanometers. Extremely low concentrations of hydrogen, H^+ and hydroxide, OH^- are in equilibrium with the undissociated water molecules, H_2O.</p> <p>Students could be asked to derive ion product for water, K_w,</p> $K_w = [H^+][OH^-]$ <p>Students should be aware that in any aqueous system, pure or not, the ionic product is always constant at a constant temperature. It doesn't matter where the hydrogen (H^+) and hydroxide (OH^-) ions come from.</p> <p>Students should appreciate that an aqueous solution can be acidic, neutral or basic. A solution is described as neutral when the hydrogen ion concentration is equal to the hydroxide ion concentration, $[H^+] = [OH^-]$. If the concentration of hydrogen ions is above 1×10^{-7} mol/l the solution is said to be acidic. If the concentration of hydrogen ions is below 1×10^{-7} mol/l (i.e. the concentration of hydroxide ions is above 1×10^{-7} mol/l) the solution is said to be alkaline or basic.</p> <p>Students should be aware that P^H scale was introduced by Danish biochemist S. P. Sorensen in 1909. Chemists use the P^H scale to measure acidity or alkalinity of solutions.</p> $P^H = -\log [H^+]$ $P^{OH} = -\log [OH^-]$ <p>Students should know that the magnitude of acid ionization constant, K_a and the base acid ionization constant, K_b are measures of the acid strength of an acid and the base strength of a base respectively.</p> <ul style="list-style-type: none"> • An acid with a higher value of K_a is a stronger acid than the acid with a lower value of K_a • An acid with a higher value of K_a dissociates more than the acid with a lower value of K_a • A base with a higher value of K_b is a stronger base than the base with a lower value of K_b • A base with a higher value of K_b dissociates more than the base with a lower value of K_b <p>Students could exercise to calculate K_a and K_b of an acid and a base from given solutions and their P^H and P^{OH} respectively.</p> |

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| <p>from a given values of K_b and initial concentration of the solution.</p> <ul style="list-style-type: none"> • define common ion effect; • explain the importance of common ion effect; • define a buffer solution; • give some common examples buffer systems; • explain the action of buffer solution and its importance in chemical processes • calculate the P^H of a given buffer solution • demonstrate the buffer action of CH_3COOH/CH_3COONa; • define hydrolysis; | <p>2.3 Common ion effect and Buffer Solution (4 periods)</p> <ul style="list-style-type: none"> • Common ion effect • Buffer solutions <p>2.4 Hydrolysis of Salts (2 periods)</p> | <p>Common ion effect could be given to the students as a group work to see if they have understood the Le Chatelier's principle discussed in grade 11 Chemistry class. Effort has to be made to see that the students understand this topic because it is a test case for many related topics.</p> <p>Students should know that buffer solutions have wide applications especially in biological systems. However since the concept is a formidable one, the discussion has to focus on applications rather than numerical problems.</p> <p>A buffer is a mixture of a weak acid and a weak base. It can be prepared by mixing a weak acid with one of its salts, or a weak base with one of its salts. A solution resists pH change in the presence of a buffer when a small amount of an acid or base is added, which otherwise would suffer pH change. Give your students how this is useful in biological systems.</p> <p>One of the first things your students learn in this section is that there is an inverse relationship between the strength of a Brønsted acid and the strength of its conjugate base. Strong acids have very weak conjugate bases; strong bases have very weak conjugate acids. Students should demonstrate the buffer action of CH_3COOH/CH_3COONa.</p> <p>Students should know that solutions of salts contain two ions. The cation of the salt can potentially serve as an acid; the anion as a base. Whether a solution of a salt is acidic, basic, or neutral depends on the strengths of these acids and bases.</p> <p>In determining what happens in solution of a salt, students have to examine the salt to determine what its 'parents' are. For example, sodium chloride, NaCl has as its parents the acid HCl and the base NaOH. Each of these are strong, so we categorize NaCl as the salt of a strong acid and a strong base. On the other hand, sodium nitrite, $NaNO_2$, is the salt of a weak acid, HNO_2 and a strong base, NaOH.</p> <p>Students could discuss hydrolysis of each type of salt. They may raise the following points.</p> <p>Salts of strong acids and strong bases. Solutions of these salts are neutral because the ions are very weak acids and bases. Hence these ions don't hydrolyze. Examples are NaCl and KNO_3. Ask students to give more examples.</p> |

| Competencies | Contents | Suggested Activities |
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| <ul style="list-style-type: none"> • explain why a salt of weak acid and strong base gives a basic solution • explain why a salt of strong acid and weak base gives an acidic solution • explain why salts of weak acids and weak bases give acidic, basic or neutral solution • define acid-base indicators • write some examples of acid-base indicators • suggest suitable indicator for a given acid-base titration | <ul style="list-style-type: none"> • Salts of weak acids and strong bases: anion hydrolysis • Salts of strong acids and weak bases: cation hydrolysis • Salts of weak acids and weak bases: cation and anion hydrolysis 2.5 Acid-base indicators and Titrations (6 periods) • Acid-base indicators | <p>Salts of weak acids and strong bases. Solutions of these salts are basic because the anion of the weak acid is a moderately strong base. Examples are $\text{NaC}_2\text{H}_3\text{O}_2$ and NaNO_2. In general, the anion of the salt reacts as follows: $\text{X}^- + \text{H}_2\text{O} \leftrightarrow \text{HX} + \text{OH}^-$. Since a component of the salt is reacting with water, the phenomenon has come to be known as hydrolysis.</p> <p>Salts of strong acids and weak bases. Solutions of these salts are acidic because the cation of the weak base is a moderately strong acid. Examples are NH_4Cl and $\text{N}_2\text{H}_5\text{Cl}$.</p> <p>Salts of weak acids and weak bases. In these solutions there is both a strong acid (the cation of the weak base) and a strong base (the anion of the weak acid). Whether the solution of such a salt is acidic, basic, or neutral depends on the relative strengths of the acidic cation and basic anion. If the acid is stronger than the base, the solution is acidic, and vice versa. If they are of equal strengths, the solution is neutral. This topic should be treated lightly and only qualitatively.</p> <p>Students should understand that acid-base indicators are organic molecules whose colors depend on the acidity or basicity of the solution in which they are dissolved. This needs to be discussed from equilibrium point of view. pH meters can also be used to measure $[\text{H}^+]$ by means of two electrodes immersed in a test solution. This should not be discussed in detail but showing how to use the apparatus would be beneficial</p> <p>The addition of a common ion to a solution containing a salt in equilibrium with its ions decreases the solubility of the salt. This is in accordance to Le Chatelier's principle applicable to systems at equilibrium.</p> <p>Acid-base indicators are organic molecules that indicate whether a solution is acidic, basic or neutral.</p> <p>An acid-base indicator is a weak acid or weak base with its corresponding conjugate pair. The molecular form of the indicator, HIn, has one color and the ionic form In^-, has a different color. The color that is observed in a solution of the indicator is controlled by the ratio of $[\text{HIn}]$ to $[\text{In}^-]$, which is determined in turn by the H^+ concentration in the solution.</p> <p>An indicator can be used to signal the completion of a reaction during a titration by changing color.</p> <p>Since students find it difficult to understand how the equivalent mass is different from the molecular mass of a substance analogies could be used.</p> |

