DEPARTMENT OF CHEMISTRY

Estd 1864

FORMAN CHRISTIAN COLLEGE
LAHORE

SOLUTION OF QUICK CHECKS AND EXERCISES

F.Sc. CHEMISTRY PART 1 (2025-26)

CHAPTER #01

PERIODIC TABLE AND PERIODIC PROPERTIES

Quick Check 1.1

a) Why are the elements in Groups 1 and 2 known as s-block elements?

The elements in Groups 1 and 2 have their valence electrons in the s subshell. For example, Group 1 elements have one valence electron in the ns^1 orbital, and Group 2 elements have two valence electrons in the ns^2 orbital. Hence, they are called s-block elements.

b) Name the elements in the chalcogen family. Give their two characteristics.

Elements: Oxygen (O), Sulfur (S), Selenium (Se), Tellurium (Te), Polonium (Po).

Characteristics:

- **i.** They typically form compounds with metals e.g., oxides and sulfides.
- **ii.** They shows variable oxidation states e.g., -2, +4, +6.

Quick Check 1.2

a) X belongs to group 14 and period 2.

i) Write electronic configuration of the element X.

Electronic configuration: $1s^2\ 2s^2\ 2p^2$.

ii. Identify block of the element. Identify this element from the periodic table.

The valence electrons are in the p subshell ($2p^2$), so it is a p-block element.

The element is Carbon (C).

b) Identify an element that is in Period 4 and Group 17?

The element is Bromine (Br).

Quick Check 1.3

a) Which factors affect atomic and ionic radii?

Nuclear charge, shielding effect, number of electronic shells effects atomic radii.

b) Predict and explain relative sizes:

- **i.** Li > F (Li has fewer protons and a larger electron cloud).
- **ii.** Li > Li^+ (Li^+ loses an electron, reducing repulsion and size).
- **iii.** O^{2-} > O (O^{2-} gains electrons, increasing repulsion and size).
- **iv.** N^{3-} > F^- (N^{3-} has more electrons and less nuclear charge than F^-).

Quick Check 1.4

a) Explain: i. Boron (B) has lower 1st IE than Be:

Be has a stable $2s^2$ configuration, while B has one electron in 2p, which is easier to remove, so it has lower first ionization energy than Be.

ii. Al has lower IE than Mg:

Mg has a stable $3s^2$ configuration, while Al's $3p^1$ electron is easier to remove, so it has lower first ionization energy than Mg.

b) Trend in Group 3 IE:

The ionization energy of group 3 elements decreases down the group due to increases in atomic size (shielding reduces effective nuclear charge).

Quick Check 1.5

a) 1st EA of O is -141 kJ/mol; 2nd EA is +844 kJ/mol:

Adding the first electron releases energy (stable half-filled p^4). Adding a second electron to O^- requires energy due to repulsion.

b) P has higher EA than N:

N's half-filled $2p^3$ is stable, making it harder to add an electron.

c) F < Cl despite smaller size:

Fluorine has a compact electron cloud, which causes repulsion, reducing electron affinity.

d) Why noble gases have positive 1st electron affinity?

Noble gases have positive 1st electron affinities because their stable and completely filled electronic configurations do not favor the addition of another electron, making the process energetically unfavorable.

Quick Check 1.6

a) Illustrate how does the metallic character vary in group 14.

Metallic character increases down in Group 14 like Carbon (C) is a nonmetal, Silicon (Si) and Germanium (Ge) are metalloids, Tin (Sn) and Lead (Pb) are metals.

b) Identify semimetals in groups 14, 15, and 16. Why are they semimetals?

Group 14: Silicon (Si), Germanium (Ge)

Group 15: Arsenic (As), Antimony (Sb)

Group 16: Tellurium (Te)

Reason: They exhibit properties intermediate between metals and nonmetals (e.g., semi conductivity).

Quick Check 1.7

a) What is the nature of oxides and hydroxides of Na and Mg?

Na_2O and NaOH are basic in nature because they form alkalis in water. MgO and $Mg(OH)_2$ are also basic in nature and $Mg(OH)_2$ is sparingly soluble.

b) Predict the reactivity of Ca (Group 2) with water and oxygen.

Calcium reacts with water to formed $Ca(OH)_2$ and H_2. On the other hand, Ca reacts with oxygen and forms CaO (calcium oxide) with a bright flame.

Quick Check 1.8

a) ZnO reactions:

With HCl: $ZnO+2HCl \longrightarrow ZnCl_2+H_2O$

With NaOH: $ZnO+2NaOH \longrightarrow Na_2ZnO_2+H_2O$

Type: Amphoteric oxide.

b) $AlCl_3$ is acidic; NaCl is not: \

Al^{3+} hydrolyzes water to form H^+, while Na^+ does not.

c) PCl_5 (acidic), NCl_3 (neutral):

PCl_5 hydrolyzes to HCl while NCl_3 does not react with water.

d) SO_2 and P_2O_5 react with NaOH (basic), not HCl/H_2SO_4 (acidic).

Quick Check 1.9

a) Oxidation number of S in SO_2 is +4 SO_3 is +6.

b) p-block elements show variable oxidation states due to vacant d-orbitals e.g., P (+3, +5), S (+4, +6).

Short Answer Questions

a. What is 1st ionization energy? Give an example.

The energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous +1 ions.

Example: $Na_{(g)} \longrightarrow Na_{(g)} + e^-$ $\Delta H_{i1}=494$ kJ/mol

b. Why does sulfur have a lower first ionization energy than phosphorus?

Phosphorus has a half-filled 3p subshell (extra stable), making electron removal harder. Sulfur's paired 3p electrons experience repulsion, reducing ionization energy.

c. Why are Group 13–17 elements called p-block elements?

Their valence electrons occupy the p subshell. For example, Group 13 has ns^2, np^1 and Group 17 has ns^2, np^5.

d. Factors affecting electronegativity:

i. **Atomic size:** Smaller atoms have higher electronegativity.
ii. **Nuclear charge:** Higher charge increases electronegativity.

e. Why do alkali metals become more reactive down the group?

Atomic size increases, reducing nuclear attraction on the single valence electron, making it easier to lose so metallic increases down the group.

f. Why do some elements show variable oxidation states?

P-block elements (e.g., P, S) can expand their octet by utilizing empty d-orbitals, enabling multiple oxidation states (e.g., +3/+5 for P).

Identify the element in Period 5, Group 15.

Antimony (Sb).

h. Why are Na_2O/MgO more ionic than N_2O_5/P_2O_5?

Na/Mg (metals) lose electrons easily, forming ionic bonds with oxygen. N/P (nonmetals) share electrons covalently with oxygen.

i. Why are Na and Mg differently reactive with O_2/Cl_2?

Na (Group 1) has 1 valence electron, losing it more readily than Mg (Group 2), which requires more energy to lose 2 electrons.

j. Why is Li's ionization energy lower than He's?

He has a fully filled $1s^2$ shell (high stability), while Li's outer electron is in 2s, farther from the nucleus and shielded by inner electrons.

k. Why is Be's ionization energy higher than B's?

Be has a stable $2s^2$ configuration, whereas B's $2p^1$ electron is easier to remove due to higher energy and less shielding.

l. Commonality in Na^+, Mg^{2+}, Al^{3+}, Ne, F^-? Arrange by size.

All have 10 electrons (isoelectronic). Size order:

$Al^{3+} < Mg^{2+} < Na^+ < F^- <$ Ne (Smaller positive charge = larger size due to reduced nuclear pull.)

m. Classify NaCl, $MgCl_2$, PCl_5:

- **NaCl:** Neutral (pH 7; ionic, no hydrolysis).
- **$MgCl_2$:** Slightly acidic (pH ~6.5; weak Mg^{2+} hydrolysis).
- **PCl_5:** Acidic (hydrolyzes to HCl and H_3PO_4).

Descriptive Questions

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Q.4 Page (16,17)

Q.5 Page (09,10)

Q.6 Page (11,12)

CHAPTER # 02

ATOMIC STRUCTURE

Quick Check 2.1

a) Calculate the number of neutrons:

Use the formula: **Neutrons = Mass number (A) – Atomic number (Z)**

- $^{39}_{19}K$ = 39 – 19 = **20**
- $^{35}_{17}Cl$ =35 – 17 = **18**
- $^{40}_{18}Ar$ =40 – 18 = **22**

b) Which of electron, proton and neutron is deflected the most in the magnetic field?

Electron and proton are charged particles, so they get deflected in a magnetic field. Neutron is neutral (no charge), so it is not deflected. Electron is much lighter than proton (about 1836 times lighter). In a magnetic field, lighter charged particles are deflected more.

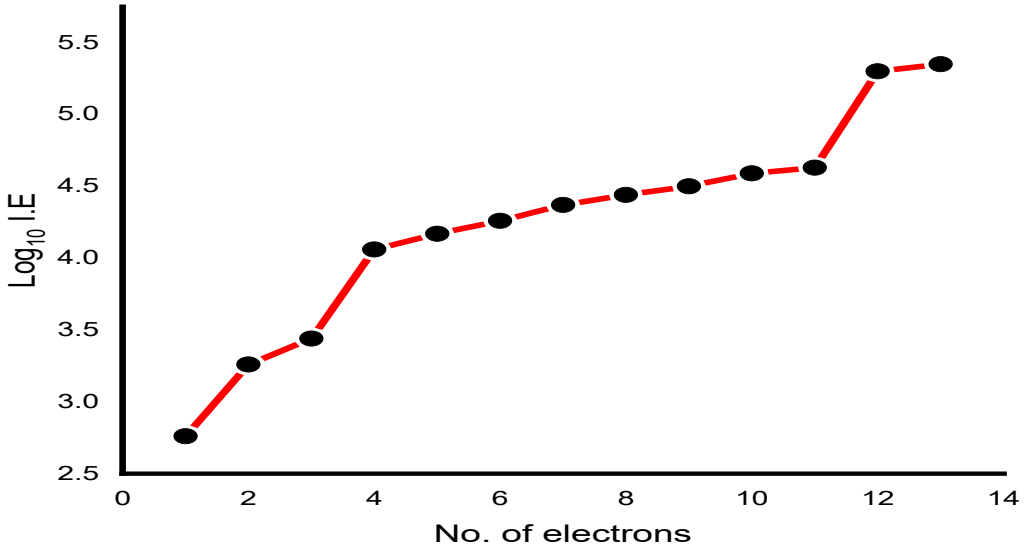
Quick Check 2.2

a) Ionization equations:

1. **1st Ionization of Ca**: $Ca \rightarrow Ca^{+} + e^{-}$
2. **3rd Ionization of K**: $K^{2+} \rightarrow K^{3+} + e^{-}$
3. **2nd Ionization of Li**: $Li^{+} \rightarrow Li^{2+} + e^{-}$
4. **5th Ionization of S**: $S^{4+} \rightarrow S^{5+} + e^{-}$

b) Sketch graph: Plot number of electrons removed on X-axis vs. $\log_{10}$ of ionization energy on Y-axis for Al (Z=13). A sharp increase after removing valence electrons indicates inner shell electrons.

Orbital	**(I.E)**	**$\log_{10}$(IE) (y)**
1st ($3p^1$)	578	2.76
2nd ($3s^2$)	1817	3.26
3rd ($3s^1$)	2745	3.44
4th ($2p^6$)	11,580	4.06
5th ($2p^5$)	14,840	4.17
6th ($2p^4$)	18,380	4.26
7th ($2p^3$)	23,260	4.37
8th ($2p^2$)	27,470	4.44
9th ($2p^1$)	31,860	4.50
10th ($2s^2$)	38,470	4.59
11th ($2s^1$)	42,680	4.63
12th ($1s^2$)	201,000	5.30
13th ($1s^1$)	222,000	5.35



c) Analyze ionization data:

i. **Reactive metal**: **II** (lowest ΔH_{i_1})

ii. **Reactive non-metal**: **IV**

iii. **Noble gas**: **I** (highest ΔH_{i_1})

iv. **Metal with formula AX_2**: **III** (moderate values)

Quick Check 2.3

a) Quantum number roles:

i. **Principal (n):** Size/energy level

ii. **Azimuthal (l):** Shape (s, p, d, f)

iii. **Magnetic (m):** Orientation in space

iv. **Spin (s):** Spin direction (±½)

b) Quantum values:

- i. n = 2 , l = 1 → m = –1, 0, +1 → **3 orientations**
- ii. n = 3 , l = 2 → **M shell, d-subshell**
- iii. L = 2 → **m = –2 , -1, 0 , +1 , +2 (5 values), max electrons = 10**

Quick Check 2.4

a) An orbital is a 3D region where the probability of finding an electron is highest.

b) l = 1 (p orbital) → m = –1, 0, +1 → **3 orientations** due to 3 values of m.

Quick Check 2.5

a) Noble gas shorthand:

i. $_{48}Cd \rightarrow [Kr]\ 4d^{10}\ 5s^2$

ii. $_{57}La \rightarrow [Xe]\ 5d^1\ 6s^2$

b) Sb (Z = 51): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^3$

c) $5p^3$ → **3 unpaired electrons**

d) **$_{13}Al$** = $1s^2\ ,2s^2\ ,2p^6\ ,3s^2\ ,3p^1$, **$_{14}Si$** = $1s^2\ ,2s^2\ ,2p^6\ ,3s^2\ ,3p^2$, **$_{28}Ni$** = $1s^2\ ,2s^2\ ,2p^6\ ,3s^2\ ,3p^6,\ 4s^2,\ 3d^8$

Quick Check 2.6

a) Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^5$

i. Block: **p-block**

ii. Group: **Group 17 (Halogens)**

iii. Period: **Period 5**

iv. Element: **Iodine (I)**

b) Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$ → Block: **d-block**

Quick Check 2.7

Write configurations:

i. **Al· (Z = 13)** → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ → One unpaired electron

ii. **O^{2-} (Z = 8)** → $1s^2\ 2s^2\ 2p^6$

iii. **Fe^{3+} (Z = 26)** → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

iv. **Cu^{2+} (Z = 29)** → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9$

v. **Cu. (Z = 29)** → One unpaired electron in $3d^{10}$ or $4s^1$ depending on configuration

SHORT ANSWER QUESTIONS

Q.2

a) P-orbital has l = 1, so m = –1, 0, +1, so it has three orientations i.e p_x, p_y, p_z.

b) Mg has filled s-orbital (stable configuration), so it requires more energy to remove electron than Al. I_3 of Mg is much higher because the third electron is removed from a new (inner) shell, closer to nucleus.

c) Lowest ionization energy = K (largest atom, least nuclear pull)
Highest ionization energy = Kr (smallest atom among them, complete shell)

d) Potassium (Z = 19)
i) Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
ii) 4s is filled before 3d because it has lower (n + l) value (Aufbau principle).

e) i) X = Cl (Z = 17, A = 35), Protons = 17, Neutrons = 35 – 17 = 18, Electrons = 17

ii) Cl^-: Electrons = 18 (gains 1), Protons = 17, Neutrons = 18

f) Mercury (Z = 80)

Configuration: [Xe] $4f^{14}\ 5d^{10}\ 6s^2$

i. Electrons with n = 3 → $3s^2\ 3p^6\ 3d^{10}$ = 18 electrons

ii. 4d orbitals → 10 electrons

iii. $4p_z$ → 2 electrons

iv. Spin up (½) = 40 electrons (half of total)

g) Sharp rise after I_2 indicates that the element likely has two valence electrons so that it will be an alkaline earth metals.

h) i) Ionization energy increases due to greater nuclear attraction after removing electrons.

(ii) I_4 is very high value because the electron is removed from a new inner shell.

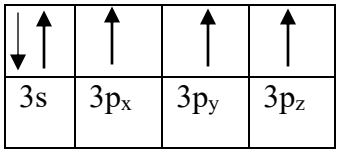
(iii) Size order: $Al^{4+} < Al^{3+} < Al^{2+} < Al^+$ (More positive charge = smaller size)

i) i) Filling order: 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p

ii) 4s has lower (n + l) than 3d, so it is filled first (Aufbau principle).

j) Phosphorus (Z = 15): Valence electrons = $3s^2\ 3p^3$

Orbital box diagram:



Descriptive Questions

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Q.4 Page (32, 33)

Q.5 Page (26, 27)

CHAPTER # 03

CHEMICAL BONDING

Solved Quick Check & Exercise

Quick Check 3.1:

a) Draw Lewis structure of N_2 & CS_2 molecule

Step 1: Count total valence electrons
- Nitrogen has 5 valence electrons.
- Total for N_2 = 5 × 2 = **10 valence electrons**

Step 2: Arrange atoms
- N ≡ N (just two nitrogen atoms)

Step 3: Form bonds and distribute electrons
- Form a **triple bond** between the two nitrogen atoms.
- Place a lone pair on each nitrogen.

:N≡N:

Step 1: Count total valence electrons
i. Carbon: 4 valence electrons
ii. Sulfur: 6 valence electrons × 2 = 12
iii. Total = 4 + 12 = **16 valence electrons**

Step 2: Arrange atoms
- Carbon is central atom: S – C – S

Step 3: Form bonds and satisfy octets
- Form **double bonds** between carbon and each sulfur.
- Distribute remaining electrons as lone pairs on sulfur atoms.

:S=C=S:

b) How many electrons are there in the valence shell of B in BF_3? Does it have the ability to accept a lone pair of electrons?

- Boron (B) has **3 valence electrons** (Group 13 of the periodic table).
- In **BF_3**, Boron forms **three single bonds** with three fluorine atoms.
- So, Boron is surrounded by **only 6 electrons** (3 bonds × 2 electrons each).

✓ **Answer:** Boron has **6 electrons** in its valence shell in **BF_3**.

- Since Boron has **only 6 electrons** (less than an octet), **it is electron-deficient**.
- This makes **BF_3** a **Lewis acid**.
- It can **accept a lone pair of electrons** from a Lewis base to complete its octet.

✓ **Answer: Yes**, Boron in **BF_3** can accept a lone pair of electrons and acts as a **Lewis acid**.

c) Show the formation of a dative bond between NH_3 and BF_3.

- **NH_3 (Ammonia)** has a **lone pair** of electrons on nitrogen.
- **BF_3 (Boron tri-fluoride)** has an **incomplete octet** — only 6 electrons around boron.
- Nitrogen can **donate** its lone pair to **boron**, forming a **coordinate covalent bond** (also called a **dative bond**).

Quick Check 3.2:

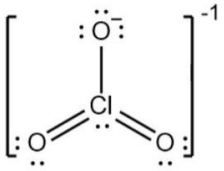
a) **Can the elements of period 2 of the periodic table have expanded octet? Explain why or why not?**

- Atoms tend to have **8 electrons** in their valence shell for stability (like noble gases).
- Some elements can have **more than 8 electrons** — this is called an **expanded octet**.

❗But Period 2 Elements (Li, Be, B, C, N, O, F, Ne) cannot do this.

No, Period 2 elements **cannot have expanded octets** because they **lack d-orbitals** to accommodate more than 8 electrons in their valence shell.

b) **Predict and explain the expanded octets in the following ions: ClO_3^-, PO_4^{3-}.**



The number of electrons in the valence shell of Cl atom can be calculated as:

No of valence electrons = 2 × (double bond electrons) + 2 × (single bond electrons) + 2(No. of lone pairs)

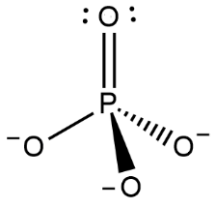
= 2 (4) + 2(1) + 2(1) = 12

Thus, Cl has 12 electrons in its valence shell and it exceeds the octet by 4 electrons. The expansion of octet is caused by the involvement of the d orbital in bonding, which can accommodate the extra electrons. The following electronic configuration of native Cl(0) atom shows that it has one unpaired electrons in the p orbitals. The presence of d-orbital allows this configuration to extend to 5 unpaired electrons by the transfer of two electrons

from the 3p pair. It explains not only the variable oxidation states of Cl, but also the possibility of accepting extra electrons in its d orbital

Cl_{17}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑					
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x^2$	$3p_y^2$	$3p_z^1$	$3d_{xy}$	$3d_{yz}$	$3d_{zx}$	$3d_{x2-y2}$	$3d_{z2}$

Cl_{17}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓ (lone pair)	↑	↑	↑	↑	↑			
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x^1$	$3p_y^1$	$3p_z^1$	$3d_{xy}^1$	$3d_{yz}^1$	$3d_{zx}$	$3d_{x2-y2}$	$3d_{z2}$



No of valence electrons = 2 × (double bond electrons) + 2 × (single bond electrons)

= 2 (2) + 2(3) = 10 electrons

Thus, P has 10 electrons in its valence shell and it exceeds the octet by 2 electrons. The expansion of octet is caused by the involvement of the d orbital in bonding, which can accommodate the extra electrons. The following electronic configuration of native P(0) atom shows that it has three unpaired electrons in the p orbitals. The presence of d-orbital allows this configuration to extend to 5 unpaired electrons by the transfer of one electrons from the 3s pair. It explains not only the variable oxidation states of P, but also the possibility of accepting extra electrons in its d orbital

P_{15}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑					
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x^1$	$3p_y^1$	$3p_z^1$	$3d_{xy}$	$3d_{yz}$	$3d_{zx}$	$3d_{x2-y2}$	$3d_{z2}$

P_{15}	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑				
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x^1$	$3p_y^1$	$3p_z^1$	$3d_{xy}^1$	$3d_{yz}$	$3d_{zx}$	$3d_{x2-y2}$	$3d_{z2}$

Quick Check 3.3:

a) **Predict with the help of electronegativity values whether the bonds in these compounds would be non-polar covalent, polar covalent, or ionic. HF, KI, CaF_2, ICl , Br_2**

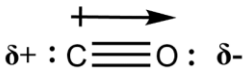
Compound	Electronegativity		ΔE.N	Bond Type
HF	H: 2.1	F: 4.0	1.9	**Polar Covalent**
KI	K: 0.8	I: 2.5	1.7	**Ionic (border line)**
CaF_2	Ca: 1	F:4.0	3.0	**Ionic**
ICl	I: 2.5	Cl: 3.0	0.5	**Polar covalent**
Br_2	Br: 2.8	Br: 2.8	0	**Non-polar**

Quick Check 3.4:

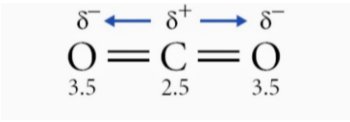
a) Can you explain why CO has a dipole moment, but CO_2 does not have any?

- **CO** has a dipole moment because it's **asymmetric** with one polar bond.

a) **CO_2** has **no dipole moment** because the **symmetry cancels out** the dipoles. The molecule is **non-polar**, despite having **polar bonds** because dipole on both sides is equal in magnitude and opposite in direction



Dipole Moment: 0.11D



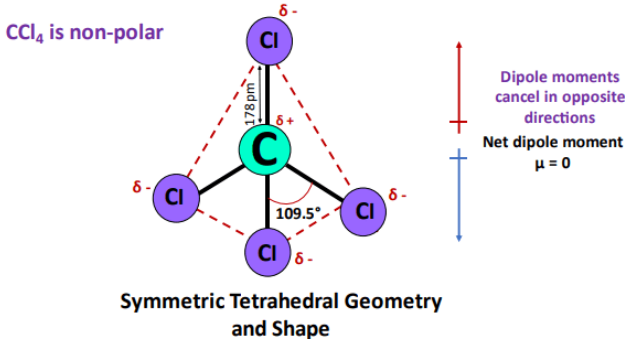
Net dipole: 0

b) Do you think that individual bonds in CCl_4 are polar? Explain in terms of the electronegativity difference. What about the polarity of overall CCl_4 molecule?

Electronegativity of carbon is 2.5 while chlorine is 3.5. ΔEN of 0.5 indicates a **polar covalent bond**.

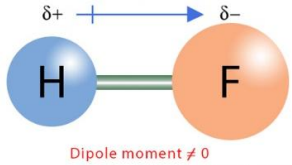
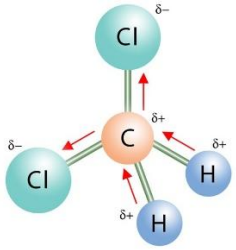
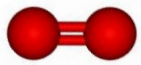
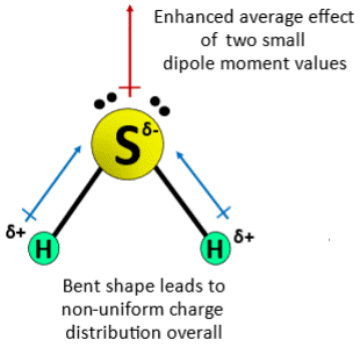
So yes, each C–Cl bond is polar, with electrons pulled slightly toward the more electronegative **chlorine.** CCl_4 has a **tetrahedral shape**, with Cl atoms arranged **symmetrically** around the central carbon.

Even though each bond is polar, the dipoles cancel out due to the **symmetrical 3D shape**.



Symmetric Tetrahedral Geometry and Shape

c) Are these molecules polar or non-polar? Briefly give reasons. HF, CH_2Cl_2, O_2, H_2S

Molecule	Electronegativity		ΔE.N	Type	Shape	Figure
HF	H: 2.1	F: 4.0	1.9	Polar	Linear	
CH_2Cl_2 Dichloromethane	H: 2.1 C: 2.5	Cl: 3.0 C: 2.5	C-H: 0.4 C-Cl: 0.5	Polar	Tetrahedral	
O_2	O: 3.44	O: 3.44	0	Non-polar	Linear	
H_2S	H: 2.1	S: 2.58	0.47	Slightly polar	V/ Bent	

Quick Check 3.5

a) HI is a stronger acid and a robust reducing agent, whereas HF is a weaker acid. Explain.

HI is a stronger acid and a better reducing agent than HF primarily due to differences in bond length and bond energy. The H–I bond is much longer and weaker (bond energy ≈ 295 kJ/mol) compared to the short and strong H–F bond (bond energy ≈ 565 kJ/mol). This weaker bond in HI breaks more easily, allowing it to release protons (H^+) more readily, which makes it a strong acid. In contrast, HF's strong bond makes proton release more difficult, making it a weaker acid despite fluorine's high electronegativity. Additionally, the

iodide ion (I^-), being large and less electronegative, loses electrons more easily than the small, highly electronegative fluoride ion (F^-), making HI a more robust reducing agent.

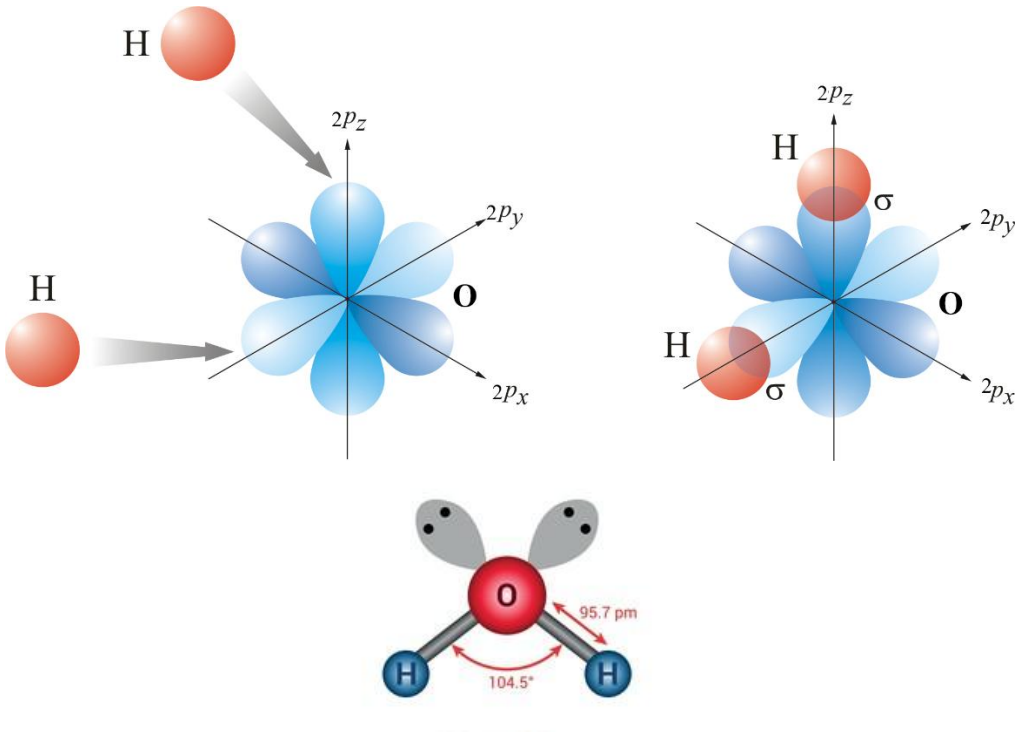
b) Acetylene (HC≡CH) is more stable than ethene (HC=CH). Can you explain why

Acetylene (HC≡CH) is more stable than ethene ($H_2C=CH_2$) primarily due to the nature of the bonds and orbital overlap involved in triple and double bonds. In acetylene, the carbon atoms are **sp-hybridized**, leading to a **linear structure** and **stronger sigma bonds** due to greater s-character (50% s, 50% p), which results in **shorter and stronger bonds**. The triple bond consists of one strong σ bond and two π bonds formed by effective sideways overlap of un-hybridized p orbitals. In contrast, in ethene, the carbon atoms are **sp^2-hybridized** with only 33% s-character, forming a weaker σ bond and a single π bond. The increased s-character in acetylene's bonds leads to **greater bond energy** and overall **stronger C–C bonding**, making acetylene more thermodynamically stable than ethene, despite having more π bonds.

Quick Check 3.6

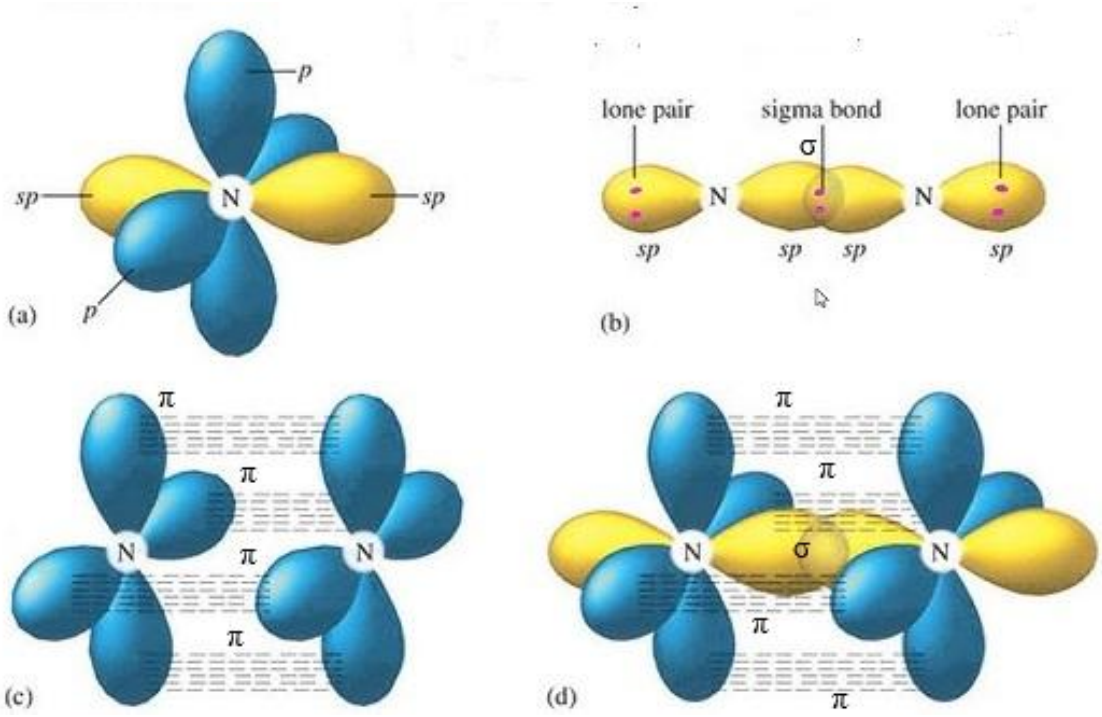
a) Draw the orbital structures of H_2O and N_2 molecules.

