

# 1 Some Basic Concepts in Chemistry

**Question:** The density of 1 M solution of a compound 'X' is  $1.25 \text{ g mL}^{-1}$ . The correct option for the molality of solution is (Molar mass of compound X = 85 g):

NEET 2023 Manipur

A 0.705 m

B 1.208 m

C 1.165 m

D 0.858 m

**Answer: D**

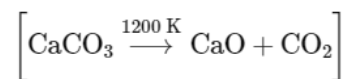
**Explanation**

$$\begin{aligned} m &= \frac{1000 \times M}{1000 \times d - MM_w} \\ m &= \frac{1000 \times 1}{1000 \times 1.25 - 1 \times 85} \\ m &= \frac{1000}{1165} = 0.858 \end{aligned}$$

**Question:**

The right option for the mass of  $\text{CO}_2$  produced by heating 20 g of 20% pure limestone is

(Atomic mass of  $\text{Ca} = 40$ )



## NEET 2023

**A** 1.76 g

**B** 2.64 g

**C** 1.32 g

**D** 1.12 g

**Answer: A**

## Explanation

Weight of impure limestone = 20 g

Weight of pure limestone ( $\text{CaCO}_3$ ) = 20% of 20 g

$$= \frac{20}{100} \times 20$$

$$= 4 \text{ g}$$

$$n_{\text{CaCO}_3} = \frac{4}{100} = 0.04$$

Image

$$n_{\text{CO}_2} = 0.04$$

$$W_{\text{CO}_2} = 0.04 \times 44$$

$$= 1.76 \text{ g}$$

2022

## MCQ (Single Correct Answer)

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**Q.1.** The density of the solution is  $2.15 \text{ g mL}^{-1}$ , then mass of  $2.5 \text{ mL}$  solution in correct significant figures is :

**A**  $53.75 \text{ g}$

**B**  $5375 \times 10^{-3} \text{ g}$

**C**  $5.4 \text{ g}$

**D**  $5.38 \text{ g}$

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**Ans. (C)**

### Explanation

Mass = Volume  $\times$  Density

$$= 2.5 \times 2.15$$

$$= 5.375 \text{ g}$$

Since 2.5 has two significant figures, so the mass of solution in correct significant figures will be  $5.4 \text{ g}$ .

**Q.2.** What fraction of Fe exists as Fe(III) in  $\text{Fe}_{0.96}\text{O}$  ?

(Consider  $\text{Fe}_{0.96}\text{O}$  to be made up of Fe(II) and Fe(III) only)

A  $\frac{1}{20}$

B  $\frac{1}{12}$

C 0.08

D  $\frac{1}{16}$

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Ans. (B)

Explanation



Let Fe(II) present in  $\text{Fe}_{0.96}\text{O} = x$

Fe(III) present =  $(0.96 - x)$

Total charge on Fe =  $2x + (0.96 - x)3$

Total charge on O =  $-2$

$$2x + (0.96 - x)3 = 2$$

$$2x + 2.88 - 3x = 2$$

$$-x = -0.88$$

$$x = 0.88$$

$$\text{Fe}^{2+} = 0.88, \text{Fe}^{3+} = 0.08$$

$$\text{Fraction of Fe}^{3+} = \frac{0.08}{0.96} = \frac{1}{12}$$

**Q.3.** In one molal solution that contains 0.5 mole of a solute, there is

**A** 500 mL of solvent

**B** 500 g of solvent

**C** 100 mL of solvent

**D** 1000 g of solvent

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**Ans. (B)**

### Explanation

Molality is the moles of solute dissolved per kg of solvent therefore 500 g, 1 molal solution contains 0.5 of solute, as

$$m = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$$

$$1 = \frac{0.5}{\text{Mass of solvent (in kg)}}$$

$$\therefore \text{Mass of solvent (in kg)} = 0.5$$

$$= 500 \text{ g}$$

**Q.4.** What mass of 95% pure  $\text{CaCO}_3$  will be required to neutralise 50 mL of 0.5 M HCl solution according to the following reaction?