| Competencies | Contents | Suggested Activities |
|---|---|---|
| <ul style="list-style-type: none"> • explain the equivalents of acids and bases • calculate the normality of a given acidic or basic solution • define acid-base titration • define end point • define equivalence point • distinguish between end point and equivalent point; • discuss the different types of titration curves | <ul style="list-style-type: none"> • Equivalents of acids and bases <ul style="list-style-type: none"> - Number of equivalents - Normality • Acid-base titrations • The equivalent point and the end point • Acid-base titration curves <ul style="list-style-type: none"> - Titration of a strong acid with a strong base - Titration using a weak acid and a strong base - Titration of a weak base with a strong acid | <p>An equivalent of an acid is the mass of the acid that releases one mole of H^+ in solution. Likewise an equivalent of a base is the mass of the base that produces one mole of OH^- or accepts one mole of H^+. Students should be able to make similar definitions for reductants and oxidants in redox reactions. Students should appreciate how the definitions relate to the various definitions of acids and bases.</p> <p>Students should be given numerous examples and problems to acquaint themselves with the difference between equivalent mass and molecular mass of a substance.</p> <p>Ask students to define molarity and normality and see if they know that the difference lies in noticing the differences between equivalent mass and molecular mass.</p> <p>Students could exercise to calculate the normality of a given acidic or basic solution.</p> <ul style="list-style-type: none"> • Students should not confuse between titration, a method of determining how much acid or base is present in a solution and pH, which is a measure of acidity or basicity of a solution. It might be helpful to give examples like the following one. A 0.1 M solutions of HCl in water and a 0.1 M solution acetic acid have pH's 1 and 2.9, respectively, that is they differ in acidity. On the other hand the total concentration of acid is the same and each of these solutions neutralizes the same amount of NaOH solution. • Students usually think at the equivalence point of a titration the pH is always 7. They need to see that this is true only if the titration is between a strong acid and strong base. They will have to relate this to hydrolysis of anions or cations. <p>Acids are titrated to the point of neutralization by adding equivalent amount of a base. An indicator is employed to determine the end point. At the endpoint or equivalence point, equal numbers of equivalents of acids and bases will have been added.</p> <p>Titration is an analytical procedure in which a solution, generally of known concentration, is added gradually from a burette to another solution where the solutes react. This is continued until the completion of the reaction is signaled by an indicator which was added at the beginning of the titration.</p> <p>Acid base titration curves are drawn by plotting points corresponding to the pH of the solution and equivalents of the titrant added.</p> <p>Students could think that at the equivalence point the pH of the solution is 7. This may not be true always. This is true if hydrolysis of the salt formed from the acid base reaction is nonexistent.</p> <p>Salts of strong acid and strong base are neutral.</p> <p>Salts of weak acids and strong base are basic and salts of strong acids and weak base are acidic. A way to remember this is to consider hybrids of characteristics of different strengths: the stronger one dominates. What is more important is for the students to be able to explain this from hydrolysis of salts point of view.</p> |

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define acids and bases by the Arrhenius, Bronsted-Lowry and Lewis concepts, and give examples for each of them; explain why proton exists bounded to water molecule, as H_3O^+ , in all acid-base reactions; explain what conjugate acids and conjugate bases are and identify the acid-base conjugate pairs from the given reaction; write an equation for self-ionization of water and ammonia; explain what is meant by amphiprotic species and give examples of their reactions; describe the ionization of water and explain why water is a weak electrolyte; derive the expression of ion product for water, K_w , and use it to calculate $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$ in aqueous solution; define P^{H} and P^{OH} , and explain the relationship between them; calculate P^{H} from $[\text{H}^+]$ and $[\text{H}^+]$ from P^{H} ; calculate P^{OH} from $[\text{OH}^-]$ and $[\text{OH}^-]$ from P^{OH} ; write an expression for the percent ionization of weak acids or weak bases and calculate them; write the expression for the acid dissociation constant, K_a , and the base dissociation constant, K_b ; calculate K_a for an acid from the concentration of a given solution and its P^{H} , and K_b for a base from the concentration of the basic solution and its P^{OH} ; Calculate the $[\text{H}^+]$ and P^{H} of an acidic solution from given values of K_a and initial concentration of the solution, and the $[\text{OH}^-]$ and P^{OH} of a basic solution from a given values of

K_b and initial concentration of the solution; define common ion effect and explain the importance of common ion effect; define buffer solution, give some common examples buffer systems, and explain the action of buffer solution and its importance in chemical processes; calculate the P^{H} of a given buffer solution; prepare a simple buffer solution and test its buffering capacity; define hydrolysis; explain why a salt of weak acid and strong base gives a basic solution, a salt of strong acid and weak base gives an acidic solution, and a salt of weak acid and weak base gives acidic, basic or neutral solution; define acid-base indicators and write some examples of them; suggest suitable indicator for a given acid-base titration; explain the equivalents of acids and bases; calculate the normality of a given acidic or basic solution; define acid-base titration, end point and equivalence point; distinguish between end point and equivalent point; discuss the different types of titration curves;

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 3: Introduction to Chemical Thermodynamics (12 periods)

Unit outcomes: Students will be able to:

- Understand the terms systems (open, closed & isolated), spontaneous process, functions(state & path), properties (intensive & extensive), internal energy, heat & work;
- Understand the first and second laws of thermodynamics;
- Compare the energy changes observed when chemical bonds are formed & broken, & relate these changes to endothermic & exothermic reactions;
- Identify ways in which the terms reactants, products & the heat are combined to form thermo chemical equations representing endothermic & exothermic chemical changes;
- Understand concepts such as enthalpy change, entropy & entropy changes & free energy change;
- Determine the spontaneity of a given reaction;
- Describe scientific enquiry skills along this unit: observing. Communicating, measuring, applying concept and asking questions.

