ACTC ADVANCED CHEMISTRY TUITION CENTRE, PWD ROAD, NAGERCOIL 9940847892

"May God's guidance be with you during the Exam and may you be able to answer each question correctly. My prayers and Blessings are with you". - ACTC EMS

"Everything is chemistry, without chemistry Nothing"

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CLASS 12 VOLUME 1 & II **(WITH TEXT BOOK PAGE NUMBER)**

UNIT 1 BASIC CONCEPTS OF CHEMISTRY AND CHEMICAL CALCULATIONS

Answer the following

1. **Define relative atomic mass. (4)**

The relative atomic mass is defined as ***the ratio of the average atomic mass factor to the unified atomic mass unit.***

$$\text{Relative atomic mass } (A_r) = \frac{Average\ mass\ of\ the\ atom}{Unified\ atomic\ mass}$$

2. **What do you understand by the term mole. (6)**

- One mole is **the amount of substance of a system**, ***which contains as many elementary particles as there are atoms in 12 g of carbon -12 isotope.***
- The elementary particles can be molecules, atoms, ions, electrons or any other specified particles.

3. **Define equivalent Mass. (8)**

Equivalent mass of an element, compound or ion is the number of parts of mass of an element combines or displaces 1.008 g hydrogen or 8 g oxygen or 35.5g chlorine.

4. **What do you understand by the term oxidation number.**

- It is defined as ***the imaginary charge left on the atom when all other atoms of the compound have been removed in their usual oxidation states that are assigned according to set of rules.***
- Oxidation number refers to the number of charges an atom would have in a molecule or an ionic compound, if electrons were transferred completely. The oxidation numbers reflect the number of electrons transferred.

5. **Distinguish between oxidation and reduction**

	Oxidation	**Reduction**
(i)	Addition of oxygen	Additional of hydrogen
ii	removal of hydrogen	removal of oxygen

(iii)	This process involves loss of electrons $Fe^{2+} \rightarrow Fe^{3+} + e^-$	This process involves gain electrons. $Cu^{2+} + 2e^- \rightarrow Cu$
(iv)	Oxidation number increases	Oxidation number decreases

6. Calculate the molar mass of the following compounds (i) urea [$CO(NH_2)_2$] (ii) acetone [CH_3COCH_3] (iii) boric acid [H_3BO_3] (iv) sulphuric acid [H_2SO_4]

(i) urea [$CO(NH_2)_2$]

C : 1 × 12 = 12
O : 1 × 16 =16
N : 2 × 14 = 28
H : 4 × 1 = 4
60g mol^{-1}

(ii) acetane [CH_3COCH_3]

C : 3 × 12 = 36
H : 6 × 1 = 6
O : 1 × 16 = 16
58 g mol^{-1}

(iii) boric acid [H_3BO_3]

H : 3 × 1 = 3
B : 1 × 11 = 11
O : 3 × 16 = 48
62g mol^{-1}

(iv) sulphuric acid [H_2SO_4]

H : 2 × 1 = 2
S : 1 × 32 = 32
O : 4 × 16 = 64
98 g mol^{-1}

7. **The density of carbon dioxide is equal to 1.965 kgm^{-3} at 273 K and 1atm pressure. Calculate the molar mass of CO_2.**

Given : The density of CO_2 at 273K and 1 atm pressure = 1.965 kgm^{-3} At 273 K and 1 atm pressure, 1 mole of CO_2 occupies a volume of 22.4L

$$\text{Mass of 1 mole of } CO_2 = \frac{1.965Kg}{1m^3} \times 22.4L = \frac{1.965 \times 10^3\, g \times 22.4 \times 10^{-3} m^3}{1m^3}$$

Molar mass of CO_2 = **44 $gmol^{-1}$**

ANOTHER METHOD:

Molecular mass = Density x Molar mass

Molar volume of CO_2 = 2.24 x 10^{-2} m^3

Density of CO_2 = 1.965kgm^{-3}

Molecular mass of CO_2 = 1.965 x 10^3gm^{-3} x 2.24 x $10^{-2}m^3$
=1.965 x 10^1 x 2.24
= 44g

8. **Which contains the greatest number of moles of oxygen atoms (i) 1 mol of ethanol (ii) 1 mol of formic acid (iii) 1 mol of H_2O**

Answer: 1 mole of formic acid contains the greatest number of oxygen atoms.

Compound	Given no. of moles	No. of oxygen atoms
Ethanol (C_2H_5OH)	1	$1 \times 6.022 \times 10^{23}$
Formic acid (HCOOH)	**1**	$\mathbf{2 \times 6.022 \times 10^{23}}$
Water (H_2O)	1	$1 \times 6.022 \times 10^{23}$

9. **Calculate the average atomic mass of naturally occurring magnesium using the following data**

Isotope	Isotopic atomic mass	Abundance (%)
Mg^{24}	**23.99**	**78.99**
Mg^{26}	**24.99**	**10.00**
Mg^{25}	**25.98**	**11.01**

$$\text{Average atomic mass} = \frac{(78.99 \times 23.99)+(10 \times 24.99)+(11.01 \times 25.98)}{100}$$

$$= \frac{2430.9}{100} = 24.31\text{u.}$$

10. **In a reaction $x + y + z_2 \rightarrow xyz_2$ identify the Limiting reagent if any, in the following reaction mixtures**

(a) 200 atoms of x + 200 atoms of y + 50 molecules of z_2

(b) 1 mol of x + 1 mol of y + 3 mol of z_2

(c) 50 atoms of x + 25 atoms of y + 50 molecules of z_2

(d) 2.5 mol of x + 5 mol of y + 5 mol of z_2

Reaction : $x + y + z_2 \rightarrow xyz_2$

Q	Number of moles of reactants allowed to react			Number of moles of reactants consumed during reaction			Limiting reagent
	X	Y	z_2	X	y	z_2	
a	200 atoms	200atoms	50 molecules	50 atoms	50atoms	50 molecules	z_2
b	1 mol	1mol	3 mol	1 mol	1 mol	1mol	x and y
c	50 atom	25atom	50 molecules	25atom	25atom	25 molecules	y
d	2.5 mol	5 mol	5 mol	2.5mol	2.5 mol	2.5 mol	x

11. **Mass of one atom of an element is 6.645×10^{-23}g. How many moles of element are there in 0. 320 kg.**

Given: Mass of one atom = 6.645×10^{-23} g

∴ mass of 1 mole of atom $= 6.645 \times 10^{-23} \text{ g} \times 6.022 \times 10^{23}$

$= 40\text{g}$

∴ number of moles of element in 0.320 kg

Number of moles = mass/molar mass

$$= \frac{320g}{40g} = 8 \text{ mol.}$$

12. What is the difference between molecular mass and molar mass? Calculate the molecular mass and molar mass for carbon monoxide.

	Molecular mass	**Molar mass**
1	Molecular mass is defined as ***the ratio of the mass of a molecule to the unified atomic mass unit***	Molar mass is defined as the ***mass of one mole of a substance.***
2	The relative molecular mass of any compound is calculated by adding the relative atomic masses of its constituent atoms.	The molar mass of a compound is equal to the sum of the relative atomic masses of its constituents.
3	Its unit is ***u or amu***	Its unit is ***g mol^{-1}***
4	Molecular mass of CO: (1× 12 amu) + (1 × 16 amu) = 28 amu	Molar mass of CO: (1×12)+ (1× 16) = 28g mol^{-1}

13. What is the empirical formula of the following?

(i) Fructose ($C_6H_{12}O_6$) found in honey

(ii) Caffeine ($C_8H_{10}N_4O_2$) a substance found in tea and coffee.

Compound	**Molecular Formula**	**Empirical Formula**
Fructose	$C_6H_{12}O_6$	CH_2O
Caffeine	$C_8H_{10}N_4O_2$	$C_4H_5N_2O$

14. The reaction between aluminium and ferric oxide can generate temperatures up to 3273 K and is used in welding metals. (Atomic mass of Al = 27u Atomic mass of O = 16u) $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$; If, in this process , 324 g of aluminium is allowed to react with 1.12 kg of ferric oxide. (i) Calculate the mass of Al_2O_3 formed. (ii) How much of the excess reagent is left at the end of the reaction?

Given: $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

	Reactants		**Products**	
	Al	**Fe_2O_3**	**Al_2O_3**	**Fe**
Stoichiometric Co-efficient	2	1	1	2
Amount of reactant allowed to react	324 g	1.12 kg	-	-
Number of moles allowed to react (No of moles = mass/molar mass)	$\frac{324}{27} = 12\ mol$	$\frac{1120}{160} = 7mol$	-	-
Number of moles consumed during reaction ratio (2:1)	12 mol	6 mol	-	-
Number of moles of **reactant unreacted** and number of **moles of product** formed	-	1mol	6mol	12mol

i) Molar mass of Al_2O_3 formed

$= 6\ \text{mol} \times 102\ \text{g mol}^{-1} = 612\text{g}$

$$\left[\begin{matrix} Al_2O_3 \\ (2\times27)+(3\ \times16) \\ 54+48=102 \end{matrix}\right]$$

ii) Excess reagent = Fe_2O_3

Amount of excess reagent left at the end of the reaction = 1mol × 160 g mol^{-1} = 160g

$$\left[\begin{matrix} Fe_2O_3 \\ (2\times56)+(3\ \times16) \\ 112+48=160 \end{matrix}\right]$$

15. **How many moles of ethane is required to produce 44 g of $CO_{2(g)}$ after combustion.**

Balanced equation for the combustion of ethane

$$C_2H_6 + 3\tfrac{1}{2}\ O_2 \rightarrow 2CO_2 + 3H_2O$$

$$\Rightarrow\ 2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

To produce 4 moles of CO_2, 2 moles of ethane is required

To produce 1 mole (44 g) of CO_2 required number of moles of ethane

$$= \frac{2\ mol\ ethane}{4\ mol\ \ CO_2} \times 1\ mol\ CO_2 \quad = \tfrac{1}{2}\ \text{mole of ethane} = 0.5\ \text{mole of ethane}$$

16. Hydrogen peroxide is an oxidising agent. It oxidises ferrous ion to ferric ion and reduced itself to water. Write a balanced equation.

$H_2O_2 \rightarrow$ Oxidising agent

-1 -2

$H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + H_2O$ (Acidic medium)

$(1e^{-} \times 2)$ $1e^{-}$

Ferrous ion is oxidized by H_2O_2 to Ferric ion.

The balanced equation is $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$ X 2

$H_2O_2 + 2H^{+} + 2e^{-} \rightarrow 2H_2O$

$\underline{2Fe^{2+} \qquad\qquad \rightarrow\ 2Fe^{3+}\ + 2e^{-}}$

$H_2O_2 + 2Fe^{2+} + 2H^{+} \rightarrow Fe^{3+} + 2H_2O$

17. Calculate the empirical and molecular formula of a compound containing 76.6% carbon, 6.38 % hydrogen and rest oxygen its vapour density is 47.

Element	**Percentage**	**Atomic mass**	**Relative number of atoms**	**Simple ratio**	**Whole no**
C	76.6	12	$\frac{76.6}{12} = 6.38$	$\frac{6.38}{1.06} = 6$	6
H	6.38	1	$\frac{6.38}{1} = 6.38$	$\frac{6.38}{1.06} = 6$	6
O	17.02	16	$\frac{17.02}{16} = 1.06$	$\frac{1.06}{1.06} = 1$	1

Empirical Formula = C_6H_6O

$$n = \frac{Molar\ mass}{Calculated\ empirical\ formula\ mass}$$

$$= \frac{2 \times vapour\ density}{94} = \frac{2 \times 47}{94} = 1,$$

since Molar mass = 2 × Vapour density

Molecular formula n × empirical formula

Molecular formula $(C_6H_6O) \times 1 = C_6H_6O$

18. A Compound o analysis gave Na = 14.31 % S = 9.97% H = 6.22% and O = 69.5% calculate the molecular formula of the compound if all the hydrogen in the compound is present in combination with oxygen as water of crystallization. (molecular mass of the compound is 322).

Element	%	Relative number of atoms	Simple ratio
Na	14.31	$\frac{14.31}{23} = 0.62$	$\frac{0.62}{0.31} = 2$
S	9.97	$\frac{9.97}{32} = 0.31$	$\frac{0.31}{0.31} = 1$
H	6.22	$\frac{6.22}{1} = 6.22$	$\frac{6.22}{0.31} = 20$
O	69.5	$\frac{69.5}{16} = 4.34$	$\frac{4.34}{0.31} = 14$

Empirical formula = $Na_2SH_{20}O_{14}$

$$n = \frac{molar\ mass}{Calculated\ empirical\ formula\ mass}$$

$$= \frac{322}{322} = 1$$

$$\left[\begin{array}{c} Na_2SH_{20}O_{14} \\ = (2 \times 23) + (1 \times 32) + (20 \times 1) + 14(16) \\ = 46 + 32 + 20 + 224 \\ = 322 \end{array}\right]$$

Molecular formula = $Na_2\ S\ H_{20}\ O_{14}$

Since all the hydrogen in the compound present as water

Molecular formula is $Na_2\ SO_4.10H_2O$.

44. Balance the following equations by oxidation number method

i) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

ii) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

iii) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

iv) $KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

+6 -1 +3 0

i) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

$2 \times \uparrow 3e^-$ $\downarrow 1e^-$

$K_2Cr_2O_7 + 6KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

$K_2Cr_2O_7 + 6KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + 3I_2 + H_2O$

$K_2Cr_2O_7 + 6KI + 7H_2SO_4 \rightarrow 4K_2SO_4 + Cr_2(SO_4)_3 + 3I_2 + 7H_2O$

+7 +4 +4 +6

ii) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

↑ ↓

$3e^-$ $2e^-$

⇨ $2KMnO_4 + 3Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

⇨ $2KMnO_4 + 3Na_2SO_3 \rightarrow 2MnO_2 + 3Na_2SO_4 + KOH$

⇨ $2KMnO_4 + 3Na_2SO_3 \rightarrow 2MnO_2 + 3Na_2SO_4 + 2KOH$

0 +5 +2 +4

iii) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

↓ ↑

$2e^-$ $1e^-$

$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

$Cu + 2HNO_3 + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

+7 +3 +2 +4

iv) $KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

↓ ↑

$5e^-$ $1e^- \times 2$

$2KMnO_4 + 5H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

$2KMnO_4 + 5H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + H_2O$

$2KMnO_4 + 5H_2C_2O_4 + 3H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + 8H_2O$

45. Balance the following equations by ion electron method.

i) $KMnO_4 + SnCl_2 + HCl \rightarrow MnCl_2 + SnCl_4 + H_2O + KCl$

ii) $C_2O_4^{2-} + Cr_2O_7^{2-} \rightarrow Cr^{3+} + CO_2$ (in acid medium)

iii) $Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + NaI$ (in acid medium)

iv) $Zn + NO_3^- \rightarrow Zn^{2+} + NO$

(i) Half reaction are:

+7

$MnO_4^- \rightarrow Mn^{2+}$ ---- (1)

And $Sn^{2+} \rightarrow Sn^{4+}$ ---- (2)

(1) ➔ $MnO_4^- + 8H^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$

(2) ➔ $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

(1)× 2 ➔ $2MnO_4^- + 16H^- + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$

(2) × 5➔ $5Sn^{2+} \rightarrow 5Sn^{4+} + 10e^-$

➔ $2MnO_4^- + 5Sn^{2+} + 16H^+ \rightarrow 2Mn^{2+} + 5Sn^{4+} + 8H_2O$

(ii) +3 +4

$C_2O_4^{2-} \rightarrow CO_2$ ---- (1)

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ ---- (2)

(1) ➔ $C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$ ---- (3)

$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$ ---- (4)

(2) × 3 ➔ $3C_2O_4^{2-} \rightarrow 6CO_2 + 6e^-$ ---- (5)

$\underline{Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O}$ ---- (4)

➔ $\underline{Cr_2O_7^{2-} + 3C_2O_4^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 6CO_2 + 7H_2O}$

(iii) $S_2O_3^{2-} \rightarrow S_4O_6^{2-}$ ---- (1)

Half reaction ➔ $I_2 \rightarrow I$ ---- (2)

(1) ➔ $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$ ---- (3)

(2) ➔ $\underline{I_2 + 2e^- \rightarrow 2I^-}$ ---- (4)

(3) + (4) ➔ $\underline{2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-}$

(iv) $Zn \rightarrow Zn^{2+}$ ---- (1)

$NO_3^- \rightarrow NO$ ---- (2)

(1) ➔ $Zn \rightarrow Zn^{2+} + 2e^-$ ---- (3)

(2) ➔ $NO_3^- + 3e^- + 4H^+ \rightarrow NO + 2H_2O$ ---- (4)

(3) × 3 ➔ $3Zn \rightarrow Zn^{2+} + 6e^-$ ---- (5)

(4) × 2 ➔ $\underline{2NO_3^- + 6e^- + 8H^+ \rightarrow 2NO + 4H_2O}$ ---- (6)

$\underline{3Zn + 2NO_3^- + 8H^+ \rightarrow 3Zn^{2+} + 2NO + 4H_2O}$

IMPORTANT QUESTION

LESSON 1 Basic Concepts of Chemistry and Chemical Calculations.

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24	M25
8	6	8	7	5	6	11	3	6	9	

1. Define matter. Explain classification of matter. (2)
2. Define relative atomic mass. (4) calculate relative atomic mass of hydrogen.
3. Define average atomic mass. Find average atomic mass of chlorine. (4)
4. Define relative molecular mass. (4) **J24 2M**
5. Find relative molecular mass of H_2, $C_6H_{12}O_6$, Ethanol….. (4,5)
6. What do you understand by term mole. (6) **(J19 2M, J23 2M, J24 5Mi**
7. What is Avogadro number. (7)
8. Define molar mass. Give an example. (8)
9. Define molar volume. Give an example. (8)
10. Define Gram equivalent mass. (8) **M22 2M, M24 2M**
11. Calculate equivalent mass of acid.(H_2SO_4**(M19 3M)**, HCl, H_3PO_4**(S20)**)(9)
12. Define basicity. Find the basicity of ortho-phosphoric acid. (9) **S20 2M**
13. Calculate equivalent mass of base.(KOH, NaOH)(9)
14. Calculate equivalent mass of Oxidising or reducing agent ($KMnO_4$)(9)
15. Define empirical formula & molecular formula.(10) Determination of Empirical formula, molecular formula **PROBLEMS & book back** (11)

16. What are the steps involving in the calculation of molecular formula from empirical formula? (10)
17. An acid found in tamarinds on analysis shows the following percentage composition: 32% Carbon, 4% Hydrogen, 64% oxygen. Find the empirical formula of the compound. (11)
18. An organic compound present in vinegar has 40% carbon, 6.6% of Hydrogen and 53.4 % of Oxygen. Find the empirical formula of the compound. (11) **(MQ19) M24 5Mi**
19. A compound having the empirical formula C_6H_6O has the vapour density 47. Find its Molecular formula. (**M19 5Mi)** (BB34)
20. Calculate the empirical and molecular formula of a compound containing 76.6% carbon, 6.38% of hydrogen and rest oxygen. Its vapour density is 47. (BB34) **S20 5M, J22 3M**
21. Define stoichiometry. (13) & Calculations based on stoichiometry problems. (15)
22. Define Limiting reagent **J22 2M** & excess reagents. (17) problems (18)
23. One the formation of SF_6 by the direct combination of S and F_2, which is the limiting reagent? Prove it.
24. Define redox reactions. (19)
25. Distinguish between oxidation and reduction. (19) **S21 3M**, **M23 2M**
26. What do you understand by the term oxidation number. (20) rules of oxidation number. (20) Oxidation number **problem** (21)
27. Calculate oxidation state of carbon in (i) CH_4 (ii) CCl_4 **(MQ19)**
28. Calculate oxidation number of oxygen in H_2O_2. (21) **(M19 5Mi)**
29. Calculate the oxidation number of underlined elements. I) $\underline{C}O_2$ ii) $H_2\underline{S}O_4$ **(21) M22 3M**
30. Redox reactions in terms of oxidation numbers. (22)
31. What is combination reaction? Give example. (22)
32. What is decomposition reaction? Give two examples. (22)

(Write the balanced equation for the action of heat on calcium carbonate. (22) **M24 2M) UNIT 5**

33. What is displacement reactions? Give its types. Explain with example.(23)
34. What is disproportionation reactions? Give example. (24)
35. What are competitive electron transfer reaction? Give example. (25)
36. What is the empirical formula of the following? BB (24) **S21 2M, J24 2M**
    a) Fructose ($C_6H_{12}O_6$) b) Caffeine ($C_8H_{10}N_4O_2$)
37. Balance the following equation using oxidation number method. (25)
    $FeSO_4 + KMnO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + MnSO_4 + K_2SO_4 + H_2O$
    $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$ **M23 3Mi**
    $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$ **M23 3Mii**
38. Balance the following equation by oxidation number method. (26)
    $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$ (Acidic medium)
39. A Compound o analysis gave Na = 14.31 % S = 9.97% H = 6.22% and O = 69.5% calculate the molecular formula of the compound if all the hydrogen in the compound is present in combination with oxygen as water of crystallization. (molecular mass of the compound is 322). BB **M23 5M**

PREOBLEM- Mole, molecular mass, equivalent mass, empirical formula, balance equation & study all book back question & answer

2. Quantum Mechanical Model of Atom

Answer the following: ***TEXT BOOK BACK ANSWER***

26. Which quantum number reveal information about the shape, energy, orientation and size of orbitals?

ANSWER: Magnetic quantum number

	Quantum Number	**Information Obtained**
(i)	Principal Quantum number (n)	Size and energy
(ii)	Azimuthal Quantum number (l)	Shape
(iii)	Magnetic Quantum number (m)	Orientation

27. How many orbitals are possible for n = 4?

n = 4, (l=n-1) l = 0, 1, 2, 3

four sub-shells => s, p, d, f

$l = 0$	$m_l = 0$;	one 4s orbital
$l = 1$	$m_l = -1, 0, +1$;	three 4p orbitals.
$l = 2$	$m_l = -2, -1, 0, +1, +2$;	five 4d orbitals.
$l = 3$	$m_l = -3, -2, -1, 0, +1, +2, +3$;	seven 4f orbitals

Over all sixteen orbitals.

(Or) $\mathbf{n^2=4^2=16}$

28. How many radial nodes for 2s, 4p, 5d and 4f orbitals exhibit? How many angular nodes?

Orbital	**n**	**l**	**Radial node n-l-1**	**Angular node l**
2s	2	0	2-0-1=1	0
4p	4	1	4-1-1=2	1
5d	5	2	5-2-1=2	2
4f	4	3	4-3-1=0	3

29. The stabilization of a half-filled d- orbital is more pronounced than that of the p-orbital why?

(i) The half-filled d- orbital is much **stabilized** than the p-orbital.

(ii) This is due to its **high exchange energy** and **symmetry** than that of p-orbital.

30. Consider the following electronic arrangements for the d^5 configuration.

⥮	⥮	1		

1	1	1	⥮	

1	1	1	1	1

(a) (b) (c)

(i) which of these represents the ground state

(ii) which configuration has the maximum exchange energy.

(i) c) Ground state

(ii) c) Maximum exchange energy :

1	1	1	1	1

31. **State and explain Pauli's exclusion principle.**

"No two electrons in an atom can have the same set of values of all four-quantum number".

For the two electrons present in **Helium**,$1s^2$

	n	l	m	s
First electron	1	0	0	+1/2
Second electron	1	0	0	-1/2

Spin quantum number can have only two values +1/2 and -1/2, only two electrons can be accommodated in a given orbital in accordance with Pauli's exclusion principle.

32. Define orbital? What are the n and l values for 3p_x and 4d_{x2-y2} electron?

An orbital is a three-dimension space which the probability of finding an electron of given energy is maximum.

Orbital	**n**	**l**
$3p_x$	3	1
$4d_{x2-y2}$	4	2

33. Explain briefly the time independent Schrödinger wave equation?

Erwin Schrodinger expressed the wave nature of electron in terms of a differential equation. This equation determines the change of wave function in space depending on the field of force in which the electron moves.

The time independent Schrodinger equation can be expressed as,

$$\hat{H}\psi = E\psi \text{ -------------------(1)}$$

Where Ĥ is called Hamiltonian operator, ψ is the wave function and is a function of position co-ordinates of the particle and is denoted as ψ (x, y, z) E is the energy of the system

$$\hat{H} = \left[\frac{-h^2}{8\pi^2 m}\left(\frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2}\right) + V\right] \text{ ---------------------------(2)}$$

Substitute equ (2) in (1) Can be written as,

$$\left[\frac{-h^2}{8\pi^2 m}\left(\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2}\right) + V\psi\right] = E\psi$$

Multiply by $-\frac{8\pi^2 m}{h^2}$ on both sides,

$$-\frac{8\pi^2 m}{h^2}\left[\frac{-h^2}{8\pi^2 m}\left(\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2}\right) + V\psi\right] = -\frac{8\pi^2 m}{h^2}E\psi$$

$$\left[-\frac{8\cancel{\pi^2}m}{\cancel{h^2}} x \frac{-\cancel{h^2}}{8\pi^2\cancel{m}}\left(\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2}\right)\right] - \frac{8\pi^2 m}{h^2}V\psi = -\frac{8\pi^2 m}{h^2}E\psi$$

$$\left[\left(\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2}\right)\right]V\psi + \frac{8\pi^2 m}{h^2}E\psi = 0$$

Rearranging the above equation,

$$\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

The above Schrödinger wave equation does not contain time as a variable and is referred to as **time independent Schrödinger wave equation**.

34. **Calculate the uncertainty in position of an electron, if Δv = 0.1% and υ = 2.2 × 10^6 ms^{-1}.**

$$\Delta x.\ \Delta p \geq \frac{h}{4\pi}$$

$$\Delta x.\ \Delta v.m = \frac{h}{4\pi}$$

$$\Delta x = \frac{h}{4\pi \Delta v.m}$$

$$\Delta v = \frac{0.1}{100} \times 2.2 \times 10^6\ ms^{-1} = 2.2 \times 10^3\ ms^{-1}$$

$$\Delta x \ = \frac{6.626 \times 10^{-34}\ Kgm^2s^{-1}}{4x3.14x\,9.1\ \times 10^{-31}\ Kg\ \times 2.2 \times 10^3 ms^{-1}}$$

$$\Delta x \ = \frac{6.626 \times 10^{-34}\ x\ 10^{-3}\ \times 10^{31}}{4x3.14x\,9.1\ \times 2.2}$$

$$\Delta x \ = \frac{6.626 \times 10^{-6}}{251.45}$$

$\Delta x = 0.02635\ x\ 10^{-6}$

$\Delta x = 2.635x10^{-8}$ m

Another method:

$$\Delta x.\ \Delta p \geq \frac{h}{4\pi}$$

$\Delta x.\ \Delta p \geq 5.28 \times 10^{-35}\ Kgm^2s^{-1}$

$\Delta x.\ (m\Delta v) \geq 5.28 \times 10^{-35}\ Kgm^2s^{-1}$

Given $\Delta v = 0.1\%$

$\upsilon = 2.2 \times 10^6\ ms^{-1}$; $m = 9.1 \times 10^{-31}$ Kg

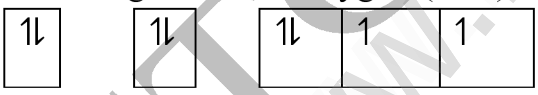
$$\Delta v = \frac{0.1}{100} \times 2.2 \times 10^6\ ms^{-1} = 2.2 \times 10^3\ ms^{-1}$$

$$\Delta x \ \geq \frac{5.28 \times 10^{-35}\ Kgm^2s^{-1}}{9.1\ \times 10^{-31}\ Kg\ \times 2.2 \times 10^3 ms^{-1}}$$

$\Delta x\ \geq 2.64 \times 10^{-8}$ m

35. Determine the values of all the four quantum numbers of the 8th electron in O- atom a 15th electron in the Cl atom.

Electronic configuration of oxygen (Z=8)= $1s^2\ 2s^2\ 2p_x^2 2p_y^1 2p_z^1$



8th electron present in **$2p_x$ orbital** and an the quantum numbers are

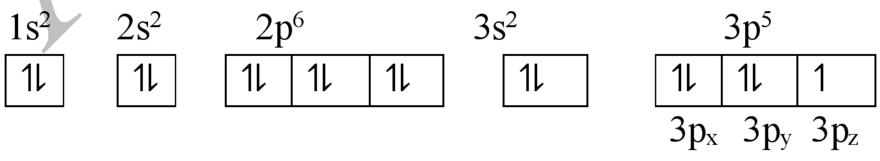
n = principal quantum number =2

l= azimuthal quantum number= 1

m_l =magnetic quantum number = either +1 or -1

m_s=spin quantum number = -1/2

Electronic configuration of chlorine (Z=17) = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^2 3p_y^{\ 2}\ 3p_z^1$



15th electron present in **$3p_z$ orbital** and

the quantum numbers are **$n = 3,\ l = 1\ m_l$ = either +1 or -1 and $m_s = +1/2$**

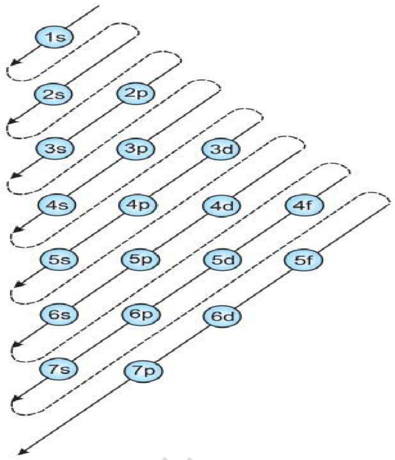


Fig.2.17 Order of filling of orbitals

36. The quantum mechanical treatment of the hydrogen atom gives the energy value: $E_n = \frac{-13.6}{n^2}$ ev atom^{-1}

(i) use this expression to find ΔE between n=3 and m=4

(ii) Calculate the wavelength corresponding to the above transition

$$E_n = \frac{-13.6}{n^2} \text{ eV atom}^{-1}$$

n = 3 $\quad E_3 = \frac{-13.6}{3^2} = \frac{-13.6}{9} = -1.51 \text{ eV atom}^{-1}$

n = 4 $\quad E_4 = \frac{-13.6}{4^2} = \frac{-13.6}{16} = -0.85 \text{ eV atom}^{-1}$

$\Delta E = (E_4 - E_3)$

$= (-0.85) - (-1.51) \text{ eV atom}^{-1}$

$= (-0.85 + 1.51)$

$= 0.66 \text{ eV atom}^{-1} \quad (1\text{eV} = 1.6 \times 10^{-19} \text{ J})$

$\Delta E = 0.66 \times 1.6 \times 10^{-19}$ J

$\Delta E = 1.06 \times 10^{-19}$ J

$hv = 1.06 \times 10^{-19}$ J $\quad \frac{hc}{\lambda} = 1.06 \times 10^{-19}$ J

$$\lambda = \frac{hc}{1.06 \times 10^{-19}\text{J}} = \frac{6.626 \times 10^{-34}\, JS \times 3 \times 10^{8}\, ms^{-1}}{1.06 \times 10^{-19} J}$$

$\lambda = 1.875 \times 10^{-6}$ m

37. how fast must a54g tennis ball in order to have a de Broglie wavelength that is equal to that a photon of green light 5400Å ?

De Broglie wavelength of the tennis ball equal to 5400 Å

m = 54 g ; v = ? $\quad \lambda = \frac{h}{mV} \qquad V = \frac{h}{m\lambda}$

$V = \frac{6.626 \times 10^{-34}\, JS}{54 \times 10^{-3}\, Kg \times 5400 \times 10^{-10} m}$ $\quad$ **$V = 2.27 \times 10^{-26}$ ms^{-1}**

38. For each of the following, give the sub level designation , the allowable m values and the number of orbitals (i) n=4, *l*=2 (ii) n=5, *l*=3 (iii) n=7, *l*=0

n	*l*	Sub Energy levels	m_l values	Number of orbitals
4	2	4d	-2, -1, 0,+1,+2	Five 4d orbitals
5	3	5f	-3,-2, -1, 0, +1,+2, +3	Seven 5f orbitals
7	0	7s	0	One 7s orbitals

39. **Give the electronic configuration of Mn^{2+} an Cr^{3+}**

Ions	No. of electrons	Electronic configuration
Mn^{2+}	23	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5$
Cr^{3+}	21	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^3$

40. Describe the Aufbau principle

The word Aufbau in German means '**building up**'.

- ***In the ground state of the atoms, the orbitals are filled in the order of their increasing energies***.
- That is the electrons first occupy the lowest energy orbitals available to them.
- Once the **lower energy orbitals are completely filled**, then the electrons enter the ***next higher energy orbitals.***
- The order of filling of various orbitals as per Aufbau principle is $1s<2s<2p<3s<3p<4s<3d<4p<5s<4d<\ldots\ldots$

For example: Electronic configuration of chromium (Z=24) = $1s^2\,2s^2\,2p^6\,3s^2\,3p^6 3d^5\,4s^1$

41. An atom of an element contains 35 electrons and 45 neutrons. Deduce

(i) the number of protons

(ii) the electronic configuration for the element

(iii) All the four quantum numbers for the last electron

Ans. (i) No of protons= 35

(ii) Electronic configuration $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^5$

(iii)

↑↓	↑↓	↑
$4p_x$	$4p_y$	$4p_z$

Last electron present in $4P_y$ orbital

n = 4, l = 1 m_l= either +1 or -1 and s = -1/2

42. Show that the circumference of the bohr's orbit for the hydrogen atom is an integral multiple of the de Broglie wave length associated with the electron revolving around the nucleus.

Ans: Quantization of angular momentum and de Broglie concept.

- According to the de Broglie concept, the electron that revolves around the nucleus exhibits both particle and wave character.
- In order for the electron wave to exist in phase , the circumference of the orbit should be an integral multiple of the wavelength of the electron wave. Otherwise, the electron wave is out of phase.

Circumference of the orbit = $n\lambda$

$2\pi r = n\lambda$ $\qquad 2\pi r = nh / mv$

Rearranging, $mvr = nh/2\pi$ Angular momentum = $nh/2\pi$

The above equation was already predicted by Bohr. Hence, De Broglie and Bohr's concepts are in agreement with each other.

43. Calculate the energy required for the process. $He^+_{(g)} \rightarrow He^{2+} + e^-$ The ionisation energy required for the H atom in its ground state is -13.6 ev atom^{-1}.

$He^+ \rightarrow He^{2+} + e^-$

$E_n = \frac{-13.6z^2}{n^2}$ $\quad E_1 = \frac{-13.6(2)^2}{(1)^2} = -54.4$

$E_\infty = \frac{-13.6(2)^2}{(\infty)^2} = 0$

∴ Required Energy for the given process

$= E_\infty - E_1 = 0 - (-54.4) = 54.4$ev.

44. An ion with the mass number 37 possesses unit negative charge . If the ion contains 11.1 % more neutrons than electrons. Find the symbol of the ion.

	Atom	**Uni-negative ion**
Number of electron	X – 1	X
Number of protons	X – 1	X - 1
Number of neutrons	Y	Y

Given that y, = x + 11.1% of x

$$= \left(x + \frac{11.1}{100}\,x\right) = x + 0.111x$$

$y = 1.111\,x$

Mass number = 37

Number of protons number of neutrons = 37

$(x-1) + 1.111\,x = 37$

$X + 1.111x = 38$

$2.111\,x = 38 \quad x = \frac{38}{2.11}\ ; \qquad x = 18.009$

x = 18 (whole number)

∴ Atomic number = x – 1 = 18 – 1 = 17

Mass number = 37

Symbol of the ion $^{37}{}_{17}Cl^-$

45. The Li ion is a hydrogen like ion that can be described by the Bohr's model. Calculated the Bohr radius of the third orbit and calculate the energy of the electron in 4th orbit

$$r_n = \frac{(0.529)n^2}{z}\,\text{Å} \qquad E_n = \frac{-13.6\,(z^2)}{n^2}\ \text{ev atom}^{-1}$$

for Li^{2+} Z = 3

Bohr radius for the third orbit (r_3)

$$= \frac{(0.529)(3)^2}{3} = 0.529 \times 3 = \mathbf{1.587\ Å}$$

Energy of an electron in the fourth orbit

$$(E_4) = \frac{-13.6(3)^2}{(4^2)} = \mathbf{-7.65\ eV\ atom^{-1}}$$

46. Protons can be accelerated in particle accelerators. Calculate the wavelength (in Å) of such accelerated proton moving at 2.85 × 10^8ms^{-1} (the mass of proton is 1.673 × 10^{-27} Kg)

$V = 2.85 \times 10^8\ ms^{-1}$; $m_p = 1.673 \times 10^{-27}$ Kg

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}\ kgm^2ms^{-1}}{1.673 \times 10^{-27} kg \times 2.85 \times 10^8\ ms^{-1}} = 1.389 \times 10^{-15}m \quad [\text{Å} = 10^{-10}m]$$

➔ $\boldsymbol{\lambda = 1.389 \times 10^{-5}}$ **Å**

47. What is the de Broglie wave length(in cm) of a 160g cricket ball travelling at 140 Km hr^{1}.

m = 160g = 160 × 10^{-3} kg

v = 140 Km hr^{-1} = $\frac{140 \times 10^3}{60 \times 60}$ ms^{-1}

v = 38.88ms^{-1}

$\lambda = \frac{h}{mv} \quad = \frac{6.626 \times 10^{-34}\ \text{kgm}^2\text{ms}^{-1}}{160 \times 10^{-3} kg \times 38.88\ ms^{-1}}$

λ = 1.065 × 10^{-34}m

wave length in cm =1.065 x 10^{-34} x 100 =1.065 x 10^{-32}cm

48. Suppose that the uncertainty in determining the position of an electron in an orbit is 0.6Å. what is the uncertainty in its momentum?

Δx = 0.6Å = 0.6 × 10^{-10}m

Δp = ? ; **Δx . Δp ≥ $\frac{h}{4\pi}$**

Δx . Δp ≥ 5.28 × 10^{-35} kgm^{2}s^{-1}

(0.6 × 10^{-10}) . Δp ≥ 5.28 × 10^{-35}

$\Delta p \geq \frac{5.28 \times 10^{-35}\ kgm^2s^{-1}}{0.6510^{-10}m}$ **Δp ≥ 8.8 × 10^{-25} kgm s^{-1}**

49. Show that if the measurement of the uncertainty in the location of the particle is is equal to its de Broglie wavelength, the minimum uncertainty in its velocity/4π.

Δx = Δv = ? **Δx . Δp ≥ $\frac{h}{4\pi}$** **λ(m** Δv) ≥ $\frac{h}{4\pi}$

$\Delta v \geq \frac{h}{4\pi\ (m\lambda)}$ $\Delta v \geq \frac{h}{4\pi \times m \times \frac{h}{mv}}$ $[\lambda = \frac{h}{mv}]$ **Δv ≥ $\frac{v}{4\pi}$** ;

Therefore, minimum uncertainty in velocity = $\frac{v}{4\pi}$

50. What is the de Broglie wavelength of an electron, which is accelerated from the rest, through a potential difference of 100V?

Potential difference = 100V

= 100 × 1.6 × 10^{-19}J

$\lambda = \frac{h}{\sqrt{2mev}} \quad = \frac{6.626 \times 10^{-34}\ \text{kgm}^2\text{ms}^{-1}}{\sqrt{2 \times 9.1 \times 10^{-31}\ Kg \times 100 \times 1.6 \times 10^{-19} J}}$

λ = 1.22 × 10^{-10}m

51. Identify the missing quantum numbers and the sub energy level.

n	*l*	m	Sub energy level
?	?	0	4d
3	1	0	?
?	?	?	5p
?	?	-2	3d

Answer:

n	*l*	m	Sub energy level

4	2	0	4d
3	1	0	3p
5	1	Any one value -1, 0, +1	5p
3	2	-2	3d

IMPORTANT QUESTION

Lesson 2 Quantum Mechanical Model of Atom

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
6	8	8	7	8	7	9	8	6	8				

1. Write note on Thomson model of an atom. (39)
2. Write note on Rutherford model of an atom. (39)
3. Write note on Bohr atom model & limitation. (39) (write the assumptions of Bohr atom model) **J24 5M**
4. Derive De –Broglie equation.(40) **(M19 5Mii)**
5. A macroscopic particle of mass one Kg is moving at a velocity 10 cms^{-1}. calculate its de Broglie wavelength. (41 model) **(MQ19)**
6. Davison and Germer experiment. (42)
7. Heisenberg uncertainty principle. (42) **S20 3M**, **J22 2M, M23 2M**
8. Calculate the uncertainty in the position of an electron, if the uncertainty in its velocity is 5.7 × 10^5 ms^{-1}. **(J19)**
9. Explain briefly the time independent Schrodinger wave equation? (43)
10. Explain main features of the quantum mechanical model of atom. (44)
11. Explain different types of quantum numbers. (44) Principal quantum number **S21 5Mii**, **M23 3M,** Azimuthal quantum number **J23 5Mii**, Magnetic quantum number **J22 5Mi, J23 5Mi**, Spin quantum number. **M24 5Mii**
12. Calculate the maximum number of electrons that can be accommodated in L shell. (44) **M22 2M**
13. Define orbital? **J23 2M** what are the n and l values for $3p_x$ and $4dx^2$-y^2 electron?(44) **(BBQ32 64)(J15), M24 3M**
14. Calculate the orbital angular momentum for d and f orbital. (45) **(J15)**
15. How many orbitals are possible for n=4? **(46) M22 5Mi , J24 2M**
16. Shapes of orbitals s, p, d, f. (49)
17. Calculate the total number of angular nodes and radial nodes present in 3d and 4f orbitals.(50) **S20 5Mi**
18. State and explain Pauli's exclusion principle.(52) **(M19 2M)** **M23 5Mi**
19. i) Describe the Aufbau principle. (52) **M21 2M (State Aufbau principle)**
    (ii) Write the electronic configuration of Fe^{2+} ion.
    (iii) How many radial nodes exist in 2s and 4f orbitals **(MQ19)**
20. State Hund's rule (53)
21. Give the electronic configuration of Mn^{2+} ,Cr^{3+}, Ni^{2+} and Fe^{3+}.(53) **J22 5Mii**
22. Write the electronic configuration and orbital diagram for nitrogen. (54) **M22 5Mii**
23. Give the electronic configuration of Copper and Chromium (55)

24. In degenerate orbitals, why do the completely filled and half-filled configurations are more stable than the partially filled configuration? **(55) (S20 2M COMPULSORY)**
25. What is exchange energy? How it is related with stability of atoms? Explain with suitable examples. (56)**S21 5Mi**

PROBLEM- De-Broglie, Heisenberg uncertainty principle & study all book back question & answer.

LESSON 3 PERIODIC CLASSIFICATIONS OF ELEMENTS

Write Brief Question with Answer

1. Define modern periodic law.

The modern periodic law states that, " the physical and chemical properties of the elements are periodic functions of their atomic numbers."

2. What are isoelectronic ions? Give examples.

Ions of different elements having the same number of electrons are called iso electronic ions.

Ions of different elements	Na^+	Mg^{2+}	Al^{3+}	F^-	O^{2-}	N^{3-}
No. of electrons	10	10	10	10	10	10

3. What is effective nuclear charge?

The net nuclear charge experienced by valence electrons in the outermost shell is called the effective nuclear charge. $\mathbf{Z_{eff} = Z - S}$

Where Z is the atomic number and 'S' is the screening constant

4. Is the definition given below for ionisation enthalpy correct? "Ionisation enthalpy is defined as the energy required to remove the most loosely bound electron from the valence shell of an atom"

No the above definition is incorrect.

The correct definition is, Ionisation energy is defined as the minimum amount of energy required to remove the most loosely bound electron from the valence shell of the isolated neutral gaseous atom in its ground state.

5. Magnesium loses electrons successively to form Mg^+, Mg^{2+} and Mg^{3+} ions. Which step will have the highest ionisation energy and why?

$Mg + I.E_1 \rightarrow Mg^+ + 1e^-$ ------ (i)

$Mg + I.E_2 \rightarrow Mg^{2+} + 1e^-$ ------- (ii)

$Mg^{2+} I.E_3 \rightarrow Mg^{3+} + 1e^-$ ------- (iii)

- The Step (iii) which involves the formation of Mg^{3+} **requires higher ionisation energy. $IE_3 > IE_2 > IE_1$**
- Mg^{2+} consist of 10 electrons (2, 8) attaining the stable noble gas configuration of argon (Z = 10)

- Since the valence orbital is completely filled, more energy will be required to remove electrons.

6. Define electronegativity

It is defined as the relative tendency of an element present in a covalently bonded molecule, to attract the shared pair of electrons towards itself.

7. How would you explain the fact that the second ionisation potential is always higher than the ionisation potential?

- Second ionisation potential is always higher than the ionisation potential.
- Removal of one electron from the valence orbit of a neutral gaseous atom is easy so first ionisation energy is less. But from a unipositive ion, removal of one more electron becomes difficult due to force of attraction between the excess of protons and less number of electrons.
- Due to **greater nuclear attraction**, second ionisation energy is higher than first ionisation energy.

The total numbers of electrons are less in the cation than the neutral atom while the nuclear charge remains the same. Therefore the effective nuclear charge of the cation is higher than the corresponding neutral atom. Thus the successive ionisation energies, always increase in the following order $\mathbf{IE_1 < IE_2 < IE_3 < \ldots\ldots}$

8. Energy of an electron in the ground state of the hydrogen atom is -2.18 × 10^{-18} J. Calculate the ionisation enthalpy of atomic hydrogen in terms of kJ mol^{-1}

Ionisation energy is the amount of energy required to remove the electron from the ground state (E_1) to excited state(E^{∞}).

$E_1 = -2.18 \times 10^{-18}$ J; $E_{\infty} = 0$

$\Delta E = E_{\infty} - E_1$

$= 0 - (-2.18 \times 10^{-18} \text{ J}) = 2.18 \times 10^{-18}$ J I.E per hydrogen atom $= 2.18 \times 10^{-18}$ J

Energy required to ionize 1 mole of hydrogen atoms, we multiply by the Avogadro constant.

I.E per mole of H-atom $= 2.18 \times 10^{-18} \text{ J} \times 6.023 \times 10^{23}$

$= 13.13 \times 10^{5} \text{ J mol}^{-1}$

$= +1313 \text{ KJmol}^{-1}$

9. Give the general electronic configuration of atom is one of the important factor which affects the value of ionisation potential and electron gain enthalpy. Explain.

- Half-filled valence shell electronic configuration and completely filled valence shell electronic configuration are more stable than partially filled electronic configuration.
- **For example: Be (Z=4) $1s^2 2s^2$ (Completely filled electronic configuration)**
  N (Z=7) $1s^2 2s^2 2p^3$ (half-filled electronic configuration)
- Beryllium, Nitrogen the addition of extra electron will disturb their stable electronic configuration and they have almost zero electron affinity.

- Ionisation energy and electron affinity is the amount of energy released or required in pulling out or adding an electron to a neutral atom. So both depend electronic configuration of the element
- Electronic configuration is the arrangement of electrons in an atom. The outermost electron shell is often referred to as the " valence shell" determines the chemical properties.
- Ionisation energy and electron affinity is the amount of energy released or required in pulling out or adding an electron to a neutral atom. So both depend electronic configuration of the element.

10. In what period and group will an element with Z = 118 will be present?

Ans: 7th period, 18th group .

Explanation:

$Z = 118$; $[_{86}Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^6$

In the periodic table the element with Z = 118 is located in p – block.

11. Justify that the fifth period of the periodic table should have 18 elements on the basis of quantum numbers.

i. According to aufbau's principle 5th period (Rb – Xe) has **nine orbital** (one 5s, five 4d and three 6p) to be filled.

ii. Nine orbitals can accommodate a maximum of 18 electrons. Hence fifth period of the periodic table should has 18 elements from rubidium (Z = 37) to Xenon (Z = 54).

12. Elements a, b, c and d have the following electronic configuration: a : $1s^2, 2s^2, 2p^6$ b : $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$ c : $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ d: $1s^2, 2s^2, 2p^1$ Which group of element among will belongs to the same group of the periodic table

- Elements of **a** and **c** belongs to **group18**;
- Elements of **b and d** belongs to **group 13**

13. Give the general electronic configuration of lanthanides and actinides.

Lanthanides : $[_{54}Xe]\ 4f^{1-14}\ 5d^{0-1}\ 6s^2$

Actinides : $[_{86}Rn]\ 5f^{0-14}\ 6d^{0-2}\ 7s^2$

14. Why halogens acts as a oxidizing agents.

i. Due to low bond dissociation enthalpy, high electronegativity and large negative electron gain enthalpy halogens have a strong tendency to accept electron and thus get reduced.

$$X_2 + 2e^- \rightarrow 2X^-$$

ii. The ready acceptance of the electron is due to the strong oxidising nature of the halogens.

iii. F_2 is the strongest oxidising halogen and it oxidises other halide ions in solution or even in the solid phase. In general, a halogen oxidises halide ions of higher atomic number.

iv. Hence, the oxidising ability of halogens decreases from fluorine to iodine as:

$$F_2 > Cl_2 > Br_2 > I_2$$

15. Mention any two anomalous properties of second period elements.

i. Lithium and beryllium form more covalent compounds, unlike the alkali and alkali earth metals which predominantly form ionic compounds.

ii. The elements of the second period have only four orbitals (2s & 2p) in the valence shell and have a maximum co-valence of 4, whereas the other members of the subsequent periods have more orbitals in their valence shell and shows higher valences.

For example, boron forms BF^-_4 and aluminium forms AlF_6^{3-}

16. Explain the pauling method for the determination of ionic radius.

Definition: Ionic radius is defined as the distance from the centre of the nucleus of the ion upto which exerts its influence on the electron cloud of the ion.

Pauling has calculated the radii of the ions on the basis of the observed internuclear distances in four crystal crystals namely NaF, KCl, RbBr and CsI. In each ionic crystal the cations and anions are isoelectronic with inert gas configuration.

NaF crystal – Na^+ - $1s^2 2s^2 2p^6$ Ne configuration

F^- - $1s^2 2s^2 2p^6$

i. Ionic radius of uni-univalent crystal can be calculated using Pauling's method from the inter ionic distance between the nuclei of the cation and anion.

ii. Pauling assumed that ions present in a crystal lattice are perfect spheres, and they are in contact with each other therefore,

$$d = r_{C+} + r_{A-} \quad \text{------ (1)}$$

Where d = distance between the centre of the nucleus of cation C^+ and anion A^- .

r_{C+} = radius of the cation

r_{A-} = radius of the anion

Pauling also assumed that the radius of the ion having noble gas electronic configuration is inversely proportional to the effective nuclear charge.

$$r_C{}^+ \alpha \frac{1}{(Z_{eff})_{C+}} \quad \text{------- (2)} \quad \text{and}$$

$$r_A{}^- \alpha \frac{1}{(Z_{eff})_{A^-}} \quad \text{------- (3)}$$

Where Z_{eff} is the effective nuclear charge and $Z_{eff} = Z - S$

Dividing the equation 1 by 3

$$\frac{r_{C^+}}{r_{A^-}} = \frac{(Z_{eff})_{A^-}}{(Z_{eff})_{C^+}} \qquad \text{-------}\ (4)$$

On solving equation and (1) and (4) the values of r_{C^+} and r_{A^-} can be obtained.