[Calculate upto second place of decimal point]

**A** 1.25 g

**B** 1.32 g

**C** 3.65 g

**D** 9.50 g

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**Ans. (B)**

**Explanation**

Let  $m$  gram mass of  $\text{CaCO}_3$  is required

$$\text{Pure CaCO}_3 \text{ in } m \text{ gram} = \frac{95}{100} \times m$$

$$\text{Moles of CaCO}_3 = \frac{95}{100} \times \frac{m}{100}$$

$$\text{Moles of HCl required} = 2 \times \text{moles of CaCO}_3$$

$$= 2 \times \frac{95}{100} \times \frac{m}{100}$$

$$2 \times \frac{95}{100} \times \frac{m}{100} = \frac{50}{1000} \times 0.5$$

$$m = 1.315 \text{ g} \approx 1.32 \text{ g}$$

## TOPIC 1

### Nature of Matter, Significant Figures and Laws of Chemical Combinations

- 01** The number of significant figures for the three numbers 161 cm, 0.161 cm, 0.0161 cm are

[CBSE AIPMT 1998]

- (a) 3, 4 and 5 respectively  
(b) 3, 4 and 4 respectively  
(c) 3, 3 and 4 respectively  
(d) 3, 3 and 3 respectively

**Ans. (d)**

- (i) All non-zero digits are significant.  
(ii) Non-zero digits to the right of the decimal point are significant.  
(iii) Zeroes to the left of the first non-zero digit in a number are not significant.

So, the number of significant figures for the numbers 161 cm, 0.161 cm and 0.0161 cm are same, i.e. 3.

- 02** 0.24 g of a volatile gas, upon vaporisation, gives 45 mL vapour at NTP. What will be the vapour density of the substance? (Density of  $H_2 = 0.089$ )

[CBSE AIPMT 1996]

- (a) 95.93 (b) 59.93  
(c) 95.39 (d) 5.993

**Ans. (b)**

Weight of gas = 0.24 g

Volume of gas ( $V$ ) = 45 mL = 0.045 L

Density of  $H_2$  ( $d$ ) = 0.089

$$\begin{aligned}\text{Weight of 45 mL of } H_2 &= V \times d \\ &= 0.045 \times 0.089\end{aligned}$$

$$= 4.005 \times 10^{-3} \text{ g}$$

Therefore, vapour density

$$\begin{aligned}&= \frac{\text{Weight of certain volume of substance}}{\text{Weight of same volume of hydrogen}} \\ &= \frac{0.24}{4.005 \times 10^{-3}} = 59.93\end{aligned}$$

- 03** In the final answer of the expression

$$\frac{(29.2 - 20.2)(1.79 \times 10^5)}{1.37}$$

the number of significant figures is

[CBSE AIPMT 1994]

- (a) 1 (b) 2  
(c) 3 (d) 4

**Ans. (c)**

On calculation we find

$$\frac{(29.2 - 20.2)(1.79 \times 10^5)}{1.37} = 1.17 \times 10^6$$

As the least precise number contains 3 significant figures, therefore answer should also contain 3 significant figures.

- 04** The molecular weight of  $O_2$  and  $SO_2$  are 32 and 64 respectively. At  $15^\circ\text{C}$  and 150 mmHg pressure, 1 L of  $O_2$  contains 'N' molecules. The number of molecules in 2L of  $SO_2$  under the same conditions of temperature and pressure will be

[CBSE AIPMT 1990]

- (a)  $N/2$  (b)  $N$   
(c)  $2N$  (d)  $4N$

**Ans. (c)**

According to Avogadro's law "equal volumes of all gases contain equal number of molecules under similar conditions of temperature and pressure." Thus, if 1 L of one gas contains  $N$  molecules, 2 L of any gas under similar conditions will contain  $2N$  molecules.

## TOPIC 2

### Atomic Mass, Molecular Mass and Formulae of Compounds

- 05** An organic compound contains 78% (by wt.) carbon and remaining percentage of hydrogen. The right option for the empirical formula of this compound is [At. wt. of C is 12, H is 1]

[NEET 2021]

- (a)  $CH$  (b)  $CH_2$   
(c)  $CH_3$  (d)  $CH_4$

**Ans. (c)**

Element	%	Atomic mass	Relative number of moles	Simple ratio of moles	Simplest whole number ratio
C	78	12	$\frac{78}{12} = 6.5$	$\frac{6.5}{6.5} = 1$	1
H	22	1	$\frac{22}{1} = 22$	$\frac{22}{6.5} = 3.3$	3

The empirical formula of the organic compound is  $CH_3$ .

