THE UNITED REPUBLIC OF TANZANIA

NATIONAL EXAMINATIONS COUNCIL OF TANZANIA

ADVANCED CERTIFICATE OF SECONDARY EDUCATION EXAMINATION

132/1 **CHEMISTRY 1**

(For Both School and Private Candidates)

Time: 3 Hours **ANSWERS** **Year: 2004**

Instructions

1. This paper consists of sections A, B and C with total of fourteen questions
2. Each question carries ten marks.

1. (a) (i) Outline the fundamental postulates of Planck's quantum theory.

i. Energy is not emitted or absorbed continuously but in discrete packets called quanta or photons.
ii. The energy of each quantum is proportional to the frequency of radiation and is given by $E = h\nu$, where h is Planck's constant and ν is the frequency.
iii. Energy is exchanged only in whole-number multiples of quanta, meaning a system can emit or absorb energy only in discrete amounts.
iv. Higher energy quanta correspond to electromagnetic waves with higher frequency.
v. Black body radiation is explained using quantization of energy, resolving the ultraviolet catastrophe issue in classical physics.

(ii) What does the term wave-particle duality nature of an electron mean?

i. Wave-particle duality means that electrons exhibit both wave-like and particle-like behavior.
ii. As a particle, an electron has mass, charge, and a definite position when measured.
iii. As a wave, an electron exhibits diffraction and interference, as demonstrated in the double-slit experiment by Thomas Young.
iv. According to de Broglie's hypothesis, the wavelength of an electron is given by $\lambda = h/mv$, where h is Planck's constant, m is mass, and v is velocity.

(iii) State the uncertainty principle.

i. Heisenberg's Uncertainty Principle states that it is impossible to simultaneously determine both the exact position and momentum of an electron with absolute certainty.
ii. Mathematically, it is given by $\Delta x \, \Delta p \geq h/4\pi$, where Δx is the uncertainty in position and Δp is the uncertainty in momentum.
iii. The more precisely we measure an electron's position, the less precisely we can measure its momentum, and vice versa.

(b) Calculate the energy of a photon of radiations whose wavelength is 5.89×10^{-7}cm emitted from sodium atoms when heated.

i. The energy of a photon is given by the equation $E = h\nu = hc/\lambda$.
ii. Given:
- Wavelength, $\lambda = 5.89 \times 10^{-7}$ cm $= 5.89 \times 10^{-9}$ m
- Planck's constant, $h = 6.626 \times 10^{-34}$ J·s
- Speed of light, $c = 3.0 \times 10^{8}$ m/s

iii. Substituting values:
$E = (6.626 \times 10^{-34} \times 3.0 \times 10^{8}) / (5.89 \times 10^{-9})$
$E = 3.37 \times 10^{-19}$ J

2. (a) (i) State Boyle's gas law.

i. Boyle's law states that at constant temperature, the pressure of a fixed amount of gas is inversely proportional to its volume.
ii. Mathematically: $P_1V_1 = P_2V_2$.
iii. This means that if pressure increases, volume decreases, and vice versa, provided temperature remains constant.

(ii) State Charles's gas law.

i. Charles's law states that at constant pressure, the volume of a fixed amount of gas is directly proportional to its absolute temperature (Kelvin scale).
ii. Mathematically: $V_1/T_1 = V_2/T_2$.
iii. This means that when temperature increases, gas volume expands, and when temperature decreases, gas volume contracts.
(b) Under what conditions do real gases fail to obey Charles's and Boyle's gas laws?

i. High pressure: At high pressures, gas molecules are forced closer together, and intermolecular forces become significant, causing deviations.
ii. Low temperature: At low temperatures, kinetic energy decreases, and intermolecular attractions cause deviations from ideal behavior.
iii. High molecular mass gases: Heavy gases with stronger intermolecular forces deviate more from ideal gas laws.
iv. Polar molecules: Polar gases (like NH_3, CO_2) experience intermolecular attractions, making them behave non-ideally.