17. Explain the periodic trend of ionisation potential.

The energy required to remove the most loosely held electron from an isolated gaseous atom is called as ionisation potential.

Variation along a period :

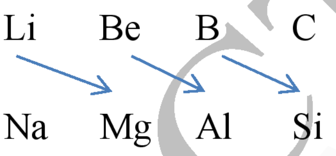
- Ionisation energy usually ***increases along a period***.
- This is due to ***increase of nuclear charge*** and ***decrease in size*** as we move from left to right in a period.

Periodic variation in group:

- Ionisation energy ***decreases down a group***.
- A gradual ***increase in atomic size***.
- As we move down a group, the valence electron occupies new shells, the distance between the nucleus and the valence electron increases.
- So, nuclear forces of attraction on valence electron decreases and hence ionisation energy also decreases down a group.

18. Explain the diagonal relationship.

- On moving diagonally across the periodic table the second and third period elements show certain similarities.
- Even though the similarity is not same as we see in a group, it is quite pronounced in the following pair of elements.



- The similarity in properties existing between the diagonally placed elements is called diagonal relationship.

19. Why the first ionisation enthalpy of sodium is lower than that of magnesium while its second ionisation enthalpy is higher than that of magnesium?

The electronic configuration of sodium (Z = 11) $1s^2\ 2s^2\ 2p^6\ 3s^1$

Magnesium (Z = 12) $1s^2\ 2s^2\ 2p^6\ 3s^2$

Magnesium atom has a ***smaller radius*** and ***higher nuclear charge*** than a sodium atom, thus more energy will be required to remove the electron from the same orbital (3s), making, The first ionisation energy of magnesium higher than that of sodium.

However, the second ionization enthalpy of sodium is higher than that of magnesium. This is because after losing 1 electron, sodium attains the stable noble gas configuration of neon ($1s^2$ $2s^2$ $2p^6$). On the other hand, magnesium, after losing 1 electron still has one electron in the 3s-orbital($1s^2$ $2s^2$ $2p^6$ $3s^1$). In order to attain the stable noble gas configuration, Thus, the energy required to remove the second electron in case of sodium is much higher than that required in case of magnesium. Hence, the second ionization enthalpy of sodium is higher than that of magnesium.

20. By using paulings method calculate the ionic radii of K^+ and Cl^- ions in the potassium chloride crystal. Given that $d_{K^+-Cl^-} = 3.14$ Å

$r(K^+) + r(Cl^-) = d(K^+ - Cl^-) = 3.14$ Å ------- (1)

The effective nuclear charge for K+ and Cl- can be calculated as follows.

K^+ =	$(1s^2)$	$(2s^2\ 2p^6)$	$(3s^2\ 3p^6)$
	Innershell	(n – 1)th shell	nth shell

$Z^*(K\text{-}) = Z - S$

$= 19 - [(0.35 \times 7) + (0.85 \times 8) + (1 \times 2)]$

$= 19 - 11.25 = 7.75$

$Z^*(Cl^-) = 17 - [(0.35 \times 7) + (0.85 \times 8) + (1 \times 2)]$

$= 17 - 11.25 = 5.75$

$$\frac{r(K^+)}{r(Cl^-)} = \frac{Z^*(Cl^-)}{Z^*(K^+)} = \frac{5.75}{7.75} = 0.74$$

$r(K^+) = 0.74\ r(Cl^-)$ **------- (2)**

Substitute (2) in (1)

$0.74\ r(Cl^-) + r(Cl^-) = 3.14$ Å

$0.74\ r(Cl^-) = 3.14$ Å;

$r(Cl^-) = \frac{3.14A}{1.74} = \mathbf{1.81}$**Å**

21. Explain the following, give appropriate reasons.

(i) Ionisation potential of N is greater than that of O.

(ii) First ionisation potential of C-atom is greater than that of B atom, where as the reverse is true is for second ionisation potential.

(iii) The electron affinity values of Be, Mg and noble gases are zero and those of N (0.02 eV) and P (0.80 eV) are very low

(iv) The formation of $F^-_{(g)}$ from $F_{(g)}$ is exothermic while that of $O^{2-}_{(g)}$ from $O_{(g)}$ is endothermic.

i. Electron configuration of nitrogen (Z = 7) $1s^2$ $2s^2$ $2p^3$

Electron configuration of oxygen (Z = 8) $1s^2\ 2s^2\ 2p^4$

Nitrogen has a ***half-filled electronic configuration*** and its ***more stable***. Due to stability , ionisation energy of nitrogen is high.

Hence nitrogen would require more ionization energy to remove an electron from it's outer shell than oxygen.

ii. Electron configuration of Carbon (Z = 6) $1s^2\ 2s^2\ 2p^2$

Electron configuration of Boron (Z = 5) $1s^2\ 2s^2\ 2p^1$

- The size of a carbon atom is smaller than boron. So the valence electron of carbon has ***greater nuclear charge*** than that of boron.
- Hence the first I.E of carbon is greater than that of boron.
- However, the second ionization enthalpy of boron is higher than that of carbon.
- This is because after losing 1 electron, Boron has a fully filled orbital ($2s^2$) than carbon ($2p^1$). ***Fully filled orbitals have more stability*** than partially filled orbitals so greater amount of energy will be needed to remove an electron from boron. So in this case, the second I.E of boron is higher than that of carbon.

iii. The electron affinities of Be, Mg and noble gases are almost zero.

Be (Z= 4; $1s^2\ 2s^2$) and

Mg (Z = 12 ; $1s^2\ 2s^2\ 2p^6\ 3s^2$) are having s ***fully filled in their valence shell***.

- Fully filled orbitals are ***most stable due to symmetry***.
- Therefore, these elements would be having least tendency to accept electron.

Hence, Be and Mg would be having zero electron affinity.

- N (Z = 7 ; $1s^2\ 2s^2\ 2px^1\ 2py^1\ 2pz^1$) and

P (Z = 15) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ is having ***half-filled 2p- subshell***.

- Half-filled sub shells are ***most stable due to symmetry*** (Hund's rule).
- Thus, nitrogen and phosphorous are having least tendency to accept electron.

Hence, have low electron affinity.

iv. F (Z = 9) $1s^2\ 2s^2\ 2p^5$

Fluorine is ***highly electro negative*** in nature therefore as it gains the one electron to attain the nearest inert gas configuration, become stable and releases the energy so exothermic.

O (Z = 8) $1s^2\ 2s^2\ 2p^4$

While in oxygen the addition of first electron is exothermic in nature but addition of second electron experiences high repulsive force. So needs extra external energy to enter outer shell, hence endothermic in nature.

22. What is screening effect?

Screening effect: The repulsive force between the inner shell electrons and the valence electrons leads to a decrease in the electrostatic attractive forces acting on the valence electrons by the nucleus. Thus, the inner shell electrons act as a shield between the nucleus and the valence electrons. This effect is called shielding effect (or) Screening effect.

23. Briefly give the basis for pauling's scale of electronegativity.

Electronegativity: It is defined as the relative tendency of an element present in a covalently bonded molecule, to attract the shared pair of electrons towards itself.

Pauling's scale: Pauling, he assigned arbitrary value of electro negativities for hydrogen and fluorine as 2.2 and 4.0 respectively. Based on this the electronegativity values for other elements can be calculated using the following expression.

$$(X_A - X_B) = 0.182\sqrt{E_{AB}} - (E_{AA} * E_{BB})^{\frac{1}{2}}$$

Where E_{AB} , E_{AA} and E_{BB} are the bond dissociation energies of AB, A_2 and B_2 molecules respectively.

The electronegativity of any given element is not a constant and its value depends on the element to which it is covalently bound.

The electronegativity values play an important role in predicting the nature of the bond.

Pauling assigned arbitrary value of electro negativities for hydrogen and fluorine as 2.1 and 4.0 respectively.

24. State the trends in the variation of electronegativity in group and periods.

(i) Variation of Electronegativity in a period:

- The electronegativity generally ***increases across a period from left to right***.
- The atomic radius decreases in a period, as the attraction between the valence electron and the nucleus increases.
- Hence the tendency to attract shared pair of electrons increases.
- Therefore, electronegativity also increases in a period.

(ii) Variation of Electronegativity in a group :

- The electronegativity generally ***decreases down a group***.
- As we move down a group the atomic radius increases and the nuclear attractive force on the valence electron decreases.
- Hence, the electronegativity decreases.

IMPORTANT QUESTION

Lesson 3 PERIODIC CLASSIFICATION OF ELEMENTS

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
9	6	7	8	8	9	6	9	8	6				

1. State Johann Dobereiner's law of triads. (69) **M19 5Mi**
2. Write a note about Chancourtois classification. (70)
3. State the Newland's law of octaves. (70)
4. State Lothar Meyer law. (70)
5. State Mendeleev's periodic law. (70)
6. Explain about the relationship between the atomic number of an element and frequency of the X-ray emitted from the elements.(72)
7. Define modern periodic law.(73) M23 5Mii

8. Nomenclature of elements with atomic number greater than 100. (75)
9. Variation of electronic configuration along the periods. (76)
10. Variation of electronic configuration along in the Groups.(76) s, p, d, f **(J22 5Mi)** block elements.
11. Give the general electronic configuration of Lanthanides and actinides. **(78) (J19 2M, M24 5Mii**
12. Define atomic radius (79) **S21 5Mi**
13. Covalent radius (79) calculation (79)
14. What are isoelectronic ions? Give examples. (bb)
15. Explain the Pauling's method of determination of ionic radius. (83) **(MQ19, S20 3M, M22 5M**
16. what is effective nuclear charge? (79)
17. Effective nuclear charge calculation (81)
18. Calculate the effective nuclear charge on 4s electron and 3d electron in scandium. (81) **J23 5M**
19. Define ionisation energy & variation of group, period. (84)
20. Describe the periodic trend of ionisation potential. (84) **M24 3M**
21. Define ionization energy. The first ionization energy of Nitrogen is greater than that of Oxygen- give appropriate reason. (84) **(MQ19) (J19)**
22. Explain the fact that the second ionisation potential is always higher than first ionisation potential. (BBQ30) (84) **J23 3M**
23. Compare the ionisation energy of beryllium and boron. (85) **J22 3M**
24. Define electron affinity **(M22 3M)** & variation of group, period. (86)
25. Explain why the electron affinity of Be and N is almost zero. (86) **(S20 5Mii)**
26. Define electronegativity. **M24 2M.** State the trends in the variation of electronegativity in group and periods. (87) **S21 3M, J22 5Mii**
27. Define – valency. (88) **(M19 2M)**
28. Explain diagonal relationship. (90) **(M19 3M, S21 5Mii, M23 3M, J24 5Mii.**
29. Why halogens act as oxidizing agents? (BB97) **J24 3M**

& study all book back question & answer

4. Hydrogen

Answer the following: BOOK BACK

22. Explain why hydrogen is not placed with the halogen in the periodic table.

- Hydrogen has similarities with alkali metals as well as halogens.
- In most of compound's hydrogen exist in +1 oxidation state.
- So, it is placed in group 1 and not placed with the halogen in periodic table.
- The **electron affinity** of hydrogen is **much less than** that of halogen atoms.

ALTERNERIVELY:

- Hydrogen has a tendency to gain one electron to from hydride ion is **low** compared to that of halogen to form halide ions.
- The electron affinity and electronegativity of hydrogen is much **less than** that of halogen atoms.
- Hence hydrogen is less reactive as compared to halogens.
- Due to this unique behavior hydrogen is not placed with halogens in the periodic table.

23. Discuss the three types of covalent hydrides.

Covalent hydrides are compounds in which hydrogen is attached to another element by sharing of electrons. Covalent hydrides are further divided into three categories, viz.,

- **Electron precise - Example: CH_4, C_2H_6, SiH_4, GeH_4**
- **Electron-deficient – Example: B_2H_6**
- **Electron-rich hydrides – Example: NH_3, H_2O**

24. Predict which of the following hydrides is a gas on a solid (a) HCl (b) NaH. Give your reason.

- NaH is a **gas on a solid** hydride.

a) HCl **Gas**- due to the absence of their molecular hydrogen bonding.

(b) NaH **Solid**- ***Reason***: NaH is an ionic hydride formed by transfer of electrons from metal to Hydrogen.

They have high lattice enthalpy and high density since hydride ions occupy holes in the lattice of metal without distorting the metal lattice.

25. Write the expected formulas for the hydrides of 4th period elements. What is the trend in the formulas? In what way the first two numbers of the series different from the others?

(i) The *first two elements of period 4, __K and Ca__* form hydrides with the formula MH and MH_2 i.e, (MH_n). They are ionic hydrides and have high melting and boiling points.

(ii) Elements from Sc to Zn form metallic or interstitial hydrides are **non-stoichiometric hydrides** which do not follow the law of constant composition.

(iii) Elements from Ga to Br, the general formula is **XH_{8-n}**, where n is the number of valence electrons.

26. Write chemical equation for the following reactions.

i) reaction of hydrogen with tungsten (VI) oxide NO_3 on heating. ii) hydrogen gas and chlorine gas.

(i) $WO_3 + 3H_2 \rightarrow W + 3H_2O$

(ii) $H_2 + Cl_2 \rightarrow 2HCl$.

27. Complete the following chemical reactions and classify them in to (a) hydrolysis (b) redox (c) hydration reactions.

1) $KMnO_4 + H_2O_2 \rightarrow 2MnO_2 + 2KOH + 2H_2O + 3O_2$ **(redox)**

2) $CrCl_3 + 6H_2O \rightarrow CrCl_3.6H_2O$ **(hydration)** or $[Cr(H_2O)_6]Cl_3$

3) $CaO + H_2O \rightarrow Ca(OH)_2$ **(hydrolysis)**

28. Hydrogen peroxide can function as an oxidising agent as well as reducing agent. substantiate this statement with suitable examples.

Hydrogen peroxide can act both as an usually performed in acidic medium while the reduction reactions are performed in basic medium.

i) **H_2O_2 act as oxidising agent in acidic medium**

For example $2FeSO_4 + H_2SO_4 + H_2O_2 \rightarrow Fe_2(SO_4)_3 + 2H_2O$

ii) **H_2O_2 act as reducing agent in acidic medium**

For example $2KMnO_{4(aq)} + 3H_2O_{2(aq)} \rightarrow MnO_2 + 2KOH + 2H_2O + 3O_{2(g)}$

29. Do you think that heavy water can be used for drinking purposes?

- Heavy water (D_2O) acts as a moderator, i.e., it ***slows the rate of a reaction.*** Due to this property of D_2O, it cannot be used for drinking purposes because it ***will slow down anabolic and catabolic reactions taking place in the body and lead to a casualty.***
- ***Molecular mass of heavy water is higher than water*** so its ***physiological action is different from ordinary water*** and because of that reason it cannot be used as drinking water.

30. What is water-gas shift reaction?

The carbon monoxide of the water gas can be converted to carbon dioxide by mixing the gas mixture with more steam at 400° C and passed over a shift converter containing iron/copper catalyst. This reaction is called as water-gas shift reaction.

$$CO + H_2O \rightarrow CO_2 + H_2$$

The CO_2 formed in the above process is absorbed in a solution of potassium carbonate.

$$CO_2 + K_2CO_3 + H_2O \rightarrow 2KHCO_3$$

31. Justify the position of hydrogen in the periodic table?

- The position of Hydrogen in the periodic table is **first** (**Period 1 group 1)** because its atomic numbe is 1. This means that hydrogen has a single electron in its outer shell. The elements in the periodi table are placed according to their electronic configuration.
- Hydrogen has the electronic configuration of $1s^1$ which resembles with ns^1 general valence shel configuration of alkali metals.

32. What are isotopes? Write the names of isotopes of hydrogen.

Isotopes: Atoms of the same element having same atomic number but different mass number are called isotopes.

	Isotopes of Hydrogen	Symbol	Atomic No	Mass No
(i)	Protium or hydrogen	$_1H^1$	1	1
(ii)	Deuterium	$_1D^2$	1	2
(iii)	Tritium	$_1T^3$	1	3

33. Give the uses of heavy water.

i. Heavy water (D_2O) is widely used as ***moderator in nuclear reactors.***
ii. It is commonly used as a tracer to study element reaction mechanisms.
iii. It is used as a coolant in nuclear reactors.

34. Explain the exchange reactions of deuterium.

Deuterium can replace reversibly hydrogen in compounds either partially or completely depending upon the reaction conditions.

$$CH_4 + 2D_2 \rightarrow \underset{\text{Deutero methane}}{CD_4} + 2H_2$$

$$2NH_3 + 3D_2 \rightarrow \underset{\text{Deutero ammonia}}{2ND_3} + 3H_2$$

35. How do you convert parahydrogen into ortho hydrogen?

i. By treatment with catalyst like Pt **or** Fe
ii. By passing an electric discharge
iii. By heating to 800°C or more
iv. By mixing with paramagnetic molecules like O_2, NO, NO_2.
v. By mixing with nascent hydrogen or atomic hydrogen.

36. Mention the uses of deuterium.

i. It is used as tracers in the study of mechanism of chemical reactions
ii. High speed deuterons are used in artificial radioactivity.
iii. Deuterium is mainly used in the preparation of heavy water (D_2O). Which is employed a moderator in nuclear reactor to slow down the speed of fast-moving neutrons.

37. Explain preparation of hydrogen using electrolysis.

Obtained by electrolysis of water

i. **Electrolyte:** Water containing traces of acids or alkali or the electrolysis of aqueous solution o sodium hydroxide or potassium hydroxide
ii. **Anode:** Nickel
iii. **Cathode:** Iron
iv. **At anode:** $2OH^- \rightarrow H_2O + ½ O_2 + 2e-$
v. **At Cathode :** $2H_2O + 2e^- \rightarrow 2OH^- + H_2$
vi. **Overall reaction:** $H_2O \rightarrow H_2 + ½ O_2$

38. A groups metal (A) which is present in common salt reacts with (B) to give compound (C) i which hydrogen is present in –1 oxidation state. (B) on reaction with a gas (D) to give universal

solvent (E). The compound (E) on reacts with (A) to give (B), a strong base. Identify A, B, C, D , E and F. Explain the reactions.

(i) Metallic **sodium (A)** reacts with **hydrogen (B)** to give **sodium hydride (C).**

$2\,Na + H_2 \rightarrow 2NaH$

(A) (B) (C)

(ii) Hydrogen reacts with a gas (oxygen) to give **water (D)**

$2H_2 + O_2 \rightarrow 2H_2O$

(B) (D) (E)

(iii) Water reacts with sodium to give **sodium hydroxide (E)** and hydrogen (B)

$2H_2O + 2Na \rightarrow 2NaOH + H_2$

(E) (A) (F) (B)

A	Na	Sodium
B	H_2	Hydrogen
C	NaH	Sodium Hydride
D	O_2	Oxygen
E	H_2O	Water
F	NaOH	Sodium hydroxide

39. An isotope of hydrogen (A) reacts with diatomic molecule of element which occupies group number 16 and period number 2 to give compound (B) is used as a modulator in nuclear reaction. (A) adds on to a compound (C), which has the molecular formula C_3H_6 to give (D). Identify A, B, C and D.

(i) Isotope of hydrogen reacts with which is oxygen to give **heavy water (B)**

$2D_2 + O_2 \rightarrow 2D_2O$

(A) (B)

(ii) Deuterium (A) undergoes addition reaction with **propane (C)** to give **propane deuteride (D)**

$CH_3CH = CH_2 + D_2 \rightarrow CH_3 - CHD - CH_2D$

A	D_2	Deuterium
B	D_2O	Heavy water or deuterium oxide
C	$CH_3 - CH = CH_2$	Propene
D	$CH_3 - CHD - CH_2D$	Propane deuteride

40. NH_3 has exceptionally high melting point and boiling point as compared to those of the hydrides of the remaining element of group 15 - Explain.

- **NH_3** has higher boiling and melting point **compared to** all other hydrides of **group 15 elements** due to ***intermolecular hydrogen bonding.***
- Each molecule can from a maximum of 4 hydrogen bonds but on average 1 hydrogen bond per molecule as there is only one lone pair on NH_3 available for hydrogen bonding.
- Hydrogen bonding is strong intermolecular attraction as H on NH_3 acts like a proton due to partial positive on it whole N atom in another NH_3 molecule, a very strong hydrogen bond is formed.
- Due to much strong intermolecular interactions compared to weaker permanent dipole-dipole interactions between other XH_3 molecules in group 15, large amount of energy are required to overcome the forces, giving it the highest point and highest melting point.

41. Why interstitial hydrides have a lower density than the parent metal.

- ***In interstitial metal hydrides, hydrogen occupies the interstitial sites (voids) of the metal lattice. Because of absorption of hydrogen, the crystal lattice expands. Hence the density of metal hydrides is lower than these of the metal.***
- Hydrogen atom being small in size occupies some in the metallic lattice producing distortion without any change in its type.

- The densities of these hydrides are lower than those of metals from which they are formed since the crystal lattice expands due to the inclusion of dihydrogen.
- ***Metallic hydrides or interstitial hydrides are less dense than the parent metal. This is due to <u>slight expansion of the lattice during the formation of their hydrides.</u>***

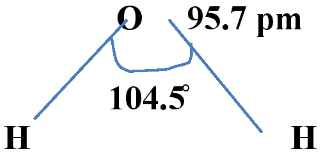
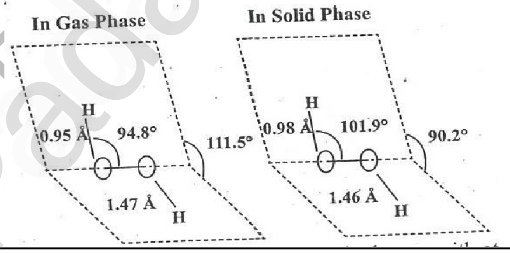
42. How do you expect the metallic hydrides to be useful for hydrogen storage?

- In some of the transition metal hydrides, ***hydrogen is absorbed as H-atoms.***
- Due to the inclusion of H-atoms, the metal lattice expands and thus becomes less stable.
- Therefore, when such metallic hydride is heated, it decomposes to release hydrogen gas and very finely divided metal.
- The hydrogen evolved in this manner can be used as a fuel.
- Thus, transition metals or their alloys can act as sponge and can be used to store and transport hydrogen to be used as a fuel.

43. Arrange NH_3, H_2O and HF in the order of increasing magnitude of hydrogen banding and explain the basis for your arrangement.

i. The order of increasing magnitude of H-bonding is, $\mathbf{NH_3 < H_2O < HF}$
ii. Strength of H bond depends upon the atomic size and electronegativity of the other atom to which H atom is covalently bonded.
iii. Smaller size and higher electronegativity favour H bonding
iv. Among N, F and O the order of electronegativity is **F > O > N.**
v. Hence HF will have highest magnitude of H bonding.

44. Compare the structures of H_2O and H_2O_2.

STRUCTURE OF H_2O	STRUCTURE OF H_2O_2
H_2O has a **bent** structure	H_2O_2 has a **open book** like structure
The **HOH** bond angle is 104.5°	The **H-O-O** bond angle is 94.8° and the dihedral angle 111.5° in gas phase.
It is a polar molecule 	 H_2O_2 has a non polar structure

IMPORTANT QUESTIONS

Lesson 4 HYDROGEN

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
6	6	9	3	3	3	7	9	7	8				

1. Hydrogen Position in periodic table. (101)
2. What are isotopes? Write the names of isotopes of hydrogen.(101) M23 5Mi
3. Write note on ortho hydrogen and para hydrogen. (102)
4. How do you convert parahydrogen into ortho hydrogen?(102) **(S20 3M, S21 2M, M23 3M (What is ortho and para hydrogen? How do you convert para hydrogen into ortho hydrogen? J24 5M)**
5. What is the difference between Ortho hydrogen and para hydrogen. (103)
6. Preparation of hydrogen i) from water ii) Lab preparation **S20 5Mi** iii) industrial production. (103)
7. What is water gas shift reaction. (103) **J23 5Mi, J24 2M**
8. Preparation of Deuterium from heavy water. (104)

9. Preparation of Tritium. (104) **(M19 3MARK, J23 2M**
10. Chemical properties of Hydrogen. I)O_2 ii) X_2 iii) Li, Na, Ca iv) ethyne (104)
11. Explain the exchange reactions of deuterium. (105) **(MQ19, S20 2M)**
12. Properties of Tritium. (105)
13. Explain uses of hydrogen. (105) **M24 3M**
14. Chemical properties of water i) Na, Ba, Fe ii) Cl_2, F_2 (108)
15. Water is an amphoteric oxide. why? (108)
16. What it temporary hardness of water? How is it removed?(109)
17. What is permanent hardness of water? How it will be removed?(109)
18. Chemical properties of heavy water. I) NaOH, HCl, NH_4Cl ii) H_3PO_2 iii)Al_4C_3, CaC_2, Mg_3N_2, Ca_3P_2(111)
19. Do you think that heavy water can be used for drinking purpose? (BB)
20. What are the uses of heavy water. (111) **J23 3M**
21. Preparation of hydrogen peroxide from BaO_2, Na_2O_2, 2-ethylanthraquinone (111)
22. Complete the following equation. $Na_2O_2 + ? \rightarrow Na_2SO_4 + H_2O_2$ (111) **M19 5Mii**
23. Hydrogen peroxide can function as an oxidizing agent as well as reducing agent. Substantiate this statement with suitable examples. BB (112)
24. What are the uses of hydrogen peroxide. (112)
25. Compare the Structures of water (106) and hydrogen peroxide. (112)
26. Mention the three types of covalent hydrides. (113) **M22 2M**
27. Give an example for ionic hydride and covalent hydride. (113) **J22 2M**
28. What are Interstitial hydrides? Give an example. (Metallic hydrides.) (113) **M24 5Mi**
29. Explain types of hydrogen bonding. (114)
30. How do you expect the metallic hydrides to be useful for hydrogen storage? (117)
31. Explain hydrogen -future fuel. (117)

& study all book back question & answer

LESSON 5 Alkali and Alkaline earth metals

26. Why sodium hydroxide is much more water soluble than chloride?

NaOH , NaCl

- In both case Na^+ is common. In NaOH, OH^- is the conjugate base of weak acid H_2O, **OH^- is strong base**.
- In NaCl, Cl^- is the conjugate base of strong acid HCl. Hence **Cl^- is weak base.**
- So, sodium hydroxide is much more water soluble than chloride

(NaOH is much more soluble than NaCl. Enthalpy of a solution can be expressed as the sum of lattice enthalpy and enthalpy of hydration of a compound. Dissolution of NaCl is accompanied by very small heat change so solubility of NaCl is less than NaOH.)

27. Write the chemical equations for the reactions involved in solvay process of preparation of sodium carbonate.

The equations involved in solvay process are,

$2NH_3 + H_2O + CO_2 \rightarrow (NH_4)_2CO_3$ Ammonium carbonate

$(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$ ammonium bicarbonate

$2NH_4HCO_3 + NaCl \rightarrow NH_4Cl + NaHCO_3$ sodium bicarbonate

$2NaHCO_3 \rightarrow \underset{\text{Sodium carbonate}}{Na_2CO_3} + CO_2 + H_2O$

The ammonia used in this process can be recovered by treating the resultant ammonium chloride solution with calcium hydroxide. Calcium chloride is formed as a by-product.

28. An alkali metal (x) forms a hydrated sulphate, $X_2SO_4 . 10H_2O$. Is the metal more likely to be sodium (or) potassium.

The metal more likely to form a hydrated sulphate is **sodium** of formula $Na_2SO_4.10H_2O$. it is otherwise called as Glauber's salt.

Reason:

Smaller the size of the ion greater is the degree of hydration. Hydration energy is in the order of $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$ so sodium is hydrated more easily than potassium.

29. Write balanced chemical equation for each of the following chemical reactions.

(i) Lithium metal with nitrogen gas **(ii) heating solid sodium bicarbonate**

(iii) Rubidum with oxygen gas **(iv) solid potassium hydroxide with CO_2**

(v) heating calcium carbonate **(vi) heating calcium with oxygen**

(i) $6Li_{(s)} + N_{2(g)} \rightarrow 2Li_3N_{(s)}$

(ii) $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

(iii) $Rb + O_2 \rightarrow Rb_2O$

(iv) $KOH + CO_2 \rightarrow KHCO_3$

(v) $CaCO_3 \leftrightarrow CaO + CO_2$

(vi) $2Ca + O_2 \rightarrow 2CaO$

30. Discuss briefly the similarities between beryllium and aluminium.

- Beryllium shows a diagonal relationship with aluminium.
- In this case the size of these ions is not as close.
- However, their charge per unit area and electro-negativity values are almost similar. Similarities between Beryllium and Aluminium.

	Properties
1	Beryllium chloride forms a **dimeric structure** like aluminium chloride with chloride bridges. Both are soluble in organic solvents and are strong Lewis acids.
2	Beryllium hydroxide dissolves in excess of alkali and gives beryllate ion $[Be(OH)_4]^{2-}$ as aluminium hydroxide which gives aluminate ion, $[Al(OH)_4]^-$.
3	Beryllium and Aluminium ions have strong tendency to form complexes, BeF_4^{2-}, AlF_6^{3-}
4	Both beryllium and aluminium hydroxides are **amphoteric** in nature.
5	Carbides of beryllium (Be_2C) like aluminium carbide (Al_4C_3) give methane on hydrolysis
6	Both beryllium and aluminium are rendered passive by nitric Acid.

31. Give the systematic names for the following (i) milk of magnesia (ii) lye (iii) lime (iv) Caustic potash (v) washing soda (vi) soda ash (v) trona

(i) Magnesium hydroxide $Mg(OH)_2$

(ii)	Sodium hydroxide	$NaOH$
(iii)	Calcium oxide	CaO
(iv)	Potassium hydroxide	KOH
(v)	Sodium carbonate decahydrate	$Na_2CO_3.10H_2O$
(vi)	Sodium carbonate	Na_2CO_3
(vii)	Sodium sesquicarbonate	$Na_2CO_3.NaHCO_3.2H_2O$ (mineral)

32. Substantiate Lithium fluoride has the lowest solubility among group one metal fluorides.

(i) The solubility of alkali metal fluorides is in the order $LiF < NaF < KF < RbF < CsF$.

(ii) The solubility of LiF is due to its Very **high lattice energy** because of small sizes of both Li^+ and F^-.

33. Mention the uses of plaster of Paris.

(i) In the building industry as well as plasters.

(ii) It is used for immobilizing the affected part of organ where there is abone fracture.

(iii) In dentistry, in ornamental work.

iv)For making casts of statues and busts.

34. Beryllium halides are Covalent whereas magnesium halides are ionic why?

- Due to small size of Be^{2+}, the charge density is very high.
- According to Fajan's Rule, cation with small size has high polarizing power and prefers to form covalent bonds.
- So beryllium halides are Covalent in nature.
- Whereas magnesium ion(Mg^{2+}) is bigger and it is involved in transfer of electrons to form ionic bond.

35. Alkaline earth metal (A), belongs to 3rd period reacts with oxygen and nitrogen to form compound (B) and (C) respectively. It undergo metal displacement reaction with $AgNO_3$ solution to form compound (D).

(i) Alkaline earth metal (A) belonging to 3rd period is magnesium

(ii) Magnesium reacts with oxygen and nitrogen to form magnesium oxide (B) and magnesium nitride (C).

$$2Mg + O_2 \rightarrow 2MgO$$

$$3Mg + N_2 \rightarrow Mg_3N_2$$

(iii) Magnesium undergoes metal displacement reaction with $AgNO_3$ to form magnesium nitrate (D). $Mg + 2AgNO_3 \rightarrow Mg(NO_3)_2 + 2Ag$

Compound / Element	**Formula**	**Name**
A	Mg	Magnesium
B	MgO	Magnesium oxide
C	Mg_2N_2	Magnesium nitride
D	$Mg(NO_3)_2$	Magnesium nitrate

36. Write balanced chemical equation for the following processes

(a) heating calcium in oxygen (b) heating calcium carbonate
(c) evaporating a solution of calcium hydrogen carbonate
(d) heating calcium oxide with carbon

(a) $2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$
(b) $CaCO_3 \leftrightarrow CaO + CO_2$
(c) $Ca(HCO_3)_2 \rightarrow CaCO_3 + CO_2 + H_2O$
(d) $2CaO_{(s)} + 5C_{(s)} \rightarrow 2CaC_{2(s)} + CO_{2(g)}$

37. Explain the important common features of Group 2 elements.

- Group 2 elements are known as alkaline earth metals.
- It includes beryllium, magnesium, calcium, strontium, barium, and radium.
- They exist in +2 oxidation states.
- Their general electronic configuration is **ns^2.**
- Atomic and ionic radii of alkaline earth metals are smaller than alkali metals, **on moving down the group, the radii increases.**
- Alkaline earth metals have **higher ionization enthalpy** values than alkali metals and they are **less electropositive than** alkali metals.
- Hydration enthalpies of alkaline earth metals decrease down the group.
- Alkaline earth metal salts moistened with conc HCl gave a characteristic coloured flame, when heated on a platinum wire in a flame.
- They are silvery, white, and hard metals. They are softer but harder than alkali metals.
- Their melting and boiling points are higher compared to alkali metals.
- They are strongly electropositive in nature Alkaline earth metals give different color with flame test.

Calcium – Brick red colour.
Strontium – Crimson red colour.
Barium – Apple Green colour.

38. Why alkaline earth metals are harder than alkali metals.

- Due to the presence of 2 electrons in its outermost shell as compared to alkali metals, which have only 1 electron in valence shell.
- The alkaline earth metals have greater nuclear charge and more valence electrons, thus metallic bonding is more effective. Due to this they are harder than alkali metals.

39. How is plaster of paris prepared?

Plaster of pairs is obtained when gypsum, $CaSO_4.2H_2O$, is heated to 393K
$2(CaSO_4.2H_2O) \rightarrow 2CaSO_4.H_2O + 3H_2O$

40. Give the uses of gypsum.

- Gypsum is used in making drywalls or plaster boards.
- It is used in the production of plaster of Paris, which is used as a sculpting material.
- It is used in making surgical and orthopedic cats.
- It plays an important role in agriculture as a soil additive, conditioner, and fertilizer.

- It is used in toothpastes, shampoos, and hair products.
- It is used in baking as a dough conditioner.
- It is used to give colour to cosmetics and drugs.
- It plays a very important role in wine making.

41. Describe briefly the biological importance of Calcium and magnesium.

(i) Magnesium plays an important role in many biochemical reactions catalyzed by enzymes.

(ii) Magnesium is the co-factor of all enzymes that utilize ATP in phosphate transfer and energy release.

(iii) Magnesium also essential for DNA synthesis and is responsible for the stability and proper functioning of DNA.

(iv) Magnesium is also used for balancing electrolytes in our body.

(v) Deficiency of magnesium results into convulsion and neuromuscular irritation.

(vi) Calcium is a major component of bones and teeth.

(vii) Calcium is also present in blood and its concentration is maintained by hormones (calcitonin and parathyroid hormone).

(viii) Deficiency of calcium in blood causes it to take longer time to clot.

(ix) Calcium is also important for muscle contraction.

(x) Chlorophyll, contains magnesium which plays an important role in photosynthesis.

42. Which would you expect to have a higher melting point, magnesium oxide or magnesium fluoride? Explain your reasoning

- Magnesium oxide has very strong ionic bonds as compared to magnesium fluoride.
- Mg^{2+} and O^{2-} have charges of +2 and -2 respectively.
- Oxygen ion is smaller than fluoride ion.
- The smaller the ionic radii, the smaller the bond length in MgO and the bond is stronger than MgF_2.
- Due to more strong bond nature in MgO, it has high melting point than MgF_2.

Lesson 5 ALKALI AND ALKALINE EARTH METALS

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
7	8	4	7	7	7	5	6	8	7				

1. General characteristics of alkali metals. **Or** 1s block elements.(126)
2. Distinctive (anomalous) behavior of Lithium. (129)
3. Discuss the Similarities between Lithium and Magnesium. (129) **J22 5M**
4. Among the alkali metal halides, which is covalent? Explain with reason.(130)**(J19)**
5. Why blue colour appears during the dissolution of alkali metals in liquid ammonia?(130) **J19 5Mi**
6. Uses of alkali metals. (131)
7. Except LiF, all other halides are soluble in water. Why? (132)
8. Write the chemical equations for the reactions involved in **Solvay process** of preparation of sodium carbonate (washing soda). (132)

9. What is soda ash? Write equations. (133)
10. Uses of washing soda? (133)
11. Explain preparation, uses of cooking salt. (133,134)
12. Explain preparation, uses of sodium hydroxide. (134)
13. What are the uses of sodium bicarbonate. (134) **J23 5Mii**
14. Explain biological importance of sodium and magnesium. (135)
15. General characteristics of alkaline earth metals. Or 2s block elements(136)
16. Distinctive (anomalous) behavior of Beryllium. (139)
    (What are the reasons for the anomalous properties of Beryllium? (139)**M22 5Mi)**
17. Explain the properties of Beryllium that are difference from other elements of the group. 139 **M22 5Mii**
18. Discuss the similarities between beryllium and aluminum. **(140) (MQ19, J19, S21 5M, J23 3M, M24 5M**
19. Among the alkaline earth metals BeO is insoluble in water but other oxides are soluble. Why? 142 **(M19 5Mi)**

    ANSWER:

    (i) **BeO is insoluble in water**
    BeO is covalent in nature, While other alkaline earth metal oxides are ionic in nature

20. Uses of alkaline earth metals Mg **(S20 5Mii),** Ca **M23 5Mii**, Sr, Ba, Ra. (141)
21. Explain preparation, properties and uses of Quick lime. (144)
22. Explain preparation, properties and uses of calcium hydroxide. (144)
23. How is bleaching powder prepared? 145 **S20 5Mi, J24 3M**
24. Explain why $Ca(OH)_2$ is used in white washing. (145) **M19 3M**
25. Explain properties and uses of Gypsum. (145)
26. Plaster of Paris preparation and uses. (147) **(MQ19) M23 2M**
27. Explain biological importance of magnesium and **calcium.** **(148) J24 5Mi**
28. Write the balanced equation for each of the following chemical reactions.
    (i) Reaction of metallic Lithium with Nitrogen gas **(MQ19) BB 154**
    (ii) Heating solid sodium bicarbonate.
    (iii) Rubidium with oxygen gas.
    (iv) Solid potassium hydroxide with CO_2
    (v) Heating calcium carbonate. **M24 2M**
    Heating calcium with oxygen. **& study all book back question & answer**

LESSON 6 GASEOUS STATE

Answer these questions briefly

26. State Boyle's law.

At a given temperature the ***volume occupied by a fixed mass of a gas is inversely proportional to its pressure.*** Mathematically, the Boyle's law can be written as

$$V \alpha \frac{1}{P} \qquad \text{------ (1)}$$

(T and n are fixed, T-temperature, n- number of moles)

$$V = k \times \frac{1}{P} \qquad \text{------ (2)}$$

k – proportionality constant When we rearrange equation (2)

$PV = k$ at constant temperature and mass

27. Name two items that can serve as a model for Gay Lusaac' law and explain.

$P \alpha T$ at constant volume.

(i) Pressure in well inflated tyre is almost constant but when temperature increases in summer days it increases the pressure and sometimes tyres may burst.

(ii) **Guns and other firing equipment's** are thrilling examples of Gay Lussac's law. When gun pin strikes, it ignites the gun powder and this increases the temperature which in turn increases the pressure and bullet is fired from the gun.

iii) **Heating a closed aerosol cane.** The increased pressure may cause the container to explode. You don't toss an "empty" can of hair spray into a fire.

iv) **The egg in the bottle experiment.**

A glass bottle is taken, inside the bottle put some pieces of cotton with fire. Then place a boiled egg (shell removed) at the top of the bottle. The temperature inside the bottle increases from the fire, rising the pressure. By sealing the bottle with egg, the fire goes on, dropping the temperature and pressure. This causes the egg to be sucked into the bottle.

28. Give the mathematical expression that relates gas volume and moles. Describe in words what the mathematical expression means.

- The Mathematical expression between the volume of gas and number of moles is

$V \alpha n$

$V_1/n_1 = V_2/n_2 =$ constant

- **where** V_1 & n_1 are the volume and number of moles of a gas,

V_2 & n_2 are a different set of values of volume and number of moles of the same gas at same temperature and pressure.

- The above relation is expressed as **"Equal volumes of all gases under the same condition of temperature and pressure contain equal number of molecules".**

29. What are ideal gases? In what way real gases differ from ideal gases.

An ideal gas a gas which obeys the gas laws and ideal gas equation

$PV=nRT$.

An ideal gas is defined as **"one in which all collisions between atoms or molecules are perfectly elastic forces".**

The difference between ideal gas and real gas is

Ideal gas	**Real gas**
1. It obeys gas laws under all conditions of temperature and pressure	It **obeys gas laws** only under ***low pressure and high temperature***
2. No gas is ideal	All gases are real
3. Volume occupied by the molecules is negligible as compared to the total volume occupied by the gas.	Volume occupied by the molecules is not negligible as compared to the total volume occupied by the gas.

4. The forces of attraction among the molecules of the gas are negligible.	The forces of attraction among the molecules cannot be neglected at high pressure and low temperature
5. It obeys ideal gas equation PV = nRT	It obeys Van der Waals equation $(P + an^2/V^2)$ (V-nb) = nRT
6. Compressibility factor for ideal gas Z=1.	Compressibility factor for real gas Z >1. Or Z <1
7. Ideal gases **does not** lose energy during collision. Linear	Real gases **lose** energy during collision

30. Can a Van der Waals gas with a=0 be liquefied? explain.

- a=0 for Van der Waals gas i.e. for a real gas.
- Van der Waals constant a=0. ***It cannot be liquefied.***
- The van der Waals constant 'a' is a measure of the attractive forces among the molecules of the gas.
- ***Greater the value of 'a' larger the intermolecular force of attraction*** and the gas can be liquefied.
- Here a=0 means there is intermolecular attraction and the gas **cannot be liquefied.**

31. Suppose there is a tiny sticky area on the wall of a container of gas. Molecules hitting this area stick there permanently. Is the pressure greater or less than on the ordinary area of walls?

- Molecule hitting the tiny sticky area on the wall of the container of gas moves faster as they get closer to adhesive surface, but this effect is not permanent.
- The pressure on the sticky wall is greater than on the ordinary area of walls.

32. Explain the following observations

a) Aerated water bottles are kept under water during summer

- In aerated water bottles the CO_2 is passed through the aqueous solution under pressure.
- The **solubility of gas is decreases with increase of temperature**.
- In summer season the **temperature is raise the solubility is decreases.**
- Due to this will increase very high pressure above the surface of the liquid inside the bottle and bottle will not able to withstand the pressure and bottle may explode.
- To avoid this Aerated water bottles are kept under water during summer.
- As a result, the temperature decreases and solubility of CO_2 is increases in aqueous solution resulting the pressure inside the bottle decreases.

b) Liquid ammonia bottle is cooled before opening the seal

- Liquid ammonia bottle contains the gas under very high pressure. If the bottle is opened, the sudden decrease in pressure will increase the volume of gas.
- As result, the gas is come out the bottle with greater force will cause breakage of bottle and accident.
- The pressure of the gas over liquid ammonia is decreased by dipping H_2O.

- Once it attains the temperature of water, it can be opened so that no gas will come out of the mouth of the tube with force.

c) The tyre of an automobile is inflated to slightly lesser pressure in summer than in winter.

- In summer due to high temperature the air expands and hence to avoid tyre burst, the pressure is kept less.
- As a result, when the vehicle runs and the tyre air gets heated and expands, resulting increased pressure is still kept below the bursting limit.

d) The size of a weather balloon becomes larger and larger as it ascends up into larger altitude.

- According Boyle's law the volume of gas is inversely proportional to the pressure at a given temperature.
- As the weather balloon ascends, the atmospheric pressure is less, pressure of the gas tends to decrease and so volume as well as the size of the balloon increases.

33. Give suitable explanation for the following facts about gases.

a) Gases don't settle at the bottom of a container

- Gases are less denser than solids and liquids. They have negligible intermolecular force of attraction between free particles, so are free to move.
- Hence gases don't settle at bottom of the container.

b) Gases diffuse through all the space available to them

- Gases have the tendency to mix with one another spontaneously and form a homogeneous mixture.
- This is due to the fact that gas particles are mainly in random at very high velocities and there in so much of inter molecular empty space in the volume of any gas.
- This permit them to mix spontaneous and this phenomenon is known as diffusion.

34. Suggest why there is no hydrogen (H_2) in our atmosphere. Why does the moon have no atmosphere?

- Under ordinary conditions on earth, hydrogen exist as diatomic (H_2). Because of its light weight, which enable into escape from earth gravity more easily than heavier gases. So, no hydrogen is there in atmosphere.
- Moon has no atmosphere because the value of acceleration due to gravity (g) on the surface of moon is small. The molecules of atmospheric gases on the surface of the moon have thermal velocities greater than escape velocity.

35. Explain whether a gas approaches ideal behavior or deviates from ideal behaviour if

a) it is compressed to a smaller volume at constant temperature.

When the gas is compressed, there is a decrease in volume and molecules are close to each other and hence inter molecular attraction becomes more and hence it deviates from ideal behavior.

b) the temperature is raised at while keeping the volume constant

- When volume constant, the pressure is directly proportional to the temperature.
- The temperature is raised the pressure is always increases which leads the intermolecular force of attraction between the gas molecules.
- So it deviates from the ideal behavior

c) more gas is introduced into the same volume and at the same temperature.

- If more gas is introduced in the same volume and temperature is constant, more will be the pressure.
- The increase in pressure is always increases which leads the intermolecular force of attraction between the gas molecules.
- So it deviates from the ideal behavior

36. Which of the following gases would you expect to deviate from ideal behaviour under conditions of low temperature F_2, Cl_2, or Br_2? Explain.

- These molecules are held together by a weak van der Waals forces.
- The forces of attraction between the molecules with increase in the size of the molecule.
- ***Br_2 deviate from ideal behavior,*** since ***the Br_2 has the biggest size*** provides maximum attraction between bromine molecules.

37. Distinguish between diffusion and effusion.

Diffusion	**Effusion**
The natural tendency of gases to be intermixed with each other thoroughly	The process in which a gas comes out of the vessel through a small opening or hole.
The spreading of the molecules of a gas throughout the available space or second substance is called diffusion.	Effusion is the escape of gas molecules through a very small hole.
Diffusion refers to the ability of the gases to mix with each other	Effusion is an ability of a gas to travel through a small pin-hole.
3. **Example:** Some particles are dissolved in a glass of water. (spreading of something such as brown tea liquid spreading through the water in a tea cup.)	**Example:** Air escaping from punctured tyre or football bladder. (pouring out something like the soap studs bubbling out from a bucket of water.)

38. Aerosol cans carry clear warning of heating of the can. Why?

- Aerosols are colloids in which air (gas) is dispensed in liquid. On heating the can, the pressure of the gas increases and it can burst out.
  - Hence they carry clear warning that they should not be heated or kept in near fire.

39. Would it be easier to drink water with a straw on the top of Mount Everest?

- Drinking through a straw is slightly more difficult on the top a mountain.
- This is because the atmospheric pressure and the temperature is low at top mountain.
- When you drink through a straw you are decreasing the pressure in your mouth, so the atmospheric pressure pushes the liquid up the straw.

- On mountain there is less pressure so there is less pressure to push the water into the straw.

40. Write the Van der Waals equation for a real gas. Explain the correction term for pressure and volume.

The Vander Waals equation for a real gas is

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

Pressure Correction:

- The *pressure of a gas is directly proportional to the force* created by the bombardment of molecules on the walls of the container.
- The speed of a molecule moving towards the wall of the container is reduced by the attractive forces exerted by its neighbours.
- Hence, the measured gas *pressure is lower than the ideal pressure* of the gas. Hence, Vander Waals introduced a correction term to this effect.
- Van der Waals found out the *forces of attraction experienced by a molecule near the wall are directly proportional to the square of the density of the gas.*

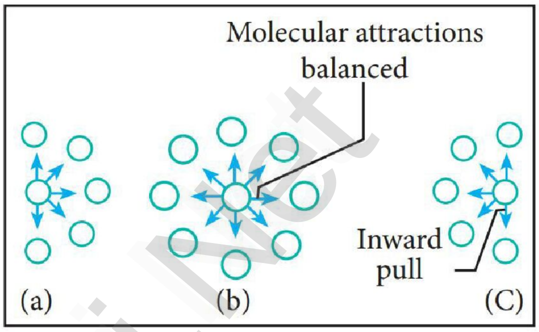


Figure. 6.10 Inter-molecular forces of attraction

$$P` \alpha \rho^2 \qquad \rho = \frac{n}{V}$$

Where n is the number of moles of gas and V is the volume of the container

$$\Rightarrow P' \alpha \frac{n^2}{V^2}$$

$$\Rightarrow P' = \frac{an^2}{V^2}$$

Where a is proportionality constant and depends on the nature of gas

Therefore, $\mathbf{P_{ideal} = P + \frac{an^2}{V^2}}$

Volume Correction:

As every individual molecule of a gas occupies a certain volume, the actual volume is less than the volume of the container, V. Van der Waals introduced a correction factor V' to this effect. Let us calculate the correction term by considering gas molecules as spheres.

V = excluded volume

Excluded volume for two molecules $= \frac{4}{3}\pi(2r)^3$

$$= 8\left(\frac{4}{3}\pi r^3\right) = 8\ V_m$$

Where V is a volume of a single molecule $= \frac{8V_m}{2} =$

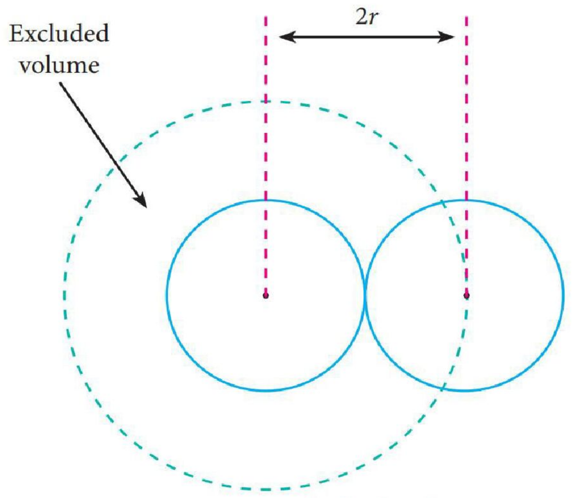


Figure. 6.11 Excluded volume

$4V_m$

Excluded volume for n molecule = $n(4V_m) = nb$

Where b is van der waals constant

Which is equal to $4V_m$ ⇒ $V = nb$

$V_{ideal} = V - nb$

Replacing the corrected pressure and volume in the ideal gas equation PV = nRT we get the van der Waals equation of state for real gases as below,

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

The constants a and b are van der Waals constants and their values vary with the nature of the gas. It is an approximate formula for the non-ideal gas.