The H_2O molecule is a non-linear molecule which is formed by the combination of one oxygen and two hydrogen atoms. The two 2p (say 2py and 2pz) orbitals of oxygen containing one electron each can overlap with the 1s orbitals of two hydrogen atoms. A v-shaped molecule is thus formed having a bond angle of 104.5° as in Fig



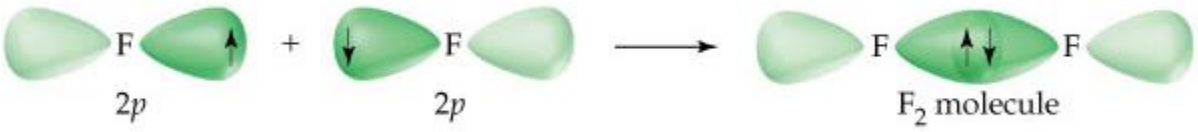
Consider the formation of a triple covalent bond between nitrogen atoms (N ≡N). There are three unpaired electrons on each atom in perpendicular p orbitals (say Px, Py and Pz). The px orbitals on the two oxygen atoms are oriented in such a way that they overlap end-to-end linearly. The

linear overlap of the px orbitals gives a σ bond. However, the py and pz orbitals on both atoms are aligned parallel to each other. They overlap in a parallel way so that the four p lobes overlap above the plane of the nuclei and the other four lobes below the plane as shown in Figure. This results in the formation of two π bond. The nitrogen atoms are triply bonded through one σ and two π bonds



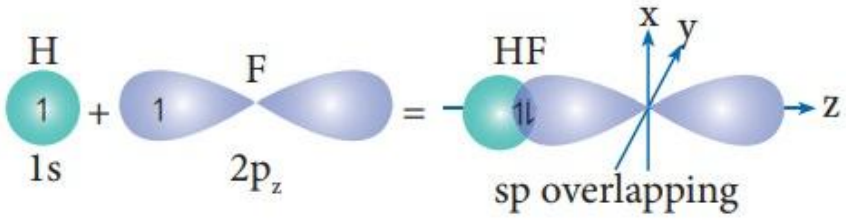
b) Draw the orbital overlap to show the formation of F_2 and HF molecules

In the **F_2 (fluorine)** molecule, each fluorine atom has the electron configuration $1s^2\ 2s^2\ 2p^5$, with one unpaired electron in the 2p orbital. According to Valence Bond Theory (VBT), the single covalent bond in F_2 is formed by the **end-to-end (linear) overlap of two 2p orbitals**, one from each fluorine atom. This results in the formation of a **sigma (σ) bond**, which is the only bond holding the two atoms together and making it a stable molecule.



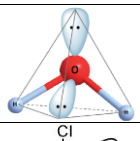
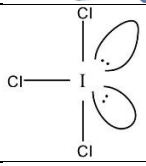
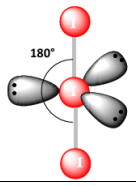
In the **HF (hydrogen fluoride)** molecule, hydrogen has one electron in its 1s orbital, and fluorine has the electron configuration $1s^2\ 2s^2\ 2p^5$, with one unpaired electron in one of its 2p

orbitals. According to Valence Bond Theory (VBT), the covalent bond in HF is formed by the **end-to-end overlap of hydrogen's 1s orbital with fluorine's 2p orbital**, resulting in the formation of a **sigma (σ) bond**.



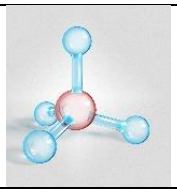
Quick Check 3.7

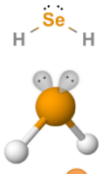
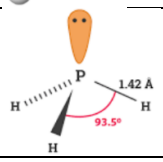
a) Calculate the number of lone pairs on the central atoms in the following:

Molecule	Group Number of central atom	Valence Electrons	Number of electrons used in bonding	Total electron pairs around central atom	Bond pair	Lone pair	Figure	Electronic Geometry	Shape	Angle
H_2O	O: 16	6 electron	2	4 AB_2L_2	2	2		Tetrahedral	V/Bent/ angular	Less than 109.5°
ICl_3	I: 17	7 electron	3	5 AB_3L_2	3	2		Trigonal bipyramidal	T-Shape	≈ 90°
I_3^-	I: 17	7electrons	1 one electron pair is gained from iodide	3 AB_2L_3	2	3		Trigonal bipyramidal	Linear	180°

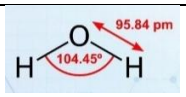
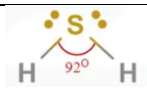
b) Calculate the number of bond pairs and lone pairs in SiH_4, H_2Se, and PH_3.

b) Predict the shapes and angles in SiH_4, H_2Se, and PH_3.

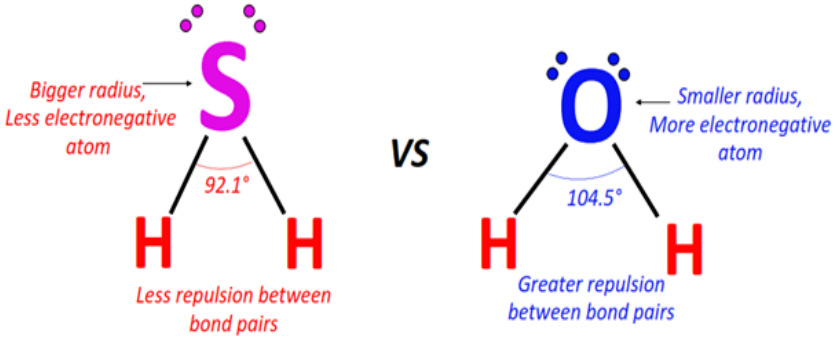
Molecule	Group Number of central atom	Valence Electrons	Number of electrons used in bonding	Total electron pairs around central atom	Bond pair	Lone pair	Figure	Electronic Geometry	Shape	Angle
SiH_4	Si: 14	4 electron	4	4 AB_4	4	-		Tetrahedral	Tetrahedral	109.5°

H_2Se	Se: 16	6 electron	2	4 AB_2L_2	2	2		Tetrahedral	V/ Bent/ angular	Less than 109.5°
PH_3	P: 15	5 electrons	3	4 AB_3L	3	1		Tetrahedral	Pyramidal	Less than 109.5°

c) Predict how the bond angle in H_2S would be different from that in H_2O

Molecule	**Group Number of central atom**	**Valence Electrons**	**Number of electrons used in bonding**	**Total electron pairs around central atom**	**Bond pair**	**Lone pair**	**Figure**	**Electronic Geometry**	**Shape**	**Angle**
H_2O	O: 16	6 electron	2	4 AB_2L_2	4	2		Tetrahedral	V/ Bent/ angular	104.5°
H_2S	S: 16	6 electron	2	4 AB_2L_2	2	2		Tetrahedral	V/ Bent/ angular	≈ 92°

Oxygen is **smaller and more electronegative** than sulfur. It pulls bonding pairs closer, increasing **lone pair–bond pair repulsion**, widening the angle in H_2O. That's why in water bond angle is greater 104.5°

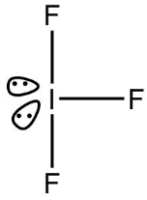
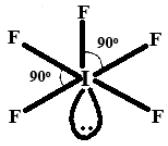


Interesting Information!

The bond angle in H_2O is 104.5°, whereas in H_2S, it is 92°. This is because the orbitals of S are larger and the lone pair exerts a stronger repulsion.

Quick Check 3.8:

a) Draw the Lewis structures for IF_3 and IF_5 and predict their geometry with reference to the VSEPR model.

Molecule	Group Number of central atom	Valence Electrons	Number of electrons used in bonding	Total electron pairs around central atom	Bond pair	Lone pair	Figure	Electronic Geometry	Shape	Angle
IF_3	I: 17	7 electron	3	5 AB_3L_2	3	2		Trigonal bipyramidal	T-Shape	≈ 90°
IF_5	I: 17	7 electron	5	6 AB_5L	5	1		Octahedral	Square pyramidal	≈ 90°

EXERCISE

MULTIPLE CHOICE QUESTIONS

Q.1 Four choices are given for each question. Select the correct choice.

I. Chemical bond formation takes place when

a) force of attraction are equal to the force or repulsion

b) force of repulsion is greater than force of attraction

c) force of attraction overcomes force of repulsion

d) none of these

II. An ionic compound A^+B^- is most likely to be formed when

a) the ionization energy of A is high and electron affinity of B is low.

b) the ionization energy of A is low and electron affinity of B is high.

c) both the ionization energy of A and electron affinity of B are high.

d) both the ionization energy of A and electron affinity of B are low.

III. Which of the following molecules has zero dipole moment?

a) NH_3 b) $CHCl_3$

c) H_2O d) BF_3

IV. The numbers of σ and π bonds in the N_2 molecule are:

a) one σ and one π bonds

b) one σ and two π bonds

c) three σ bonds only

d) two σ and one π

V. Which of the following species has unpaired electrons in antibonding molecular orbitals?

a) O_2^{2+} b) N_2^{2-}

c) B d) F_2

VI. The shape of ICl_3 according to the VSEPR model is:

a) Tetrahedral b) Trigonal planar

c) Trigonal bipyramidal d) T-shape

VII. Which of the following molecules has a net dipole moment?

a) CO_2 b) CS_2

c) SO_2 d) CCl_4

VIII. How many electrons are present in the valence shell of P in PO_4^{3-}?

a) 8 b) 10

c) 12 d) 14

IX. How many extra electrons than a normal octet are there in the valence shell of I in ICl_5? No of valence electrons = 2(No. of single bond electrons) + 2(No. of lone pairs)

a) 2 b) 3

c) 4 d) 5

X. What is the type and shape of $[ICl_4]^-$ according to the VSEPR model?

a) AB_4, tetrahedral b) AB_4, pyramidal

c) AB_5, trigonal bipyramidal d) AB_6, square planar

XI. Which of the following molecules has a central atom with sp3 hybridization and a tetrahedral electron pair geometry?

(a) BF_3 b) SO_2

(c) CCl_4 d) PCl_5

XII. Which of the following species contains a dative bond?

a) CH_4 b) NaCl

c) NH_4^+ d) O_2

Q No. 02: SHORT ANSWER QUESTIONS

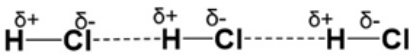
a. Define the following

(i) **Dipole**: A molecule with δ+ charge on one part and δ- charge on the other part is called a dipole

(ii) **Bond Order:** The number of bonds formed between two atoms after the atomic orbitals overlap, is called the bond order and is taken as half of the difference between the number of bonding electrons (say a) and anti-bonding electrons (say b).

$$\text{Bond order} = \frac{a - b}{2}$$

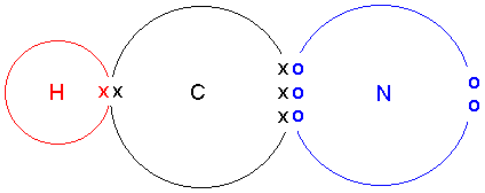
(iii) **Permanent dipole-permanent dipole force:** The force of attraction between the positive end of a polar molecule and the negative end of a nearby polar molecule is called permanent dipole-permanent dipole force.



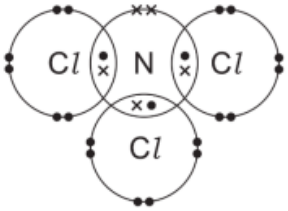
(iv) **London dispersion force:** The momentary force of attraction between an instantaneous dipole and an induced dipole is called instantaneous dipole-induced dipole force, also known as London dispersion forces

b. Draw the Lewis (electron dot) structures for the following species:

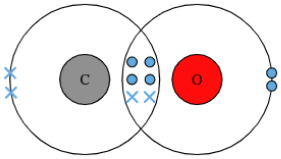
(i) **HCN**



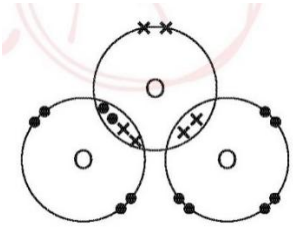
(ii) **NCl_3**



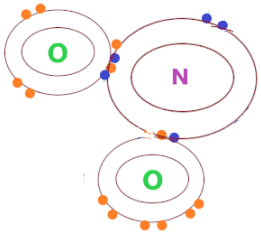
(iii) **CO**



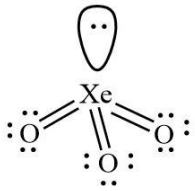
(iv) **O_3**



(v) **NO_2**

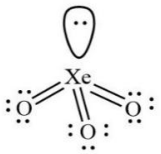


c. Xenon is a noble gas (group 18); xenon trioxide has the following structure:



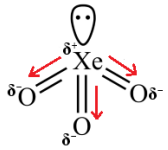
i. By counting electron pairs around the central atom, explain why xenon trioxide has this shape

According to VSEPR Theory, the two electron pairs of a double bond and three electron pairs of a triple bond, contain a higher electronic charge density, but they are regarded as single pairs.

Molecule	Group Number of central atom	Valence Electrons	Number of electrons used in bonding	Total electron pairs around central atom	Bond pair	Lone pair	Figure	Electronic Geometry	Shape	Angle
XeO_3	Xe: 18	8 electron	6	4 AB_3L	3	1		Tetrahedral	Pyramidal	Less than 109.5°

ii. Draw a structure of xenon trioxide showing partial charges on the atoms and the direction of the dipole in the molecule.

Due to the asymmetrical shape caused by the lone pair, these dipoles do not cancel out, making XeO_3 a polar molecule.

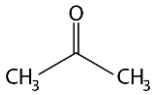


pyramidal

d. Explain the difference between the formation of σ and π bonds.

Page number 59-61

e. The structure of propanone (acetone) is:



i. Show how the central carbon atom forms σ and π bonds through hybridization.

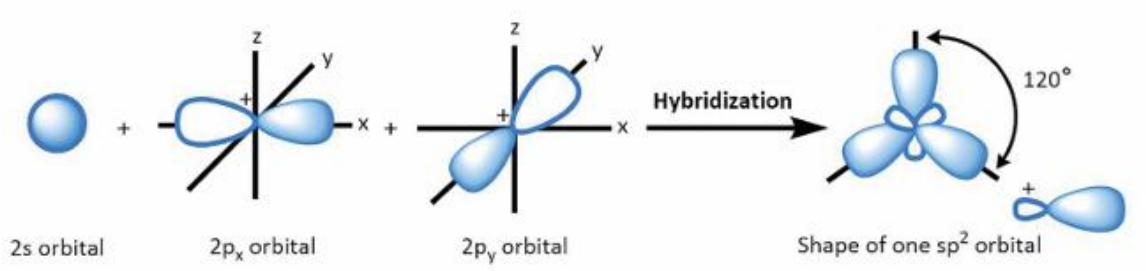
In **acetone**, the central carbon atom in the carbonyl group is **sp² hybridized**. It uses its three sp² orbitals to form **three sigma (σ) bonds**—two with methyl carbons and one with oxygen. The **remaining un-hybridized p orbital** overlaps with a p orbital from oxygen to form a **pi (π) bond**, completing the **C=O double bond**. This results in a **planar structure** around the central carbon with approximately **120° bond angles**

Electronic configuration of carbon:

C_6	↑↓	↑↓	↑	↑	
Ground State	$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$

C_6	↑↓	↑	↑	↑	↑
Excited State	$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$

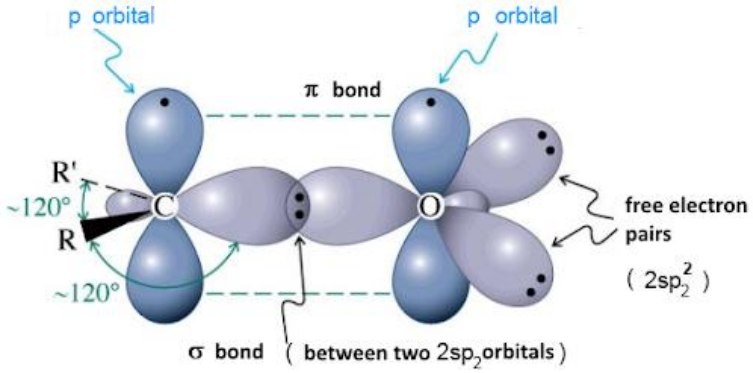
$2s^1$, $2p_x^1$,$2p_y^1$ will undergo sp^2 hybridization as follows:



Electronic configuration of Oxygen:

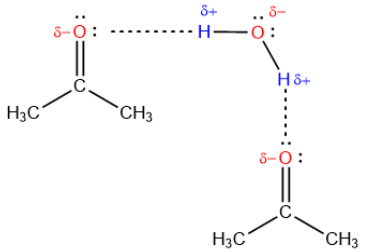
O_8	↑↓	↑↓	↑↓	↑	↑
	$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$

Oxygen is divalent atom, has ability to make double bond.



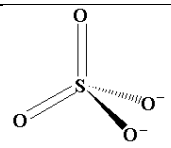
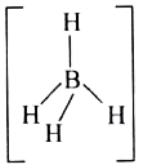
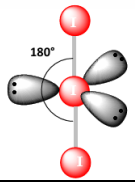
R=R/ : -CH_3 / methyl group

ii. **Can propanone make a hydrogen bond with water when both are intermixed?**
Yes, propanone can form hydrogen bonds with water, with water donating a hydrogen (from its –OH group) to the lone pair on propanone's carbonyl oxygen. This interaction helps explain why propanone is **highly soluble in water**.



f. Predict the shapes of sulfate ($[SO_4]^{2-}$, borate $[BH_4]^-$and tri-iodide ions $[I_3]^-$ according to the VSEPR model.

Ion	Group Number of central atom	Valence Electrons	Number of electrons used in bonding	Total electron pairs around	Bond pair	Lone pair	Figure	Electronic Geometry	Shape	Angle

				central atom						
SO_4^{2-}	S: 16	6 electron	6	4 AB_4	4	0		Tetrahedral	Tetrahedral	109.5°
BH_4^-	B: 13	3 electron	3 one electron pair is gained from hydride	4 AB_4	4	0		Tetrahedral	Tetrahedral	109.5°
I_3^-	I: 17	7electrons	1 one electron pair is gained from iodide	3 AB_2L_3	2	3		Trigonal bipyramidal	Linear	180°

g. Sketch the molecular orbital pictures of π(2p) and π*(2p)

Page number 65

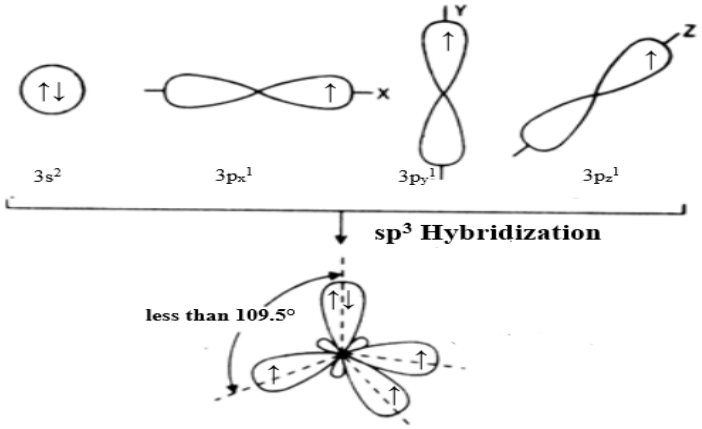
h. Sketch the hybrid orbitals and bond formation in PCl_3, $SiCl_4$, and NH_4^+.

i. PCl_3

Phosphorus belongs to group number 15. It has 5 valence electron out of these 3 are used in bond formation with each chlorine (via single bond) and one lone pair is found. It shows sp^3 hybridization.

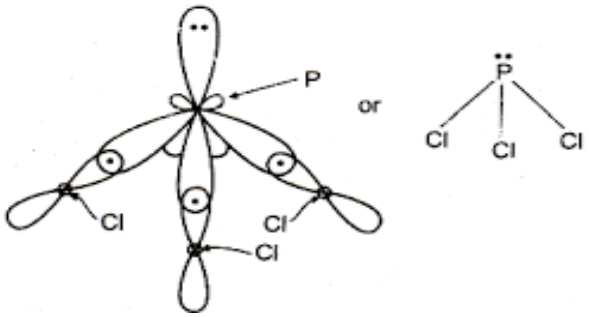
P_{15}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x$	$3p_y$	$3p_z$

$3s^2$,$3p_x^1$,$3p_y^1$ and $3p_z^1$ will undergo hybridization having one lone pair



Cl_{17}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x$	$3p_y$	$3p_z$

Each chlorine will make single bond



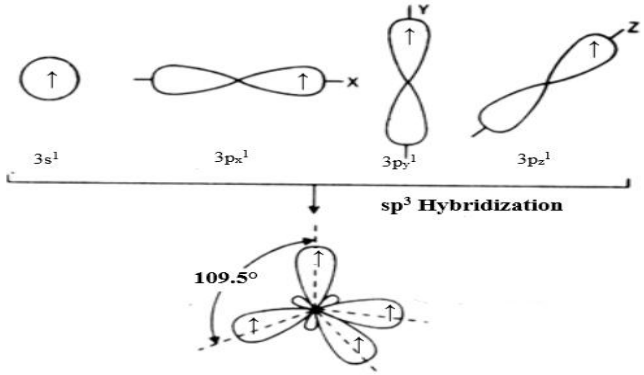
- Electronic Geometry: Tetrahedral
- Shape: Pyramidal
- Angle: less than 109.5°

ii. $SiCl_4$

Silicon belongs to group number 14. It has 4 valence electrons all are involved in bond formation with each chlorine (via single bond).

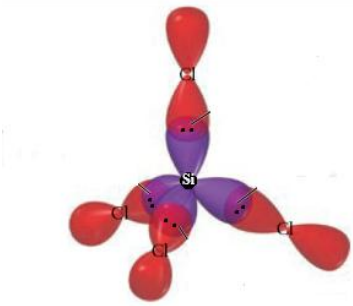
Si_{14}	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x$	$3p_y$	$3p_z$

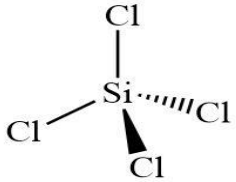
$3s^1$,$3p_x^1$,$3p_y^1$ and $3p_z^1$ will undergo sp^3 hybridization



Cl_{17}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x$	$3p_y$	$3p_z$

Each chlorine will make single bond





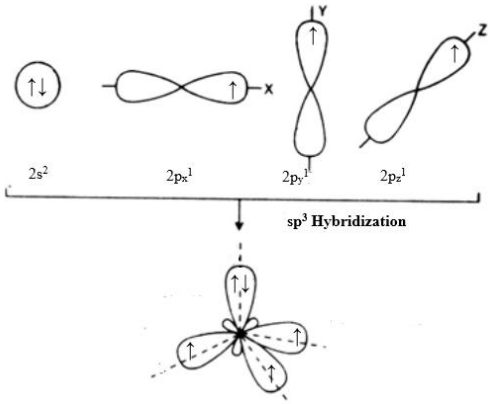
- Shape: Tetrahedral
- Angle: 109.5°

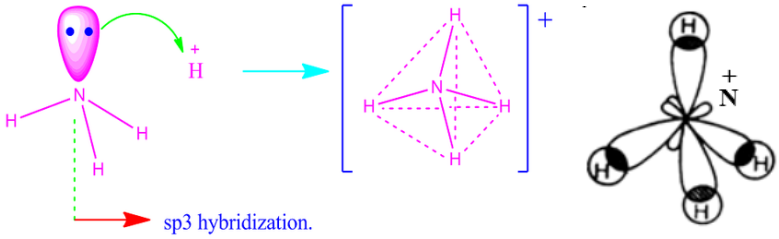
iii. NH_4^+

Nitrogen belongs to group number 15. It has 5 valence electrons, out of five one electron pair is donated to H^+, rest of the electrons make bond with each hydrogen (via single bond).

N_7	↑↓	↑↓	↑	↑	↑
	$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$

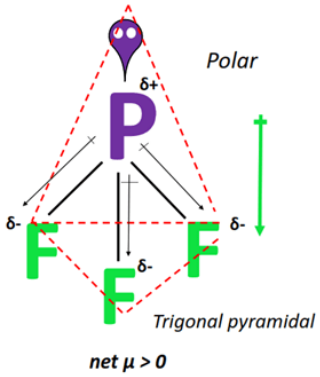
$3s^2$,$3p_x^1$,$3p_y^1$ and $3p_z^1$ will undergo sp^3 hybridization

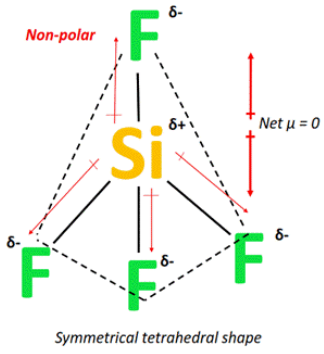




- Shape: Tetrahedral
- Angle: 109.5°

i. The structures of PF_3 and SiF_4 are given below. Redraw these with partial charges and state which is polar and which is non-polar.

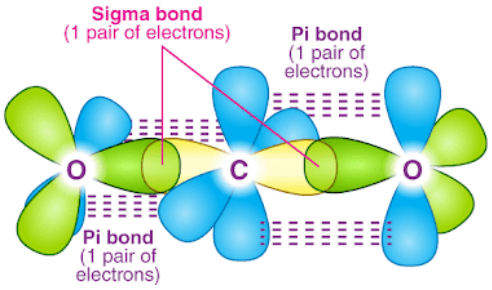


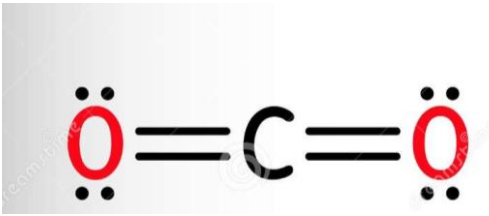


PF_3 is **polar** because it has **polar bonds** and a **non-symmetrical (trigonal pyramidal) geometry** that leads to a **net dipole moment**.

While **SiF_4 is non-polar** because it has a **symmetrical tetrahedral geometry** that causes the **individual bond dipoles to cancel**, resulting in **no net dipole moment**.

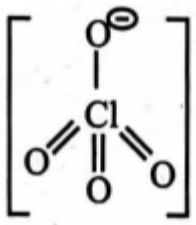
j. Draw the orbital structures of the CO_2 molecule in terms of VBT.





k. Draw the Lewis structures and tell whether these ions involve expanded octets?

i) **ClO_4^-**



The number of electrons in the valence shell of Cl atom can be calculated as:

No of valence electrons = 2 × (double bond electrons) + 2 × (single bond electrons) + 2(No. of lone pairs)

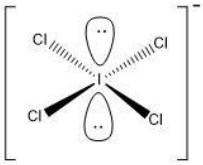
= 2 (6) + 2(1) + 2(0) = 14

Thus, Cl has 14 electrons in its valence shell and it exceeds the octet by 6 electrons. The expansion of octet is caused by the involvement of the d orbital in bonding, which can accommodate the extra electrons. The following electronic configuration of native Cl(0) atom shows that it has one unpaired electrons in the p orbitals. The presence of d-orbital allows this configuration to extend to 7 unpaired electrons by the transfer of two electrons from the 3p pair and one electron from 3s pair. It explains not only the variable oxidation states of Cl, but also the possibility of accepting extra electrons in its d orbital

Cl_{17}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑					
	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p_x^2$	$3p_y^2$	$3p_z^1$	$3d_{xy}$	$3d_{yz}$	$3d_{zx}$	$3d_{x2-y2}$	$3d_{z2}$

Cl_{17}	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	↑	↑		
	$1s^2$	$2s^2$	$2p^6$			$3s^1$	$3p_x^1$	$3p_y^1$	$3p_z^1$	$3d_{xy}^1$	$3d_{yz}^1$	$3d_{zx}^1$	$3d_{x2-y2}$	$3d_{z2}$

ii) ICl_4^-



The number of electrons in the valence shell of Cl atom can be calculated as:

No of valence electrons = 2 × (double bond electrons) + 2 × (single bond electrons) + 2(No. of lone pairs)

= 2 (0) + 2(4) + 2(2) = 12

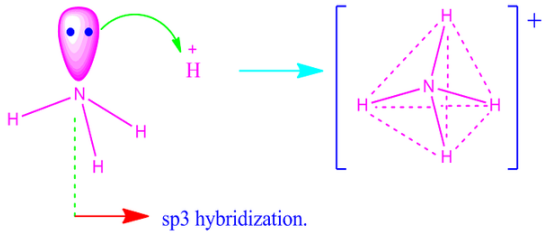
Thus, Iodine has 12 electrons in its valence shell and it exceeds the octet by 4 electrons. The expansion of octet is caused by the involvement of the d orbital in bonding, which can accommodate the extra electrons. The following electronic configuration of native I(0) atom shows that it has one unpaired electrons in the p orbitals. The presence of d-orbital allows this configuration to extend and the possibility of accepting extra electrons in its d orbital

I_{53}	$Kr_{36},4d^{10}$	↑↓	↑↓	↑↓	↑					
		$5s^2$	$5p_x^2$	$5p_y^2$	$5p_z^1$	$5d_{xy}$	$5d_{yz}$	$5d_{zx}$	$5d_{x2-y2}$	$5d_{z2}$

I_{53}	$Kr_{36},4d^{10}$	↑↓ (lone pair)	↑↓ (lone pair)	↑x	↑x	↑x	xx			
		$5s^2$	$5p_x^2$	$5p_y^2$	$5p_z^1$	$5d_{xy}$	$5d_{yz}$	$5d_{zx}$	$5d_{x2-y2}$	$5d_{z2}$

x: upcoming electrons of Cl

iii) NH_4^+



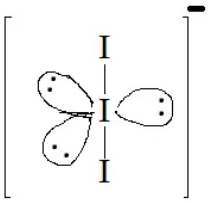
No of valence electrons = 2 × (double bond electrons) + 2 × (single bond electrons) + 2(No. of lone pairs)

2(0) + 2(4) +2(0) = 08

This **perfectly satisfies the octet rule — no expansion of the octet is needed or possible**. No empty d-orbital is available

N_7	↑↓	↑↓	↑	↑	↑
	$1s^2$	$2s^2$	$2p_x^1$	$2p_y^1$	$2p_z^1$

iv) I_3^-



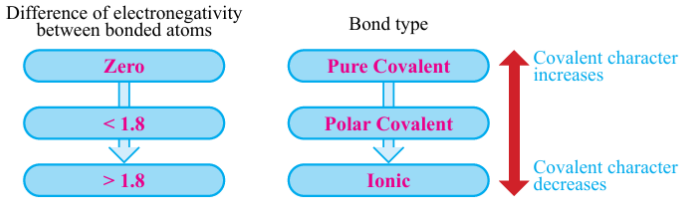
No of valence electrons = 2 × (double bond electrons) + 2 × (single bond electrons) + 2(No. of lone pairs)

2(0) + 2(2) +2(3) = 10

Thus, Iodine has 10 electrons in its valence shell and it exceeds the octet by 2 electrons. The expansion of octet is caused by the involvement of the d orbital in bonding, which can accommodate the extra electrons. The following electronic configuration of native I(0) atom shows that it has one unpaired electrons in the p orbitals. The presence of d-orbital allows this configuration to extend and the possibility of accepting extra electrons in its d orbital

I_{53}	$Kr_{36},4d^{10}$	↑↓	↑↓	↑↓	↑					
		$5s^2$	$5p_x^2$	$5p_y^2$	$5p_z^1$	$5d_{xy}$	$5d_{yz}$	$5d_{zx}$	$5d_{x2-y2}$	$5d_{z2}$

l. The bond between K and Cl is ionic but that between Si and Cl is polar covalent. Explain why.