(c) With the help of Amagat's curves, show how carbon dioxide gas deviates from ideal gas behaviour.

i. Amagat's curves are graphs showing the relationship between volume and pressure for gases at different temperatures.
ii. For ideal gases, the volume decreases linearly with increasing pressure (Boyle's law).
iii. For real gases (like CO_2):
- At high pressure, the volume is smaller than predicted by Boyle's law due to attractive forces.
- At very high pressure, the volume becomes larger than expected due to the finite size of molecules.

iv. This deviation is explained by the van der Waals equation: $(P + a/V^2)(V - b) = RT$, where a accounts for intermolecular forces, and b accounts for molecular volume.

3. (a) (i) State Raoult's law.

i. Raoult's law states that the partial vapor pressure of a component in a solution is directly proportional to its mole fraction.
ii. Mathematically: $P_1 = X_1 P_1^0$, where P_1 is the vapor pressure of the component in solution, X_1 is its mole fraction, and P_1^0 is the vapor pressure of the pure component.

(b) Under what conditions does a solution formed upon mixing two liquids A and B behaves as an ideal solution?

i. No volume change on mixing: The total volume of A and B remains unchanged after mixing.
ii. No enthalpy change: The heat of mixing (ΔH) is zero.
iii. Obeys Raoult's Law: Partial vapor pressures follow the equation $P\_total = P\_A + P\_B$.
iv. Similar intermolecular forces: The forces between A-A, B-B, and A-B are nearly the same.
v. No azeotrope formation: The solution does not deviate positively or negatively from Raoult's law.

4. (a) (i) State a homogeneous catalyst.

i. A homogeneous catalyst is a catalyst that exists in the same phase as the reactants, usually in a single-phase system (liquid or gas).
ii. Example: The use of H_2SO_4 in the esterification of ethanol and acetic acid.

(ii) State a heterogeneous catalyst.

i. A heterogeneous catalyst is a catalyst that exists in a different phase from the reactants, typically a solid catalyst in a liquid or gas reaction.
ii. Example: The use of Fe in the Haber process ($N_2 + 3H_2 \rightarrow 2NH_3$).

(iii) State an autocatalyst.

i. An autocatalyst is a product of the reaction that acts as a catalyst, increasing the rate of the reaction as more product is formed.
ii. Example: Mn^{2+} ions acting as an autocatalyst in the reaction of MnO_4^- with oxalic acid.

(b) Deduce the relationship between Kp and Kc for the following gaseous equilibria:
$2NO_2(g) + 1/2O_2(g) \rightleftharpoons 2NO(g)$

i. The general relationship between Kp and Kc is given by $Kp = Kc(RT)\hat{}(\Delta n)$, where Δn is the change in the number of moles of gaseous products and reactants.
ii. In this reaction:
  - Reactant moles = (2 + 1/2) = 2.5
  - Product moles = 2
  - $\Delta n = 2 - 2.5 = -0.5$
iii. Thus, $Kp = Kc(RT)\hat{}(-0.5)$.

(c) Consider the following equilibrium:
$Cr_2O_7^{2-}(aq) + H_2O(l) \rightleftharpoons 2CrO_4^{2-}(aq) + 2H^+(aq)$

(i) What would you expect to see if sodium hydroxide is added to the equilibrium mixture?

i. Adding NaOH increases OH^- concentration, which reacts with H^+ to form water, reducing H^+ concentration.
ii. This shifts the equilibrium to the right, increasing CrO_4^{2-} formation, changing the solution color from orange ($Cr_2O_7^{2-}$) to yellow (CrO_4^{2-}).

(ii) What would you expect to see if dilute hydrochloric acid is added to the equilibrium mixture?

i. Adding HCl increases H^+ concentration, shifting the equilibrium to the left to favor $Cr_2O_7^{2-}$ formation.
ii. The solution turns from yellow (CrO_4^{2-}) to orange ($Cr_2O_7^{2-}$).