41. Derive the values of van der Waals equation constants in terms of critical constants.

The van der waals equation for n moles is

$$\left(p + \frac{an^2}{V^2}\right)(V - nb) = nRT \qquad \text{------- (1)}$$

For 1 mole

$$\left(p + \frac{a}{V^2}\right)(V - b) = RT \qquad \text{------- (2)}$$

From the equation we can derive the values of critical constants Pc, Vc and Tc in terms of a and b, the van der waals constant, on expanding the above equation

$$PV + \frac{a}{V} - Pb - \frac{ab}{V^2} - RT = 0 \qquad \text{--------- (3)}$$

Multiply equation (3) by $\frac{V^2}{P}$

$$\frac{V^2}{P}\left(PV + \frac{a}{V} - Pb - \frac{ab}{V^2} - RT\right) = 0$$

$$V^3 + \frac{aV}{P} + - bV^2 - \frac{ab}{V^2} - \frac{RTV^2}{P} \qquad \text{------- (4)}$$

When the above equation is rearranged in powers of V.

$$V^3 + \left[\frac{RT}{P} + b\right]V^2 + \left[\frac{a}{P}\right]V - \left[\frac{ab}{P}\right] = 0 \qquad \text{------- (5)}$$

When equation (5) is a cubic equation in V. On solving this equation,

We will get three solutions. At the critical point all these three solutions of V are equal to the critical volume Vc. The pressure and temperatures becomes Pc and Tc respectively

i.e., V = Vc;

V – Vc=0

$(V - Vc)^3=0$ $[(a-b)^3 = a3-3a^2b+3ab^2-b^3]$

$$V^3 - 3VcV^2 + 3Vc^2V - Vc^3 = 0 \qquad \text{------- (6)}$$

As equation (5) is identical with equation (6), we can equate the coefficients of V2, V and constant terms in (5) and (6)

$$-3VcV^2 = -\left[\frac{RT_c}{P_c} + b\right]V^2$$

$$3Vc = \frac{RT_c}{P_c} + b \qquad \text{--------- (7)}$$

$$3Vc^2 = \frac{a}{P_c} \qquad \text{--------} \ (8)$$

$$3Vc^2 = \frac{ab}{P_c} \qquad \text{-------}(9)$$

Divide equation (9) by equation (8)

$$\frac{V_c{}^3}{3V_c{}^2} = \frac{ab/P_c}{a\,/P_c}$$

$$\frac{V_c}{3} = b$$

i.e., $Vc = 3b$ -------- (10)

when equation (10) is substituted in (8)

$$3Vc^2 = \frac{a}{P_c}$$

$$Pc = \frac{a}{3Vc^2} = \frac{a}{3(3b^2)} = \frac{a}{3 \times 9b^2} = \frac{a}{27b^2}$$

$$Pc = \frac{a}{27b^2} \qquad \text{-------} \ (11)$$

Substituting the values of Vc and Pc in equation (7),

$$3Vc = b + \frac{RT_c}{P}$$

$$3(3b) = b + \frac{RT_c}{\left(\frac{a}{27b^2}\right)}$$

$$9b - b = \left(\frac{RT_c}{a}\right) 27b^2$$

$$8b = \frac{T_c R 27b^2}{a}$$

$$Tc = \frac{8ab}{27Rb^2} = \frac{8a}{27Rb} \qquad \text{--------} \ (12)$$

The critical constants can be calculated using the values of van der waals constant of a gas and vice versa. $a = 3Vc^2P_c$ and $b = \frac{V_c}{3}$

42. Why do astronauts have to wear protective suits when they are on the surface of moon?

- Astronauts must wear space suits since the surface of moon, there is no air to breath and no air pressure.
- Space is extremely cold and filled with dangerous radiation.
- Space suits are specially designed to protect astronauts from the cold, radiation and low pressure in space. It also provide air to breathe.

43. **When ammonia combines with HCl, NH_4Cl is formed as white dense fumes. Why do more fumes appear near HCl?** $NH_3 + HCl \rightarrow NH_4Cl$

- The *rate of diffusion is inversely proportional to the molecular weight of the gas.*
- Lower the molecular weight faster is the diffusion.
- The molar mass of HCl is 36.5 $g.mol^{-1}$ while the molar mass of NH_3 is 17 $g.mol^{-1}$.
- Hence NH_3 diffuses faster than HCl. Hence white fumes appear near HCl.

44. A sample of gas at 15 °C at 1 atm. has a volume of 2.58 dm³. When the temperature is raised to 38 °C at 1 atm does the volume of the gas increase? If so, calculate the final volume.

$T_1 = 15\ °C + 273$ $T_2 = 38 + 273$

$T_1 = 288\ K$ $T_2 = 311\ K$

$V_1 = 2.58\ dm^3$ $V_2 = ?$

($P = 1$ atm constant)

$$\frac{V_1}{T_1} = \frac{V_2}{V_1}; \qquad V_2 = \left(\frac{V_1}{T_1}\right) \times T_2$$

$$= \frac{2.58 dm^3}{288\ K} \times 311\ K$$

$V_2 = 2.78\ dm^3$ i.e. volume increased from 2.58 m³ to 2.78 dm³

45. A sample of gas has a volume of 8.5 dm³ at an unknown temperature. When the sample is submerged in ice water at 0 °C, its volume gets reduced to 6.37 dm³. What is its initial temperature?

$V_1 = 8.5\ dm^3$ $V_2 = 6.37\ dm^3$

$T_1 = ?$ $T_2 = 0\ °C = 273\ K$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}; \qquad V_1 \times \left(\frac{T_1}{V_2}\right) = T_2$$

$$T_1 = 8.5\ dm^3 \times \frac{273\ K}{6.37\ dm^3} \qquad \mathbf{T_1 = 364.28K}$$

46. Of two samples of nitrogen gas, sample A contains 1.5 moles of nitrogen in a vessel of volume of 37.6 dm³ at 298K, and the sample B is in a vessel of volume 16.5 dm³ at 298K. Calculate the number of moles in sample B.

$n_A = 1.5$ mol $n_B = ?$

$V_A = 37.6\ dm^3$ $V_B = 16.5\ dm^3$

($T = 298$ K constant)

$$\frac{V_A}{n_A} = \frac{V_B}{n_B}; \qquad n_A = \left(\frac{n_A}{V_A}\right) V_B$$

$$= \frac{1.5\ mol}{37.6\ dm^3} \times 16.5\ dm^3 \qquad \mathbf{n_A = 0.66\ mol.}$$

47. Sulphur hexafluoride is a colourless, odourless gas; calculate the pressure exerted by 1.82 moles of the gas in a steel vessel of volume 5.43 dm³ at 69.5 °C, assuming ideal gas Behavior

$n = 1.82$ mole; $V = 5.43\ dm^3$; $T = 69.5 + 273 = 342.5K$ $P = ?$ $PV = nRT$

$$P = \frac{nRT}{V}$$

$$P = \frac{1.82\ \text{mole} \times 0.821 dm^3\ \text{atm}\ mol^{-1} K^{-1} \times 342.5\ K}{5.43\ dm^3} \quad \mathbf{P = 9.425\ atm}$$

48. **Argon is an inert gas used in light bulbs to retard the vaporization of the tungsten filament. A certain light bulb containing argon at 1.2 atm and 18°C is heated to 85°C at constant volume. Calculate its final pressure in atm**.

$P_1 = 1.2$ atm ; $T_1 = 18°C + 273 = 291\ K$

$P_2 = ?$ $T_2 = 85°C + 273 = 358\ K$

$$\frac{\mathbf{P_1}}{\mathbf{T_1}} = \frac{\mathbf{P_2}}{\mathbf{T_2}}$$

$$P_2 = \left(\frac{P_1}{T_1}\right) \times T_2 = \frac{1.2\ atm}{291\ K} \times 358\ K \quad \mathbf{P_2 = 1.48\ atm.}$$

49. A small bubble rises from the bottom of a lake where the temperature and pressure are 6°C and 4 atm. to the water surface, where the temperature is 25°C and pressure is 1 atm. Calculate the final volume in (mL) of the bubble, if its initial volume is 1.5 mL.

$T_1 = 6°C + 273 = 279\ K$;

$T_2 = 25°C + 273 = 298\ K$

$P_1 = 4$ atm $V_1 = 1.5$ ml $P_2 = 1$ atm $V_2 = ?$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2} = \frac{4\ atm \times 1.5 ml \times 298\ K}{279\ K \times 1\ atm} \qquad \mathbf{V_2 = 6.41\ ml.}$$

50. Hydrochloric acid is treated with a metal to produce hydrogen gas. Suppose a student carries out this reaction and collects a volume of 154.4 × 10^{-3} dm^3 of a gas at a pressure of 742 mm of Hg at a temperature of 298 K. What mass of hydrogen gas (in mg) did the student collect?

V = 154.4 × 10^{-3} dm^3; P = 742 mm of Hg; T = 298 K ; m = ?

$$\mathbf{n = \frac{PV}{RT}} = \frac{742\ mm\ \ Hg \times 154.4 \times 10{-}3\ dm^3 L}{62mm\ Hg\ LK^{-1}\ mol^{-1} \times\ 298K}$$

$$n = \frac{Mass}{Molar\ Mass}$$

Mass = n × Molar Mass

= 0.006 × 2.016 = 0.0121g **= 12.1 mg.**

51. It takes 192 sec for an unknown gas to diffuse through a porous wall and 84 sec for N_2 gas to effuse at the same temperature and pressure. What is the molar mass of the unknown gas?

$$\frac{\gamma_{unknown}}{\gamma_{N_2}} = \frac{t_{N_2}}{t_{unknown}} = \sqrt{\frac{m_{N_2}}{m_{unknown}}}$$

$$\frac{84\ sec}{192\ sec} = \sqrt{\frac{14g\ mol^{-1}}{m_{unknown}}} = \frac{14g\ mol^{-1}}{m_{unknown}}$$

$$m_{unknown} = 14g\ mol^{-1} \times \left(\frac{192\ sec}{84\ sec}\right)^2$$

$\boldsymbol{m_{unknown}}$ **= 146.28g mol^{-1}**

52. **A tank contains a mixture of 52.5 g of oxygen and 65.1 g of CO_2 at 300 K the total pressure in the tanks is 9.21 atm. Calculate the partial pressure (in atm.) of each gas in the mixture.** m_{o_2} = 52.5 g ; P_{o_2} = ?

m_{co_2} = 65.1 g ; P_{co_2} = ?

T = 300 K P = 9.21 atm

$P_{o_2} = X_{o_2}$ × Total Pressure

$$\boldsymbol{X_{o_2} = \frac{n_{o_2}}{n_{o_2}+n_{co_2}}}$$

$$n_{o_2} = \frac{Mass\ of\ O_2}{Molar\ Mass\ of\ O_2}$$

$$= \frac{52.5\ g}{32g\ mol^{-1}} = 1.64\ \text{mol}$$

$$n_{co_2} = \frac{Mass\ of\ CO_2}{Molar\ Mass\ CO_2}$$

$$= \frac{65.1\ g}{44g\ mol^{-1}} = 1.48\ \text{mol}$$

$$n_{o_2} = \frac{P_1V_1}{P_2} = \frac{1.64}{3.12} = 0.53$$

$$\boldsymbol{X_{co_2}} = \frac{\boldsymbol{n_{co_2}}}{\boldsymbol{n_{o_2}+n_{co_2}}} = \frac{1.48}{3.12} = 0.47$$

$P_{o_2} = X_{o_2}$ × Total Pressure

= 0.53 × 9.21 atm = 4.88 atm

$P_{co_2} = X_{co_2}$ × Total Pressure

= 0.47 × 9.21 atm = 4.33 atm

53. A combustible gas is stored in a metal tank at a pressure of 2.98 atm at 25 °C. The tank can withstand a maximum pressure of 12 atm after which it will explode. The building in which the tank has been stored catches fire. Now predict whether the tank will blow up first or start melting? (Melting point of the metal = 1100 K).

Pressure of the gas in the tank at its melting point

T_1 = 2.98K ; P_1 = 2.98 atm ; T_2 = 1100 K ; P_2 = ?

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$; $P_2 = \frac{P_1}{T_1} \times T_2$ $P_2 = \frac{2.98\ atm}{298\ K} \times 1100K = 11$ atm

At 1100 K the pressure of the gas inside the tank will become 11 atm. Given that tank can withstand a maximum pressure of 12 atm, the tank will start melting first

IMPORTANT QUESTION

Lesson 6 GASEOUS STATE

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
5	9	5	4	5	4	4	11	6	8				

1. What is difference between gas and vapour. (159)
2. Boyle's experiment. (160)
3. State Boyle's law. (160) **J24 5Mii**
4. All passenger aeroplane cabins have to be artificially pressurized? (161)
5. State Charles law.(162)
6. Inside a certain automobile engine, the volume of air in a cylinder is 0.375dm^3 when the pressure is 1.05atm. when the gas is compressed to a volume of 0.125dm^3 at the same temperature. What is the pressure of the compressed air? (162) **S21 3M COMPULSORY**
7. Find the missing parameters (164) **(MQ19)**

P=1 atm	P=1 atm	P= 1 atm
V_1=0.3dm^3	V_2 = ?	V_3=0.15dm^3
T_1 = 200K	T_2=300K	T_3=? K

8. State Gay Lusaac' law. (164)
9. State Avogadro's law. (165)
10. Give the mathematical expression that relates gas volume and moles. (165) **J24 2M**
11. Derive ideal gas equation. (165) **J19 5M, M23 3M**
12. What are ideal gases? (165) **(M19 2M)**
13. State Dalton's law of partial pressures. (166) **M22 3M, M24 2M**
14. State Grahams law of diffusion. (168) (state Diffusion law). **M19 5Mii, J23 2M**
15. Distinguish between diffusion and effusion. (168) **MQ19, J22 3M**
16. Define compressibility factor. **(169)**
17. Write the mathematical formula for compressibility factor Z. (169) **S20 5Miii**
18. Define Boyle temperature or Boyle point. (171)
19. Compressibility factor for real gases. (171)
20. Write the Vander Waals equation for a real gas. Explain the correction term for pressure and volume. (171) **M24 5Mii**(volume correction)
21. Write the Vander Waals equation for real gases and explain the terms involved. (171) **J24 3M**
22. Define critical temperature, critical pressure, critical volume. (173)
23. Derive the values of critical constants in terms of Vander Waals constants. (174) **M23 5M**
24. Define Joule-Thomson effect. (175) **J23 5Mi**

25. Define inversion temperature. (175)
26. What are the methods used for liquefaction of gases. (175) **S20 5Mii, J23 3M**
27. A sample of gas at 15 °C at 1 atm. has a volume of 2.58 dm^3. When the temperature is raised to 38 °C at 1 atm does the volume of the gas increase? If so, calculate the final volume. (183 BBQ44) **J23 5Mii**

Study lesson EXAMPLE problem and Evaluation & study all book back question & answer

7. THERMODYNAMICS

Question and Answer

26. State the first law of thermodynamics.

- **"The total energy of an isolated system remains constant though it may change from one form to another."**
- Energy can neither be created nor destroyed, but may be converted from one form to another.

The mathematical statement of the Frist Law is:

$$\Delta U = q + w$$

Where, q- the amount of heat
w – work done on the system

27.Define Hess's law of constant heat summation.

The enthalpy changes of a reaction either at constant volume or constant pressure is the same whether it takes place in a single or multiple steps provided the initial and final states are same.

$$\begin{tikzcd} A \arrow[r, "\Delta H_r"] \arrow[d, "\Delta H_1"'] & B \arrow[d, "\Delta H_3"] \\ X \arrow[r, "\Delta H_2"] & Y \end{tikzcd}$$

$$\Delta H_r = \Delta H_1 + \Delta H_2 + \Delta H_3$$

28. Explain intensive properties with two examples.

The property that is ***independent of the mass or the size of the system*** is called an intensive property.

Examples: Density, Temperature, Refractive index

29. Define the following terms:

(a) isothermal process (b) adiabatic process (c) isobaric process (d) isochoric process

(a) Isothermal process: An isothermal process is defined as ***one in which the temperature of the system remains constant, during the change from its initial to final state.***

For an isothermal process dT = 0

(b) Adiabatic process

An adiabatic process is defined as one in which there is ***no exchange of heat*** (q) between the system and surrounding during the process.

For an adiabatic process q = 0

(c) Isobaric process

An isobaric process is defined as one in which the ***pressure of the system remains constant*** during its change from the initial to final state.

For an Isochoric process dP = 0

(d) Isochoric process

An isochoric process is defined as the one in which the ***volume of system remains constant*** during its change from initial to final state.

For an isochoric process, dV = 0

30. What is the usual definition of entropy? What is the unit of entropy?

(i) Entropy is a measure of the molecular disorderliness (randomness) of a system. $dS = dq_{rev} / T$

(ii) The entropy (S) is equal to heat energy exchanged (q) divided by the temperature (T) at which the exchange takes place.

The SI unit of entropy is **JK^{-1}**

31.**Predict the feasibility of a reaction when i) both ΔH and ΔS positive ii) both ΔH and ΔS negative iii) ΔH decreases but ΔS increases**

(i) When both ΔH and ΔS are positive, **the reaction is not feasible.** (Non-spontaneous)

(ii) When both ΔH and ΔS are negative, **the reaction is not feasible.** (Non-spontaneous)

(iii) When ΔH decreases but ΔS increases, **the reaction is feasible.** (Spontaneous)

32.Define is Gibb's free energy.

- The isothermally available energy to do work in a system.
- This quantity is the energy associated with a chemical reaction that can be used to do work.
    - Gibb's free energy is defined as **G = H – TS**

Where, H = enthalpy or heat content of the system, T = Temperature in Kelvin and S = Entropy

33.Define enthalpy of combustion.

The heat of combustion of a substance is defined as

"The change in enthalpy of a system when one mole of the substance is completely burnt in excess of air or oxygen".

For example, the heat of combustion of methane is $-87.78 kJmol^{-1}$

It is denoted by ΔH_c

34.Define molar heat capacity. Give its unit.

Molar heat capacity is defined as "***The amount of heat absorbed by one mole of the substance to raise its temperature by 1 kelvin".***

The SI unit of molar heat capacity is $JK^{-1} mol^{-1}$

35. Define the calorific value of food. What is the unit of calorific value?

The calorific value is defined as ***"The amount of heat produced in calories (or joules) when one gram of the substance is completely burnt."***

The SI unit of calorific value is **J Kg^{-1}.** It is usually expressed in cal g^{-1}

36. Define enthalpy of neutralization.

The heat of neutralization is defined as ***"The change in enthalpy when one gram equivalent of an acid is completely neutralized by one gram equivalent of a base or vice versa in dilute solution".***

37. What is lattice energy?

Lattice energy is defined as ***"The amount of energy required to completely remove the constituent ions from its crystal lattice to an infinite distance".*** It is also referred as lattice enthalpy.

38. What are state and path functions? Give two examples.

(i) State Function: A state function is a thermodynamic property of a system, which has a ***specific value*** for a given state and does not depend on the path (or manner) by which the particular state is reached.

Example: Pressure (P), Volume (V), Temperature (T)

(ii) Path Functions: A path function is a thermodynamic property of the system whose value depends on the path by which the system changes from its initial to final states.

Examples: Work (w), Heat (q).

39. Give Kelvin statement of second law of thermodynamics.

It is impossible to construct a machine that absorbs heat from a hot source and converts it completely into work by a cyclic process without transferring a part of heat to a cold sink.

40. The equilibrium constant of a reaction is 10, what will be the sign of ΔG? Will this reaction be spontaneous?

Given $K_{eq} = 10$

Gas constant $R = 8.314\ JK^{-1}\ mol^{-1}$

$T = 300K$

The relationship between Free energy change ΔG and equilibrium constant K is **$\Delta G = -RT \ln K$**

Since K, T and Rare positive values, **ΔG** will be negative. When ΔG is –ve, the process is ***spontaneous and feasible.***

41. Enthalpy of neutralization is always a constant when a strong acid is neutralized by a strong base: account for the statement.

Strong acids and strong bases exist in the fully ionized from in aqueous solutions as below:

$H_3O^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl + 2H_2O$

(or)

$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2H_2O\ (1)^{\circ} \qquad \Delta H = -57.32\ KJ$

The H^+ ions produced in water by the acid molecules exist as H_3O^+. Thus, enthalpy change per mole of water formed from H_3O^+ and OH^- ions. Therefore, irrespective of the chemical nature, the enthalpy of neutralization of strong acid by strong base is a constant value which is equal to -57.32 KJ.

42. State the third law of thermodynamics.

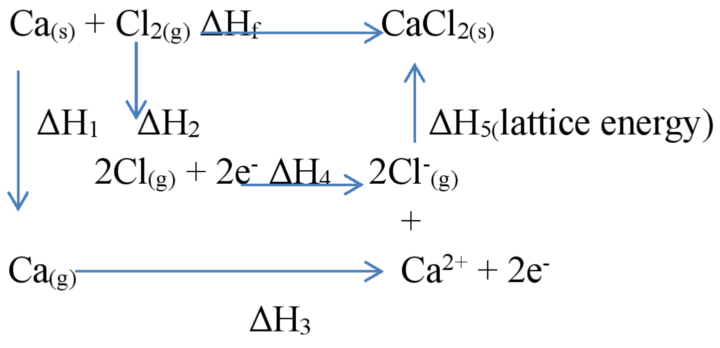
(i) The third law of thermodynamics states that the entropy of pure crystalline substance at absolute zero is zero.

(ii) It can also be stated as it is impossible to lower the temperature of an object to absolute zero in a finite number of steps.

(iii) Mathematically, $\lim_{T\to 0} S = 0$ for a perfectly ordered crystalline state.

43. Write down the Born-Haber cycle for the formation of $CaCl_2$

- Born – Haber cycle for the formation of $CaCl_2$
- Born – Haber cycle is used to calculate the lattice enthalpy of $CaCl_2$.



$\Delta H_f = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5$

44.Identify the state and path functions out of the following: a)Enthalpy b)Entropy c) Heat d)Temperature e) Work f)Free energy.

State Function: Enthalpy, entropy, temperature, free energy

Path Function: Heat, **Work**.

45. State the various statements of second law of thermodynamics.

(i) Kelvin-Planck statement: It is impossible to construct a machine that absorbs heat from a hot source and converts it completely into work by a cyclic process without transferring a part of heat to a cold sink.

(ii) **Clausius statement:** It is impossible to transfer heat from a cold reservoir to a hot reservoir without doing some work.

(iii) **Entropy statement:** The entropy of an isolated system increases during a spontaneous process.

46. What are spontaneous reactions? What are the conditions for the spontaneity of a process?

(i) Spontaneous reaction: A reaction that occurs under the given set of conditions without any external driving force is called a spontaneous reaction.

Increase in randomness favours a spontaneous change.

(ii) **Criteria for spontaneity of a process:** The spontaneity of any process depends on three different factors. ΔH = -ve, ΔS = +ve, ΔG = -ve.

47. List the characteristics of internal energy.

- The internal energy of a system is an extensive property

- The internal energy of a system is a state function
- The change in internal energy of a system is expressed as $\Delta U = U_f - U_i$
- In a cyclic process, there is no internal energy change. $\Delta U_{(cyclic)} = 0$
- If the internal energy of the system in the final state (U_f) is less than the internal energy of the system in its initial state (U_i), then ΔU would be **negative.**

$$\Delta U = U_f - U_i = -ve\ (U_f < U_i)$$

- If the internal energy of the system in the final state (U_f) is greater than the internal energy of the system in its initial state (U_i), then ΔU would be **positive.**

$$\Delta U = U_f - U_i = +ve\ (U_f > U_i)$$

23. Explain how heat absorbed at constant volume is measured using bomb calorimeter with a neat diagram.

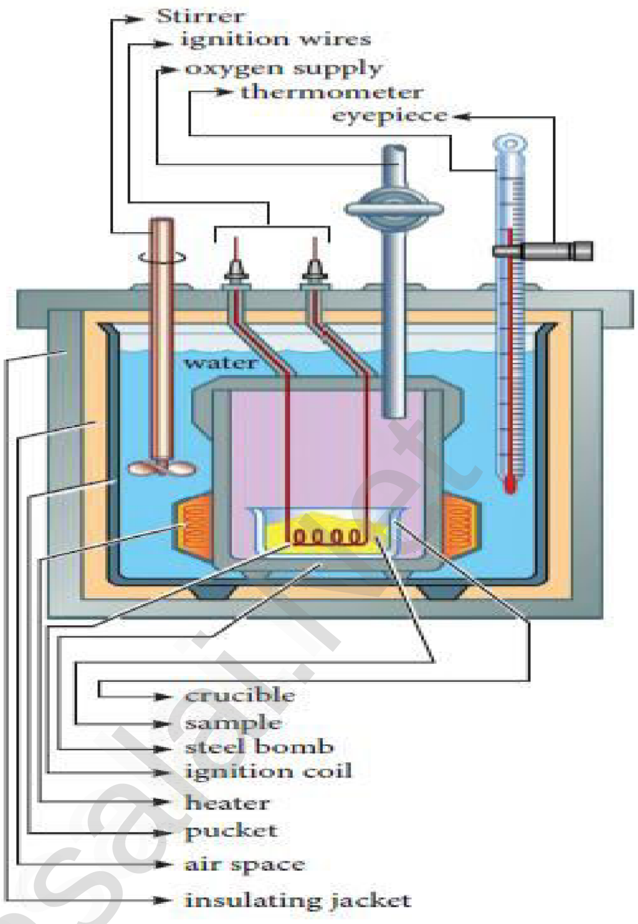


Figure 7.6 Bomb calorimeter

(i) Heat evolved at constant volume, is measured in a bomb calorimeter.

(ii) Apparatus setup: The inner vessel (the bomb) and its cover are made of strong steel. The cover is fitted tightly to the vessel by means of metal lid and screws.

(iii) Experiment: A weighed amount of the substance is taken in a platinum cup connected with electrical wires for striking an arc instantly to kindle combustion. The bomb is then tightly closed and pressurized with excess oxygen. The bomb is immersed in water, in the inner volume of the calorimeter. A stirrer is placed in the space between the wall of the calorimeter and the bomb, so that water can be stirred the substance through electrical heating.

(iv) **Calculation:** A known amount of combustible substance is burnt oxygen in the bomb. Heat evolved during the reaction is absorbed by the calorimeter as well as the water in which the bomb is immersed. The change in temperature is measured using a Backman thermometer. Since the bomb is sealed its volume does not change and hence the heat measurement is equal to the heat of combustion at a constant volume $(\Delta U)_c$.

The amount of heat produced in the reaction (ΔU)c is equal to the sum of the heat absorbed by the calorimeter and water.

Heat absorbed by the calorimeter $q_1 = k.\Delta T$

Where m_w is molar mass of water C_w is molar heat capacity of water (4,184 kJ K^{-1} mol^{-1})

Therefore $\Delta U_c = q_1 + q_2$

$$= k.\Delta T + m_w C_w \Delta T$$

$= (k + m_w C_w)\,\Delta T$

Calorimeter constant can be determined by burning a known mass of standard sample (benzoic acid) for which the heat of combustion is known (-3227 kJmol-1) The enthalpy of combustion at constant pressure of the substance is calculated from the equation

$$\Delta H^0_{C(pressure)} = \Delta U^0_{C(vol)} + \Delta n_g RT$$

49. Calculate the work involved in expansion and compression process.

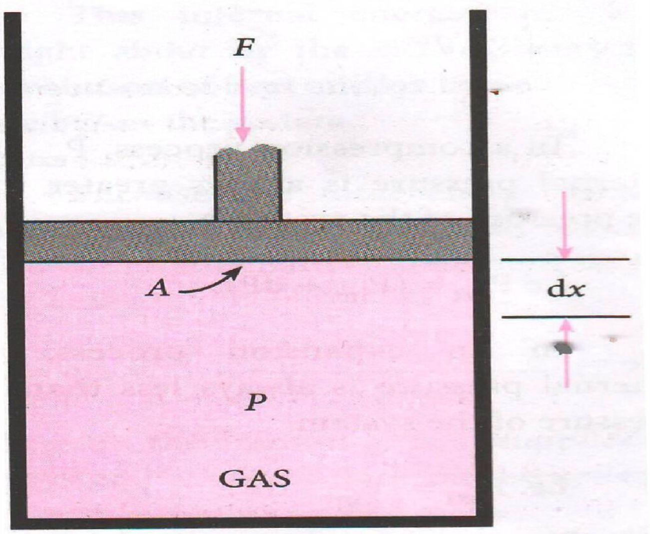


Figure 7.3 showing work involved in compression processes

i. For understanding pressure volume work, let us consider a cylinder which contains 'n' moles of an ideal gas fitted with a frictionless piston of cross sectional area A. The total volume of the gas inside is V*i* and pressure of the gas inside is P_{int}.

ii. If the external pressure P_{ext} is greater than P_{int}, the piston moves inward till the pressure inside becomes equal to P_{ext}. Let this change be achieved in a single step and the final volume be V_f.

iii. In this case, the work is done on the system (+w). It can be calculated as follows

$$w = -F\,.\Delta x \quad \text{------------- (1)}$$

iv. where dx is the distance moved by the piston during the compression and F is the force acting on the gas.

$$F = P_{ext}\,A \quad \text{---- (2)}$$

Substituting 2 in 1

$$w = -P_{ext}\,.\,A.\,\Delta x$$

$$A\ .\Delta x = \text{change in volume} = V_f - V_i$$

$$w = -P_{ext}\,.\,(V_f - V_i) \quad \text{——— (3)}$$

$$w = -P_{ext}\,.\,(-\Delta V) \quad \text{——— (4)}$$

$$= P_{ext}\,.\,\Delta V$$

v. Since work is done on the system, it is a positive quantity.

vi. If the pressure is not constant, but changes during the process such that it is always infinitesimally greater than the pressure of the gas, then, at each stage of compression, the volume decreases by an infinitesimal amount, dV. In such a case we can calculate the work done on the gas by the relation

$$w_{rev} = \int_{v_i}^{v_f} P_{int}\,dV$$

vii. In a compression process, P_{ext} the external pressure is always greater than the pressure of the system. **i.e** $P_{ext} = (P_{int} + dP)$.

viii. In an expansion process, the external pressure is always less than the pressure of the system i.e. $P_{ext} = (P_{int} - dP)$.

ix. When pressure is not constant and changes in infinitesimally small steps (reversible conditions) during compression from V_i to V_f , the P-V plot looks like in fig Work done on the gas is represented by the shaded area. In general case we can write,

x. $P_{ext} = (P_{int} + dP)$. Such processes are called reversible processes. For a compression process work can be related to internal pressure of the system under reversible conditions by writing equation

$$w_{rev} = \int_{v_i}^{v_f} P_{int}\, dV$$

For a given system with an ideal gas

$$P_{int} V = nRT$$

$$P_{int} = \frac{nRT}{V}$$

$$W_{rev} = \int_{v_i}^{v_f} \frac{nRT}{V} dV$$

$$W_{rev} = -\, nRT \int_{v_i}^{v_f} \left(\frac{dV}{V}\right)$$

$$W_{rev} = -\, nRT \ln \left(\frac{V_f}{V_i}\right)$$

$$W_{rev} = -\,2.303\ nRT \log \left(\frac{V_f}{V_i}\right)$$

xi. If $V_f > V_i$ (expansion), the sign of work done by the process is negative.

xii. If $V_f < V_i$ (compression) the sign of work done on the process is positive.

50. Derive the relation between ΔH and ΔU for an ideal gas. Explain each term involved in the equation.

When the system at constant pressure undergoes changes from an initial state with H_1, U_1 and V_1 to a final state with H_2, U_2 and V_2 the change in enthalpy ΔH, can be calculated as follows:

$$H = U + PV$$

In the initial state

$$H_1 = U_1 + PV_1 \qquad \text{------} \ (1)$$

In the final state

$$H_2 = U_2 + PV_2 \qquad \text{-----} \ (2)$$

change in enthalpy is (2) - (1)

$$(H_2 - H_1) = (U_2 - U_1) + P(V_2 - V_1)$$

$$\Delta H = \Delta U + P\Delta V \qquad \text{---------} \ (3)$$

As per first law of thermodynamics,

$$\Delta U = q + w$$

Equation (3) becomes

$$\Delta H = q + w + P\Delta V$$

$$W = -P\Delta V$$

$$\Delta H = q_p - P\Delta V + P\Delta V$$

$$\Delta H = q_p \qquad \text{———} \ (4)$$

q_p – is the heat absorbed at constant pressure and is considered as heat content.

Consider a closed system of gases which are chemically reacting to form gaseous products at constant temperature and pressure with V_i and V_f as the total volumes of the reactant and product gases respectively, and n_i and n_f as the number of moles of gaseous reactants and products, then,

For reactants (initial state) :

$$PV_i = n_i RT \qquad \text{---------} \ (5)$$

For products (final state) :

$$PV_f = n_f RT \qquad \text{---------} \ (6)$$

(6) - (5)

$$P(V_f - V_i) = (n_f - n_i) RT$$

$$P\Delta V = \Delta n_{(g)} RT \qquad \text{——} \ (7)$$

Substituting in (7) in (3)

$$\Delta H = \Delta U + \Delta n_{(g)} RT \qquad \text{——} \ (8)$$

51.Suggest and explain an indirect method to calculate lattice enthalpy of sodium chloride crystal.

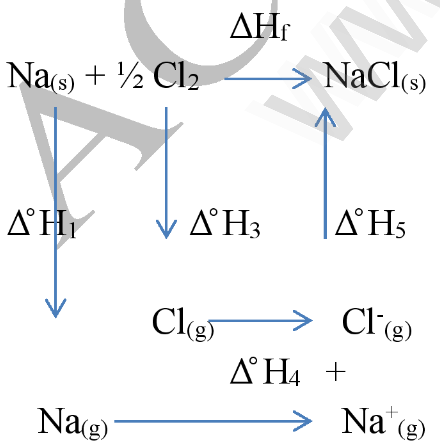
Let us use the Born - Haber cycle for determining the lattice enthalpy of NaCl as follows:

Since the reaction is carried out with reactants in elemental forms and products in their standard states, at 1 bar, the overall enthalpy change of the reaction is also the enthalpy of formation for NaCl.

Also, the formation of NaCl can be considered in 5 steps.

The sum of the enthalpy changes of these steps is equal to the enthalpy change for the overall reaction from which the lattice enthalpy of NaCl is calculated.

Let us calculate the lattice energy of sodium chloride using Born-Haber cycle



$\Delta^\circ H_2$

$\Delta^\circ H_f$ = heat of formation of sodium chloride = – 411.3 kJ mol^{-1}

$\Delta^\circ H_1$ = heat of sublimation of $Na_{(S)}$ = 108.7 kJ mol^{-1}

$\Delta^\circ H_2$ = ionisation energy of $Na_{(S)}$ = 495.0 kJ mol^{-1}

$\Delta^\circ H_3$ = dissociation energy of $Cl_{2(S)}$ = 244 kJ mol^{-1}

$\Delta^\circ H_4$ = Electron affinity of $Cl_{(S)}$ = – 349.0 kJ mol^{-1}

$\Delta^\circ H_f = \Delta^\circ H_1 + \Delta^\circ H_2 + ½\, \Delta^\circ H_3 + \Delta^\circ H_4 + \Delta^\circ H_5$

$\therefore \Delta^\circ H_5 = (\Delta^\circ H_f) - (\Delta^\circ H_1 + \Delta^\circ H_2 + \Delta^\circ H_3 + \Delta^\circ H_4)$

$\Rightarrow \Delta^\circ H_5 = (-411.3) - (108.7 + 495.0 + 122 - 349)$

$\Delta^\circ H_5 = (-411.3) - (376.7)$

$\therefore \Delta^\circ H_5 = -788$ kJ mol–1

This negative sign in lattice energy indicates that the energy is released when sodium is formed from its constituent gaseous ions Na^+ and Cl^-

52. List the characteristics of Gibbs free energy.

- Gibbs free energy is defined as the part of total energy of a system that can be converted (or) available for conversion into work.

  Free energy is defined as **G = H – TS.**

  Where, H-enthalpy T –Temperature and S=entropy
- Unit of Gibbs free energy **$Jmol^{-1}$**
- 'G' is a **state function**
- G is an **Extensive property.**
- ΔG becomes intensive property for a closed system. When mass remains constant between initial and final states of system
- 'G' has a single value for the thermodynamic state of the system
- G and ΔG values correspond to the system only.

Process	Spontaneous ΔG<0	Equilibrium ΔG=0	Non-spontaneous ΔG>0
ΔG	negative	Zero	positive

(v) Gibbs free energy and the network done by the system:

For any system at constant pressure and temperature

$$\Delta G = \Delta H - T\Delta S \text{ ----- } (1)$$

We know that,

$$\Delta H = \Delta U + P\Delta V$$

$$\therefore \Delta G = \Delta U + P\Delta V - T\Delta S$$

from first law of thermodynamics

$$\Delta U = q + w$$

from second law of thermodynamics

$$\Delta S = \frac{q}{T} \qquad \Delta G = q + w + P\Delta V - T\left(\frac{q}{T}\right)$$

$$\Delta G = w + P\,\Delta V$$

$$-\Delta G = -w - P\Delta V \quad \text{-------} \ (2)$$

But –PΔV represents the work done due to expression against a constant external pressure. Therefore, it is clear that the decrease in free energy (-ΔG) accompanying a process taking place at constant temperature and pressure is equal to the maximum work obtained from the system other than the work of expansion.

53. Calculate the work done when 2 moles of an ideal gas expands reversibly and isothermally from a volume of 500 ml to a volume of 2 L at 25°C and normal pressure.

Given :

$n = 2$ moles; $V_i = 500$ ml $= 0.5$ lit ; $V_f = 2$ lit

$T = 25\ \ C = 298$ K

$$w = -2.303\ nRT\log\left(\frac{V_f}{V_i}\right)$$

$$w = -2.303 \times 2 \times 8.314 \times 298 \times \log\left(\frac{2}{0.5}\right)$$

$$w = -2.303 \times 2 \times 8.314 \times 298 \times \log(4)$$

$$w = -2.303 \times 2 \times 8.314 \times 298 \times 0.6021$$

$$w = -\,6871\ J$$

$$w = -6.871\ Kj$$

55. Calculate the entropy change in the system, and surroundings, and the total entropy change in the universe during a process in which 245 J of heat flow out of the system at 77°C to the surrounding at 33°C.

Given :

$T_{sys} = 77\ \ C = (77 + 273) = 350$ K

$T_{surr} = 33\ \ C = (33 + 273) = 306$ K

$q = 245$ J

$\Delta S_{sys} = q$

$$T_{sys} = \frac{q}{T_{sys}} = \frac{-245}{350} = -\,0.7\ JK^{-1}$$

$$\Delta S_{surr} = \frac{q}{T_{sys}} = \frac{+245}{306} = +\,0.8\ JK^{-1}$$

$$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr}$$

$$\Delta S_{univ} = -\,0.7\ JK^{-1} + 0.8\ JK^{-1}$$

$$\mathbf{\Delta S_{univ} = 0.1\ JK^{-1}}$$

56. 1 mole of an ideal gas, maintained at 4.1 atm and at a certain temperature, absorbs heat 3710J and expands to 2 litres. Calculate the entropy change in expansion process.

Given :

n = 1mole; P = 4.1 atm; V= 2 Lit; T = ?; q=3710 J

$\Delta S = \frac{q}{T}$; $\Delta S = \frac{q}{\left(\frac{PV}{nR}\right)}$; $\Delta S = \frac{nRq}{PV}$

$$\Delta S = \frac{1 \times 0.082\ lit\ atm\ K^{-1} \times 3710 J}{4.1\ atm \times 2\ lit}$$

$\mathbf{\Delta S = 37.10\ JK^{-1}}$

57. 30.4 kJ is required to melt one mole of sodium chloride. The entropy change during melting is 28.4 JK^{-1} mol^{-1}. Calculate the melting point of sodium chloride.

ΔH_f (NaCl) = 30.4 kJ = 30400 J mol^{-1}

ΔS_f(NaCl) = 28.4 JK^{-1} mol^{-1}; T_f = ?

$\Delta S_f = \frac{\Delta H_f}{T_f}$; $T_f = \frac{\Delta H_f}{\Delta S_f}$

$T_f = \frac{30400\ J\ mol^{-1}}{28.4\ J\ K^{-1}\ mol^{-1}}$ **T_f = 1070.4 K**

33. Calculate the standard heat of formation of propane, if its heat of combustion is −2220.2 kJ mol^{-1}. the heats of formation of $CO_{2(g)}$ and H_2O(l) are −393.5 and −285.8 kJ mol^{-1} respectively.

SOLUTION :

Given

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

ΔH_c = -2220.2 kJmol^{-1} ———(1)

$C + O_2 \rightarrow CO_2$

ΔH_f = −393.5 kJ mol^{-1} ———(2)

$H_2 + ½ O_2 \rightarrow H_2O$

ΔH_f = −285.8 kJ mol^{-1} ——— (3)

$3C + 4H_2 \rightarrow C_3H_8$

ΔH_c = ?

(2) × 3 ⇒ $3C + 3O_2 \rightarrow 3CO_2$

ΔH_f = −1180.5 kJ ———(4)

(3)× 4 ⇒ $4 H_2 + 2O_2 \rightarrow 4H_2O$

ΔH_f = −1143.2 kJ ———(5)

(4) + (5) − (1) ⇒ $3C + 3O_2 + 4H_2 +$

$2O_2 + 3CO_2 + 4H_2O \rightarrow 3CO_2 + 4H_2O + C_3H_8 + 5O_2$

ΔH_f = −1180.5 − 1143.2 − (−2220.2) kJ

3C + 4 H_2 → C3H8

ΔH_f = −103.5 kJ

Standard heat of formation of propane is **ΔH_f (C_3H_8) = −103.5 kJ**

34. You are given normal boiling points and standard enthalpies of vapourisation. Calculate the entropy of vapourisation of liquids listed below.

S.No	**Liquid**	**Boiling points (C)**	**ΔH (KJ mol^{-1})**
1.	Ethanol	78.4	+ 42.4
2.	Toluene	110.6	+ 35.2

For ethanol :

Given : $T_b = 78.4^{\circ}C = (78.4 + 273)$

$= 351.4$ K

ΔH_V (ethanol) = + 42.4 kJ mol^{-1}

$$\Delta S_v = \frac{\Delta H_V}{T_b} \quad ; \quad \Delta S_v = \frac{+\,42.4\ KJ\ mol^{-1}}{351.4\ K}$$

$$\Delta S_v = \frac{+42400\ J\ mol^{-1}}{351.4\ K};$$

$$\mathbf{\Delta S_v = +\,120.66\ JK^{-1}\ mol^{-1}}$$

For Toluene :

Given :

$T_b = 110.6\ \ C = (110.6+273)$

$= 383.6$ K

ΔH_V (toluene) = + 35.2 kJ mol^{-1}

$$\Delta S_v = \frac{\Delta H_V}{T_b}$$

$$\Delta S_v = \frac{+35.2\ \text{KJ}\ mol^{-1}}{383.6K}; \quad \Delta S_v = \frac{+\,35200\ \text{J}\ mol^{-1}}{383.6\ K}$$

$$\mathbf{\Delta S_v = +\,91.76\ JK^{-1}}$$

35. For the reaction $Ag_2O(s) \rightarrow 2Ag(s) + 1/2\ O_2(g)$: ΔH = 30.56 kJ mol^{-1} and ΔS=6.66JK^{-1}mol^{-1}(at1 atm). Calculate the temperature at which ΔG is equal to zero. Also predict the direction of the reaction (i) at this temperature and (ii) below this temperature.

Solution :

Given :

$\Delta H = 30.56$ kJ mol^{-1}

$= 30560$ J mol^{-1}

$\Delta S = 6.66 \times 10^{-3}$ kJK^{-1}mol^{-1}

T = ? at which $\Delta G=0$

$\mathbf{\Delta G = \Delta H - T\Delta S}$

$0 = \Delta H - T\Delta S$

$$T = \frac{\Delta H}{\Delta S} \quad ; \quad T = \frac{30.56\ kJ\ mol^{-1}}{6.66 \times 10^{-3} kJ\ K^{-1} mol^{-1}}$$

T = 4589K

(i) At 4589K ; $\Delta G = 0$ the reaction is in equilibrium.

(ii) at temperature below 4598 K , $\Delta H > T\Delta S$ $\Delta G = \Delta H - T\Delta S > 0$, the reaction in the forward direction, is nonspontaneous. In other words the reaction occurs in the backwarddirection.

37. Cyanamide (NH_2CN) is completely burnt in excess oxygen in a bomb calorimeter, ΔU was found to be −742.4 kJ mol^{-1}, calculate the enthalpy change of the reaction at 298K. $NH_2CN(s) + 3/2O_2(g) \rightarrow N_2(g) + CO_2(g) + H_2O(l)$ ΔH= ?

Solution :

Given

$T = 298K$; $\Delta U = -742.4$ kJ mol^{-1}

$\Delta H = ?$

$\Delta H = \Delta U + \Delta n_g RT$

$\Delta H = \Delta U + (n_p - n_r) RT$

$\Delta H = -742.4 + \left(2 - \frac{3}{2}\right) \times 8.314 \times 10^{-3} \times 298$

$= -742.4 + (0.5 \times 8.314 \times 10^{-3} \times 298)$

$= -742.4 + 1.24$

$\Delta H = -741.16$ kJ mol^{-1x}

38. Calculate the enthalpy of hydrogenation of ethylene from the following data. Bond energies of C− H, C − C , C = C and H − H are 414, 347, 618 and 435 kJ mol^{-1}.

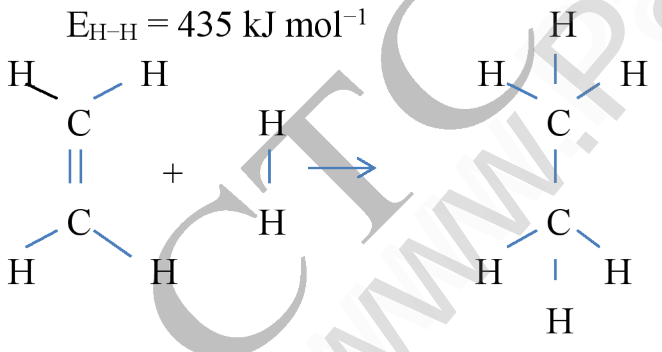
Given :

$E_{C-H} = 414$ kJ mol^{-1}

$E_{C-C} = 347$ kJ mol^{-1}

$E_{C=C} = 618$ kJ mol^{-1}

$E_{H-H} = 435$ kJ mol^{-1}



Ethylene Ethane

$\Delta H_r = \Sigma\,(\text{Bond energy})_r - \Sigma\,(\text{Bond energy})p$

$\Delta H_r = (E_{C=C} + 4E_{C-H} + E_{H-H}) - (E_{C-C} + 6E_{C-H})$

$\Delta H_r = (618 + (4 \times 414) + 435) - (347 + (6 \times 414)\,)$

$\Delta H_r = 2709 - 2831$; $\Delta H_r = -122$ kJ mol^{-1}

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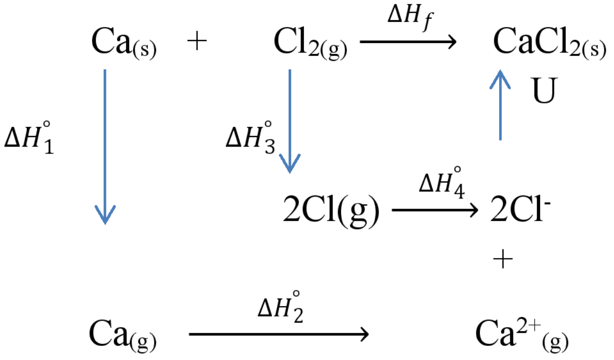
39. Calculate the lattice energy of $CaCl_2$ from the given data $Ca(s)+Cl_2(g) \rightarrow CaCl_2(s)$; $\Delta H_f^0 = - 795$ kJ mol^{-1}

Atomisation : $Ca(s) \rightarrow Ca(g)$; $\Delta H_1^0 = + 121$ kJ mol^{-1}

Ionisation : $Ca(g) \rightarrow Ca_2+(g) + 2e^-$; $\Delta H_2^0 = + 2422$ kJ mol^{-1}

Dissociation : $Cl_2(g) \rightarrow 2Cl(g)$; $\Delta H_3^0 = + 242.8$ kJ mol^{-1}

Electron affinity : $Cl(g) + e^- \rightarrow Cl^-(g)$; $\Delta H_4^0 = -355$ kJ mol^{-1}



$\Delta H_f = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + U$

$-795 = 2422 + 242.8 + (2 \times -255) + U$

$-795 = 2785.8 - 710 + U$

$-795 = 2075.8 + U$

$U = -795 - 2075.8$

$\mathbf{U = -2870.8\ KJmol^{-1}}$

40. Calculate the enthalpy change for the reaction $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$ from the following data.

$2Fe + 3/2\ O_2 \rightarrow Fe_2O_3$; $\Delta H = -741$ kJC $+1/2\ O_2 \rightarrow CO$; $\Delta H = -137$ kJ $C + O_2 \rightarrow CO_2$; $\Delta H = - 394.5$ kJ

Given:

$\Delta H_f (Fe_2O_3) = -741\ kJmol^{-1}$

$\Delta H_f (CO) = -137\ kJmol^{-1}$

$\Delta H_f (CO_2) = -394.5\ kJmol^{-1}$

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

$\Delta H_f = ?$

$\Delta H_r = \Sigma (\Delta H_f)products - \Sigma (\Delta H_f)reactants$

$\Delta H_r = [2\Delta H_f(Fe) + 3\Delta H_f(CO_2)] - [\Delta H_f (Fe_2O_3) + 3\ \Delta H_f (CO)]$

$\Delta H_r = [0 + 3 (-394.5)] - [-741 + 3 (-137)]$

$\Delta H_r = [-1183.5] - [-1152]$

$\Delta H_r = -1183.5 + 1152$

ΔH_r = -31.5kJmol^{-1}

44. For the reaction at 298 K: 2A +B → C ΔH=400 J mol^{-1} ; ΔS = 0.2 JK^{-1} mol^{-1} Determine the temperature at which the reaction would be spontaneous.