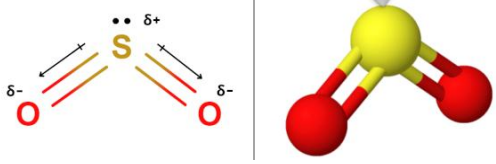


Compound	Electronegativity		ΔE.N	Bond Type
KCl	K: 0.8	Cl: 3.0	2.2	**Ionic**
$SiCl_4$	Si: 1.9	Cl: 3.0	1.1	**Polar covalent**

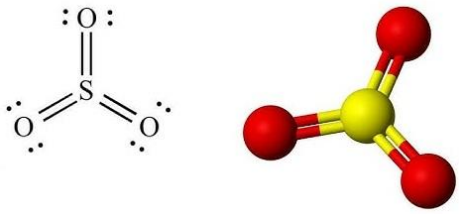
m. SO_2 is a polar molecule but SO_3 not. Justify.

Molecule	Shape	Bond Polarity	Symmetry	Net Dipole	Polarity
SO_2	Bent	Polar	Asymmetric	Yes	**Polar**
SO_3	Trigonal planar	Polar	Symmetric	No	**Nonpolar**

SO_2



SO_3



n. Which of O_2^{2+}, and O_2^{2-} would be paramagnetic? Give reason in the light of MOT.

O_2^{2+} is diamagnetic because there is no unpaired electron in $\pi^*2p_y^0$ and $\pi^*2p_Z^0$

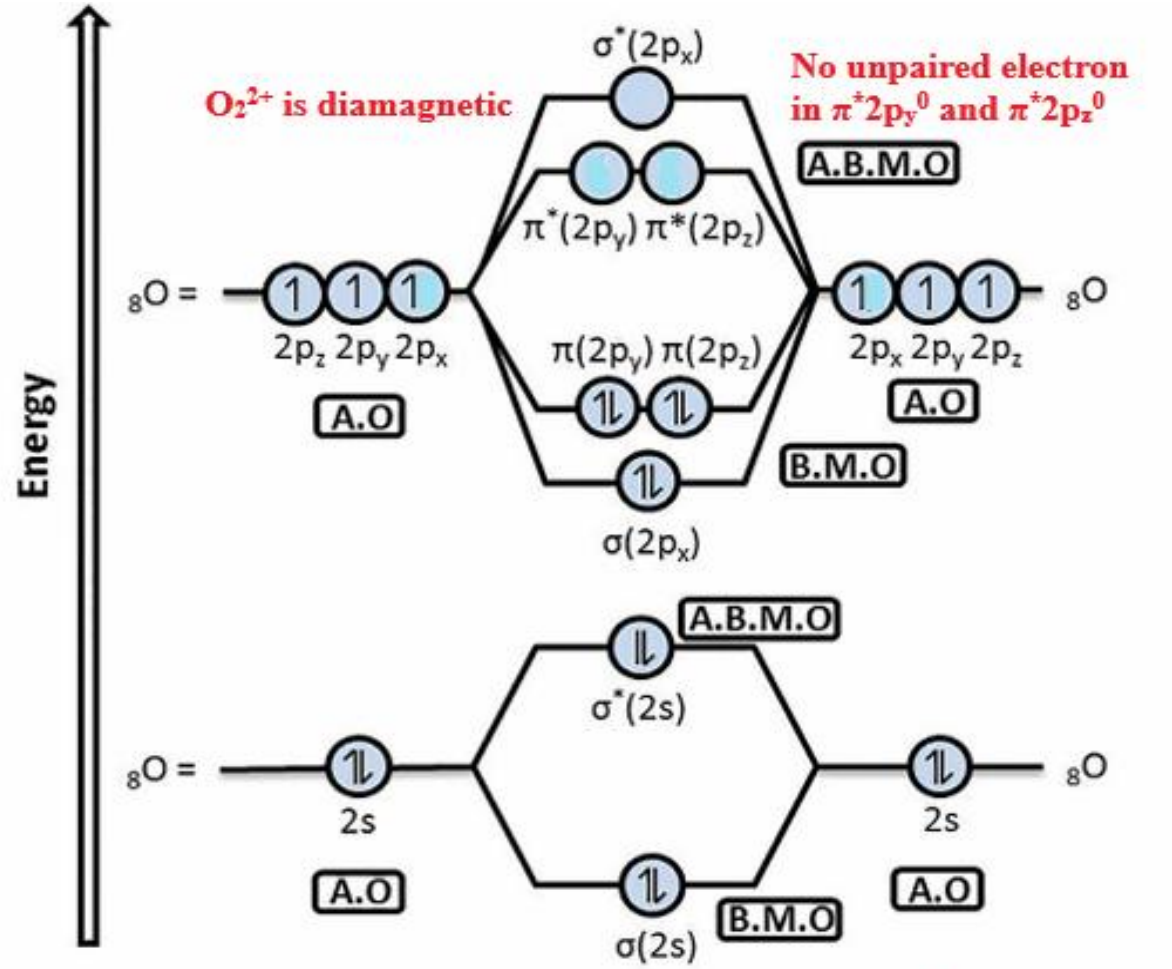


Figure Molecular orbital diagram of O_2^{2+} molecule

O_2^{2-} is diamagnetic because there is no unpaired electron in $\pi^*2p_y^2$ and $\pi^*2p_Z^2$

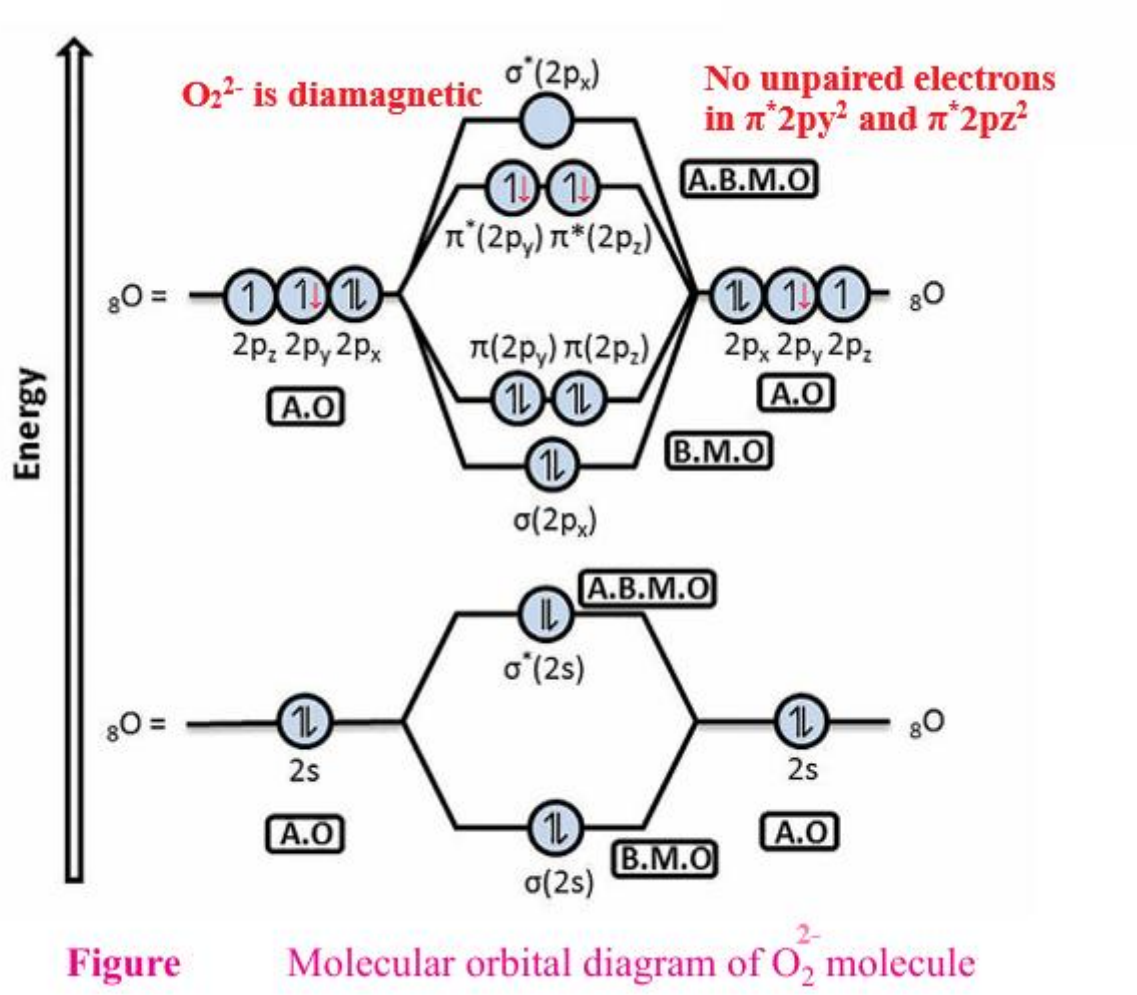


Figure Molecular orbital diagram of O_2^{2-} molecule

o. Which of the following bonds would be most polar?

Bonding atoms	Electronegativity		ΔE.N	Bond polarity
C-Cl	C: 2.5	Cl: 3.0	0.5	**Slightly Polar**
Si-F	Si: 1.9	F: 4.0	2.1	**Most Polar**
Se-F	Se: 2.5	F: 4.0	1.5	**Moderately polar**

Polarity order: Si-F > Se-F > C-Cl

p. Compare the bond energies of single, double, and triple bonds between the same two atoms (e.g., H-H, O=O, O≡O). Explain the trend in terms of the number of shared electrons

The bond energies of single, double, and triple bonds between the same two atoms increase with the number of shared electron pairs. A single bond involves one pair of electrons (σ bond) and is the weakest; a double bond has two pairs (one σ and one π bond), making it stronger; and a triple bond has three pairs (one σ and two π bonds), making it the strongest of the three. For example, between oxygen atoms:

- H-H (single bond) ≈ 436 kJ/mol

- O=O (double bond) ≈ 498 kJ/mol
- O≡O (triple bond, hypothetical) would be even stronger if stable.

This trend occurs because more shared electrons mean stronger electrostatic attraction between the nuclei and the bonding electrons, resulting in a shorter, stronger bond. However, the increase in bond strength is not perfectly proportional due to electron repulsion and orbital overlap limitations.

CHAPTER #0 4

STOICHIOMETRY

Quick Check 4.1

a) Calculate the molar mass of $KMnO_4$.

Molar mass of $KMnO_4$ = M = 39 g + 55 g + 16 g x 4 = 158g

b) Calculate the number of moles in 0.23 g of sodium.

$$\text{Moles or Gram atoms of Na} = \frac{\text{Given mass}}{\text{Molar Mass}} = \frac{0.23\text{g}}{23\ \text{g.mol}^{-}} = 0.01\ \text{mol}$$

c) Calculate the mass of 1.5 moles of $Ca(OH)_2$.

Mass = moles x molar mass

= 1.5 mol x 74.1 $g.mol^{-}$ = 111.15 g

d) The given mass of $KClO_3$ is 24.5 g. Calculate its number of moles.

$$\text{Moles or Gram atoms of } KClO_3 = \frac{\text{Given mass}}{\text{Molar Mass}} = \frac{24.5\ \text{g}}{122.5\ \text{g.mol}^{-}} = 0.2\ \text{mol}$$

e) How many molecules are present in 1.75 g of H_2O_2?

$$\text{Number of } H_2O_2 \text{ molecules} = \frac{\text{Given mass}}{\text{Molecular mass}} \times \text{Avogadro's Number}$$

$$= \frac{10.75\ \text{g}}{34.0\ \text{g mol}^{-}} \times 6.02 \times 10^{23}\ \text{molecules.mol}^{-}$$

$$= 3.1 \times 10^{22}\ \text{molecules}$$

e) How many atoms are present in 15 g of a gold ring?

$$\text{Number of Au atoms} = \frac{\text{Given mass}}{\text{Atomic mass}} \times \text{Avogadro's Number}$$

$$= \frac{15\ \text{g}}{197\ \text{g mol}^{-}} \times 6.02 \times 10^{23}\ \text{atoms.mol}^{-}$$

$$= 4.58 \times 10^{22}\ \text{atoms}$$

Quick Check 4.2

a) A copper wire contains 27.10×10^{25} atoms of copper. Calculate the number of moles of copper.

$$\text{Moles of Cu} = \frac{\text{Number of atoms}}{\text{Avogadro's Number}} = \frac{27.10 \times 10^{25}\ \text{atoms}}{6.02 \times 10^{23}\ \text{atoms.mol}^{-}} = 450.2\ \text{mol}$$

b) Calculate the molecules of 1×10^{-6} g of isopentyl acetate, $C_7H_{14}O_2$ which are released in a typical bee sting. How many atoms of carbon, hydrogen and oxygen are present in it?

$$\text{Number of } C_7H_{14}O_2 \text{ molecules} = \frac{\text{Given mass}}{\text{Molecular mass}} \times \text{Avogadro's Number}$$

$$= \frac{1 \times 10^{-6}\ \text{g}}{130.18\ \text{g mol}^{-}} \times 6.02 \times 10^{23}\ \text{molecules.mol}^{-}$$

$$= 4.63 \times 10^{15}\ \text{molecules}$$

C-atoms = $4.63 \times 10^{15} \times 7 = 3.24 \times 10^{16}$

H-atoms = $4.63 \times 10^{15} \times 14 = 6.48 \times 10^{16}$

O-atoms = $4.63 \times 10^{15} \times 2 = 9.26 \times 10^{15}$

Quick Check 4.3

Calculate the molar mass of a gas which has density of 1.34 $g.dm^{-3}$ at STP.

Molar Mass = Density × Molar Volume

Molar Mass = $1.34\ g.dm^{-3} \times 22.4\ dm^3.mol^{-}$

$= 30.02\ g.mol^{-}$

Quick Check 4.4

Calculate the molar concentration of a solution containing 7.9 g of $KMnO_4$ dissolved in 1 dm^3 of the given solution. The molar mass of $KMnO_4$ is 158 g mol-1.

$$\text{Moles or Gram molecules of } KMnO_4 = \frac{\text{Given mass}}{\text{Molar Mass}} = \frac{7.9\ \text{g}}{158\ \text{g.mol}^{-}} = 0.05\ \text{mol}$$

$$\text{Concentration} = \frac{\text{Moles}}{\text{Volume in dm}^3} = \frac{0.05\ \text{mol}}{1\ \text{dm}^3} = 0.05\ \text{mol.dm}^{-3}$$

Quick Check 4.5

How many moles of carbon dioxide are produced when 2.25 moles of glucose are used by a person? The oxygen is in excess. The equation for the combustion of glucose is:

$$C_6H_{12}O_6(S) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

Moles of $C_6H_{12}O_6$: Moles of CO_2

1 : 6

2.25 : $\frac{6}{1} \times 2.25 = 13.5$ mol

Quick Check 4.6

Fe_2O_3, an ore of iron is called Hematite. CO can reduce it to get free Fe as below:

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe(s) + 3CO_{2(g)}$$

How much Fe can be produced from 160 g of Fe_2O_3?

Moles of Fe_2O_3

$$\text{Moles} = \frac{\text{Given mass}}{\text{Molar Mass}} = \frac{160\ g}{160\ g.mol^{-1}} = 1\ mol$$

$$Fe_2O_{3\,(s)} + 3CO_{(g)} \rightarrow 2Fe(s) + 3CO_{2(g)}$$

Moles and Mass of Fe

Moles of Fe_2O_3 : Moles of Fe

1 : 2

Mass of Fe = Moles x Molar mass

= 2 mol x 56 $g.mol^{-1}$ = **112 g**

Quick Check 4.7

Calculate the volume of carbon dioxide produced at STP when 4.5 dm^3 of methane is burnt by a person. The oxygen is in excess. The equation for the reaction is:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Moles of CH_4

$$\text{Moles} = \frac{\text{Given volume}}{\text{Molar volume}} = \frac{4.5\ dm^3}{22.4\ dm^3.mol^{-1}} = 0.2\ mol$$

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Moles and volume of CO_2

Moles of CH_4	:	Moles of CO_2
1	:	1
0.2	:	0.2 mol

Volume of CO_2 = Moles x Molar volume

= 0.2 mol x 22.4 $dm^3.mol^{-1}$= 4.5 dm^3

Quick Check 4.8

Calculate the mass of sodium hypochlorite (NaOCl), a household bleach, produced by the reaction of 2.25 moles of chlorine with excess sodium hydroxide. The balanced equation is

$$2\ NaOH_{(aq)} + Cl_{2(g)} \rightarrow NaOCl_{(aq)} + NaCl_{(aq)} + H_2O_{(l)}$$

Moles of NaOCl

Moles of Cl_2	:	Moles of NaOCl
1	:	1
2.25	:	2.25 mol of NaOCl

Mass of NaOCl

Mass of NaOCl= Moles x Molar mass

= 2.25 mol x 74.44 $g.mol^{-1}$ = 167.49 g

Quick Check 4.9

Which of the following reaction mixtures could produce the greatest amount of product when they combine according to the reaction given below?

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

a) 1 mole of N_2 and 3 moles of H_2
b) 2 moles of N_2 and 3 moles of H_2
c) 1 mole of N_2 and 5 moles of H_2
d) 3 moles of N_2 and 3 moles of H_2
e) Each produces the same amount of product.

Option	N_2 (mol)	H_2 (mol)	Limiting Reactant	Excess Reactant (mol left)	NH_3 Produced (mol)	Explanation
A	1	3	Neither	None	2	Both reactants fully used.
B	2	3	H_2	$N_2 = 2 - 1 = 1$ mol left	2	Only 1 mol N_2 reacts with 3 mol H_2.
C	1	5	N_2	$H_2 = 5 - 3 = 2$ mol left	2	Only 3 mol H_2 reacts with 1 mol N_2.
D	3	3	H_2	$N_2 = 3 - 1 = 2$ mol left	2	Only 1 mol N_2 reacts with 3 mol H_2.
E	**All produce same amount of product.**					

Quick Check 4.10

When limestone ($CaCO_3$) is roasted, quicklime (CaO) is produced according to the following equation.

$$CaCO_{3(s)} \; CaO_{(s)} + CO_{2(g)}$$

The actual yield of CaO is 2.5 kg, when 4.5 kg of limestone is roasted. What is the percentage yield of this reaction?

Data

Given mass of $CaCO_3$ = 4.5 kg = 4500 g

Molar mass of $CaCO_3$ = 40g+12g+16gx3 = 100 $g.mol^{-}$

Actual yield of Cao = 2.5 kg = 2500 g

Molar mass of CaO = 40g+16g = 56 $g.mol^{-}$

Solution

Theoratical Yield

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

Mass of $CaCO_3$	:	**Mass of CaO**
100 g	:	56 g
1 g	:	56/100
4500 g	:	56 x 4500

$= 2520$ g Theoretical yield of CaO

Actual Yield

Actual yield of CaO as given in statement is 2500 g

Percentage Yield

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

$$\text{Percentage yield} = \frac{2500}{2520} \times 100 = 99.2\ \%$$

EXERCISE NUMERICALS

Q.5 A solution of sodium hydroxide (NaOH) is prepared by dissolving 2.00 g of solid sodium hydroxide in water to make a final volume of 250 cm^3.

a) Determine the molar mass of sodium hydroxide.

Molar mass of NaOH = 23g + 16g + 1g = 40 $g.mol^{-}$

b) Calculate the number of moles of sodium hydroxide used.

$$\text{Moles of NaOH} = \frac{\text{Given mass}}{\text{Molar Mass}} = \frac{2\ g}{40\ g.mol^{-}} = 0.05\ \text{mol}$$

c) Calculate the concentration of the sodium hydroxide solution in mol dm^{-3}.

$$\text{Concentration} = \frac{\text{Moles}}{\text{Volume in } dm^3} = \frac{0.05\ \text{mol}}{0.25\ dm^3} = 0.2\ \text{mol.dm}^{-3}$$

d) If more water is added to the above solution to raise the volume of solution to 500 cm^3, what would be the concentration now?

$$C_1V_1 = C_2V_2$$

$$C_2 = \frac{C_1V_1}{V_2} = \frac{0.2 \times 0.25}{0.5} = 0.1\ \text{mol.dm}^{-3}$$

Q.6 Ammonia gas (NH_3) reacts with oxygen gas (O_2) according to the following balanced equation:

$$4NH_3\ (g) + 5O_2\ (g) \longrightarrow 4NO\ (g) + 6H_2O\ (g)$$

In an experiment, 34.0 g of ammonia is reacted with 96 g of oxygen.

a) Determine the limiting reactant.

Moles of Reactants

$$\text{Moles of } NH_3 = \frac{\text{Mass}}{\text{Molar mass}} = \frac{34\ g}{17\ g.mol^{-1}} = 2\ \text{mol}$$

Moles of O_2 = Mass = 96 g = 3 mol

Molar mass 32 $g.mol^{-1}$

Moles of Product NO

$$4\ NH_3\ (g) + 5O_2\ (g) \longrightarrow 4NO\ (g) + 6H_2O\ (g)$$

i) From NH_3

Moles of NH_3 : Moles of NO

4 : 4

2 : 4 x 2 = 2 mol

4

ii) From O_2

Moles of O_2 : Moles of NO

5 : 4

3 : 4 x 3 = 2.4 mol

5

As NH_3 gives less amount of the product NO so it is the limiting reactant and 2 mol NO will be formed.

b) Calculate the maximum mass of nitrogen monoxide (NO) that can be formed.

Mass of NO

Mass of NO = Moles x Molar mass

= 2 mol x 30 $g.mol^{-1}$ = 60 g

c) Calculate the mass of the excess reactant remaining after the reaction is complete. (Relative atomic masses: H = 1.0, N = 14.0, O = 16.0)

Required Moles of O_2

Moles of NH_3 : Moles of O_2

4 : 5

2 : 5 x 2 = 2.5 mol

4

Excess Moles of O_2

= 3 mol – 2.5 mol = 0.5 mol

Excess Mass of O_2

Mass = Moles x Molar mass

= 0.5 mol x 32 $g.mol^{-1}$ = 16 g

Q.7 When iron(III) oxide (Fe_2O_3) reacts with carbon monoxide (CO) in a blast furnace, iron metal (Fe) is produced according to the following equation:

$$Fe_2O_3\ (s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2\ (g)$$

If 1.00 kg of iron(III) oxide is reacted with excess carbon monoxide, and 650 g of iron is obtained, what is the percentage yield of iron? (Ar of O = 16.0, Fe = 55.8)

Theoretical Yield

Moles of Fe_2O_3

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{1000\ g}{159.6\ g.mol^{-1}} = 6.265\ mol$$

$$Fe_2O_3\ (s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2\ (g)$$

Moles of Fe

Moles of Fe_2O_3 : Moles of Fe

1 : 2

6.265 : $\frac{2}{1}$ x 6.265 = 12.53 mol

Mass of Fe

Mass = Moles x Molar mass

= 12.53 mol x 56 $g.mol^{-1}$ = 670 g

Actual Yield

Amount of Iron obtained in reaction is 650 g

Percentage Yield

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Percentage yield = $\frac{650}{670}$ x 100 = 92.96 %

Q.8 $PbSO_4$ is precipitated when aqueous solutions of Na_2SO_4 and $Pb(NO_3)_2$ are mixed. Calculate the mass of $PbSO_4$ formed when 1.25 dm^3 of 0.05 M $Pb(NO_3)_2$ and 2 dm^3 of 0.025 M Na_2SO_4 are mixed.

$$Na_2SO_4\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$$

Moles Of Reactant

Moles of $Pb(NO_3)_2$

Moles = Molarity × Volume of solution =0.05 $mol.dm^{-3}$×1.25 dm^3 = 0.0625 mol

Moles of Na_2SO_4

Moles = Molarity × Volume of solution = 0.025 $mol.dm^{-3}$ × 2.00 dm^3 = 0.050 mol

Moles and mass of $PbSO_4$

i) **From $Pb(NO_3)_2$**

Moles of **$Pb(NO_3)_2$**	:	Moles of $PbSO_4$
1	:	1
0.0625	:	0.0625 mol

ii) **From Na_2SO_4**

Moles of **Na_2SO_4**	:	Moles of $PbSO_4$
1	:	1
0.05	:	0.05 mol

As **Na_2SO_4 has produced less moles of** $PbSO_4$ **so it is the limiting reactant and 0.05 mol of** $PbSO_4$ **will be produced**

Mass of $PbSO_4$

Mass of $PbSO_4$ = moles x molar mass = 0.05 mol x 303.3 $g.mol^{-1}$ = 15.165 g

CHAPTER # 05

STATES AND PHASES OF MATTER

Chapter 5

STATES AND PHASES OF MATTER

Quick Check 5.1

a) Explain why gases can be compressed easily?

Ans: The molecules of gases are widely separated from one another and most of the volume of the gas is empty space (nearly 99.9%). That is why gases can be compressed easily.

b) What is Joule-Thomson effect?

Ans: When sudden expansion of gases occurs cooling takes place. It is called Joule- Thomson effect.

c) The volume of 21g of a gas is 8 dm^3 at -90°C under 7atm pressure. Calculate the relative molecular mass of the gas.

Ans:

Mass of the gas = m = 21 g

Volume of the gas = V = 8 dm^3

Temperature = T = -90°C = -90°C +373 = 183 K

Pressure = 7 atm

R = 0.0821 atm dm^3 mol^{-1} K^{-1}

Relative molecular mass of the gas = Mr = ?

$$M = \frac{mRT}{PV}$$

$$M = \frac{21 \times 0.0821 \times 183}{7 \times 8}$$

$$M = \frac{315.5103}{56}$$

$$M = 5.63 \quad gmol^{-1}$$

Quick Check 5.2

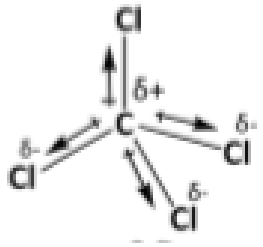
a) To improve the quality of gasoline (petrol), straight chain hydrocarbons in the gasoline fraction of petroleum are converted to branched chain ones? What could be the possible reason?

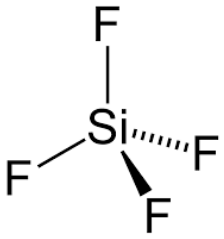
Ans: The conversion of straight-chain hydrocarbons to branched-chain ones in gasoline improves fuel quality. Especially by increasing the octane number and reducing engine knock. Branched

chain hydrocarbons, like isooctane are more resistant to engine knocking than their straight chain counterparts.

b) Which forces are present among the molecules of the following substances? CCl_4, SiF_4.

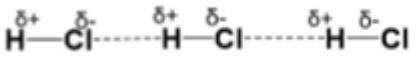
Ans: The only forces present between molecules of CCl_4 and SiF_4 are London dispersion forces or id-id forces. Both the molecules are nonpolar molecules due to their symmetrical tetrahedral structures, meaning the dipoles cancel out each other. As a result, the only intermolecular force present is the weakest type, called London dispersion forces, which arise from temporary fluctuations in electron distribution.





c) Differentiate id-id and pd-pd forces with examples.

Ans:

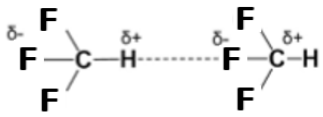
id-id Forces	**pd-pd Forces**
The momentary force of attraction between an instantaneous dipole and an induced dipole is called instantaneous dipole-induced dipole force. For example, Forces among non-polar molecules like F_2, I_2 etc.	The force of attraction between the positive end of a polar molecule and the negative end of a nearby polar molecule is called permanent dipole-permanent dipole force. For example, Forces among HCl molecules in liquid state. 

Quick Check 5.3

a) Can the CHF_3 molecule make a hydrogen bond? Explain why or why not?

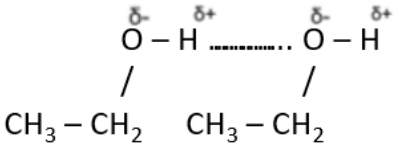
Ans: Yes. The molecules of CHF_3 make a hydrogen bond.

The molecule of fluoroform has a positive center at the H atom and the negative center on the end with F atoms.



b) Show a hydrogen bond between two molecules of ethanol.

Ans:



c) Describe which forces are present in the following and arrange them in increasing order of boiling point. **i. $CH_3CH_2CH_3$** **ii. CH_3CH_2OH** **iii. CH_3CH_2Cl**

Ans:

Molecules	Forces
$CH_3CH_2CH_3$	Instantaneous dipole-Induced dipole Forces
CH_3CH_2OH	Hydrogen Bonding
CH_3CH_2Cl	Permanent dipole-Permanent dipole forces

Increasing order of boiling point

$CH_3CH_2CH_3$ < CH_3CH_2Cl < CH_3CH_2OH

d) The boiling point difference in each of the following pairs is given

i. CH_3CH_3 (-89 ℃) and CH_3OH (65 ℃), difference = 154 ℃ <

ii. $CH_3CH_2CH_2CH_4$ (0 ℃) and $CH_3CH_2CH_2OH$ (97 ℃), difference = 97 ℃

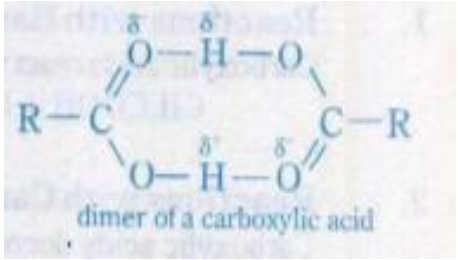
Explain why the difference decreases as the size of the molecules increases.

Ans: the boiling pont difference decreases as the size of the hydrocarbon chain increases due to:

1. **Increased London dispersion forces:** larger hydrocarbon molecules have stronger London dispersion forces, which increases their boiling points.
2. **Hydrogen boniding remains relatively constant:** The strength of hydrogen boning in alcohols remains relatively consistent, regardless of chain length.

e) Molecules of ethanoic acid (acetic acid) exist in the form of dimers in pure form but not in aqueous solution. How hydrogen bond can explain this?

Ans: Hydrogen Bonding between molecules: The carboxyl (-COOH) groups of two acetic acid molecules from strong hydrogen bonds, creating a dimer.



In aqueous solution, the dimers break apart because Water forms hydrogen bonds with the –COOH group of acetic acid, disrupting the intermolecular hydrogen bonds between acid molecules.