- 06** The number of protons, neutrons and electrons in  ${}^{175}_{71}\text{Lu}$ , respectively, are [NEET (Sep.) 2020]

(a) 104, 71 and 71 (b) 71, 71 and 104  
(c) 175, 104 and 71 (d) 71, 104 and 71

**Ans. (d)**

In  ${}^{175}_{71}\text{Lu}$ ,

Mass number ( $A$ ) = 175 =  $n + p$

Atomic number ( $Z$ ) = 71 =  $p = e^-$

$\therefore$  Number of protons = 71

Number of neutrons

$$= A - Z = 175 - 71 = 104$$

Number of electrons = 71

- 07** Suppose the elements X and Y combine to form two compounds  $\text{XY}_2$  and  $\text{X}_3\text{Y}_2$ . When 0.1 mole of  $\text{XY}_2$  weighs 10 g and 0.05 mole of  $\text{X}_3\text{Y}_2$  weighs 9 g, the atomic weights of X and Y are

[NEET Phase II 2016]

(a) 40, 30 (b) 60, 40 (c) 20, 30 (d) 30, 20

**Ans. (a)**

Let atomic masses of X and Y be  $A_X$  and  $A_Y$ , respectively

$$\text{For } \text{XY}_2, n_{\text{XY}_2} = 0.1 = \frac{10}{A_X + 2A_Y}$$

$$\text{or } A_X + 2A_Y = 100 \quad \dots(i)$$

$$\text{For } \text{X}_3\text{Y}_2, n_{\text{X}_3\text{Y}_2} = 0.05 = \frac{9}{3A_X + 2A_Y}$$

$$\text{or } 3A_X + 2A_Y = 180 \quad \dots(ii)$$

On solving Eqs. (i) and (ii), we get,

$$A_X = 40 \text{ g mol}^{-1} \Rightarrow A_Y = 30 \text{ g mol}^{-1}$$

- 08** An organic compound contains carbon, hydrogen and oxygen. Its elemental analysis gave C, 38.71% and H, 9.67%. The empirical formula of the compound would be

[CBSE AIPMT 2008]

(a)  $\text{CH}_3\text{O}$  (b)  $\text{CH}_2\text{O}$  (c)  $\text{CHO}$  (d)  $\text{CH}_4\text{O}$

**Ans. (a)**

Element	% abundance	At. wt.	Molar ratio	Simple ratio
C	38.71	12	$\frac{38.71}{12} = 3.23$	$\frac{3.23}{3.23} = 1$
H	9.67	1	$\frac{9.67}{1} = 9.67$	$\frac{9.67}{3.23} = 3$
O	$[100 - (38.71 + 9.67)]$ = 51.62	16	$\frac{51.62}{16} = 3.23$	$\frac{3.23}{3.23} = 1$

Thus, the empirical formula of the compound is  $\text{CH}_3\text{O}$ .

- 09** An element, X has the following isotopic composition:

${}^{200}\text{X} : 90\%$ ,  ${}^{199}\text{X} : 8.0\%$ ,

${}^{202}\text{X} : 2.0\%$

The weighted average atomic mass of the naturally occurring element X is closest to

[CBSE AIPMT 2007]

(a) 201 u

(b) 202 u

(c) 199 u

(d) 200 u

**Ans. (d)**

Weight of  ${}^{200}\text{X} = 0.90 \times 200 = 180.00 \text{ u}$

Weight of  ${}^{199}\text{X} = 0.08 \times 199 = 15.92 \text{ u}$

Weight of  ${}^{202}\text{X} = 0.02 \times 202 = 4.04 \text{ u}$

Total weight =  $199.96 \approx 200 \text{ u}$

- 10** Which of the following is

isoelectronic? [CBSE AIPMT 2002]

(a)  $\text{CO}_2$ ,  $\text{NO}_2$

(b)  $\text{NO}_2^-$ ,  $\text{CO}_2$

(c)  $\text{CN}^-$ ,  $\text{CO}$

(d)  $\text{SO}_2$ ,  $\text{CO}_2$

**Ans. (c)**

$\text{CN}^-$  and  $\text{CO}$  are isoelectronic because they have equal number of electrons.