Solution:

Given:

T = 298 K

$\Delta H = 400$ J mol^{-1}

$\Delta S = 0.2$ JK^{-1}mol^{-1}

$\Delta G = \Delta H - T\Delta S$

If T = 2000K

$\Delta G = 400 - (0.2 \times 2000) = 0$

If T > 2000 K

ΔG will be negative

The reaction would be spontaneous only beyond 2000 K.

46. A gas mixture of 3.67 lit of ethylene and methane on complete combustion at 25°C and at 1 atm pressure produce 6.11 lit of carbon dioxide. Find out the amount of heat evolved in kJ, during this combustion. (ΔHC(CH_4)= − 890 kJ mol^{-1} and (ΔHC(C_2H_4)= −1423 kJ mol^{-1}

Solution:

Given:

ΔH_c (CH_4)= -890 KJmol^{-1}

ΔH_c (C_2H_4) = -1423 KJmol^{-1}

Let the mixture contain x lit of CH4 and (3.67 – x) lit of ethylene.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

x lit x lit

$C_2H_2 + 3O_2 \rightarrow 2CO_2 + 2H_2O$

(3.67 – x) lit 2 (3.67 – x)

Volume of carbon dioxide formed = x + 2(3.67 – x) = 6.11 lit

= x + 7.34 – 2x = 6.11

= 7.34 – x = 6.11

X = 1.23 lit

Given mixture contains 1.23 lit of Methane and 2.44 lit of ethylene, hence

$$\Delta H_c = \left[\frac{\Delta H_c\,(CH_4)}{22.4\ lit} \times (x)lit\right] + \left[\frac{\Delta H_c\,(C_2H_4)}{22.4\ lit} \times (3.67 - x)lit\right]$$

$$\Delta H_c = \left[\frac{-890\ KJmol^{-1}}{22.4\ lit} \times 1.23\ lit\right] + \left[\frac{-1423}{22.4\ lit} \times (3.67 - 1.23)lit\right]$$

ΔH_c =[-48.87kJmol^{-1}]+[-155kJmol^{-1}]

ΔH_c= -203.87 kJmol^{-1}

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Lesson 7 THERMODYNAMICS

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
9	6	11	8	10	10	11	8	9	11				

1. Write note on system. (187)
2. Define Surrounding & Boundary. (188)
3. What are the types of system. (188)
4. What is isolated system? Give example. (188)
5. Explain closed system with an example. (188)
6. What is open system? Give example. (188)
7. Explain intensive properties with two examples. (189) **J23 2M**
8. Explain extensive properties with two examples. (189)
9. Distinguish between extensive and intensive property? (189) **(S20 3M)**
10. What is reversible process? Give an ex (189)
11. What is an irreversible process? Give an example. (189)
12. Define adiabatic process? (190)
13. Define isothermal process. (190)
14. Define isobaric process. (190)
15. Define isochoric process. (190)
16. Define cyclic process. (190)
17. What is state function? Give two examples. (190) **(MQ19)** **M23 3Mi**
18. What is path function? Give two examples. (190) **J22 2M** **M23 3Mii**
19. What is meant by internal energy? (191)
20. List the characteristics of internal energy. (191) **M22 5M, M24 5M**
21. Calculate the work involved in expansion and compression process. (193)
22. Explain sign convention of heat and work. (194) **(M22 2M)**
23. Define Zeroth law of thermodynamics (or) Law of thermal equilibrium. (195) **(S20 2M)**
24. State first law of thermodynamics. (195) **S21 5Mi**
25. Derive the various mathematical statements of the first law.(196)
26. Explain the relation between enthalpy (H) and internal energy (U). (197) (Derive the relation between enthalpy ΔH and internal energy ΔU for an ideal gas. Explain each term involved in the equation. BBQ50 **(S20 5Mi, J23 5M**
27. Define standard heat of formation.(198)
28. Explain thermochemical Equations. (198)
29. Define heat of combustion. (200)
30. Calculate ΔH_f^0 for the reaction $CO_2(g) + H_2(g) \rightarrow CO(g) + H_2O(g)$, given that ΔH_f^0 for $CO_2(g), CO(g) and H_2O(g) are -393.5\ , -111.31\ and -242 kJmol^{-1}$ respectively. (200)**(MQ19)**
31. Define molar heat capacity. Give its Unit.(201) **(J19) 3mark**
32. Relationship between Cp and Cv (201)
33. Explain Bomb calorimeter. (203)
34. Explain coffee cup calorimeter. (204)
35. What are the applications of the heat of combustion. (205)

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36. Define heat of solution. Give an example. (206)
37. Define heat of neutralization. Give an example. (206)
38. Define molar heat of fusion. Give an example (206)
39. Define heat of vaporization. Give an example (207)
40. Define heat of sublimation. Give an example. (207)
41. Define heat of transition. Give an example. (207)
42. State Hess's law of constant heat summation. (207) **S21 2M, M24 3M**
43. Define lattice energy. (208)
44. Explain Born-Haber cycle. (208)
45. Explain various statement of second law thermodynamics. (210) **M23 5M**
46. Define entropy. (210) **J22 5Mi**
47. If an automobile engine burns petrol at a temperature of 816°C (1089K) and if the surrounding temperature is 21°C (294K). Calculate its maximum possible efficiency. (211) **J22 2M COM**
48. An engine operating between 127°C and 47°C takes some specified amount of heat from a high temperature reservoir. Assuming that there are no fractional losses, calculate the percentage efficiency of the engine. **(211) J24 5Mii**
49. Write note on standard entropy change (212)
50. Define standard entropy of formation. (212)
51. Write note on Entropy change accompanying change of phase. (212)
52. $C_{(s)} + O_{2(s)} \rightarrow CO_{2(g)}$ Calculate the standard entropy change for the above reaction, given the standard entropies of $CO_{2(g)}$, $C_{(s)}$, $O_{2(g)}$ are 213.6, 5.740 and 205 JK^{-1} respectively. (212) **(M19 5Mi)**
53. Calculate the entropy change during the melting of one mole of ice into water at 0°C. Enthalpy of fusion of ice is 6008 J mol^{-1}. (213) **(M19 5Mi, S20 5Mii) M23 2M compulsory**
54. Define Gibbs free energy & character (214) **S21 2M, J22 5Mii, J24 2M**
55. What are the Conditions (Criteria) for spontaneity of a process. (215) **M22 2M, S21 5Mii**
56. State the third law of Thermodynamics. (218) **(M19 2M, J24 5Mi**

& study all book back question & answer

Unit 8: Physical and Chemical Equilibrium

26. If there is no change in concentration, why is the equilibrium state considered dynamic?

Rate of forward reaction=Rate of backward reaction

Chemical reactions which are reversible do not cease, when equilibrium is attained. At equilibrium the forward and the backward reactions are proceeding at the same rate and no macroscopic change is observed. So chemical equilibrium is in a state of dynamic equilibrium.

27. For a given reaction at a particular temperature, the equilibrium constant has constant value. Is the value of Q also constant? Explain.

- No.
- The value of Q is determined by ***the concentration of product and reactants that are not necessarily equilibrium concentration.***
- Thus, its value is not a constant. At equilibrium $K_C = Q$
- In the chemical reaction, as the reaction proceeds, there is a continuous change in the concentration of reactants and products and also the Q value until the reaction reaches the equilibrium.
- So even at particular temperature, Q is not constant.

28. What the relation between K_P and K_C. Give one example for which K_P is equal to K_C.

Relation between K_P and K_C

- **$K_p = K_c (RT)^{\Delta ng}$**

Where, *Kp= equilibrium constant is terms of partial pressure*

Kc= equilibrium constant is terms of concentration

R= gas constant, T= Temperature

Example: $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

$\Delta n_g = 2\text{-}2=0$

$K_p = K_c(RT)^0 = K_c$

$K_p=K_c$

29. For a gaseous homogeneous reaction at equilibrium, number of moles of products are greater than the number of moles of reactants. Is K_C is larger or smaller than K_P.

For a gaseous homogeneous reaction at equilibrium, number of moles of products are greater than the number of moles of reactants, then $\Delta n_g = +ve$

- **$K_p = K_c (RT)^{\Delta ng}$**

When $\Delta n_g = +ve$

$K_P = K_C (RT)^1$; **$K_P > K_C$**

Example: $PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

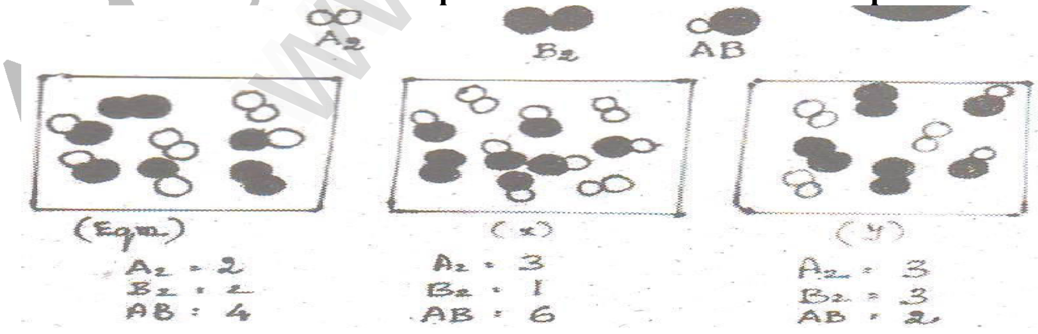
Δn_g = No. of moles of product – No. of moles of reactant

$\Delta n_g = 2 - 1 = 1$

30. When the numerical value of the reaction quotient (Q) is greater than the equilibrium constant (K), in which direction does the reaction proceed to reach equilibrium?

If $Q > K_c$, the reaction will proceed in the **reverse direction** i.e., formation of reactants.

31. For the reaction, $A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$; ΔH is –ve. the following molecular scenes represent different reaction mixture (A – green, B – blue) i) Calculate the equilibrium constant K_P and (K_C). ii)For the reaction mixture represented by scene (x), (y) the reaction proceed in which directions ? iii)What is the effect of increase in pressure for the mixture at equilibrium.



$K_C = \frac{[AB]^2}{[A_2][B_2]}$ A – green ; B – blue

Given that ‘V’ is constant (closed system) At equilibrium,

$$K_C = \frac{\left[\frac{4}{V}\right]^2}{\left(\frac{2}{V}\right)\left(\frac{2}{V}\right)} = \frac{16}{4} = 4$$

$$K_P = K_C\,(RT)^{\Delta ng}\ ;\ K_P = 4(RT)^0 = 4$$

ii) At stage ‘x’

$$Q = \frac{\left(\frac{6}{V}\right)^2}{\left(\frac{2}{V}\right)\left(\frac{1}{V}\right)} = \frac{36}{2} = 18$$; $Q > K_C$ (ie.), reverse reaction is favoured.

At Stage ‘y’

$$Q = \frac{\left(\frac{3}{V}\right)^2}{\left(\frac{3}{V}\right)\left(\frac{3}{V}\right)} = \frac{9}{3 \times 3} = 1$$; $K_C > Q$ (ie.), forward reaction is favoured.

iii) since $\Delta ng = 2-2 = 0$, thus, pressure has no effect. So by increasing the pressure, equilibrium will not be affected.

32. State Le-Chatelier principle.

It states that "if a system at equilibrium is disturbed, then the system shifts itself in a direction that nullifies the effect of that disturbance."

33. Consider the following reactions,

a) $H_{2(g)} + I_{2(g)} \rightleftharpoons 2\,HI_{(g)}$ b) $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$ c) $S_{(s)} + 3F_{2(g)} \rightleftharpoons SF_{6(g)}$

In each of the above reaction find out whether you have to increase (or) decrease the volume to increase the yield of the product.

a) $H_{2(g)} + I_{2(g)} \rightleftharpoons 2\,HI_{(g)}$

1V 1V 2V

- ***Ans: volume change has no effect.***
- ***$\Delta n_g = 2-2=0$***

- According to Le Chatelier's principle increase in pressure will shift the equilibrium a direction that has lesser number of moles.
- In the above equilibrium, ***pressure and volume has no effect, since number of product is equal to number of moles of reactant.***

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

0 0 1V

Ans: Increase in volume (decreasing the pressure) will yield more products.

- In this equilibrium number of moles of product is greater than number of moles of reactant.
- So increase of pressure will shift the equilibrium towards the backward reaction.
- To increase the yield of product, pressure should be decrease and hence volume increases.

c) $S_{(s)} + 3F_{2(g)} \rightleftharpoons SF_{6(g)}$

0 3V 1V

Ans: Decrease in volume (high pressure) will yield more products.

In the above equilibrium increase in pressure favours formation of product hence volume should be decreased.

34. State law of mass action.

The law states that, "*At any instant, the rate of a chemical reaction at a given temperature is directly proportional to the product of the active masses of the reactants at that instant*".

$\text{Rate}\ \alpha\ [\ \text{Reactant}]^x$

Where, x is the stoichiometric coefficient of the reactant.

35. Explain how will you predict the direction of an equilibrium reaction.

Consider a general homogeneous reversible reaction,

$xA + yB \rightleftharpoons lC + mD$

For the above reaction under non equilibrium conditions, **reaction quotient 'Q'** is defined *as the ratio of the product of active masses of reaction products raised to the respective stoichiometric coefficients in the balanced chemical equation to that of the reactants.*

By knowing the Q value, we can predict the direction of the reaction by comparing it with Kc.

- If Q = Kc, the reaction is in **equilibrium** state.
- If Q > Kc, the reaction will proceed in the **reverse direction** i.e., formation of reactants.
- If Q < Kc, the reaction will proceed in the **forward direction** i.e., formation of products.

36. Derive a general expression for the equilibrium constant K_P and K_C for the reaction.

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

Let us consider the formation of ammonia in which, 'a' moles nitrogen and 'b' moles hydrogen gas are allowed to react in a container of volume V. Let 'x' moles of nitrogen react with 3x moles of hydrogen to give 2x moles of ammonia.

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

	N_2	H_2	NH_3
Initial number of moles	**a**	**b**	**0**
number of moles reacted	x	3x	0
Number of moles at equilibrium	a-x	b-3x	2x
Active mass or molar concentration at equilibrium(n/V)	$\frac{a-x}{V}$	$\frac{b-3x}{V}$	$\frac{2x}{V}$

Applying law of mass action,

$$K_C = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K_C = \frac{\left(\frac{2x}{V}\right)^2}{\left(\frac{a-x}{V}\right)\left(\frac{b-3x}{V}\right)^3} = \frac{\left(\frac{4x^2}{V}\right)}{\left(\frac{a-x}{V}\right)\left(\frac{b-3x}{V}\right)^3} = \frac{4x^2V^2}{(a-x)(b-3x)^3}$$

The equilibrium constant K_P can also be calculated as follows:

$$\mathbf{K_P = K_C\ (RT)^{(\Delta ng)}}$$

$$\Delta n_g = n_p - n_r = 2 - 4 = -2$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^2}(RT)^{-2}$$

Total number of moles at equilibrium,

$$n = a-x + b-3x + 2x = \mathbf{a+b-2x}$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^3}\left(\frac{PV}{n}\right)^{-2}$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^3}\left(\frac{n}{PV}\right)^{2}$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^3}\left(\frac{a+b-2x}{PV}\right)^{2}$$

$$K_P = \frac{4x^2V(a+b-2x)^2}{P^2(a-x)(b-3x)^3}$$

37. Write a balanced chemical equation for a equilibrium reaction for which the equilibrium constant is given by expression $K_C = \frac{[NH_3]^4[O_2]^5}{[NO]^4[H_2O]^6}$

Chemical equation is,

$$4NO + 6H_2O \rightleftharpoons 4NH_3 + 5O_2$$

38. What is the effect of added inert gas on the reaction at equilibrium.

- When an **inert gas** (i.e, a gas which does not react with any other species involved in equilibrium) is added to an equilibrium system at constant volume, the total number of moles of gases present in the container increases, that is, the total pressure of gases increases.
- The partial pressure of the reactants and the products or the molar concentration of the substance involved in the reaction remains unchanged.
- Hence at constant volume, addition of **inert gas has no effect on equilibrium.**

39. Derive the relation between K_P and K_C.

Let us consider the general reaction in which all reactants and products are ideal gases.

$$xA + yB \rightleftharpoons lC + mD$$

The equilibrium constant, K_C is, $\mathbf{K_C} = \frac{[C]^l[D]^m}{[A]^x[B]^y}$ ---- (1)

And K_P is, $K_P = \frac{P_C^l P_D^m}{P_A^x P_B^y}$ ---- (2)

The ideal gas equation is

$PV = nRT$ or $P = \frac{n}{V} RT$

Since Active mass = molar concentration = n/V

P = active mass × RT

Based on the above expression the partial pressure of the reactants and products can be expressed as,

$P_A^x = [A]^x (RT)^x$

$P_B^y = [B]^y (RT)^y$

$P_C^l = [C]^l (RT)^l$

$P_D^m = [D]^m (RT)^m$

On substitution in Eqn. 2,

$k_P = \frac{[C]^l[RT]^l[D]^m[RT]^m}{[A]^x[RT]^x[B]^y[RT]^y}$ ---- (3)

$k_P = \frac{[C]^l[D]^m[RT]^{l+m}}{[A]^x[B]^y[RT]^{x+y}}$

$k_P = \frac{[C]^l[D]^m}{[A]^x[B]^y} RT^{(l+m)-(x+y)}$ ---- (4)

By comparing equation (1) and (4), we get

$k_P = K_C (RT)^{(\Delta ng)}$ ---- (5)

where, Δn_g is the difference between the sum of number of moles of products and the sum of number of moles of reactants in the gas phase.

40. One mole of PCl_5 is heated in one litre closed container. If 0.6 mole of chlorine is found at equilibrium, calculate the value of equilibrium constant.

Given that $[PCl_5]_{initial} = \frac{1\ mole}{1\ dm^3}$

$[Cl_2]_{eq} = 0.6$ mole dm^{-3}

$PCl_5 \rightleftharpoons PCl_3 + Cl_2$

$[PCl_3]_{eq} = 0.6$ mole dm^{-3}

$[PCl_5]_{eq} = 0.4$ mole dm^{-3}

$K_C = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{0.6 \times 0.6}{0.4} = \mathbf{0.9}$

41. For the reaction $SrCO_3 (s) \rightleftharpoons SrO (s) + CO_2(g)$, the value of equilibrium constant $K_P = 2.2 \times 10^{-4}$ at 1002 K. Calculate K_C for the reaction.

For the reaction, $SrCO_3 (s) \rightleftharpoons SrO (s) + CO_2(g)$

$\Delta n_g = 1 - 0 = 1$

$\therefore K_P = K_C (RT)$

$2.2 \times 10^{-4} = K_C (0.0821) (1002)$

$K_C = 2.674 \times 10^{-6}$

42. To study the decomposition of hydrogen iodide, a student fills an evacuated 3 litre flask with 0.3 mol of HI gas and allows the reaction to proceed at 500 0C. At equilibrium he found the concentration of HI which is equal to 0.05 M. Calculate K_C and K_P.

$V = 3L$; $[HI]_{\text{initial}} = \frac{0.3\,mol}{3L} = 0.1M$

$[HI]_{eq} = 0.05M$

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

	$HI(g)$	$H_2(g)$	$I_2(g)$
Initial Concentration	0.1	-	-
Reacted	0.05	-	-
Equilibrium Concentration	0.05	0.025	0.025

$$K_C = \frac{[H_2][I_2]}{[HI]^2} = K_C = \frac{0.025 \times 0.025}{0.05 \times 0.05} = 0.25$$

$K_P = K_C\,(RT)^{\Delta ng}$

$\Delta n_g = 2\text{-}2 = 0$

$K_P = 0.25\,(RT)^0 = \mathbf{0.25}$

43. 1 mol of CH_4, 1 mole of CS_2 and 2 mol of H_2S are 2 mol of H_2 are mixed in a 500 ml flask. The equilibrium constant for the reaction $K_C = 4 \times 10^{-2}$ mol^2 lit^{-2}. In which direction will the reaction proceed to reach equilibrium ?

$CH_4(g) + 2H_2S(g) \rightleftharpoons CS_2(g) + 4H_2(g)$

$K_C = 4\times 10^{-2}$ mol lit^{-2}

Volume = 500ml = ½ L

$[CH_4]_{in} = \frac{1mol}{\frac{1}{2}L}$ $[CS_2]_{in} = \frac{1mol}{\frac{1}{2}L}$

$= 2$ mol L^{-1} $= 2$ mol L^{-1}

$[H_2S]_{in} = \frac{2mol}{\frac{1}{2}L}$ $[H_2] = \frac{2mol}{\frac{1}{2}L}$

$= 4$ mol L^{-1} $= 4$ mol L^{-1}

$$Q = \frac{[CS_2][H_2]^4}{[CH_4][H_2S]^2} = \frac{2 \times (4)^4}{(2)(2)^2} = 64; \qquad \mathbf{Q > K_C}$$

44. At particular temperature $K_C = 4 \times 10^{-2}$ for the reaction $H_2S(g) \rightleftharpoons H_2(g) + ½\, S_2(g)$ Calculate KC for each of the following reaction i) $2H_2S\,(g) \rightleftharpoons 2H_2\,(g) + S_2\,(g)$

ii) $3H_2S\,(g) \rightleftharpoons 3H_2\,(g) + 3/2\, S_2(g)$

$K_C = 4 \times 10^{-2}$ for the reaction,

$H_2S(g) \rightleftharpoons H_2(g) + ½\, S_2(g)$

$$K_C = \frac{[H_2][S_2]^{\frac{1}{2}}}{[H_2S]}$$

$$4 \times 10^{-2} = \frac{[H_2][S_2]^{\frac{1}{2}}}{[H_2S]}$$

For the reaction, **i) $2H_2S\,(g) \rightleftharpoons 2H_2\,(g) + S_2\,(g)$**

$$K_C = \frac{[H_2]^2[S_2]}{[H_2S]^2} = (4 \times 10^{-2})^2 = 16 \times 10^{-4}$$

For the reaction, **ii) $3H_2S\,(g) \rightleftharpoons 3H_2\,(g) + 3/2\, S_2(g)$**

$$K_C = \frac{[H_2]^3[S_2]^{\frac{3}{2}}}{[H_2S]^3} = (4 \times 10^{-2})^2 = 64 \times 10^{-6}$$

45. 28 g of Nitrogen and 6 g of hydrogen were mixed in a 1 litre closed container. At equilibrium 17 g NH_3 was produced. Calculate the weight of nitrogen, hydrogen at equilibrium.

Given $m_{N_2} = 28g;\ m_{H_2} = 6g;\qquad V = 1L$

$$(n_{N_2})_{initial} = \frac{28}{28} = 1 \text{ mol}$$

$$(n_{N_2})_{initial} = \frac{6}{2} = 3 \text{ mol}$$

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

	$N_2(g)$	$H_2(g)$	$NH_3(g)$
Initial Concentration	1	3	-
Reacted	0.5	1.5	-
Equilibrium Concentration	0.5	1.5	1

$$[NH_3] = \left(\frac{17}{17}\right) = 1 \text{ mol}$$

Weight of N_2 = (no.of moles of N_2) × molar mass of N_2

= 0.5 × 28 = **14g**

Weight of H_2 = (no.of moles of H_2) × molar mass of H_2

= 1.5 × 2 = **3g**

46. The equilibrium for the dissociation of XY_2 is given as, $2XY_2 (g) \rightleftharpoons 2XY (g) + Y_2(g)$ if the degree of dissociation x is so small compared to one. Show that $2 K_P = PX^3$ where P is the total pressure and K_P is the dissociation equilibrium constant of XY_2.

$2XY_2 (g) \rightleftharpoons 2XY (g) + Y_2(g)$

	XY_2	XY	Y_2
Initial no.of moles	1	-	-
No. of moles dissociated	X	-	-
No. of moles at equilibrium	$(1 - X) \cong 1$	X	X/2

Total no. of moles = 1-X + X + X/2 = 1 + X/2 ≅1

[therefore Given that x<< 1; 1 – x ≅1 and 1 + X/2 ≅ 1]

$$K_p = \frac{(P_{XY})^2(P_{Y_2})}{(P_{XY_2})^2} = \frac{\left(\frac{x}{1}\times P\right)^2\left(\frac{\frac{X}{2}}{1}\times P\right)}{\left(\frac{1}{1}\times P\right)^2}$$

$$K_P = \frac{X^2P^2\times P}{2P^2} \quad \mathbf{= 2K_P = X^3P}$$

47. A sealed container was filled with 1 mol of A_2 (g), 1 mol B_2 (g) at 800 K and total pressure 1.00 bar. Calculate the amounts of the components in the mixture at equilibrium given that K = 1 for the reaction $A_2 (g) + B_2 (g) \rightleftharpoons 2AB (g)$

$A_2 (g) + B_2 (g) \rightleftharpoons 2AB (g)$

	A_2	B_2	AB
Initial Concentration	1	1	-
No. of moles dissociated	X	X	-
No. of moles at equilibrium	1 - x	1- x	2x

Total no. of moles = 1- x + 1 – x + 2x = 2

$$K_p = \frac{(P_{AB})^2}{(P_{A_2})(P_{B_2})} = \frac{\left(\frac{2x}{2}\times P\right)^2}{\left(\frac{(1-x)}{2}\times P\right)\left(\frac{1-x}{2}\times P\right)}$$

$$K_P = \frac{4x^2}{(1-x)^2}$$

Given that $K_P = 1$; $\frac{4x^2}{(1-x)^2} = 1$

⇨ $4x^2 = (1 - x)^2$

⇨ $4x^2 = 1 + x^2 - 2x$

3x2 + 2x -1 = 0

$$x = \frac{-b\pm\sqrt{b^2-4ac}}{2a} = \frac{-2\pm\sqrt{4-4\times3\times-1}}{2(3)} = \frac{-2\pm\sqrt{4+12}}{6} = \frac{-2\pm\sqrt{16}}{6} \qquad = \frac{-2+4}{6};\frac{-2-4}{6} \qquad = \frac{2}{6};\frac{-6}{6}$$

x = 0.33; -1 (not possible)

$[A_2]_{eq} = 1 - x = 1 - 0.33 = 0.67$

$[B_2]_{eq} = 1 - x = 1 - 0.33 = 0.67$

$[AB]_{eq} = 2x = 2 \times 0.33 = 0.66$

48. Deduce the Vant Hoff equation.

This equation gives the quantitative temperature dependence of equilibrium constant (K). The relation between standard free energy change (ΔG) and equilibrium constant is

$$\Delta G = -RT\ln K \qquad \text{-------- (1)}$$

We know that

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \qquad \text{-------- (2)}$$

Substituting (1) in equation (2)

$$-RT\ln K = \Delta H^\circ - T\Delta S^\circ$$

Rearranging (both side divide -RT)

$$\ln K = \frac{-\Delta H^\circ}{RT} + \frac{\Delta S^\circ}{R} \qquad \text{-------- (3)}$$

Differentiating equation (3) with respect to temperature,

$$\frac{d(\ln K)}{dT} = \frac{\Delta H^\circ}{RT^2} \qquad \text{-------- (4)}$$

Equation 4 is known as differential form of van't Hoff equation. On integrating the equation 4, between T_1 and T_2 with their respective equilibrium constants K_1 and K_2.

$$\int_{K_1}^{K_2} d(\ln K) = \frac{\Delta H^\circ}{R}\int_{T_1}^{T_2}\frac{dT}{T^2}$$

$$[\ln K]_{K_1}^{K_2} = \frac{\Delta H^\circ}{R}\left[-\frac{1}{T}\right]_{T_1}^{T_2}$$

$$\ln K_2 - \ln K_1 = \frac{\Delta H^\circ}{R} - \left[\frac{1}{T_2} + \frac{1}{T_1}\right]$$

$$\ln\frac{K_2}{K_1} = \frac{\Delta H^\circ}{R}\left[\frac{T_2 - T_1}{T_2T_1}\right]$$

$$\log\frac{K_2}{K_1} = \frac{\Delta H^\circ}{2.303R}\left[\frac{T_2 - T_1}{T_2T_1}\right] \qquad \text{-------- (5)}$$

Equation 5 is known as integrated form of van't Hoff equation.

49. The equilibrium constant K_P for the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is 8.19×10^2 at 298 K and 4.6×10^{-1} at 498 K. Calculate ΔH^0 for the reaction.

K_{P_1} = 8.19 × 102 T1= 298K

K_{P_2} = 4.6 × 10-2 T2 = 498K

$$\log\left(\frac{K_{P_2}}{K_{P_1}}\right)=\frac{\Delta H^\circ}{2.303R}\left[\frac{T_2-T_1}{T_2T_1}\right]$$

$$\log\left(\frac{4.6\times10^{-1}}{8.19\times10^{2}}\right)=\frac{\Delta H^\circ}{2.303\times8.314}\left(\frac{498-298}{498\times298}\right)$$

$$\frac{-3.2505\times2.303\times8.314\times498\times298}{200}=\Delta H^\circ$$

ΔH = - 46181 J mol^{-1}

ΔH = - 46.18 kJ mol^{-1}

50. The partial pressure of carbon dioxide in the reaction $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ is 1.017×10^{-3} atm at 500^0 C. Calculate K_P at 600^{0}C C for the reaction. ΔH for the reaction is 181 KJ mol^{-1} and does not change in the given range of temperature.

$P_{CO_2} = 1.017\times10^{-3}$ atm ; T = 500^0 C

$K_P = P_{CO_2}$

$K_{P_1} = 1.017\times10^{-3}$; T = 500 + 273 = 773K

K_{P_1} = ? T = 600 + 273 = 873K

ΔH = 181 kJ mol^{-1}

$$\log\left(\frac{K_{P_2}}{K_{P_1}}\right)=\frac{\Delta H^\circ}{2.303R}\left[\frac{T_2-T_1}{T_2T_1}\right]$$

$$\log\left(\frac{K_{P_2}}{1.017\times10^{-3}}\right)=\frac{181\times10^3}{2.303\times8.314}\left(\frac{873-773}{873\times773}\right)$$

$$\log\left(\frac{K_{P_2}}{1.017\times10^{-3}}\right)=\frac{181\times10^3\times100}{2.303\times8.314\times873\times773}$$

$$\frac{K_{P_2}}{1.017\times10^{-3}} = \text{anti log of }(1.40)$$

$$\frac{K_{P_2}}{1.017\times10^{-3}} = 25.12$$

$K_{P_2} = 25.12\times1.017\times10^{-3}$

$\boldsymbol{K_{P_2} = 25.54\times10^{-3}}$

LESSON 8 CHEMICAL EQUILIBRIUM

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
6	6	5	6	6	11	6	6	9	11				

1. Why the chemical equilibrium is referred to as Dynamic equilibrium. (5)
2. Explain Homogeneous equilibrium and Heterogeneous equilibrium give an example. (5) **S20 2M, S21 3M**
3. State law of mass action and mention its unit. (5) M23 5Mi
4. Define equilibrium constant. Give any one application of equilibrium constant. **(MQ19)**
5. Derive the relation between Kp and Kc for a general homogeneous gaseous reaction. (6)**J19 5Mi**
6. What is the relation between Kp and Kc? Give one example for which Kp is equal to Kc? (7) **S21 2M**
7. What is equilibrium constant? (6)
8. If Δn_g = 0, +ve, -ve mention the relationship between Kp and Kc (7)

9. Write a balanced chemical equation for a equilibrium reaction for which the equilibrium constant is given by expression $K_C = \frac{[NH_3]^4[O_2]^5}{[NO]^4[H_2O]^6}$ (BB) **M22 2M**
10. Write Kp, Kc and Δng (i)$H_{2(g)}+I_{2(g)} \rightleftharpoons 2HI_{(g)}$ (ii) $N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$. (iii) $2CO_{(g)} \rightleftharpoons CO_{2(g)} + C_{(S)}$. **J23 2M**
11. Write the Balanced chemical equation for the $K_c = \frac{[CaO_{(s)}][CO_{2(g)}]}{[CaCO_{3(s)}]}$ (8) **M19 5Mii**
12. What is reaction quotient? (11)**S20 5Mii, J23 5Mi, (MQ19 2M)**
13. Application of equilibrium constant. (10)
14. How equilibrium constant help to find the direction of chemical equilibrium? (10)
15. How will you predict the feasibility of a reaction using Q value? (Q compare Kc.) (11)
16. Derive the K_P and Kc value for formation of HI (12) **M24 3M**
17. Derive the K_P and Kc value for Dissociation of PCl_5 (13)
18. Derive the K_P and Kc value for synthesis of ammonia (14) **J22 5M, J24 5M**
19. One mole of H_2 and one mole of I_2 are allowed to attain equilibrium mixture contains 0.4mole of HI. Calculate the equilibrium constant. (13)
20. The equilibrium concentrations of NH_3, N_2 and H_2 are 1.8 x 10^{-2}M and 3x10^{-2}M respectively. Calculate the equilibrium constant for the formation of NH_3 from N_2 andH_2.(14) **M22 3M C**
21. State Le- Chatelier Principle(16) **(GMQ19, M19 3MARK M23 2M**
22. Effect of concentration. (16)
23. If there is no change in concentration, why is the equilibrium considered dynamic? (16) **J24 2M**
24. Effect of Pressure. (17)
25. Effect of Temperature. (18)
26. Effect of catalyst.(18)
27. What is the effect of added inert gas on the reaction at equilibrium? (19) **J19 2MARK**
28. Deduce the Van't Hoff equation. (20) **M24 5M**
29. At particular temperature $K_C=4x10^{-2}$ for the reaction $H_2S_{(g)} \rightleftharpoons H_{2(g)} + ½ S_{2(g)}$ .Calculate the K_C for each of the following reactions. **(27 BBQ$_{45}$) J22 3M**
    (i) $2H_2S_{(g)} \rightleftharpoons 2H_{2(g)} + ½ S_{2(g)}$ (ii) $3H_2S_{(g)} \rightleftharpoons 3H_{2(g)}+3/2S_{2(g)}$

Unit 9:Solutions

31. Define (i) molality (ii) Normality

(i) molality: Molarity is defined as the number of moles of the solute per kilogram of the solvent

$$\textbf{Molality} = \frac{No.\ of\ moles\ of\ solute}{Mass\ of\ the\ solvent\ (in\ kg)}$$

(ii) Normality: Normality is defined as number of gram equivanlents of solute present in one litre of solution.

$$\text{Normality} = \frac{No.\ of\ gram\ equivalent\ of\ solute}{Volume\ of\ solution\ (in\ L)}$$

32. What is a vapour pressure of liquid? What is relative lowering of vapour pressure?

Vapour pressure of liquid :

The pressure of the vapour in equilibrium with its liquid is called vapour pressure of the liquid at the given temperature.

Relative lowering of vapour pressure:

The ratio of lowering of vapour pressure to vapour pressure of pure solvent.

Relative lowering of vapour pressure$= \frac{P^{\circ}_{solvent} - P_{solution}}{P^{\circ}_{solvent}}$

33. State and explain Henry's law

Henry's law states that, "the partial pressure of the gas in vapour phase is directly proportional to the mole fraction(x) of the gaseous solute in the solution at low concentrations".

Henry's law can be expressed as, $P_{\text{solute}} \ \alpha \ x_{\text{solute in solution}}$

$$P_{\text{solute}} = K_H \, x_{\text{solute in solution}}$$

34. State Raoult law and obtain expression for lowering of vapour pressure when nonvolatile solute is dissolved in solvent.

Raoult law states that "in the case of a solution of volatile liquids, the partial vapour pressure of each component (A & B) of the solution is directly proportional to its mole fraction".

According to Raoult's law

$P_A \ \alpha \ x_A$; $P_A = k\, x_A$

when $x_A = 1$, $k = P^{\circ}_A$

where P°_A is the vapour pressure of pure component 'A' at the same temperature.

Therefore, $P_A = P^{\circ}_A \, x_A$

When a nonvolatile solute is dissolved in a pure solvent, the vapour pressure of the pure solvent will decrease.

Volatile solvent Particles

Nonvolatile solute particles

$P_{\text{solution}} \propto x_A$; Where x_A is the mole fraction of the solvent

$P_{\text{solution}} = k\, x_A$ When $x_A = 1$, $K = P^{o}_{\text{solvent}}$

(P^{o}_{solvent} is the partial pressure of pure solvent)

$P_{\text{solution}} = P_{\text{ solvent}}\, x_A$

$\frac{P_{solution}}{P^{\circ}_{solvent}} = x_A$

$$1 - \frac{P_{solution}}{P^{\circ}_{solvent}} = 1 - x_A$$

$$\frac{P^{\circ}_{solvent} - P_{solution}}{P^{\circ}_{solvent}} = x_B$$

Where x_B is the fraction of the solute

($\therefore x_A + x_B = 1, x_B = 1 - x_A$)

The above expression gives the relative lowering of vapour pressure. Based on this expression, Raoult's Law can also be stated as "the relative lowering of vapour pressure of an ideal solution containing the nonvolatile solute is equal to the mole fraction of the solute at a given temperature"

35. What is molal depression constant? Does it depend on nature of the solute ?

$$\Delta T_f = K_f \,.\, m$$

If m=1 then $\Delta T_f = K_f$

The molal depression constant K_f is equal to the depression is freezing point for 1 molal solution. Depression constant **does not depend on the nature of the solute**. (but depends on the number of solution particles in the solution solvent).

36. What is osmosis?

Osmosis is a spontaneous process by which the solvent molecules pass through a semi permeable membrane from a solution of lower concentration to a solution of higher concentration.

37. Define the term 'isotonic solution'.

Two solutions having same osmotic pressure at a given temperature are called isotonic solutions.

38. You are provided with a solid 'A' and three solutions of A dissolved in water – one saturated, one unsaturated, and one super saturated. How would you determine which solution is which ?

Saturated Solution: It can dissolve salt an additional to it.

(A Solution in which no more solute can be dissolved in a definite amount of solvent at a given temperature is called a saturated solution.)

Unsaturated solution: Further solubility of salt do not takes place but solubility can takes place on heating.

(A solution which has more of solute than the saturated at a given temperature is called super saturated solution.)

Super saturated solution: Solubility of salt do not takes place on even an further heating.

(A Solution which has more of solute than the saturated solution at a given temperature is called super saturated solution.)

39. Explain the effect of pressure on the solubility.

The change in pressure does not have any significant effect in the solubility of solids and liquids as they are not compressible. However, **the solubility of gases generally increases with increase of pressure.**

Consider a saturated solution of a gaseous solute dissolved in a liquid solvent in a closed container. In such a system, the following equilibrium exists.

Gas (in gaseous state) $\leftrightarrows$ Gas (in solution)

According to Le-Chatelier principle, **the increase in pressure will shift the equilibrium in the direction which will reduce the pressure.** Therefore, more number of gaseous molecules dissolves in the solvent and the solubility increases.

40. A sample of 12 M Concentrated hydrochloric acid has a density 1.2 gL^{-1} Calculate the molality.

Molarity = 12 M HCl

Density of the solution = 1.2 gL^{-1}

In 12 M HCl solution, there are 12 moles of HCl in 1 litre of the solution

$$\textbf{Molarity} = \frac{\textit{No. of moles of solute}}{\textit{Mass of the solvent (in kg)}}$$

Calculate mass of water (solvent)

Mass of 1 litre HCl solution = density × volume

= 1.2 gmL^{-1} × 1000mL

= 1200 g

Mass of HCl = No. of moles of HCl × molar mass of HCl

= 12 mol × 36.5g mol^{-1} = 438 g

Mass of water = mass of HCl solution – mass of HCl

Mass of water = 1200 – 438 = 762 g

$$\text{Molarity} = \frac{12}{0.762} = \mathbf{15.75\ m}$$

41. A 0.25 M glucose solution at 370.28 K has approximately the pressure as blood does what is the osmotic pressure of blood ?

C = 0.25 M; T = 370.28 K

$(\pi)_{\text{glucose}}$ = CRT

(π) = 0.25 mol L^{-1} × 0.082L atm K^{-1} mol^{-1} × 370.28 K **(π) = 7.59 atm**

42. Calculate the molality of a solution containing 7.5 g of glycine (NH_2-CH_2 -COOH) dissolved in 500 g of water.

$$\textbf{Molarity} = \frac{\textit{No. of moles of solute}}{\textit{Mass of the solvent (in kg)}}$$

$$\textbf{No. of moles of glycine} = \frac{\textit{mass of glycine}}{\textit{molar mass of glycine}} = \frac{7.5}{75} = \mathbf{0.1}$$

$$\text{Molarity} = \frac{0.1}{0.5\ Kg} = 0.2\text{m}$$

43. Which solution has the lower freening point?10 g of methanol (CH_3OH) in 100g of water (or) 20 g of ethanol (C_2H_5OH) in 200 g of water. $\Delta T_f = K_f\ m$

i.e., $\Delta T_f\ \alpha\ m$

$$m_{CH3\ -\ OH} = \frac{\left(\frac{10}{32}\right)}{0.1} = 3.125\text{m} \quad m_{C2H5\ -\ OH} = \frac{\left(\frac{20}{46}\right)}{0.2} = 2.174\text{m}$$

∴ *Depression* in freezing point is more in methanol solution and it will have lower freezing point.

44. How many moles of solute particles are present in one litre of 10^{-4} M potassium sulphate?

In 10^{-4} M K_2SO_4 solution, there are 10^{-4} moles of potassium sulphate.

K_2SO_4 molecule contains 3 ions ($2K^+$ and $1SO_4^{2-}$)

1moleof K_2SO_4 contains $3 \times 6.023 \times 10^{23}$ ions

10^{-4} mole of K_2SO_4 contains $3 \times 6.023 \times 10^{23} \times 10^{-4}$ ions

$= 18.069 \times 10^{19}$ [therefore $n_1 >> n_2$; $n_1 + n_2 \approx n_1$]

45. Henry's law constant for solubility of methane in benzene is $4.2\text{x}10^{-5}$ mm Hg at a particular constant temperature At this temperature. Calculate the solubility of methane at i) 750 mm Hg ii) 840 mm Hg

$(K_H)_{benzene} = 4.2 \times 10^{-5}$ mm Hg

Solubility of methane = ? ; P = 750 mm Hg ; P = 840 mm Hg

According to Henrys Law,

$P = K_H\ .\ X_{\text{in solution}}$

750 mm Hg = $4.2 \times 10_{-5}$ mm Hg . $X_{\text{in solution}}$

$$\rightarrow X_{\text{in solution}} = \frac{750}{4.2 \times 10^{-5}}$$

i.e, solubility = 178.5×10^5

similarly at P = 840 mm Hg

$$\text{solubility} = \frac{840}{4.2 \times 10^{-5}} = 200 \times 10^{-5}$$

46. The observed depression in freezing point of water for a particular solution is 0.093^0C. Calculate the concentration of the solution in molality. Given that molal depression constant for water is 1.86 KKg mol^{-1}

$\Delta T_f = 0.093$ C = 0.093K; m = ?

$K_f = 1.86$ C kg mol^{-1}

$\Delta T_f = K_f\ .\ m$ $$m = \frac{0.093\ K}{1.86K\ kg\ mol^{-1}} = 0.05\ \text{mol kg}^{-1} = 0.05\text{m}$$

47. The vapour pressure of pure benzene (C_6H_6) at a given temperature is 640 mm Hg. 2.2 g of non-volatile solute is added to 40 g of benzene. The vapour pressure of the solution is 600 mm Hg. Calculate the molar mass of the solute?

P_{C6H6} = 640mm Hg

W_2 = 2.2 g (non volabile solute); W_1 = 40 g (benzene)

$P_{solution}$ = 600mm Hg; M_2 = ?

$\frac{P^\circ - P}{P^\circ} = x_2$; $\frac{640-600}{640} = \frac{n_2}{n_1+n_2}$; $\frac{40}{640} = \frac{n_2}{n_1}$

$0.0625 = \frac{W_2 \times W_1}{M_2 \times M_1}$ $M_2 = \frac{2.2 \times 78}{0.0625 \times 40}$ = **68.64 g mol^{-1}**

IMPORTANT QUESTION 9. Solutions

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
7	9	6	11	11	6	5	6	5	4				

1. Define solution, solute, solvent. (31)
2. Explain the different type of solutions based on the physical state of the solute and solvent. (31)
3. Define Molality **J23 3M** and Molarity (32)
4. Calculate the molality of the solution containing 90g of glucose dissolved in 2kg of water. (32) **M24 2M Compulsory**
5. Define Normality and Formality (32)
6. Define Mole fraction and Mass percentage (33)
7. Calculate the mole fraction of methanol and water when 0.5mole of methanol is mixed with 1.5moles of water. (33) **S20 3MARK**
8. Define Molarity. If 5.6 g of KOH is present in 250 ml of the solution, calculate the molarity of the solution. (34) **(MQ19)**
9. Define volume percentage and mass by volume percentage (34)
10. Define parts per million (34)
11. 50g of tap water contains 20mg of dissolved solids. what is the TDS value in ppm?(34) **J22 2M**
12. What are the advantages of using standard solutions? (35)
13. What are standard and working solutions? (35)
14. Define solubility (36)
15. What are the factors influencing the solubility? (36)
16. What is the nature of solute and solvent? (36)
17. How does temperature affect the solubility? (36)
18. Draw and explain the graph obtained by plotting solubility versus temperature for calcium chloride. (37) **J19 5Mii**
19. Explain the effect of pressure on solubility? (38) **J24 5Mi**
20. State Henry's law (38)
21. NH_3 and HCl do not obey Henry's law. Why?(38) **M19 5Mi**

22. What are the limitations of Henry's law? (40) M23 5Mii
23. Define vapour pressure (41)
24. State Raoult's law (43) **M24 5Mi**
25. How will you compare Raoult's law with Henry's law? (45)
26. What are ideal solutions? Give example. (45) **S21 3M**
27. What are conditions when a solution tends to behave like an ideal solution? (46) **J22 3M**
28. What are non-ideal solutions? Give example. (46)
29. What are the conditions for Non ideal solutions? (46)
30. Explain the positive deviation of non-ideal solutions (46)
31. Explain the negative deviation of non-ideal solutions (47)
32. Explain the factors responsible for deviation from Raoult's law (48)
33. What are colligative properties? (49) **J23 5Mii**
34. What is relative lowering of vapour pressure? (49)
35. Determination of Molar mass weights from relative lowering of vapour pressure (50) Or Write the formula to calculate the molar mass of a solute from relative lowering of vapour pressure values. (50) **(M22 3M)**
36. Determination of molar mass of solute from elevation of boiling point (52) **M22 5M**
37. What is Ebullioscopic constant? (52)
38. Determination of molar mass of solute from depression in freezing point (54)
39. What is molal depression constant? (54) **S21 2M**
40. What is Cryoscopic constant? (54)
41. Define osmosis (55)
42. Define osmotic pressure? (55) M23 2M
43. What is isotonic solution? (56) **M19 3MARK, M22 2M, J24 3M**
44. How will you determine the molar mass of a solute from osmotic pressure (56) **S21 5M**
45. Explain the application of reverse osmosis in water purification (57)
46. Define reverse Osmosis (57)
47. What is abnormal molar mass? (58)
48. What is Van't Hoff factor 'I'? (58) **S20 5Mi**
49. Calculate the mass of non-volatile solute (molar mass 80 $gmol^{-1}$) which should be dissolved in 92 g of toluene is reduced to its Vapour pressure to 90%. BBQ20 63**(MQ19)**

Unit: 10 Chemical Bonding

31. Define the following i) Bond order ii) Hybridisation iii) σ- bond

i) Bond order: ***The number of bonds*** formed between the two bonded atoms in a molecule is called the bond order.

ii) Hybridisation: Hybridisation is the process of mixing of atomic orbitals of the same atom with comparable energy to form equal number of new equivalent orbitals with same energy.

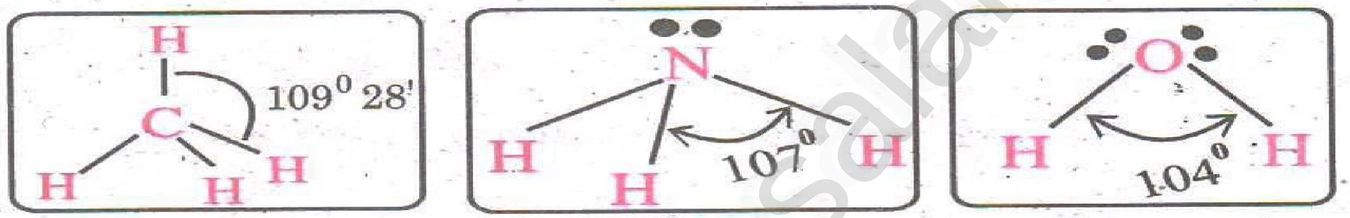
iii) σ- bond: When two atomic orbitals ***overlap linearly along the axis***, the resultant bond is called a sigma **(σ)** bond.

32. What is a pi bond?

When two atomic orbitals ***overlap sideways***, the resultant covalent bond is called a pi (π) bond.

33. In CH_4, NH_3 and H_2O, the central atom undergoes sp^3 Hybridisation - yet their bond angles are different. why?

- Bond angles are different due to **the presence of lone pair of electrons.**
- In case of **CH_4,** there are 4 bond pairs and no lone pair of electrons. So it remains in its regular geometry. i.e., tetrahedral with bond angle =**109°28'.**
- **H_2O** has 2 bond pairs and 2 lone pairs. There is large repulsion between lp-lp. Again repulsion between lp-lp is more than that of 2 bond pairs. So 2 bonds are more restricted to form inverted V shape (or) bent shape molecule with a bond angle of **104°35'**
- **NH_3** has 3 bond pairs and 1 lone pair. There is repulsion between lp-bp. So 3 bonds are more restricted to form pyramidal shape with bond angle equal to **107°18'**.



(no lone pair)

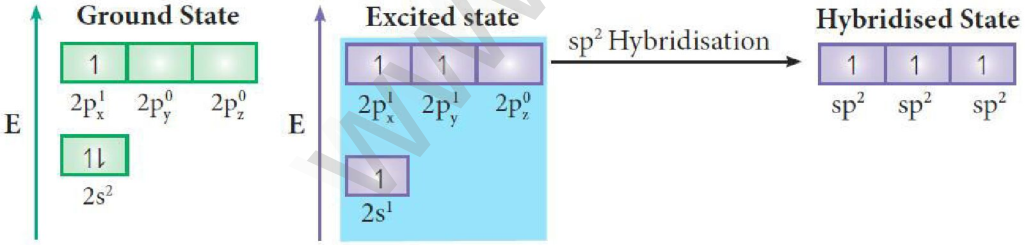
one lone pair so less than tetra hedral (lp - bp repulsion)

two lone pairs (lp-lp and lp-bp repulsions)

34. **Explain Sp^2 Hybridisation in BF_3.**

sp^2 Hybridisation: Consider boron trifluoride molecule.