Quick Check 5.4

a) Arrange the following liquids is increasing order of surface tension, give reason:

Acetone (CH_3COCH_3), ethanol (C_2H_5OH), methoxy methane (dimethyl ether, CH_3OCH_3)

Ans: Increasing order of surface tension is:

Dimethyl ether (CH_3OCH_3) < Acetone (CH_3COCH_3) < ethanol (C_2H_5OH)

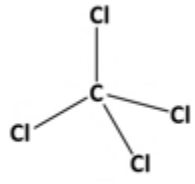
Reason:

- Dimethyl ether: Weak intermolecular forces (dipole-dipole interactions), resulting in low surface tension.
- Acetone: Dipole-dipole interactions, slightly stronger than dimethyl ether, resulting in moderate surface tensions.
- Ethanol: Hydrogen bonding between molecules, resulting in stronger intermolecular forces and higher surface tension.

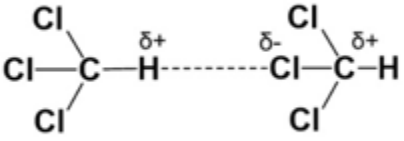
b) Why do you think tetrachloromethane (carbon tetrachloride, CCl_4) has higher viscosity than chlroform ($CHCl_3$) but less than ethanol (C_2H_5OH)?

Ans: CCl_4 has weaker London Dispersion forces compared to the dipole-dipole interactions in $CHCl_3$, so it is unexpected for CCl_4 to have higher viscosity than $CHCl_3$ having stronger forces. However, molecular shape, sized and polarizability also play roles.

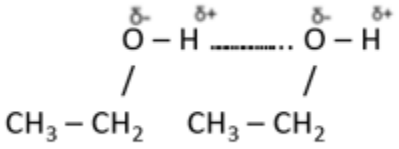
Ethanol has hydrogen bonding that leads to stronger intermolecular forces, resulting in higher viscosity compared to both CCl_4 and $CHCl_3$.



Carbon tetrachloride



Chloroform



Ethanol

e) The viscosity of honey is higher than water, explain why.

Ans: Honey contains complex sugars (fructose and glucose) that form hydrogen bonds, leading to stronger intermolecular forces. Secondly the sugar molecules in honey are larger and more complex, increasing resistance to flow. Both the hydrogen bonding and complexity of molecules make honey more viscous than that of water.

f) Which of the following is more viscous: glycerin ($CH_2OHCH_2OHCH_2OH$) or hexane (C_6H_6)? Why?

Ans: The stronger hydrogen bonding in glycerin leads to higher internal friction, making it more viscous than hexane which is nonpolar molecule with only London dispersion forces and lower viscosity.

Quick Check 5.5

a) Which of the liquids in each of the following pairs has a higher vapour pressure?

i) Alcohol, glycerine ii) Petrol, kerosene, iii) Mercury, water?

Ans: i) Alcohol, glycerine: Alcohol has weaker intermolecular forces (hydrogen bonding) compared to glycerine, which has stronger hydrogen bonding due to its three –OH groups. Weaker forces in alcohol result in higher vapor pressure.

ii) Petrol, kerosene: Petrol is more volatile and has a lower boiling point than kerosene indicating weaker intermolecular forces, thus has higher vapor pressure.

iii) Mercury, water: Mercury has very strong metallic bonds, resulting in very low vapor pressure. Water with hydrogen bonding still has a higher vapor pressure compared to mercury.

b) Which one in each of the following pairs is more viscous: Glycerin, kerosene?

Ans: The stronger hydrogen bonding in glycerin leads to higher internal friction, making it more viscous than kerosene which is nonpolar molecule with only London dispersion forces and lower viscosity.

c) Separate portions of acetone and water at the same temperature are poured on your hands. The acetone feels colder. Account for this in terms of attractive forces.

Ans: The reason, acetone feels colder than water is due to differences in intermolecular forces. Acetone has weaker intermolecular forces (dipole-dipole interactions) compared to water having hydrogen bonding. So acetone has higher rate of evaporation than that of water, resulting in colder sensation.

d) Why evaporation gets faster at higher temperatures?

Ans: With increase in temperature, the average kinetic energy of molecules increases which weakens the strength of intermolecular forces, resulting in faster evaporation.

e) Why do we feel cool near the bank of river?

Ans: Due to continuous evaporation of water of river, cooling takes place so we feel cool near the bank of river.

Quick Check 5.6

a) Why food cooking is difficult in areas with high altitudes?

Ans: at higher elevations, atmospheric pressure is lower, leading to a decrease in the boiling point of water. Water boils at a lower temperature, resulting in slower cooking times and potentially undercooked food.

b) The food cooks faster in the pressure cooker, explain.

Ans: We can increase the external pressure artificially on the surface of boiling water by using a pressure cooker. Pressure cooker is a closed container. The vapour of water formed is not allowed to escape. In this way, it exerts they develop more pressure on the water surface in the cooker and the boiling temperature increases. As more heat is absorbed in water, so food is cooked quickly under increased pressure.

c) Why the boiling point of water (100 ℃) is higher than that of ethanol (78 ℃), although both have hydrogen bonds?

Ans: Stronger the intermolecular forces, the lower the vapor pressure and higher will be the boiling point. Higher boiling point of H2O indicates stronger intermolecular forces than that of ethanol and methanol

Quick Check 5.7

a) Why solids have very low compressibility and expansion?

Ans: The atoms, molecules or ions of a solid substance are closely packed. The particles of solids cannot move closer to each other unlike gases. So solids cannot be compressed easily. As a result, solids have low compressibility.

In case of solids the forces of attractions are so strong that increase of temperature hardly affects their relative positions. That is why solids do not show a reasonable increase in volume, resulting very low expansion.

b) What is meant by habit of a crystal.

Ans: The shape of a crystal in which it usually grows is called habit of a crystal. If the conditions for growing a crystal are maintained, then the shape of the crystal always remains the same. If the conditions are changed the shape of the crystal may change. For example, a cubic crystal of NaCl becomes needle like when 10% urea is present in its solution as an impurity.

c) Why solids do not undergo translatory motion?

Ans: The atoms, ions and molecules that make up a solid are closely packed. They are held together by strong cohesive forces. The constituent atoms, ions or molecules of solids cannot move at random. The particles in a solid have vibrational motions about their mean positions.

Quick Check 5.8

a) Name the properties of liquid crystals in which they resemble solids.

Ans: Anisotropic properties: Liquid crystals exhibit anisotropic properties, such as refractive indices in different directions.

Ordered structure: liquid crystals have a degree of molecular order, similar to solids which is responsible for their anisotropic properties.

b) Mention the properties of liquid crystals in which they resemble liquids.

Ans: fluidity: liquid crystals can flow like the liquids allowing them to be easily deformed and changing shape in response to external stimuli.

Ability to take the shape of their container: like liquids, liquids crystals can be contained and take the shape of their container.

c) Which property of liquid crystals make it possible to use them in temperature sensing devices?

Ans: Liquid crystals exhibit changes in their optical properties such as color of reflectivity, in response to changes in temperature. This property makes them suitable for use in temperature sensing devices, such as room thermometers.

EXERCISE

MULTIPLE CHOICE QUESTIONS

Q.1 Four choices are given for each question. Select the correct choice.

I. London dispersion forces are the only forces present among:

a) molecules of water in liquid state

b) atoms of helium in gaseous state at high temperature

c) molecules of solid iodine ✓

d) molecules of hydrogen chloride gas

II. When the vapour pressure of a liquid equals the external pressure, what phenomenon occurs?

a) Sublimation b) Condensation c) Boiling ✓ d) Freezing

III. When water freezes at 0 °C, its density decreases due to:

a) cubic structure of ice

b) empty spaces present in the structure of ice ✓

c) decrease in volume

d) decrease in viscosity

IV. Which of following is the correct sequence of increasing ΔH_v values of substances mentioned?

a) $H_2O > NH_3 > F_2$ ✓ b) $F_2 > NH_3 > H_2O$

c) $NH_3 > H_2O > F_2$

d) $H_2O > F_2 > NH_3$

V. Surface tension of a liquid is due to:

a) inward pull of surface molecules ✓

b) upward pull from the surface

c) collision of molecules

d) repulsive forces

VI. Which change of state involves overcoming only London dispersion forces?

a) Melting of ice ($H_2O_{(s)} \rightarrow H_2O_{(l)}$)

b) Boiling of ethanol ($CH_3CH_2OH_{(l)} \rightarrow CH_3CH_2OH_{(g)}$)

c) Sublimation of iodine ($I_{2(s)} \rightarrow I_{2(g)}$) ✓

d) Dissolving sodium chloride in water

VII. In which of the following substances are London dispersion forces the only significant intermolecular forces present?

a) Ammonia (NH_3)

b) Water (H_2O)

c) Methane (CH_4) ✓

d) Hydrogen fluoride (HF)

VIII. Which of the following is a characteristic property of crystalline solids?

a) They have a range of melting points.

b) They are isotropic.

c) They have a definite and sharp melting point. ✓

d) They lack a regular arrangement of particles.

IX. Which of the following liquids would you expect to have the highest viscosity at a given temperature?

a) Water (H_2O)

b) Ethanol (CH_3CH_2OH)

c) Diethyl ether ($CH_3CH_2OCH_2CH_3$)

d) Glycerol ($CH_2(OH)CH(OH)CH_2(OH)$) ✓

X. Which type of intermolecular force is present in all types of molecule regardless of their polarity?

a) Dipole-dipole forces

b) Hydrogen bonds

c) London dispersion forces ✓

d) Ion-dipole forces

XI. Liquid crystals exhibit properties:

a) Only like solids.

b) Only like liquids.

c) Between solids and liquids. ✓

d) Unlike solids or liquids.

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a. Explain, at a molecular level, why evaporation leads to a cooling effect.

Ans: When high energy molecules leave the liquid and low energy molecules are left behind, the temperature of the liquid falls. The heat moves from the surrounding to the liquid and then the temperature of the surrounding also falls. This phenomenon helps to understand that evaporation causes cooling.

b. Explain why liquids with stronger intermolecular forces tend to have lower rates of evaporation at a given temperature compared to liquids with weaker intermolecular forces.

Ans: Stronger intermolecular forces require more energy for molecules to escape the liquid's surface and transition into the vapor phase. So fewer molecules have enough energy to break free and evaporate at a given temperature.

c. One feels sense of cooling under the fan after bath.

Ans: A person after bath feels a sense of cooling due to evaporation of water from his body when exposed to air. The molecules of H2O take away the energy of body.

d. How a dynamic equilibrium is established during evaporation of a liquid in a closed vessel at constant temperature.

Ans: The molecules of a liquid which leave the open surface are mixed up with air above the liquid. If the vessel is open these molecules go on leaving the surface of liquid. But if we close the system the molecules of liquid start gathering above the surface. These molecules collide with the walls of the container, and also with the surface of the liquid as well. There are chances that these molecules are recaptured by the surface of liquid. This process is called condensation. The two processes, i.e., evaporation and condensation continue till a stage reaches when the rate of evaporation becomes equal to the rate of condensation, the state is called dynamic equilibrium state.

e. The boiling point of water is different at Lahore and Murree hills.

Ans: The boiling of water differs between Lahore and Murree hills due to the difference in atmospheric pressure. At higher elevations like Murree, atmospheric pressure is lower, resulting in a lower boiling point of water compared to lower elevation like Lahore.

f. Discuss two significant consequences of the lower density of ice compared to liquid water in natural environments.?

Ans: Two significant consequences of ice being less dense than liquid water are:

1. **Insulation of aquatic ecosystems:** Ice floats on top of liquid water insulating aquatic life from extreme cold temperatures and allowing them to survive during winter.
2. **Frost heaving and rock formation damage:** Water expanding as it freezes can cause rocks to crack and soil to heave, leading to changes in landscapes and ecosystems.

g. Why B.P of a liquid increases when the external pressure rises?

Ans: As already explained that when vapour pressure of a liquid becomes equal to the external pressure then the liquid boils. It means that when external pressure is changed, its boiling point will also change. When the external pressure is high the liquid requires greater amount of heat to equalize its vapour pressure to external pressure. In this way boiling point is raised. Similarly, at a lower external pressure a liquid absorbs less amount of heat and it boils at a lower temperature.

h. Mention four items in which liquid crystals are used.

Ans:

1. Special liquid crystal devices can be used to diagnose the tumors and infections in the human body.
2. Liquid crystal temperature sensors can also be used to find faulty connections on a circuit board by detecting the characteristic higher temperature.
3. As temperature sensors, they are being used in thermometers.
4. These are used in Liquid crystal displays (LCDs).

i. How do you differentiate between crystalline solids and amorphous solids?

Ans:

Crystalline Solids	**Amorphous Solids**
Crystalline solids show characteristic geometrical shapes.	Amorphous solids generally appear in a lump or in a fine powder form.
Crystalline solids melt sharply at their melting points.	Amorphous solids do not have sharp melting points and melt over a range of temperatures.
Crystalline solids are anisotropic in nature. It means that their properties depend upon the direction along which the measurements are made.	Amorphous solid are isotropic. Their properties do not depend upon the direction of measurement.
Crystalline solids have long range order.	In amorphous solids, long range order is absent.

j. Propanone (CH_3COCH_3), propanol ($CH_3CH_2CH_2OH$) and butane ($CH_3CH_2CH_2CH_3$) have very similar relative molecular masses. List them in the expected order of increasing boiling points. Explain your answer.

Ans:

butane ($CH_3CH_2CH_2CH_3$): Weakest intermolecular forces i.e. Londons Dispersion forces has the lowest boiling point.

Propanone (CH_3COCH_3): Moderate intermolecular forces i.e. dipole-dipole interactions has moderate boiling point.

propanol ($CH_3CH_2CH_2OH$): Strongest intermolecular forces i.e hydrogen bonding has the highest boiling point.

k. Discuss how hydrogen bonding is responsible for the relatively high surface tension of water.

Ans: Hydrogten bonds between water molecules create a strong attraction at the surface. These bonds form a cohesive 'skin' on the surface, resulting in higher surface tension.

l. What type of intermolecular forces will dominate in the following liquids?

(i) NH_3 (ii) Ar (iii) CH_3COCH_3 (iv) CH_3OH

Ans:

(i) **NH_3 (Ammonia):** Hydrogen bonding (due to N –H bonds)

(ii) **Ar (Argon):** London dispersion forces (weak van der Waals forces due to nonpolar nature).

(iii) **CH_3COCH_3 (Propanone/Acetone):** Dipole-dipole interactions (due to polar C = O bond)

(iv) **CH_3OH (Methanol):** Hydrogen bonding (due to O – H bond)

m. The boiling points and molar masses of hydrides of some first-row elements are tabulated below:

Substance	**Boiling Point (K)**	**Molar Mass (g mol^{-1})**
CH_4	109	16
NH_3	240	17
H_2O	373	18

Suggest reasons for the difference in their boiling points in terms of the type of molecules involved and the nature of the forces present between them.

Ans:

Substance	**Boiling Point (K)**	**Molar Mass (g mol^{-1})**	**Reason for boiling points**

CH_4	109	16	London dispersion forces (weak) due to non-polar nature resulting in low boiling point.
NH_3	240	17	Hydrogen bonding (moderate) due N –H bonds but weaker than that in water, because N –H bond is less polar due to less electronegativity of N. So it has moderate boiling point.
H_2O	373	18	Hydrogen bonding due to O – H bonds, resulting high boiling point.

DESCRIPTIVE QUESTIONS

Q.3. What are London dispersion forces? Give examples, and discuss the factors affecting these forces.

Ans: page 98

Q.4 Hydrogen bonding is present in H_2O, NH_3, HF, $(CH_3)_2CO$ and $CHCl_3$ molecules. Sketch structures and discuss briefly.

Ans: Page 100

Q.5 Discuss the structural changes when water turns into ice. Justify the empty spaces in its crystals as compared to H_2O at 4°C and lower density of ice.

Ans: Page 102

Q.6 How liquid crystals resemble liquids and solids? Give their uses in daily life.

Ans: Page 112

Q.7 Describe the following properties of crystalline solids.

i) Geometrical shape ii) Melting point iii) Cleavage plan iv) Habit of a crystal

Ans: Page 110

NUMERICAL PROBLEMS

Q8. A sample of an unknown gas has a mass of 0.560 g. It occupies a volume of 2.87×10^{-4} m^3 at a temperature of 300 K and a pressure of 1.01×10^5 Pa. Calculate the molar mass of the gas. (Gas constant, R = 8.31$JK^{-1}mol^{-1}$).

Solution

Mass of the gas = m = 0.560 g

Volume of the gas = V = 2.87×10^{-4} m^3

Temperature = T = 300 K

Pressure = 1.01×10^5 Pa

R = 8.31 $JK^{-1}mol^{-1}$

Relative molecular mass of the gas = Mr = ?

$$M = \frac{mRT}{PV}$$

$$M = \frac{0.560 \times 8.31 \times 300}{1.01 \times 10^5 \times 2.87 \times 10^{-4}}$$

$$M = \frac{1396.08}{28.987}$$

$$M = 48.16 \quad \text{gmol}^{-1}$$

Q9. In a laboratory experiment, 150 cm³ of a volatile liquid was completely vaporized at 98 °C and a pressure of 1.01 × 10^5 Pa. The mass of the vapor was found to be 0.495 g. Determine the molecular mass of the liquid (R = 8.314 J K^{-1} mol^{-1}).

Solution

Mass of the gas = m = 0.495 g

Volume of the gas = V = 150 cm^3 = 150 × 106 = 1.5 × 10^{-4} m^3

Temperature = T = 98 °C = 98 + 273 = 371 K

Pressure = 1.01 × 10^5 Pa

R = 8.31 $JK^{-1}mol^{-1}$

Relative molecular mass of the gas = Mr = ?

$$M = \frac{mRT}{PV}$$

$$M = \frac{0.495 \times 8.31 \times 371}{1.01 \times 10^5 \times 1.5 \times 10^{-4}}$$

$$M = \frac{1526.82}{15.15}$$

$$M = 100.78 \quad \text{gmol}^{-1}$$

CHAPTER # 06

CHEMICAL ENERGETICS

Quick Check 6.1

i) Exothermic: energy of products < reactants; diagram dips after peak.
ii) Endothermic: energy of products > reactants; diagram ends higher.

Quick Check 6.2

a) Atomization equations:
(i) $O_2(g) \rightarrow 2O(g)$
(ii) $Ba(s) \rightarrow Ba(g)$
(iii) $Br_2(l) \rightarrow 2Br(g)$
b) Classification:
i) Exothermic
ii) Exothermic
iii) Exothermic
c) $S(g) + 2e^- \rightarrow S^{2-}(g)$: $\Delta H = -200 + 640 = +440$ kJ/mol

Quick Check 6.3

$q = mc\Delta T = 100 \times 4.18 \times 4 = 1672$ J = 1.672 kJ
$n = 0.050 \times 1.5 = 0.075$ mol
$\Delta H = -1.672 / 0.075 = -22.29$ kJ/mol

Quick Check 6.4

$\Delta Hf = -3267 - (-2364 + -858) = -45$ kJ/mol

Quick Check 6.5

Draw Hess cycle using combustion products and find ΔHf of propane.

Quick Check 6.6

a) $\Delta H = 2253 - 2346 = -93$ kJ/mol
b) $\Delta H = 4221 - 6010 = -1789$ kJ/mol

Quick Check 6.7

a) Born–Haber cycle includes atomization, ionization, EA, lattice energy.
b) ΔHf = –1107 + 161 + 520 + 79 – 328 = –675 kJ/mol

Quick Check 6.8

i) $Br_2(l) > I_2(s)$
ii) $CH_4(g) > H_2(g)$
iii) $Hg(l) > Na(s)$
iv) $SO_2(g) > SO_3(l)$

Quick Check 6.9

a)
i) Negative
ii) Positive
b)
i. Yes
ii. No
iii. Yes
iv. Yes
v. No

Exercise Answers

Multiple Choice Questions

1. b
2. d
3. c
4. a
5. c
6. c
7. d
8. a
9. c
10. c
11. c
12. b

Short Questions

a. Exothermic: releases heat (ΔH negative). Endothermic: absorbs heat (ΔH positive).
b. Enthalpy is the total heat content (H) of a system at constant pressure.

c. Entropy (S): disorder; Gibbs Free Energy (G): usable energy (G = H – TS).
d. ΔH°r: heat of reaction; ΔH°f: heat when 1 mol of compound forms from elements.
e. i) Solution: $NaCl(s) \rightarrow Na^+ + Cl^-$;
ii) Hydration: $Na^+(g) \rightarrow Na^+(aq)$;
iii) Atomization: $Br_2(l) \rightarrow 2Br(g)$;
iv) Combustion: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
f. Ionic bonds are strong; high energy is needed to form lattice.
g. Influenced by ionic charge and size.
h. Ion-dipole interactions release energy, making hydration exothermic.
i. Bonds broken: C–H, O=O; Bonds formed: C=O, O–H.
j. ΔG < 0 for spontaneity. Always spontaneous if ΔH < 0 and ΔS > 0.
k. +ΔHsol: endothermic (absorbs energy), –ΔHsol: exothermic (releases energy).
l. Higher charge density = more exothermic hydration.

Descriptive Questions

Q3. Hess's Law: Total ΔH is independent of the path. (On page 126-127)

Applications:
1. Determine ΔH for complex reactions.
2. Calculate ΔH using ΔHf or ΔHc.

Q4. Lattice Energy: Energy to form 1 mol ionic solid from gaseous ions.
Born–Haber Cycle helps calculate it using ΔHat, IE, EA, and ΔHf. (On page 132, 134-135)

Numerical Problems (Solved)

Q.5 (a) and (b)

Given:
Mass of water (m) = 100.0 g
Initial temperature (Ti) = 25.00°C
Final temperature (Tf) = 26.03°C
Specific heat capacity (c) = 4.18 J/g°C
ΔT = Tf - Ti = 26.03 - 25.00 = 1.03°C

i) $q = mc\Delta T = 100 \times 4.18 \times 1.03 = 430.54\ J \approx 430\ J$

ii) Moles of NaOH = 0.400 g / 40 g/mol = 0.01 mol
$\Delta H = -q/n = -0.430\ kJ / 0.01\ mol = -43.0\ kJ/mol$

Q.6

Use Hess's Law: $\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$
(i) $NH_3(g) + aq \rightarrow NH_3(aq)$; $\Delta H_1 = -35.16$ kJ/mol
(ii) $HCl(g) + aq \rightarrow HCl(aq)$; $\Delta H_2 = -72.41$ kJ/mol
(iii) $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$; $\Delta H_3 = -51.48$ kJ/mol

Total $\Delta H = -35.16 - 72.41 - 51.48 = -159.05$ kJ/mol

Q.7

Combustion: $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
Given:
$\Delta Hc(C_2H_5OH) = -1367$ kJ/mol
$\Delta Hf(CO_2) = -393.7$ kJ/mol
$\Delta Hf(H_2O) = -285.8$ kJ/mol

$\Delta Hf(C_2H_5OH) = [2\times(-393.7) + 3\times(-285.8)] - (-1367)$
$= (-787.4 - 857.4) - (-1367) = -1644.8 + 1367 = -277.8$ kJ/mol

Q.8

Use Born–Haber cycle:
$\Delta Hf = \Delta Hat(K) + IE(K) + \Delta Hat(Br_2) + EA(Br) + \text{Lattice Energy}$
$-392 = 90 + 420 + 112 - 342 + L.E$
$L.E = -392 - 280 = -672$ kJ/mol

Q.9

$\Delta S°surroundings = -\Delta H / T$
$\Delta H = -1270.2$ kJ/mol $= -1270200$ J/mol
$T = 298$ K
$\Delta S° = -(-1270200) / 298 = 4262.4$ J/K

Q.10

Question truncated in image, assumed to be continuation of ΔG° calculation.
Use: $\Delta G = \Delta H - T\Delta S$
Given: ΔH and ΔS; convert ΔS to kJ/K and substitute into equation.
Example: $\Delta G = -17.8 - (298 \times -0.1447) = -17.8 + 43.1 = +25.3$ kJ
Conclusion: Non-spontaneous at 25°C

CHAPTER # 07

REACTION KINETICS

Quick checks & exercises

Quick Check 7.1

a) The collision frequency and the orientation of molecules are necessary conditions a reaction to occur. Justify the statement.

Ans. For a reaction to occur, particles must **collide with enough energy** (activation energy) and in the **correct orientation**. Without proper alignment, even high-energy collisions may not lead to product formation.

b) What role does the activation energy play in chemical reactions?

Ans. It is the **minimum energy** that reacting particles must possess for a successful collision leading to a reaction. Hence, the particles must have energy equal to or greater than activation energy to carry out a chemical reaction.

c) How does the activation energy affect the rate of reaction?
Higher activation energy means **fewer effective collisions**, hence a **slower reaction rate**.

Quick Check 7.2

The reaction of hydrogen and iodine to make hydrogen iodide at a particular temperature,

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

was studied at various times. At 100.0 s after the start of the reaction, the iodine concentration had fallen from 0.010 mol dm-3 to 0.0080 mol dm-3. What is the average rate of reaction during this period?

Solution:

Rate of reaction = $\Delta C / \Delta t$

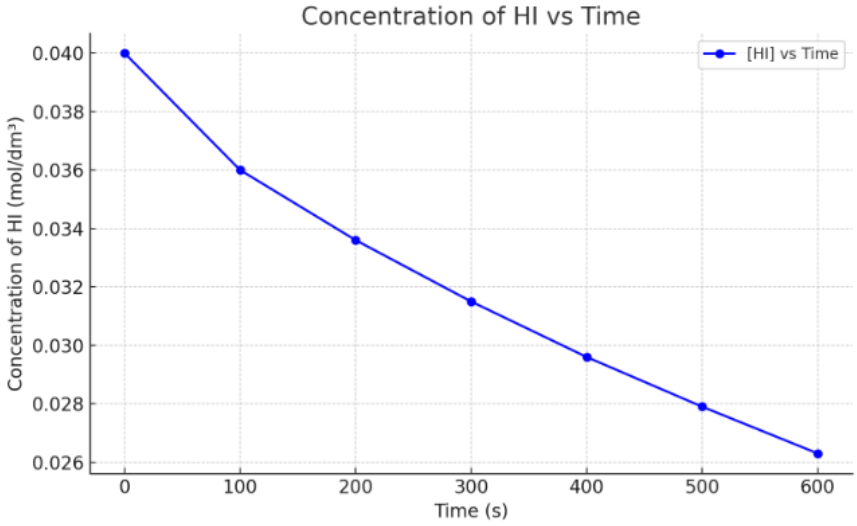
Rate = $0.010 - 0.0080/100$

$= 0.0020/100$

$= 2.0 \times 10^{-5}$ mol/dm^3/s

Quick Check 7.3

a) Plot the data in Table 7.1 for HI



b) Calculate the rate after 300 sec (when the concentration is 0.03 mol dm–3) by drawing a tangent.

Solution:

From tangent at 300 s:

Approximate change: $\Delta C = 0.0336 - 0.0296$

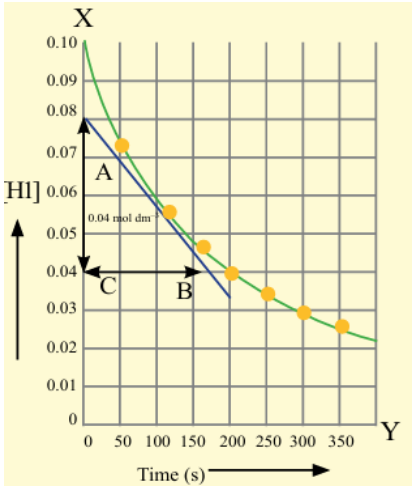
$= 0.0040$ mol/dm^3

$\Delta T = 50$ s

Rate = $\Delta C / \Delta t$

Rate = 0.0040/50

= 8.0×10^{-5} mol/dm^3/s



c) Use the same method to calculate the rate of reaction at HI concentrations of 0.10 mol dm^{-3}, 0.050 mol dm^{-3} and 0.02 mol dm^{-3}.

Use similar method to solve

d) What do you deduce about the rate of the reaction with time from these calculations?

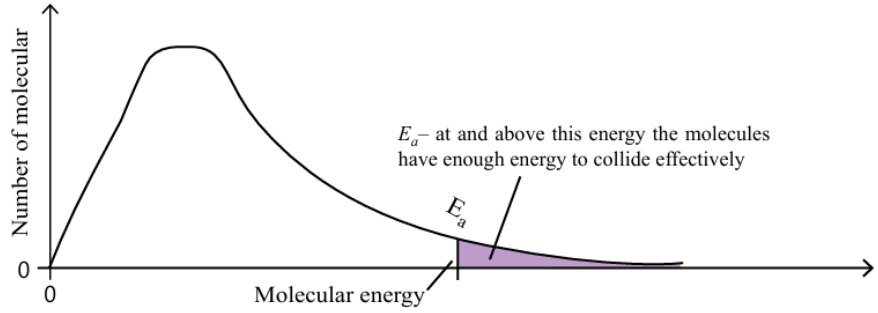
The rate of reaction is maximum at highest concentration of [HI] and decrease with the passage of time as the [HI] decreases.

e) At which concentration, the rate is highest, and lowest?

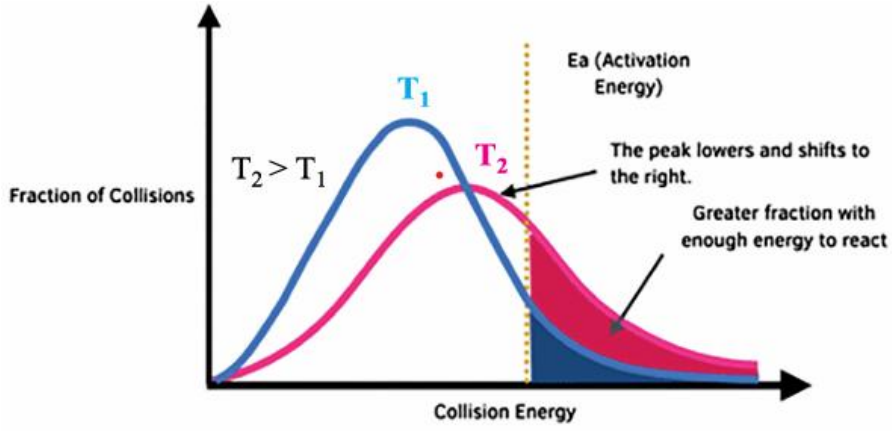
The rate of reaction is highest at maximum concentration i.e 0.1 mol/dm^3 and lowest at minimum concentration i.e 0.01 mol/dm^3

Quick Check 7.4

a) What is the Boltzmann distribution curve?
Boltzmann distribution curve is a curve showing **energy distribution** among molecules.



b) Explain why a 10 °C rise in temperature approximately doubles the rate of a reaction.
A 10 °C rise of temperature increases number of particles with energy ≥ Ea, thus doubling the rate of reaction as is shown in the following graph.



Quick Check 7.5

a) Can a catalyst be consumed in a chemical reaction? Why or why not?
No, because It's regenerated and unchanged overall.