5. (a) (i) What does the term colligative properties mean?

i. Colligative properties are properties of solutions that depend only on the number of solute particles in the solution and not on their identity.
ii. Examples include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

(ii) Outline the limitations of colligative properties.

i. They assume ideal solution behavior, which is not always accurate in real solutions.
ii. They do not consider ion pairing or complex formation in ionic solutions.
iii. The presence of electrolytes that dissociate into multiple ions leads to deviations from expected values.
iv. Non-volatile solutes may interact with solvents, causing deviations from Raoult's law.

(b) Study the table below and answer the questions that follow:

Solution	Concentration (mol/L)	Freezing Point (°C)
Cane sugar ($C_{12}H_{22}O_{11}$)	0.030	-0.054
Glycerol ($C_3H_8O_3$)	0.132	-0.200
Potassium bromide (KBr)	0.084	-0.300

i. Explain why potassium bromide has the most negative freezing point.

- Potassium bromide dissociates into two ions (K^+ and Br^-) in solution, increasing the number of solute particles.
- More particles lead to a greater colligative effect, further depressing the freezing point.

(c) The vapor pressure of 1% urea solution at 38°C is 49.85 mmHg. If the vapor pressure of pure water is 50 mmHg and 250 mmHg at 38°C and 72°C, respectively, calculate the vapor pressure of 1% urea solution at 72°C.

i. The relative lowering of vapor pressure is given by Raoult's law: $(P_0 - P) / P_0 = X\_solute$
ii. Using the data at 38°C: $(50 - 49.85) / 50 = X\_solute$
iii. Using the same X_solute at 72°C: $(250 - P) / 250 = X\_solute$
iv. Solving for P at 72°C, P = 249.25 mmHg.

6. (a) (i) Write down a balanced equation for the dissociation of phosphorus pentachloride gas and show the ratio of volumes of reactants to the volumes of products.

- $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$
- The mole ratio is 1:1:1, meaning 1 volume of PCl_5 produces 1 volume of PCl_3 and 1 volume of Cl_2.

(ii) Write down the expression which relates the degree of dissociation of phosphorus pentachloride to its relative densities before and after dissociation.

- The degree of dissociation (α) is given by:
  $\alpha = (D\_initial - D\_final) / (D\_initial - D\_products)$

(b) Define the following:

(i) Normal density of a substance.
- Normal density is the mass per unit volume of a substance under standard conditions of temperature and pressure.

(ii) Relative density of a gas.
- Relative density (or vapor density) is the ratio of the density of a gas to the density of hydrogen or air under the same conditions.

(iii) The mass of $243cm^3$ of a gas at 273K and at an atmospheric pressure of 1 atmosphere is 0.162g. Calculate the normal density and relative density of the gas.

i. Density = mass/volume = $0.162g / 243cm^3 = 0.000667 g/cm^3$.
ii. Relative density (compared to hydrogen) = density of gas / density of H_2.

7. Explain the following

(a) Although water and hydrogen sulphide are the hydrides of elements of the same group, the boiling point of water is higher than that of hydrogen sulphide.

Water (H_2O) and hydrogen sulphide (H_2S) are both hydrides of group 16 elements, oxygen and sulphur, respectively. The significant difference in their boiling points arises due to intermolecular forces. Water molecules are held together by strong hydrogen bonds, which require more energy to break, leading to a higher boiling point. In contrast, hydrogen sulphide molecules experience only weak van der Waals forces and dipole-dipole interactions, requiring less energy to separate them, resulting in a much lower boiling point.

(b) Iodine and chlorine both are found in group seven of the periodic table but hydrogen iodide is a stronger reducing agent than hydrogen chloride.

Hydrogen iodide (HI) and hydrogen chloride (HCl) are halogen acids, but their reducing abilities differ due to bond strength and electronegativity. The H-I bond is weaker than the H-Cl bond due to the larger atomic size of iodine compared to chlorine. As a result, HI dissociates more readily, releasing hydrogen atoms that act as reducing agents. Chlorine, being more electronegative, holds onto the hydrogen more tightly, making HCl a weaker reducing agent.