The valence shell electronic configuration of boron atom is [He] $2s^2$ $2p^1$



In boron, the s orbital and two p orbitals (p_x and p_y) in the valence shell hybridises, to generate three equivalent sp^2 orbitals. These three orbitals lie in the same xy plane and the angle between any two orbitals is equal to 120^0.

Overlap with $2p_z$ orbitals of fluorine:

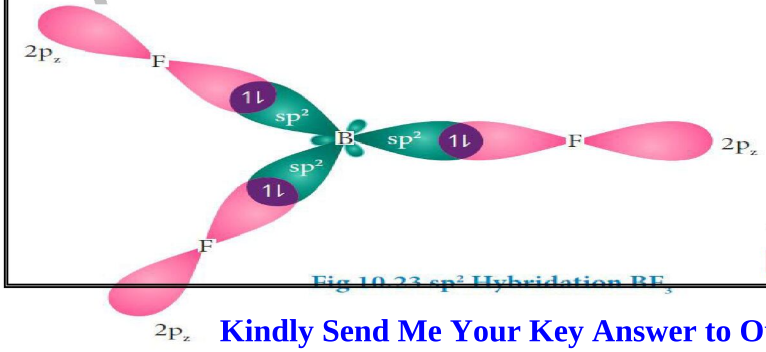


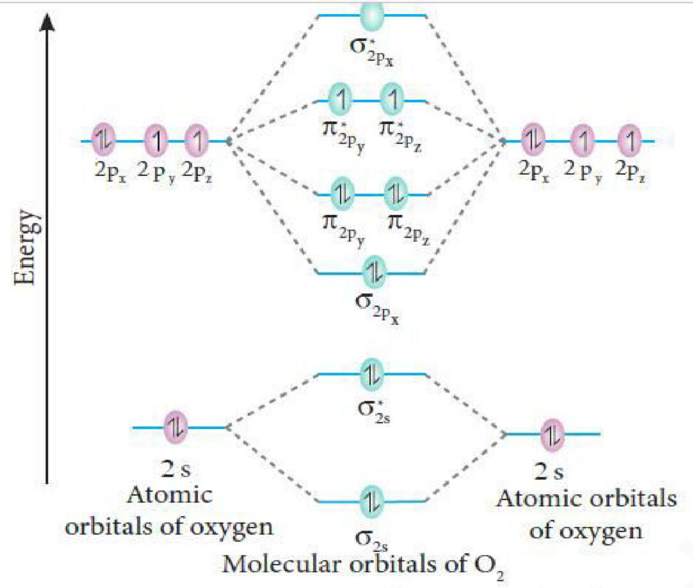
Fig 10.23 sp^2 Hybridisation BF_3

The three sp^2 hybridised orbitals of boron now overlap with the 2pz orbitals of fluorine (3 atoms). This overlap takes place along the axis.

35. Draw the M.O diagram for oxygen molecule calculate its bond order and show that O_2 is paramagnetic.

Molecular orbital diagram of oxygen molecule (O_2)

Electronic configuration of O atom $1s^2\, 2s^2\, 2p^4$



Molecular orbital diagram of oxygen molecule (O_2)

Electronic configuration of O atom $1s^2\, 2s^2\, 2p^4$

Electronic configuration of O_2 molecule $\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \sigma_{2p_x}^2\ \pi^2_{2p_y}, \pi^2_{2p_z}, \pi^{*1}_{2p_y}, \pi^{*1}_{2p_z},$

Bond order $= \frac{N_b - N_a}{2} = \frac{10-6}{2} = 2$

Molecule has two unpaired electrons hence it is paramagnetic.

Fig 10.35 MO Diagram for O_2 molecule

Bond order $= \frac{N_b - N_a}{2} = \frac{10-6}{2} = 2$

Molecule has two unpaired electrons hence it is paramagnetic.

36. Draw MO diagram of CO and calculate its bond order.

Bonding in some heteronuclear di-atomic molecules

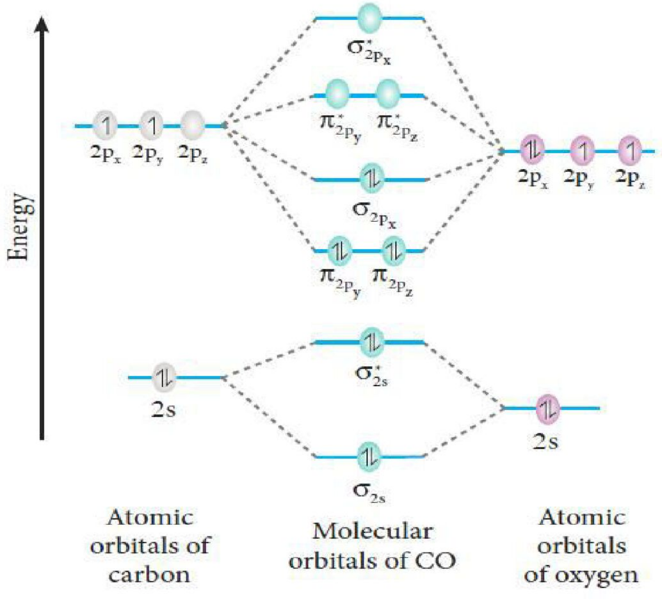


Fig 10.36 MO Diagram for CO molecule

Bonding in some heteronuclear di-atomic molecules

Molecular orbital diagram of Carbon monoxide molecule (CO)

Electronic configuration of C atom $1s^2\ 2s^2\ 2p^2$

Electronic configuration of O atom $1s^2\ 2s^2\ 2p^4$

Electronic configuration of CO molecule $\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \pi^2_{2p_y}, \pi^2_{2p_z}, \sigma^2_{2p_x}$

Bond order $= \frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

Molecule has no unpaired electrons hence it is diamagnetic.

37. What do you understand by Linear combination of atomic orbitals in MO theory?

i) The wave functions for the molecular orbitals can be obtained by solving Schrodinger wave equation for the molecule.

To obtain the ***wave function for molecular orbitals***.

The most common method is the linear combination of atomic orbitals (LCAO).

ii) **atomic orbitals** are represented by the **wave function Ψ.**

Let us consider two atomic orbitals represented by the wave function ψ_A and ψ_B with comparable energy, combines to form two molecular orbitals.

iii) One is bonding molecular orbital($\psi_{bonding}$) and the other is antibonding molecular orbital($\psi_{antibonding}$).

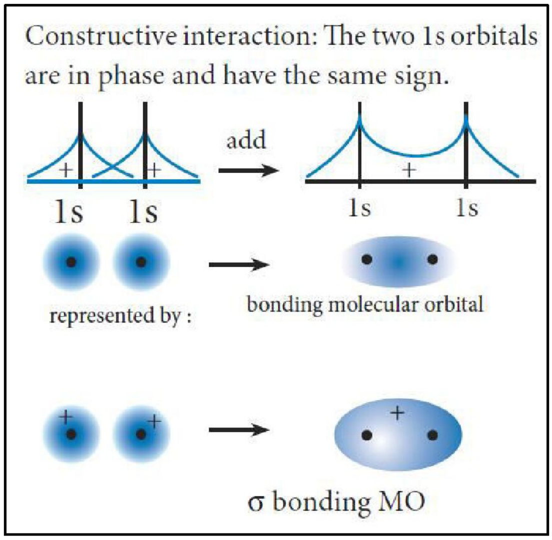
iv) The wave functions for these two molecular orbitals can be obtained by the linear combination of the atomic orbitals ψ_A and ψ_B as below.

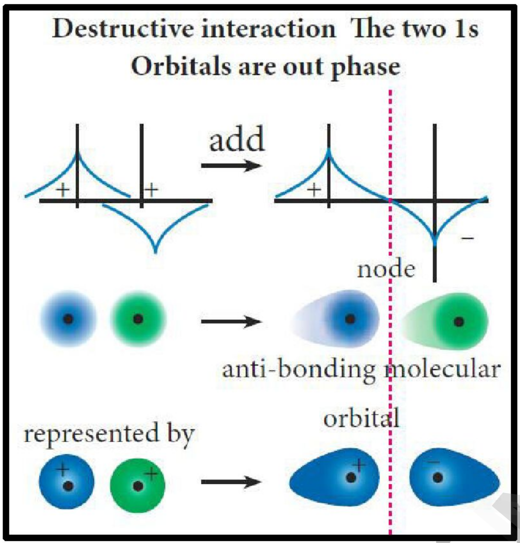
$$\psi_{bonding} = \psi_A + \psi_B$$

$$\psi_{antibonding} = \psi_A - \psi_B$$

v) The formation of **bonding molecular orbital** can be considered as the result of **constructive interference of the atomic orbitals** and the formation of **anti-bonding molecular orbital** can be the result of the **destructive interference** of the atomic orbitals.

vi) The formation of the two molecular orbitals from two 1s orbitals is shown below.





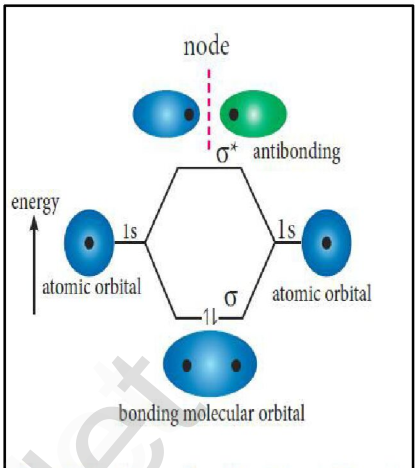


Fig 10.29 Linear Combination of atomic orbitals

38. Discuss the formation of N_2 molecule using MO Theory

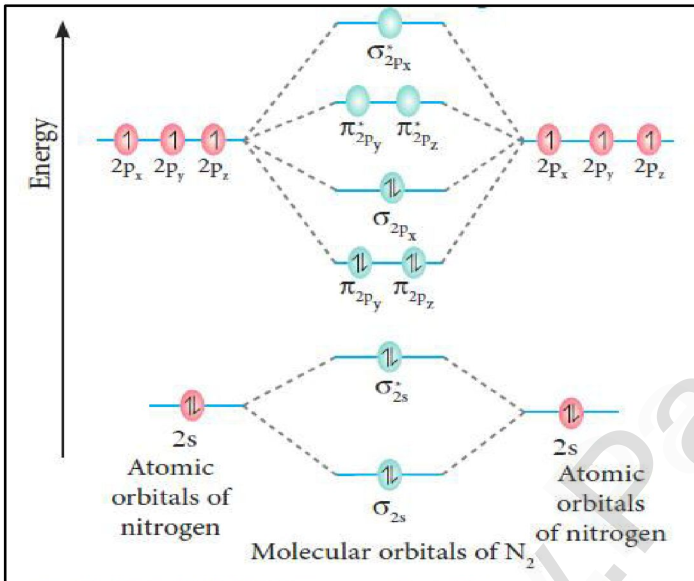


Fig 10.34 MO Diagram for N_2 molecule

Molecular orbital diagram of nitrogen molecule (N_2)

Electronic configuration of N atom $1s^2 2s^2 2p^3$

Electronic configuration of N_2 molecule $\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \pi^2_{2p_y}, \pi^2_{2p_z} \sigma_2{}^2px$

$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$$

Molecule has no unpaired electrons hence it is diamagnetic.

39. What is dipole moment?

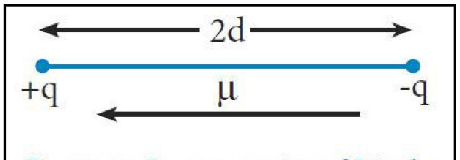


Fig 10. 15 Representation of Dipole

i)The polarity of a covalent bond can be measured in terms of dipole moment which is defined as **$\mu = q \times 2d$**

Where, μ - dipole moment,

q - the charge and

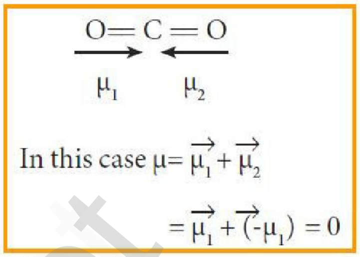
2d - distance between the two charges.

ii) The dipole moment is a vector quantity and the direction of the dipole moment points from the negative charge to positive charge.

iii) The unit of dipole moment is coulomb metre (Cm) (Debye= 3.336×10^{-30} C m).

40. Linear form of carbon dioxide molecule has two polar bonds. yet the molecule has Zero dipole moment why?

- The **linear form** of carbon dioxide has **zero dipole moment**, even though it has **two polar bonds.**
- In CO_2, the dipole moments of two polar bonds (C=O) are equal in magnitude but have opposite direction.
- Hence, the net dipole moment of the CO_2 is, $\mu = \mu_1 + \mu_2 = \mu_1 + (-\mu_1) = 0$



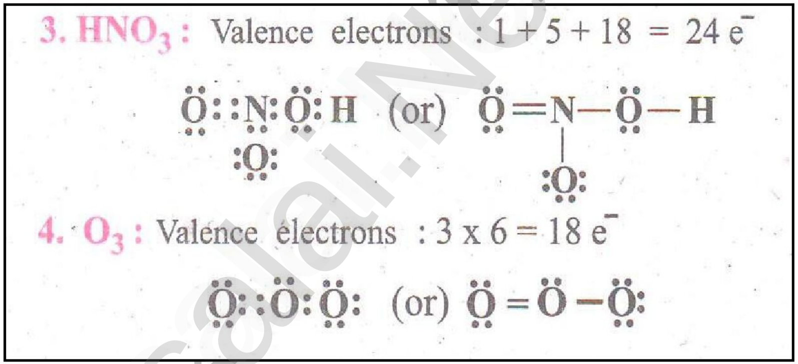
In this case $\mu = \vec{\mu_1} + \vec{\mu_2}$

$= \vec{\mu_1} + (\vec{-\mu_1}) = 0$

41. Draw the Lewis structures for the following species.

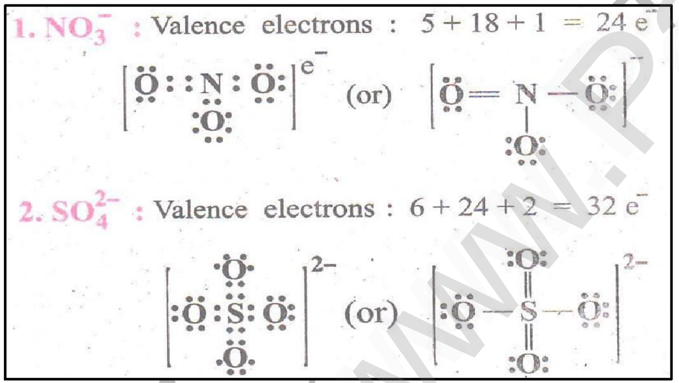
i) NO_3^- **ii) SO_4^{2-}**

iii) HNO_3 **iv) O_3**



3. HNO_3 : Valence electrons : 1 + 5 + 18 = 24 e^-

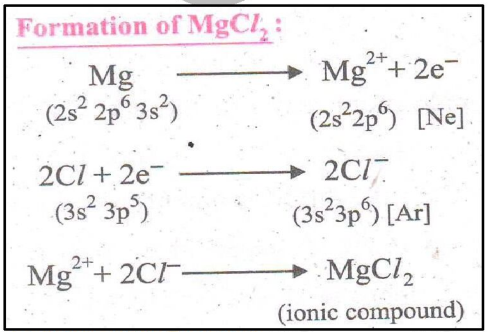
4. O_3 : Valence electrons : 3 x 6 = 18 e^-



1. NO_3^- : Valence electrons : 5 + 18 + 1 = 24 e^-

2. SO_4^{2-} : Valence electrons : 6 + 24 + 2 = 32 e^-

42. Explain the bond formation in $BeCl_2$ and $MgCl_2$.



Formation of $MgCl_2$:

$$Mg\ (2s^2\,2p^6\,3s^2) \longrightarrow Mg^{2+} + 2e^-\ (2s^2 2p^6)\ [Ne]$$

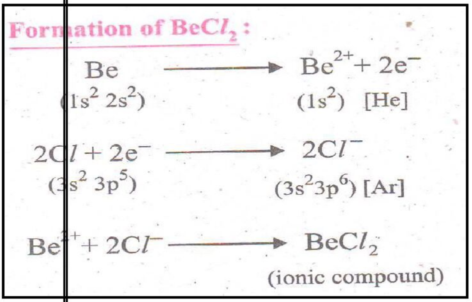
$$2Cl + 2e^-\ (3s^2\,3p^5) \longrightarrow 2Cl^-\ (3s^2 3p^6)\ [Ar]$$

$$Mg^{2+} + 2Cl^- \longrightarrow MgCl_2$$

(ionic compound)

Bond formation in $MgCl_2$:

Mg is a metal with 2 valence electrons. Cl is a non-metal with 7 valence electrons and hence it needs one electron to complete the octet. Thus Mg donates one electron to each Cl and becomes positively charged. Each Cl accepts one electron and becomes negatively charged.



Formation of $BeCl_2$:

$Be \longrightarrow Be^{2+} + 2e^-$
$(1s^2\,2s^2) \quad (1s^2)\ [He]$

$2Cl + 2e^- \longrightarrow 2Cl^-$
$(3s^2\,3p^5) \quad (3s^2 3p^6)\ [Ar]$

$Be^{2+} + 2Cl^- \longrightarrow BeCl_2$
(ionic compound)

Bond formation in $BeCl_2$:

In AB_2 type the central atom A has two bond pairs only the two bond pairs orient themselves at 180 to have minimum repulsion among themselves.

$BeCl_2$ is linear and forms a pair of covalent bonds instead for forming ionic bonds. Be uses its two hybrid orbitals to form two Be – Cl bonds. The bond angle is 180 .

43. Which bond is stronger σ or π? Why?

- Sigma (σ) bonds are strong than pi (π) bonds.
- Sigma bonds are formed by ***head on overlap*** of atomic orbitals so extent of overlapping is maximum. ***(axial overlapping).***
- Pi bonds are weaker, since according to quantum mechanics, the orbital path are parallel to there is much less overlap between the p orbitals. ***(lateral overlapping).***
- Pi bonds are more diffused bonds than sigma bonds.

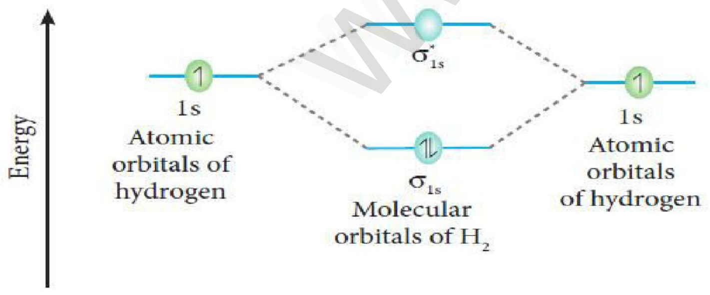
44. Define bond energy.

The bond enthalpy (bond energy) is defined as **the minimum amount of energy required to break one mole of a particular bond in molecules in their gaseous state.**

The unit of bond energy is **$kJmol^{-1}$.**

45. Hydrogen gas is diatomic whereas inert gases are monoatomic – explain on the basis of MO theory.

- Hydrogen gas is diatomic whereas inert gases are monoatomic because noble gases are having its valence shell filled with octet electrons.
- So there is no necessity for the noble gases to involve in the formation of covalent bonds.
- This could be further explained by MO diagram of H_2 and He_2
  - **Molecular orbital diagram of hydrogen molecule (H_2):**



Molecular orbital diagram of hydrogen molecule (H_2)

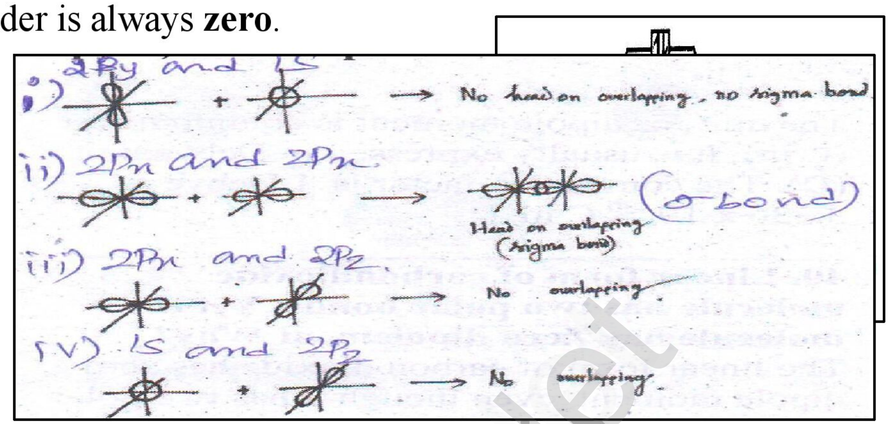
Electronic configuration of H atom $1s^1$

Electronic configuration of H_2 molecule σ_{1s}^2

Bond order $= \dfrac{N_b - N_a}{2} = \dfrac{2-0}{2} = 1$

Molecule has no unpaired electrons hence it is diamagnetic.

- In inert gases, total number of bonding electrons will exactly be equal to be total number of antibonding electrons. Their bond order is always **zero**.
  - They are monoatomic.



Inert gas example - He₂ molecule:

- The electronic configuration of He (Z = 2) in the ground state is $1s^2$
- $He_2 : (\sigma_{1s})^2 (\sigma^*_{1s})^2$

N_b =2 and N_a = 2 Bond order = $\frac{N_b - N_a}{2} = \frac{2-2}{2} = 0$

46. What is Polar Covalent bond? Explain with example.

Polar covalent bond is formed between atoms having different electronegativities, the atom with higher electronegativity will have greater tendency to attract the shared pair of electrons towards itself than other atom.

Example: HF

The electronegativities of hydrogen and fluorine on Pauling's scale are 2.1 and 4 respectively. It means that fluorine attracts the shared pair of electrons approximately twice as much as the hydrogen which leads to partial negative charge on fluorine and partial positive charge on hydrogen. Hence, the H-F bond is said to be polar covalent bond.

47. Considering x- axis as molecular axis, which out of the following will form a sigma bond?

i) 1s and $2p_y$ ii) $2P_x$ and $2P_x$ iii) $2P_x$ and $2P_z$ iv) 1s and $2P_z$

i) 1s and $2p_y$ – No sigma bond

ii) $2P_x$ and $2P_x$ – sigma bond

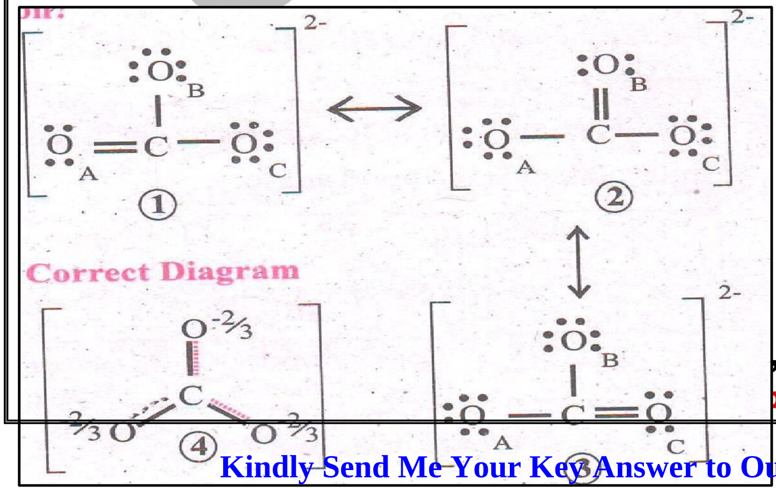
iii) $2P_x$ and $2P_z$– No sigma bond

iv) 1s and $2P_z$– No sigma bond

Along x- axis as molecular axis, only $2p_x$ and $2p_x$ can form σ- bond.

48. Explain resonance with reference to carbonate ion?

The skeletal structure of carbonate ion (The



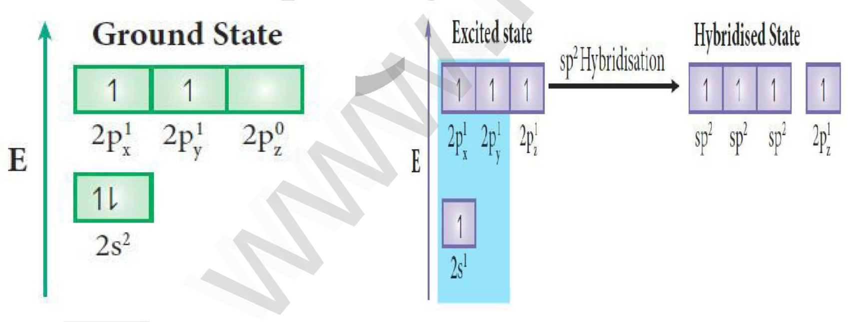
xygen atoms are denoted as O_A, O_B & O_C)

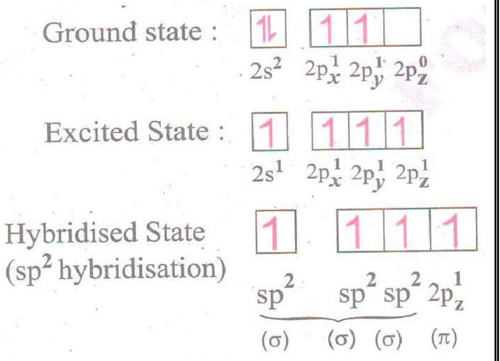
Total number of valence electrons = [1 x 4(carbon)] + [3 x 6 (oxygen)] + [2 (charge)] = 24 electrons.

- They only differ in the position of bonding and lone pair of electrons. Such structures are called resonance structures (canonical structures) and this phenomenon is called **resonance.**
- It is evident from the experimental results that all carbon-oxygen bonds in carbonate ion are equivalent. The actual structure of the molecules is said to be the resonance hybrid, an average of these three resonance forms.
- However, the following structure (4) gives a qualitative idea about the correct structure.
- It is found that the energy of the resonance hybrid (structure 4) is lower than that of all possible canonical structures (Structure 1, 2 & 3). The difference in energy between structure 1 or 2 or 3, (most stable canonical structure) and structure 4 (resonance hybrid) is called resonance energy.

49. Explain the bond formation in ethylene and acetylene.

- The bonding in ethylene can be explained using hybridisation concept.
- The molecular formula of ethylene is C_2H_4. The valency of carbon is 4.
- The electronic configuration of valence shell of carbon in ground state is [He]$2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0$.
- To satisfy the valency of carbon promotes an electron from 2s orbital to 2pz orbital in the excited state.
- In ethylene both the carbon atoms undergoes sp^2 hybridisation involving 2s, $2p_x$ and $2p_y$ orbitals, resulting in three equivalent sp^2 hybridised orbitals lying in the xy plane at an angle of 120^0 to each other. The unhybridised $2p_z$ orbital lies perpendicular to the xy plane.





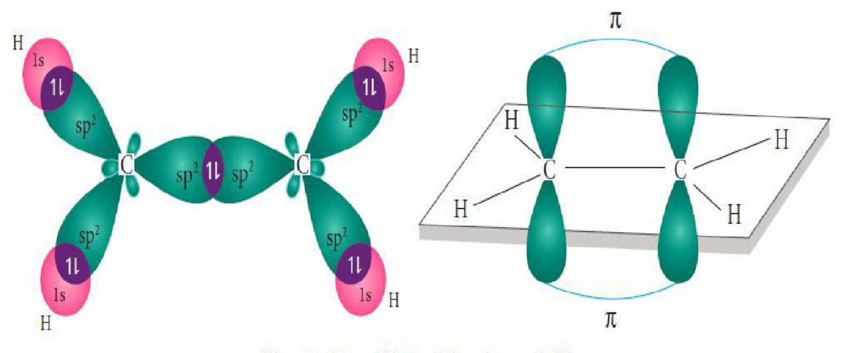


Fig 10.27 sp^2 Hybridisation : C_2H_4

Formation of sigma bond:

One of the sp^2 hybridised orbitals of each carbon lying on the molecular axis (x-axis) linearly overlaps with each other resulting in the formation a **C-C sigma bond.**

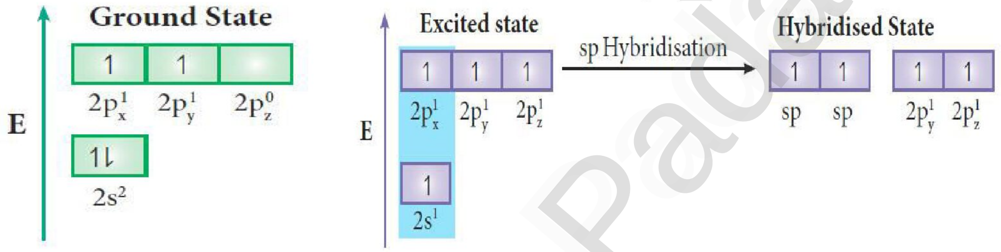
Other two sp^2 hybridised orbitals of both carbons linearly overlap with the four 1s orbitals of four hydrogen atoms leading to the formation of **two C-H sigma bonds on each carbon.**

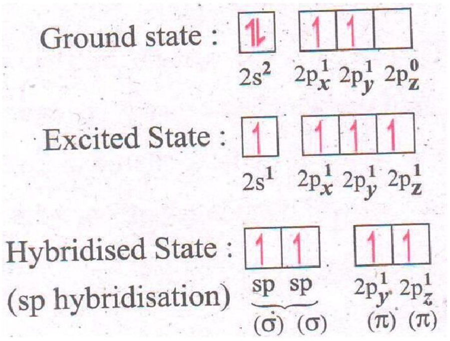
Formation of pi bond:

The unhybridised $2p_z$ orbital of both carbon atoms can overlap only sideways as they are not in the molecular axis. This lateral overlap results in the formation a pi bond between the two carbon atoms.

Bonding in acetylene:

- The bonding in acetylene can be explained using hybridization concept.
- The molecular formula of acetylene is C_2H_2.
- The electronic configuration of valence shell of carbon in ground state is $[He]2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0$.
- To satisfy the valency of carbon promotes an electron from 2s orbital to $2p_z$ orbital in the excited state.
- In acetylene molecule, both the carbon atoms are in sp hybridised state.
- The 2s and $2p_x$ orbitals, resulting in two equivalent sp hybridised orbitals lying in a straight line along the molecular axis (x-axis).
- The unhybridised $2p_y$ and $2p_z$ orbitals lie perpendicular to the molecular axis.





Formation of sigma bond:

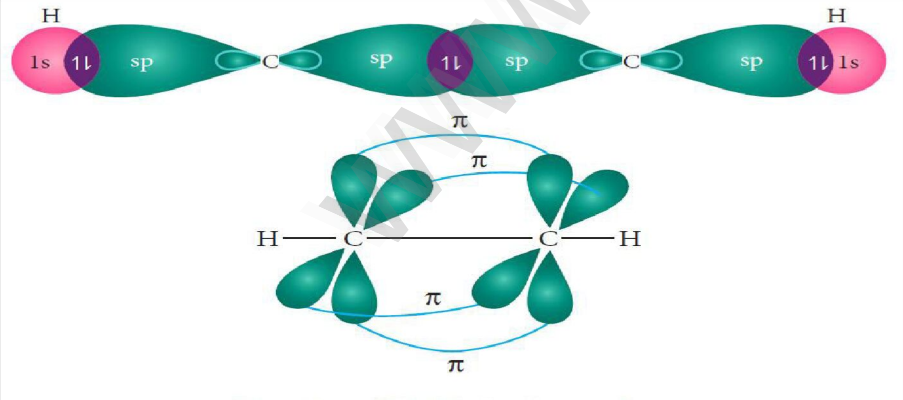


Fig 10.28 sp Hybridisation in acetylene

One of the two sp hybridised orbitals of each carbon linearly overlaps with each other resulting in the **formation a C-C sigma bond.** The other sp hybridised orbital of both carbons linearly overlap with the two 1s orbitals of two hydrogen atoms leading to the formation of one C-H sigma bonds on each carbon.

Formation of pi bond:

The unhybridised $2p_y$ and $2p_z$ orbitals of each carbon overlap sideways. This lateral overlap results in the formation of two pi bonds (p_y-p_y and p_z-p_z) between the two carbon atoms.

50. What type of hybridisations are possible in the following geometeries?

a) octahedral b) tetrahedral c) square planer.

Molecular Geometry	**Hybridisation**	**Example**
Octahedral	sp^3d^2	SF_6
Tetrahedral	sp^3	CH_4
Square planner	dsp^2	$[Ni(CN)_4]$

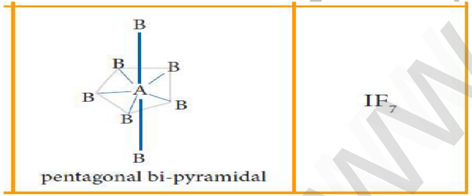
51. Explain VSEPR theory. Applying this theory to predict the shapes of IF_7, and SF_6.

1. The shape of the molecules depends on the number of valence shell electron pair around the central atom.
2. There are two types of electron pairs namely **bond pairs** and **lone pairs**.
3. The **bond pair** of electrons are those ***shared between two atoms***, while the **lone pairs** are the valence electron pairs that are ***not involved in bonding.***
4. Each pair of valence electrons around the central atom repels each other and hence, they are located as far away as possible in three dimensional spaces to minimize the repulsion between them.
5. The repulsive interaction between the different types of electron pairs is in the following order.

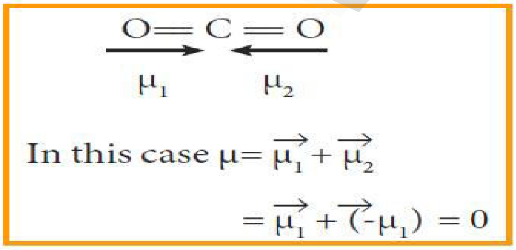
   lp - lp > lp - bp> bp-bp lp- lone pair; bp- bond pair
6. The lone pairs occupy more space and have greater repulsive power than the bond pairs in a molecule.

SF_6 shape - ***Octahedral***

IF_7 Shape – ***pentagonal bi-pyramidal***

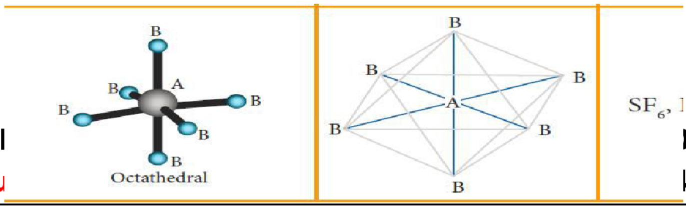


52. CO_2 and H_2O both are triatomic molecule but their dipole moment values are different. Why?



$$\mu = \mu_1 + \mu_2 = \mu_1 + (-\mu_1) = 0$$

In case of **water** net dipole moment is the



vector sum of $\mu_1 + \mu_2 = 1.85D$

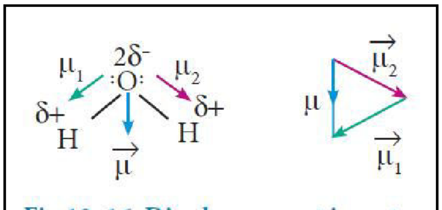


Fig 10. 16 Dipole moment in water

53. Which one of the following has highest bond order? N_2, N_2^+ or N_2^-

Bond order $= \frac{N_b - N_a}{2}$

Where N_b = Number of electrons in bonding molecular orbitals

N_a = Number of electrons in anti-bonding molecular orbitals.

Bond Order of N_2:

Total no. of electrons = 14

Electronic configuration = $\sigma^2_{1s}, \sigma^{*2}_{1s}, \sigma^2_{2s}, \sigma^{*2}_{2s}, \pi^2_{2py}, \pi^2_{2pz}, \sigma^2_{2px}$

Bond order $= \frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

Bond Order of N_2^+ :

Total no. of electrons = 13

Electronic configuration = $\sigma^2_{1s}, \sigma^{*2}_{1s}, \sigma^2_{2s}, \sigma^{*2}_{2s}, \pi^2_{2py} = \pi^2_{2pz}, \sigma^2_{px}{}^1$

Bond order $= \frac{N_b - N_a}{2} = \frac{9-4}{2} = 2.5$

Bond Order of N_2^- :

Total no. of electrons = 15

Electronic configuration = $\sigma^2_{1s}, \sigma^{*2}_{1s}, \sigma^2_{2s}, \sigma^{*2}_{2s}, \pi^2_{2py} = \pi^2_{2pz}, \sigma^2_{2px}\, \sigma^2_{py}{}^1$

Bond order $= \frac{N_b - N_a}{2} = \frac{10-5}{2} = 2.5$

Stability of the molecule is directly proportional to bond order.

Hence, $N_2 > N_2^+ = N_2^-$

54. Explain the covalent character in ionic bond.

- The partial covalent character in ionic compounds can be explained on the basis of a phenomenon called polarisation.
- We know that in an ionic compound, there is an electrostatic attractive force between the cation and anion.
- The **positively charged cation** attracts the valence electrons of anion while repelling the nucleus.
- This causes a **distortion** in the electron cloud of the anion and its electron density drifts towards the cation, which results in some sharing of the valence electrons between these ions.

- Thus, a **partial covalent character** is developed between them. This phenomenon is called polarisation.
- **Example:** LiCl shows covalent character and is soluble in ethanol.

55. Describe Fajan's rule.

The ability of a cation to polarize an anion is called its polarising ability and the tendency of an anion to get polarized is called its polarisability. The extent of polarisation in an ionic compound is given by the Fajans rules

Fajans Rules:

i. To show greater covalent character, both the cation and anion should have high charge on them.

Higher the positive charge on the cation, greater will be the attraction on the electron cloud of the anion.

Similarly higher the magnitude of negative charge on the anion, greater is its polarisability.

Hence, ***the increase in charge on cation or in anion increases the covalent character.***

Let us consider three ionic compounds aluminum chloride, magnesium chloride and sodium chloride.

Since the charge of the cation increase in the order $Na^+ < Mg^{2+} < Al^{3+}$, the covalent character also follows the same order $NaCl < MgCl_2 < AlCl_3$.

ii. The ***smaller cation and larger anion show greater covalent character due to the greater extent of polarisation.***

Lithium chloride is more covalent than sodium chloride. (The size of Li^+ is smaller than Na^+ and hence the polarising power of Li^+ is more. Lithium iodide is more covalent than lithium chloride as the size of I^- is larger than the Cl^-). Hence I^- will be more polarized than Cl^- by the cation, Li^+.

iii. Cations having $ns^2\ np^6\ nd^{10}$ configuration show ***greater polarizing power*** than the cations with $ns^2\ np^6$configuration. Hence, they show ***greater covalent character.***

CuCl is more covalent than NaCl. Compared to Na^+ (1.13 Å) . Cu^+ (0.6 Å) is small and have $3s^2\ 3p^6\ 3d^{10}$ configuration.

Electronic configuration of Cu^+

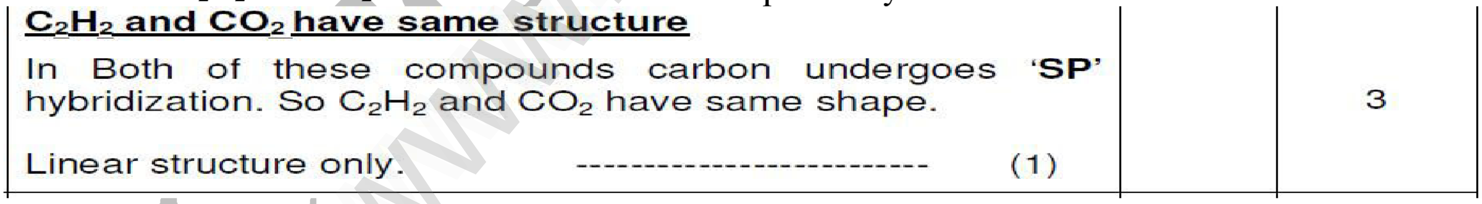
$[Ar]3s^2, 3p^6, 3d^{10}$

Electronic Configuration of Na^+

$[He]2s^2, 2p^6$

1. State Octet rule (69) , What is covalent bond? (69)

2. Draw the Lewis dot structures for sulphur trioxide. (71) **(MQ19)** Study Lewis dot structure of H_2O, NH_3, CH_4, N_2O_5, HNO_2, H_3PO_4, NO_3^-, SO_4^{2-}, HNO_3, O_3. **M23 2M (water, nitric acid)**
3. How will you find formal charge of an atom? (72)
4. What is ionic or electrovalent bond? (74)
5. Explain the ionic bond formation in KCl, MgO and CaF_2. (75)
6. What is coordinate covalent bond? (75)
7. What is bond order? (76) **S21 5Mi,** Define bond length? (76) **M22 5Mi**
8. What is bond length? Name the techniques through which the length of a bond can be determined. (76) **M19 2M**
9. Define bond angle. **(77) M22 5Mii**
10. Define bond enthalpy (77) **M22 5Miii**
11. What is resonance? (78) **S21 3M**
12. Explain the resonance structure of CO_3^{2-} (78)
13. What is dipole moment? (79) **J24 5Mii**
14. What is polar covalent bond? Give an example (79)
15. CO_2 has zero dipole moment even though two polar bonds. Why? (80)
16. How will you find ionic character? (80)
17. What is polarization? (80)
18. Linear form of carbon dioxide molecule has two polar bonds. Yet the molecule has zero dipole moment. Why? (80) **J19 2M**
19. How will you determine the ionic character in covalent bond using electronegativity values? (80) **M22 5M**
20. State Fajan's rule (81) **MQ19, J22 3M, J23 3M**
21. What are the important principles of VSEPR theory? (81)
22. Write the shape and molecular geometry for BF_3. (82) **S20 2M**
23. Give the shapes of molecules predicated by VSEPR theory (82) **S21 3M** a)$BeCl_2$ b) NH_3 c) H_2O
24. Predict the shape of $ClF_3 and NH_3$ using VSEPR theory. (83)**(MQ19)**
25. Write the structure of the following compounds. (84) **M19 5Mii** a) NH_3 (B) BF_3 **(S20 2M)**
26. Mention the shape of following molecule base on VSEPR theory. (82-84) **J22 5M**
    (i)BF_3 (ii)BrF_3 (iii)PCl_5 (iv)SF_6 (v)IF_7
27. Both C_2H_2 and CO_2 have the same structure. Explain why?. **M19 3MARK ANS:**



C_2H_2 and CO_2 have same structure In Both of these compounds carbon undergoes 'SP' hybridization. So C_2H_2 and CO_2 have same shape. Linear structure only. ------------------------- (1)		3

28. Explain the salient features of VB theory (86) **S21 5Mii**
29. Define sigma, Pi bond? (87) **J23 3M COMPULSORY, M24 3M(Pi bond)**
30. Explain the formation of H_2 **S21 5M**, F_2, HF **(M22 3M),** O_2 molecule by overlapping of orbitals (87-89)
31. Draw the hybridization in $BeCl_2$ (90)
32. What is hybridization? (89) Mention the type of hybridisation found in CH_4.(92) **S20 3M**
33. Write the hybridisation present in the following compounds. **M24 2M**
    a)BF_3 b)CH_4 c) PCl_5 d) SF_6
34. Draw the hybridization in BF_3 (91), CH_4(92), PCl_5 (93), SF_6 (94), ethylene molecule(95), acetylene molecule (96)
35. Explain the salient features of Molecular orbital Theory (99) **S20 5Mii, M23 5M**

36. Explain the molecular formation of the following by MOT i) H_2 ii) Li_2 iii) B_2 iv) C_2 **J22 5M** (99) v) N_2 **J20 5Mi , J23 5M, M24 5M** vi) O_2 **J24 3M COMPULSORY** vii) CO viii) NO (100)

10. Chemical Bonding

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
8	8	9	14	14	14	8	12	11	6				

37. State Octet rule (69)
38. What is covalent bond? (61)
39. Draw the Lewis dot structures for sulphur trioxide. (71) **(MQ19)** Study Lewis dot structure of H_2O, NH_3, CH_4, N_2O_5, HNO_2, H_3PO_4, NO_3^-, SO_4^{2-}, HNO_3, O_3. **M23 2M (water, nitric acid)**
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66. Both C_2H_2 and CO_2 have the same structure. Explain why?. **M19 3MARK**

ANS:

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IMPORTANT QUESTION 10. Chemical Bonding

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102. Predict the shape of $ClF_3 and NH_3$ using VSEPR theory. (83)**(MQ19)**

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ANS:

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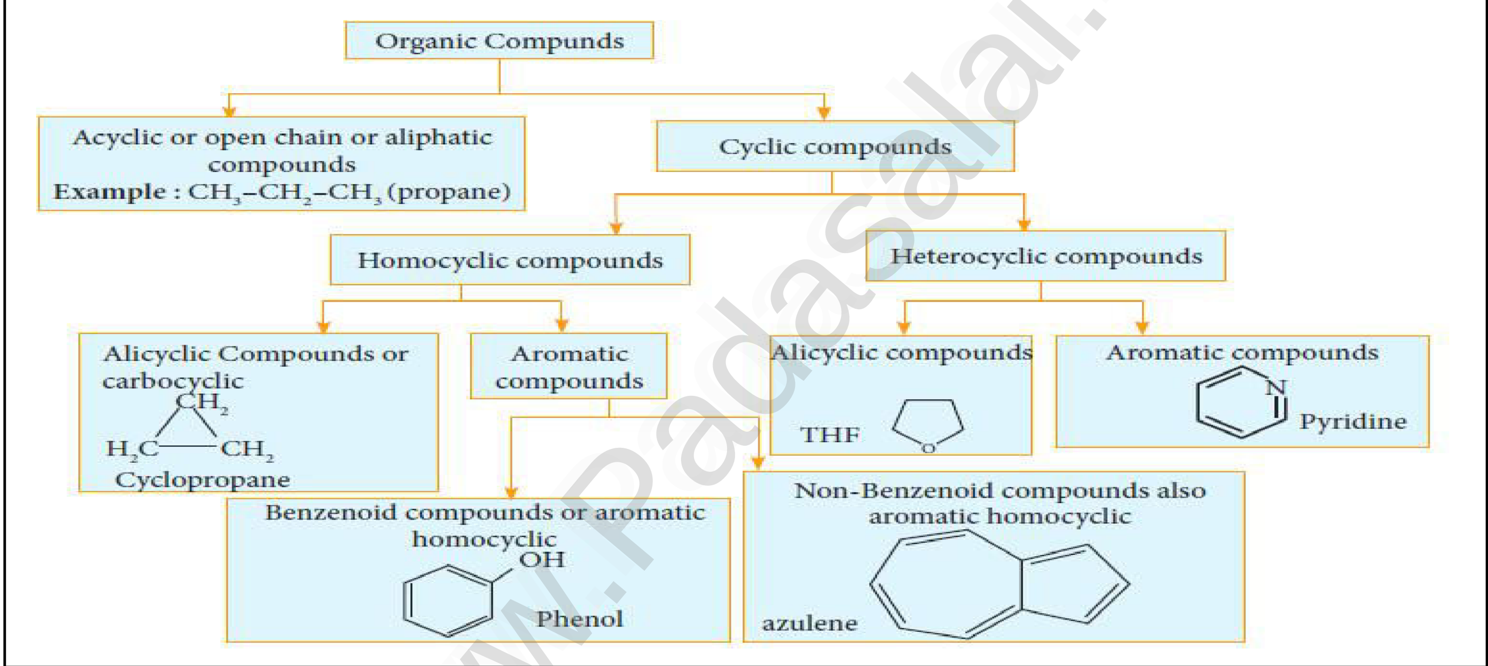
Unit 11: Fundamentals of Organic Chemistry

31.Give the general characteristics of organic compounds?

- They are **covalent compounds** of carbon.
- Generally **insoluble in water.**
- Readily **soluble in organic solvent** such as benzene, toluene, ether, chloroform etc….
- They are highly **inflammable in nature.** (except CCl_4).
- They possess **low boiling and melting points** due to their covalent nature.
- They are characterized by **functional groups.**
- They **exhibit isomerism** which is a unique phenomenon.
- Organic compounds have a high molecular weight and a complex structure.
- They are **volatile** in nature.

32.Describe the classification of organic compounds based on their structure.

11.2.1 Classification based on the structure:



33.Write a note on homologous series.

- **Homologous series:** A series of organic compounds each containing a characteristic functional group and the successive members differ from each other in molecular formula by a **$-CH_2$ - group** is called homologous series.
- All members of homologous series contain the same elements and functional group.
- Compounds of the homologous series are represented by a general formula.
  Alkanes C_nH_{2n+2}, Alkenes C_nH_{2n}, Alkynes C_nH_{2n-2}.
- They can be prepared by general methods.
- They show **regular gradation in physical properties** but have almost **similar chemical property.**
- **Example- Alkanes:** Methane (CH_4), Ethane (C_2H_6), Propane (C_3H_8) etc.

Alcohols: Methanol (CH_3OH), Ethanol (C_2H_5OH), Propanol (C_3H_7OH) etc,,

34.What is meant by a functional group? Identify the functional group in the following compounds. (a) acetaldehyde (b) oxalic acid (c) di methyl ether (d) methylamine

A functional group is an atom or a specific combination of bonded atoms that react in a characteristic way, irrespective of the organic molecule in which it present.

(a) **acetaldehyde**: (CH_3CHO) **functional group -CHO**

b) oxalic acid : HOOC-COOH **functional group -COOH**

c) di methyl ether: $CH_3–O—CH_3$ **functional group - O –**

d)methylamine : CH_3- NH_2 **functional group –NH_2**

35.Give the general formula for the following classes of organic compounds. (a) Aliphatic monohydric alcohol (b) Aliphatic ketones. (c) Aliphatic amines.

(a) Aliphatic monohydric alcohol : $C_nH_{2n+1}OH$ or $C_nH_{2n+2}O$

(b) Aliphatic ketones : $C_nH_{2n}O$

(c) Aliphatic amines: $C_nH_{2n+1}NH_2$

36.Write the molecular formula of the first six members of homologous series of nitro alkanes.

General formula $C_nH_{2n+1}NO_2$

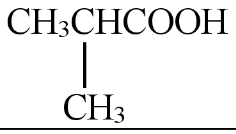
The first six members of nitro alkanes are

	Molecular formula	**Structural formula**	**Name**
1	CH_3NO_2	CH_3NO_2	Nitro methane
2	$C_2H_5NO_2$	$CH_3CH_2NO_2$	Nitro ethane
3	$C_3H_7NO_2$	$CH_3CH_2CH_2NO_2$	1-Nitro propane
4	$C_4H_9NO_2$	$CH_3CH_2CH_2CH_2NO_2$	1-Nitro butane
5	$C_5H_{11}NO_2$	$CH_3CH_2CH_2CH_2CH_2NO_2$	1-Nitro pentane
6	$C_6H_{13}NO_2$	$CH_3CH_2CH_2CH_2CH_2CH_2NO_2$	1-Nitro hextane

37.Write the molecular and possible structural formula of the first four members of homologous series of carboxylic acids.