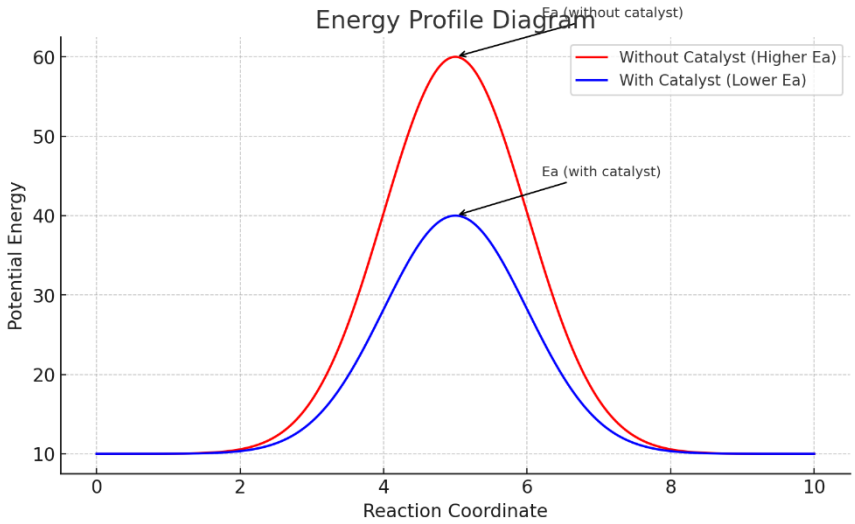
b) Explain whether the reaction below an example of heterogeneous or homogeneous catalysis is:

$$2SO_2(g) + O_2(g) \rightarrow 2SO$$

Above equation is a Homogenous catalysis. nitric oxide (NO) acts as a catalyst and is also in the gaseous state like the reactants. Since both the catalyst and

the reactants are in the same phase, this is an example of homogeneous catalysis. In such catalysis, the catalyst forms an intermediate with the reactants, lowers activation energy, and is regenerated at the end. This makes the reaction proceed faster without being consumed.

c) Draw an energy profile diagram to show a typical uncatalyzed reaction and an enzyme



The red curve represents the reaction without a catalyst, showing a higher activation energy (Ea.).
The blue curve represents the reaction with a catalyst, showing a lower activation energy.

Quick Check 7.6

a) How is order of reaction derived from the rate law?
The order of reaction is derived by adding the exponents to which the molar concentrations are raised to in rate law of the slow rate determining step.

b) Explain what is meant by the specific rate (rate constant) of a reaction and how it is represented in rate equation.
The specific rate constant of a chemical reaction is the rate of reaction when the concentrations of the reactants are unity i.e. 1 mol/dm^3.
In rate law:

$$\text{Rate} = k[A]^x[B]^y$$
$$\text{Rate} = k\ (1x1)$$
$$\text{Rate} = k$$

Quick Check 7.7

a) Calculate the overall order of reactions which have the rate expressions:

i) rate = $k[NO]^2\,[NH3]^0$

Rate = $k[NO]^2[NH_3]^0 \rightarrow$ Order = 2

ii) rate = $k[BrO_3]\,[Br^-]\,[H^+]^2$

Rate = $k[BrO_3]\,[Br^-]\,[H^+]^2 \rightarrow$ Order = 1+1+2=4

b) Why do you think chemists want to know the order of a reaction and the rate constant for a reaction?

Knowing order and rate constant helps understand reaction mechanism and optimize conditions.

c) Why the sum of the coefficients of a balanced chemical equation is not necessarily important to give the order of a reaction?

The order of a reaction is an experimentally determined quantity and cannot be inferred simply by looking at the reaction equation i.e. Coefficients ≠ order unless proven by experiment.

Quick Check 7.8

Consider a first-order reaction with a half-life of 15 minutes. If the initial concentration of the reactant is 0.100 mol dm-3, calculate the rate constant (k) for the reaction.

First-order reaction,

$t_{1/2}$ = 15 min = 900s

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{900} = 7.7 * 10^{-4} s^{-1}$$

Quick Check 7.9

An acidified solution of hydrogen peroxide reacts with iodide ions.

$H_2O_{2(aq)} + 2H^+_{(aq)} + 2I^-_{(aq)}\ 2H_2O(l) + I_{2(aq)}$

The rate equation for this reaction is

rate = $[H_2O_2]\,[I^-]$

The mechanism below has been proposed for this reaction.

$H_2O_2 + I^- \xrightarrow{\text{Slow}} H_2O + IO$

$H^+ + IO^- \xrightarrow{\text{Fast}} HIO$

$HIO + H^+ + I^- \xrightarrow{\text{Fast}} I^2 + H_2O$

Explain why this mechanism is consistent with the rate equation.

In the reaction involving H_2O_2 and I^-, the slow step is the reaction between H_2O_2 and I^-, making it the rate-determining step. Since H^+ only appears in

the fast steps, it does not affect the rate law. This explains why the observed rate law is Rate = $k[H_2O_2]\,[I^-]$, which matches the proposed mechanism.

Short Answer Questions

a. What do you understand by the rate of a reaction?
The rate of a reaction is the change in concentration of reactants or products per unit time. It shows how fast or slow a chemical reaction occurs.

b. Give the difference between enthalpy change of reaction and energy of activation of reaction.

Enthalpy Change (ΔH)	**Activation Energy (Ea.)**
Heat released or absorbed in a reaction	Minimum energy needed to start a reaction
Can be positive (endothermic) or negative (exothermic)	Always positive
Difference between reactant and product energy	Energy between reactants and transition state on a diagram

c. Differentiate clearly between order and molecularity of a reaction.

Order of Reaction	**Molecularity**
Sum of powers of concentration terms in rate law	Number of species colliding in an elementary step
Determined experimentally	Determined from the reaction mechanism
Can be fractional	Always a whole number
Applies to overall reaction	Applies to individual steps only

d. Why the instantaneous rate changes during a reaction?
Because reactant concentrations decrease over time, the frequency of effective collisions also decreases, leading to a decline in reaction rate.

e. Briefly summarize the effects of temperature and surface area on the rates of reactions.

- Temperature: Increase in temperature increases particle kinetic energy → more frequent and energetic collisions → faster rate.

- Surface Area: Larger surface area of solid reactants exposes more particles → higher collision rate → faster reaction.

f. Justify that the radioactive decay is always a first order reaction.
Radioactive decay follows the form:

$$Rate = k[N]$$

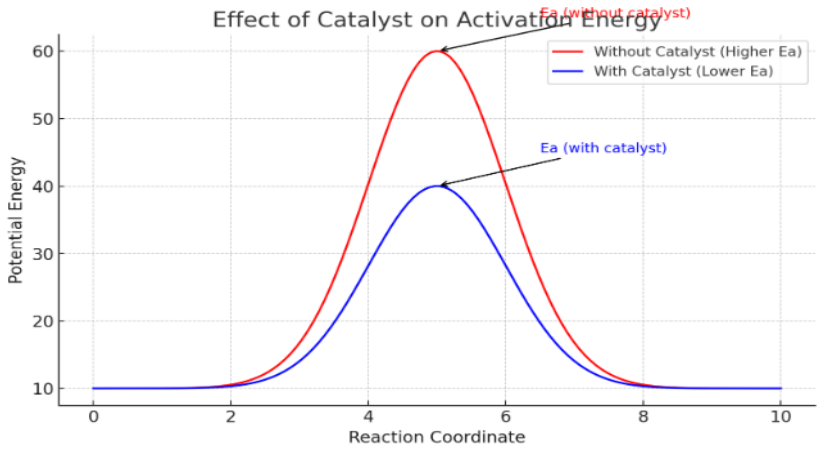
where [N] is the number of undecayed nuclei. The rate depends only on the amount remaining, not on any other substance, making it a first-order process.

g. A reaction is second order with respect to a reactant. How is the rate of reaction affected if the concentration is doubled and reduced to half?
If concentration is doubled:
Rate $\propto [A]^2$ => Rate increases by 4x
If concentration is halved:
Rate $\propto (\frac{1}{2}[A]^2) = \frac{1}{4}$ => Rate decreases by 4x

h. What is meant by half-life and what is it used for?
Half-life is the time required for the concentration of a reactant to reduce to half of its initial value. It's used in radioactive dating, pharmacokinetics, and reaction monitoring.

i. Why does wood burn more rapidly in pure oxygen than in air?
Air contains only about 21% oxygen, while pure oxygen increases the concentration and the number of effective collisions, leading to faster combustion.

h. A catalyst lowers the activation energy of a chemical reaction. Illustrate it.
A catalyst provides an alternative pathway with lower activation energy, allowing more particles to react successfully.
In an energy profile diagram, the peak (Ea.) with catalyst is lower than without catalyst.



k. The rate constant for a certain reaction is 3.5×10−4 s−1 at 25◦C. What is the order of the reaction.? Explain based on the units of the rate constant.
The units of the rate constant help determine the order of a reaction:
For zero-order: mol $dm^{-3}s^{-1}$
For first order: s^{-1}
For second order: mol $dm^{-3}s^{-1}$
Here, the unit is s^{-1}, which means that it is a first order reaction. This is because only first-order reactions have a rate constant with units of s^{-1}

m. If the initial concentration of the reactant is 0.50 mol dm−3, calculate the initial rate of the reaction.

For a first order reaction:

$$\text{Rate} = k[\text{A}]$$

$$\text{Rate} = (3.5 \times 10^{-4}\,\text{s}^{-1}) \times (0.50\,\text{mol dm}^{-3}) = 1.75 \times 10^{-4}\,\text{mol dm}^{-3}\text{s}^{-1}$$

n. How would the rate of this reaction change if the concentration of the reactant were doubled?

For first-order reactions:

$$\text{Rate} \propto [\text{A}]$$

If [A] is doubled, the rate also doubles.

o. A certain first-order reaction has a rate constant of 2.5×10−3 s−1. Calculate the half-life of the reaction in minutes.
For the first order reactions:

$$t_{1/2} = \frac{0.693}{k}$$

$$t_{1/2} = \frac{0.693}{2.5 \times 10^{-3}} = 277.2\,\text{seconds}$$

Convert into minutes:

$$t_{1/2} = \frac{277.2}{60} \approx 4.62\,\text{minutes}$$

p. A radioactive isotope decays by a first-order process with a half-life of 12 hours. Calculate the rate constant for the decay in s−1.
Convert 12 hours into seconds:

$$12 \times 3600 = 43200\,\text{seconds}$$

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{43200} \approx 1.60 \times 10^{-5}\,\text{s}^{-1}$$

Descriptive Questions

Q. 6 The reaction between hydrogen peroxide (H_2O_2) and iodide ions (I^-) in acidic
solution is believed to occur via the following mechanism:
Step 1: $H_2O_2 + I^- \longrightarrow H_2O + OI^-$ (slow)
Step 2: $OI^- + H^+ \longrightarrow HOI$ (fast)
Step 3: $HOI + I^- + H^+ \longrightarrow H_2O + I_2$ (fast)

i) Write the overall balanced equation for the reaction.

$$\boxed{H_2O_2 + 2I^- + 2H^+ \rightarrow 2H_2O + I_2}$$

ii) Identify any intermediates and catalysts in this mechanism.
OI^- and HOI are intermediates.
No catalyst is used here.

iii) What is the rate-determining step?
$H_2O_2 + I^- \rightarrow H_2O + OI^-$ is the rate-determining step.

iv) Write the rate equation for the reaction, expressing it in terms of the reactants in the overall reaction.

The rate law depends on the rate-determining step

$$\text{Rate} = k[\text{H}_2\text{O}_2][\text{I}^-]$$

This is expressed using only the reactants from the overall balanced equation.

Numerical Problems

Q.6 Calculate the reaction rate if the concentration of A is 0.5 M, the concentration of B is 0.2 and the rate constant k is 4.0M-2s-1. Given the rate law for a reaction: Rate = k[A][B]2.

Solution:

$$\text{Rate} = (4.0\,\text{M}^{-2}\text{s}^{-1}) \times (0.5\,\text{M}) \times (0.2\,\text{M})^2$$

$$= 4.0 \times 0.5 \times 0.04 = 0.08\,\text{mol dm}^{-3}\text{s}^{-1}$$

Q.7 A first order reaction is found to have a rate constant, k = 5.5 × 10-14 s-1. Find the half-life of the reaction.

Temp. (K)	Rate constant ($cm^3\ mol^{-1}\ s^{-1}$) (K)
500	6.814×10^{-4}
550	2.64×10^{-2}
600	0.56×10^{0}
650	7.31×10^{0}
700	66.67×10^{0}

Solution:

$$t_{1/2} = \frac{0.693}{5.5 \times 10^{-14}} \approx 1.26 \times 10^{13}\,\text{s}$$

Q.8 Three experiments that have identical conditions were performed to measure the initial rate of the reaction. 2HI(g) → H2(g) + I2(g)

Experiment	[HI] (M)	Rate (M/s)
1	0.015	1.1×10^{-3}
2	0.030	4.4×10^{-3}
3	0.045	9.9×10^{-3}

Write the rate law for the reaction. Find the value and units of the specific rate constant, k.

Solution:

$$\text{Rate} = k[\text{HI}]^2$$

$$k = \frac{\text{Rate}}{[\text{HI}]^2}$$

$$k = \frac{1.1 \times 10^{-3}}{(0.015)^2} = \frac{1.1 \times 10^{-3}}{2.25 \times 10^{-4}} = 4.89\, M^{-1}s^{-1}$$

$$k = \frac{9.9 \times 10^{-3}}{(0.045)^2} = \frac{9.9 \times 10^{-3}}{2.025 \times 10^{-3}} \approx 4.89\, \text{M}^{-1}\text{s}^{-1}$$

CHAPTER # 08

CHEMICAL EQUILIBRIUM

Quick Check 8.1

a)

- **Macroscopic events**: Observable phenomena (e.g., color change, gas evolution).
- **Microscopic events**: Atomic/molecular level processes (e.g., bond breaking/forming).

b)

- Dynamic equilibrium means that although concentrations are constant, at the **microscopic level**, forward and reverse reactions continue **at equal rates**.

c)

- **i)** Plot shows product concentration increasing and leveling off; reactant decreasing and leveling off.
- **ii)** Forward rate starts high and decreases; reverse starts low and increases until both rates become equal.

Quick Check 8.2

a) Yes, at 100°C, water evaporates and condenses at the same rate in a closed container → dynamic equilibrium. **b)** Yes, ice melts and water freezes at the same rate at 0°C → dynamic equilibrium.

Quick Check 8.3

a)

- **Homogeneous equilibrium**: All reactants/products in same phase.
- **Heterogeneous equilibrium**: Reactants/products in different phases.

i) Closed bottle prevents escape of gas, allowing equilibrium to establish. **ii)** Removing the cap disturbs equilibrium → shifts to produce more gas (reverse direction).

b) Forward rate slows down because reactants are being consumed, lowering collision frequency.

c) $CO2\ (g) \leftrightarrow HCO3^{-}\ (aq) + H^{+}\ (aq) \rightarrow$ Gas dissolved in water forms dynamic equilibrium with aqueous ions.

🧮 Quick Check 8.4

a) Write Kc expressions:

- i. $Kc = [Sn^{4+}][Fe^{2+}]^2 / [Sn^{2+}][Fe^{3+}]^2$
- ii. $Kc = [NO]^4[H_2O]^6 / [NH_3]^4[O_2]^5$
- iii. $Kc = [PCl_3][Cl_2] / [PCl_5]$

b) Given:

- Initial $[H_2] = [CO_2] = 10.00\ mol/dm^3$
- $[CO] = 9.47\ mol/dm^3 \rightarrow$ At equilibrium, $[H2O] = 9.47\ mol/dm^3$
- Remaining $[H_2] = [CO_2] = 10.00 – 9.47 = 0.53\ mol/dm^3$

$Kc = [H_2O][CO] / [H_2][CO_2] = (9.47)^2 / (0.53)^2 = 318.3$

🧮 Quick Check 8.5

a) Comparative Kc vs Kp:

- i. $\Delta n = -2 \rightarrow Kp < Kc$
- ii. $\Delta n = +1 \rightarrow Kp > Kc$

b) $Kp = [NH3]^2 / [N2][H2]^3 = ?$

Given:

- Total $P = 2.0 \times 10^7\ Pa$
- $P_{N2} = 1.5 \times 10^7\ Pa$
- $P_{H2} = 0.4 \times 10^7\ Pa$

$PNH3 = P_{total} - P_{N2} - P_{H2} = 0.1 \times 10^7\ Pa$

$Kp = (0.1)^2 / (1.5)(0.4)^3 = 0.01 / (1.5 \times 0.064) = 0.01 / 0.096 = 0.104\ Pa^{-2}$

■ EXERCISE SOLUTIONS

📉 MCQs

I. c II. a III. b IV. c V. c VI. d VII. d VIII. c IX. b X. c XI. a

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📌 SHORT QUESTIONS

a. Chemical equilibrium is a state where the forward and reverse reactions occur at the same rate and concentrations remain constant.

b. Reversible reactions proceed in both directions. Example: $N_2(g) + 3H_2(g) \leftrightarrow 2NH3(g)$

c.

- i) $\Delta n = -2 \rightarrow Kp < Kc$
- ii) $\Delta n = +1 \rightarrow Kp > Kc$

d. Volume/pressure changes affect equilibrium position when gaseous reactants/products are involved and $\Delta n \neq 0$, but Kc stays constant.

e. Characteristics:

- Rate of forward = reverse
- Concentrations are constant
- Reached from either side
- Catalyst doesn't affect equilibrium position

f. Dynamic equilibrium: Reactions continue at molecular level; macroscopically nothing changes.

g. Forward rate decreases as reactants are consumed.

h. Pressure shifts equilibrium to melt ice without heat (volume decrease favored).

i. Equilibrium constant conditions:

- Reaction must be reversible
- System must be closed

j. $Qc = [H_2][I_2]/[HI]^2 = (0.5)(0.1)/(2.0)^2 = 0.05/4 = 0.0125$

- If $Kc > 0.0125$, then: $Qc < Kc \rightarrow$ more HI forms → Answer: iii)

k. Let x be moles of SO_3 at equilibrium = 0.040 mol/dm³ From stoichiometry:

$2SO_2 + O_2 \rightarrow 2SO_3$

$\Delta[SO_2] = x = 0.040 \times (2/2) = 0.040$

$[SO_2] = 0.060 - 0.040 = 0.020$ mol/dm³

$[O_2] = 0.050 - 0.5 \times x = 0.050 - 0.020 = 0.030$ mol/dm³

Answer: $[O_2] = 0.030$ mol/dm³

CHAPTER # 9

ACID-BASE CHEMISTRY

Solution of Quick Checks

Quick Check 9.1

a) Identify the conjugate acid-base pairs in the following reactions:

i. $HCN(g) + H_2O \rightleftharpoons H_3O^+(aq) + CN^-(aq)$

Ans. $HCN(g) + H_2O \rightleftharpoons H_3O^+(aq) + CN^-(aq)$

Acid base conj. Acid conj. base

ii. $CH_3COOH + H_2 \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

Ans. $CH_3COOH + H_2O \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

Acid base conj. Acid conj. base

iii. $H_2S(g) + H_2O \rightleftharpoons H_3O^+(aq) + HS^-(aq)$

Ans. $H_2S(g) + H_2O \rightleftharpoons H_3O^+(aq) + HS^-(aq)$

Acid base conj. Acid conj. base

iv. $HCl(g) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq) + Cl^-(aq)$

Ans. $HCl(g) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq) + Cl^-(aq)$

Acid base conj. Acid conj. base

b) Identify the Lewis acid and Lewis base in the reaction between:

i. $SO_3 + O^{2-} \rightleftharpoons SO_4^{2-}$

Ans. L. Acid L. Base

ii. $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$

Ans. L. Acid L. Base

iii. $BF_3 + NH_3 \rightleftharpoons NH_3 \rightarrow BF_3$

Ans. L. Acid L. Base

Also write down the balanced chemical equation for the reaction and explain

Quick Check 9.2	a) **A solution is prepared by mixing equal volumes of two solutions: one with a pH of 4.0 and another with a pH of 10.0. Calculate the Kw for this mixture at 25°C.** b) **At a specific temperature, the ionic product of water Kw is 1.0×10^{-14}. If the concentration of OH^- ions in a solution is 2.5×10^{-8} M. Calculate the concentration of H+ ions and the pH of the solution.** c) **c) Copperplate etching solutions are prepared by diluting concentrated HNO_3 to 0.30 M HNO_3. Calculate $[H^+]$, pH, $[OH^-]$ and pOH of this solutions at 25 ^oC.**

Solution: (a)

Solution 1 has pH=4, solution 2 has pH=10

pOH= 14^{-10} =4

$[H^+]= 10^{-pH} = 10^{-4}$

Since volumes are equal so each concentration is diluted by a factor of 2

$[H^+]= 10^{-pH} = 10^{-4}/2= 5x10^{-5}M$

$[OH^-]= 10^{-pOH} = 10^{-4}/2= 5x10^{-5}M$

$Kw= [H^+] [OH^-]= (5x10^{-5}M) (5x10^{-5}M) = 25x10^{-10}= 2.5x10^{-9}$ **Ans.**

Solution: (b)

$Kw= [H^+] [OH^-]$

$1.0x10^{-14}= [H^+] (2.5x10^{-8})$

$[H^+] = 1.0x10^{-14} /(2.5x10^{-8})= 4.0x10^{-7}M$

pH= -log $[H^+]$ = -log $(4.0x10^{-7})$=6.40 **Ans.**

Solution: (c)

$$\mathbf{HNO_3} \rightleftharpoons \mathbf{H^+ + NO_3^-}$$

0.30M 0.30M 0.30M

pH= -log $[H^+]$= -log (0.30)= 0.52

pOH= 14-0.52=13.48

$Kw= [H^+] [OH^-]$

$1x10^{-14}= 0.30x [OH^-]$

$[OH^-] = 10^{-14}/0.30= 3.33 \times 10^{-14}$ **An**s.

Quick Check 9.3	**a) The pH of a 0.10 M solution of formic acid, HCOOH, at 25°C is 2.38. Calculate Ka for formic acid at this temperature.** **b) Calculate the concentration of $[H_3O^+]$ in a 0.1 M solution of nitrous acid $[HNO_2]$, given that the acid dissociation constant Ka is 4×10^{-4}.** **c) A vinegar sample is found to have 0.837 M CH_3COOH. Its hydronium ion concentration is found to be 3.86×10^{-3} mol dm^{-3}. Calculate Ka for acetic acid.**

Solution (a): Calculating H^+ ions concentration from pH. Taking antilog

Antilog pH= antilog (2.38)

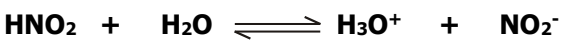
$[H^+]= 10^{-pH} = 10^{-2.38}= 4.17x10^{-3}M$

0.1 mol		0 mol	0 mol
(0.1-x)mol		x mole	x mole

(0.1-x)mol is approximately equal to 0.1 as value of x is so negligible due to limited ionization.

$Ka= [HCOO-]]H^+]/[HCOOH]= x.x/0.1= (4.17 \times 10^{-3})^2/0.1= 1.74 \times 10^{-4}$ **Ans.**

Solution (b):

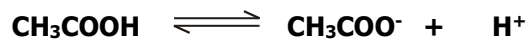


t=0	0.1mol	0 mol	0 mol
t=eq	(0.1-x=0) mol	x mol	x mol
	as x is very less		

$Ka= [NO_2-][H_3O^+]/[HNO_2]= x. x/10$

$X^2= (4 \times 10^{-4})x0.1= 4 \times 10^{-5}$ **Ans.**

Solution (c):



t=0	0.837 mol	0 mol	0 mol
t=eq	(0.837-x=0.837)mol	x mole	x mole
	as x is very less		

$Ka= [CH_3COO^-]]H^+]/[CH_3COOH]= x.x/0.837= (3.86 \times 10^{-3})^2/0.837= 1.78 \times 10^{-5}$ **Ans.**

Quick Check 9.4	**a) Explain the impact of common ion effect on solubility.** **b) To a saturated solution of AgCl, some of NaCl solution was added.** **i. State the effect of the concentration of Ag^+ on the equilibrium.** **ii. Explain your answer with respect to the common ion effect.** **c) How does a buffer maintain pH stability?** **d) Calculate the pH of a buffer consisting of 0.50 M HF and 0.45 M of a fluoride (F^-) salt before and after addition of 0.40 g NaOH to 1.0 dm^3 of the buffer (Ka of HF = 6.8×10^{-4}).** **e) Calculate the pH of a buffer solution in which 0.11 molar CH_3COONa and 0.09 molar acetic acid solutions are present. Ka for CH_3COOH is 1.85×10^{-5}**

Solution: (a) Because of the presence of common ions, the solubility of a substance always decreases at the temperature. On adding common ions, the precipitates of the already dissolved salt start reappearing.

Solution: (b)

i) The concentration of Silver ions will decrease
ii) The solubility of less soluble salt AgCl in water is suppressed by the addition of more soluble salt NaCl because of common ion effect.

Solution: (c) A buffer maintains pH stability when small amounts of strong acids or bases are added to it from outside.

Solution: (d)

pKa= -log Ka= -log (6.8×10^{-4})= 3.17

pH= pKa + log $[F^-]/[HF]$= 3.17+ log (0.45)/(0.50)= 3.17-0.0458=3.12 **Ans**

Solution: (e)

Given Information:

0.11M CH_3COONa solution means that 0.11 moles are dissolved in 1 dm^3 of solution.

$[CH_3COONa] = 0.11M$

$[CH_3COOH] = 0.09M$

K_a of $CH_3COOH = 1.85 \times 10^{-5}$

Requirement:

pH of buffer solution = ?

Formula Applied:

$$pH = pK_a + \log\frac{[salt]}{[acid]}$$

Calculation and Result:

$pK_a = \_ \log K_a$

$pK_a = -\log(1.8\times10^{-5}) = 4.74$

$pH = 4.74 + \log\frac{0.11}{0.09}$

$pH = 4.74 + 0.087 = 4.83$

Answer: **4.83**

Quick Check 9.5	**a) The solubility product constant (Ksp) of silver chloride (AgCl) is 1.77×10^{-10} at 25 °C. A solution already contains 0.10 M of sodium chloride (NaCl). Find the solubility of AgCl in a solution that contains 0.10 NaCl.** **b) Predict whether $CaSO_4$ will be precipitated or not when 10^{-3} mol dm^{-3} of each of Ca^{2+} and SO_4^{2-} is mixed.** **c) Will the aqueous solution of ammonium oxalate be acidic or basic? Explain your answer by giving the equation.** **d) Solution of potassium carbonate is basic.** **i. Explain why the solution is alkaline.** **ii. Also give the equation for the hydrolysis of the above salt.**

Solution: (a)

$AgCl \rightleftharpoons Ag^+ + Cl^-$

$[Ag^+]= s$ $[Cl^-]= 0.10$ M (From NaCl, assumed to stay constant)

$Ksp = [Ag^+][Cl^-]=s \times 0.10$

$S= 1.77\times 10^{-10}/0.10= 1.77 \times 10^{-9}$ **Ans.**

Solution: (b)

Ionic Product $= [Ca^{+2}][SO_4^{-2}]=(1 \times 10^{-3})(1\times10^{-3})= 1X10^{-6}$

Now we compare Ionic product with Ksp

Ionic product, ksp means No precipitation occurs

Solution: (c)

In water both the ammonium ions and oxalate ions undergo hydrolysis. Since NH_4^+ ion is stronger acid than $C_2O_4^{-2}$ ions which is weak base so the resultant solution will be acidic

Solution: (d) i: Potassium carbonate is a basic salt as it is formed from strong base KOH and weak carbonic acid. Upon hydrolysis K_2CO_3 makes strong base KOH and weak acid H_2CO_3, resulting in basic solution

Solution: (d)ii: $K_2CO_3 + 2H_2O \rightleftharpoons 2KOH + H_2CO_3$

Quick Check 9.6	**a) Differentiate end point and equivalent point.** **b) Explain how an indicator changes its colour in acidic and basic solution.** **c) Suggest a suitable indicator for weak acid and strong base titration.** **d) Suggest a suitable indicator for weak acid and weak base titration.**

Solution: (a) The equivalence point is the stage in a titration where the amount of OH- added is stoichiometrically equal to the amount of H+ ions originally present in the acid.

The end point is the stage when the indicator changes color. It is visual clue to the completion of neutralization process.

Solution: (b) The indicator changes its color due to pH changes in the solution. Indicators show different colors in acidic and basic solutions.

Solution: (c) For such titration, Phenolphthalein is the suitable indicator.

SOLVED EXAMPLES

Question No. 8:

A buffer solution has a pH of 5.0. It is made from a weak acid HA with a pKa of 4.8. What is the ratio of the concentration of the conjugate base [A^-] to the concentration of the weak acid [HA] in this buffer?

Solution:

Given Information:

pH = 5.0

pKa of solution = 4.8

Requirement:

[A-]/[HA] = ?

Formula Applied:

$pH = pKa + \log [A^-]/[HA]$

$\log [A^-]/[HA] = pH - pKa = 5 - 4.8 = 0.2$

$[HA][A^-] = 10^{0.2} \approx 1.58$

Answer: **1.58**

Question No. 9:

Calculate the solubility of a sparingly salt lead (II) iodide (PbI_2) in water. It has Ksp= 1.4×10^{-8}.

Solution:

Given Information:

Solubility product of **PbI_2** = K_{sp} = 1.4×10^{-8}

Required:

Solubility of Ag_2CrO_4 = S = ?

Calculation and Result:

Let solubility of **PbI_2** be 'S'

$$\mathbf{PbI_2} \rightleftharpoons Pb^{2+} + 2I^-$$

Initial stage	S	0	0
Equilibrium stage	S	2S	S

$Ksp = [Pb^{2+}][I^-]^2$

$1.4 \times 10^{-8} = (S)(2S)^2$

$1.4 \times 10^{-8} = 4S^3$

Or $S^3 = \frac{(1.4 \times 10^{-8})^{1/3}}{4}$

$S = 1.5 \times 10^{-3}$

Question No. 10:

The molar solubility of silver chromate (Ag_2CrO_4) in pure water at 298 K is 6.5×10^{-5} mol dm^{-3}. Calculate the Ksp of silver chromate at this temperature.

Solution:

Given Information:

Solubility or concentration of $\mathbf{Ag_2\ CrO_4}$ = $\mathbf{6.5\times10^{-5}}$ mol dm^{-3}

Requirement:

Value of K_{sp} = ?

Calculation and Result:

$$\mathbf{Ag_2\ CrO_4} \rightleftharpoons 2Ag^{+} + CrO_4^{-2}$$

$$\mathbf{6.5\times10^{-5}}\ M \qquad 2x\ \mathbf{6.5\times10^{-5}} \qquad \mathbf{6.5\times10^{-5}}$$

$$K_{sp} = [Ag^{+}]^2[CrO_4^{-2}]$$

$$= (2x\ \mathbf{6.5\times10^{-5}})^2\ (\mathbf{6.5\times10^{-5}})$$

$$= \mathbf{1.1 \times 10^{-12}}$$

SOLUTION OF TEXTBOOK EXERCISE

MULTIPLE CHOICE QUESTIONS

Q.1 Four choices are given for each question. Select the correct choice.

SOLUTION OF MCQS OF TEXT BOOK'S EXERCISE

Q.#	Answers
1.	b
2.	b
3.	b
4.	b
5.	a
6.	c

7.	**b**
8.	**a**
9.	**b**
10.	**a**
11.	**b**

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a. Define the following with an example for each:

- **i) Ionization constant**
- **ii) Solubility product**
- **iii) Common ion effect**
- **iv) Acid-base Indicator**

Solution:

(a) Ionization Constant: ionization constant of an acid is also called dissociation constant of acid. It is a measurement of the strength of acid. Its symbol is Ka.

