# CHRISTIAN SOCIAL SERVICES COMMISSION (CSSC)
## NORTHERN ZONE JOINT EXAMINATIONS SYNDICATE (NZ-JES)

# FORM SIX PRE-NATIONAL EXAMINATIONS 2026
# 132/1 CHEMISTRY 1

# MARKING SCHEME

1. **(a) (i) Definition of Isoelectronic Species**

Isoelectronic species are atoms, ions, or molecules that have the same number of electrons. The species $K^+$, $Cl^-$, $Ca^{2+}$, and $S^{2-}$ all have 18 electrons, making them isoelectronic.

**(ii) Physical Significance of $n = \infty$ and $E = 0$ for Hydrogen Atom**

- $n = \infty$: This represents a state where the electron is no longer bound to the nucleus and is at an infinite distance from the nucleus, meaning the atom is ionized.
- $E = 0$: In the Bohr model, energy is measured relative to a free electron at rest infinitely far from the nucleus. When $E = 0$, the electron is completely free, meaning it has escaped the hydrogen atom, leading to ionization.

**(iii) Why Sigma Bonds Are Stronger Than Pi Bonds**

Sigma ( ) bonds are stronger than pi ( ) bonds because sigma bonds result from the head-on (axial) overlap of atomic orbitals, which allows for greater electron density between the nuclei. In contrast, pi bonds result from the side-by-side overlap of orbitals, which leads to weaker bonding due to less effective orbital overlap.

**(b) Why the Given Sets of Quantum Numbers Are Not Allowed**

(i) $= 1, \quad = 1, \quad = 0$

a. For $= 1$, the only allowed value of is 0. Since $= 1$ is given, this set is not valid.

(ii) $= 1, \quad = 0, \quad = 2$

a. The magnetic quantum number ($m_l$) ranges from $-l$ to $+l$, so for $l = 0$, the only possible value of $m_l$ is 0. Since $m_l = 2$ is given, this set is invalid.

(iii) $n = 2, l = -2, m_l = 1$

a. The azimuthal quantum number ($l$) must be in the range $0 \leq l \leq n - 1$. Since $l = -2$ is given (which is negative and invalid), this set is not allowed.

(iv) $= 0, \quad = 0, \quad = 0$

a. The principal quantum number ($n$) must be a positive integer ($n = 1,2,3, ...$), but $= 0$ is given, which is not physically meaningful. Therefore, this set is not allowed.

**(c) Proof that the Bohr Orbit Circumference is an Integral Multiple of the De Broglie Wavelength**

According to the Bohr model and De Broglie hypothesis, an electron in a stable orbit around the nucleus behaves as a standing wave. This condition requires that the circumference of the orbit be an integer multiple of the electron's De Broglie wavelength.

***Step-by-Step Proof:***

The De Broglie wavelength of an electron is given by:

$$= \frac{h}{\quad}$$

where $h$ is Planck's constant, $m$ is the mass of the electron, and $v$ is its velocity.

The Bohr quantization condition states that the angular momentum of the electron is quantized:

$$mvr = n\hbar$$

where $\hbar = \frac{h}{2}$ and    is the principal quantum number.

Solving for  ,

$$= \frac{n\hbar}{\quad}$$

Substituting into the De Broglie wavelength formula:

$$= \frac{h}{\quad} \times \frac{1}{\quad} = \frac{h}{m} \times \frac{mr}{n\hbar} = \frac{2}{\quad}$$

Rearranging, we get:

$$2 \quad =$$

This equation shows that the circumference of the Bohr orbit is an integral multiple of the De Broglie wavelength, proving the required result

2. **(a) Explanation of Given Phenomena**

(i) **Why is water ($H_2O$) a liquid while hydrogen sulfide ($H_2S$) is a gas?**

- Water has strong **hydrogen bonding** due to the highly electronegative oxygen atom. These intermolecular forces require significant energy to break, leading to a high boiling point (100°C).
- In contrast, hydrogen sulfide ($H_2S$) lacks strong hydrogen bonding since sulfur is less electronegative than oxygen. Instead, it experiences weaker dipole-dipole interactions and Van der Waals forces, resulting in a much lower boiling point (25°C), making it a gas at room temperature.

(ii) **Why does oxygen ($O_2$) have a lower boiling point than ozone ($O_3$)?**

- $O_2$ is a **nonpolar diatomic molecule**, meaning it relies only on weak Van der Waals forces (London dispersion forces) for intermolecular interactions.
- $O_3$, on the other hand, is a **polar molecule** with dipole-dipole interactions in addition to dispersion forces. The stronger intermolecular forces in ozone require more energy to break, leading to a higher boiling point compared to $O_2$.