General formula $C_nH_{2n}O_2$

The first four members of homologous series of carboxylic acids are

Molecular Formula	**Structural formula**	Name
CH_2O_2	H COOH	Formic acid
$C_2H_4O_2$	CH_3 COOH	Acetic acid
$C_3H_6O_2$	CH_3CH_2COOH	Propionic acid
$C_4H_8O_2$	$CH_3CH_2CH_2COOH$	1-Butanoic acid
	 $\underset{\displaystyle CH_3}{CH_3\underset{\vert}{C}HCOOH}$	2-Methyl-1-propanoic acid

38.Give the IUPAC names of the following compounds

1. $(CH_3)_2CH - CH_2 - CH(CH_3) - CH(CH_3)_2$

$$\overset{6}{C}H_3\text{—}\overset{5}{C}H(CH_3) - \overset{4}{C}H_2 - \overset{3}{C}H(CH_3) - \overset{2}{C}H(\overset{1}{C}H_3)CH_3$$

Ans : 2,3,5 - trimethyl hexane

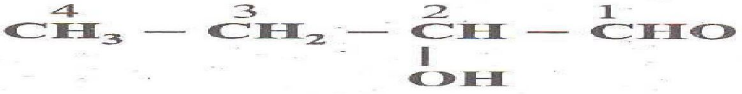
2. 

$$\overset{4}{C}H_3 - \underset{CH_3}{\overset{3}{C}H} - \underset{Br}{\overset{2}{C}H} - \overset{1}{C}H_3$$

Ans : 2-bromo 3-methyl butane

3. $CH_3 - O - CH_3$

Ans : methoxy methane

4. 

$$\overset{4}{C}H_3 - \overset{3}{C}H_2 - \underset{OH}{\overset{2}{C}H} - \overset{1}{C}HO$$

Ans : 2-hydroxy butanal

5. $CH_2 = CH - CH = CH_2$

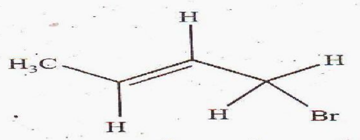
Ans : 1, 3 - butadiene

6. 

$$\overset{1}{C}H_3 - \overset{2}{C} \equiv \overset{3}{C} - \underset{Cl}{\overset{4}{C}H} - \overset{5}{C}H_3$$

Ans : 4-chloro pent - 2 - yne

7.



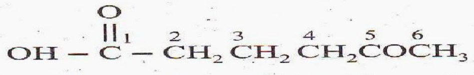
(or)

$\overset{4}{C}H_3 - \overset{3}{C}H = \overset{2}{C}H - \overset{1}{C}H_2Br$

Ans : 1 - bromo 2-butene

8.

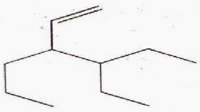
(or)



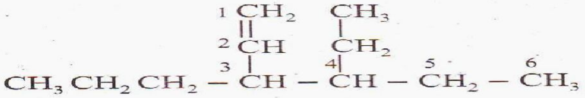
$$\begin{array}{ccccc} & & O & & \\ & & \|_1 & & \\ OH & - & C & - & \overset{2}{C}H_2\overset{3}{C}H_2\overset{4}{C}H_2\overset{5}{C}O\overset{6}{C}H_3 \end{array}$$

Ans : 5 - oxo hexanoic acid

9.



(or)

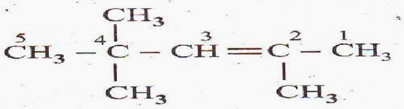


$$\begin{array}{ccccccc} & & {}^1CH_2 & & CH_3 & & \\ & & \| & & | & & \\ & & {}^2CH & & CH_2 & & \\ & & {}^3| & & {}^4| & & \\ CH_3CH_2CH_2 & - & CH & - & CH & - & \overset{5}{C}H_2 - \overset{6}{C}H_3 \end{array}$$

Ans : 3 - ethyl - 4- ethenyl heptane

10.

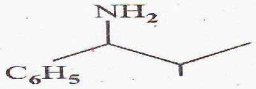
(or)



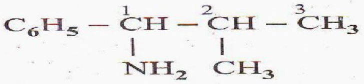
$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & {}^4| & & & & \\ \overset{5}{C}H_3 & - & C & - & \overset{3}{C}H = \overset{2}{C} & - & \overset{1}{C}H_3 \\ & & | & & | & & \\ & & CH_3 & & CH_3 & & \end{array}$$

Ans : 2,4,4 - trimethyl pent - 2 ene

11.



(or)

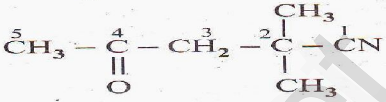


$$\begin{array}{ccccccc} C_6H_5 & - & \overset{1}{C}H & - & \overset{2}{C}H & - & \overset{3}{C}H_3 \\ & & | & & | & & \\ & & NH_2 & & CH_3 & & \end{array}$$

Ans : 2- methyl -1- phenyl - 1 - amino propane

12.

(or)



$$\begin{array}{ccccccccc} & & & & & & CH_3 & & \\ & & & & & & {}^2| & & \\ \overset{5}{C}H_3 & - & \overset{4}{C} & - & \overset{3}{C}H_2 & - & C & - & \overset{1}{C}N \\ & & \| & & & & | & & \\ & & O & & & & CH_3 & & \end{array}$$

Ans : 2, 2-dimethyl - 4 - oxo pentanenitrile

13.

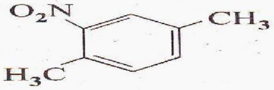
(or)

$$\begin{array}{ccc} CH_3 - CH_2 - O - & \overset{2}{C}H & - \overset{3}{C}H_3 \\ & {}^1| & \\ & CH_3 & \end{array}$$

Ans : 2- ethoxy propane

14.

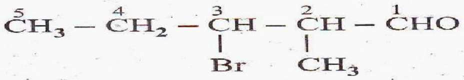
(or)



Ans : 2, 5 - dimethyl - 1-nitro benzene

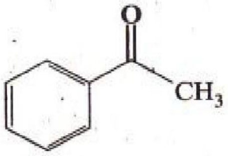
15.

(or)

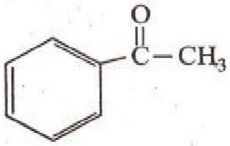


$$\begin{array}{ccccccccc} \overset{5}{C}H_3 & - & \overset{4}{C}H_2 & - & \overset{3}{C}H & - & \overset{2}{C}H & - & \overset{1}{C}HO \\ & & & & | & & | & & \\ & & & & Br & & CH_3 & & \end{array}$$

Ans : 3 - bromo 2 - methyl pentanal

16. 

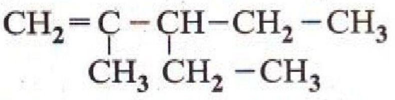
(or)



Ans : 1-phenyl ethanone

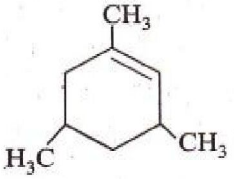
9. **Give the structure for the following compound.**

1) **3 - ethyl - 2 - methyl - 1- pentene**

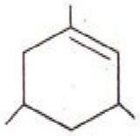


$CH_2=\underset{CH_3}{C}-\underset{CH_2-CH_3}{CH}-CH_2-CH_3$

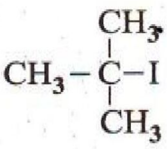
2) **1,3,5 - trimethyl cyclohex - 1 - ene**



(or)

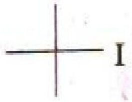


3) **Tertiary butyl iodide**

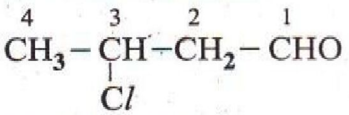


$CH_3-\overset{CH_3}{\underset{CH_3}{C}}-I$

(or)



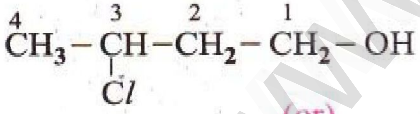
4) **3 - Chloro butanal**



$\overset{4}{CH_3}-\underset{Cl}{\overset{3}{CH}}-\overset{2}{CH_2}-\overset{1}{CHO}$

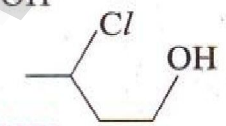
(or)

5) **3 - Chloro butanol**

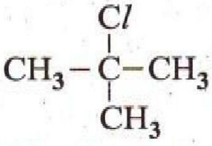


$\overset{4}{CH_3}-\underset{Cl}{\overset{3}{CH}}-\overset{2}{CH_2}-\overset{1}{CH_2}-OH$

(or)



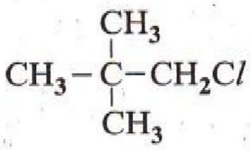
6) **2 - Chloro - 2 - methyl propane**



$CH_3-\overset{Cl}{\underset{CH_3}{C}}-CH_3$

(or)

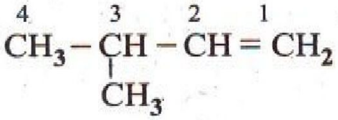
7) **2,2 - dimethyl - 1 - chloro propane**



$CH_3-\overset{CH_3}{\underset{CH_3}{C}}-CH_2Cl$

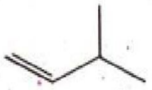
(or)

8) **3 - methylbut - 1 - ene**

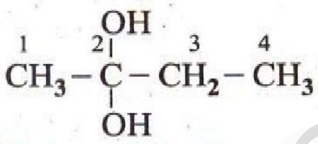


$\overset{4}{CH_3}-\underset{CH_3}{\overset{3}{CH}}-\overset{2}{CH}=\overset{1}{CH_2}$

(or)

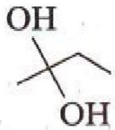


9) **Butan - 2,2 - diol**



$\overset{1}{CH_3}-\overset{OH}{\underset{OH}{\overset{2}{C}}}-\overset{3}{CH_2}-\overset{4}{CH_3}$

(or)

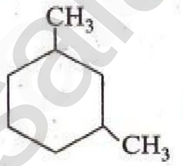


10) **Octane - 1,3 - diene**

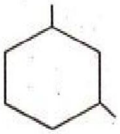
$\overset{1}{CH_2}=\overset{2}{C}=\overset{3}{CH}-\overset{4}{CH_2}\overset{5}{CH_2}\overset{6}{CH_2}\overset{7}{CH_2}\overset{8}{CH_3}$

(or)

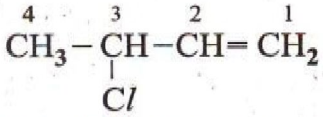
11) **1,3 - dimethyl cyclo hexane**



(or)

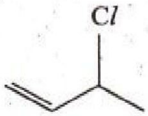


12) **2 - Chloro but - 1 - ene**

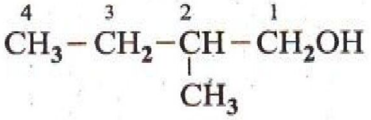


$\overset{4}{CH_3}-\underset{Cl}{\overset{3}{CH}}-\overset{2}{CH}=\overset{1}{CH_2}$

(or)

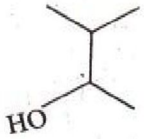


13) **2 - methyl butan - 1 - ol**



$\overset{4}{CH_3}-\overset{3}{CH_2}-\underset{CH_3}{\overset{2}{CH}}-\overset{1}{CH_2OH}$

(or)



14) **Acetaldehyde :** $CH_3\,CHO$ (or)

10. **Describe the reactions involved in the detection of nitrogen in an organic compound by Lassaigne's method.** **(Mar-19)**

If nitrogen is present in hydrocarbon it gets converted to sodium cyanide which reacts with **ferrous sulphate** and **ferric ion** followed by conc.HCl and gives a **Prussian blue color** (or) **green color** (or) precipitate.

40) Describe the reactions involved in the detection of nitrogen in an organic compound by Lassaigne method.

$Na + C + N \rightarrow NaCN$

From organic compounds

$FeSO_4 + 2NaOH \rightarrow Fe(OH)_2 + Na_2SO_4$

(from excess of sodium)

$6NaCN + Fe(OH)_2 \rightarrow Na_4[Fe(CN)_6] + 2NaOH$

Sod.ferrocyanide

$3Na_4[Fe(CN)_6] + FeCl_3 \rightarrow Fe_4[Fe(CN)_6]_3 + 12NaCl$

Ferric ferrocyanide Prussian
blue or green ppt

41) Give the principle involved in the estimation of halogen in an organic compound by carius method.

A known mass of the organic compound is heated with fuming HNO_3 and $AgNO_3$. C, H & S get oxidized to CO_2, H_2O & SO_2 and halogen combines with $AgNO_3$ to form a precipitate of silver halide.

$$X \xrightarrow[AgNO_3]{fum.HNO_3} AgX \downarrow$$

The precipitate of AgX is filtered, washed, dried and weighed. From the mass of AgX and the mass of the organic compound taken, percentage of halogens are calculated.

42) Give a brief description of the principles of

i) Fractional distillation ii)Column Chromatography

i) Fractional distillation:

- This is one method to purify and separate liquids present in the mixture having their boiling point close to each other.
- In the fractional distillation, a fractionating column is fitted with distillation flask and a condenser.
- A thermometer is fitted in the fractionating column near the mouth of the condenser.
- This will enable to record the temperature of vapour passing over the condenser.
- ***The process of separation of the components in a liquid mixture at their respective boiling points in the form of vapours and the subsequent condensation of those vapours is called fractional distillation***. The process of fractional distillation is repeated.

ii) Column chromatography:

- This is the simplest chromatographic method carried out in long glass column having a stop cock near the lower end.
- This method involves separation of a mixture over a column of adsorbent (Stationery phase) packed in a column.
- In the column a plug of cotton or glass wool is placed at the lower end of the column to support the adsorbent powder.

- The tube is uniformly packed with suitable absorbent constitute the stationary phase. (Activated aluminum oxides (alumina), Magnesium oxide, starch are also used as absorbents). top of the adsorbent column.
- Eluent which is a liquid or a mixture of liquids is allowed to flow down the column slowly.
- Different components depending upon the degree to which the components are adsorbed and complete separation takes place.
- The most readily adsorbed substances are retained near the top and others come down to various distances in the column.

43) Explain paper chromatography

Partition chromatography:

- Paper chromatography (PC) is an example of partition chromatography.
- The same procedure is followed as in thin layer chromatography except that a strip of paper acts as an adsorbent.
- This method involves continues differential portioning of components of a mixture between stationary and mobile phase.
- In paper chromatography, a special quality paper known as chromatography paper is used. This paper act as a stationary phase.
- A strip of chromatographic paper spotted at the base with the solution of the mixture is suspended in a suitable solvent which act as the mobile phase.
- The solvent rises up and flows over the spot. The paper selectively retains different components according to their different partition in the two phases where a chromatogram is developed.
- The spots of the separated colored compounds are visible at different heights from the position of initial spots on the chromatogram.
- The spots of the separated colorless compounds may be observed either under ultraviolent light or by the use of an appropriate spray reagent.

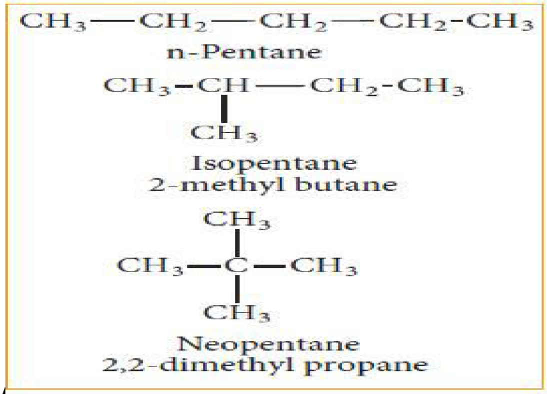
44. Explain various types of constitutional isomerism (structural isomerism) in organic compounds.

Isomerism: Isomers are compounds having **same molecular formula** but differ in physical or chemical or both physical and chemical properties.

Structural isomerism: Isomers have same molecular formula but differ in their bonding sequence. Structural or constitutional isomerism is further classified into following types.

(a) Chain or nuclear or skeletal isomerism :

These isomers differ in the way in which the carbon atoms are bonded to each other in a carbon chain or in other words **isomers have similar molecular formula but differ in the nature of the carbon skeleton** (ie. Straight or branched)

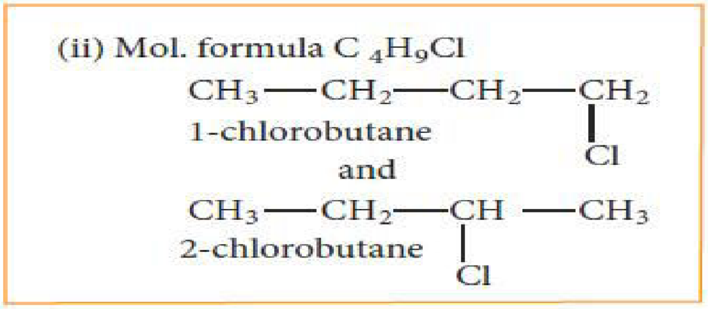


$CH_3 - CH_2 - CH_2 - CH_2 - CH_3$ n-Pentane

$CH_3 - CH(CH_3) - CH_2 - CH_3$ 2- methyl butane

b) Position isomerism: If different compounds belonging to same molecular formula and carbon skeleton, but **differ in the position of substitutent or functional group or an substitutent** or functional group or an unsaturated linkage are said exhibit position isomerism.

Example: $CH_3 - CH_2 - CH_2 - CH_2 = CH_2$ Pent-1-ene **and** $CH_3 - CH_2 - CH = CH - CH_3$ Pent-2-ene


(ii) Mol. formula C_4H_9Cl

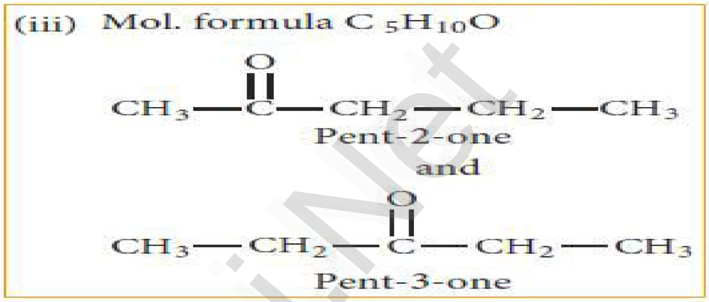
$CH_3-CH_2-CH_2-\underset{\displaystyle Cl}{\underset{|}{CH_2}}$

1-chlorobutane

and

$CH_3-CH_2-\underset{\displaystyle Cl}{\underset{|}{CH}}-CH_3$

2-chlorobutane


(iii) Mol. formula $C_5H_{10}O$

$CH_3-\overset{\displaystyle O}{\overset{\|}{C}}-CH_2-CH_2-CH_3$

Pent-2-one

and

$CH_3-CH_2-\overset{\displaystyle O}{\overset{\|}{C}}-CH_2-CH_3$

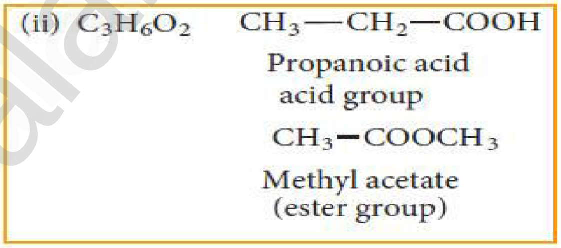
Pent-3-one

c) Functional isomerism:

Different compounds having **same molecular formula but different functional groups** are said to exhibit functional isomerism.

Example: $CH_3 - CH_2 - CHO$ Propanal **and**

$CH_3 - CO - CH_3$ (Acetone)


(ii) $C_3H_6O_2$ CH_3-CH_2-COOH

Propanoic acid

acid group

$CH_3-COOCH_3$

Methyl acetate

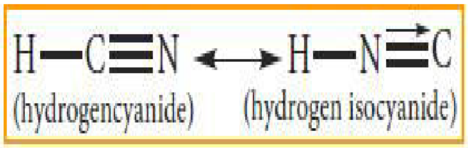
(ester group)

d) **Metamerism:** This type of isomerism is a special kind of structural isomerism arises due to the unequal distribution of carbon atoms on either side of the functional group or **different alkyl groups attached to the either side of the same functional group and having same molecular formula.**

$CH_3 - O - CH_2 - CH_2 - CH_3$ Methyl propyl ether

$CH_3 - CH_2 - O - CH_2 - CH_3$ Diethyl ether

(e) Tautomerism: It is a special type of functional isomerism in which a single compound exists in two readily inter convertible structures that differ markedly in the relative position of atleast one atomic nucleus, generally hydrogen. The different structures are known as tautomerism.


$H-C\equiv N \longleftrightarrow H-N\overset{\rightarrow}{\equiv}C$

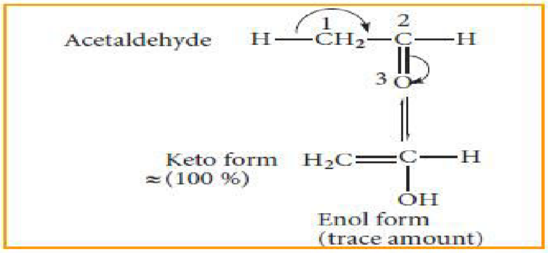
(hydrogencyanide) (hydrogen isocyanide)

Dyad system: In this system hydrogen atom oscillates between two directly linked polyvalent atoms.

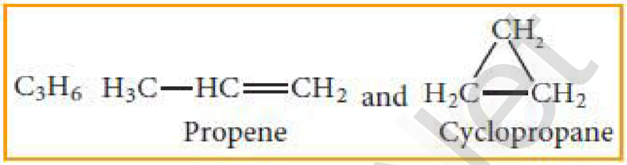
Triad system: In this system hydrogen atom oscillates between three polyvalent atoms.

It involves 1,3 migration of hydrogen atom from one polyvalent atom to other within the molecule.

Example:



f) Ring chain isomerism: In this type of isomerism, compounds having same molecular formula but differ in terms of bonding of carbon atom to form open chain and cyclic structures for eg:



45. Describe optical isomerism with suitable example.

Compounds having same physical and chemical property but differ only in the rotation of plane of the polarized light are known as optical isomers and the phenomenon is known as optical isomerism.

Example: Glucose, Lactic acid, Tartaric acid

d-Lactic acid: It rotates of the plane polarised light to the right or in clockwise direction is said to be dextrorotary (dexter means right) denoted by the sign (+).

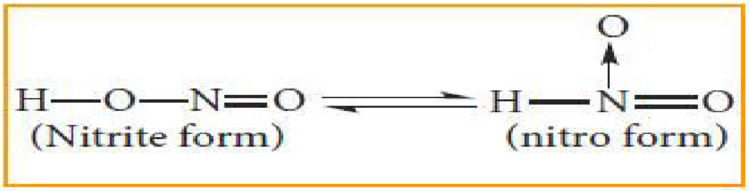
Dextrorotatory compounds are represented as 'd' or by sign (+).

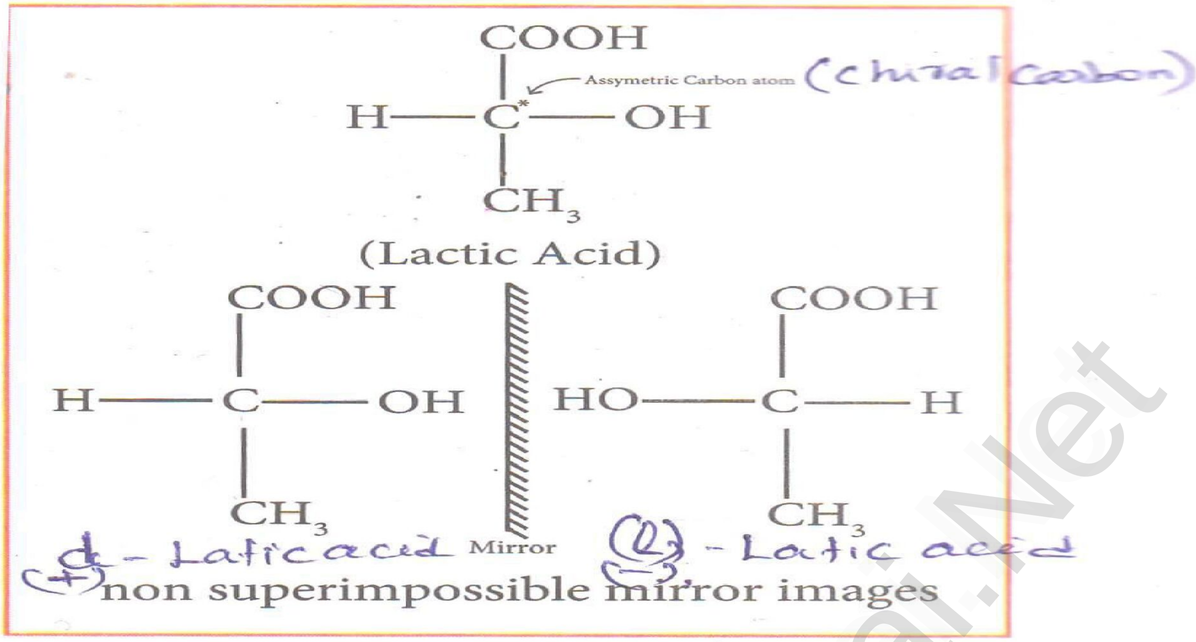
l-Lactic acid: It rotates of the plane polarised light to the left or anticlockwise is said to be laevorotatory (laevues means left) denoted by sign(-).

Laevorotatory compounds are represented as 'l' or by sign (-).

The optical isomers which are non-superimposable mirror images of each other are called **entantiomers.**

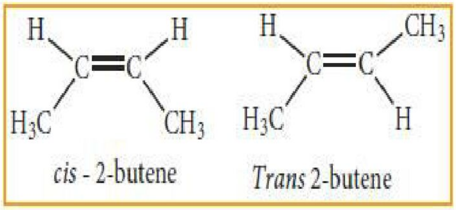
Nitro-aci tautomerism.





46. Briefly explain geometrical isomerism in alkene by considering 2- butene as an example.

- *2-butene Geometrical isomerism: CH_3-CH=CH-CH_3*



cis - 2-butene *Trans* 2-butene

- Geometrical isomers are the stereoisomers which have different arrangement of groups or atoms around a rigid frame work of double bonds.
- This type of isomerism occurs due to restricted rotation of double bonds, or about single bonds in cyclic compounds.
- In alkenes, the carbon-carbon double bond is sp^2 hybridized. The carbon-carbon double bond consists of a σ bond and a π bond.
- The cis isomer is one in which **two similar groups are on the same side** of the double bond.
- The trans isomers is that in which the **two similar groups are on the opposite side** of the double bond, hence this type of isomerism is often called cis-trans isomerism.

47) 0.30 g of a substance gives 0.88 g of carbon dioxide and 0.54 g of water calculate the percentage of carbon and hydrogen in it.

Weight of organic compound = 0.30g
Weight of Carbon dioxide = 0.88g
Weight of water = 0.54g

Percentage of Carbon:

44g if carbondioxide contains, carbon = 12g

0.88g of carbon dioxide contains,

Carbon = $\frac{12 \times 0.88}{44}$ g

0.30g substance contains, carbon = $\frac{12 \times 0.88}{44}$ g

100g substance contains = $\frac{12 \times 0.88}{44} \times \frac{100}{0.30}$ = 80g of carbon

Percentage of carbon = 80%

Percentage of hydrogen :

18g of water contains, hydrogen = 2g

0.54 g of water contains = hydrogen = $\frac{12 \times 0.54}{18}$ g

0.30 g of substance contains hydrogen = $\frac{12 \times 0.54}{18 \times 0.30}$ g

100g of substance contains = $\frac{12 \times 0.54}{18 \times 0.30} \times 100$ g

= 20g of hydrogen

Percentage of hydrogen =20%

48) The ammonia evolved form 0.20 g of an organic compound by kjeldahl method neutralised 15ml of N/20 sulphuric acid solution. Calculate the percentage of Nitrogen.

Weight of organic compound = 0.20g

Normality of acid = N/20

Volume of standard acid neutralized by ammonia = 15 ml

1000 ml of 1N acid = 14g of nitrogen

15 ml of acid of normality $\frac{N}{20}$ contains nitrogen = $\frac{14 \times 15 \times 1}{1000 \times 20}$

0.20 g of compound contains nitrogen = $\frac{14 \times 15}{1000 \times 20 \times 0.20}$

100 g of compound contains nitrogen = $\frac{14 \times 15 \times 100}{1000 \times 0.30 \times 0.20}$

= 5.25g

Percentage of nitrogen = 5.25%

49) 0.32 g of an organic compound, after heating with fuming nitric acid and barium nitrate crystals is a sealed tube game 0.466 g of barium sulphate. Determine the percentage of sulphur in the compound.

Mass of solution substance taken = 0.32g

Mass of $BaSO_4$ formed = 0.466g

Molecular mass of $BaSO_4$ = 137 + 32 + 64 = 233

Then mass of 0.466 g of $BaSO_4$ = $\frac{0.466+32}{233}$ g

Percentage of S in compound = $\frac{0.466+32 \times 100}{233 \times 0.32}$ = 20%

50) 0.24g of an organic compound gave 0.287 g of silver chloride in the carius method. Calculate the percentage of chlorine in the compound.

Weight of the organic substance = 0.24 g

Weight of AgCl is = 0.287 g

143.5 g of AgCl contains 35.5 g of chlorine

0.287 g of AgCl contains $\frac{35.5}{143.5} \times \frac{0.287}{0.284}$

% of chlorine is $\frac{35.5}{143.5} \times \frac{0.287}{0.284} \times 100$ = **24.56%**

51) In the estimation of nitrogen present in an organic compound by Dumas method 0.35 g yielded 20.7 mL of nitrogen at 15^0 C and 760 mm pressure. Calculate the percentage of nitrogen in the compound.

Weight of the organic compound = 0.359 Volume = 20.7 ml

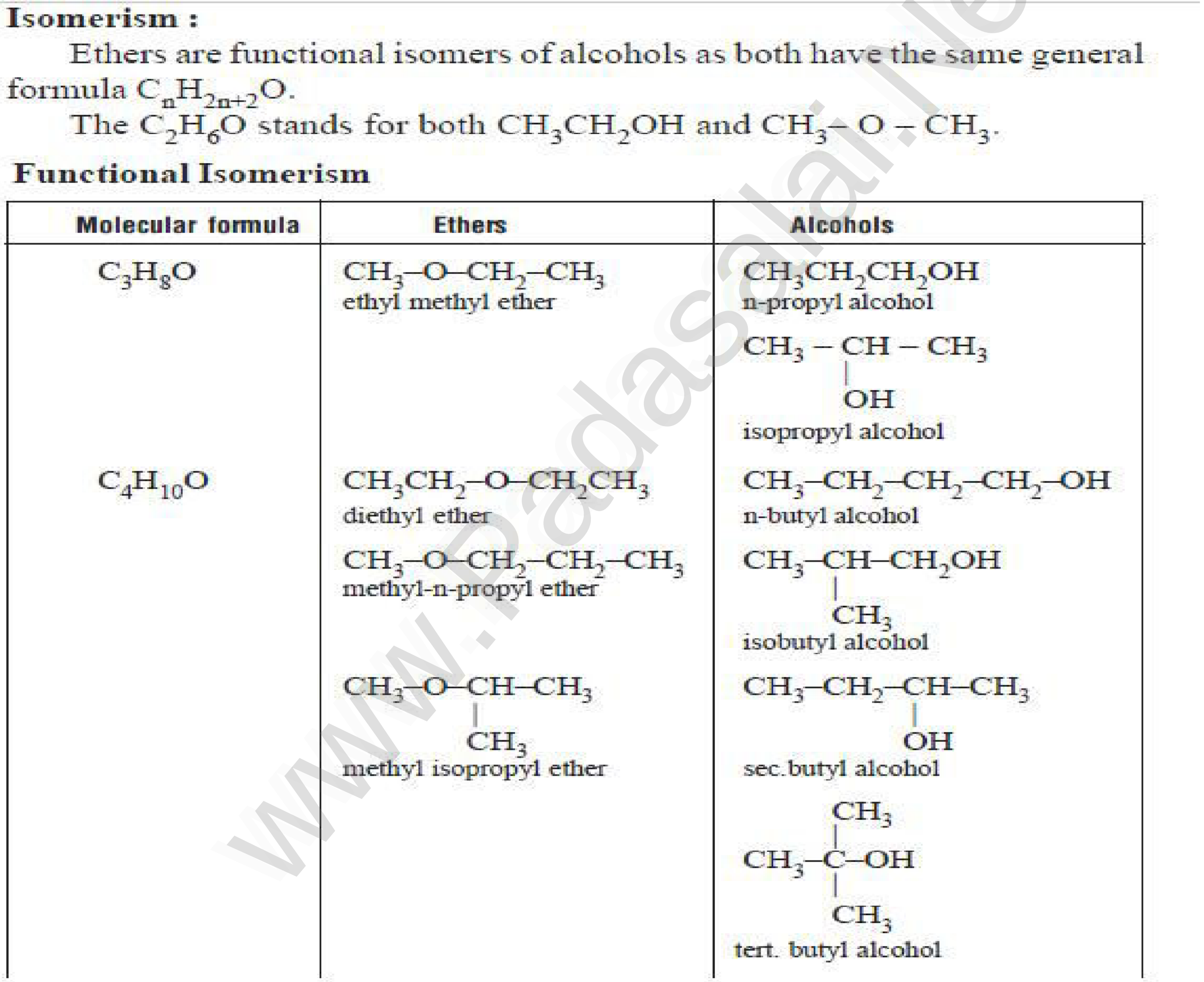
Temperature T = 15 + 273 = 288K

Pressure P = 760 mm

Percentage of N = $\frac{28 \times v}{22400} \times \frac{100}{w}$ = $\frac{28 \times 20.7}{22400} \times \frac{100}{0.359}$ = 7.208 Percentage of N = **7.208%**

Evaluate yourself:

4) Write all the possible isomers of molecular formula $C_4H_{10}O$ and identify the isomerisms found in them.



Isomerism :

Ethers are functional isomers of alcohols as both have the same general formula $C_nH_{2n+2}O$.

The C_2H_6O stands for both CH_3CH_2OH and CH_3-O-CH_3.

Functional Isomerism

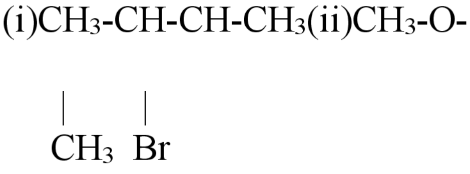
Molecular formula	Ethers	Alcohols
C_3H_8O	$CH_3-O-CH_2-CH_3$ ethyl methyl ether	$CH_3CH_2CH_2OH$ n-propyl alcohol $CH_3-\underset{\substack{\vert\\ OH}}{CH}-CH_3$ isopropyl alcohol
$C_4H_{10}O$	$CH_3CH_2-O-CH_2CH_3$ diethyl ether $CH_3-O-CH_2-CH_2-CH_3$ methyl-n-propyl ether $CH_3-O-\underset{\substack{\vert\\ CH_3}}{CH}-CH_3$ methyl isopropyl ether	$CH_3-CH_2-CH_2-CH_2-OH$ n-butyl alcohol $CH_3-\underset{\substack{\vert\\ CH_3}}{CH}-CH_2OH$ isobutyl alcohol $CH_3-CH_2-\underset{\substack{\vert\\ OH}}{CH}-CH_3$ sec.butyl alcohol $CH_3-\overset{\substack{CH_3\\ \vert}}{\underset{\substack{\vert\\ CH_3}}{C}}-OH$ tert. butyl alcohol

LESSON 11. FUNDAMENTALS OF ORGANIC CHEMISTRY

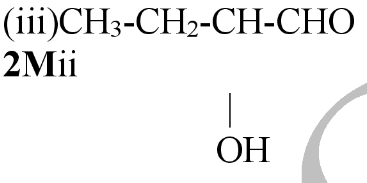
M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
11	10	13	10	9	10	9	8	6	8				

1. Which element exhibits maximum catenation and why? 110 **S20 2M**

2. What are characteristics of organic compounds?(111) **M23 5Mi**
3. What is meant by homologous series. (112) **S21 5Mi, M24 2M**
4. Gives the general formula for the following class of organic compounds. (112) **S21 3M**
   a)Alkanes b) alkenes c) alkynes
5. How organic compounds are classified. Based on structure (112) **J23 5M**
6. Give an example for each of the following type of organic compounds. (112) **J22 3M COMPUL**
   (i) Non benzenoid aromatic compound.
   (ii) Aromatic heterocyclic compound
   (iii) Carbocyclic compound.
7. How organic compounds are classified. Based on functional groups(112)
8. Find the functional group of the following compounds.BB **M23 5Mii**
   a) Acetaldehyde b) Oxalic acid c) Dimethyl ether d) methylamine
9. Give the structure for the following compounds: **J24 2M**
   i)1,3-dimethyl cyclohexane ii) 3-ethyl-2-methyl-1-pentene
10. Give the structural formula for the following compounds **S21 5Mii**
    a) 3-methyl pentane b)2-methylpropan-2-ol c) propanone
11. Give the IUPAC name for the following compounds. **M24 5Mi**
    a)CH_3CH_2COOH b)$CH_3CH_2COCH_2CH_3$ c) $CH_3CH_2CH_2N(CH_3)_2$
12. Give the IUPAC name of the following compounds **M22 5M** (i)

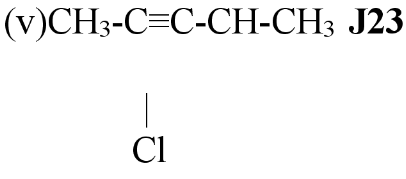


$CH_3-\underset{CH_3}{CH}-\underset{Br}{CH}-CH_3$ (ii)CH_3-O-CH_3



(iii) $CH_3-CH_2-\underset{OH}{CH}-CHO$

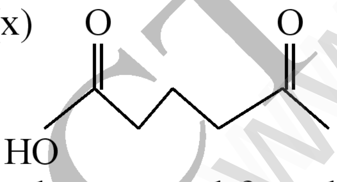
(iv)$CH_2=CH-CH=CH_2$ **J23 2Mi**

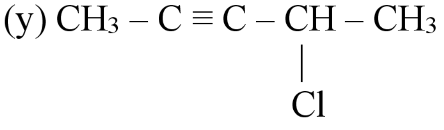


(v)$CH_3-C\equiv C-\underset{Cl}{CH}-CH_3$ **J23 2Mii**

13. Write the IUPAC names for the following compounds. **J19 5Mi**

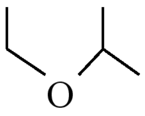
(x)





(y) $CH_3-C\equiv C-\underset{Cl}{CH}-CH_3$

(z)

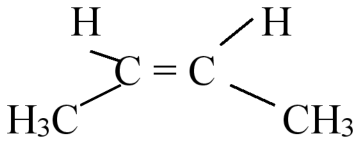


14. Give the structural formula for the following compounds. (124) **(M19 3MARK Compulsory, M23 3M compulsory**
    a) m – dinitrobenzene b) p – dichlorobenzene c) 1,3,5 Tri – methyl Benzene
15. Define isomerism. (132)
16. Describe any two types of constitutional isomers. (132) **J22 5M**
17. Define chain isomers. Give an example.(132)
18. Define position isomers. Give an example. (132)
19. Define functional isomers. Give an example. (133)

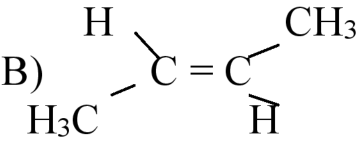
20. Define metamerism. Give an example. (133)
21. Define tautomerism. Give an example. (134)
22. Define Ring chain isomers. Give an example. (134)
23. Define stereoisomerism. (134)
24. Define geometrical isomerism. Draw 2-butene cis, trans isomerism. (135) **(S20) 3mark comp**

Identify the cis and trans isomer for the following compounds. (135) **M19 5Mi**

a)



(B)



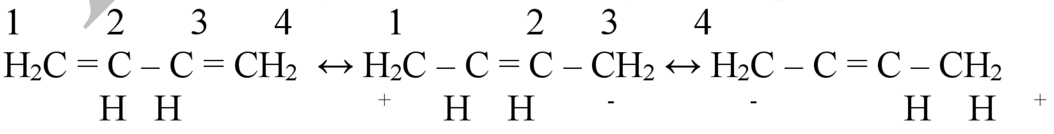
25. Draw Cis – Trans isomers for 2,3- dichloro -2-butene. (135) **(MQ19)**
26. Draw oximes and azo compounds cis, trans isomerism. (136)
27. Define optical isomerism. Give an example. (137) **M22 3M**
28. Explain optical isomerism of Lactic acid.(137)
29. Define enantiomerism.(137)
30. Describe the reactions involved in the detection of nitrogen in an organic compound by Lassaigne method. (138) **M19 2M**
31. Test for sulphur. (139)
32. Test for halogen. (139)
33. Test for phosphorus. (139)
34. Estimation of carbon and hydrogen. (140)
35. Give the principle involved in the estimation of halogen in an organic compound by carius method. (142) **J24 5M**
36. Estimation of phosphorus. (143)
37. Which is the suitable method for detection of nitrogen present in food and fertilizers? **(146) M19 2M COMPULSORY**
38. Explain the purification of solid organic compound by crystallization method. (148) **S20 5M**
39. Write note on column chromatography, thin layer chromatography, paper chromatography. (151)
40. Define retention factor. (152)

REVISE ALL Structure IUPAC NAME, problem & study well BOOK BACK QUESTION & ANSWER

Unit: 12 Basic Concepts of Organic reactions

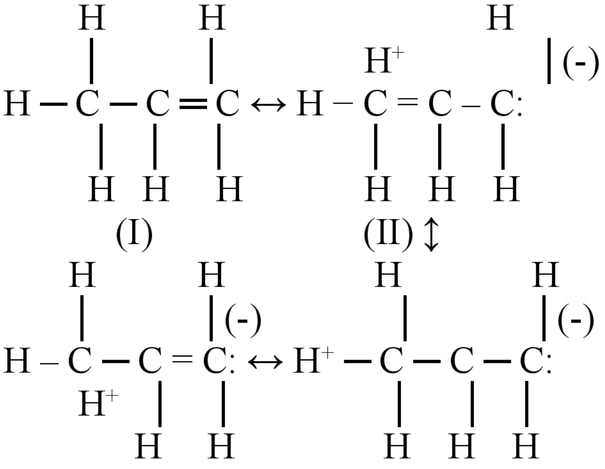
16. Write short notes on(a) Resonance (b) Hyperconjugation

(a) Resonance : Certain organic compounds can be represented by more than one structure and they dier only in the position of bonding and lone pair of electrons. Such structures are called resonance structures (canonical structures) and this phenomenon is called resonance.



$$\overset{1}{H_2C} = \underset{H}{\overset{2}{C}} - \underset{H}{\overset{3}{C}} = \overset{4}{CH_2} \leftrightarrow \overset{1}{H_2\overset{+}{C}} - \underset{H}{\overset{2}{C}} = \underset{H}{\overset{3}{C}} - \overset{4}{\overset{-}{C}H_2} \leftrightarrow H_2\overset{-}{C} - \underset{H}{C} = \underset{H}{C} - \overset{+}{C}H_2$$

(b) Hyper conjugation: The delocalisation of electrons of **σ** bond is called as hyper conjugation. It is a special stabilising effect that results due to the interaction of electrons of a **σ**-bond (usually **C-H** or **C-C**) with the adjacent, empty non-bonding p-orbital or an antibonding **σ*** or **π***-orbitals resulting in an extended molecular orbital. Unlike electromeric effect, hyper conjugation is a permanent effect. In propene, the *σ*-electrons of C-H bond of methyl group can be delocalized into the *π*-orbital of doubly bonded carbon as represented below.



$$H-C-C=C \leftrightarrow H^{+}\ H-C=C-C^{(-)}: \leftrightarrow H-C-C=C^{(-)}: \ H^{+} \leftrightarrow H^{+}-C-C-C^{(-)}:$$

17. What are electrophiles and nucleophiles? Give suitable examples for each.

Nucleophiles:

Phile means affinity: Nucleophile means nucleus loving species ie species which that are attracted towards positive charge. Nucleophiles are electron rich species. They may be neutral, or negatively charged. Neutral nucleophiles possess one or more lone pair of electrons. All lewis bases are nucleophiles.

Neutral Nucleophiles: NH_3, RNH_2, R_2NH and R_3N

Negatively Charged Nucleophiles: Cl^-, Br^-, CN^-

Electrophiles: Electron

Electrophiles means : electron loving species ie species which that are attracted towards negative charge. Electrophiles are electron deficient species. They may be neutral, or positively charged. Neutral electrophiles possess electron deficient centre. All Lewis acids are Electrophiles.

Neutral Electronphiles : $AlCl_3$, BF_3 and $FeCl_3$

Positively charged electrophiles: R^+, H^+, RX, H_3O^+, NO^+, etc.,

18. Show the heterolysis of covalent bond by using curved arrow notation and complete the following equations. Identify the nucleophile is each case.

(i) CH_3 - Br + KOH →

(ii) CH_3 - OCH_3 + HI →

(i) CH_3 - Br + KOH → CH_3 – OH + KBr

Methyl bromide Methanol

OH^- from KOH is the nuclephile.

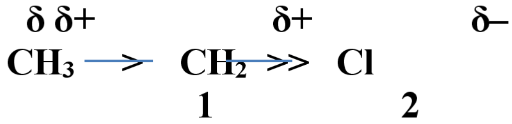
(ii) CH_3 – O - CH_3 + HI → CH_3I + CH_3OH

Dimethyl ether Methyl iodide Methanol

I^- from HI acts as a nucleophile

19. Explain inductive effect with suitable example.

- Inductive effect is defined as the change in the polarisation of a covalent bond due to the presence of adjacent bonds, atoms or groups in the molecule. This is a permanent phenomenon.
- Let us explain the inductive effect by considering ethane and ethyl chloride as examples. The C-C bond in ethane is non polar while the C-C bond in ethyl chloride is polar. We know that chlorine is more electronegative than carbon, and hence it attracts the shared pair of electron between C-Cl in ethyl chloride towards itself. This develops a slight negative charge on chlorine and a slight positive charge on carbon to which chlorine is attached. To compensate it, the C1 draws the shared pair of electron between itself and C2. This polarisation effect is called inductive effect.
- This effect is greatest for the adjacent bonds, but they also be felt farther away. However, the magnitude of the charge separation decreases rapidly, as we move away from C1 and is observed maximum for 2 carbons and almost insignificant after 4 bonds from the active group.

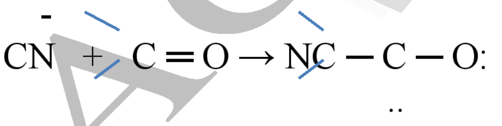


- It is important to note that the inductive effect does not transfer electrons from one atom to another but the displacement effect is permanent. The inductive effect represents the ability of a particular atom or a group to either withdraw or donate electron density to the attached carbon. Based on this ability the substituents are classified as +I groups and –I groups. Their ability to release or withdraw the electron through sigma covalent bond is called +I effect and -I effect respectively.

20. Explain electromeric effect.

Electromeric effect (E) : Electromeric is a temporary effect which operates in unsaturated compounds (containing >C=C<, >C=O, etc...) in the presence of an attacking reagent. Let us consider two different compounds

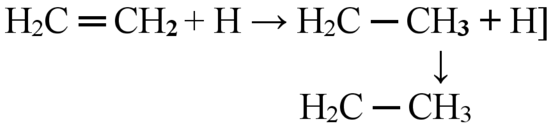
(i) compounds containing carbonyl group (>C=O) : When a nucleophile approaches the carbonyl compound, the **π** electrons between C and O is instantaneously shifted to the more electronegative oxygen. This makes the carbon electron deficient and thus facilitating the formation of a new bond between the incoming nucleophile and the carbonyl carbon atom.



(ii) unsaturated compounds such as alkenes (>C=C<) :

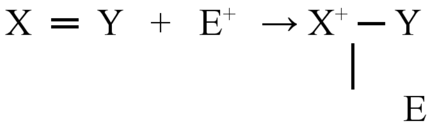
when an electrophile such as H+ approaches an alkene molecule, the **π electrons** are instantaneously shifted to the electrophile and a new bond is formed between carbon and hydrogen. This makes the other carbon electron deficient and hence it acquires a positive charge.