(b) Solubility Product: It is defined as the product of the concentration of oppositely charged ions of a sparingly soluble salt at a given temperature. Its symbol is Ksp.

(c) Common ion effect: It is defined as the suppression of ionization of a weak electrolyte because of the addition of strong electrolyte that contains common ions.

(d) Acid-Base Indicator: A substance which describes the end point of an acid-base t titration is called acid-base indicator. e.g. Phenolphthalein and Methyl orange

b. Differentiate between:

i) Hydrolysis and dissolution ii) Acidic and basic buffer solutions

Ans.

	Hydrolysis	**Dissolution**
1	In this process a compound is broken down because of the reaction to water. e.g. hydrolysis of NaCl yields NaOH and HCl	In this process a solid solute dissolves into a liquid solution. e.g. dissolving sugar into water
2	It involves hydrolytic reactions	It only involves formation of certain attractive forces among solute and solvent particles

	Acidic Buffer Solution	Basic Buffer Solution
1	A buffer solution with pH less than 7 is called an acidic buffer solution.	A buffer solution with pH more than 7 is called Basic Buffer solution
2	**Example:** the buffer solution containing acetic acid and sodium acetate	**Example:** the buffer solution containing Ammonium hydroxide and Ammonium chloride

c. Explain the concept of conjugate acid-base pairs. How are they related in terms of proton transfer?

Ans. An acid after ionization changes to conjugate base. A base after ionization changes to conjugate acid. They differ in the presence and absence of H^+ ions

d. What is the relationship between the strength of an acid and the strength of its conjugate base?

Ans. A strong acid after ionization changes to weak conjugate base and vice versa.

e. For the following three reactions, identify the reactants that are Arrhenius bases, Bronsted-Lowry bases, and/or Lewis bases. State which type(s) of bases each reactant is. Explain your answers.

Ans.

- **$NaOH(s) \rightleftharpoons Na^+(aq) + OH^-(aq)$**

  Arrhenius base

- **$HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$**

  Bronsted acid br. Base conj. Base conj. acid

- **$H+(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq)$**

  Bronsted base

f. An amphoteric substance can behave as either an acid or a base. Identify whether water behaves as an acid or a base in each of the following reactions.

Ans.

i) $H_2O + HCl \rightleftharpoons H_3O^+ + Cl^-$

base acid conj. acid conj. base

ii) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

base acid conj. acid conj. base

iii) $HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$

Acid base conj. acid conj. base

iv) $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

Acid base conj. base conj. acid

g. Which salt would you expect to dissolve more readily in an acidic solution: Barium sulfate or Barium fluoride? Explain.

Ans. BaF_2 has greater Ksp so it will dissolve more into acidic solution.

h. Why does common ion effect decrease the solubility of a less soluble salt?

Ans. This happens because the addition of common ions from outside increases the rate of formation of sparingly salt.

i. State the basic principle of solubility product. Mention factors affecting solubility product.

Ans. Solubility indicates the maximum concentration of the sparingly soluble salt that can dissolve into water at a particular temperature. It depends upon temperature, pH and presence of common ions etc.

j. Why titration of weak acids with weak bases are not performed for volumetric analysis?

Ans. This is because the sharp end points are not obtained and completion of neutralization process cant be indicated.

k. Prove by equations what happens when Na_2CrO_4 is added to saturated solution of $PbCrO_4$.

Ans. When Na_2CrO_4 is added to saturated solution of $PbCrO_4$, the solubility of is decreased and its crystals start reappearing. This is because of the common ion effect.

$$PbCrO_4 \rightleftharpoons Pb^{2+} + CrO_4^{2-}$$

$$Na_2CrO_4 \rightleftharpoons 2Na^{+} + CrO_4^{2-}$$

l. According to the Lewis acid-base concept, Boron trifluoride (BF_3) can act as an acid. Is this statement correct?

Ans. Yes, it is correct because BF_3 acts as Lewis acid. Boron has incomplete octet and tends to receive lone pair from any electron rich species.

m. If the concentration of hydrogen ions in a solution is 1×10^{-5} M, what is the pH of the solution?

Ans. $[H^+] = 10^{-5}$

pH = $-\log [H^+] = -\log [10^{-5}] = 5$ **Ans.**

CHAPTER # 10

ELECTROCHEMISTRY

Quick Check 10.1

a) Oxidation: Loss of electrons. Example: $Zn \rightarrow Zn^{2+} + 2e^-$
Reduction: Gain of electrons. Example: $Cl_2 + 2e^- \rightarrow 2Cl^-$
b) Oxidation: Increase in oxidation number.
Reduction: Decrease in oxidation number.
Example: $Fe^{2+} \rightarrow Fe^{3+}$ (oxidation), $Mn^{7+} \rightarrow Mn^{2+}$ (reduction)
c) Determine Ox. No.:
i) $Na_2O \rightarrow$ Na = +1, So, 2(1)+x=0, x=-2
ii) $ICl_3 \rightarrow$ Cl = -1, So, x+3(-1)=0, x=-3
iii) $NO_3^- \rightarrow$ O = -2, So, 3(-2)+x=-1, -6+x=-1, x=6-1=+5
iv) $K_2Cr_2O_7 \rightarrow$ K = +1, O = -2, So, 2(1)+2x+7(-2)=0, 2+2x-14=0, x=+6
v) $KMnO_4 \rightarrow$ K = +1, O = -2, So, 1+x+4(-2)=0, 1+x-8=0, x= +7

Quick Check 10.2

Part (a)
Identify the oxidized species and reduced species in the following equation. Also identify the oxidizing and reducing agents on the reactant side.

- i. $Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$

Fe in Fe_2O_3: Oxidation state = +3
Fe in Fe: Oxidation state = 0
→ Reduction (Gain of electrons): $Fe^{3+} + 3e^- \rightarrow Fe$

C in CO: Oxidation state = +2
C in CO_2: Oxidation state = +4
→ Oxidation (Loss of electrons): $C^{2+} \rightarrow C^{4+} + 2e^-$

• Oxidized species: CO
• Reduced species: Fe_2O_3
• Oxidizing agent: Fe_2O_3 (it gets reduced)
• Reducing agent: CO (it gets oxidized)

- ii. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

C in CH_4: Oxidation state = -4
C in CO_2: Oxidation state = +4
→ Oxidation: $C^{4-} \rightarrow C^{4+} + 8e^-$

O_2: Oxidation state = 0
O in H_2O and CO_2: Oxidation state = -2
→ Reduction: $O_2 + 4e^- \rightarrow 2O^{2-}$

• Oxidized species: CH_4
• Reduced species: O_2
• Oxidizing agent: O_2
• Reducing agent: CH_4

Part (b)
Identify the species in the following reaction which undergoes both oxidation and reduction.

- i. $2Na_2O_2(s) + 2H_2O(l) \rightarrow 4NaOH(aq) + O_2(g)$

O in Na_2O_2: Oxidation state = -1
O in NaOH: Oxidation state = -2 (Reduction)
O in O_2: Oxidation state = 0 (Oxidation)

• Species undergoing both oxidation and reduction: O^{-1} in Na_2O_2 (Disproportionation reaction)

- ii. $P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$

P in P_4: Oxidation state = 0
P in PH_3: Oxidation state = -3 (Reduction)
P in $H_2PO_2^-$: Oxidation state = +1 (Oxidation)

• Species undergoing both oxidation and reduction: P_4 (Disproportionation reaction)

Quick Check 10.3

(i) **$Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$**

$$\overset{+3\ -2}{Fe_2O_3}(s) + \overset{+2-2}{CO}(g) \rightarrow \overset{0}{Fe}(s) + \overset{+4\ -2}{CO_2}(g)$$

Each Fe atom gains 3 electrons: $Fe^{3+} + 3e^- \rightarrow Fe$
Each C atom loses 2 electrons: $C^{2+} \rightarrow C^{4+} + 2e^-$
Fe: +3 → 0 ($3e^-$ gain, Reduction)
C: +2 → +4 ($2e^-$ loss, Oxidation)

$$\overset{+2\times 3=+6}{\overset{+3\ -2}{Fe_2O_3}(s) + \overset{+2-2}{CO}(g) \rightarrow \overset{0}{Fe}(s) + \overset{+4\ -2}{CO_2}(g)}$$
$$\underset{-3\times 2=-6}{}$$

2 Fe^{3+} will gain 6 electrons → 2 Fe
3 C^{2+} will lose 6 electrons → 3 CO_2
$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$

(ii) **$MnO_2 + HCl \longrightarrow MnCl_2 + Cl_2 + H_2O$**

Write oxidation number of each element.

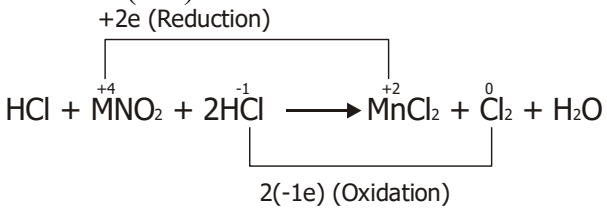
$\overset{+4}{Mn}\overset{-2}{O_2} + \overset{+1}{H}\overset{-1}{Cl} \longrightarrow \overset{+2}{Mn}\overset{-1}{Cl_2} + \overset{0}{Cl_2} + \overset{+1}{H_2}\overset{-2}{O}$

Identify the elements whose oxidation number have changed consider the change per atom.

Also write HCl twice because Cl^{-1} change to Cl^{-1} and Cl_2^0

$HCl + \overset{+4}{Mn}O_2 + HCl \longrightarrow \overset{+2}{Mn}Cl_2 + \overset{0}{Cl_2} + H_2O$

Mention the change in oxidation number with arrow. Also balance the same atom on both sides. Multiply HCl and (−1e) with 2.



Gain of electrons is equal to loss of electrons so

$HCl + MnO_2 + 2HCl \longrightarrow MnCl_2 + Cl_2 + H_2O$

Balance rest of species by general inspection

$2HCl + MnO_2 + 2HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$

or

$4HCl + MnO_2 \longrightarrow MnCl_2 + Cl_2 + 2H_2O$

(iii) **$Fe^{2+} + MnO_4^- \longrightarrow Fe^{3+} + Mn^{2+}$**

Splitting the equations into half-reactions:

Split the equation into oxidation half and reduction half.

Oxidation half reaction

$Fe^{2+} \longrightarrow Fe^{3+}$

Reduction half reaction

$MnO_4^- \longrightarrow Mn^{2+}$

Balancing of reduction half

$Fe^{2+} \longrightarrow Fe^{3+} + 1e^-$ ---------- (1)

Balancing of reduction half

$MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$

$MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$

$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ ---------- (2)

Multiply equation (1) by 5 than add both equations.

$5Fe^{2+} \longrightarrow 5Fe^{3+} + 5e^-$

$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$

$5Fe^{2+} + MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

(iv) **$FeSO_4 + K_2Cr_2O_7 + H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$**

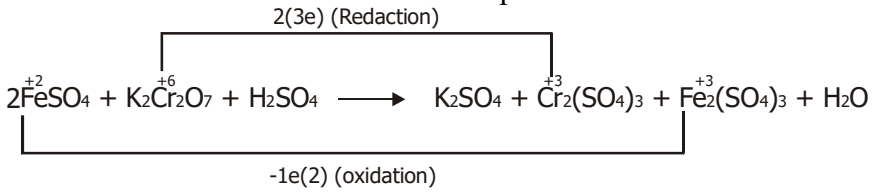
Write the oxidation number of each element.

$\overset{+2}{Fe}\overset{+6}{S}\overset{-8}{O_4} + \overset{+2}{K_2}\overset{+12}{Cr_2}\overset{-14}{O_7} + \overset{+2}{H_2}\overset{+6}{S}\overset{-8}{O_4} \longrightarrow \overset{+2}{K_2}\overset{+6}{S}\overset{-8}{O_4} + \overset{+6}{Cr_2}(\overset{+18}{S}\overset{-24}{O_4})_3 + \overset{+6}{Fe_2}(\overset{+18}{SO_4}\overset{-24}{)_3} + \overset{+2}{H_2}\overset{-2}{O}$

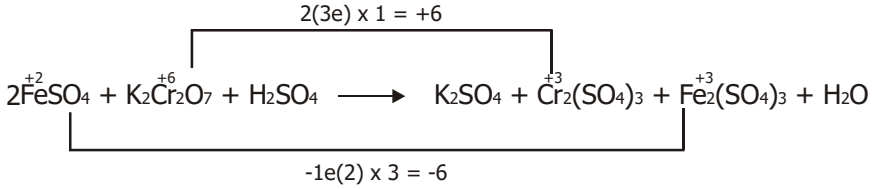
Identify the elements whose condition numbers have changed. Consider the change per atom.

$\overset{+2}{Fe}SO_4 + K_2\overset{+6}{Cr_2}O_7 + H_2SO_4 \longrightarrow K_2SO_4 + \overset{+3}{Cr_2}(SO_4)_3 + \overset{+3}{Fe_2}(SO_4)_3 + H_2O$

Mention the change with an arrow. Also balance the same atoms on both sides. Multiply (3e) by 2 on both sides because two atoms are present on both sides.



Multiple 2 by 3 to get a common number because gain of electrons is equal to loss of electron



Use multipliers as co-efficient will respective species.

$6FeSO_4 + K_2Cr_2O_7 + H_2SO_4 \longrightarrow 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$

Balance the rest of species by general inspection.

$\mathbf{6FeSO_4 + K_2Cr_2O_7 + 7H_2SO_4 \longrightarrow 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + K_2SO_4 + 7H_2O}$

Quick Check 10.4

i. Cu^{2+} and Br^-
ii. Cathode (Negative electrode)
iii. $Cu^{2+} + 2e^- \rightarrow Cu(s)$
iv. $2Br^- \rightarrow Br_2 + 2e^-$

Quick Check 10.5

a) Mass of Ag deposited:
$Q = I \times t = 0.18 \times 35 \times 60 = 378$ C
$Ag^+ + e^- \rightarrow Ag$; 1 mol = 96500 C = 108 g
378 C → $(378 \times 108) / 96500 \approx 0.42$ g
b) i. $Q = I \times t = 1.04 \times 6 \times 60 = 374.4$ C
ii. $H^+ + e^- \rightarrow H_2$ (1 mol = 96500 C)

Quick Check 10.6

i. Salt bridge completes circuit and maintains charge balance.
ii. $E°cell = 0.34 - (-0.44) = 0.78$ V

Quick Check 10.7

a) i. I_2 can be oxidized because Br_2 has higher E°.
ii. Br_2 cannot oxidize Cl^- because E° of Cl_2 is higher.

Quick Check 10.8

a)
The Nernst equation relates the electrode potential of a half-cell to the standard electrode potential, temperature, and the activities (or concentrations) of the chemical species involved.

It is given by:
$E = E° - (0.0591/n) \log([Red]/[Ox])$ at 25°C

Significance: The Nernst equation allows us to calculate the cell potential under non-standard conditions. It helps understand how cell potential changes with varying ion concentrations, which is crucial in designing and analyzing electrochemical cells.
b)
Variation in ion concentration affects the electrode potential of a half reaction. As the concentration of ions increases, the electrode potential changes according to the Nernst equation. Higher concentration of reduced species increases reduction potential, and higher concentration of oxidized species decreases it. Thus, the cell potential deviates from the standard potential.
c)
Given:
$E°(Ag^+/Ag) = +0.80$ V
$E°(Cr^{2+}/Cr) = -0.91$ V
$E°(Fe^{2+}/Fe) = -0.44$ V
1. i) Oxidizing Strength

The strongest oxidizing agent is the one with the highest (most positive) standard electrode potential.
Ag^+ (+0.80 V) > Fe^{2+} (−0.44 V) > Cr^{2+} (−0.91 V)
Therefore, Ag^+ is the strongest oxidizing agent, and Cr^{2+} is the weakest.
2. ii) Increasing Order of Reducing Power

Reducing power is opposite to oxidizing power. Metals with lower (more negative) E° values are stronger reducers.
Ag < Fe < Cr (increasing order of reducing power)
d)
Given:
$[Cu^{2+}] = 0.01$ M, $E°(Cu^{2+}/Cu) = +0.34$ V
Using Nernst Equation:
$E = E° - (0.0591/n) \log(1/[Cu^{2+}])$
n = 2 (since $Cu^{2+} + 2e^- \rightarrow Cu$)
$E = 0.34 - (0.0591/2) \log(1/0.01)$

$E = 0.34 - 0.02955 \times \log(100)$
$E = 0.34 - 0.02955 \times 2 = 0.34 - 0.0591 = 0.2809$ V
e)
Given: $[Cl^-] = 0.05$ M, $E°(Cl_2/Cl^-) \approx +1.36$ V
Reaction: $Cl_2 + 2e^- \rightleftharpoons 2Cl^-$
$E = E° - (0.0591/2) \log(1/[Cl^-]^2)$
$E = 1.36 - 0.02955 \times \log(1/0.0025)$
$E = 1.36 - 0.02955 \times \log(400)$
$\log(400) \approx 2.602$
$E = 1.36 - 0.02955 \times 2.602 \approx 1.36 - 0.0769 \approx 1.283$ V

Solution of the Exercise

Solutions of MCQ's (Q.1)

1	d	5	d	9	b
2	c	6	d	10	b
3	b	7	d	11	c
4	b	8	b	12	d

Solutions to Short Answer Questions (Q.2)

a. How and why electrical double layer is formed?

An electrical double layer forms at the interface between an electrode and an electrolyte due to the separation of charges. When an electrode is dipped in an electrolyte solution, ions from the solution get adsorbed on the electrode surface, causing a layer of opposite charges to accumulate in the solution near the electrode. This results in two parallel layers of opposite charges, known as the electrical double layer. This layer affects electrode potential and reaction kinetics.

b. Why is the electrode potential of Cu called reduction potential?

The electrode potential of Cu is referred to as the reduction potential because it represents the tendency of Cu^{2+} ions in solution to gain electrons and get reduced to metallic Cu. The standard electrode potential for Cu is measured under standard conditions and is positive, indicating a strong tendency to undergo reduction:
$Cu^{2+} + 2e^- \rightarrow Cu$

c. What are the advantages of salt bridge in a galvanic cell?

The salt bridge in a galvanic cell:
• Maintains electrical neutrality by allowing the flow of ions.

• Prevents the mixing of different solutions which can lead to precipitation.
• Completes the electrical circuit by connecting the two half-cells.

d. How can we predict the feasibility of a chemical reaction using the cell voltage?

The feasibility of a redox reaction is predicted using the standard cell potential (E°cell). If E°cell > 0, the reaction is spontaneous and feasible. It is calculated using:
$E°cell = E°cathode - E°anode$

e. During electrolysis of aqueous NaCl, why Na is not liberated at the cathode?

In aqueous NaCl electrolysis, H_2O is reduced instead of Na^+ due to its higher reduction potential.
Standard reduction potentials:
$Na^+ + e^- \rightarrow Na$; $E° = -2.71$ V
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$; $E° = -0.83$ V
Hence, H_2 gas is liberated at the cathode.

f. Calculate the Ox. No. of chromium (Cr) in the following compounds:
(i) $CrCl_3$ (ii) $Cr_2(SO_4)_3$ (iii) $Cr_2O_7^{2-}$

(i) $CrCl_3$: Let oxidation state of Cr = x
$x + 3(-1) = 0 \rightarrow x = +3$
(ii) $Cr_2(SO_4)_3$: Let oxidation state of Cr = x
$2x + 3(-2) = 0 \rightarrow x = +3$
(iii) $Cr_2O_7^{2-}$: Let oxidation state of Cr = x
$2x + 7(-2) = -2 \rightarrow x = +6$

g. Reactivity and reactions of metals:

Order of decreasing reactivity: K > Mg > Zn > Fe > Cu
i) $Cu + MgSO_4 \rightarrow$ No reaction (Cu is less reactive than Mg)
ii) $Fe + 2HCl \rightarrow FeCl_2 + H_2\uparrow$ (Fe displaces H_2 from acid as it is more reactive)

h. Explain why some metals higher in the activity series can displace hydrogen from acids, while others lower in the series cannot.

Metals higher in the activity series have a greater tendency to lose electrons and form positive ions (oxidation). Therefore, they can displace H^+ ions from acids, releasing hydrogen gas. Metals lower in the series have less tendency to be oxidized and hence cannot displace hydrogen.

i. Calculate Number of Faradays required to deposit 108 g of Ag^+, 63.5 g of Cu^{2+} and 27 g of Al^{3+}.

Using n = weight / (equivalent weight), where:
Equivalent weight = molar mass / n
Ag: 108 g / (108/1) = 1 F
Cu: 63.5 g / (63.5/2) = 2 F

Al: 27 g / (27/3) = 3 F
Total Faradays = 1 + 2 + 3 = 6 F

j. Electrolysis experiment and calculation of Avogadro's number:

Given:
I = 0.500 A, t = 30.0 min = 1800 s, m = 0.503 g, M(Ag) = 107.87 g/mol, z = 1
Using: $m = (M \times I \times t) / (n \times F)$
→ $F = (M \times I \times t) / (m \times z)$
→ $F = (107.87 \times 0.5 \times 1800) / (0.503 \times 1) = 193019.28$ C/mol
→ Avogadro's Number = F / charge on one electron = $193019.28 / (1.602 \times 10^{-19}) = 1.205 \times 10^{24}$ atoms/mol (approx.)

k. A cell is set up with a standard nickel electrode ($Ni^{2+} + 2e^- \rightleftharpoons Ni(s)$, $E° = -0.25$ V) and a standard cobalt electrode ($Co^{2+} + 2e^- \rightleftharpoons Co(s)$, $E° = -0.28$ V).

i) The metal with a higher (less negative) reduction potential acts as the cathode. Since
$E°(Ni^{2+}/Ni) = -0.25$ V and $E°(Co^{2+}/Co) = -0.28$ V,
Nickel will be the cathode (gets reduced), and Cobalt will be the anode (gets oxidized).

ii) Overall cell reaction:
$Co(s) + Ni^{2+}(aq) \rightarrow Co^{2+}(aq) + Ni(s)$

iii) Standard cell potential ($E°cell$) = $E°cathode - E°anode$ = $(-0.25\text{ V}) - (-0.28\text{ V}) = +0.03$ V

Solutions to Descriptive Questions (Q.3-6)

Q.3. How electrode potential varies with concentration of an aqueous solution? Use the NERNST equation to explain this variation.

The electrode potential varies with ion concentration in solution. According to the Nernst equation:
$E = E° - (0.0591/n) \log([Red]/[Ox])$
As the concentration of oxidized or reduced species changes, the value of log term changes, thus altering the potential.
For example, increasing the concentration of metal ion increases the reduction potential.

Q.4. How Avogadro's number can be derived using an electrolytic cell?

From Faraday's laws:
1 Faraday = 96500 C = 1 mole of electrons
By measuring the amount of substance deposited (mass), the current and time used, we calculate the total charge ($Q = I \times t$).
Knowing the number of atoms deposited per mole and using the relationship:
Avogadro's Number = Total Charge / Charge on one electron
$N_A = Q / e$
Where $e = 1.602 \times 10^{-19}$ C.

Q.5. Describe the construction and working principle of the Zn-Cu Galvanic cell.

Construction:
• Zinc electrode in $ZnSO_4$ solution (anode)
• Copper electrode in $CuSO_4$ solution (cathode)
• Salt bridge connecting the two half-cells

Working:
At anode (oxidation): $Zn(s) \rightarrow Zn^{2+} + 2e^-$
At cathode (reduction): $Cu^{2+} + 2e^- \rightarrow Cu(s)$
Electrons flow from Zn to Cu through external circuit, generating electricity.

Q.6. What is meant by Standard Hydrogen Electrode (SHE)? How it is used to measure the electrode potential of another electrode?

SHE is the reference electrode with 0 V potential under standard conditions:
• H_2 gas at 1 atm
• H^+ concentration = 1 M
• Platinum electrode
Reaction: $2H^+ + 2e^- \rightleftharpoons H_2(g)$

To measure potential of another electrode:
• Connect it with SHE in a galvanic cell
• Measure the EMF of the cell
• Determine whether the electrode acts as anode or cathode.
The EMF equals the standard electrode potential of the unknown electrode.

Solutions of Numericals (Q.7-9)

Q.7 Calculate the electrode potential for a zinc electrode immersed in a 0.010 mol dm^{-3} solution of zinc sulfate ($ZnSO_4$) at 298 K.

Given:
$E°(Zn^{2+}/Zn) = -0.76$ V
$[Zn^{2+}] = 0.010$ M
T = 298 K
n = 2
R = 8.31 J K^{-1} mol^{-1}
F = 96500 C mol^{-1}

Use the Nernst Equation:
$E = E° - (RT / nF) \times \ln([Zn^{2+}])$
$= -0.76 - (8.31 \times 298) / (2 \times 96500) \times \ln(0.010)$
$= -0.76 - (0.0128) \times (-4.605)$
$= -0.76 + 0.059$

= -0.701 V

Electrode potential = **−0.70 V** (approx.)

Q.8 A constant current of 2.00 A is passed through a solution of copper(II) sulfate ($CuSO_4$) for 30.0 minutes. Calculate the mass of copper deposited at the cathode.

Given:
I = 2.00 A, t = 30.0 min = 1800 s
F = 96500 C mol^{-1}, Molar mass of Cu = 63.5 g mol^{-1}, n = 2

Charge Q = I × t = 2 × 1800 = 3600 C
Moles of electrons = Q / F = 3600 / 96500 = 0.0373 mol
$Cu^{2+} + 2e^- \rightarrow Cu(s)$
Moles of Cu deposited = 0.0373 / 2 = 0.01865 mol
Mass of Cu = 0.01865 × 63.5 = **1.18 g**

Q.9 A galvanic cell consists of a standard hydrogen electrode (SHE) and a Ni^{2+}/Ni half-cell. The measured cell potential at 298 K is 0.25 V, and the nickel electrode is the negative terminal.

a) Cell reaction:
$Ni(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + H_2(g)$

b) $E°cell = E°cathode - E°anode$
Given: SHE (cathode) = 0 V, Ni electrode is anode

$E°(Ni^{2+}/Ni) = E°cathode - E°cell = 0 - 0.25 = **-0.25 V**$

c) Anode = Ni electrode (oxidation), Cathode = SHE (reduction)

CHAPTER # 11

HYDROCARBONS

Multiple Choice Questions

1. C
2. C
3. C
4. B
5. C
6. A
7. D
8. D
9. A
10. C
11. C
12. C

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a. Define the following:
i) Cycloalkanes ii) Isomerism iii) Conjugated dienes iv) Inductive effect
b. Differentiate between:
i) Aliphatic and Aromatic hydrocarbons
ii) Homolytic and Heterolytic Fission
iii) Electrophile and Nucleophile
c. Explain why alkanes do not undergo addition reactions.
d. How elimination reaction is considered the opposite of an addition reaction.
e. Compare the carbocation stability in propene and 2-butene.
f. Given the molecular formula C_5H_{10}, list all possible structural isomers that are alkenes.
g. When propene (C_3H_6) undergoes electrophilic addition with HBr, it forms 2-bromopropane as the major product. Explain why 2-bromopropane is favored over 1-bromopropane, using the concept of carbocation stability.
h. Explain why conjugated alkenes may show different reactivity compared to isolated alkenes.
i. Explain how inductive effects from alkyl groups stabilize carbocations in alkenes.
j. Write the equation for each reaction.
i. $CH_3CH_2CH=CH_2$ with H_2 (Ni catalyst)
ii. $CH_3CH=CH_2$ with Cl_2

iii. $CH_3CH_2CH=CHCH_2CH_3$ with H_2O (H_2SO_4 catalyst)
k. Write structural formulas for each of the following compounds.
i) Isobutylene ii) 2,3,4,4-Tetramethyl-2-pentene
iii) 2 ,5-Heptadiene iv) 4,5-Dimethyl-2-hexene
v) Vinylacetylene vi) 1,3-Pentadiene

vii) 1-Butyne viii) 3-n-Propyl-1, 4-pentadiene
ix) Vinyl bromide x) But-1 -en.3 -yne
xi) 4-Methyl-2-pentyne xii) Isopentane

I. Write down names of the following compounds according to IUPAC system:

i) $H_3C—CH=CH(CH_2)_2CH_3$

ii) $(CH_3)\ C=CH_2$

iii) $CH_3—CH_2—CH_2—C(CH(CH_3)_2)=CH_2$

iv) $CH_2=CH—CH=CH_2$

v) $CH_2=C(C_2H_5)—CH_2CH_2CH_3$

vi) $(CH_3CH_2)_3CH$

vii) $CH_3CH_2C(CH_3)_2CH(CH_2CH_3)CH_3$

Viii) $CH_2=CH—C\equiv C—CH=CH_2$

ix) $(CH_3)\ C—CH_2—C(CH_3)_3$

x) $CH_2=CH—C\equiv CH$

xi) $CH_3C(CH_3)_2(CH_2)CH_3$

a) $H_3C—CH_2—CH(H_2C—CH_3)—CH_2—CH_2CH_3$

b) $(CH_3)_2CH—CH(CH_3)—CH\ (CH_3)_2$

c) $H3C—CH—CH2—CH(CH_3)(CH_3)$

a. Definitions

i) Cycloalkanes:
Saturated hydrocarbons with carbon atoms arranged in a ring, containing only single bonds (e.g., cyclopentane).

ii) Isomerism:
The existence of two or more compounds with the same molecular formula but different structural or spatial arrangements.

iii) Conjugated Dienes:
Dienes with two double bonds separated by a single bond, allowing delocalization of π electrons (e.g., 1,3-butadiene).

iv) Inductive Effect:
The electron-withdrawing or -donating effect transmitted through sigma bonds due to electronegativity differences.

b. Differences

i) Aliphatic vs Aromatic Hydrocarbons:

Aliphatic	Aromatic
Open-chain or non-aromatic rings	Contain benzene ring or its derivatives
Undergo substitution and addition	Mostly substitution reactions

ii) Homolytic vs Heterolytic Fission:

Homolytic	Heterolytic
Each atom gets one electron	One atom gets both bonding electrons
Forms free radicals	Forms ions (cation + anion)

iii) Electrophile vs Nucleophile:

Electrophile	Nucleophile
Electron-deficient (accepts electrons)	Electron-rich (donates electrons)
E.g., H^+, Br^+	E.g., OH^-, NH_3

c. Why alkanes do not undergo addition reactions?