(iii) **Why does molten Aluminum fluoride ($AlF_3$) conduct electricity, while Aluminum chloride ($AlCl_3$) does not?**

- **Molten $AlF_3$ is ionic**, meaning it dissociates into free-moving ions ($Al^{3+}$ and $F^-$) that conduct electricity.
- **$AlCl_3$, however, is covalent** in the molten state, meaning it does not break into free ions, making it a poor conductor of electricity.

(iv) **Why is sulfur dioxide ($SO_2$) a polar molecule?**

- $SO_2$ has a **bent molecular shape** due to the presence of a lone pair on the sulfur atom. This asymmetrical shape leads to an uneven distribution of charge, making the molecule polar.

**(b) Orbital Overlapping and Molecular Formation**

(i) **sp Orbital Overlapping (Example: Acetylene, $C_2H_2$)**

- In molecules like **$C_2H_2$**, each carbon undergoes **sp hybridization**, forming two sp orbitals. One sp orbital overlaps head-on with another sp orbital from a neighboring atom, forming a **sigma (σ) bond**.

(ii) **Two p Orbitals Overlapping (Example: Ethene, $C_2H_4$)**

- In molecules like **$C_2H_4$**, unhybridized p orbitals from adjacent carbon atoms overlap **sideways**, forming a **pi (π) bond**.

(iii) **Formation of Oxygen ($O_2$) Molecule**

- Oxygen molecules ($O_2$) form by the **side-by-side overlap of unpaired p orbitals**, creating both a sigma (σ) and a pi (π) bond (a double bond).

(iv) **Formation of Nitrogen ($N_2$) Molecule**

- In the nitrogen molecule ($N_2$), **three p orbitals** overlap, forming a **triple bond** (one sigma and two pi bonds), making it very strong and stable.

**(c) Hybridization and VSEPR Theory**

(i) **What is hybridization?**

Hybridization is the process of **mixing atomic orbitals** to form new hybrid orbitals with equivalent energy, which helps explain molecular shapes and bonding.

(ii) **Three Assumptions of VSEPR Theory:**

(i) **Electron pairs repel each other** and arrange themselves as far apart as possible.
(ii) **Lone pairs take up more space than bonding pairs**, affecting molecular shapes.
(iii) **The shape of a molecule depends on the number of bonding and lone electron pairs** around the central atom.

**3. (a) (i) Definition of Effusion**

Effusion is the process by which gas particles pass through a small hole or narrow opening **without collisions** between the molecules and **without significant interaction** with the container walls. This process is described by **Graham's Law of Effusion**, which states that the rate of effusion is inversely proportional to the square root of the molar mass of the gas.

**(ii) Ratio of the Rate of Diffusion of Hydrogen to Oxygen**

According to **Graham's Law of Diffusion**, the rate of diffusion of gases is inversely proportional to the square root of their densities:

$$\frac{r_1}{2} = \sqrt{\frac{d_2}{1}}$$

where:

- $r_1$ and $r_2$ are the diffusion rates of $H_2$ and $O_2$, respectively.
- $d_1$ and $d_2$ are the densities of $H_2$ and $O_2$, respectively.

Substituting the given values:

$$\frac{r_{H_2}}{O_2} = \sqrt{\frac{1.43}{0.0899}}$$

$$= \sqrt{15.91}$$

$$\approx 3.99$$

Thus, **hydrogen diffuses about 4 times faster than oxygen**.

**(b) Volume of Ozone ($O_3$) Produced**

The reaction for ozone formation from oxygen is:

$$3O_2 \rightarrow 2O_3$$

Since the reaction occurs at **constant temperature and pressure**, the volumes of gases are proportional to their moles.

Given:

- **Volume of $O_2$** = 15.5 L
- **Moles of $O_2$** = 0.75 moles

Using the molar volume ratio from the equation:

$$\frac{\text{Volume of }_3}{\text{Volume of }_2} = \frac{2}{3}$$

$$V_{O_3} = \frac{2}{3} \times 15.5$$

$$= 10.33\ \text{L}$$

Thus, the volume of **ozone ($O_3$) produced is 10.33 L**.

**(c) Gas Density Calculations**

Given:

- **Volume** = 1000 cm³ = 1 L
- **Mass of gas** = 1.19 g
- **Temperature** = 20°C = **293 K**
- **Pressure** = **1 atm**

*(i) Relative Density*

Relative density is the **ratio of the gas density to the density of hydrogen**.

Using **density formula**:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{1.19}{1}$$

$$= 1.19 \text{ g/L}$$

Density of **$H_2$ at STP** = 0.0899 g/L

$$\text{Relative Density} = \frac{1.19}{0.0899} \approx 13.24$$

Thus, the **relative density is 13.24**.