[δ+ δ- +



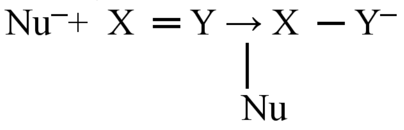
$$H_2C=CH_2 + H \rightarrow H_2C-CH_3 + H]$$
$$\downarrow$$
$$H_2C-CH_3$$

When the π electron is transferred towards the attacking reagent, it is called + E (positive electromeric) effect.



$$X = Y + E^+ \rightarrow \underset{\displaystyle E}{\underset{|}{X^+}} - Y$$

When the π electron is transferred away from the attacking reagent, it is called, -E (negative electromeric) effect

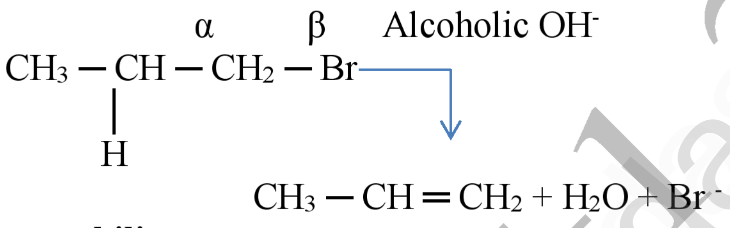


$$Nu^- + X = Y \rightarrow \underset{\displaystyle Nu}{\underset{|}{X}} - Y^-$$

21. Give examples for the following types of organic reactions

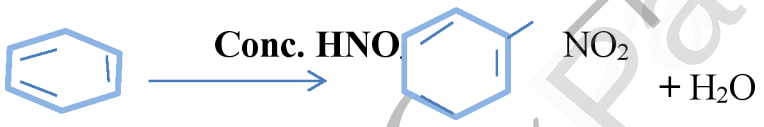
(i) β – elimination (ii) electrophilic

(i) β – elimination:



$$CH_3 - \underset{\displaystyle H}{\underset{|}{CH}} - \overset{\alpha}{CH_2} - \overset{\beta}{Br} \xrightarrow{\text{Alcoholic } OH^-} CH_3 - CH = CH_2 + H_2O + Br^-$$

(ii) electrophilic:



LESSON 12 Basic concepts of Organic reactions

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
7	7	2	4	5	4	4	1	6	3				

1. What are homolytic and heterolytic cleavages? (162)
2. What is the hybridization of carbon in carbocation? (163)
3. What are Nucleophiles and electrophiles? Give an example (164) **(M19 5Mii, M22 3M, M24 3M**
4. Explain Inductive effect in the organic compounds (166) **(J19)** M23 3M
5. Explain Electromeric effect (167) **(M19) J24 5Mi**
6. What is resonance. **(168) S21 3M**
7. Explain resonance or mesomeric effect and its types (168).
8. Explain with example the positive Mesomeric effect. (169) **M19 5Mii**

9. The bond length between all the four carbon atoms is same in 1, 3- butadiene. Explain with reason.(168) **J19 3MARK Compulsory**

10. Write no bond resonance structure shown by propene. (170) **(S20 2M)**

11. Explain hyper conjugation effect (170) **(MQ19, J22 3M, J24 3M**

12. Write ß-elimination reaction. (172) **M24 5Mii**

13. Identify the compound (A) and (B). (173) **M19 5Mii**

$$R - C \equiv N \xrightarrow{H_2O/H^+} (A) \xrightarrow{H_2O/H^+} B$$

& study well BOOK BACK QUESTION & ANSWER

Unit 13: Hydrocarbons

BOOK BACK

31. Give IUPAC names for the following compounds

1) $CH_3–CH=CH–CH=CH–C≡C–CH_3$ Octa-4,6-diene-3-yne

C_2H_5 CH_3

| |

2) $CH_3– C – C – C ≡ C – CH_3$

|

CH_3 4,5,5- trimethylhepta-2-yne

3) $(CH_3)_3 C – C ≡ C – CH (CH_3)_2$

2,2,5 – trimethyl hex-3-yne

|

CH_3

4) ethyl isopropyl acetylene

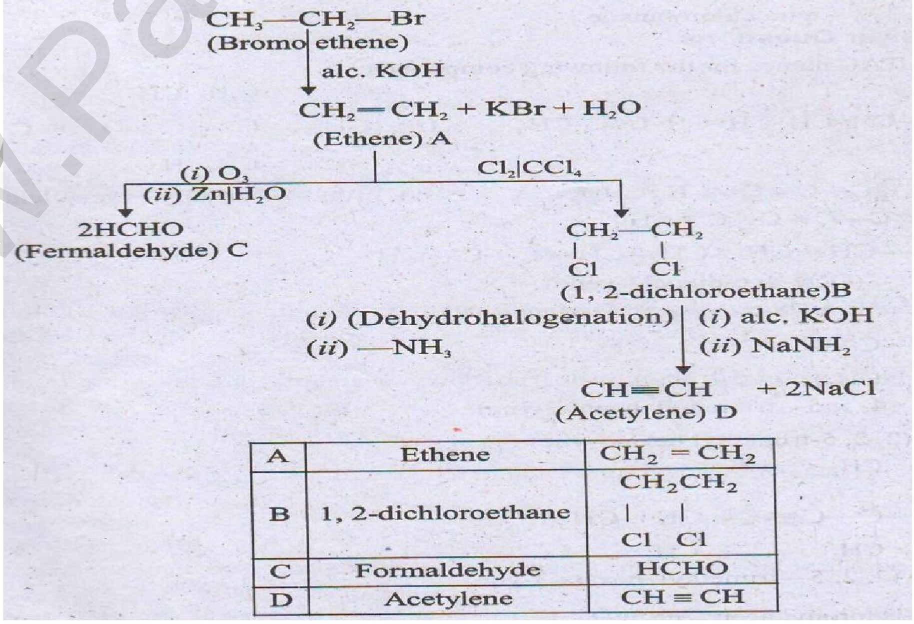
CH_3 CH_2 – C ≡ C – CH $(CH_3)_2$

2 – methyl hex-3-yne

5) CH ≡ C – C ≡ C – C ≡ CH

1,3,5- Hexa triyne

32. Identify the compound A, B, C and D in the following series of reactions



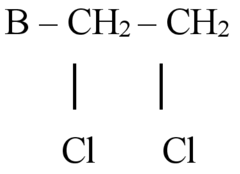
A	Ethene	$CH_2 = CH_2$
B	1, 2-dichloroethane	CH_2CH_2 | | Cl Cl
C	Formaldehyde	HCHO
D	Acetylene	CH ≡ CH

i) $CH_3 – CH_2–Br$

↓

$$\begin{tikzcd}
A \arrow[r, "\text{alc. KOH}"', "Cl_2 / CCl_4"] \arrow[d, "\text{i) } O_3 \quad \text{ii) } Zn/H_2O"'] & B \arrow[d, "NaNH_2"] \\
C & D
\end{tikzcd}$$

A – $CH_2 = CH_2$ – Ethene



B – $\underset{Cl}{CH_2} - \underset{Cl}{CH_2}$ - 1, 2 – dichloro ethane

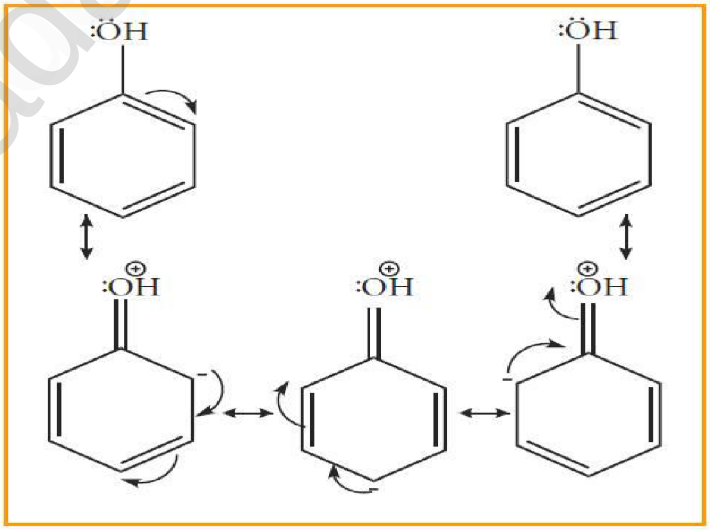
C – HCHO - Formaldehyde

D – $CH \equiv CH$ - Acetylene

33. write a short note on ortho-para directors in aromatic electrophilic substitution reactions?

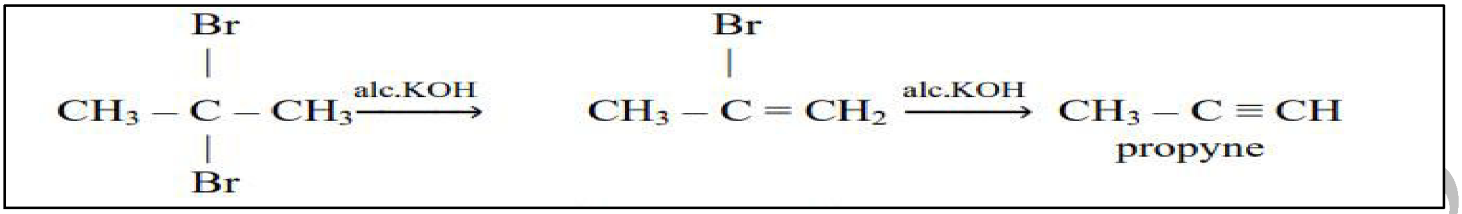
All the activating groups are 'orthopara' directors.

Example –OH, $-NH_2$, -NHR, $-NHCOCH_3$ $-OCH_3$, $-CH_3$ – C_2H_5 etc. Let us consider the directive influences of phenolic (-OH) group Phenol is the resonance hybrid of following structures.



In these resonance structures, the (-) charge residue is present on ortho and para position of ring structure. It is quite evident that the lone pair of electron on the atom which is attached to the ring involves in resonance and makes the ring more electron rich than benzene.

33. How is propyne prepared from an alkyenedihalide ?

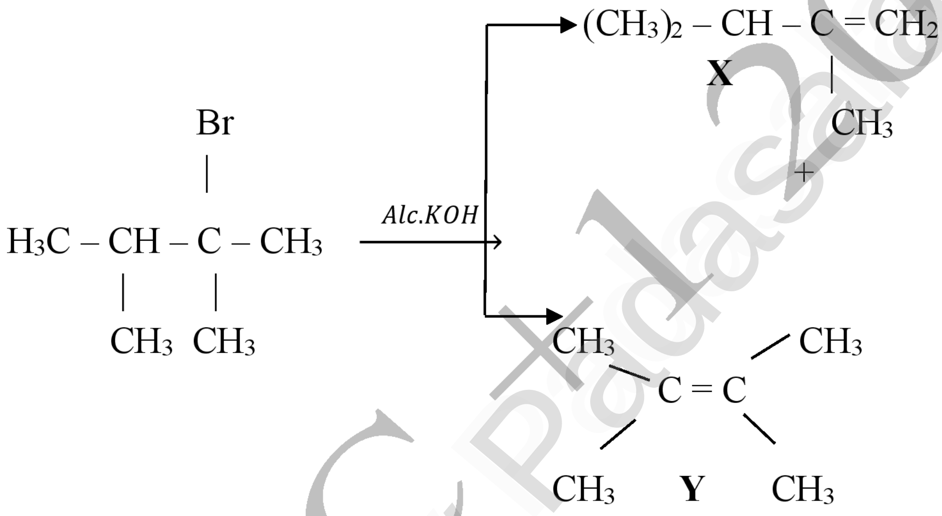


$$CH_3-\underset{Br}{\overset{Br}{C}}-CH_3 \xrightarrow{alc.KOH} CH_3-\overset{Br}{C}=CH_2 \xrightarrow{alc.KOH} \underset{\text{propyne}}{CH_3-C\equiv CH}$$

35.An alkylhalide with molecular formula $C_6H_{13}Br$ on dehydrohalogenation gave two isomeric alkenes X and Y with molecular formula C_6H_{12}. On reductive ozonolysis, X and Y gave fourcompounds CH_3COCH_3, CH_3CHO, CH_3CH_2CHO and $(CH_3)_2$ CHCHO. Find the alkylhalide.

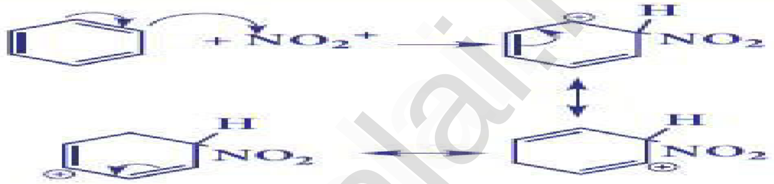
(i) The alkyl halide with molecular formula $C_6H_{13}Br$ is bromohexane (2 – Bromo – 2, 3-dimethyl butane)

(ii) $C_6H_{13}Br$ on dehydro halogenation gives two isomeric alkenes 2,3 – dimethyl – 1- butane (X) and 2,3 – dimethyl – 2 – butene (y)



(iii) X and Y on reductive ozonolysis gave four compounds CH_3COCH_3, CH_3CHO, CH_3CH_2CHO and $(CH_3)_2$ CHCHO.

36. Describe the mechanism of Nitration of benzene.

Types of reaction	NITRATION
Reagents	Con HNO_3 + Con H_2SO_4
Electrophile	$—NO_2^+$
Over all reaction	
Mechanism Step 1	$NO_3 + H_2SO_4 \longrightarrow \overset{\oplus}{N}O_2 + HSO_4 + H_2O$
Step 2	
STEP 3	

37. How does Huckel rule help to decide the aromatic character of a compound.

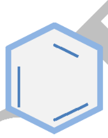
A compound may be aromatic, if it obeys Huckel rules

(i) The molecule must be co-planner

(ii) Complete delocalization of π electron in the ring.

(iii) Presence of (4n +2) π electrons in the ring where n is an integer(n = 0,1,2…..)

Eg:



Benzene

(i) The benzene is a planner molecule

(ii) It has six deloclaised π electrons

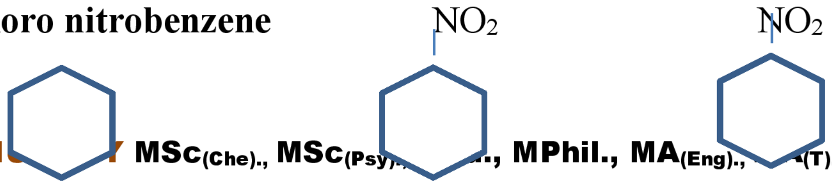
(iii) 4n + 2 = 6; 4n = 6 -2; 4n= 4; n = 1

It obeys Huckel's (4n + 2) π electron rule with n = 1 hence , benzene is aromatic

38 Suggest the route for the preparation of the following from benzene.

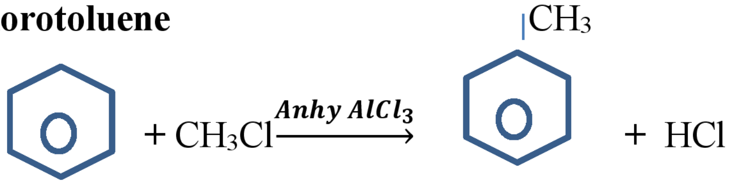
1) 3 – chloro nitrobenzene 2) 4 – chlorotoluene 3) Bromobenzene 4) m - dinitro benzene

1) 3 – chloro nitrobenzene

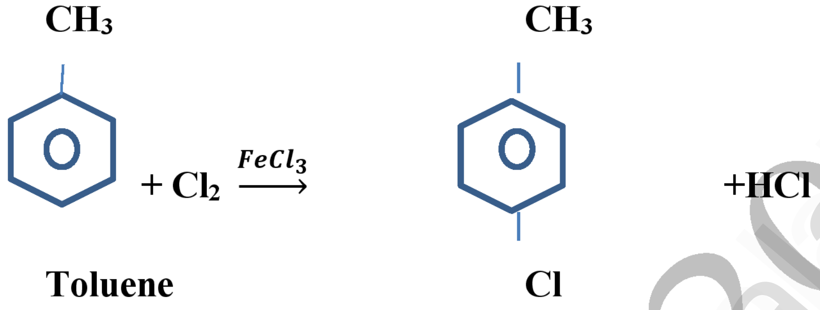


Benzene nitrobenzene 3- Chloronitro benzene

2) 4 – chlorotoluene



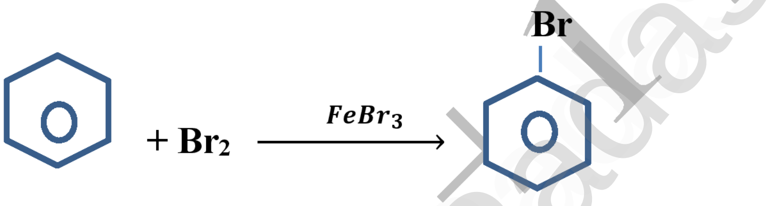
Benzene **Toluene**



Toluene

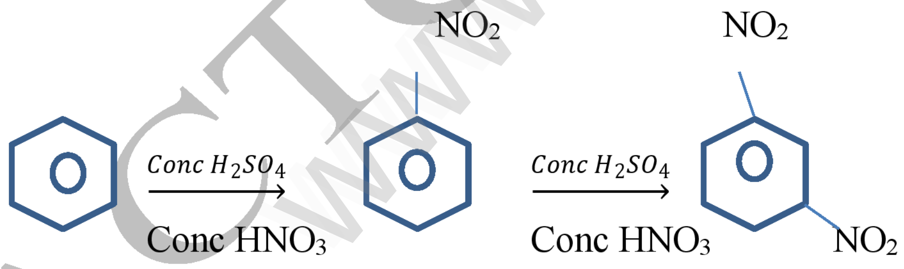
4 – Chloro toluene

3) Bromo benzene



Benzene **Bromo benzene**

4) m - dinitro benzene



Benzene Nitro benzene m – dinitro benzene

39. Suggest a simple chemical test to distinguish propane and propene.

Test to distinguish propane and propene

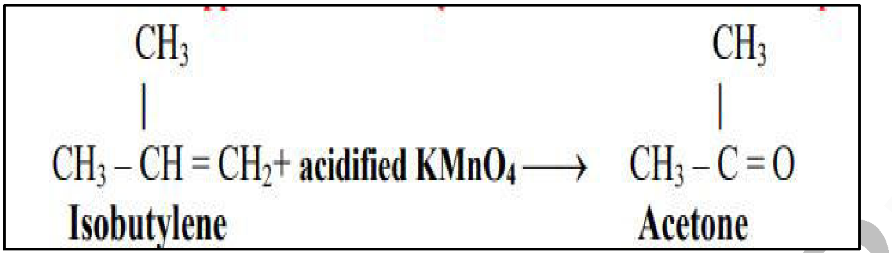
(i) Alkenes (propene) decolorize Br_2/H_2O whereas alkanes (propane) does not undergo this reaction.

(ii) De – colorization of acidified $KMnO_4$, propene de-colorises acidified $KMnO_4$ whereas propane does not.

- Both test for unsaturation.

TEST	PROPANE (CH_3-CH_2-CH_3)	PROPENE (CH_3-CH=CH_2)
1,With bromine water	No colour change	Colourless
2, Addition Reaction	No reaction	Takes place
3, With Ozone	No reaction	Aldehyde Form

40. What happens when isobutylene is treated with acidified potassium permanganate ?



$$\underset{\text{Isobutylene}}{CH_3 - \overset{\overset{\displaystyle CH_3}{|}}{CH} = CH_2} + \text{acidified } KMnO_4 \longrightarrow \underset{\text{Acetone}}{CH_3 - \overset{\overset{\displaystyle CH_3}{|}}{C} = O}$$

41. How will you convert ethyl chloride in to i) ethane ii) n – butane

(i) Ethyl Chloride → Ethane:

$$\underset{\text{Ethyl Chloride}}{CH_3 - CH_2 - Cl} \xrightarrow[\text{Zn/ HCl}]{2[H]} \underset{\text{Ethane}}{CH_3 - CH_3} + HCl$$

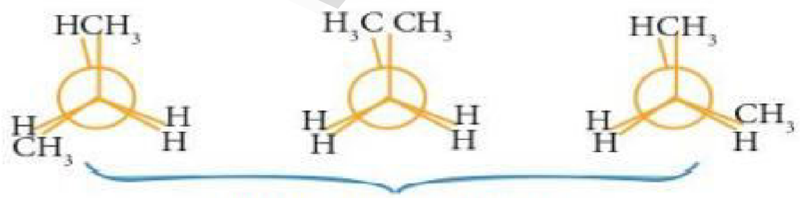
(ii) Ethyl Chloride → n – Butane (Wurtz reaction):

$$\underset{\text{Ethyl chloride}}{CH_3 - CH_2 - Cl} + 2Na + Cl - CH_2 - CH_3 \xrightarrow{Dry\ Ether} \underset{\text{n – Butane}}{CH_3 - CH_2 - CH_2 - CH_3} + 2NaBr$$

41. Describe the conformers of n - butane.

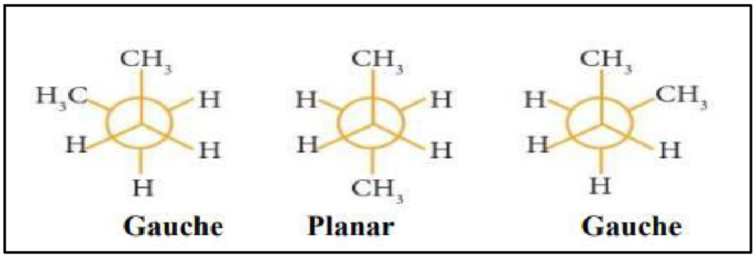
Conformations Of n – Butane: n – butane may be considered as a derivative of ethane, as one hydrogen on each carbon is replaced by a methyl group

Eclipsed Conformation: In this conformation, the distance between the two methyl group is minimum. So there is maximum repulsion between them and it is the least stable conformer.



Eclipsed conformations

Anti or Staggered form: In this conformation, the distance between the two methyl groups is maximum. So there is minimum repulsion between them and it is the most stable conformer



43. Write the chemical equations for combustion of propane.

Propane burns in excess of oxygen to form water and carbon di oxide.

$$\underset{\text{Propane}}{CH_3CH_2CH_3} + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

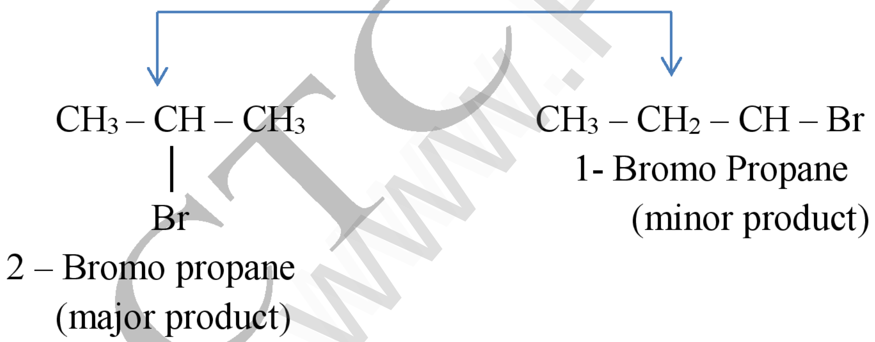
44. Explain Markownikoff's rule with suitable example.

Markovikoff's rule: " When an unsymmetrical alkene reacts with hydrogen halide, the hydrogen adds to the carbon that has more number of hydrogen and halogen add to the carbon having fewer hydrogen".

Eg: Addition HBr to unsymmetrical alkene: In the addition of hydrogen halide to an unsymmetrical alkene, two products are obtained.

$CH_3 = CH – CH_2 + HBr$

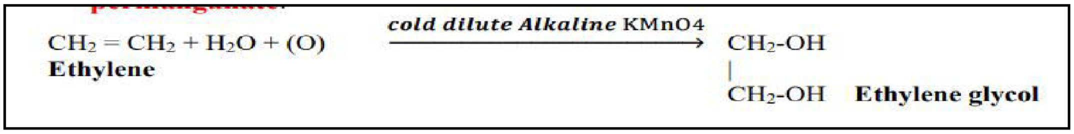
Propane



$$\underset{\substack{\text{2 – Bromo propane}\\\text{(major product)}}}{CH_3 – \underset{\displaystyle Br}{\underset{|}{CH}} – CH_3} \qquad \underset{\substack{\text{1- Bromo Propane}\\\text{(minor product)}}}{CH_3 – CH_2 – CH – Br}$$

45. What happens when ethylene is passed through cold dilute alkaline potassium permanganate.

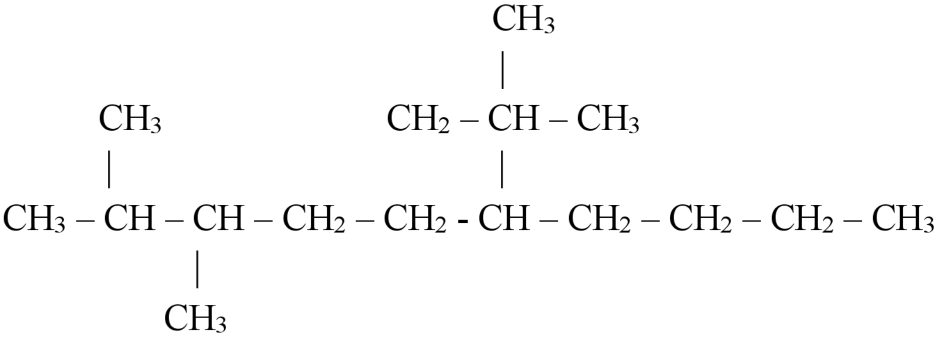
Ethene reacts with cold alk $KMnO_4$ (Balyer's regent to give ethane 1,2 – diol



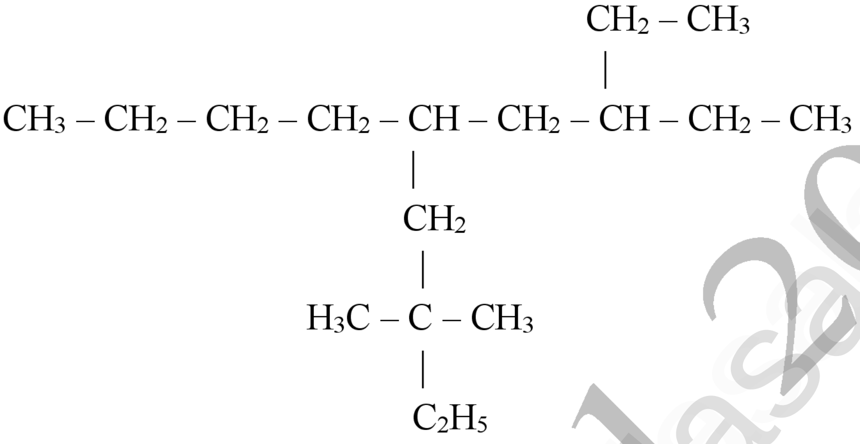
$$\underset{\text{Ethylene}}{CH_2 = CH_2 + H_2O + (O)} \xrightarrow{\text{cold dilute Alkaline } KMnO4} \underset{\displaystyle CH_2\text{-}OH}{\overset{\displaystyle CH_2\text{-}OH}{|}} \quad \text{Ethylene glycol}$$

46. Write the structures of folowing alkanes.

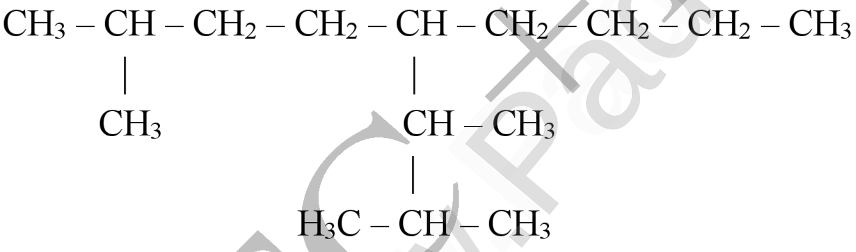
1) 2, 3 – Dimethyl – 6 – (2 – methyl propyl) decane



2) 5 – (2 – Ethyl butyl) – 3, 3 – dimethyldecane



3) 5 – (1, 2 – Dimethyl propyl) – 2 – methylnonane



47. How will you prepare propane from a sodium salt of fatty acid ?

$$CH_3CH_2CH_2COONa + NaOH \xrightarrow[\Delta]{CaO} CH_3CH_2CH_3 + Na_2CO_3$$

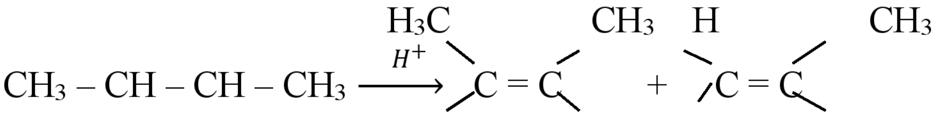
Sodium salt of butanaic acid Propane

Heating sodium salt of butanoic acid (Sodium butanoate) with soda lime gives propane.

48. $CH_3 – CH (CH_3) – CH (OH) – CH_3$

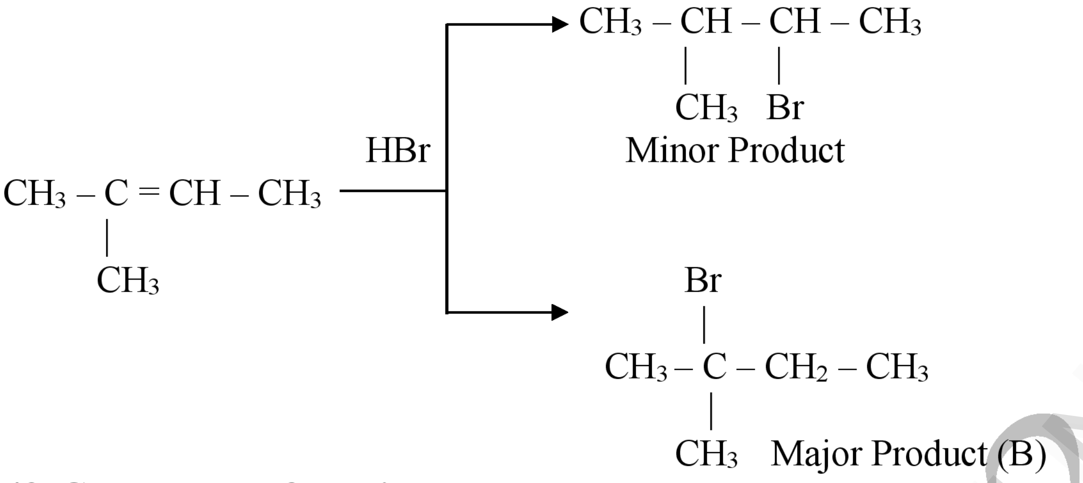
$\downarrow$ H^+/heat

(A) major product $\xrightarrow{HBr}$ (B) major product Identify A and B



| | CH3 OH Δ H3C H H CH2 – CH3

(A) Major Product Minor Product



$$CH_3-\underset{CH_3}{\underset{|}{C}}=CH-CH_3 \xrightarrow{HBr} CH_3-\underset{CH_3}{\underset{|}{CH}}-\underset{Br}{\underset{|}{CH}}-CH_3$$

Minor Product

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

Major Product (B)

49. Complete the following :

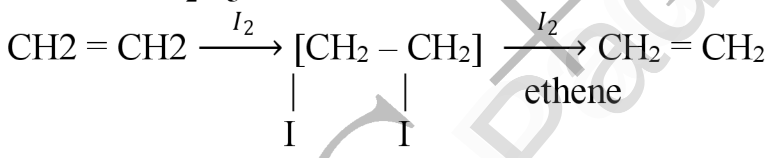
i) 2 – butyne $\xrightarrow{\textbf{Lindlar Catalyst}}$

$$\underset{2-Butyne}{CH_3-C\equiv C-CH_3+H_2} \xrightarrow[\text{Catalyst}]{Lindlar} \underset{2-Butene}{CH_3-CH=CH-CH_3}$$

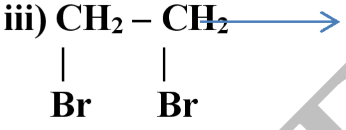
Lindlar catalyst consist of pd deposited on $CaCO_3$ and then poisoned by lead on sulphur.

ii) $CH_2=CH_2$ $\xrightarrow{I_2}$

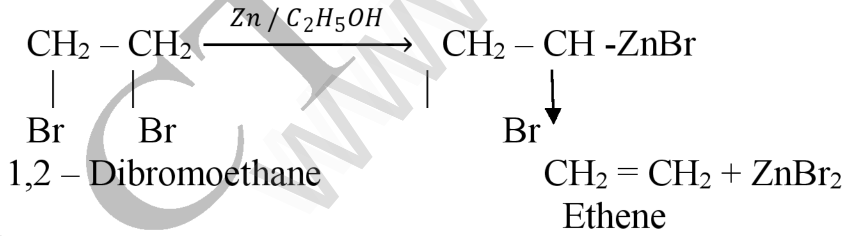
Zn/C_2H_5OH



$$CH2=CH2 \xrightarrow{I_2} [\underset{I}{\underset{|}{CH_2}}-\underset{I}{\underset{|}{CH_2}}] \xrightarrow{I_2} \underset{ethene}{CH_2=CH_2}$$



iii) $\underset{Br}{\underset{|}{CH_2}}-\underset{Br}{\underset{|}{CH_2}} \longrightarrow$



$$\underset{Br}{\underset{|}{CH_2}}-\underset{Br}{\underset{|}{CH_2}} \xrightarrow{Zn/C_2H_5OH} \underset{Br}{\underset{|}{CH_2}}-CH\text{-}ZnBr \longrightarrow CH_2=CH_2+ZnBr_2$$

1,2 – Dibromoethane Ethene

iv) $CaC_2 \xrightarrow{H_2O} CH\equiv CH+Ca(OH)_2$

$$\underset{\text{Calcium Carbride}}{CaC_2} \xrightarrow{H_2O} \underset{ethyne}{CH\equiv CH}+Ca(OH)_2$$

50. How will you distinguish 1 – butyne and 2 – butyne?

An alkyne shows acidic nature only if it contains terminal hydrogen.

$CH_3 - CH_2 - C \equiv CH + 2AgNO_3 + 2NH_4OH \rightarrow CH_3 - CH_2 - C \equiv C - Ag + 2NH_4NO_3 + 2H_2O$

1 – butyne Silver butynide

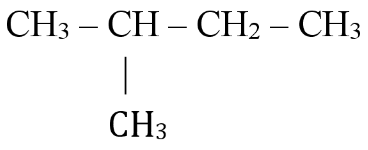
$CH_3 - C \equiv C - CH_3 + 2AgNO_3 + 2NH_4OH \rightarrow$ No Reaction due to absence of acidic hydrogen

Test	1-butyne	2-butyne
1, With $AgNO_3 + NH_4OH$	It gives silver *butynide*	No reaction
2, With $O_3 + H_2O/H_2O_2$	It gives *formic acid* and *propanoic acid*	It gives two molecules *Acetic acid*.

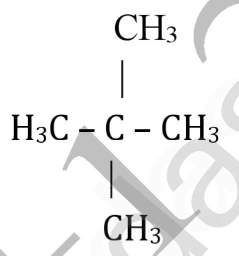
13.Hydrocarbons

M19	**J19**	**S20**	**S21**	**M22**	**J22**	**M23**	**J23**	**M24**	**J24**				
8	**9**	**9**	**11**	**11**	**11**	**8**	**9**	**9**	**6**				

1. How are hydrocarbons classified? (179)
2. Write the IUPAC name for the following compound. 181 **M19 5Mi**

(A) 

$CH_3 - \underset{\displaystyle CH_3}{\underset{|}{CH}} - CH_2 - CH_3$

(B) 

$H_3C - \overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_3}{\underset{|}{C}}}} - CH_3$

3. Write the structural formula and carbon skeleton for all possible chain isomers of C_4H_{10}, C_5H_{12}, C_6H_{14}. (181,182)
4. What is Sabatier Sendersen reaction? (184)
5. What is meant by Decarboxylation? decarboxylation of sodium acetate. (184)
6. Wurtz reaction. (184)
7. How will you convert ethyl chloride in to i) ethane **(M22 2M)** ii) n – butane (bb) (184)
8. Corey house mechanism. (185)
9. Write note on Kolbe's electrolytic method of preparation of alkanes (184) **J22 2M**
10. How is **Alkane** (methane) prepared from Grignard reagent. (185) **M19 2M**
11. Complete the following: (184, 185) **S21 2M COMPULSORY**
    a) $CH_3\text{-}CH{=}CH_2 + H_2 \rightarrow$
    b) $CH_3MgCl + H_2O \rightarrow$
12. Write note on confirmation of ethane, Butane. (186, 187)
13. Write the chemical equations for combustion of propane.(bb)(188)
14. Define pyrolysis. Give an example. (189)
15. Uses of alkane. (190)
16. Explain Geometrical isomerism of 2-butene. (191) **S20 3M COMPULSORY**
17. How are alkenes prepared from alkynes by Lindlar's catalyst? (192)

18. $CH_3CH_2Cl \xrightarrow{Conc H_2SO_4\ 430-440K} A \xrightarrow{HBr\ Benzoyl\ peroxide} B$ 192, 195 **J23 2MCOM**

19. Suggest a simple chemical test to distinguish propane and propene. (bb)(194) **S21 3M**

20. Explain Markownikoff's rule with suitable example. (194)

21. An organic compound (A) C_2H_4 decolourises bromine water.(A) on reaction with chlorine gives (B). (A) reacts with HBr to give (C). Identify (A), (B) and (C). Explain the reactions. (194) **J22 5M**

22. What happens when isobutylene is treated with acidified potassium permanganate? (bb)(198)

23. What happens when ethylene is passed through cold dilute alkaline potassium permanganate. (197-198) **M22 3M**

24. An organic compound (A) of molecular formula C_2H_6O, on heating with conc. H_2SO_4 gives compound (B). (B) on treating with cold dilute alkaline $KMnO_4$ gives compound (C). Identify (A), (B) and (C) and explain the reactions. (197) **J20 5Mi**

25. Ozonolysis of alkene. (198) **J19 2M**

26. What is polymerization? preparation of polyethene. (199) **J19 3M**

27. Find A,B and C. (200) **M24 3M Compulsory**

$$CH_2{=}CH_2 + Br_2 \rightarrow A \xrightarrow{alcKOH} B \xrightarrow{NaNH_2} C$$

28. Complete the reaction: $CaC_2 \xrightarrow{H_2O}$ 202 **J23 5Mi**

29. Explain the different types of polymerisations in ethyne. 204 **S20 3M**

30. How will you distinguish 1-Butyne and 2-Butyne? **J24 5Mii**

31. How does Huckel rule help to decide the aromatic character of a compound. (205) **S21 5Mi, J23 3M, J24 2M**

32. Explain the evidence of structure of benzene (207) **M23 5M**

33. Write any two different components you gets during fractional distillation of Coal tar at any two different temperatures. (209) **M19 5Mii**

34. Explain preparation of benzene (3methods) (210)

35. Electrophilic substitution reaction of benzene
(Nitration, halogenation, Sulphonation, Methylation, Acetylation) (211-212)

36. How will you get the following products with the given reactants? (210) **M19 5Mi**
    - (A) Acetylene → Benzene **S21 5Mii, M24 5Mia**
    - (B) Phenol → Benzene **M24 5Mib**
    - (C) Benzene → Toluene

37. How will you prepare the following compounds from benzene? **M22 5M**
(i)Nitrobenzene (ii)Benzene sulphonic acid (iii)BHC

38. The simple Aromatic Hydrocarbon compound (A) reacts with Bromine to give (B). Compound (A) reacts with Raney Ni and gives (C). Identify (A), (B) and (C). (211,215) **M19 5Mii**

39. Wurtz-Fittig reaction.(210)

40. Fridel crafts reaction. (210) **M23 2M**

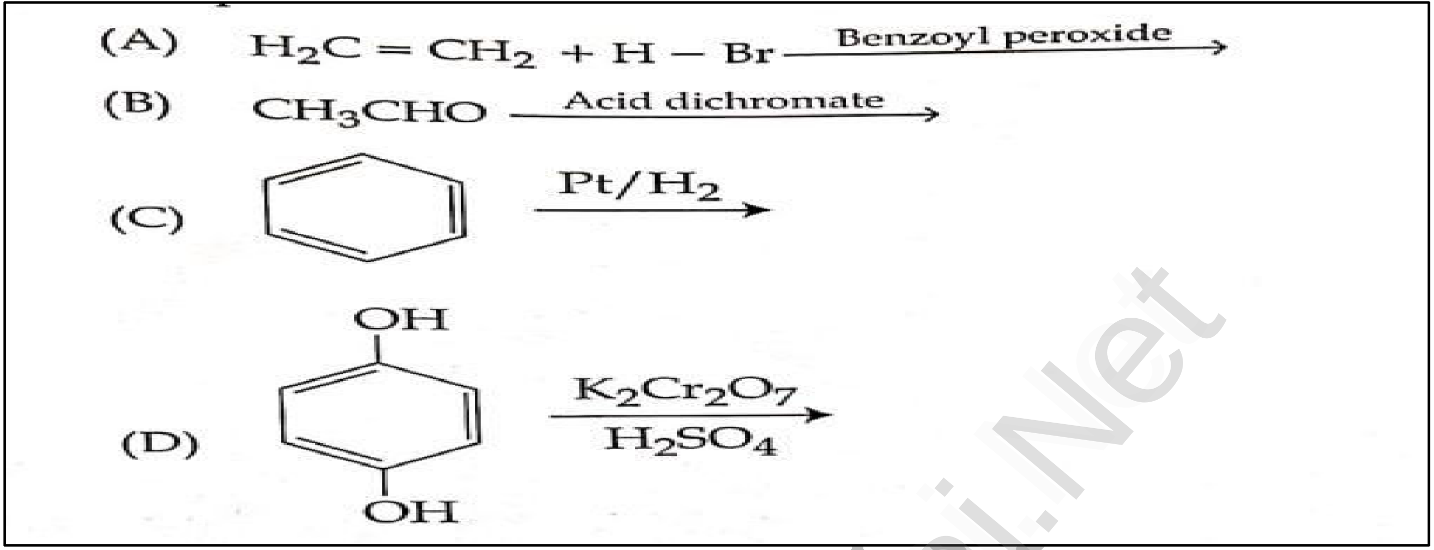
41. How will you convert Benzene to BHC? (215) **M24 2M**

42. Describe the mechanism of nitration of benzene (211,212,213)

43. Describe the mechanism of Nitration of benzene. (214)

44. Brich reduction(215) **J19 5Mii, S20 5Mii, J22 3M**

45. Complete: 195, 215 **S20 5Mi**



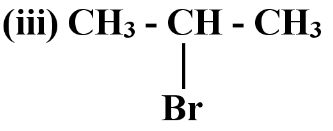
46. Write short notes on ortho, para directors in aromatic electrophilic substitution reactions. (216)

47. Write short notes on meta directors in aromatic electrophilic substitution reactions. (216)

&PRACTICE WORK SHEET study well BOOK exercise questions.

Unit 11: Haloalkanes and Haloarenes

26. Classify the following compounds in the form of alkyl, allylic, vinyl, benzylic halides. (i) $CH_3 – CH = CH – Cl$ (ii) $C_6H_5CH_2I$

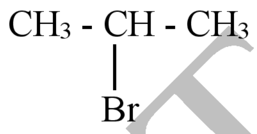


(iii) $CH_3 - \underset{\underset{Br}{|}}{CH} - CH_3$ (iv) $CH_2 = CH – Cl$

($CH_2 =$ CH $-CH_2 – Cl$ - allyl chloride → Allylic halide)

i) $CH_3 – CH = CH – Cl$ - VINYL (sp^2)

(i) $C_6H_5CH_2I$ - Benzyl iodide → **Benzylic Halide**



(iii) $CH_3 - \underset{\underset{Br}{|}}{CH} - CH_3$ - 2-bromo propane → **Alkyl halide**

(iv) $CH_2 = CH – Cl$ – Vinyl chloride → **Vinyl halide**

27. Why chlorination of methane is not possible in dark?

- Reaction of methane with chlorine proceeds by **free radical mechanism.**
- The initiation step in the free radical chain reaction is $Cl_2 \rightarrow 2Cl^{\cdot}$. This step requires chain more energy.
- The excess energy is provided by heat or light.
- **In dark, chlorine free radicals' formation is not possible and so chlorination of methane is not possible in dark.**
- The ultraviolet light is a source of energy and it being used to break of Cl-Cl and produce Cl free radical's which can attack methane. In dark this is not possible.

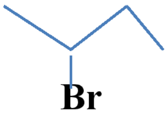
28. How will you prepare n- propyl iodide from n-propyl bromide?

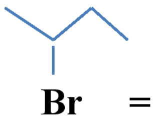
n-propyl iodide from n-propyl bromide. n-propyl bromide on heating with concentrated solution of sodium iodide in acetone gives n-propyl Iodide. This reaction is called **Finkelstein reaction.**

$CH_3CH_2CH_2Br + NaI \xrightarrow{acetone} CH_3CH_2CH_2I + NaBr$

n-propyl bromide n-propyl iodide

29. Which alkyl halide from the following pair is i) chiral ii) undergoes faster S_N^2 reaction?



(i) 

= chiral (the chiral carbon atom and it is surrounded by four different group).

(ii) 

i.e. 1-chlorobutane undergoes faster S_N^2 reaction.

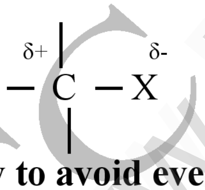
30. How does chlorobenzene react with sodium in the presence of ether? What is the name of the reaction?

Haloarenes react with sodium metal in dry ether, two aryl groups combine to give biaryl products is reaction is called **fittig reaction.**

$C_6H_5Cl + 2Na + Cl - C_6H_5 \xrightarrow{ether} C_6H_5 - C_6H_5 + 2NaCl$

Chlorobenzene Biphenyl

31. Give reasons for polarity of C-X bond in halo alkane.

- Carbon halogen bond is a polar bond as halogens are more electro negative than carbon.
- The carbon atom exhibits a partial positive charge (δ+) and halogen atom a partial negative charge (δ-)
- 

32. Why is it necessary to avoid even traces of moisture during the use of Grignard reagent?

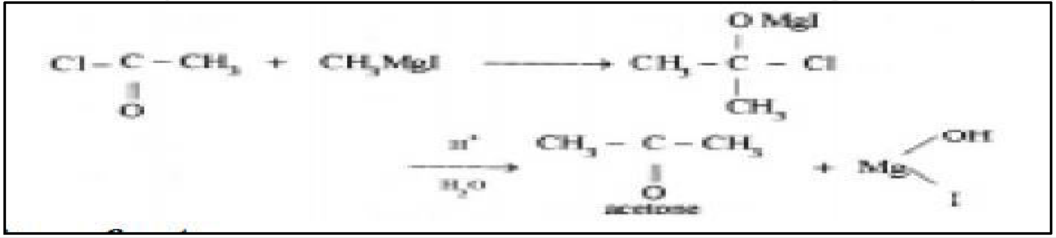
- Grignard reagents are highly reactive substances.
- The Grignard carbon is highly basic and reacts with acidic protons of polar solvents like water to form an alkane so all reagents should be pure and dry.
- Even water is sufficiently acidic to convert it into the corresponding hydrocarbon.
- So it is necessary to avoid even traces of moisture with the Grignard reagent as they are highly reactive.

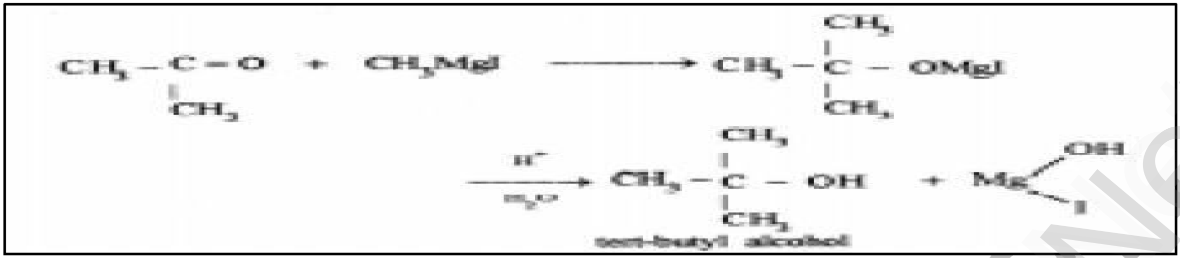
$RMgX + H_2O \rightarrow RH + Mg(OH)X$

Grignard Alkane

33. What happens when acetyl chloride is treated with excess of CH_3MgI?

When acetyl chloride is treated with excess of CH_3MgI, tertiary butyl alcohol are formed.





34. Arrange the following alkyl halide in increasing order of bond enthalpy of RX CH_3Br, CH_3F, CH_3Cl, CH_3I.

The bond strength of C-X bond decreases form C-F to C-I in CH_3X.

Increasing order of bond enthalpy.

$CH_3I < CH_3Br < CH_3Cl < CH_3F$

35. What happens when chloroform reacts with oxygen in the presence of sunlight?

Chloroform undergoes oxidation in the presence of light and air to form phosgene (carbonyl chloride).