Alkanes are saturated hydrocarbons with strong sigma (σ) bonds and lack π bonds, making them unreactive toward addition reactions.

d. Elimination vs Addition Reactions

In **elimination**, atoms/groups are removed to form double/triple bonds; in **addition**, atoms/groups are added across double/triple bonds — opposite processes.

e. Carbocation Stability: Propene vs 2-Butene

The carbocation formed in 2-butene is more stable (3°) than in propene (2°), due to greater alkyl group stabilization via hyperconjugation.

f. Structural Isomers of C_5H_{10} (Alkenes)

1. Pent-1-ene
2. Pent-2-ene (cis & trans)
3. 2-Methylbut-1-ene
4. 2-Methylbut-2-ene
5. 3-Methylbut-1-ene

g. Propene + HBr → 2-Bromopropane (Major Product)

Follows **Markovnikov's rule**: H^+ adds to carbon with more H atoms → forms **2° carbocation**, which is more stable than 1° carbocation → gives **2-bromopropane**.

h. Reactivity of Conjugated vs Isolated Alkenes

Conjugated alkenes have **delocalized π electrons**, leading to **greater stability and unique reactivity**, such as resonance-stabilized intermediates.

i. Inductive Effects and Carbocation Stability

Alkyl groups donate electron density via the inductive effect, helping to stabilize the positive charge on a carbocation, especially in 2° and 3° carbocations.

j. Write Equations

i. $CH_3CH_2CH=CH_2 + H_2$ (Ni catalyst)
→ $CH_3CH_2CH_2CH_3$ (Butane)

ii. $CH_3CH=CH_2 + Cl_2$
→ $CH_3CHClCH_2Cl$ (1,2-Dichloropropane)

iii. $CH_3CH_2CH=CHCH_2CH_3 + H_2O$ (H_2SO_4 catalyst)
→ $CH_3CH_2CH(OH)CH_2CH_2CH_3$ (Hexan-3-ol)

k. Structural Formulas

i) Isobutylene: $CH_2=C(CH_3)_2$
ii) 2,3,4,4-Tetramethyl-2-pentene:
$CH_3C(CH_3)=C(CH_3)CH(CH_3)_3$
iii) 2,5-Heptadiene: $CH_3CH=CHCH_2CH=CHCH_3$
iv) 4,5-Dimethyl-2-hexene: $CH_3CH=CH_2CH(CH_3)CH_2(CH_3)$
v) Vinylacetylene: $CH_2=CH–C\equiv CH$
vi) 1,3-Pentadiene: $CH_2=CH–CH=CH–CH_3$
vii) 1-Butyne: $CH\equiv C–CH_2–CH_3$
viii) 3-n-Propyl-1,4-pentadiene: $CH_2=CH–CH(CH_2CH_2CH_3)–CH=CH_2$
ix) Vinyl bromide: $CH_2=CH–Br$
x) But-1-en-3-yne: $CH_2=CH–C\equiv CH$
xi) 4-Methyl-2-pentyne: $CH_3C\equiv C–CH(CH_3)CH_3$
xii) Isopentane: $(CH_3)_2CHCH_2CH_3$

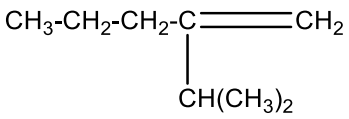
l. Write IUPAC names.

(i) $H_3C–CH=CH(CH_2)_2CH_3$

Name: **1-Hexene**

(ii) $(CH_3)\ C=CH_2$

Name: **propene**

(iii) 

Name: 2-isopropyl-1-pentene

(iv) $CH_2=CH–CH=CH_2$

Name: **1,3-Butadiene**

(v) $CH_2=C(C_2H_5)(CH_2CH_2CH_3)$

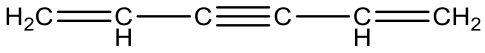
Name: 2-Ethyl-1-pentene

(vi) $(CH_3CH_2)_3CH$

Name: **3-Ethylpentane**

(vii) $CH_3CH_2C(CH_3)_2CH(CH_2CH_3)CH_3$

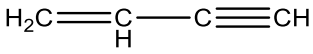
Name: 3,3,4-Trimethyl hexane

(viii) 

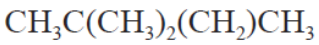
Name: **hexa-1,5-dien-3-yne**

(ix) $(CH_3)_3$ C–CH_2–$C(CH_3)_3$

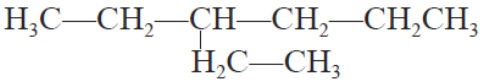
Name: 2,2,4,4-Tetramethylpentane

(x) 

Name: but-1-en-3-yne

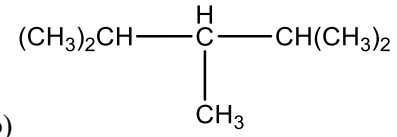
(xi) 

Name: 2,2-Dimethylbutane

a) 

Name: 3-Ethylhexane

b)



Name: 2,3,4-trimethylpentane

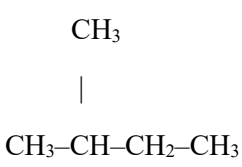
c) CH_3-CH-CH_2-$CH(CH_3)_2$

Name: 4-Methylpentane

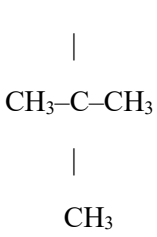
Quick Check 11.1

a) Write down the structural formulas for 2-methylbutane and 2,2-dimethylpropane.

1. 2-methylbutane

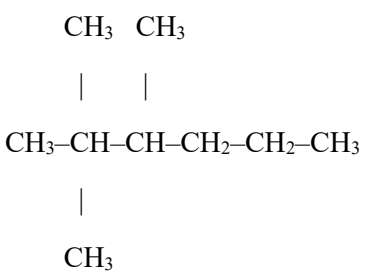


2. 2,2-dimethylpropane

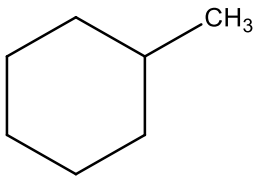


b) Write down the displayed formula of 2,3,3-trimethylhexane and methylcyclohexane.

1. 2,3,3-trimethylhexane



2. Methylcyclohexane



c) Give two differences between molecules of cyclopentane and pentane?

Feature	Cyclopentane	Pentane
Structure	Ring structure (5 carbon atoms in a ring)	Straight-chain structure (5 carbon atoms)
Molecular Formula	C_5H_{10} (due to ring – fewer hydrogens)	C_5H_{12} (saturated open-chain alkane)

d) Eicosane is a straight-chain alkane whose molecules contain 20 carbon atoms. What is the molecular formula of eicosane?

Molecular Formula of Eicosane

Eicosane is a straight-chain alkane with 20 carbon atoms.

The general formula for alkanes is:

C_nH_{2n+2}

For eicosane (n = 20)

$C_{20}H_{2(20)+2}=C_{20}H_{42}$

Quick Check 11.2

a) Why Do Branched Alkanes Have Lower Boiling Points Than Straight-Chain Alkanes?

1. Reduced Surface Area

- **Straight-chain alkanes** have a more extended structure, allowing **greater surface contact** between molecules.
- **Branched alkanes** are more compact and **spherical**, reducing the surface area for **intermolecular van der Waals forces** (London dispersion forces).

2. Weaker Intermolecular Forces

- With **less surface contact**, branched alkanes experience **weaker London dispersion forces**, which require **less energy (heat)** to overcome during boiling.
- Hence, they boil at **lower temperatures**.

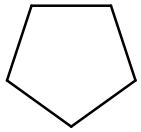
b) Explain why alkanes have a tetrahedral shape.

1. Carbon forms four single covalent bonds in alkanes (C–C and C–H).
2. In these bonds, each carbon atom uses sp^3 hybrid orbitals.
3. sp^3 hybridization creates four orbitals that point toward the corners of a tetrahedron, each separated by **109.5°** to minimize electron pair repulsion.

c) Draw the shapes of cyclopropane and cyclopentane.

1. cyclopropane

2. **cyclopentane**



Quick Check 11.3

a) Explain why alkanes have high stability?

Alkanes are considered highly stable compounds due to the following reasons:

1. Strong C–C and C–H Bonds

- Alkanes contain only single covalent bonds (sigma bonds), which are strong and stable.
- These bonds require a lot of energy to break, contributing to alkane stability.

2. Lack of Functional Groups

- Alkanes are non-polar and do not contain reactive functional groups like double bonds, hydroxyls, or halogens.
- This makes them less reactive under normal conditions.

3. Saturated Hydrocarbons

- Alkanes are saturated (each carbon has the maximum number of hydrogen atoms bonded).
- This saturation makes them less prone to addition reactions and more resistant to chemical attack.

b) What are the major types of reactions that alkanes undergo?

Although alkanes are relatively unreactive, they can undergo a few important types of reactions, primarily under specific conditions:

1. Combustion Reactions

- Alkanes burn in the presence of oxygen to produce carbon dioxide, water, and heat.
- Complete combustion (sufficient oxygen).

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

- Incomplete combustion (limited oxygen) produces carbon monoxide or carbon (soot).

2. Halogenation (Substitution Reaction)

- Alkanes react with halogens (e.g., Cl_2 or Br_2) under UV light or heat.
- Hydrogen atoms are replaced by halogen atoms:

$$CH_4 + Cl_2 \xrightarrow{UV} CH_3Cl + HCl$$

- This is a free radical substitution mechanism.

3. Cracking

- Long-chain alkanes are broken into shorter alkanes and alkenes by heating in the presence of a catalyst.
- Used in the petroleum industry to produce more useful fuels.

4. Isomerization

- Straight-chain alkanes are converted into branched isomers using a catalyst and heat.
- This improves the octane number of fuels.

c) How the termination step occurs in the halogenation of alkenes?

Termination Step

The termination step occurs when two free radicals combine to form a stable molecule, stopping the chain reaction.

Common Termination Reactions:

1. **Two halogen radicals combine**

$$\cdot Cl + \cdot Cl \rightarrow Cl_2$$

2. **Two alkyl radicals combine**

$$\cdot CH_3 + \cdot CH_3 \rightarrow C_2H_6$$

3. **An alkyl and a halogen radical combine**

$$\cdot CH_3 + \cdot Cl \rightarrow CH_3Cl$$

d) State the conditions under which a mixture of halogenoalkanes is obtained from the halogenation of alkanes.

1. **Presence of UV Light or Heat**
   - UV light or high temperature is required to initiate the free radical chain reaction by homolytic cleavage of the halogen molecule (e.g., $Cl_2 \rightarrow 2 \cdot Cl$).
2. **Excess of Alkane or Halogen**
   - Depending on the proportions:
     - Excess halogen leads to multiple substitutions (e.g., $CH_4 \rightarrow CH_3Cl \rightarrow CH_2Cl_2 \rightarrow CHCl_3 \rightarrow CCl_4$).
     - Excess alkane can reduce multiple substitutions but still may produce isomers.
3. **Use of Branched or Larger Alkanes**
   - If the alkane has more than one type of hydrogen (e.g., primary, secondary, tertiary), different isomers of halogenoalkanes will form due to the different stabilities of the intermediate radicals.

Resulting Mixture Includes:

- Mono-, di-, tri-halogenated alkanes (based on degree of substitution)
- Structural isomers (e.g., 1-chloropropane and 2-chloropropane)

e) Predict the products of the free radical chlorination of ethane.

Free radical chlorination of ethane (C_2H_6) in the presence of chlorine (Cl_2) and UV light or heat leads to substitution of hydrogen atoms with chlorine atoms.

Stepwise Substitution:

1. Monochlorination:

Only one hydrogen is replaced:

$$C_2H_6 + Cl_2 \xrightarrow{UV} C_2H_5Cl + HCl$$

Product: **Chloroethane** (C_2H_5Cl)

2. Dichlorination:

Two hydrogens replaced (can be on the same or different carbon atoms):

$$C_2H_5Cl + Cl_2 \xrightarrow{UV} C_2H_4Cl_2 + HCl$$

Possible isomers:

- 1,1-Dichloroethane

- 1,2-Dichloroethane

3. Further Substitution:

- Trichloroethane, tetrachloroethane, ... up to C_2Cl_6 (hexachloroethane) if excess chlorine is present.

Summary of Main Products:

Substitution Level	Product Name	Formula
Monosubstitution	Chloroethane	C_2H_5Cl
Disubstitution	1,1- or 1,2-Dichloroethane	$C_2H_4Cl_2$
Trisubstitution	Trichloroethane (various isomers)	$C_2H_3Cl_3$
Full substitution	Hexachloroethane	C_2Cl_6

Quick Check 11.4

a) Give the reactions of propanol with H_2SO_4, write down all the reactions involved.

a) Reactions of Propanol with Sulfuric Acid (H_2SO_4)

Dehydration to Form Alkene (Elimination Reaction)

Condition: Excess concentrated H_2SO_4, **170 °C**
Type: Elimination (E1 or E2 depending on the alcohol type)

Reaction (for **1-propanol**)

$$CH_3CH_2CH_2OH \xrightarrow[170^\circ C]{\text{conc. } H_2SO_4} CH_3CH = CH_2 + H_2O$$

Product: **Propene (alkene)**

2. Formation of Alkyl Hydrogen Sulfate (Substitution Reaction)

Condition: Lower temperature (~**20–50 °C**)
Type: Electrophilic substitution of –OH by $–OSO_3H$

Reaction (for **1-propanol**):

$$CH_3CH_2CH_2OH + H_2SO_4 \rightarrow CH_3CH_2CH_2OSO_3H + H_2O$$

Product: **Propyl hydrogen sulfate**

3. Formation of Ether (Symmetrical Etherification)

Condition: **Mild heating (~140 °C)** with excess alcohol
Type: Condensation reaction (intermolecular dehydration)

$$2CH_3CH_2CH_2OH \xrightarrow[140^\circ C]{H_2SO_4} CH_3CH_2CH_2OCH_2CH_2CH_3 + H_2O$$

Product: **Dipropyl ether**

4. Oxidation (only with hot conc. H_2SO_4 + oxidizing agent like $K_2Cr_2O_7$)

Not just H_2SO_4 alone—needs an oxidizer.
1-propanol can be oxidized to:

- **Propanal (aldehyde)** [mild oxidation]
- **Propanoic acid (carboxylic acid)** [strong oxidation]

$$CH_3CH_2CH_2OH + [O] \rightarrow CH_3CH_2COOH$$

b) Name the following alkenes

i) $CH_3C(CH_3)=CH-CH_3$ ii) $CH_2=CHCH(CH_3)_2$

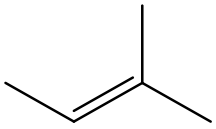
i) 2-methyl-2-butene

ii) 3-methyl-1-butene

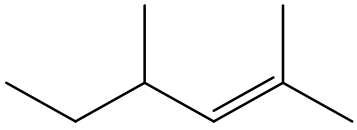
c) Draw the structural formulas of the following alkenes.

i) 3-Methyl-2-butene ii) 2-Methyl-4-ethyl-2-pentene

i)



ii)



d) Give the dehydrohalogenation reaction of bromopropane?

$$CH_3CH_2CH_2Br \xrightarrow[heat]{\text{alc. KOH}} CH_3CH = CH_2 + HBr$$

e) Can propanaol undergo dehydration? If yes give reactions involving dehydration.

1. 1-propanol

$$CH_3CH_2CH_2OH \xrightarrow{H_2SO_4,\,170^\circ C} CH_3CH = CH_2 + H_2O$$

2. 2-propanol

$$CH_3CHOHCH_3 \xrightarrow{H_2SO_4,\,170^\circ C} CH_2 = CHCH_3 + H_2O$$

Quick Check 11.5

a) How does an alkane differ from an alkene in stability?

Property	Alkanes	Alkenes
Saturation	Saturated hydrocarbons (single bonds)	Unsaturated hydrocarbons (double bonds)
Bond Type	Only **σ (sigma) bonds**	Contains one **π (pi) bond** + σ bonds
Stability	**More stable**	**Less stable** (due to π bond)
Reason	σ-bonds are strong and less reactive	π-bond is weaker and more reactive
Reactivity	Chemically less reactive	More reactive due to double bond
Tendency to react	Requires strong conditions to react	Easily undergoes addition reactions

Alkanes are more stable than alkenes because they only contain strong sigma (σ) bonds, whereas alkenes have a weaker pi (π) bond that is more reactive, making alkenes less thermally and chemically stable.

b) Which of the following species is likely to act as a nucleophile?

i) H_2 ii) H^+ iii) OH^-

OH- is the nucleophile

Explanation: A **nucleophile** is a species that donates an electron pair to form a new bond (i.e., "nucleus-loving"). **OH^-** fits this definition because it can **donate its lone pair** to electrophiles.

d) Which of the following species is likely to act as an electrophile?

i) H_2 ii) H^+ iii) OH

H^+ is an electrophile

Explanation: An **electrophile** is a species that **accepts an electron pair** (i.e., "electron-loving"). **H^+ (proton)** is highly electron-deficient and **readily accepts electrons**, making it a **strong electrophile**.

Quick Check 11.6

a) How do alkenes react with an electrophile?

Alkenes react with electrophiles via electrophilic addition, where the double bond breaks and new atoms/groups are added to the carbon atoms. The π bond acts as a nucleophile, allowing alkenes to react readily with electron-deficient electrophiles.

$$CH_2 = CH_2 + HBr \rightarrow CH_3CH_2Br$$

b) Why the order of stability of carbocation is 3° > 2° > 1°?

Reasons for This Stability Order:

1. Inductive Effect (Electron-Donating Effect of Alkyl Groups)

- **Alkyl groups** push electron density toward the positively charged carbon through the **+I effect**.
- More alkyl groups = more electron donation = greater stabilization.
- A 3° carbocation has **three alkyl groups**, so it's more stabilized than 2° or 1°.

2. Hyperconjugation

- Delocalization of electrons from **C–H σ bonds** of adjacent alkyl groups into the **empty p-orbital** of the carbocation.
- More alkyl groups = more hyperconjugation structures = more stable carbocation.

3. Delocalization and Charge Dispersion

- In 3° carbocations, the positive charge is spread out over several atoms, reducing **charge density** and increasing **stability**.

Carbocation Type	Structure	Relative Stability
Tertiary (3°)	$(CH_3)_3C^+$	Most stable
Secondary (2°)	$(CH_3)_2CH^+$	Moderate stability
Primary (1°)	$CH_3CH_2^+$	Least stable
Methyl	CH_3^+	Very unstable

Quick Check 11.7

a) Compare and explain the relative rates of addition to alkenes (reactivities) of HCl, HBr and HI

a) Comparison of the Relative Rates of Addition of HCl, HBr, and HI to Alkenes

Relative Reactivity Order:

$$HI > HBr > HCl$$

Explanation of the Trend:

1. Bond Strength (H–X bond dissociation energy)

- The H–X bond must break in the first step to release the proton (H^+).
- Weaker bonds break more easily, facilitating faster reaction.

Acid	H–X Bond Strength (kJ/mol)	Trend
HI	~299	Weakest bond → breaks easily
HBr	~366	Moderate bond
HCl	~431	Strongest bond → slowest

So, HI reacts fastest because it has the weakest H–I bond, making H^+ more available.

2. Polarity of the H–X Bond

- HCl is more polar than HI, which could help in bond cleavage.
- However, bond strength dominates over polarity in determining reactivity here.

3. Nucleophilicity of Halide Ion (X^-)

- I^- is a better nucleophile than Br^- or Cl^- due to its larger size and higher polarizability, so it stabilizes the intermediate more effectively.

Conclusion:

The rate of addition of hydrogen halides to alkenes follows the order:

$$HI > HBr > HCl$$

Because:

- HI has the weakest H–X bond, making proton release easier.
- I^- is a better nucleophile, facilitating faster completion of the reaction.

b) Explain the difference between an addition reaction and an elimination reaction?

Feature	Addition Reaction	Elimination Reaction
Definition	Two or more atoms/groups are **added** to a molecule	Atoms/groups are **removed** from a molecule
Reactants	Typically involves **unsaturated compounds** (alkenes, alkynes)	Typically involves **saturated compounds** with good leaving groups (alkyl halides, alcohols)

Products	Forms **one product** with **more atoms/groups**	Forms a **double/triple bond** or a **ring** by **loss of atoms/groups**
Bond Change	**π bond breaks** (C=C → C–C)	**π bond forms** (C–C → C=C)
Common Example	Alkene + HBr → Alkyl halide	Alcohol + H_2SO_4 → Alkene + H_2O
Type of Reaction	Increases saturation	Decreases saturation

Addition Reaction Example:

$$CH_2=CH_2+Br_2\rightarrow CH_2Br–CH_2Br$$

Elimination Reaction Example:

$$CH_3CH_2Br+\text{alc. KOH}\rightarrow CH_2=CH_2+HBr$$

c) How alkenes react with an electrophile?

Alkenes have a carbon–carbon double bond (C=C), consisting of a σ-bond and a π-bond. The π-bond is electron-rich, making it attractive to electrophiles (electron-deficient species).

Mechanism: Electrophilic Addition

Step 1: Attack of the Electrophile

- The π electrons of the double bond attack the electrophile (e.g., H^+), forming a carbocation intermediate.
- One carbon from the double bond bonds to the electrophile.

$$CH_2=CH_2+H^+\rightarrow CH_3–CH_2^+$$

Step 2: Nucleophilic Attack

- A nucleophile (e.g., Br^- or Cl^-) attacks the carbocation, forming the final addition product.

$$CH_3–CH_2^+ +Br^-\rightarrow CH_3CH_2Br$$

Example: Ethene + HBr

$$CH_2=CH_2+HBr\rightarrow CH_3CH_2Br$$

Alkenes react with electrophiles by breaking the π bond and adding new atoms/groups across the double bond. The reaction proceeds via a carbocation intermediate and is called electrophilic addition.

d) Explain how Markovnikove's rule is applied in addition of HBr to 2-pentene.

Markovnikov's Rule:

"In the addition of HX to an unsymmetrical alkene, the hydrogen (H^+) attaches to the carbon with more hydrogen atoms, and the halide (X^-) attaches to the carbon with fewer hydrogen atoms."

Structure of 2-Pentene:

$$CH_3–CH=CH–CH_2–CH_3$$

This is an unsymmetrical alkene, with the double bond between C-2 and C-3.

Addition of HBr to 2-Pentene:

Step 1: Protonation (H^+ adds to the alkene)

- According to Markovnikov's rule, H^+ adds to C-2 (the carbon with more H atoms).
- This forms a carbocation at C-3, which is more stable due to alkyl group support.

Step 2: Bromide ion (Br^-) attacks the carbocation

- Br^- adds to C-3, forming the major product.

Product Formed:

$CH_3–CHBr–CH_2–CH_2–CH_3$ **3-Bromopentane**

Quick Check 11.8

a) Write the equation for the ozonolysis of ethene.

$$C_2H_4 + O_3 \longrightarrow C_2H_4O_3 \longrightarrow 2CH_2O$$

b) Give the epoxidation reaction of 2-Butene.

$$CH_3–CH=CH–CH_3 + RCO_3H \rightarrow CH_3–CH(O)–CH(O)–CH_3$$

c) How ozonolysis can indicate the position of a double bond in an alkene? Explain with the help of an example.

Ozonolysis is used to determine the position of a double bond in an alkene by cleaving the C=C bond and forming carbonyl compounds (aldehydes or ketones).

General Reaction:

$$RCH=CHR'+O_3 \rightarrow RCHO+R'CHO \text{ (or ketone)}$$

The double bond is split, and each carbon of the double bond becomes the carbonyl carbon of a new compound. By analyzing the carbonyl products, you can reconstruct the original alkene and locate the C=C position.

Example: Ozonolysis of 2-Butene

$$CH_2 = CH{-}CH_2{-}CH_3 \xrightarrow{O_3/(Zn,H_2O)} CH_2O(formaldehyde) + CH_3{-}CH_2{-}CHO \text{ (propionaldehyde)}$$

Quick Check 11.9

i. Ethene is a monomer used in the polymerization process to create polyethylene.

(a) Discuss how the chemical structure of ethene makes it suitable for polymerization.
(b) Discuss the change in hybridization and bond angle during this reaction.

Why Ethene is Suitable for Polymerization:

Ethene has a reactive double bond (C=C) that can break and form new bonds with other ethene molecules, allowing it to link repeatedly into long chains (polyethylene).

Change in Hybridisation and Bond Angle:

Before polymerization: Carbon atoms are sp^2 hybridized, with 120° bond angles. After polymerization: Carbon atoms become sp^3 hybridized, with 109.5° bond angles.

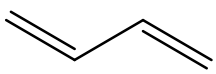
ii. Explain why an addition reaction increases the saturation of a molecule.

An addition reaction increases saturation by breaking a double or triple bond and adding atoms, turning it into single bonds, which makes the molecule more saturated and less reactive.

$$CH_2 = CH_2 + H_2 \rightarrow CH_3{-}CH_3$$

Quick Check 11.10

a) Draw the structure of simple conjugated dienes such as Buta-1,3-diene.



b) Illustrate the delocalization of electrons in conjugated dienes.

Quick Check 11.11

a) Draw the displayed formula and name an isomer of C_4H_8 that could be an example of positional isomerism?

One example is
But-2-ene
Displayed formula: $CH_3–CH=CH–CH_3$

b) What type of isomers are ethoxy methane ether and propanol?

They are functional group isomers

They have the **same molecular formula**: C_3H_8O

But they contain **different functional groups**:

Ethoxy methane → Ether group (–O–)

Propanol → Alcohol group (–OH)

c) What are the different types of structural isomerism in alkenes?

Alkenes can show the following **types of structural isomerism**:

1. Chain Isomerism

- Different **carbon chain arrangements**.
- Example:
  - **But-1-ene**: $CH_2=CH–CH_2–CH_3$
  - **2-Methylpropene**: $CH_2=C(CH_3)–CH_3$

2. Position Isomerism

- Same carbon chain, but **different position of the double bond**.
- Example:
  - **But-1-ene**: $CH_2=CH–CH_2–CH_3$
  - **But-2-ene**: $CH_3–CH=CH–CH_3$

3. Functional Group Isomerism *(less common in simple alkenes)*

- In larger molecules, alkenes may be isomers with **cycloalkanes**.
- Example:
  - **Butene (C_4H_8)** and **cyclobutane (C_4H_8)**

Conclusion:
Alkenes show **chain**, **position**, and occasionally **functional group** isomerism.

d) Write down different isomers of the compound $C_4H_9NH_2$.

Molecular Formula: $C_4H_9NH_2$
These are amine isomers with the same formula but different structures (both chain and position isomerism).

1. Primary (1°) Amines – R–NH_2

Butan-1-amine

$CH_3–CH_2–CH_2–CH_2–NH_2$

Butan-2-amine

$CH_3–CH_2–CH(NH_2)–CH_3$

2-Methylpropan-1-amine

$(CH_3)_2CH–CH_2–NH_2$

2-Methylpropan-2-amine

$(CH_3)_3C–NH_2$

2. Secondary (2°) Amines – R–NH–R' *(Not typically asked unless stated)*

But if included:

Methylpropylamine

$CH_3–NH–CH_2CH_2CH_3$

Ethylethylamine

$CH_3CH_2–NH–CH_2CH_3$

Conclusion:

There are four main structural isomers of $C_4H_9NH_2$ as primary amines, differing in chain branching and position of the $–NH_2$ group.

CHAPTER # 12

NITROGEN AND SULPHUR

Chapter no 12:

Short and Descriptive Questions

Q.2 Attempt the following short-answer questions:

a. Two reasons for the inertness of N_2:

1. N_2 has a triple covalent bond (N≡N), which is very strong and requires high energy to break.
2. It has a complete octet and lacks polarity, reducing its reactivity.

b. How is nitrogen isolated from air?

Nitrogen is obtained by fractional distillation of liquid air. Oxygen liquefies first (−183°C), followed by nitrogen (−196°C), allowing their separation.

c. Why ammonia (NH_3) is a weak base?

Ammonia has a lone pair on nitrogen that can accept protons (Bronsted base), but it's not fully ionized in water, so it is a weak base.

d. How does NO_2 catalyze the formation of SO_3?

NO_2 oxidizes SO_2 to SO_3 and itself gets reduced to NO. NO is then re-oxidized to NO_2 by O_2, making it a catalyst:

$SO_2 + NO_2 \rightarrow SO_3 + NO$

$2NO + O_2 \rightarrow 2NO_2$

e. Reactions of photochemical smog formation:

$NO_2 \rightarrow NO + O$

$O + O_2 \rightarrow O_3$

$O_3 + NO \rightarrow NO_2 + O_2$

f. Construction and function of catalytic converter:

The catalytic converter significantly reduces the emission of three major pollutants: carbon monoxide (CO), nitrogen oxides (NO_x), and unburnt hydrocarbons. Its advanced structure and the presence of catalytic metals ensure efficient chemical conversion, making modern vehicles much cleaner and environmentally friendly.
Reaction: $CO + 2NO \rightarrow N_2 + CO_2$

g. Why sulfur is unreactive at room temperature?

Sulfur exists as stable S_8 rings with strong S–S bonds and shows low reactivity at room temperature.

h. Most stable oxidation states of sulfur at pH = 0 and pH = 14:

At pH = 0: +6 (e.g., H_2SO_4)

At pH = 14: –2 (e.g., H_2S)

i. How does sulfur react with halogens?

Forms sulfur halides: $S + Cl_2 \rightarrow S_2Cl_2$ or SCl_2

j. Structures of S_8 and H_2SO_4:

S_8: Crown-shaped cyclic structure.

H_2SO_4: Tetrahedral with two –OH and two =O bonded to sulfur.

k. Role of sulfur in vulcanization of rubber:

Sulfur forms cross-links (S–S bonds) between polymer chains, improving elasticity and strength.

l. Composition and reaction of gunpowder combustion:

Gunpowder is formed by the combination of potassium nitrate, sulfur, and carbon.

$2KNO_3 + S + 3C \rightarrow K_2S + N_2 + 3CO_2$

m. Importance of disulfide bridges:
Disulfide bonds (–S–S–) stabilize protein structure by linking different parts of polypeptides.

n. Self-ionization of sulfuric acid

(i) In pure form (concentrated H_2SO_4):

Self-ionization occurs between molecules of sulfuric acid themselves: $2H_2SO_4 \rightleftharpoons H_3SO_4^+ + HSO_4^-$

In very concentrated or pure H_2SO_4, there's no water present. Hence, proton transfer occurs between sulfuric acid molecules. One molecule donates a proton (H^+), and the other accepts it.
- $H_3SO_4^+$ is a protonated form of sulfuric acid (a very strong acid).
- HSO_4^- is the hydrogen sulfate ion, a conjugate base.

(ii) In aqueous solution (in water):
$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$

Sulfuric acid is a diprotic acid, meaning it can donate two protons. In water:
- The first proton is fully ionized, giving H_3O^+ and HSO_4^-.
- The second ionization is partial:

$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$

The first dissociation is strong (complete), while the second is weak (incomplete), making H_2SO_4 a strong acid overall.

o. Sulfuric acid as a dehydrating agent (with two examples)

Sulfuric acid (H_2SO_4) has a strong affinity for water because it is hygroscopic—it absorbs water vigorously. This makes it a powerful dehydrating agent.

Example 1: Dehydration of glucose

$C_6H_{12}O_6 \rightarrow 6C + 6H_2O$

Glucose contains hydrogen and oxygen in a 2:1 ratio, just like water. Concentrated H_2SO_4 removes water elements from glucose, leaving behind elemental carbon (black). This reaction is highly exothermic.