*(ii) Normal Density*

Density at standard conditions is given by the **ideal gas law**:

$$=$$

Using **molar mass formula**:

$$\text{Density} = \frac{}{}$$

Where:

- **P** = 1 atm
- **M** = ? (Molar mass of gas)
- **R** = 0.0821 L·atm/mol·K
- **T** = 293 K

Rearrange for :

$$= \frac{}{}$$

$$= \frac{(1.19)(0.0821)(293)}{1}$$

$$= 28.62 \text{ g/mol}$$

Thus, the **normal density is 1.19 g/L**.

**4. (a) (i) Effect of Degree of Dissociation on Boiling Point**

Boiling point elevation is a colligative property, meaning it depends on the **number of solute particles** in solution rather than their identity. When a solute dissolves and **dissociates** into multiple ions, the number of dissolved particles increases, leading to a greater boiling point elevation. The relationship is given by:

$$=$$

where:

- $\Delta T_b$ = Boiling point elevation
- $i$ = Van't Hoff factor (number of particles after dissociation)
- $K_b$ = Boiling point elevation constant
- $m$ = Molality of the solution

A **higher degree of dissociation ( ) increases** , leading to a **higher boiling point**.

**(ii) Derivation of the Van't Hoff Factor ( ) in Terms of Degree of Dissociation ( )**

When an electrolyte dissolves, it **partially or completely dissociates** into ions. Suppose one mole of a solute dissociates as:

$$AB \rightleftharpoons A^{+} + B^{-}$$

Let be the **degree of dissociation** (fraction of molecules that dissociate). If we start with **1 mole** of AB:

- **Before dissociation:** 1 mole of AB, 0 moles of $A^{+}$ and $B^{-}$.
- **After dissociation:** $1-\alpha$ moles of AB remain, and $\alpha$ moles each of $A^{+}$ and $B^{-}$ are formed.

Total moles after dissociation:

$$(1-\alpha)+\alpha+\alpha=1+\alpha$$

For a general electrolyte that dissociates into ions:

$$i=1+\alpha(n-1)$$
$$=\frac{i-1}{n-1}$$

where:

- $i$ = Van’t Hoff factor
- $\alpha$ = Degree of dissociation
- $n$ = Total number of ions per formula unit

**(b) Calculation of the Degree of Dissociation of NaCl**

**Given data:**

- **Freezing point of solution** = -0.604°C
- **Freezing point depression constant ($K_f$)** = 1.86°C kg/mol
- **Mass percent of NaCl** = 1% (1 g of NaCl in 100 g of water)
- **Molar mass of NaCl** = 58.5 g/mol

***Step 1: Calculate Molality***

Molality ( ) is given by:

$$=\frac{\text{moles of solute}}{\text{kg of solvent}}$$

Moles of NaCl:

$$=\frac{1}{58.5}=0.0171\text{ moles}$$

Mass of solvent in kg:

$$=\frac{100}{1000}=0.1\text{ kg}$$

Molality:

$$= \frac{0.0171}{0.1} = 0.171 \text{ mol/kg}$$

***Step 2: Use Freezing Point Depression Formula***

$$= \quad 0.604$$

$$= \ (1.86)(0.171)$$

$$i = \frac{0.604}{(1.86 \times 0.171)}$$

$$= \frac{0.604}{0.318} = 1.9$$

***Step 3: Calculate  Using the Van't Hoff Equation***

For NaCl (  = 2):

$$i = 1 + \alpha(2 - 1) = 1 + \alpha$$

$$1.9 = 1 +$$

$$= 0.9$$

Thus, the **degree of dissociation of NaCl is 0.9 (or 90%)**.

**(c) Applications of Colligative Properties in Real Life**

(i) **Antifreeze in Vehicles**

- Ethylene glycol is added to radiator water to **lower the freezing point**, preventing the engine from freezing in cold weather.

(ii) **Boiling Point Elevation in Cooking**

- Adding salt to water **raises its boiling point**, allowing food to cook at a slightly higher temperature.

(iii) **IV Fluids in Medicine**

- IV solutions are prepared to **match the osmotic pressure of blood**, preventing cell damage from excessive water movement.

(iv) **Food Preservation**

- Adding salt or sugar to food creates a **hypertonic solution**, drawing water out of bacteria and preventing spoilage.

(v) **Desalination of Water**

- Reverse osmosis uses osmotic pressure to **remove salt from seawater**, making it drinkable.

**5. (a) (i) Definition of Ionization Energy**

**Ionization energy** is the amount of energy required to remove the **outermost electron** from a neutral atom in its gaseous state, forming a positively charged ion.

$$\text{X(g)} \rightarrow \text{X}^{+}(g) + e^{-}$$

- **First Ionization Energy**: Energy required to remove the first electron.
- **Second Ionization Energy**: Energy needed to remove a second electron, which is always higher than the first.