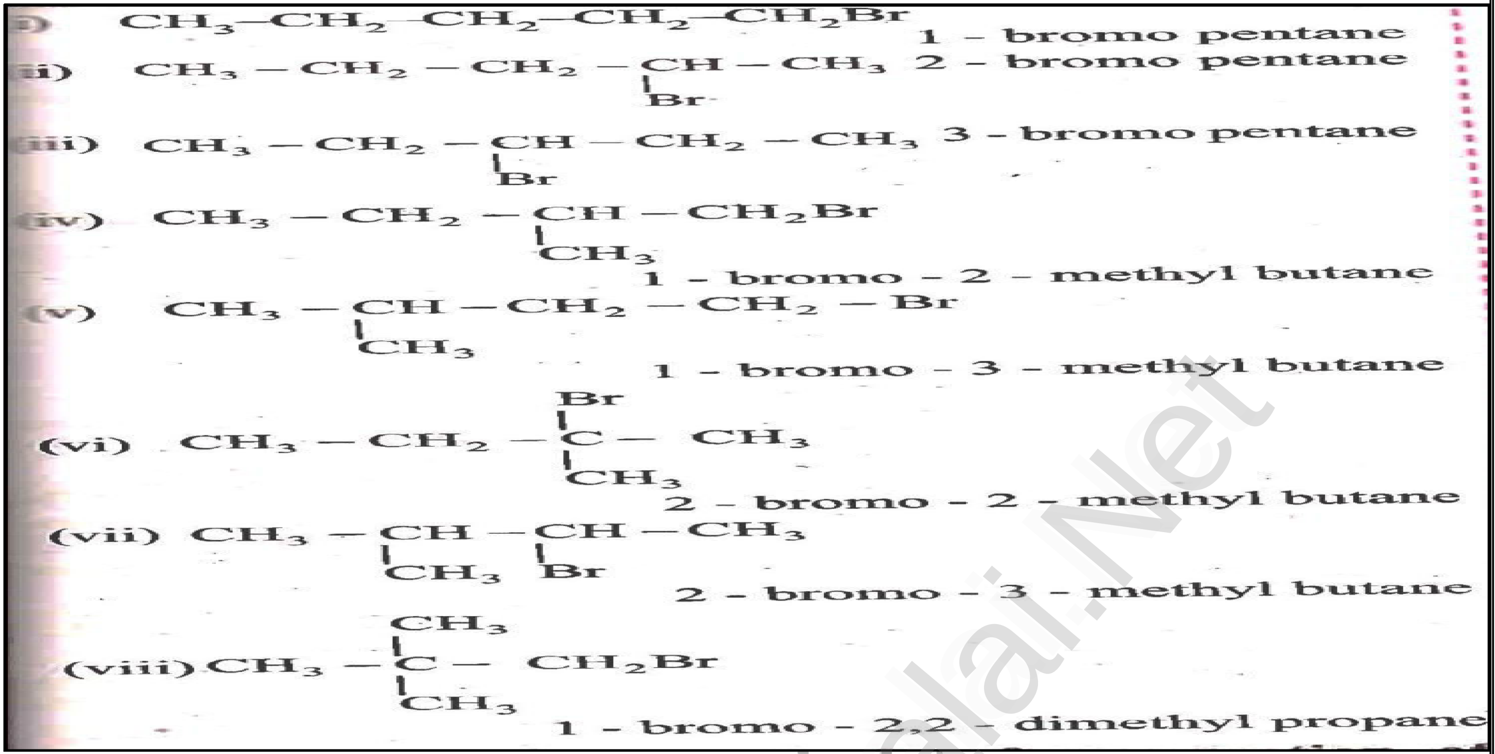
$CHCl_3 + \frac{1}{2} O_2 \xrightarrow{Air/light} COCl_2 + HCl$

Chloroform Phosgene

Since phosgene is very poisonous, its presence makes chloroform unfit for use as **anaesthetic.**

36. Write down the possible isomers of $C_5H_{11}Br$ and give their IUPAC and common names.

All possible isomers of $C_5H_{11}Br$. There are 8 isomers that have the molecule formula $C_5H_{11}Br$.



i) $CH_3-CH_2-CH_2-CH_2-CH_2Br$ 1 – bromo pentane

ii) $CH_3-CH_2-CH_2-\underset{Br}{\underset{|}{CH}}-CH_3$ 2 – bromo pentane

iii) $CH_3-CH_2-\underset{Br}{\underset{|}{CH}}-CH_2-CH_3$ 3 – bromo pentane

iv) $CH_3-CH_2-\underset{CH_3}{\underset{|}{CH}}-CH_2Br$ 1 – bromo – 2 – methyl butane

(v) $CH_3-\underset{CH_3}{\underset{|}{CH}}-CH_2-CH_2-Br$ 1 – bromo – 3 – methyl butane

(vi) $CH_3-CH_2-\overset{Br}{\overset{|}{\underset{CH_3}{\underset{|}{C}}}}-CH_3$ 2 – bromo – 2 – methyl butane

(vii) $CH_3-\underset{CH_3}{\underset{|}{CH}}-\underset{Br}{\underset{|}{CH}}-CH_3$ 2 – bromo – 3 – methyl butane

(viii) $CH_3-\overset{CH_3}{\overset{|}{\underset{CH_3}{\underset{|}{C}}}}-CH_2Br$ 1 – bromo – 2,2 – dimethyl propane

37. Mention any three methods of preparation of haloalkanes from alcohols.

- **Reaction with HCl(lucas reagent)**

$$CH_3CH_2OH + HCl \xrightarrow[\Delta]{AnhydrousZnCl2} CH_3CH_2Cl + H_2O$$

Ethanol chloro ethane

- **Reaction with PCl_5**

$CH_3CH_2OH + PCl_5 \rightarrow CH_3CH_2Cl + POCl_3 + HCl$

Ethanol chloro ethane

- **Reaction with $SOCl_2$**

$$CH_3CH_2OH + SOCl_2 \xrightarrow{pyridine} CH_3CH_2Cl + SO_2 + HCl$$

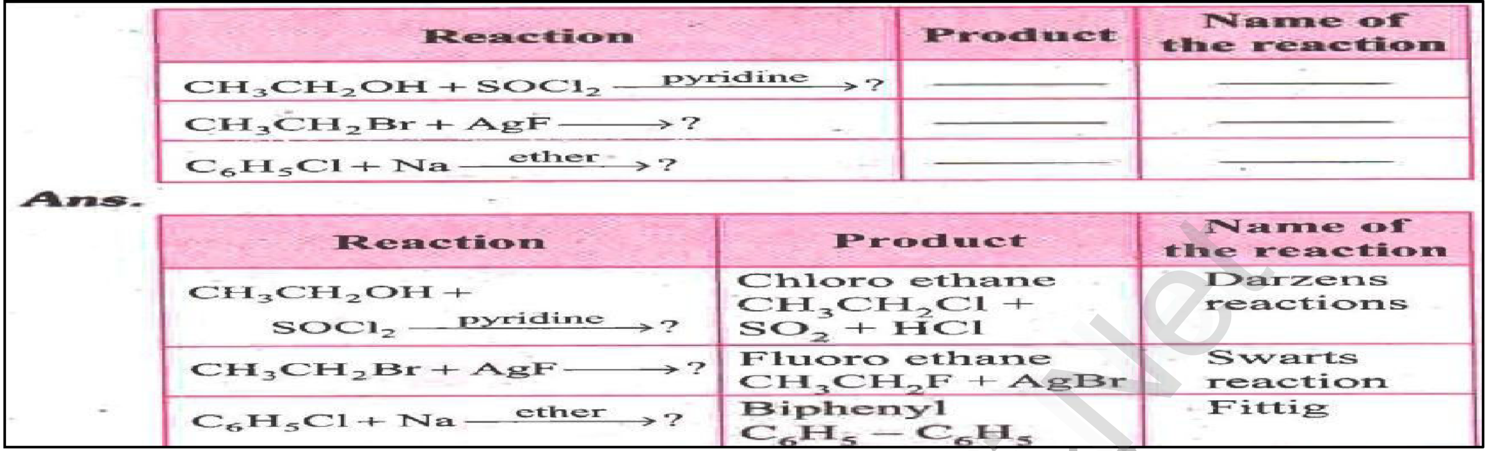
Ethanol chloro ethane

38. Compare SN_1 and SN_2 reaction mechanisms.

SN^1	**SN^2**
It is **Unimolecular Nucleophilic substitution** reaction.	It is a **bimolecular Nucleophilic substitution** reaction.
Its mechanism occurs in **two steps**.	It is a **one step** process.
It follows **first order** kinetics.	It follows **second order** kinetics.
It involves the **formation of an intermediate.**	It involves the **formation of transition state.**
Rate =k[Alkyl halide]	**Rate =k[Alkyl halide][nucleophile]**
Products have both retained and inverted configuration.	**Products have inverted configuration.**
Carbocation rearrangement occurs.	No carbocation rearrangement occurs

Reactivity: methyl <1°< 2°<3°	**Reactivity:** methyl >1°>2°>3°

39. Reagents and the conditions used in the reactions are given below. Complete the table by writing down the product and the name of the reaction.



Reaction	Product	Name of the reaction
$CH_3CH_2OH + SOCl_2 \xrightarrow{pyridine} ?$	________	________
$CH_3CH_2Br + AgF \longrightarrow ?$	________	________
$C_6H_5Cl + Na \xrightarrow{ether} ?$	________	________

Ans.

Reaction	Product	Name of the reaction
$CH_3CH_2OH + SOCl_2 \xrightarrow{pyridine} ?$	Chloro ethane $CH_3CH_2Cl + SO_2 + HCl$	Darzens reactions
$CH_3CH_2Br + AgF \longrightarrow ?$	Fluoro ethane $CH_3CH_2F + AgBr$	Swarts reaction
$C_6H_5Cl + Na \xrightarrow{ether} ?$	Biphenyl $C_6H_5 - C_6H_5$	Fittig

40. Discuss the aromatic Nucleophilic substitution's reaction of chlorobenzene.

Halo arenes do not undergo nucleophilic substitution reaction readily. This is due to C–X bond aryl halide is short and strong and also the aromatic ring is a centre of high electron density. The halogen of haloarenes can be substituted by OH^- , NH_2 –, or CN^- with appropriate nucleophilic reagents at high temperature and pressure.

For Example

$$\underset{\text{Chlorobenzene}}{C_6H_5Cl} + NaOH \xrightarrow[300\ atm]{350°C} \underset{\text{Pheno1}}{C_6H_5OH} + NaCl$$

This reaction is known as **Dow's Process.**

41. Account for the following (i) t-butyl chloride reacts with aqueous KOH by SN_1 mechanism while n-butyl chloride reacts with SN_2 mechanism. (ii) p-dichloro benzene has higher melting point than those of o-and m-dichloro benzene.

i) This is because rate of SN_2 reaction is directly proportional to delta inversely proportional to steric crowding. Tert-butyl chloride has vast steric crowding and thus it is more stable for SN_1 reaction rather then SN_2 reaction.

ii)The higher melting point of p-isomer is due to its **symmetry** which leads to more close packing of its molecules in the crystal lattice and consequently strong intermolecular attractive force which requires more energy for melting.

p –Dihalo benzene > o- Dichloro benzene > m-Dichloro benzene

42. In an experiment ethyl iodide in ether is allowed to stand over magnesium pieces. Magnesium dissolves and product is formed a) Name the product and write the equation for the reaction.

b) Why all the reagents used in the reaction should be dry? Explain

c) How is acetone prepared from the product obtained in the experiment.

a) The product formed is ethyl magnesium iodide.(Grignard reagent)

$$C_2H_5I + Mg \xrightarrow{dry\ ether} C_2H_5MgI$$

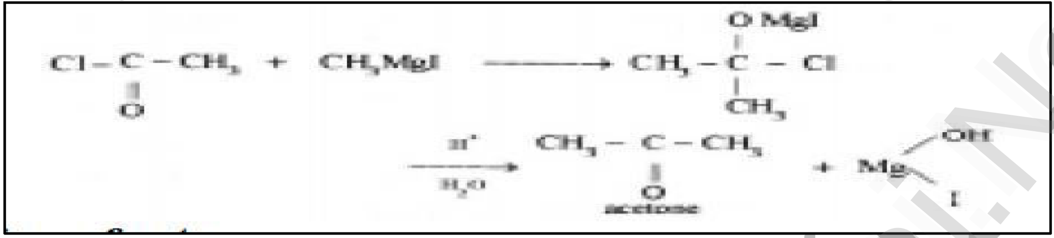
b) Grignard reagents are highly reactive substances.

- The Grignard carbon is highly basic and reacts with acidic protons of polar solvents like water to form an alkane so all reagents should be pure and dry.
- Even water is sufficiently acidic to convert it into the corresponding hydrocarbon.
- So it is necessary to avoid even traces of moisture with the Grignard reagent as they are highly reactive.

$RMgX + H_2O \rightarrow RH + Mg(OH)X$

Grignard Alkane

When acetyl chloride is treated with CH_3MgI, acetone are formed.



43. **Write a chemical reaction useful to prepare the following: i) Freon-12 from Carbon tetrachloride ii) Carbon tetrachloride from carbon disulphide.**

i) Freon – 12 is prepared by the action of hydrogen fluoride on carbon tetrachloride in the presence of catalytic amount of antimony patachloride. This is called **Swartz reaction.**

$CCl_4 + 2HF \xrightarrow{SbCl5} 2HCl + CCl_2F_2$

Carbon Freon - 12

Tetrachloride

ii) Carbon disulphide reacts with chlorine gas in the presence of anhydrous $AlCl_3$ as catalyst giving carbon tetrachloride.

$CS_2 + 3Cl_2 \xrightarrow{\text{anhydrous AlCl3}} CCl_4 + S_2Cl_2$

carbon disulphide carbon tetrachloride

44. **What are Freons? Discuss their uses and environmental effects**

i) Freons are a used as refrigerants in refrigerators and air conditioners.

(ii) It is used as a propellant for aerosols and foams.

(iii) It is used as propellant for foams to spray out deodorants, shaving creams, and insecticides.

45. Predict the products when bromo ethane is treated with the following i) KNO_2ii) $AgNO_2$

i) Haloalkanes react with alcoholic solution of $NaNO_2$ or KNO_2 to form alkyl nitrites.

$CH_3CH_2Br + KNO_2 \rightarrow CH_3CH_2–O–N=O + KBr$

Bromoethane Ethyl nitrite

ii) Haloalkanes react with alcoholic solution of $AgNO_2$ to form nitro alkanes.

$CH_3CH_2Br + AgNO_2 \rightarrow CH_3CH_2NO_2 + AgBr$

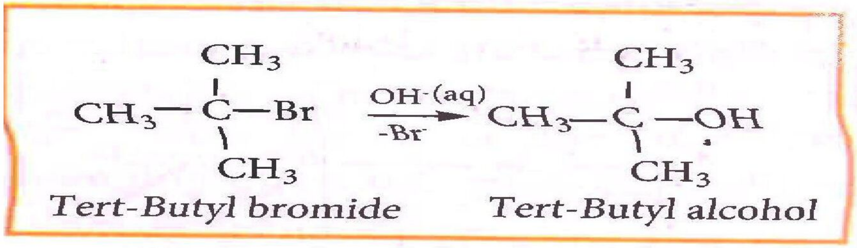
Bromo ethane Nitro ethane

46. **Explain the mechanism of SN_1 reaction by highlighting the stereochemistry behind it.**

S_N1 stands for Unimolecular Nucleophilic Substitution. The rate of the following S_N1 reaction depends upon the concentration of alkyl halide (RX) and is independent of the concentration of the nucleophile (OH^-)

Hence Rate of the reaction = k[alkyl halide]

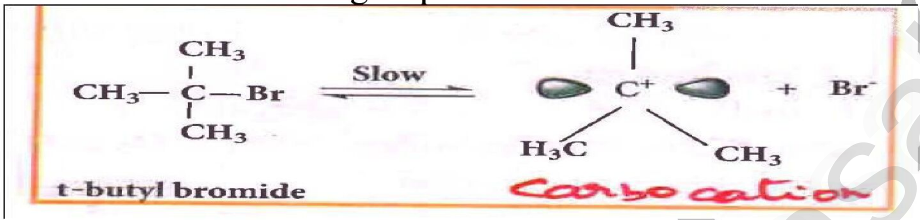
SN1 reaction mechanism by taking a reaction between tertiary butyl bromide with aqueous KOH.



This reaction takes place in two steps as shown below

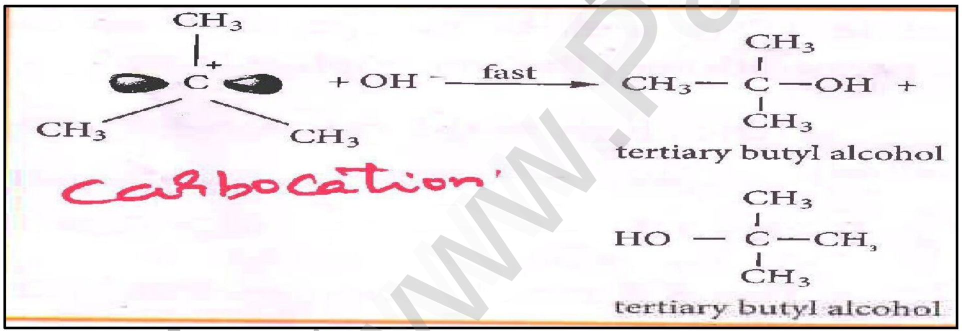
Step - 1 Formation of carbocation

The polar C - Br bond breaks forming a carbocation and bromide ion. This step is slow and hence it is the rate determining step.



The carbocation has 2 equivalent lobes of the vacant 2p orbital, so it can react equally rapidly from either face.

Step – 2 The nucleophile immediately reacts with the carbocation. This step is fast and hence does not affect the rate of the reactions.

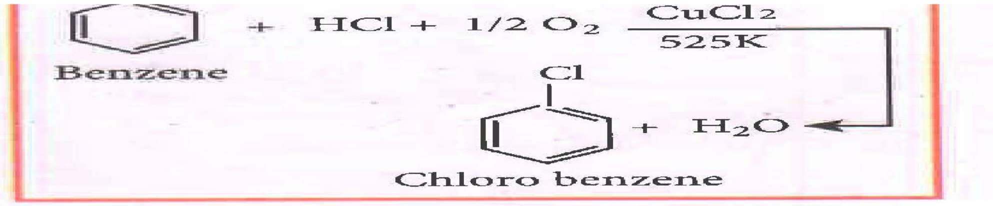


As shown above, the Nucleophilic reagent OH- can attack carbocation from both the sides.

47. Write short notes on the the following

i) Raschig process ii) Dows Process iii) Darzens process

i) **Raschig process**: Chloro benzene is commercially prepared by passing a mixture of benzene vapour, air and HCl over heated cupric chloride .This reaction is called Raschig process.



ii) Dows process:

$$C_6H_5Cl + NaOH \xrightarrow[300\ atm]{350°C} C_6H_5OH + NaCl$$

Chlorobenzene Phenol

This reaction is known as **Dow's Process.**

iii) Darzens process

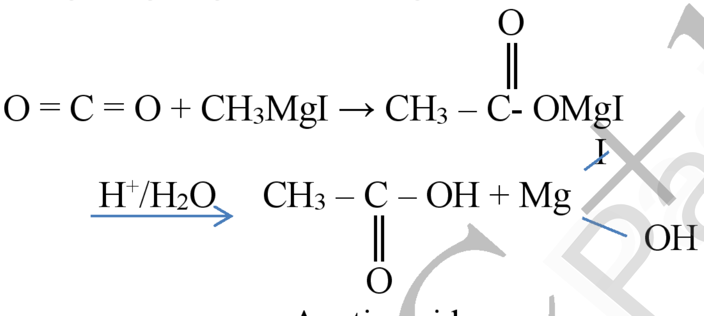
$$CH_3CH_2OH + SOCl_2 \xrightarrow{pyridine} CH_3CH_2Cl + SO_2 + HCl$$

Ethanol chloro ethane

48. Starting from CH_3MgI, How will you prepare the following?

i) Acetic acid ii) Acetone iii) Ethyl acetate iv) Iso propyl alcohol v) Methyl cyanide

i) Acetic acid : Solid carbon dioxide reacts with Grignard reagent to form addition product which on hydrolysis yields carboxylic acids.



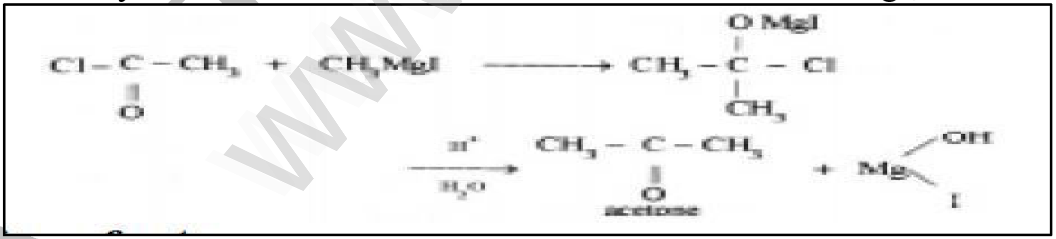
$$O = C = O + CH_3MgI \rightarrow CH_3 - \overset{\overset{O}{\|}}{C} - OMgI$$

$$\xrightarrow{H^+/H_2O} CH_3 - \underset{\underset{O}{\|}}{C} - OH + Mg\begin{smallmatrix} I \\ OH \end{smallmatrix}$$

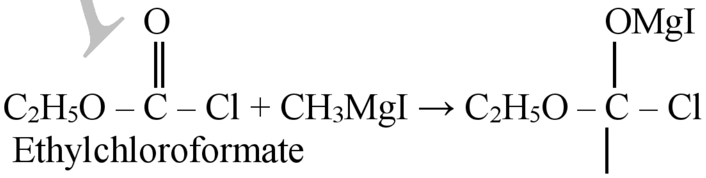
Acetic acid

ii) Acetone:

When acetyl chloride is treated with CH_3MgI, acetone are formed.

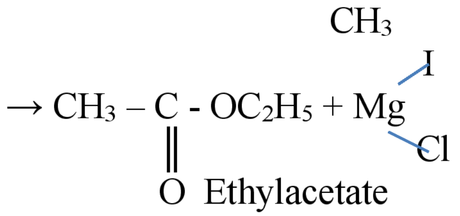


iii) Ethyl acetate: Ethylchloroformate reacts with Grignard reagent to form esters.



$$C_2H_5O - \overset{\overset{O}{\|}}{C} - Cl + CH_3MgI \rightarrow C_2H_5O - \underset{|}{\overset{\overset{OMgI}{|}}{C}} - Cl$$

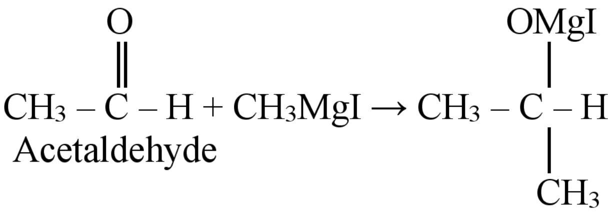
Ethylchloroformate



$$CH_3 \qquad \rightarrow CH_3-\underset{\underset{\displaystyle O}{\|}}{C}-OC_2H_5 + Mg\begin{matrix}\diagup I\\ \diagdown Cl\end{matrix}$$

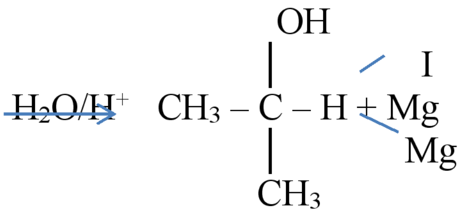
Ethylacetate

iv) Iso propyl alcohol : Aldehydes other than formaldehyde, react with Grignard reagent to give addition product which on hydrolysis yields secondary alcohol.



$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-H + CH_3MgI \rightarrow CH_3-\overset{\overset{\displaystyle OMgI}{|}}{\underset{\underset{\displaystyle CH_3}{|}}{C}}-H$$

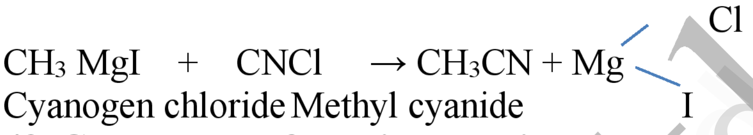
Acetaldehyde



$$\xrightarrow{H_2O/H^+} CH_3-\overset{\overset{\displaystyle OH}{|}}{\underset{\underset{\displaystyle CH_3}{|}}{C}}-H + Mg\begin{matrix}\diagup I\\ \diagdown Mg\end{matrix}$$

Isopropylalcohol(2)

v) Methyl cyanide : Grignard reagent reacts with cyanogen chloride to form alkyl cyanide



$$CH_3MgI + CNCl \rightarrow CH_3CN + Mg\begin{matrix}\diagup Cl\\ \diagdown I\end{matrix}$$

Cyanogen chloride Methyl cyanide

49. Complete the following reactions

i) $CH_3-CH=CH_2 + HBr \xrightarrow{Peroxide}$

ii) $CH_3-CH_2-Br + NaSH \xrightarrow[H_2O]{alcohol}$

iii) $C_6H_5Cl + Mg \xrightarrow{THF}$ iv) $CHCl_3 + HNO_3 \xrightarrow{} \Delta$ v) $CCl_4 + H_2O \xrightarrow{} \Delta$

i) The addition of HBr to an alkene in the presence of organic peroxide, gives the anti Markovniko's product. Is effect is called peroxide effect.

$CH_3-CH=CH_2 + HBr \xrightarrow[(C_6H_5CO)_2O_2]{Peroxide} CH_3-CH_2-CH_2-Br$

Propene 1- bromopropane

ii) Haloalkanes react with sodium or potassium hydrogen sulphide to form thio alcohols.

$CH_3-CH_2-Br + NaSH \xrightarrow[H_2O\ \Delta]{alcohol} CH_3CH_2SH + NaBr$

Bromo ethane Ethane thiol

iii) Chloroform reacts with nitric acid to form chloropicrin. (Trichloro nitro methane) Δ

$CHCl_3 + HNO_3 \rightarrow CCl_3NO_2 + H_2O$

Chloroform Chloropicrin

iv) Carbon tetrachloride reacts with hot water or with hot water vapour producing the poisonous gas, phosgene.

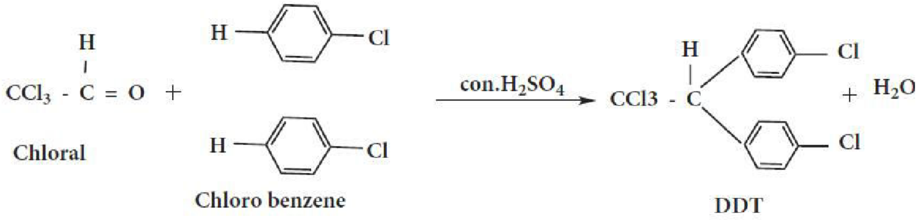
$CCl_4 + H_2O(g) \xrightarrow{\Delta} COCl_2 + 2H_2O$

Carbon tetrachloride Phosgene

50. Explain the preparation of the following compounds

i) DDT ii) Chloroform iii) Biphenyl iv) Chloropicrin v) Freon-12

i) DDT: DDT can be prepared by heating a mixture of chlorobenzene with chloral (Trichloro acetaldehyde) in the presence of con.H_2SO_4.



ii)Chloroform : The reaction of methane with excess of chlorine in the presence of sunlight will give carbon tetrachloride as the major product. $CH_4 + 4Cl_2$ hγ 4HCl

Methane $\longrightarrow$ carbon tetrachloride

iii) Biphenyl:

$$C_6H_5Cl + 2Na + Cl - C_6H_5 \xrightarrow[\Delta]{Ether} C_6H_5 - C_6H_5 + 2NaCl$$

Chlorobenzene Biphenyl

iv) Chloropicrin: Chloroform reacts with nitric acid to form chloropicrin.(Trichloro nitro methane)

$$CHCl_3 + HNO_3 \xrightarrow{\Delta} CCl_3NO_2 + H_2O$$

Chloroform Chloropicrin

v) Freon-12: Freon-12 is prepared by the action of hydrogen fluoride on carbon tetrachloride in the presence of catalylic amount of antimony pentachloride. It is called Swartz reaction.

$$CCl_4 + 2HF \xrightarrow{SbCl_5} 2HCl + CCl_2F_2$$

Carbon tetrachloride Freon - 12

51. An organic compound (A) with molecular formula C_2H_5Cl reacts with KOH gives compounds (B) and with alcoholic KOH gives compound (C). Identify (A), (B), and (C)

$CH_3CH_2Cl + KOH_{(Aq)} \rightarrow CH_3CH_2OH$

(A) (B)

Chloroethane Ethanol

\+

Alc KOH

↓

$CH_2 = CH_2$

(C)

Ethene

A	CH_3CH_2Cl	Chloroethane
B	CH_3CH_2OH	Ethanol
C	$CH_2 = CH_2$	Ethene

52. Simplest alkene (A) reacts with HCl to form compound (B).Compound (B) reacts with ammonia to form compound (C) of molecular formula C_2H_7N.Compound (C) undergoes carbylamine test. Identify (A), (B), and (C).

$CH_2 = CH_2 + HCl \rightarrow CH_3 - CH_2Cl$

(A) (B)

Ethene Chloroethane

$CH_3CH_2Cl + NH_3 \rightarrow CH_3CH_2NH_2$

(B) (C)

Chloroethane Ethyl amine

A $CH_2 = CH_2$ Ethene
B CH_3CH_2Cl Chloroethane
C $CH_3CH_2NH_2$ Ethylamine

53. A hydrocarbon C_3H_6 (A) reacts with HBr to form compound (B). Compound (B) reacts with aqueous potassium hydroxide to give (C) of molecular formula C_3H_6O.what are (A) (B) and (C). Explain the reactions.

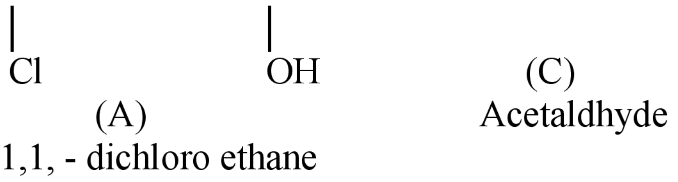
$CH_3 - CH = CH_2 + HBr \rightarrow CH_3 - CH - CH_3$
(A) |
Propene Br
(B)
2 – bromo propane
↓KOH aq
$CH_3 - CH - CH_3$
|
OH
Iso propyl alcohol

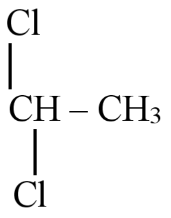
54. Two isomers (A) and (B) have the same molecular formula $C_2H_4Cl_2$.Compound (A) reacts with aqueous KOH gives compound (C) of molecular formula C_2H_4O.Compound (B) reacts with aqueous KOH gives compound (D) of molecular formula $C_2H_6O_2$

Cl
|
$CH - CH_3$ (A) 1,1, - dichloro ethane
|
Cl

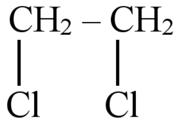
$CH_2 - CH_2 \xrightarrow[aq\ KOH]{} CH_2OH - CH_2OH$
| |
Cl Cl
(B) 1,2, dichloro ethane (D) ETHYLENE GLYCOL

Cl OH
| |
$CH - CH_3 \xrightarrow[aq\ KOH]{} CH - CH_3 \xrightarrow[-H_2O]{} CH_3CHO$



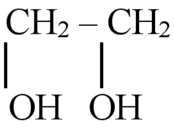
(A) 

$CH(Cl)_2-CH_3$ - 1,1, - dichloro ethane

(B) 

$CH_2(Cl)-CH_2(Cl)$ - 1,2, dichloro ethane

(C) CH_3CHO - Acetaldehyde

(D) 

$CH_2(OH)-CH_2(OH)$ - Ethylene glycol

14. Haloalkanes and Haloarenes

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24				
9	6	9	9	8	8	7	6	6	8				

1. How are organic halogen compounds classified? (226)
2. IUPAC NAME, common name (228)
3. Nature of C-X bond in haloalkane. (229)
4. How are hydrogen halides prepared using Lucas reagent? (230)
5. Mention any two methods of preparation of haloalkanes from alcohols? (230) **J22 2M**
6. What is Darzen's halogenation? (231)
7. Write note on Finkelstein reaction or (How will you prepare n propyl Iodide from n propyl bromide?) (231)
8. Chlorination of methane. (231)
9. Write note on Swartz reaction (231) **S21 2M**
10. Why chlorination of methane is not possible in dark? (231)
15. What is Hunsdiccker reaction? (231)
16. Physical properties. (Boiling point and melting point, solubility, density) 232
17. What is ammonolysis? (233)
18. How does haloalkanes reacts with silver nitrite? (234)
19. What is Williamson ether synthesis? (234) **M19 3M**
20. Explain S_N2 mechanism. (234) **M24 3M**
21. Explain S_N1 mechanism. (234)

22. Explain E2 mechanism. (236)
23. Explain E1 mechanism. (236) (Explain the mechanism involved in the elimination reaction of tertiary butyl chloride with alcoholic KOH. **S20 5M)**
24. How is Grignard reagent prepared? ALL USES (239)
    Starting from CH_3MgI, how will you prepare the following?
    a) Ethyl alcohol b) Acetaldehyde c)ethyl methyl ether **M23 5Mi**
25. What happens when acetyl chloride is treated with excess of CH_3MgI. (239) **J23 2M**
26. Starting from CH_3MgI, how will you prepare the following? (239) **J22 5M**
    (i) Acetaldehyde (ii) Acetone (iii) Methane
27. What happens when acetyl chloride is treated with excess of CH_3MgI. (239) **J24 3M**
27. Write note on sandmeyer reaction (242)
28. What is Balz Schiemann reaction? (242)
29. Raschig process.(242)
30. Among the following compounds, o-dichloro benzene and p-dichloro benzene, which has higher melting point? Explain with reason.(243) **J19 5Mii**
31. What is Dow's process? (243)
32. Write note on Wurtz Fittig reaction (244)
33. Write Fittig reaction. (How does chlorobenzene react with sodium in the presence of ether?) (244)
34. Complete the following reactions. (243,244) **J24 2M COMPULSORY**
    i) $C_6H_5Cl + 2NH_3 \xrightarrow{THF}$
    ii) $CCl_4 + H_2O \xrightarrow{\Delta}$
35. Complete the following reactions. **M22 2M COMPULSORY**
    iii) $C_6H_5Cl + Mg \xrightarrow{250°C\ 50atm}$
    iv) $C_6H_5Cl + 2Na + ClC_6H_5 \xrightarrow{Ether}$
36. Discuss the aromatic nucleophilic substitution reaction of chlorobenzene (243)
37. What are the uses of chloro benzene? (244)
38. A simple aromatic hydrocarbon (A) reacts with chlorine to give Compound (B). Compound (B) reacts with ammonia to give Compound (C) which undergoes carbylamine reaction. Identify (A), (B) and (C) and explain the reactions. (242, 243) **J20 5Mii**
39. How will you prepare gem dihalides? (245)
40. Which reaction is used to distinguish gem dihalides and Vic-Dihalides? (246)
41. What is dehalogenation? (246)
42. What is dehydrogenation? (246)
43. How will you prepare chloroform? (247)
44. How is phosgene prepared from chloroform? (248)
45. How will you prepare chloropicrin? (248)
46. What is carbylamine reaction? (248)
47. What happens when chloroform reacts with oxygen in the presence of sunlight? (248)

48. What are Freons? Discuss their uses (250)

49. How is DDT prepared? (250)

50. Give the structure and uses of DDT? (251) **S20 2M, J23 5Mii, M24 5Mii (uses)**

51. Explain the preparation of the following compounds BB
i) DDT ii) Chloroform iii) Biphenyl iv) Chloropicrin v) Freon-12

52. An organic compound (A) with molecular formula C_2H_5Cl reacts with KOH gives compounds (B) and with alcoholic KOH gives compound (C). Identify (A),(B), and (C). BB **M23 3M**

53. Simplest alkene (A) reacts with HCl to form compound (B).Compound (B) reacts with ammonia to form compound (C) of molecular formula C_2H_7N.Compound (C) undergoes carbylamine test. Identify (A), (B), and (C). BB **S21 5M, M22 5M**

54. A hydrocarbon C_3H_6 (A) reacts with HBr to form compound (B). Compound (B) reacts with aqueous potassium hydroxide to give (C) of molecular formula C_3H_6O.what are (A) (B) and (C). Explain the reactions. BB

55. Two isomers (A) and (B) have the same molecular formula $C_2H_4Cl_2$. Compound (A) reacts with aqueous KOH gives compound (C) of molecular formula C_2H_4O. Compound (B) reacts with aqueous KOH gives compound (D) of molecular formula $C_2H_6O_2$. Identify (A),(B),(C) and (D).

Unit 15: Environmental Chemistry

18. Dissolved oxygen in water is responsible for aquatic life. What processes are responsible for the reduction in dissolved oxygen in water?

- The oxygen is dissolved in water either from atmosphere or by photosynthesis, however at night, there is no photosynthesis, so amount of oxygen dissolved in water reduces.
- Thus, water contains only a limited amount of dissolved oxygen.
- Thus, the decomposition of a moderate amount of organic matter by aerobic bacteria can make water deficient in dissolved oxygen.

19. What would happen, if the greenhouse gases were totally missing in the earth's atmosphere?

In the absence of greenhouse gases, the average temperature of earth will decrease drastically earth average temperature would be bear of on earth would be impossible.

20. Define smog.

Smog is a combination of smoke and fog which forms droplets that remain suspended in the air.

21. Which is considered to be earth's protective umbrella? Why?

- **Ozone layer** in the upper atmosphere is considered to be earth's protective umbrella.

- It protects us from harmful effect such as skin cancer.
- The ozone layer acts as a fitter for the shorten wavelength and highly hazardous ultraviolet radiation from the sun, protecting life on earth.

22. What are degradable and non-degradable pollutants?

Bio-degradable pollutants: The pollutants which can be easily decomposed by the natural biological processes are called bio-degradable pollutants. ***Examples:*** plant wastes, animal wastes….

Non- degradable pollutants: The pollutants which cannot be decomposed by the natural biological processes are called Non bio-degradable pollutants.

Examples: metal wastes (mainly Hg and Pb), D.D.T, plastic, nuclear wastes etc.

23. From where does ozone come in the photo chemical smog?

In the **earth's lower atmosphere** ozone is formed when pollutants emitted by cars, power plants, industrial boilers, refineries etc chemically react in the presence of sunlight. It is formed by reaction between oxides of nitrogen and volatile organic compounds. (Or)

$$NO_2 \xrightarrow{Sun\ light} NO + (O)$$

O_3 are strong oxidizing agent and can react with unburnt hydrocarbons in polluted air to form formaldehyde, acrolein and peroxy acetyl nitrate(PAN).

24. A person was using water supplied by corporation. Due to shortage of water he started using underground water. He felt laxative effect. What could be the cause?

Excessive concentration (> 500ppm) of sulphates in drinking water causes laxative effect.

25. What is green chemistry?

Green chemistry means science of environmentally favorable chemical synthesis.

Green chemistry is a chemical philosophy encouraging the design of products and processes that reduce or eliminate the use and generation of hazardous substances.

26. Explain how does greenhouse effect cause global warming.

Greenhouse effect may be defined as "***the heating up of the earth surface due to trapping of infrared radiations reflected by earth's surface by CO_2 layer in the in the atmosphere***".

- The heating up of earth through the greenhouse effect is called global warming.

27. Mention the standards prescribed by BIS for quality of drinking water

Standard characteristics prescribed for deciding the quality of drinking water by BIS

S.No	Characteristics	Desirable limit
I	**Physico-chemical Characteristics**	
i)	P^H	6.5 to 8.5
ii)	Total Dissolved Solids(TDS)	500 ppm
iii)	Total Hardness (as $CaCO_3$)	300 ppm
iv)	Nitrate	45ppm
v)	Chloride	250 ppm
vi)	Sulphate	200 ppm
vii)	Fluoride	1 ppm
II	**Biological Characteristics**	
i)	Escherichia Coli (E.Coli)	Not at all
ii)	Coliforms	Not to exceed 10 (In 100 ml water sample)

28. How does classical smog differ from photochemical smog?

	Classical smog **(London smog)**	**Photochemical smog** **(Los Angel smog)**
1.	It occurs in ***cool humid*** climate.	It occurs in ***warm*** and dry climate.
2.	It consists of **coal smoke and fog.** The Chemical composition in the mixture of ***SO_2, SO_3 and humidity.***	It is formed by the combination of **smoke, fog, dust and air pollutants** like ***oxides of nitrogen and hydrocarbons*** in the presence of sunlight.
3.	It is reducing in nature due to high concentration SO_2 and called **reducing smog.**	It is oxidizing in nature due to high concentration of oxidizing agents like NO_2 and O_3 and is called **oxidizing smog**.
4.	Classical smog is responsible for ***acid rain*** and causes ***bronchial irritation***.	Photo chemical smog causes **irritation of eyes, skin** and lungs, increase in chances of asthma.

29. What are particulate pollutants? Explain any three.

Particulate pollutants are small solid particles and liquid droplets suspended in air.

(i) Smoke:

Smoke particulate consists of solid particles (or) mixture of solid and liquid particles formed by combustion of organic matter. **For example**, cigarette smoke, oil smoke, smokes from burning of fossil fuel, garbage and dry leaves.

(ii) Dust:

Dust composed of fine solid particles produced during crushing and grinding of solid materials.

For example, sand from sand blasting, saw dust from wood works, cement dust from cement factories and fly ash from power generating units.

(iii) Mists:

They are formed by particles of spray liquids and condensation of vapours in air. **For example**, sulphuric acid mist, herbicides and insecticides sprays can form mists.

30. Even though the use of pesticides increases the crop production, they adversely affect the living organisms. Explain the function and the adverse effects of the pesticides.

Pesticides are the chemicals that are used to kill or stop the growth of unwanted organisms. But these pesticides can affect the health of human beings.

These are further classified as

- **Insecticides:**

It like DDT, BHC, aldrin etc., can stay in soil for long period of time and are absorbed by soil. They contaminate root crops like carrot, raddish etc.,

- **Fungicide:** Organo mercury compounds are used as most common fungicide. They dissociate in soil to produce mercury which is highly toxic.
- **Herbicides:** Herbicides are the chemical compounds used to control unwanted plants. They are otherwise known as weed killers.

31. Ethane burns completely in air to give CO_2, while in a limited supply of air gives CO. The same gases are found in automobile exhaust. Both CO and CO_2 are atmospheric pollutants i) What is the danger associated with these gases

ii) How do the pollutants affect the human body?

(i)What is the danger associated with these gases

(a) Carbon monoxide:

Carbon monoxide is a poisonous gas produced as a result of incomplete combustion of coal are firewood. It is released into the air mainly by automobile exhaust. It binds with haemoglobin and form carboxy haemoglobin which impairs normal oxygen transport by blood and hence the oxygen carrying capacity of blood is reduced. This oxygen deficiency results in headache, dizziness, tension, Loss of consciousness, blurring of eye sight and cardiac arrest.

(b) Carbon dioxide:

Carbon dioxide is released into the atmosphere mainly by the process of respiration, burning of fossil fuels, forest fire, decomposition of limestone in cement industry etc.

Green plants can convert CO_2 gas in the atmosphere into carbohydrate and oxygen through a process called photosynthesis. The increased CO_2 level in the atmosphere is responsible for global warming. It causes headache and nausea.

ii) a) How do the pollutants affect the human body?

i) Particulate pollutants bigger than 5 microns are likely to settle in the nasal passage whereas particles of about 10 micron enters the lungs easily and causes scaring or fibrosis of lung lining. They irritate the lungs and causes cancer and asthma

ii) Lead particulates affect children's brain, interferes maturation of RBCs and even cause cancer

iii) Acid rain causes respiratory ailment in humans and animals.

iv) Hydrocarbons are potential cancer causing (carcinogenic) agents

v) The increased CO_2 level in the atmosphere causes headache and nausea

vi) Carbon monoxide reduces the oxygen carrying capacity of blood. This oxygen deficiency results in headache, dizziness, tension, Loss of consciousness, blurring of eye sight and cardiac arrest.

b) Harmful effects of water pollutants:

i) Cadmium and mercury can cause kidney damage.

ii) Lead poisoning can leads to the severe damage of kidneys, liver, brain etc. It also effects central nervous system

iii) Polychlorinated biphenyls (PCBs) causes skin diseases and are carcinogenic in nature.

Harmful effects of soil pollutants:

Pesticides used in soil for food cultivation can affect the health of human beings.

32. On the basis of chemical reactions involved, explain how do CFC's cause depletion of ozone layer in stratosphere?

In the presence of uv radiation, CFC's break up into chlorine free radical

$CF_2Cl_2 \xrightarrow{h\upsilon} CF_2Cl + Cl$

$CFCl_3 \xrightarrow{h\upsilon} CFCl_2 + Cl$

$Cl + O_3 \rightarrow ClO + O_2$

$ClO + O \rightarrow Cl + O_2$

Chlorine radical is regenerated in the course of reaction. Due to this continuous attack of Cl thinning of ozone layer takes place which leads to formation of ozone hole.

33. How is acid rain formed? Explain its effect.

Burning of fossil fuels (coal and oil) in power stations, furnaces and petrol, diesel in motor engines produce sulphur dioxide and nitrogen oxides. The main contributors of acid rain are SO_2 and NO_2. They are converted into sulphuric acid and nitric acid respectively by the reaction with oxygen and water.

$2SO_2 + O_2 + 2H_2O \rightarrow 2H_2SO_4$

$4NO_2 + O_2 + 2H_2O \rightarrow 4HNO_3$

The pH of rain water drops to 5.6 and hence its is called acid rain

Harmful effects of acid rain:

i)Acid rain causes extensive damage to buildings and structural materials of marbles. This attack on marble is termed as Stone leprosy. $CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2\uparrow$

ii)Acid rain affects plants and animal life in aquatic ecosystem.

iii)It is harmful for agriculture, trees and plants as it dissolves and removes the nutrients needed for their growth.

iv)It corrodes water pipes resulting in the leaching of heavy metals such as iron, lead and copper into the drinking water which have toxic effects.

v)It causes respiratory ailment in humans and animals.

34. Differentiate the following (i) BOD and COD (ii) Viable and non-viable particulate pollutants

Biochemical oxygen demand (COD	Chemical oxygen demand (COD
The total amount of oxygen in milligrams consumed by microorganisms in decomposing the waste in one litre of water at 20°C for a period of 5 days is called biochemical oxygen demand (BOD)	It is defined as the amount of oxygen required by the organic matter in a sample of water for its oxidation by a strong oxidizing agent like $K_2Cr_2O_7$ in acid medium for a period of 2 hrs.
Its value is expressed in ppm.	Its value is expressed in mg/litre

BOD is used as a measure of degree of water pollution.	COD is a measure of amount of organic compounds in a water sample.
BOD is only a measurement of consumed oxygen by microorganisms to decompose the organic matter.	COD refers to the requirement of dissolved oxygen for both oxidation of organic and inorganic constituents.
Clean water would have BOD value less than 5 ppm.	Clean water would have COD value greater than 250 mg/litre.

viable particulate pollutants	**non-viable particulate pollutants**
The viable particulates are the small size living organisms such as bacteria, fungi, moulds, algae, etc. Which are dispersed in air.	The non-viable particulates are small solid particles and liquid droplets suspended in air. **Example:** Smoke, dust, mists, fumes etc.
They are all organic particulates.	They are all inorganic particulates.
They contain living organisms.	They contain non-living organisms.
Eg: bacteria, fungi, moulds, algae, etc.	**Example:** Smoke, dust, mists, fumes etc.

35. Explain how oxygen deficiency is caused by carbon monoxide in our blood? Give its effect

Carbon monoxide is a poisonous gas produced as a result of incomplete combustion of coal are firewood. It is released into the air mainly by automobile exhaust.

- Carbon dioxide binds with hemoglobin and form carboxy hemoglobin which impairs normal oxygen transport by blood and hence the oxygen carrying capacity of blood is reduced.
- This oxygen deficiency results in headache, dizziness, tension, loss of consciousness, blurring of eye sight and cardiac arrest.

36. What are the various methods you suggest to protect our environment from pollution?

Strategies to control environment pollution.

- ***Waste Management:*** *Environmental pollution can be controlled by proper disposal of wastes.*
- *Recycling reused by recycling the waste material can be reused by recycling the waste, thus it reduces the land fill and converts waste into useful forms.*
- *Substitution of less toxic solvents for highly toxic ones used in certain industrial processes.*

- *Use of fuels with lower sulphur content (e.g., washed coal)*
- *Growing more trees.*
- *Control measures in vehicle emissions are adequate.*

15. Environmental Chemistry

M19	J19	S20	S21	M22	J22	M23	J23	M24	J24			
3	3	7	Corona	corona	corona	5	9	8	7			

1. What is environmental chemistry? (260)
2. What is environmental pollution? (260)
3. What are biodegradable and non-biodegradable pollutants? (260)
4. Differentiate viable and non-viable particulate pollutants. **J24 5Mii**
5. What is greenhouse effect? Name the gases that cause greenhouse effect. (263) **S20 3M, M24 2M**
6. What is global warming? (263)
7. What is acid rain? **M22 2M** What are harmful effects of acid rain? (264) **M24 5M**
8. What are particulate pollutants? Give example (265) **M23 2M**
9. What are the health effects of particulate pollutants. (265)
10. What is smog? (266)
11. What is classical smog or London smog? (266)
12. What are the effects of classical smog? (267)
13. What is photochemical smog or Los Angel Smog? (267)
14. What are the effects of photochemical smog? (268)
15. What is depletion of ozone layer or ozone hole? (268)
16. What is the environmental impact of ozone depletion? (269)
17. Which is considered to be earth's protective Umberlla? Why? (269) **J24 3M**
18. What is meant by water pollution? What are the causes, harmful effects of chemical water pollution. (271)
19. Define Eutrophication (271) **M23 5Mii**
20. What is Biochemical oxygen demand (BOD)? (271)
21. What is chemical oxygen demand (COD)? (271)
22. Differentiate BOD and COD. (271) **J23 5Mi**
23. What are the harmful effects of chemical water pollutants? (271)
24. Mention the standards prescribed by BIS for qualities of drinking water. (272)
25. What are the total dissolved solids (TDS)? (273)
26. What are the sources of soil pollution. (273)
27. Explain the strategies to control environmental pollution (274) **S20 5Mii**
28. What is Green Chemistry? **J23 5Mii**, **J24 5Mii** Green chemistry in day-to-day life. (275)

NAME REACTIONS – ORGANIC CHEMISTRY

1. Sabatier sendersens reaction.(184)
2. Kolbe's electrolytic method. (184)
3. Wurtz reaction (184)
4. Corey house reaction. (185)
5. Grignard reagent preparation. (185
6. Aromatization (189)
7. Pyrolysis (189)
8. Markonikoff's rule. (194 & 231)
9. Anti –markonvnikoff's rule or peroxide effect or kharasch addition. (196)
10. Action of bayer reagent(197)
11. Ozonolysis (198)
12. Polymerization (199)
13. Wurtz- fittig reaction (210 & 244)
14. Friedel crafts reaction (210)
15. Friedel crafts reaction acylation (212)
16. BHC (215) **M24 2M**
17. Brich reduction(215)
18. Lucas test (230)
19. Dorzens halogenation reaction
20. Finkelstein reaction (231)
21. Swarts reaction (231 & 249) **S21 2M**
22. Hunsidiccker reaction (231)
23. Williamson ether synthesis. (234)
24. TEL (238)
25. Sandmeyer reaction. (242)
26. Gattermann reaction (242)
27. Balz schiemann reaction (242)
28. Raschig process. (242)
29. Dow process. (243)
30. Fittig reaction(244)
31. Haloform reaction (247)
32. Carbylamine reaction. (248)
33. Chloropicrin (248)
34. Freons (249)
35. DDT (250)

1. An organic compound (A) C_2H_4 decolourises bromine water.(A) on reaction with chlorine gives (B). (A) reacts with HBr to give (C). Identify (A) , (B) and (C).Explain the reactions. (194) **J22 5M**
2. An organic compound (A) of molecular formula C_2H_6O, on heating with conc. H_2SO_4 gives compound (B). (B) on treating with cold dilute alkaline $KMnO_4$ gives compound (C). Identify (A), (B) and (C) and explain the reactions. (197) **J20 5Mi**
3. The simple Aromatic Hydrocarbon compound (A) reacts with Bromine to give (B). Compound (A) reacts with Raney Ni and gives (C). Identify (A), (B) and (C). (211,215) **M19 5Mii**
4. $CH_3CH_2Cl \xrightarrow{Conc H2SO4\ 430-440K} A \xrightarrow{HBr\ Benzoyl\ peroxide} B$ 192, 195 **J23 2MCOM**
5. A simple aromatic hydrocarbon (A) reacts with chlorine to give Compound (B). Compound (B) reacts with ammonia to give Compound (C) which undergoes carbylamine reaction. Identify (A), (B) and (C) and explain the reactions. (242, 243) **J20 5Mii**
6. An organic compound (A) with molecular formula C_2H_5Cl reacts with KOH gives compounds (B) and with alcoholic KOH gives compound (C). Identify (A),(B), and (C). BB **M23 3M**
7. Simplest alkene (A) reacts with HCl to form compound (B).Compound (B) reacts with ammonia to form compound (C) of molecular formula C_2H_7N.Compound (C) undergoes carbylamine test. Identify (A), (B), and (C). BB **M22 5M**
8. A hydrocarbon C_3H_6 (A) reacts with HBr to form compound (B). Compound (B) reacts with aqueous potassium hydroxide to give (C) of molecular formula C_3H_6O.what are (A) (B) and (C). Explain the reactions. BB
9. Two isomers (A) and (B) have the same molecular formula $C_2H_4Cl_2$. Compound (A) reacts with aqueous KOH gives compound (C) of molecular formula C_2H_4O. Compound (B) reacts with aqueous KOH gives compound (D) of molecular formula $C_2H_6O_2$. Identify (A),(B),(C) and (D).
10. Complete the following reactions. (243,244) **M22 2M COMPULSORY**

    i) $C_6H_5Cl + 2NH_3 \xrightarrow{250°C\ 50atm}$

    ii) $C_6H_5Cl + 2Na + ClC_6H_5 \xrightarrow{Ether}$

“NO PAIN, NO GAIN”.

Never Dreamed about success, Worked for it.

WISH U ALL THE BEST ACTC

“May God's guidance be with you during the Exam and may you be able to answer each question correctly. My prayers and Blessings are with you”.
- ACTC EMS

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