(c) The acidity of the solutions of the hydrides of elements from sodium to chlorine across the third period increases from left to right.

Across the third period, the hydrides shift from basic to acidic nature due to the increasing electronegativity of the central atoms. Sodium hydride (NaH) is highly basic because sodium is a metal, while silicon, phosphorus, sulphur, and chlorine form progressively more acidic hydrides due to their increased ability to accept electron pairs and donate protons in aqueous solutions. The trend results from a combination of bond polarity and bond dissociation energy, making chlorine hydride (HCl) the most acidic.

(d) Hydrogen iodide does not form hydrogen bonding while hydrogen fluoride does although both are the hydrides of group seven.

Hydrogen bonding occurs when hydrogen is bonded to highly electronegative atoms such as fluorine, oxygen, or nitrogen. In hydrogen fluoride (HF), the high electronegativity and small size of fluorine create strong hydrogen bonds between molecules. However, in hydrogen iodide (HI), iodine is larger and less electronegative, leading to weaker dipole-dipole interactions instead of hydrogen bonding.

(e) The passage of carbon dioxide through calcium hydroxide solution changes the latter into white precipitate which dissolves into a clear solution in excess of carbon dioxide.

When carbon dioxide (CO_2) is bubbled through calcium hydroxide ($Ca(OH)_2$), a white precipitate of calcium carbonate ($CaCO_3$) forms due to the reaction:

$Ca(OH)_2 + CO_2$ -------> $CaCO_3 + H_2O$

Upon further passage of CO_2, the precipitate dissolves as calcium carbonate reacts with excess carbon dioxide to form soluble calcium bicarbonate:

$CaCO_3 + CO_2 + H_2O$ -------> $Ca(HCO_3)_2$

This results in a clear solution.

8. Below is a portion of the electrochemical series. Study it carefully and answer the questions that follow:

$Fe \text{------>} Fe^{2+} + 2e^-$ -0.44
$Sn \text{----->} Sn^{2+} + 2e^-$ -0.14
$Pb \text{------>} Pb^{2+} + 2e^-$ -0.13
$1/2H_2 \text{---->} H^+ + e^-$ 0.00
$Cu \text{---->} Cu^{2+} + 2e^-$ +0.34
$2Cl^- + 7H_2O \text{----->} 6e^- + Cr_2O_7^{2-} + 14H^+$ +1.33
$Cl^- \text{---->} 1/2Cl_2 + e^-$ +1.36
$Mn^{2+} + 4H_2O \text{---->} 5e^- + MnO_4^- + 8H^+$ +1.69

(a) Hydrogen gas can be prepared from the reaction of dilute hydrochloric acid and iron but not from the reaction of dilute hydrochloric acid and copper. Explain.

Iron (Fe) is more reactive than hydrogen, as seen from its lower electrode potential (-0.44 V). It can displace hydrogen ions from dilute hydrochloric acid:

$Fe + 2HCl \text{---->} FeCl_2 + H_2$

Copper (Cu), however, has a positive electrode potential (+0.34 V), making it less reactive than hydrogen. It cannot displace hydrogen from dilute hydrochloric acid, so no reaction occurs.

(b) Although lead could be a perfect coating material for iron metal against corrosion, food cans made of iron are coated with tin metal and not lead. Explain.

Lead (Pb) is toxic and can contaminate food, making it unsuitable for coating food cans. Additionally, lead is amphoteric and reacts with acids found in food, forming soluble lead compounds. Tin (Sn), on the other hand, is non-toxic, has a stable oxide layer, and provides effective corrosion resistance.

(c) During potassium (i) permanganate (vii) titrations, dilute hydrochloric acid cannot be used as an acidic medium while during the titration of potassium (i) dichromate (vi), dilute hydrochloric acid is used. Explain.