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|---|--|---|
| <p><i>Students will be able to:</i></p> <ul style="list-style-type: none"> • define chemical thermodynamics; • define systems; • give examples of systems; • classify system into open, closed and isolated systems' • explain open closed and isolated system • define state function and path function • distinguish between intensive and extensive properties, • define spontaneous and | <p>3. Introduction to Chemical Thermodynamics</p> <p>3.1 Common thermodynamic terms (2 periods)</p> <ul style="list-style-type: none"> • Systems, Surroundings, Functions, Properties and Processes | <p>Students should know two factors that control whether the products of a chemical reaction will form are kinetics and thermodynamics. Thermodynamics controls the feasibility of a reaction in the sense that it determines whether a reaction is possible and how much products can be formed. It deals with energy changes and is applied to physical as well as chemical changes. Students find it easier to comprehend the principles of chemical thermodynamics if the principles are developed using physical systems followed by the extension of the principles to chemical systems.</p> <ul style="list-style-type: none"> • Common thermodynamic terms need to be elaborated with examples and question & answer methodology. Students should not move on before mastering these essential terms. Ask students to read and understand the meanings of the following terms and write the definitions in their own words. It would be good if they try to explain them verbally to their class mates. <p>Thermodynamics System Surroundings Isothermal Adiabatic State Equation of state Heat capacity Specific heat State function State variable</p> |

| Competencies | Contents | Suggested Activities |
|--|--|---|
| <p>non-spontaneous process</p> <ul style="list-style-type: none"> • explain internal energy in relation to the concepts of thermodynamics; • explain heat in relation to the concepts of thermodynamics, ▪ explain work in relation to the concepts of thermodynamics, <ul style="list-style-type: none"> • state first law of thermodynamics, • explain first law of thermodynamics, • calculate change in internal energy of a system from a given information, <ul style="list-style-type: none"> • define enthalpy change, ΔH • derive an expression for the enthalpy change of an ideal gas based on first law of thermodynamics, • calculate enthalpy change for ideal gases from given information, | <p>3.2 First Law of thermodynamics and some thermodynamic quantities (3 periods)</p> <ul style="list-style-type: none"> • Internal energy (E) • Heat (q) • Work (W) • First Law of Thermodynamics <p>3.3 Thermo chemistry (4 periods)</p> <ul style="list-style-type: none"> • Heats of reaction (Enthalpy changes) | <ul style="list-style-type: none"> • First law of thermodynamics has to be kept to its bare minimum; however, the various forms of its definition need to be discussed in groups for the purpose of clarity. During a chemical change the energy gained by the system, (the reaction mixture inside the vessel) is equal to the energy lost by the surroundings to the system and vice versa. In other words energy cannot simply appear or disappear- it is neither created nor destroyed. The total energy of the universe is constant. These are different forms of the first law of thermodynamics. <p>Ask students if a glass of water at room temperature will freeze. If it does freeze, it has lost energy and the surrounding will gain equal amount of energy. This does not happen not because the first law is not obeyed. They should consider the enthalpy change and see if it makes sense from what they see. How about an ice cube at room temperature melting? Is its enthalpy increasing or decreasing? Encourage them to argue among themselves by citing similar examples. Are they convinced that decrease in enthalpy of a system by itself alone is not enough to be taken as an indicator of spontaneity? When a system undergoes a series of changes that ultimately brings it back to its original state, the net energy change for the system is zero. This is a formal statement of the law of conservation of energy, and is the basis of Hess's law. Often stated mathematically as $\Delta E = q-w$, where ΔE is the change in the internal energy, q is the heat added to the system, and w is the work done by the system. Internal energy, E is the total kinetic and potential energy of a system.</p> |

| Competencies | Contents | Suggested Activities |
|--|---|--|
| <ul style="list-style-type: none"> calculate the bond energies of substances in a given chemical reaction' calculate the standard enthalpy changes of a reaction from given enthalpy changes of reactants and products; explain entropy and entropy change; calculate the entropy change, from the given standard entropies of substances, state the second law of thermodynamics; explain the second law of thermodynamics; calculate the entropy changes from the given enthalpy change of the system and absolute temperature; explain free energy and free energy change; calculate the standard free energy change, ΔG^0, from given standard free energies | <ul style="list-style-type: none"> Bond Energies <p>3.4 Entropy and Second Law of Thermodynamics (3 periods)</p> <ul style="list-style-type: none"> Entropy(s) and Spontaneous process Second law of Thermodynamics <ul style="list-style-type: none"> - Free energy - Criterion for | <p>Students should know that thermodynamic data can be used to obtain information about strengths of bonds. Likewise bond energies can be utilized to obtain estimate of the heat of formation of a compound. Students should be able to do numerical problems that will help them understand the concept better.</p> <p>Students should appreciate that the second law of thermodynamics is the most applied tool in chemical thermodynamics; it is the criterion of spontaneity. Since spontaneity is easily confused with speed by students, enough time should be allotted to explain what it exactly means. The $\Delta G = \Delta H - T\Delta S$ criterion has to be explored using a table in group discussions. This has to be limited, though, to qualitative discussions. A spontaneous change is one that occurs without continued outside help. It may need a push to get it started (the mixture of H_2 and O_2, for instance) but once started, it continues on its own. Students should understand that there are two thermodynamic quantities that determine whether events are spontaneous. Spontaneity is favored if there is an energy decrease (if ΔH is negative, corresponding to an exothermic change). It is also favored if there is an increase in the degree of randomness of a system. The thermodynamic quantity related to randomness is entropy, S. Spontaneity is favored by an entropy increase (if ΔS is positive). If both changes in enthalpy and entropy are favorable, i.e ΔH is negative and ΔS is positive then definitely the process is favored. However, if ΔH is positive and ΔS is negative the process is not favored. Students have to be aware that chemical reactions that are not favored can take place if energy from outside is continually supplied. A corollary to this is that reactions that are thermodynamically allowed may not take place until the initial push is provided. The initial push could be small or big which was discussed in Grade 11, under the topic chemical kinetics.</p> <p>Students should know that the free energy change, ΔG is equal to the maximum work that can be obtained from a spontaneous change. It is also an indicator for the spontaneity of a process; if ΔH is < 0, spontaneous if ΔH is > 0, non-spontaneous and</p> |

| Competencies | Contents | Suggested Activities |
|---|--|---------------------------------------|
| of reactants and products <ul style="list-style-type: none"> describe the relationship between standard free energy change, ΔG°, standard enthalpy change, ΔH°, and standard entropy change, ΔS°, of a reaction, determine the spontaneity of a given reaction. | spontaneous process: ΔS , ΔG , ΔH | if ΔH is = 0, at equilibrium. |