**(ii) Four Applications of Hess's Law**

**Hess's Law** states that the **total enthalpy change** of a reaction is independent of the reaction pathway, depending only on the initial and final states.

**Applications**

(i) **Determining Enthalpy of Reactions**

- Used to calculate the enthalpy of reactions that are difficult to measure directly, such as **combustion and formation reactions**.

(ii) **Calculation of Bond Enthalpies**

- Helps determine **bond dissociation energies** by breaking complex reactions into simpler steps.

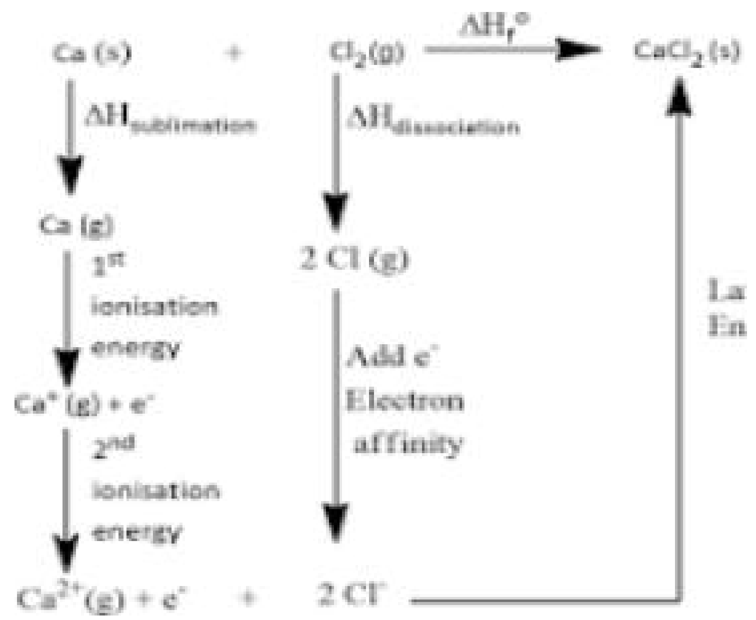
(iii) **Determination of Lattice Energy**

- Used in the **Born-Haber cycle** to calculate the lattice energy of ionic compounds.

(iv) **Energy Changes in Industrial Processes**

- Used in industries like **Haber's process for ammonia synthesis** and **metallurgy** to optimize energy usage.

**(b) Born-Haber Cycle for $CaCl_2$ and Lattice Energy Calculation**




**Given Data**

- **Sublimation energy of Ca = +193 kJ/mol**
- **First ionization energy of Ca = +590 kJ/mol**
- **Second ionization energy of Ca = +1145 kJ/mol**
- **Bond dissociation energy of $Cl_2$ = +242 kJ/mol**
- **Electron affinity of Cl = -348 kJ/mol**
- **Standard enthalpy of formation of $CaCl_2$ = -763 kJ/mol**

**Step 1: Write the Born-Haber Cycle**

- **Sublimation of Ca:**

- Ca (s) → Ca (g), $\Delta H = +193$ kJ/mol
- **First Ionization of Ca:**
- Ca (g) → $Ca^{+}(g) + e^{-}$, $\Delta H = +590$ kJ/mol
- **Second Ionization of Ca:**
- $Ca^{+}(g) \rightarrow Ca^{2+}(g) + e^{-}$, $\Delta H = +1145$ kJ/mol
- **Dissociation of $Cl_2$:**
- $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$, $\frac{+242}{2} = +121$ kJ/mol
- **Electron Affinity of Cl (2 atoms):**
- $2 \times (-348) = -696$ kJ/mol
- **Formation of $CaCl_2$:**

$$Ca^{2+} + 2Cl^{-} \rightarrow CaCl_2 \quad \Delta H = -763 \text{ kJ/mol}$$

Using Hess's Law:

$$_{f} = \; _{sub} + \; _{ion\ 1} + \; _{ion\ 2} + \; _{diss} + \; _{EA} +$$

$$-763 = (193 + 590 + 1145 + 121 - 696 + U)$$

$$U = -2250 \text{ kJ/mol}$$

Thus, the **lattice energy of $CaCl_2$ is -2250 kJ/mol**.

**(c) Heat of Formation of Hypothetical CaCl**

Given lattice energy of **CaCl** = **-155 kJ/mol**, we apply Hess's Law:

$$_{f} = \; _{sub} + \; _{ion\ 1} + \; _{diss} + \; _{EA} +$$

$$\Delta H_f = (193 + 590 + 121 - 348 - 155)$$

$$_{f} = 401 \text{ kJ/mol}$$

Thus, the **heat of formation of CaCl is +401 kJ/mol**, indicating an **unstable compound**.