In potassium permanganate ($KMnO_4$) titrations, hydrochloric acid (HCl) cannot be used because permanganate (MnO_4^-) oxidizes chloride ions to chlorine gas (Cl_2), leading to side reactions:

$MnO_4^- + Cl^- + H^+ \text{---->} Mn^{2+} + Cl_2 + H_2O$

In potassium dichromate ($K_2Cr_2O_7$) titrations, hydrochloric acid is used because dichromate does not oxidize chloride ions significantly, allowing the reaction to proceed correctly.

(d) Do you expect any of the following pairs of reactants to react?

$Fe(s) + CuSO_4(aq)$ ---> $FeSO_4 + Cu$
$Cu(s) + SnCl_2(aq)$ -----> No reaction

Iron is more reactive than copper and has a more negative electrode potential, so it displaces copper from copper sulfate. However, copper is less reactive than tin and cannot displace tin from tin(II) chloride.

(e) Give the symbol of a cell of the highest e.m.f. which can be constructed from the above portion of the electrochemical series. What is the value of its e.m.f?

The highest e.m.f. is obtained from the combination of MnO_4^- (+1.69 V) and Fe^{2+} (-0.44 V).

E_cell = E_cathode - E_anode
= (+1.69) - (-0.44)
= +2.13 V

The cell symbol is:

$MnO_4^-/Mn^{2+} \,||\, Fe^{2+}/Fe$

I will continue answering the remaining questions in the next response.

9. (a) Write down the stable electronic configuration of the following:

(i) Fe^{2+} = [Ar] $3d^6$
(ii) Na^+ = [Ne]
(iii) Mn^{2+} = [Ar] $3d^5$
(iv) Mg^{2+} = [Ne]
(v) Mn^{3+} = [Ar] $3d^4$
(vi) Fe^{3+} = [Ar] $3d^5$

(b) Explain the following in terms of electronic configuration:

(i) In manganese, the oxidation state of +2 is more stable than the oxidation state of +3.

Manganese has an atomic number of 25 with an electronic configuration of [Ar] $3d^5$ $4s^2$. In Mn^{2+}, the configuration is [Ar] $3d^5$, which is a half-filled d-orbital and provides extra stability due to exchange energy. In Mn^{3+}, the configuration is [Ar] $3d^4$, which is less stable because it disrupts the half-filled d-orbital arrangement.

(ii) The size of Na^+ ion is larger than that of Mg^{2+} ion although both have the same number of electrons.

Both Na^+ and Mg^{2+} have 10 electrons (isoelectronic), but Mg^{2+} has a higher nuclear charge (+12) compared to Na^+ (+11). The increased nuclear charge in Mg^{2+} pulls the electrons closer to the nucleus, reducing its atomic size, whereas Na^+ experiences a weaker attraction, making it larger in size.

(c) Give the coordination number, IUPAC names, and the names of geometrical shapes of the following complex compounds

(i) $[FeSO_4(NO)_2COBrOH]$

Coordination number: 6
IUPAC name: Iron(II) sulphato dinitrosyl carbonyl bromo hydroxide
Geometrical shape: Octahedral

(ii) $[NiCNH_3OC_2O_4NH_2SO_3]^{4-}$

Coordination number: 6
IUPAC name: Nickel(II) ammine oxalato amido sulphito complex
Geometrical shape: Octahedral

10. (a) Define the following terms:

(i) Equivalent weight

The equivalent weight of a substance is the mass of the substance that reacts with or replaces one mole of hydrogen ions (H^+) or one mole of electrons in a redox reaction. It is calculated as:

Equivalent weight = Molecular weight / Valency

(ii) Molality

Molality (m) is the number of moles of solute per kilogram of solvent. It is given by:

Molality = Moles of solute / Mass of solvent (kg)

(iii) Normality

Normality (N) is the number of gram-equivalents of solute per liter of solution. It is given by:

Normality = Equivalent weight of solute / Volume of solution (L)

(iv) Mole fraction

Mole fraction is the ratio of the number of moles of a particular component to the total number of moles in the mixture. It is given by:

Mole fraction of component A = Moles of A / Total moles of all components

(b) A solution of sulphuric acid contains 571.60 g of the acid per dm^3 of the solution at 20°C has a density of 1.3294 g/cm^3. Calculate:

(i) Normality of the acid solution

Molar mass of H_2SO_4 = 98 g/mol
Number of moles of H_2SO_4 in 571.60 g = 571.60 / 98 = 5.83 moles
Since H_2SO_4 provides 2 H^+ per molecule, its normality is:

Normality = Molarity × n-factor = 5.83 × 2 = 11.66 N

(ii) Molality of the acid solution

Mass of solution in 1 dm^3 = 1.3294 × 1000 = 1329.4 g
Mass of water in solution = 1329.4 - 571.6 = 757.8 g = 0.7578 kg

Molality = Moles of solute / Mass of solvent (kg)
Molality = 5.83 / 0.7578 = 7.69 mol/kg

(iii) Percentage by mass of sulphuric acid in the solution

Percentage by mass = (Mass of solute / Mass of solution) × 100
Percentage by mass = (571.60 / 1329.4) × 100 = 43.0%

(iv) Mole fraction of sulphuric acid and water in the solution

Moles of H_2SO_4 = 5.83
Moles of water = Mass of water / Molar mass of water
= 757.8 / 18 = 42.1

Total moles = 5.83 + 42.1 = 47.93

Mole fraction of H_2SO_4 = 5.83 / 47.93 = 0.122
Mole fraction of water = 42.1 / 47.93 = 0.878

11. A hydrocarbon compound M of molecular formula C_4H_8 reacts with one mole of hydrogen gas in presence of nickel catalyst to form compound P. Compound P can react with one mole of hydrogen at a temperature of 80 – 120°C in presence of nickel catalyst to form compound Q. Compound M reacts readily

with bromine gas and it can decolourize bromine water. Compound P reacts readily with bromine gas to form compound R but P can not decolourize bromine water.

(a) Write the structural formulae of compounds M, P, Q, and R.

M = Butene (C_4H_8)
P = Butane (C_4H_{10})
Q = Cyclobutane (C_4H_8)
R = 1,2-Dibromobutane ($C_4H_8Br_2$)

(b) Write the equations of reaction between compound:

(i) M and hydrogen gas in presence of nickel catalyst to form P

$C_4H_8 + H_2$ -----> C_4H_{10}

(ii) P and hydrogen gas in presence of nickel catalyst to form Q

C_4H_{10} -----> $C_4H_8 + H_2$

(iii) P and bromine gas to form compound R

$C_4H_8 + Br_2$ ------> $C_4H_8Br_2$

12. (a) With the help of chemical equations give one chemical test used to distinguish the following pairs of compounds:

(i) CH_3-CH(Cl)-CH_3 (2-Chloropropane) from $CH_3CH_2CH_2Cl$ (1-Chloropropane)

The silver nitrate test can be used to distinguish these compounds. The tertiary alkyl halide (2-Chloropropane) undergoes a rapid reaction with aqueous silver nitrate in ethanol, forming a precipitate of silver chloride instantly, while the primary alkyl halide (1-Chloropropane) reacts much slower.
Reaction:
CH_3-CH(Cl)-CH_3 + $AgNO_3$ -----> AgCl (white precipitate) + CH_3-CH(OH)-CH_3

$CH_3CH_2CH_2Cl + AgNO_3$ ----->Slow reaction, delayed AgCl precipitate formation

(ii) $CH_3CH_2CH_2OH$ (Propan-1-ol) from CH_3CH_2OH (Ethanol)

Oxidation using acidified potassium dichromate ($K_2Cr_2O_7$) can distinguish these compounds. Primary alcohols oxidize to aldehydes and then carboxylic acids.