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define chemical thermodynamics; define systems and give examples of systems; classify systems into open, closed and isolated systems and explain them; define state function and path function; distinguish between intensive and extensive properties; define spontaneous and non-spontaneous processes; explain internal energy, in relation to the concepts of thermodynamics; explain heat, internal energy and work in relation to the concepts of thermodynamics; state and explain first law of thermodynamics; calculate change in internal energy of a system from a given information; define enthalpy change, ΔH , derive an expression for the enthalpy change of an ideal gas based on first law of thermodynamics and calculate enthalpy change for ideal gases from given information; define standard state, standard molar enthalpy of combustion, $\Delta H^\circ_{\text{comb}}$, standard molar enthalpy of formation, $\Delta H^\circ_{\text{form}}$, and standard molar enthalpy of neutralization, $\Delta H^\circ_{\text{neutralization}}$; describe how heats of combustion can be used to estimate the energy available from foods; carry out an activity to measure standard molar enthalpy of neutralization; state Hess's law and apply it to solve problems on enthalpy changes of chemical reactions;

explain bond energy and calculate the bond energies of substances in a given chemical reaction; calculate the standard enthalpy changes of a reaction from given enthalpy changes of reactants and products; explain entropy and entropy change, and calculate the entropy change from the given standard entropies of substances; state and explain the second law of thermodynamics; calculate the entropy changes from the given enthalpy change of the system and absolute temperature; explain free energy and free energy change; calculate the standard free energy change, ΔG° , from given standard free energies of reactants and products; describe the relationship between standard free energy change, ΔG° , standard enthalpy change, ΔH° , and standard entropy change, ΔS° , of a reaction; determine the spontaneity of a given reaction.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 4: Electrochemistry (21 periods)

Unit outcomes: Students will be able to:

- Understand the fundamental concepts related to oxidation- reduction reaction;
- Know the application of redox reactions in production of new substances and energy;
- Demonstrate an understanding of fundamental concepts related to the interconversion of chemical & electrical energy;
- Understand the difference between metallic conduction & electrolytic conduction;
- Identify & describe the functioning of the components of electrolytic & galvanic cells;
- Understand the difference between electrolytic & galvanic cells;
- Know how to solve problems based on Faraday’s law;
- Measure through experimentation the mass of metal deposited by electroplating (e.g.:- copper from copper(II) sulphate), & apply Faraday’s law to relate the mass of metal deposited to the amount of charge passed;
- Predict the spontaneity of redox reactions & overall cell potentials by studying a table of half-cell reduction potentials;
- Determine the emf of an electrochemical cell experimentally or from given data;
- Explain the application of electrochemistry in our daily lives & in industry;
- Explain corrosion as an electrochemical process, & describe corrosion- inhibiting techniques (e.g. painting, galvanizing, cathodic protection);
- Describe examples of common galvanic cells & evaluate their environmental & social impact;
- Describe scientific enquiry skills along this unit: observing, classifying, comparing & contrasting, communicating, asking question, measuring, relating cause & effect and problem solving.

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|---|---|---|
| <p><i>Students will be able to:</i></p> <ul style="list-style-type: none"> • define redox reaction • define oxidation in terms of electron transfer and change in oxidation number. • define reduction in terms of electron transfer and change in oxidation number. • describe the oxidizing and reducing agents • identify the species that are oxidized and reduced in a given redox reaction and determine the oxidizing and reducing agents | <p>1. Electrochemistry 4.1. Reduction-oxidation reactions (5 periods)</p> <ul style="list-style-type: none"> • Oxidation • Reduction | <p>Since Redox Reactions play important role in all areas of chemistry especially Biological sciences, students should be encouraged to think of chemical reactions of relevance to plant and animal physiology.</p> <p>Remind students about rules to assign oxidation numbers. Let them understand the terminologies like redox reaction, oxidation ,reduction half-reaction, oxidizing agent and reducing agent.</p> <p>Ask students to identify the oxidized and reduced species and, the oxidizing and reducing agents. To help students the concept that an oxidizing agent it self reduced, you can discuss the action of a bleaching agent on clothing in a washing machine. A bleaching agent, which causes a whitening to occur, does not itself become whiter. An oxidizing agent is not oxidized.</p> |