**(d) Stability of $CaCl_2$ vs. CaCl**

$CaCl_2$ is **more stable** than CaCl because:

- The **lattice energy of $CaCl_2$ (-2250 kJ/mol) is much more negative than that of CaCl (-155 kJ/mol)**, indicating a more **stable structure**.
- The formation of **$Ca^{2+}$ in $CaCl_2$** achieves a **more stable noble gas configuration**, while CaCl with **$Ca^{+}$ is less stable**.

6. **(a) Preparation of Zinc Oxide (ZnO)**

(i) **By Heating Zinc Carbonate**

$$ZnCO_3 \rightarrow ZnO + CO_2$$

(ii) **By Burning Zinc in Air**

$$2Zn + O_2 \rightarrow 2ZnO$$

**6. (b) Chemical Equations and Observations**

(i) **Sodium Carbonate Heated in Air**

$$Na_2CO_3 \xrightarrow{heat} Na_2O + CO_2$$

**Observation:** White solid forms, and $CO_2$ gas is released.

(ii) **Aluminium Hydroxide Amphoteric Nature**

- Reaction with Acid:

$$Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$$

- Reaction with Base:

$$Al(OH)_3 + NaOH \rightarrow NaAlO_2 + 2H_2O$$

**Observation:** Dissolves in both acids and bases.

(iii) **Magnesium Nitride Dissolved in Water**

$$Mg_3N_2 + 6H_2O \rightarrow 3Mg(OH)_2 + 2NH_3$$

**Observation:** Ammonia gas ($NH_3$) is released, producing a pungent smell.

(iv) **Concentrated HCl Added to $PbCl_2$**

$$PbCl_2 + 2HCl \rightarrow PbCl_4^{2-}$$

**Observation:** White $PbCl_2$ precipitate dissolves, forming a clear solution.

**(c) Uses of Metal Carbonates in Daily Life**

(i) **Limestone ($CaCO_3$) in Cement & Construction**

- Used in making cement, concrete, and building materials.

(ii) **Baking Soda ($NaHCO_3$) in Cooking**

- Acts as a **leavening agent** to make bread rise.

(iii) **Antacids ($MgCO_3$, $CaCO_3$) for Indigestion**

- Neutralizes stomach acid to relieve heartburn.

(iv) **Sodium Carbonate in Water Softening**

- Helps remove **hardness-causing ions** in water.

7. **(a) Expressions for Equilibrium Constants**

The equilibrium constant expression depends on the phases of the substances involved. **Solids and pure liquids do not appear in the expression**.

(i) **For**

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

Since only $CO_2$ is a gas, we use **Kp**:

$$= \quad {}_2$$

(ii) **For**

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$

Since AgCl is a solid and the other substances are aqueous, we use **Kc**:

$$= \frac{[Ag^+][Cl^-]}{}$$

(iii) **For**

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

Since all are gases, we use **Kp or Kc**:

$$= \frac{\ ^{2}}{\ _{2}\ \ _{2}}$$

OR

$$= \frac{[\ \ ]^{2}}{[\ \ _{2}][\ \ _{2}]}$$

**(b) Calculation of Kp for $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$**

***Step 1: Define the Initial and Equilibrium Moles***

The reaction is:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

**Given data:**

- **Initial moles:** $SO_2$ = 10.0 mol, $O_2$ = 5.0 mol
- **Conversion of $SO_2$** = 90% → Reacted = **0.9 × 10.0 = 9.0 mol**

| Species | Initial (mol) | Change (mol) | Equilibrium (mol) |
|---|---|---|---|
| $SO_2$ | 10.0 | -9.0 | 1.0 |
| $O_2$ | 5.0 | -4.5 | 0.5 |
| $SO_3$ | 0.0 | +9.0 | 9.0 |

***Step 2: Calculate Mole Fractions***

Total moles at equilibrium:

$$1.0 + 0.5 + 9.0 = 10.5$$

Mole fractions:

$$_2 = \frac{1.0}{10.5} = 0.0952, \qquad _2 = \frac{0.5}{10.5} = 0.0476, \qquad _3 = \frac{9.0}{10.5} = 0.857$$

***Step 3: Calculate Partial Pressures***

Total pressure = 200 kPa.

$$P_i = X_i \times P_{\text{total}}$$

$$P_{SO_2} = 0.0952 \times 200 = 19.05 \text{ kPa}$$

$$P_{O_2} = 0.0476 \times 200 = 9.52 \text{ kPa}$$

$$P_{SO_3} = 0.857 \times 200 = 171.43 \text{ kPa}$$

***Step 4: Calculate Kp***

$$= \frac{\ ^{2}\ _{3}}{\ ^{2}\ _{2}\ \ _{2}}$$

$$K_p = \frac{(171.43)^2}{(19.05)^2 \times 9.52}$$

$$= \frac{29388.2}{3447.5} = 8.52$$

Thus, **Kp = 8.52**.