In  $\text{CN}^-$  the number of electrons  
=  $6 + 7 + 1 = 14$

In  $\text{CO}$  the number of electrons  
=  $6 + 8 = 14$

- 11** An organic compound containing C, H and N gave the following results on analysis C = 40%, H = 13.33%, N = 46.67%. Its empirical formula would be

[CBSE AIPMT 2002, 1999, 98]

(a)  $\text{C}_2\text{H}_7\text{N}_2$

(b)  $\text{CH}_5\text{N}$

(c)  $\text{CH}_4\text{N}$

(d)  $\text{C}_2\text{H}_7\text{N}$

**Ans. (c)**

Table for empirical formula

Element	%	At. wt.	Molar ratio	Simple ratio
C	40.00	12	$\frac{40}{12} = 3.33$	$\frac{3.33}{3.33} = 1$
H	13.33	1	$\frac{13.33}{1} = 13.33$	$\frac{13.33}{3.33} = 4$
N	46.67	14	$\frac{46.67}{14} = 3.33$	$\frac{3.33}{3.33} = 1$

Hence, empirical formula is  $\text{CH}_4\text{N}$ .

- 12** An organic compound contains C = 40%, O = 53.34% and H = 6.60%. The empirical formula of the compound is

[CBSE AIPMT 1994]

(a)  $\text{CH}_2\text{O}$

(b)  $\text{CHO}$

(c)  $\text{CH}_4\text{O}_2$

(d)  $\text{C}_2\text{H}_2\text{O}$

**Ans. (a)**

Element	%	At. wt.	Molar ratio	Simple ratio
C	40	12	$\frac{40}{12} = 3.33$	$\frac{3.33}{3.33} = 1$
H	6.60	1	$\frac{6.60}{1} = 6.60$	$\frac{6.60}{3.33} = 2$
O	$\frac{53.34}{4}$	16	$\frac{53.34}{16} = 3.33$	$\frac{3.33}{3.33} = 1$

Hence, empirical formula is

C : H : O = 1 : 2 : 1 =  $\text{CH}_2\text{O}$

- 13** Boron has two stable isotopes,  ${}^{10}\text{B}$  (19%) and  ${}^{11}\text{B}$  (81%). Calculate average atomic weight of boron in the periodic table.

[CBSE AIPMT 1990]

(a) 10.8

(b) 10.2

(c) 11.2

(d) 10.0

**Ans. (a)**

Average of atomic weight

% of  ${}^{10}\text{B} \times \text{atomic mass of } {}^{10}\text{B} + \%$  of  ${}^{11}\text{B}$

$$= \frac{\times \text{atomic mass of } {}^{11}\text{B}}{\% \text{ of } {}^{10}\text{B} + \% \text{ of } {}^{11}\text{B}}$$

$$= \frac{19 \times 10 + 81 \times 11}{19 + 81}$$

$$= \frac{190 + 891}{100} = 10.81$$

- 14** While extracting an element from its ore, the ore is grind and leached with dil. KCN solution to form the soluble product potassium argento- cyanide. The element is

[CBSE AIPMT 1989]

(a) lead

(b) chromium

(c) manganese

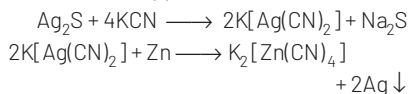
(d) silver

**Ans. (d)**

Silver metal is extracted from the argentite ore  $\text{Ag}_2\text{S}$  by cyanide process. In this method, the concentrated ore is treated with dilute solution of potassium



cyanide, then a soluble complex potassium dicyanoargentate(I) is formed which when reacted with zinc, silver is extracted as a ppt.

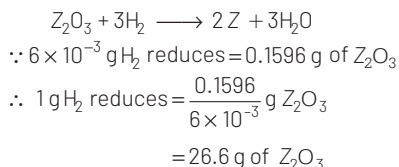


- 15** A metal oxide has the formula  $\text{Z}_2\text{O}_3$ . It can be reduced by hydrogen to give free metal and water. 0.1596 g of the metal oxide requires 6 mg of hydrogen for complete reduction. The atomic weight of the metal is

[CBSE AIPMT 1989]