| Competencies | Contents | Suggested Activities |
|--|---|--|
| <ul style="list-style-type: none"> • balance a given redox reaction using oxidation number method • balance a given redox reaction using ion-electron method • Recall metallic conductivity, electrolytic conductivity and electrolysis of molten electrolytes. • draw labeled diagrams of an electrolytic cell; • define preferential discharge • explain factors that affect preferential discharge. • describe the effect of nature of the ions on the electrolysis of aqueous solutions of dil. H_2SO_4 and NaOH. • describe the effect of concentration of the ions on the electrolysis of dilute and concentrated solutions of NaCl | <ul style="list-style-type: none"> • Balancing redox reactions <ul style="list-style-type: none"> - Change in Oxidation number method - Ion-electron method 4.2 Electrolysis of Aqueous Solutions (6 periods) <ul style="list-style-type: none"> • Electrolytic cell • Preferential discharge • Electrolysis of some selected aqueous solutions | <p>Students should be guided step by step on how to use the Change in Oxidation Number method and Ion-Electron method to balance equations where the simpler methods do not help. It is important to point out to students that oxidation numbers used in balancing redox equations are arbitrary values obtained by using arbitrary rules. Students should be warned against thinking that the oxidation number represents the real ion charge in the molecule.</p> <p>Students could discuss the metallic and electrolytic conductivities and electrolysis of molten electrolytes.</p> <p>All chemical changes, including electrochemical ones, involve energy changes. The main focus of this unit is to show how redox reactions can be caused to occur by the action of electricity, and the way electricity can be obtained from redox reactions that occur spontaneously.</p> <p>Students should appreciate the processes that take place at the cathode and anode in an electrolytic cell. The terminologies-cathode, cation, anode and anion are easily confused by students. Students should discuss in groups until they have them right.</p> <p>Students should learn from examples to be given how the net reaction in the electrolysis of aqueous solutions is controlled by which redox reactions occur most easily.</p> <p>Students should be asked to draw and label electrolytic cells. They should also show at which electrodes reduction and oxidation reactions take place.</p> <p>Students should know that the electrode can discharge only one type of species in preference to the other at a time</p> <p>Students should be able to confirm that electrons lost during oxidation half reaction are equal to electrons gained during reduction. The net reaction in a cell is the sum of cathode (reduction) and anode (oxidation) half reactions. Here the connection to the ion-electron method of balancing equations.</p> <p>Students should be assisted to conduct experiments on electrolysis of aqueous solutions and confirm the effects of different factors on preferential discharge of ions. They should write the anode, cathode and overall reactions name the product for each types of activities.</p> |

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
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| <ul style="list-style-type: none"> • describe the effects of types of electrodes on the electrolysis of CuSO_4 solution using Ni electrodes and Cu electrodes. • write electrode half reactions and overall reactions for the electrolysis of dil. H_2SO_4, NaOH, dil NaCl, conc. NaCl and CuSO_4 solutions • distinguish between molten electrolytes and aqueous electrolytic solutions, <ul style="list-style-type: none"> • State Faraday's first law of electrolysis. • write the mathematical expression for the Faraday's first law of electrolysis, • do calculations related to Faraday's first law of electrolysis, • state Faraday's second law of electrolysis. • write the mathematical expression for the Faraday's second law of electrolysis. • do calculations related to Faraday's second law of electrolysis. | <p style="text-align: center;">4.3 Quantitative Aspects of Electrolysis (3 periods)</p> <ul style="list-style-type: none"> • Faraday's First Law of Electrolysis • Faraday's Second Law of Electrolysis | <p>Ask students to tell the difference between the electrolysis of molten electrolytic and aqueous electrolytic solutions.</p> <p>Students should know that electrolytic reactions are no exception to the general quantitative relationship of the amount of energy involved and the amount of chemical produced. Students should be asked to formulate the two Faraday's laws of electrolysis. Quantitative and qualitative significance of the laws have to be explored from numerous qualitative and quantitative problems. Students should be asked to practice calculations involving Faraday's laws of electrolysis as class work and assignment.</p> |

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|--|---|---|
| <ul style="list-style-type: none"> • mention industrial applications of electrochemistry, • explain electroplating and electro refining, • demonstrate copper refining, • explain how electrolysis is used in the production of some metals, nonmetals and compounds, • define voltaic cell and salt bridge • draw and label Zn - Cu voltaic cell, • define electrode potential and cell potential, • construct Zn - Cu voltaic cell • measure the cell potential of Zn - Cu cell using voltmeter • explain how standard electrode potential is measured • describe the reactivity of a metal from its position in the activity series or electro motive series | <p>4.4 Industrial Application of Electrolysis (2 periods)</p> <ul style="list-style-type: none"> • Electroplating and Electro refining • Extraction of metals • Manufacturer of Non-metals and some compounds <p>4.5 Voltaic Cells (5 periods)</p> <ul style="list-style-type: none"> • Construction of Zn - Cu voltaic cell • Cell diagram (Notation for Voltaic cell) • Electromotive Force • Measuring standard electrode potentials | <p>Electrolysis has wide applications. Students should be asked to list down some of them in class. They should also be asked to explain to class how they are related to electrolysis. Let the students conduct an experiment on refining of copper. When they discuss electroplating they should know which electrode is the metal to be plated and vice versa. Ask students which elements are extracted from their ores only through electrolysis and why.</p> <p>Students should know that voltaic cells have important roles to play in society. The cells are divided into primary and secondary. Students have to be careful when handling such devices. Ask the students to write chemical equations for the reactions they think take place. They should be able to identify the various components of the cells. Ask students to draw diagrams of the various types of voltaic cells.</p> <p>Students should construct a simple galvanic cell (Daniel's cell). They should identify the parts and measure the cell potential.</p> <p>They should also measure potentials of half-cell using standard half-cells. Finally students can make their own EMS of given elements.</p> |

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|--|---|--|
| <ul style="list-style-type: none"> • calculate cell potential • decide whether a given redox reaction is spontaneous or not. • explain the effect of concentration on cell potential • Mention the different types of voltaic cells • give examples of each type of voltaic cell • describe how Hydrogen-Oxygen Fuel cell operates • distinguish between primary and secondary cells • identify the cathode, anode, and the electrolyte of a given voltaic cell, • compare and contrast electrolytic and voltaic cells • explain metallic corrosion in terms of redox reaction • explain the negative effects of corrosion, • explain the different methods of prevention of corrosion | <ul style="list-style-type: none"> • Calculation of cell potential • Dependence of emf on concentration. • Types of voltaic cells, • Corrosion of metals. | <p>Students should exercise calculating the cell potential and decide the spontaneity of the reactions. Assign students to collect and bring used and disposed batteries and ask them to identify their types. Students have to be careful when handling such devices.</p> <p>Students should understand how hydrogen-oxygen fuel cell operates. Students should be able to tell the differences between primary, secondary and fuel cells.</p> <p>Help students design experiments that will help them identify the factors responsible for corrosion of metals. Can the students see that this is a spontaneous reaction. Would we have to worry if it was a non-spontaneous reaction?</p> |