**(c) Effect of Changes on Kp for $CH_4 + 3H_2O \rightleftharpoons CO_2 + 3H_2$**

$$CH_4(g) + 3H_2O(g) \rightleftharpoons CO_2(g) + 3H_2(g)$$

***(i) Increasing Pressure***

- The reaction involves **4 gas molecules (reactants) ⇌ 4 gas molecules (products)**.
- Since there is **no net change in the number of moles of gas**, **Kp remains unchanged**.

***(ii) Increasing Temperature***

- The reaction is **endothermic** (absorbs heat).
- **Le Chatelier's Principle**: Increasing temperature **shifts equilibrium to the right**, increasing product concentration.
- **Kp increases**.

***(iii) Using a Catalyst***

- A catalyst **speeds up the reaction rate** but does **not affect the equilibrium position**.
- **Kp remains unchanged**.

**8. (a) Structural Formulas of Given IUPAC Names**

(i) **3,4-Diethyl-2,2-dimethylheptane**

- **Base chain:** Heptane (7 carbon atoms)
- **Substituents:**
  - 2,2-Dimethyl (two methyl groups at carbon 2)
  - 3,4-Diethyl (ethyl groups at carbons 3 and 4)

**Structural formula:**

CH3

|

CH3-C-CH-C2H5 | | CH3 C2H5 | CH2-CH2-CH3

(ii) 2-Methylpent-2-ene

- Base chain: Pent-2-ene (five carbon atoms with a double bond at C2)

- Substituent: Methyl group at C2

**Structural formula:**

CH3

|

CH3-C=CH-CH2-CH3

(iii) Pent-1-ene-3-yne

Base chain: Pentane (five carbon atoms)

- Functional groups:

- Alkene (-ene) at C1

- Alkyne (-yne) at C3

Structural formula:

CH2=CH-C≡C-CH3

(b) Observations from Organic Experiments

(i) 1-Bromopropane heated with alcoholic caustic potash (KOH in ethanol):

**Observation**: Formation of a gas that decolorizes bromine water.

**Explanation**: This is an elimination reaction, forming propene ($CH_3$-CH=$CH_2$) and HBr.

(ii) Ethene treated with cold, alkaline potassium permanganate ($KMnO_4$):

**Observation**: The purple color of $KMnO_4$ disappears, and a brown precipitate ($MnO_2$) may form.

**Explanation**: Ethene undergoes oxidation to form ethylene glycol (HO-$CH_2$-$CH_2$-OH).

(iii) Addition of bromine to hexane under sunlight:

Observation: The red-brown color of bromine slowly fades.

Explanation: A substitution reaction occurs, forming bromohexane and HBr in the presence of UV light.

(c) Organic Reactions and Predictions

(i) Ozonolysis Experiment Producing Ethanal ($CH_3CHO$):

- The two possible structures for the original compound are:
- But-2-ene ($CH_3CH=CHCH_3$)
- 2,3-Dimethylbut-2-ene ($(CH_3)_2C=CH(CH_3)_2$)
- Explanation: Both structures cleave at the double bond to form only ethanal.

(ii) Addition of HCl to 2-Methylbut-2-ene:

- Reaction follows Markovnikov's Rule (H attaches to the carbon with more H's).
- Product: 2-Chloro-2-methylbutane ($CH_3C(Cl)(CH_3)$-$CH_2$-$CH_3$).

Grignard Reagent Preparation in Anhydrous Conditions:

**Reason**: Grignard reagents (R-MgX) react violently with water:

$R\text{-}MgX + H_2O \rightarrow RH + Mg(OH)X$

**Effect**: The reagent is destroyed, making it useless for organic synthesis.

(d) Distinguishing Between Pairs of Compounds

(i) $CH_3CH_2C{\equiv}CCH_3$ vs. $CH_3CH_2CH_2C{\equiv}CH$

-Test: Silver nitrate in ammonia solution (Tollen's Test).

- Observation: $CH_3CH_2CH_2C{\equiv}CH$ (terminal alkyne) forms a white precipitate , while $CH_3CH_2C{\equiv}CCH_3$ does not.

(ii) $CH_2=CH_2$ vs. $CH_3CH_3$ (Ethene vs. Ethane)

- Test: Bromine water.
- Observation: Ethene decolorizes bromine rapidly (addition reaction), whereas ethane has no reaction .

(iii) $CH_3CH_2Cl$ vs. $CH_3CH_3$ (Ethyl chloride vs. Ethane)
- Test: Silver nitrate ($AgNO_3$) in ethanol.
- Observation:Ethyl chloride forms a white precipitate (AgCl), while ethane shows no reaction .