Example 2: Dehydration of ethanol

$C_2H_5OH \rightarrow C_2H_4 + H_2O$ (with conc. H_2SO_4 and 170°C)

Ethanol loses a water molecule in the presence of hot concentrated H_2SO_4, forming ethene (C_2H_4). This is an elimination reaction where H_2SO_4 acts as both a catalyst and a dehydrating agent.

p. How V_2O_5 catalyzes the formation of SO_3

Contact process for making sulfuric acid involves:
$2SO_2 + O_2 \rightarrow 2SO_3$ (catalyzed by V_2O_5 at 450°C)
Vanadium(V) oxide (V_2O_5) is a solid catalyst used in the contact process. It provides an alternate pathway with lower activation energy for the reaction between SO_2 and O_2.
1. SO_2 reacts with $V_2O_5 \rightarrow V_2O_4 + SO_3$
2. V_2O_4 is deoxidized by $O_2 \rightarrow V_2O_5$
Thus, V_2O_5 is regenerated and acts as a true catalyst, speeding up the reaction without being consumed.

q. Purpose of forming Oleum

Oleum is also known as fuming sulfuric acid, and its formula is:
$H_2S_2O_7$ or ($H_2SO_4 \cdot SO_3$)

Formation:
$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$

Purpose and advantages:

1. Safe transport of SO_3:
SO_3 is volatile and reactive with water, forming corrosive acid mist. To avoid this, it is converted into oleum, which is a liquid and easier to handle safely.
2. Direct addition of sulfur trioxide (SO_3) to water is highly exothermic and leads to a violent

reaction that produces a dense, corrosive mist of sulfuric acid, which is difficult to condense and control. The intense heat causes rapid vaporization of water, resulting in splattering and potential safety hazards. To avoid this, industries first absorb SO_3 in concentrated sulfuric acid to form oleum ($H_2S_2O_7$), which can then be safely diluted with water to produce sulfuric acid. This method allows better control, prevents mist formation, and ensures safer handling and transportation of SO_3.

DESCRIPTIVE QUESTIONS

Q.3 Explain the preparation and basicity of ammonia.

Preparation:

Ammonia (NH_3) is prepared industrially by the Haber process, which involves the direct combination of nitrogen and hydrogen gases in the presence of an iron (Fe) catalyst at high temperature and pressure:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ (iron catalyst, 450°C, 200 atm)

This is a reversible reaction and is exothermic. According to Le Chatelier's principle, low temperature favors the forward reaction, but a compromise temperature of 450°C is used to achieve a reasonable yield and rate.

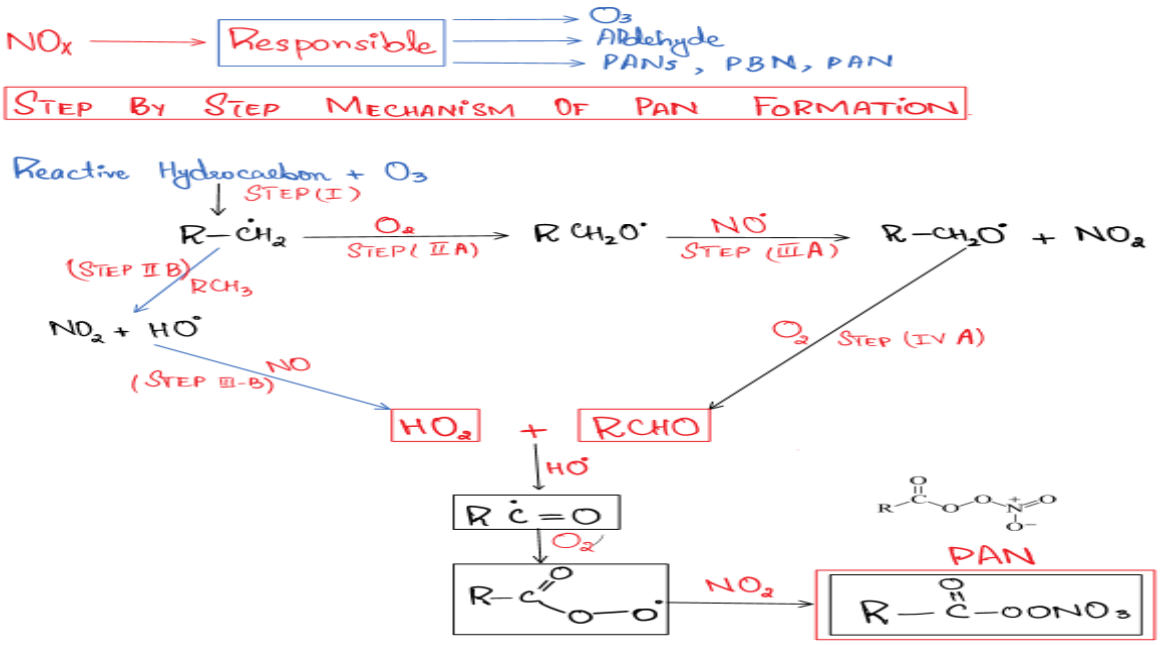
Basicity:

Ammonia is a weak base. It has a lone pair of electrons on the nitrogen atom, which allows it to accept a proton (H^+). In aqueous solution, ammonia reacts with water to form ammonium ion (NH_4^+) and hydroxide ion (OH^-):

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Since the equilibrium lies largely to the left, ammonia is classified as a weak base. Its basicity is due to the availability of the lone pair on nitrogen for coordination with protons.

Q.4 How oxides of nitrogen (NO_x) cause the formation of photochemical smog and PAN? Give its mechanism.



STEP I: Oxidation of Hydrocarbons

Reactive Hydrocarbon + O_3 → R–CH_2• (radical)

- A hydrocarbon (e.g., ethane, propane) reacts with **ozone (O_3)**.
- A hydrogen atom is abstracted, forming an **alkyl radical**: R–CH_2•

STEP II-A: Formation of Peroxy Radical

R–CH_2• + O_2 → R–CH_2–O• (peroxy radical)

- The radical formed in step I reacts rapidly with **molecular oxygen (O_2)**.
- This produces a **peroxy radical**.

STEP III-A: Reaction with NO

R–CH_2–O• + NO → R–CH_2–O• + NO_2

- The peroxy radical reacts with **NO** to give an **alkoxy radical (R–CH_2–O•)** and **NO_2**.

STEP IV-A: Further oxidation

R–CH_2–O• + O_2 → R–CHO + HO_2•

The **alkoxy radical** is oxidized by oxygen to form:

Aldehyde (R–CHO)

Hydroperoxyl radical (HO_2•)

STEP II-B (Alternative Pathway): Formation of Aldehyde & HO_2

$NO_2 + HO\bullet \rightarrow HO_2\bullet \rightarrow$ reacts with $RCH_3 \rightarrow RCHO + HO_2\bullet$

Hydroxyl radical **(OH•)** initiates the conversion of hydrocarbons into aldehydes (e.g., **R–CHO**) and **HO_2•**.

STEP III-B: Aldehyde → Acyl Radical

$RCHO + OH\bullet \rightarrow R–C\bullet=O + H_2O$

Aldehyde reacts with a hydroxyl radical **(OH•)** to form an **acyl radical (R–C•=O)**.

STEP IV-B: Peroxyacyl Radical Formation

$R–C\bullet=O + O_2 \rightarrow R–C(O)–OO\bullet$

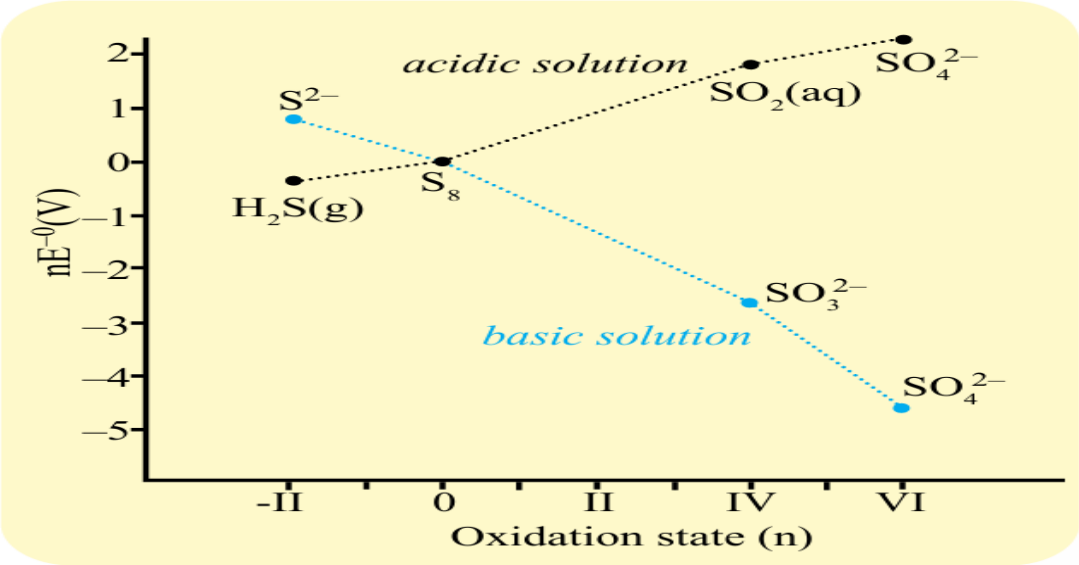
The **acyl radical** reacts with oxygen to form a **peroxyacyl radical**.

Final Step: PAN Formation

$R–C(O)–OO\bullet + NO_2 \rightleftharpoons R–C(O)–OONO_2$ (PAN)

The peroxyacyl radical reacts with **NO_2** to form **Peroxyacetyl Nitrate (PAN)**.

Q.5 Frost diagram and sulfur oxidation states:



The Frost diagram is a graphical representation used to evaluate the relative stabilities of different oxidation states of an element, in this case, sulfur. It plots the product of the number of electrons transferred and the standard electrode potential (nE^0) against the oxidation state (n). This plot is essentially proportional to the standard Gibbs free energy change ($-\Delta G^0$), as described by the equation $-\Delta G^0 = nFE^0$. A lower nE^0 value, or more negative ΔG^0, corresponds to greater

thermodynamic stability, meaning the oxidation state that lies at the lowest point in the Frost diagram is considered the most stable.

According to the Frost diagram for sulfur, the +6 oxidation state, represented by the sulfate ion (SO_4^{2-}), is the most thermodynamically stable under **basic conditions** (pH = 14). This is indicated by its position at the bottom of the diagram under those conditions. It is completely non-oxidizing in a basic environment, suggesting that it is highly resistant to further oxidation or reduction. However, under **acidic conditions** (pH = 0), the Frost diagram shows that all oxidized forms of sulfur, including SO_4^{2-}, become less stable than elemental sulfur (oxidation state 0), meaning that in acidic environments, elemental sulfur is more stable than its oxidized forms.

In conclusion, the most stable oxidation state of sulfur is +6 in strongly **basic** environments, while **elemental sulfur (oxidation state 0)** is the most stable form under strongly **acidic** conditions. This illustrates how pH influences the redox behavior and stability of sulfur compounds.

Q.6 Sulfuric acid as oxidizing and dehydrating agent (3 reactions): Sulfuric acid (H_2SO_4) serves as both a strong oxidizing agent and a powerful dehydrating agent, depending on the conditions and the reactants involved.

1. Oxidizing Agent:

$Cu + 2H_2SO_4$ (conc.) $\rightarrow CuSO_4 + SO_2 + 2H_2O$

Explanation: Concentrated sulfuric acid oxidizes copper metal to copper(II) sulfate while being reduced itself to sulfur dioxide gas. This is a redox reaction.

2. Dehydrating Agent (Carbohydrates):

$C_6H_{12}O_6 + H_2SO_4 \rightarrow 6C + 6H_2O$

Explanation: Sulfuric acid removes water (H and OH) from glucose, leaving behind elemental carbon as a black residue. The acid acts as a dehydrating agent, absorbing water formed during the reaction.

3. Dehydrating Agent (Alcohols):

$C_2H_5OH \rightarrow C_2H_4 + H_2O$ (with conc. H_2SO_4, 170°C)

Explanation: Ethanol loses a molecule of water when heated with concentrated sulfuric acid, forming ethene (C_2H_4). This is an elimination reaction where H_2SO_4 acts as a catalyst and dehydrating agent.

Chapter 13

Halogens

Quick Checks

Quick Check 13.1

a) Which halogens elements are radioactive?

Ans: Astatine (At) and tennessine (Ts)

b) What is the reason behind the different colours of halogens?

Ans: The trends of colour changes from chlorine to iodine is due to changes in the absorption of light as a result of electron transitions within the molecules of respective halogens.

c) Why chlorine is more volatile than bromine and iodine?

Ans: Generally, volatility decreases from chlorine to iodine. This trend is due to the increasing molecular mass, increase in size of outer shell and stronger intermolecular forces (London dispersion forces)

Quick Check 13.2

a) The F-F bond is weaker than Cl-Cl bond although fluorine is the most electronegative element, Explain.

Ans: In the case of fluorine, the bond strength is relatively weak because fluorine atoms are very small. Due to the small size, there is significant electron–electron repulsion between the lone pairs on the fluorine atoms, which weakens the bond despite the high electronegativity of fluorine.

b) Is the reaction between $NaCl_{(aq)}$ and F_2 gas possible?

i) Give reason whether yes or no.

Ans: Yes the reaction is possible because Fluorine has the highest tendency to acquire an electron and form fluoride. Fluorine molecule can oxidize and displace all the halide ions from their solutions (Cl^-, Br^- and I^-) to free halogens.

ii) If yes, write the equation for this reaction.

Ans: $F_{2(g)} + Cl_{(aq)}{}^{1-} \rightarrow F_{(aq)}{}^{1-} + Cl_{2(g)}$

c) What is the relationship between the oxidizing power of halogens and their standard reduction potential values?

Ans: The oxidizing power of halogens can be related to standard electrode potential (E^o) values. Fluorine is the most reactive halogen and the most powerful oxidizing agent. The standard electrode potential E°

(X^2/X^-) for halogens become less positive from fluorine to iodine. This reflects the decreasing oxidizing power with decreasing standard electrode potential.

Quick Check 13.3

a) The reaction between H2 and F2 is explosive but that between H2 and I2 is slow and reversible. Explain why.

Ans: The difference in reactivity between hydrogen and fluorine compared to hydrogen and iodine stems from the bond energies involved and the resulting activation energies of the reactions. Fluorine's high electronegativity and small atomic size lead to a very strong bond with hydrogen (HF) and a relatively weak F-F bond, making the reaction highly exothermic and explosive. In contrast, the I-I bond is much stronger, and the HI bond is weaker than HF, leading to a slower, less exothermic reaction.

b) Refer to Table 13.4 and 13.5 to predict which of the following reactions would be more exothermic.

Explain your answer.

H2 (g) + Cl 2 (g) 2HCl (g) and H2 (g) + Br 2 (g) 2HBr (g)

Ans: The reaction between hydrogen and chlorine will be more exothermic because H-Cl bond is stronger than H-Br bond. Additionally, chlorine is stronger oxidizing agent than bromine so it releases more energy while forming the bond with hydrogen.

b) How is the thermal stability of hydrogen halides related to their bond dissociation energies?

Ans: The thermal stability of hydrogen halides is decreased due to decrease in bond dissociation energies as given below:

HF> HCl> HBr>HI

HF is the most thermally stable hydrogen halide due to the strongest bond formation between hydrogen and fluorine and it is proved by its highest bond dissociation energy 569kJ/mole. and vise versa.

c) HF is the most thermally stable hydrogen halide. Give reasons.

Ans: HF is the most thermally stable hydrogen halide due to the strongest bond formation between hydrogen and fluorine and it is proved by its highest bond dissociation energy 569kJ/mole

Quick Check 13.4

a) F- is a weaker reducing agent than Cl-. Explain why.

Ans: F^- is a weak reducing agent because its small size and stable electronic configuration make it less likely to lose electrons.

b) What is the cause of the different solubilities of silver halides in aqueous ammonia?

Ans: The different solubilities of silver halides (AgCl, AgBr, and AgI) in aqueous ammonia are primarily due to the varying abilities of these halides to form stable complex ions with ammonia, specifically the $[Ag(NH_3)_2]^+$ complex. AgCl is soluble in dilute ammonia because it readily forms a stable, soluble complex ion with ammonia, $[Ag(NH_3)_2]^+$. AgBr is less soluble in ammonia than AgCl whereas AgI is practically insoluble in aqueous ammonia. The stability of the $[Ag(NH_3)_2]^+$ complex is not strong enough to overcome the strong lattice energy of AgI.

c) Write down the equation for the reaction of KI with Ag+ followed by aq. NH3. What would you observe at the completion of this reaction?

Ans: KI reacts with Ag+1 to form a precipitate of Siver iodide.

$$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

The AgI precipitate reacts with aqueous ammonia to form a soluble complex ion.

$$AgI(s) + 2NH_3(aq) \rightarrow [Ag(NH_3)_2]^+(aq) + I^-(aq)$$

Quick Check 13.5

a) How would KI react with conc. H2SO4. What does this reaction indicate about the reducing power of iodide?

Ans: KI reacts with Conc. H2SO4 to produce hydrogen iodide.

$$KI(s) + H_2SO_4(l) \rightarrow HI(g) + KHSO_4(s)$$

KI formed then reacts with Conc. H2SO4 to produce iodine with other by- products.

$$2HI(g) + H_2SO_4(l) \rightarrow I_2(s) + SO_2(g) + 2H_2O(l)$$

This reaction shows that iodine is a strong reducing agent.

b) Show that the reaction of Cl2 with cold and hot aqueous KOH is a disproportionation reaction.

Ans: See 13.10 of text book

c) HI acts as strong reducing agent. Explain it with chemical reactions.

Ans: See 13.5 (a).

Quick Check 13.6

a) Why HOCl is more effective disinfectant than OCl- to kill bacteria in water?

Ans: HOCl is more effective disinfectant due to its neutral charge which allows it to penetrate the cell walls of microorganisms easily. HOCl oxidizes essential cellular components which disrupt the cellular function leading to cell death.

b) What are the factors that affect disinfection of bacteria in water?

Ans: See 13.11.3 of text book

c) What are the primary active species in the chlorination of water? Give equation that shows their formation.

Ans: The primary active species in the chlorination of water are HCl and HOCl.

$$Cl_{2\,(aq)} + H_2O_{\,(l)} \leftrightarrow HOCl_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)}$$

Exercise

MULTIPLE CHOICE QUESTIONS

Q.1 Four choices are given for each question. Select the correct choice.

I. Which halogen molecule has the strongest bond?

a) F2 b) Br2 c) I2 ✓ d) Cl2

II. The volatility of the halogens (Group 17) generally _____ as you move down the group (from Fluorine to Iodine).

a) Increases b) Decreases ✓

c) Remains the same d) Fluctuates unpredictably

III. Which one of the following halogen molecules has strongest oxidizing power?

a) Br2 b) F2 ✓ c) I2 d) Cl2

IV. The decreasing thermal stability of the halogens down the group is primarily due to the:

a) Increasing electronegativity of the atoms.

b) Decreasing bond length between the halogen atoms.

c) Increasing atomic radius, leading to a weaker covalent bond.

d) Increasing strength of van der Waals forces ✓

V. Which one of the following halides has strongest reducing power?

a) F- b) Cl- b) Br- d) I- ✓

VI. Which statement about the reaction between halogens and hydrogen is correct?

a) Iodine reacts most vigorously with hydrogen.

b) Chlorine and hydrogen explode in darkness.

c) Fluorine combines explosively with hydrogen even in cold and dark conditions. ✓

d) Bromine and hydrogen do not react at all.

VII. How does the acidic strength of hydrogen halides change as you move down the group?

a) It remains constant.

b) It decreases from HF to HI.

c) It increases from HF to HI ✓

d) It fluctuates erratically.

VIII. Why is fluorine the most reactive halogen?

a) Bond length in the halogen molecule

b) Bond strength in the halogen molecule

c) Electronegativity of the halogen ✓

d) Number of electrons in the halogen molecule

IX. When aqueous silver nitrate is added to a solution containing bromide ions, a cream precipitate forms. What is the solubility of this precipitate in aqueous ammonia solution?

a) Soluble in dilute ammonia solution

b) Partially soluble in dilute ammonia solution

c) Insoluble in dilute ammonia solution ✓

d) Soluble only upon heating with ammonia

X. Concentrated sulfuric acid is added to solid sodium chloride. What is the initial observation?

a) Reddish-brown fumes are evolved.

b) A purple vapor is evolved.

c) Steamy white fumes of hydrogen chloride are evolved ✓

d) A black solid is formed.

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a. Which halogen is the least reactive, which is the most reactive? Give reason.

Ans: See 13.4 and 13.7 of text book

b. The ionic equation for a reaction is: $Br_2 + 2I^- \rightarrow Br^- + I_2$

Explain which species is oxidized in this reaction. Why?

Ans: Iodide ion is a strong reducing agent in this reaction as it gives electrons to bromine and gets oxidized itself.

d. What is role of London dispersion forces in the trend of volatility of halogens?

Ans: See 13.2 of text book

e. How does the reactivity of halogens with hydrogen vary?

Ans: See 13.5 of text book

f. Which halogen is used as an antiseptic? How does it work?

Ans: See 13.11 of text book

g. What is the colour change when chlorine displaces bromine?

Ans: When chlorine displaces bromine, the color change is from colorless to orange or brown. This is because chlorine is more reactive than bromine and will displace it from a solution of a bromide salt, like potassium bromide. The displaced bromine will then be in its elemental form (Br2), which has an orange-brown color.

h. How are the halogen acids ionized in water?

Ans: Halogen acids (like HCl, HBr, and HI) ionize in water by dissociating into hydrogen ions (H^+) and halide ions (X^-) due to the polar nature of both the acid and water molecules. The water molecules surround and solvate these ions, effectively stabilizing them in solution. This process is driven by the strong attraction between the polar water molecules and the charged ions produced.

i. Why HF is weaker acid than HCl?

Ans: HF is a weaker acid than HCl primarily due to the stronger H-F bond and the influence of hydrogen bonding. The H-F bond is shorter and stronger than the H-Cl bond, making it more difficult to break and release a proton (H+). Additionally, hydrogen bonding in HF further stabilizes the molecule and hinders dissociation in solution.

j. Describe a simple chemical test that could be used to distinguish between aqueous solutions of potassium bromide and potassium iodide. Include the reagents and expected observations.

Ans: See 13.8 from text book

k. Explain the chemical principles behind the use of chlorine as a disinfectant in water purification. Include relevant chemical equations in your explanation.

Ans: See 13.11 of text book

l. Describe one significant disadvantage associated with the use of chlorine in water purification.

Ans: Chlorine is commonly used as a disinfectant in public water supplies to kill harmful bacteria and other microbes. This safeguards our health from waterborne diseases. However, the downside is that chlorine in drinking water can also have potential health risks if the levels are too high.

m. What is disproportionation reaction? Give an example.

Ans: See 13.10 of textbook

n. Chlorine gas reacts differently with sodium hydroxide solution depending on the temperature and concentration.

Ans: See 13.10 of textbook

o. Write balanced chemical equations for the reaction of chlorine (Cl2) with:

i) Cold, dilute sodium hydroxide (NaOH).

ii) Hot, concentrated sodium hydroxide (NaOH).

Ans: See 13.10 of text book

p. For each reaction in question l), identify the oxidation states of chlorine in the reactant (Cl2) and in each of the chlorine-containing products. Use these oxidation states to explain why both reactions are classified as disproportionation reactions.

Ans: See 13.10 of text book

DESCRIPTIVE QUESTIONS

Q.3 Describe and explain the relative thermal stabilities of the halogen hydrides in terms of bonds strength.

Q 4. Discuss the relative reactivity of the halogen elements as oxidizing agents. Arrange F_2, Cl_2, Br_2, I_2 in increasing order of the oxidizing power.

Q 5. Describe the reactions that occur when chlorine is bubbled through

(i) Cold and (ii) Hot, aqueous sodium hydroxide (NaOH**).**

Q.6 Discuss the reducing power of halide ions with relevant reactions. Also explain the factors affecting it.

CHAPTER # 14

ATMOSPHERE

Quick Check 14.1

a) Mention important Air Pollutants.

Ans. The important air pollutants include: Carbon Monoxide, NO and NO_2, SO_2 and SO_3 etc.

b) Give the equations for the formation and depletion of ozone in the stratosphere.

Ans. Formation of ozone takes place in the middle stratosphere where UV rays are present. Here monoatomic and diatomic oxygen recombine to make Ozone O_3. It is an exothermic reaction.

$$O_2 + O \longrightarrow O_3$$

c) Write down the names and approximate height of different layers of atmosphere.

Layer	Altitude range
Troposphere	Earth surface to 12Km
Stratosphere	12Km to 50Km
Mesosphere	50Km to 85Km
Thermosphere	85Km to 600Km

Quick Check 14.2

a) Mention man-made sources of Air Pollution.

Ans. They include vehicles means cars, bikes, trucks air crafts etc. and factors. Fossil fuels (Wood, Petrol, Diesel, Kerosene, Coal) burning is also a big source.

b) How do polycyclic aromatic hydrocarbons (PAHs) primarily enter the atmosphere

Ans. Natural Sources: some PAHs in the environment originate from natural source such as open burning, Natural loss or seepage of petroleum and coal deposits

Human Activities: They are generated during the incomplete combustion of fossil fuels, vehicle emissions and industrial processes.

c) How do volatile organic compounds (VOC) affect air quality?

Ans. Exposure to VOCs can cause a variety of health effects including throat, eye and nose irritation. Headache and nausea are short term exposures and damage to liver, kidney and central nervous system are long term exposures.

c) What are the major sources of heavy metals in the atmosphere?

Ans. They come into atmosphere from various industrial processes, transportation and other human activities. Metallurgy, battery wastes and incineration are the major sources.

Quick Check 14.3

a) Differentiate classical and photochemical smog?

Ans.

Classical Smog	**Photochemical Smog**
Main source is burning coal, industrial emissions	Main source is vehicle exhaust, industrial fumes, VOCs
SO_2, Carbon Soot, H_2SO_4 are the key pollutants	No, NO_2, O_3 and PAN are the main pollutants.

b) Name Green House Gases (GHGs). How do these gases cause global warming?

Ans. Geen House gases include CO_2, CH_4, NO and water vapors. The short wavelength UV and Visible radiations pass through the earth atmosphere and heat it. At night the hot earth surface emits long wave length Infra-red radiations which are mostly absorbed by green house gases. So, these IR radiations don't reach into space and are caught up. This increases the temperature of earth atmosphere. It is Green House effect or Global warming.

c) Write a balanced equation including state symbols, showing the formation of sulphuric acid from atmospheric Sulphur trioxide SO_3.

Ans.

$$O_2 + S \longrightarrow SO_2$$

$$2SO_2 + O_2 \longrightarrow 2SO_3$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

d) How HNO_3 is formed from NO in the atmosphere?

Ans.

$$2NO + O_2 \longrightarrow 2NO_2$$

$$2NO_2 + H_2O \longrightarrow HNO_2 + HNO_3$$

Quick Check 14.4

a) Explain the impact of particulate matter (PM) on the air quality.

Ans. Particulate matter degrades the air quality by causing haze, reducing visibility and forming smog. It harms human health leading to respiratory and cardiovascular diseases.

b) What is AQI? How does it measure air quality?

Ans. Air Quality Index (AQI) is a scale (0-500) that reports daily air pollution levels and linked health risks. It tracks major pollutants like Pm and Ozone. Grater the AQI means dangerous conditions.

c) What are different levels of AQI. Mention the safest and most hazardous ranges of AQI

Ans.

AQI value	Levels of health concern
0 to 50	Good
51-100	Moderate
101-150	Unhealthy for sensitive groups
151-200	Unhealthy
201-300	Very unhealthy
301-500	Hazardous

SOLUTION OF EXERCISE:

Answers of MCQs

Q.#	Answers
1.	C
2.	d
3.	c
4.	c
5.	b
6.	b
7.	b
8.	b
9.	b
10.	c
11.	b

SOLUTION OF SHORT QUESTIONS:

Q.2(a) Identify and briefly explain three major sources of air pollutants.

Ans. Air pollutants: Three major air pollutants are as follows;

Particulate Matter: Naturally occurring particulate matter includes dust from earth's surface. Pollens, spores and animal debris also produces PM.

Oxides of Nitrogen: Thunders in the sky produces significant amounts of oxides of nitrogen in the air.

Hydrogen Sulphide: algae on the surface of oceans which gives hydrogen sulphide gas.

Q.2(b) How can deforestation impact air quality?

Ans. Deforestation increases CO_2 levels in air which may cause green house effect. It reduces oxygen production in the air. It releases lot of air pollutants.

Q.2 (c) Explain the reasons for the temperature trends observed in the troposphere and stratosphere.

Ans. Troposphere: In troposphere, the temperature decreases with altitude. (17°C to -58°C) due to decrease in the concentration of greenhouse gases.

Stratosphere: In stratosphere, the temperature increases with altitude. (-58°C to -2°C) because ozone absorbs UV rays in the middle stratosphere.

Q.2(d) Describe four significant anthropogenic (caused by humans) activities that contribute to the deterioration of the air quality. For each activity name atleast one major pollutant released.

Ans.

ii. Vehicle Emissions causing Smog and acid rains.
iii. Factories emitting oxides of Sulphur and nitrogen contributing to air pollution.
iv. Coal and fossil fuel burning
v. Agricultural burnings

Q.2(e) What are the environmental impacts of Persistent Air Pollutants (POP's)?

Ans. POPs accumulate in the ecosystems poisoning wild life and entering the food chain. They resist degradation causing long term contamination of soil and water.

Q.2(f) How Polycyclic aromatic hydrocarbons affect human health?

Ans. They are carcinogenic and cause lung and bladder cancer. They cause respiratory and cardiovascular diseases.

Q.2(g) What is photochemical smog, under what conditions it is formed?

Ans. It is like a harmful brownish haze in the atmosphere usually seen in the evening time. It contains, Nox, VOC's, PANs and Ozone means ita has oxidizing agents. It formed in the following conditions.

i. Sunlight, Intense UV rays,
ii. High Nox and VOcs emissions
iii. Warm temperature
iv. Stagnant air means low wind

Q.2(h) What type of data do air quality index systems provide?

Ans. AQI systems provide levels of $PM_{2.5}$, ozone and NO2. The values range from 0 to 500 marking them from Good to hazardous with color codes. They also indicate health risks.

Q.2(i) Distinguish between PM_{10} and $PM_{2.5}$. specifying the size, range and describing why PM2.5 is considered more harmful to human health.

Ans. PM_{10} Contains bigger particles like dust and pollen while $PM_{2.5}$ consists finer particles like smoke and soot. These particles penetrate deep into lungs and blood increasing risk of respiratory and cardiovascular diseases.

Q.2(j) What are the main chemical processes involved in the formation of acid rains.

Ans. Production of Sulphur Oxides

$$S + O_2 \rightarrow SO_2$$

$$2SO_2 + O_2 \rightarrow 2SO_3$$

Production of nitrogen oxides

$$N_2 + O_2 \rightarrow 2NO$$

$$2NO + O_2 \rightarrow 2NO_2$$

Formation of acids

$$SO_3 + H_2O \rightarrow H_2SO_4$$

$$2NO_2 + H_2O \rightarrow HNO_2 + HNO_3$$

Q.2(k) What are the Specific measures to control smog?

Ans. These include as follows;

- Vehicular emission controls
- Industrial effluents regulation laws enforcement
- Public awareness
- Technological solutions

Q.2(l) How does a catalyst convertor reduces harmful vehicle emissions?

Ans. The most cars are equipped with catalyst converter to convert harmful pollutants to harmless substances. Unburnt hydrocarbon and CO is oxidized to CO_2, NO_2 reduced to N_2.

Q.2(m) Describe the sources of Lead and mercury emissions

Ans. These include transportation, metallurgical operations, battery waste and incinerations are the major sources.