Reaction:

CH_3CH_2OH + [O] -----> CH_3CHO + H_2O (color change from orange to green)

$CH_3CH_2CH_2OH$ + [O] -----> CH_3CH_2CHO + H_2O (color change from orange to green, further oxidation possible)

(iii) $CH_3C{\equiv}CCH_3$--->(2-Butyne) from $CH_3CH_2C{\equiv}CH$ (1-Butyne)

Reaction with ammoniacal silver nitrate (Tollens' reagent) can be used to distinguish terminal alkynes from internal alkynes. Terminal alkynes react to form a white precipitate of silver acetylide, while internal alkynes do not react.

Reaction:
$CH_3CH_2C{\equiv}CH + AgNO_3/NH_3$ -----> $AgC{\equiv}CCH_2CH_3$ (white precipitate)

$CH_3C{\equiv}CCH_3 + AgNO_3/NH_3$ ------> No reaction

(iv) Phenol (C_6H_5OH) from Benzyl alcohol ($C_6H_5CH_2OH$)

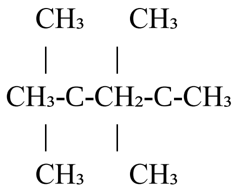
The Ferric chloride ($FeCl_3$) test can distinguish phenols from benzyl alcohols. Phenols give a violet color with $FeCl_3$, whereas benzyl alcohol does not.
Reaction
$C_6H_5OH + FeCl_3$ -----> Violet color

$C_6H_5CH_2OH + FeCl_3$ ----> No color change

13. (a) Correcting the incorrect IUPAC names

(i) 1,1,1,1-Tetraethylmethane (Incorrect)
- The structure implies a central carbon bonded to four ethyl ($-C_2H_5$) groups, which is not possible for methane.
- The correct structure corresponds to 2,2,3,3-Tetramethylbutane.

Correct Structure:



$$CH_3-\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{C}}-CH_2-\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{C}}-CH_3$$

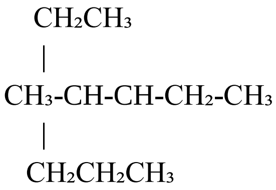
(ii) 1-Chloropentan-4-ol (Incorrect)
- The correct numbering should give priority to the hydroxyl (-OH) over the halogen (-Cl).
- The correct name is 4-Chloropentan-1-ol.

Correct Structure:

$CH_3-CH_2-CH(Cl)-CH_2-CH_2OH$

(iii) 3-Propyl-2-ethylpentane (Incorrect)
- The correct IUPAC naming rule states that substituents must be listed in alphabetical order.
- The correct name is 2-Ethyl-3-propylpentane.

Correct Structure:



$$\begin{array}{l} \quad CH_2CH_3 \\ \quad | \\ CH_3-CH-CH-CH_2-CH_3 \\ \quad | \\ \quad CH_2CH_2CH_3 \end{array}$$

13. (b) IUPAC names of the given compounds

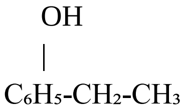
(i) Triple bond and hydroxyl group present
- The structure contains a triple bond and an alcohol (-OH) functional group.
- The longest chain must include both with the lowest numbering priority.
- The correct name is 5-Hexyn-2-ol.

Correct Structure:

$CH_3-C\equiv C-CH_2-CH(OH)-CH_2-CH_3$

(ii) Benzene ring with an ethyl alcohol ($-CH_2CH_2OH$) group
- The correct name is 2-Phenylethanol.

Correct Structure:

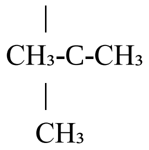


$$\begin{array}{l} \quad OH \\ \quad | \\ C_6H_5-CH_2-CH_3 \end{array}$$

(iii) Branched alkane (Neopentane)
- The correct IUPAC name is 2,2-Dimethylpropane.

Correct Structure:

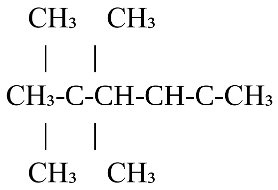
CH_3



$$\begin{array}{c} | \\ CH_3-C-CH_3 \\ | \\ CH_3 \end{array}$$

(iv) Long-chain alkane with four methyl groups
- The correct name is 3,3,4,4-Tetramethylhexane.