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define redox reaction, oxidation and reduction; identify the species that are oxidized and reduced in a given redox reaction, and describe the oxidizing and reducing agents; balance a given redox reaction using oxidation number method and ion-electron method; distinguish between molten electrolytes and aqueous electrolytic solutions; draw labeled diagrams of an electrolytic cell; define preferential discharge and explain factors that affect it; describe the reactivity of a metal from its position in the activity series or electro motive series; describe the effect of nature of the ions on the electrolysis of aqueous solutions of dil. H_2SO_4 and NaOH ; describe the effect of concentration of the ions on the electrolysis of dilute and concentrated solutions of NaCl ; describe the effects of types of electrodes on the electrolysis of CuSO_4 solution using Ni electrodes and Cu electrodes; write electrode half reactions and overall reactions for the electrolysis of dil. H_2SO_4 , NaOH , dil NaCl , conc. NaCl and CuSO_4 solutions; state Faraday's first law of electrolysis, write its mathematical expression and do calculations related to it; state Faraday's second law of electrolysis, write its mathematical expression and do calculations related to it; mention industrial applications of electrochemistry; explain electroplating and electro refining;

explain how electrolysis is used in the production of some metals, nonmetals and compounds; define voltaic cell, salt bridge, electrode potential and cell potential; draw, label and construct $\text{Zn} - \text{Cu}$ voltaic cell; measure the cell potential of $\text{Zn} - \text{Cu}$ cell using voltmeter; explain how standard electrode potential is measured and calculate it from the given information; decide whether a given redox reaction is spontaneous or not; explain the effect of concentration on cell potential; mention different types of voltaic cells and give examples of each type of them; distinguish between primary and secondary cells; describe how Hydrogen-Oxygen Fuel cell operates; identify the cathode, anode, and the electrolyte of a given voltaic cell; compare and contrast electrolytic and voltaic cells; explain metallic corrosion in terms of redox reaction and the negative effects of corrosion; explain the different methods of prevention of corrosion.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 5: Some Elements in Nature and Industry (17 periods)

Unit outcomes: Students will be able to:

- Describe occurrence and abundance of the elements in nature;
- Explain how carbon, nitrogen and phosphorus cycle in nature;
- Understand the metallurgical processes;
- Understand the occurrence, extraction & chemical properties of Sodium, Calcium, Tin, Lead, Zinc & Chromium;
- Explain the steps in industrial production of ammonia, nitric acid, sulfuric acid and Diammonium phosphate.
- Describe scientific enquiry skills along this unit: classifying, communicating, asking questions, relating cause & effect and making generalization.

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
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| <p><i>Students will be able to:</i></p> <ul style="list-style-type: none"> • define the terms mineral and ore • describe the forms of occurrences of metals • discuss the distribution and relative amounts of the elements in the earth's crust • define the term fixation • discuss the carbon cycle, the nitrogen cycle and the phosphorus cycle | <p>5. Some Elements in Nature and Industry</p> <p>5.1 Some Elements in Nature (5 periods)</p> <p>5.1.1 Occurrence of elements</p> <ul style="list-style-type: none"> • Sources of elements • Abundance of elements <p>5.1.2 The recycling of elements in nature</p> <ul style="list-style-type: none"> • The carbon cycle | <p>This unit is believed to be handled with ease if portions are assigned to students for class presentation. The students need to be encouraged to bring metal pieces and non-metals from their neighborhoods for viewing.</p> <p>There are 92 elements found in nature. All natural elements are made out of them.</p> <p>Students should be aware that elements can be found either pure or as part of a compound. The elements, in their pure form, are found in one of the three states at ambient conditions of temperature and/or pressure. Most non-metallic elements are gases; whereas most metallic elements are solids. Variety of techniques can be applied to separate elements in their pure form from their sources based on differences in physical properties. Some elements are not found abundantly in pure form; therefore they are extracted from compounds where they are abundant utilizing physical and chemical methods. Most metallic elements, however, are found in ores, substances that contain desirable constituents in concentrations large enough to make their recovery economically feasible.</p> <p>The financial costs of mining, separating and purifying an element are considered before a process is chosen for implementation.</p> <p>Students already know matter is conserved; this of course does not preclude elements from undergoing chemical and physical changes that affect the amount of a certain element in a given form and at a given location without affecting the total sum in the different layers of the crust: lithosphere, hydrosphere and atmosphere.</p> <p>Students should appreciate that carbon is found in all layers of the earth's crust and living things. It is found as CO₂ (g) in the air which can be removed in the form of glucose during photosynthesis. It then becomes part of the plants and animals that feed on plants. The dissolved CO₂ in the ocean becomes part of the various carbonates. Ask students what impact afforestation, deforestation and building plenty of cement factories will have on global</p> |

| Competencies | Contents | Suggested Activities |
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| <ul style="list-style-type: none"> • define metallurgy, • explain the major steps in metallurgical processes, • describe the manufacture of sodium by the Down's cell, • explain the chemical properties of sodium • describe the uses of sodium, | <ul style="list-style-type: none"> • The nitrogen cycle • The phosphorus cycle 5.2 Some Elements in Industry (12 periods) 5.2.1 Metallurgy <ul style="list-style-type: none"> • Pretreating the ore • Converting the mineral to the element • Refining and alloying 5.2.2 Extraction, properties and uses of some selected metals • Sodium <ul style="list-style-type: none"> - Extraction - Chemical properties - Uses | <p>warming and why. Describe the role of photosynthesis, respiration, and decay in the carbon cycle.</p> <p>Students should appreciate that nitrogen is 80% of the air in the atmosphere but is not usable in this form by living things. Students should appreciate this from the type of bond elemental nitrogen has. Oxides of nitrogen, which eventually are converted to nitrates become water soluble, making them available for plant usage. The plants, when eaten by animals pass the nitrogen to them. Synthesis of biologically important molecules such as amino acids and nucleic acids require nitrogen.</p> <p>Put students into two groups to debate why fertilizers like urea and diammonium phosphate are extensively used. What role do the nitrogen-fixing bacteria in legumes play in the nitrogen cycle?</p> <p>Students should know that there is no gaseous form of pure or in a combined form of phosphorus. It is found only in water and land. Students should appreciate how overuse of phosphorus in the form of fertilizers and detergents has affected the environment negatively.</p> <p>Students should be able to see how metallurgy, the science and technology of metals, is used in the commercial production of metals. The three principal steps in metallurgy are concentrating and pretreating the ore, converting the mineral to the element by reduction and refining and alloying.</p> <p>Discuss each step why and how it is done and ask students to support their reasoning with chemical principles they already know.</p> <p>For all elements discussed in this section, chemical principles have to be highlighted why certain extraction methods but not others have been used.</p> <p>Ask first the students to identify the methods of extraction, the chemical they have and their uses for all six elements being considered here.</p> <p>Students should know that among the alternative reduction methods, extraction of sodium and calcium uses electrochemical method of reducing fused NaCl and fused CaCl₂ respectively, but using coke (elemental C) does not work, why?</p> <p>Calcium can also be prepared by reduction of CaO by aluminum in a vacuum, where the calcium produced distills off.</p> |