9. **(a) Explanation of Terms**

(i) **Farm Yard Manure (FYM)**

A mixture of **animal dung, urine, and crop residues** from farm animals like cows, goats, and sheep.

It improves soil **structure, water retention, and microbial activity**.

(ii) **Organic Fertilizers**

Fertilizers derived from **natural sources** such as plant residues, compost, and manure.

They release nutrients **slowly** and **enhance soil fertility** in the long run.

(iii) **Artificial Fertilizers**

Also called **chemical or synthetic fertilizers**, they are industrially manufactured and provide specific nutrients like **NPK (Nitrogen, Phosphorus, Potassium)**.

They provide **rapid nutrient supply** but may **degrade soil quality** if overused.

(iv) **Compost Manure**

- Made from **decomposed organic waste** (vegetable peels, grass, leaves, etc.).
- It **enriches the soil** with essential nutrients and **improves soil structure**.

**(b) Calculation of Percentage Saturation and $H^+$ Ion Concentration**

**Given data:**

- **Mass of soil** = 5 g
- **Volume of NaOH used** = 10 cm³ = 0.01 L
- **Concentration of NaOH** = 0.1 M
- **Cation exchange capacity (CEC)** = 25 meq/100 g soil

***Step 1: Calculate the exchangeable $H^+$ in meq per 5 g soil***

From the neutralization reaction:

$$\text{moles of NaOH} = C \times V = (0.1 \times 0.01) = 0.001 \text{ moles}$$

Since NaOH neutralizes $H^+$ ions in a **1:1 ratio**,

$$\text{milliequivalents of H}^+ \text{ in 5g soil} = 0.001 \times 1000 = 1.0 \text{ meq}$$

***Step 2: Calculate Percentage Saturation***

$$\text{Percentage Saturation} = \left(\frac{\text{Exchangeable H}^+ \text{ (meq/100g)}}{\text{Total CEC (meq/100g)}}\right) \times 100$$

$$= (\frac{(1.0 \times 20)}{25}) \times 100 = (\frac{20}{25}) \times 100 = 80\%$$

Thus, **Percentage Saturation = 80%**.

***Step 3: Calculate Concentration of $H^+$ in 75g of Soil***

$$H^+ \text{ concentration in 75 g} = (\frac{1.0}{5}) \times 75 = 15 \text{ meq}$$

**(c) Calculation of Calcium Concentration in meq/100g**

**Given data:**

- **Mass of soil** = 20 g
- **Mass of calcium** = 0.015 g
- **Atomic weight of Ca** = 40 g/mol
- **Charge of $Ca^{2+}$** = 2

***Step 1: Convert Mass of Ca to Moles***

$$\text{Moles of Ca} = \frac{0.015}{40} = 0.000375 \text{ moles}$$

***Step 2: Convert to Milliequivalents***

$$\text{Milliequivalents of } Ca^{2+} = 0.000375 \times 1000 \times 2 = 0.75 \text{ meq}$$

***Step 3: Convert to meq per 100g Soil***

Since the given mass is **20 g**, we scale it to **100 g**:

$$\text{Calcium concentration} = (\frac{0.75}{20}) \times 100 = 3.75 \text{ meq/100g}$$

Thus, **Calcium concentration = 3.75 meq/100g**.

**(d) Soil Colloids and Their Role in Soil Science**

***(i) What are Soil Colloids?***

- Soil colloids are **microscopic particles** (clay and humus) that are **negatively charged** and can **retain and exchange cations (nutrients).**

***(ii) How Soil Colloids Contribute to CEC***

- Soil colloids **attract cations ($Ca^{2+}$, $Mg^{2+}$, $K^+$, $Na^+$, $H^+$, etc.)** due to their **negative charge**.
- The **higher the colloid content**, the **higher the CEC**, improving nutrient-holding capacity.

***(iii) Role of Soil Colloids in Nutrient Retention and Availability***

- **Prevent Nutrient Leaching:**

  Hold essential nutrients and **release them gradually**.

- **Improve Soil Fertility:**
  - Store nutrients like **$K^+$, $Mg^{2+}$, $Ca^{2+}$**, making them available to plants.
- **Regulate Soil pH:**

  Buffer against **acidic or alkaline conditions** by **absorbing or releasing $H^+$ and $OH^-$**.