Correct Structure:



$$\begin{array}{l} \quad CH_3 \quad CH_3 \\ \quad\;\; | \qquad | \\ CH_3-C-CH-CH-C-CH_3 \\ \quad\;\; | \qquad | \\ \quad CH_3 \quad CH_3 \end{array}$$

14. (a) Complete the following conversions by giving the structural formulae and the IUPAC names of compounds R, S, and T

Step 1: $CH_3CH_2CH_2Cl + NaOH (aq) \text{----->} R + NaCl$

R is formed by the elimination reaction of 1-chloropropane with aqueous sodium hydroxide. The product is propene.

Structural formula of R:

$CH_3-CH=CH_2$

IUPAC name: Propene

Step 2: R + Na ------> S + NaCl

S is formed when propene reacts with sodium metal, leading to the formation of sodium prop-2-enide.

Structural formula of S:

$CH_2=CH-CH_2Na$

IUPAC name: Sodium prop-2-enide

Step 3: $S + CH_3I \text{----->} T$

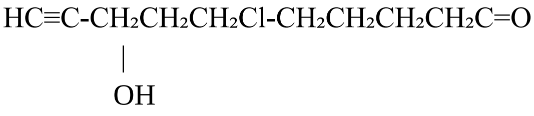
T is formed by the alkylation of sodium prop-2-enide with methyl iodide, producing but-2-ene.

Structural formula of T:

$CH_3\text{-}CH{=}CH\text{-}CH_3$

IUPAC name: But-2-ene

(b) You are provided with an organic compound L whose molecular structural formula is:



$$HC{\equiv}C\text{-}\underset{\substack{|\\ OH}}{CH_2}CH_2CH_2Cl\text{-}CH_2CH_2CH_2CH_2C{=}O$$

IUPAC names of the functional groups present in compound L:
- Alkyne ($-C{\equiv}C$)
- Hydroxyl ($-OH$)
- Carbonyl ($-C{=}O$)
- Alkyl halide ($-Cl$)

(c) Write down the chemical equation for the reaction which will take place when the following reagents are added to compound L

(i) Cl_2 / $FeCl_3$

This is an electrophilic substitution reaction where chlorine replaces a hydrogen atom.

Reaction:

$HC{\equiv}C\text{-}CH_2CH_2CH_2Cl\text{-}CH_2CH_2CH_2CH_2C{=}O+Cl_2\text{----->}HC{\equiv}C\text{-}CH_2CH_2CH_2Cl\ CH_2CH_2CH_2CH_2C{=}O\text{-}Cl$

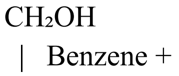
(ii) $NaOCH_3$ / 25°C

This reaction leads to nucleophilic substitution where $NaOCH_3$ replaces the chlorine atom.

Reaction:

$HC{\equiv}C\text{-}CH_2CH_2CH_2Cl\text{-}CH_2CH_2CH_2CH_2C{=}O + NaOCH_3 \text{------>} HC{\equiv}C\text{-}CH_2CH_2CH_2OCH_3\text{-}CH_2CH_2CH_2CH_2C{=}O$

(iii)



$$\overset{\substack{CH_2OH\\ |}}{\text{Benzene}}$$

CH_2OH
| Benzene + dil. H_2SO_4 ------> Oxidation of the $-CH_2OH$ group into an aldehyde ($-CHO$)

Reaction:

$C_6H_5\text{-}CH_2OH + [O] \text{----->} C_6H_5\text{-}CHO + H_2O$

(iv)

$CH_2{=}O$
|
Benzene + dil. H_2SO_4 -----> Oxidation of the -CHO group into a carboxyl (-COOH)

Reaction:

$C_6H_5\text{-}CHO + [O] \text{--->} C_6H_5\text{-}COOH$