| Competencies | Contents | Suggested Activities |
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| <ul style="list-style-type: none"> • describe the manufacture of calcium • explain the chemical properties of calcium, • describe the uses of calcium, • describe the manufacture of tin, • explain the chemical properties of tin, • describe the uses of tin, • describe the manufacture of lead, • explain the chemical properties of lead, • describe the uses of lead, • describe the manufacture of zinc, • explain the chemical properties of zinc, • describe the uses of zinc, • describe the manufacture of chromium, • explain the chemical properties of chromium, • describe the uses of chromium. • Describe production of silicon • Explain the chemical properties of silicon • Describe uses of silicon | <ul style="list-style-type: none"> • Calcium <ul style="list-style-type: none"> - Extraction - Chemical properties - Uses • Tin <ul style="list-style-type: none"> - Extraction - Chemical properties - Uses • Lead <ul style="list-style-type: none"> - Extraction - Chemical properties - Uses • Zinc <ul style="list-style-type: none"> - Extraction - Chemical properties - Uses • Chromium <ul style="list-style-type: none"> - Extraction - Chemical properties - Uses 5.2.3. Silicon <ul style="list-style-type: none"> • Production • Chemical properties • Uses | <p>Students should know the chemical properties and common uses of sodium and calcium.</p> <p>Students should know that tin, lead, zinc and chromium are extracted from their ores by using a reducing agent coke (elemental carbon) in furnace, unlike sodium and calcium which can't.</p> <p>They should also know that lead and zinc are produced in two steps. They are:</p> <ul style="list-style-type: none"> • First, roasting: oxidation of sulfides to oxides. Because it is easier to reduce oxides. • Next, reduction: reduction of oxides to their elements using a reducing agent coke(elemental carbon) <p>Students should know the chemical properties and common uses of Tin, lead, zinc and chromium.</p> <p>Students should appreciate how silicon is prepared by reducing quartz sand, SiO₂, with coke, C, in electric furnace at 3000 °C . The teacher should emphasize the semi-metallic nature of silicon and its modern uses based on semi-conductivity.</p> |

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|---|--|---|
| <ul style="list-style-type: none"> • Explain the steps in Haber process, the industrial production of ammonia and its uses • Explain the steps in Ostwald's process, the industrial production of nitric acid and its uses • Explain the steps in Contact process, the industrial production of sulphuric acid and its uses • Explain the steps in the industrial production of Diammonium Phosphate and its uses | <p>5.2.4. Some important compounds of selected nonmetals</p> <ul style="list-style-type: none"> • Ammonia • Nitric acid • Sulphuric acid • Diammonium Phosphate (DAP) | <p>One could always select different molecules for discussion based on the criteria used for selection. But these ones are among the many considered to be important in today's world. In each one of them go through the following points:</p> <ul style="list-style-type: none"> • elemental composition, structure, polarity, properties and uses • use the methodical approach and work your way by using the elemental composition as an explanation for the structure and so on |

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define the terms mineral, ore and metallurgy; describe the forms of occurrences of metals; discuss the distribution and relative amounts of the elements in the earth's crust; define the term fixation; discuss the carbon cycle, the nitrogen cycle and the phosphorus cycle; explain the major steps in metallurgical processes; describe the manufacture of sodium, calcium, tin, lead, zinc and chromium; explain the chemical properties of sodium, calcium, tin, lead, zinc and chromium; describe the uses of sodium, calcium, tin, lead, zinc and chromium; explain the steps in Haber process, the industrial production of ammonia, the Ostwald's process, the industrial

production of nitric acid, the contact process, industrial production of sulphuric acid, the industrial production of Diammonium Phosphate, and the industrial production of Silicates ;

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 6: Polymers (14 periods)

Unit outcomes: Students will be able to:

- Understand the process of addition & condensation polymerization;
- List a variety of synthetic polymers & natural polymers & explain their differences;
- Classify synthetic polymers as addition or condensation polymers & identify their monomers;
- Describe the monomers, properties & uses of plastics, rubbers, carbohydrates & proteins;
- Analyze the risks & benefits of the development & application of synthetic polymers(e.g. plastics) & suggest possible methods of control;
- Describe scientific enquiry skills along this unit: classifying, comparing & contrasting, communicating, asking questions, applying concept and making generalization.

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|--|---|--|
| <p><i>Students will be able to:</i></p> <ul style="list-style-type: none"> • define the terms monomer and polymer; • classify polymers into synthetic and natural polymers; • give examples of synthetic and natural polymers; <ul style="list-style-type: none"> • explain polymerization; • mention the two types of polymerizations; • explain how addition polymerization takes place; • explain how condensation polymerization takes place; <ul style="list-style-type: none"> • give examples of addition polymers; • tell the monomers of each of the given addition polymers | <p>6. Polymers</p> <p>6.1 Polymerization (2 periods)</p> <ul style="list-style-type: none"> • Addition polymerization • Condensation polymerization <p>6.2 Synthetic polymers (5 periods)</p> <ul style="list-style-type: none"> • Addition polymers <ul style="list-style-type: none"> - Polyethylene - Polypropylene - Polyvinyl chloride | <p>Students should know that macromolecules are made of smaller units(monomers). They could be made of identical monomers- homopolymers or different types of monomers-copolymers.</p> <p>Polymers could be:</p> <ul style="list-style-type: none"> • Synthetic • Natural <p>Ask students to list down examples of polymers from their surroundings and classify them as synthetic or natural polymers.</p> <p>Students should be aware that polymerization could be effected either through:</p> <ul style="list-style-type: none"> • Addition polymerization or • Condensation polymerization <p>Give students reading assignment on the difference between addition and condensation polymerization to be presented to fellow students. Help them identify the functional groups involved in each case. This is a good opportunity to show students how applied organic chemistry is in every day life both synthetically and naturally.</p> <p>Students should know that the number of synthetic polymers is increasing rapidly in response to the needs of mankind. Ask if students know what problems are arising with this development.</p> <p>Do they have opinions about what needs to be done to protect the environment? You could tell them at this point there are fields of study known as Environmental Science and even Environmental Chemistry.</p> |