10. **(a) Suggested Structures of the Compounds**

(i) **2-chloro-3-methylpentane**

- The structure of this compound is as follows:
  - ✓ A **pentane** chain (5 carbon atoms).
  - ✓ A **methyl group** ($-CH_3$) at position 3.
  - ✓ A **chloro group** (-Cl) at position 2.
- The structure is:

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

(ii) **4-tertbutyl-3-iodoheptane**

- The structure of this compound is as follows:
  - ✓ A **heptane** chain (7 carbon atoms).
  - ✓ A **tert-butyl group** ($-C(CH_3)_3$) at position 4.
  - ✓ An **iodo group** (-I) at position 3.
- The structure is:

  $CH_3-CH_2-CH(I)-C(C_6H_{19})-CH_2-CH_3$

(iii) **1,4-dibromobut-2-ene**

- The structure of this compound is as follows:
  - ✓ A **butene** chain (4 carbon atoms with a double bond at position 2).
  - ✓ A **bromo group** (-Br) at positions 1 and 4.
- The structure is:

  $Br-CH_2-CH=CH-CH_2-Br$

(iv) **1-bromo-4-sec-butyl-2-methylbenzene**

- The structure of this compound is as follows:
  - ✓ A **benzene ring** ($C_6H_6$) with:
    1. A **bromo group** (-Br) at position 1.
    2. A **methyl group** ($-CH_3$) at position 2.
    3. A **sec-butyl group** ($-C_4H_9$) at position 4.
- The structure is:

  $Br-C_6H_4-CH_2CH(CH_3)_2$ (with the sec-butyl group at position 4).

**(b) Structure and Reactions of Compound X ($C_5H_6$)**

- **Given Data:**

  Compound X takes up **two moles of chlorine**.

  Compound X **decolorizes $Br_2$** in $CCl_4$.

  Compound X **forms no precipitates with ammoniacal silver nitrate**.

- **Analysis:**

- The molecular formula **$C_5H_6$** suggests an **unsaturated compound**, likely containing a **double bond** or a **triple bond**.
- The fact that **two moles of chlorine** are taken up suggests the presence of a **diene** or a **compound with conjugated double bonds**.
- The **decolorization of $Br_2$** indicates the presence of **C=C** bonds (typical of alkenes or dienes).
- The absence of a precipitate with **ammoniacal silver nitrate** suggests the compound is **not an alkyne** (as alkynes typically form precipitates with this reagent due to their ability to undergo addition reactions with silver salts).

- **Possible Structure of X:**
  - The structure that fits this data is **1,3-pentadiene** (a diene), which can react with chlorine in a **1,4-addition** to form a dichloride, decolorizing the bromine solution.

**Reaction 1: Addition of Chlorine:**

$$CH_2 = CH - CH = CH - CH_3 + Cl_2 \rightarrow CH_2Cl - CHCl - CH = CH - CH_3$$

**Reaction 2: Decolorization of $Br_2$:**

$$CH_2 = CH - CH = CH - CH_3 + Br_2 \rightarrow \text{Decolorization of } Br_2$$

The compound reacts with **$Br_2$** due to the presence of conjugated double bonds.

**Reaction 3: No reaction with Ammoniacal Silver Nitrate:**

$$CH_2 = CH - CH = CH - CH_3 + AgNO_3 \text{ (no precipitate)}$$

**No reaction** indicates it's **not an alkyne**.

Thus, the structure of **X** is **1,3-pentadiene**.

## 10. (c) Isomeric Compounds from 1,2-Dibromodecane

**Given Data:**

- **1,2-dibromodecane** undergoes reaction with **potassium hydroxide** in **aqueous ethanol**, yielding two isomeric compounds of molecular formula **$C_{10}H_{19}Br$**.
- These isomers are then converted to **dec-1-yne** upon reaction with **sodium amide**.

- **Step 1: Reaction of 1,2-Dibromodecane with KOH (Ethanol)**
  - When **1,2-dibromodecane** reacts with **KOH** in ethanol, a **dehydrohalogenation** reaction occurs, forming **alkenes**.
  - The two possible elimination products are:
    - **1-bromodec-1-ene** (from the elimination of one HBr molecule).
    - **dec-2-ene** (from the elimination of the other HBr molecule).
- **Step 2: Conversion of Alkenes to Alkynes**
  - Both of these alkenes undergo **double dehydrohalogenation** when treated with **sodium amide ($NaNH_2$)**.
  - The alkenes undergo **elimination** of HBr in two steps to form the alkyne **dec-1-yne**.

**Reactions Involved:**

- **1,2-dibromodecane** with KOH in ethanol:

$$CH_3(CH_2)_6CHBr\text{-}CH_2Br \xrightarrow{KOH} \text{dec-1-ene (or dec-2-ene)}$$

- Conversion to **dec-1-yne** using sodium amide:

$$\text{dec-1-ene (or dec-2-ene)} \xrightarrow{NaNH_2} \text{dec-1-yne}$$

- **Identifying the Isomers:**
    - The two possible isomeric compounds are:
        - **dec-1-ene** (from elimination at the 1-position of the 1,2-dibromodecane).
        - **dec-2-ene** (from elimination at the 2-position of the 1,2-dibromodecane).

Both are converted to **dec-1-yne**.