| Competencies | Contents | Suggested Activities |
|--|--|--|
| <ul style="list-style-type: none"> • explain the common uses of addition polymers; • give examples of condensation polymers; • tell the monomers of each of the given condensation polymers; • explain the uses of condensation polymers: nylon, polyester, and bakelite; • describe thermoplastics and thermo set polymers • give examples of thermoplastics and thermo set polymers; • describe natural rubber; • tell the monomer of natural rubber; • explain vulcanization • explain the use of natural rubber; • give three examples of synthetic rubbers; • explain uses the examples of synthetic rubbers; • compare and contrast synthetic and natural rubber; | <ul style="list-style-type: none"> - Polymethylmetacrylate (Perspex) - Polystyrene - PolyTetra Flouro Ethylene (PTFE) or Teflon • Condensation polymers <ul style="list-style-type: none"> - Nylon - Polyester (Dacron) - Bakelite • Thermoplastics and thermo set polymers <p>6.3 Natural polymers (7 periods)</p> <ul style="list-style-type: none"> • Rubber <ul style="list-style-type: none"> - Natural Rubber - Synthetic Rubber | <p>Students should understand problems arise when plastics are discarded into the environment. Assign students to collect synthetic polymers which they think are useful or with serious drawbacks. Could they identify the polymers, the corresponding monomers and whether they are addition polymers or condensation polymers? You should appreciate the effort they put but be ready to answer the questions you asked them; they may not have the answers. You may not know what they have collected. It is not always easy. You could bring properly identified samples.</p> <p>Students should be aware one good thing about synthetic polymers is that they can design and produce materials with properties of their desire. Students may want flexible but of considerable strength or hard and brittle polymers. Students may want to reuse old polymers as starting materials for new ones; in other wards to recycle substances. Ask the students what uses recycling has. What makes thermoplastic polymers recyclable whereas thermoses ones are not? Ask students if they can see similarity about intermolecular forces in solids covered in grade 11 and thermoplastic and thermo set polymers.</p> <p>There is a popular saying, 'Nature had it first'. Long before mankind started making synthetic polymers there were natural polymers everywhere. Ask students to list down as many natural polymers as they could find around them or even within them.</p> <p>Can they see that the number of natural polymers is almost constant but every day new synthetic polymers are made.?</p> <p>Students should know the monomers for both synthetic and natural rubber polymers. Discuss the role of sulfur in making rubber harder (vulcanization). You may bring this up when you discuss proteins that use similar bonding with cysteine amino acid.</p> |

Chemistry: Grade 12

| <i>Competencies</i> | <i>Contents</i> | <i>Suggested Activities</i> |
|--|---|--|
| <ul style="list-style-type: none"> • define the term carbohydrate • describe mono saccharides; • draw the structure of monosaccharide, • give examples of mono saccharides • describe disaccharides; • draw the structure of disaccharides; • give examples of disaccharides; • describe poly saccharides; • draw structures of starch and cellulose; • explain the difference between starch, glycogen, and cellulose; • define the terms amino acid, peptide and protein; • describe the structure of amino acids; • explain how proteins are formed; list types of proteins; | <ul style="list-style-type: none"> • Carbohydrates <ul style="list-style-type: none"> - Monosaccharide - Disaccharide - Polysaccharide • Proteins <ul style="list-style-type: none"> - Amino acid - Polypeptides | <p>Carbohydrates are compounds that have C, H and O. You may tell your students that carbohydrates are poly hydroxyl aldehydes or polyhydroxy ketones. Most carbohydrates are homopolymers of simple sugars known as monosaccharides. The monosaccharides are the monomers for polysaccharides which could contain thousands of monomers. Most polysaccharides are homopolymers. Warn your students to pay special attention to the monomer, glucose, as it has a bigger role than any other monosaccharide in biochemistry. Ask students to draw various structures of few monosaccharides including glucose. In water solution the cyclic conformations are dominant.</p> <p>Ask the students whether there is any similarity between starch, cellulose and glycogen. Which ones can be used as energy sources by mammalian cells? Do they give the comparable amount of energy if combusted in a crucible? Do you see reasons for the differences, if any? Discuss this from bonding point of view. Ask students if they can identify synthetic fibers used as clothing that were developed on principles learned from carbohydrates? Are carbohydrates condensation or addition polymers?</p> <p>Students should know that proteins got their names from a similar Greek word meaning of primary importance. They are made by polymerization of twenty naturally existing amino acids. Students should know why they are called amino acids. Discuss some of their properties that are important for polymerization.</p> <p>Try to show similarity in the way amino acids are polymerized to form proteins with nylon. Why do amino acids and monomers of nylon have two functional groups? Ask students to see that proteins are condensation polymers. What small molecule is released as a byproduct here?</p> <p>Limit the discussion to introducing that they are polymers of nucleotides, which themselves consist of an organic base, the ribose (or deoxy) sugar and at least one phosphate group. Ask the students if they already know the two types of nucleic acids, RNA and DNA. They are called nucleic acids because they were isolated and identified as acids because of the phosphate group, from the nucleus of eukaryotic cells.</p> |

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define the terms monomer and polymer; classify polymers into synthetic and natural polymers and give examples for each of them; explain polymerization and mention the two types of polymerizations; explain how addition and condensation polymerizations take place; give examples of addition polymers and tell the monomers of each of them; explain the common uses of addition polymers; give examples of condensation polymers and tell the monomers of each of them; explain the uses of condensation polymers: nylon, polyester, and bakelite; describe thermoplastics and thermo set polymers and give examples for each of them; describe natural rubber and tell its monomer; explain vulcanization and the uses of natural rubber; give some examples of synthetic rubbers and

explain their uses; compare and contrast synthetic and natural rubber; define the term carbohydrate; describe monosaccharide, draw its structure and give examples; describe disaccharides, draw their structure and give examples; describe polysaccharides; draw structures of starch and cellulose; explain the difference between starch, glycogen, and cellulose; define the terms amino acid, peptide and protein; describe the structure of amino acids; explain how proteins are formed; list types of proteins.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

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