- (a) 27.9 (b) 159.6  
(c) 79.8 (d) 55.8

**Ans. (d)**



$\therefore$  Equivalent weight of  $\text{Z}_2\text{O}_3 = 26.6$

Equivalent weight of Z + Equivalent weight of O = 26.6

Equivalent weight of Z + 8 = 26.6

Equivalent weight of Z = (26.6 - 8) = 18.6

Valency of metal in  $\text{Z}_2\text{O}_3 = 3$

$$\text{Equivalent weight} = \frac{\text{Atomic weight}}{\text{Valency}}$$

$$\text{Atomic weight} = 18.6 \times 3 = 55.8$$

## TOPIC 3

### Mole Concept and Concentration Terms

- 16** One mole of carbon atom weighs 12 g, the number of atoms in it is equal to, (Mass of carbon -12 is  $1.9926 \times 10^{-23} \text{ g}$ ) [NEET (Oct.) 2020]

- (a)  $1.2 \times 10^{23}$  (b)  $6.022 \times 10^{22}$   
(c)  $12 \times 10^{22}$  (d)  $6.022 \times 10^{23}$

**Ans. (d)**

1 mole of carbon atoms weight 12 g, its contains Avogadro number of carbon atoms, i.e.  $6.022 \times 10^{23}$  number of carbon atoms.

- 17** Which one of the followings has maximum number of atoms?

[NEET (Sep.) 2020]

- (a) 1 g of Mg(s) [Atomic mass of Mg = 24]  
(b) 1 g of  $\text{O}_2$ (g) [Atomic mass of O = 16]  
(c) 1 g of Li(s) [Atomic mass of Li = 7]  
(d) 1 g of Ag(s) [Atomic mass of Ag = 108]

**Ans. (c)**

$$\text{Number of atoms (n)} = \frac{\text{Mass in g (1 g)} \times \text{Atomicity of the molecule}}{\text{Gram molar mass (M)}} \times N_A$$

[ $\therefore N_A = \text{Avogadro's number}$ ]

$$\Rightarrow n \propto \frac{\text{Atomicity}}{M}$$

$$(a) n_{\text{Mg}} = \frac{1}{24}$$

$$(b) n_{\text{O}} = \frac{2}{32} = \frac{1}{16}$$

$$(c) n_{\text{Li}} = \frac{1}{7}$$

$$(d) n_{\text{Ag}} = \frac{1}{108}$$

$$\text{So, } n_{\text{Li}} > n_{\text{O}} > n_{\text{Mg}} > n_{\text{Ag}}$$

- 18** In which case is the number of molecules of water maximum?

[NEET 2018]

- (a) 0.00224 L of water vapours at 1 atm and 273 K  
(b) 0.18 g of water  
(c) 18 mL of water  
(d)  $10^{-3} \text{ mol}$  of water

**Ans. (c)**

Number of molecules = Mole  $\times$  Avogadro's number ( $N_A$ )

The number of molecules of water in each of the given options is calculated as

- (i) 18 mL of water

Number of moles ( $n_{\text{H}_2\text{O}}$ )

$$= \frac{\text{Mass of substance in g (} w_{\text{H}_2\text{O}} \text{)}}{\text{Molar mass in g mol}^{-1} (M_{\text{H}_2\text{O}})}$$

$$w_{\text{H}_2\text{O}} = 18 \text{ g}$$

$$[\therefore \text{Density of water (} d_{\text{H}_2\text{O}} \text{)} = 1 \text{ g L}^{-1}]$$

$$\therefore n_{\text{H}_2\text{O}} = \frac{18}{18} = 1$$

Number of molecules of water =  $1 \times N_A$

- (ii) 0.18 g of water

$$n_{\text{H}_2\text{O}} = \frac{w_{\text{H}_2\text{O}}}{M_{\text{H}_2\text{O}}} = \frac{0.18}{18} = 0.01$$

$$\text{Number of molecules of water} = 0.01 \times N_A$$

- (iii) 0.00224 L of water vapours at 1 atm and 273 K. At STP [1 atm and 273 K], Number of moles [with reference to volume]

$$= \frac{\text{Volume of gas in litres}}{22.4}$$

$$= \frac{0.00224}{22.4} = 0.0001$$

$$\text{Number of molecules of water} = 0.0001 \times N_A$$

- (iv)  $10^{-3} \text{ mol}$  of water

$$\text{Number of molecules of water} = 10^{-3} \times N_A$$

$\therefore$  Among the given options, option (i) contains the maximum number of water molecules.

- 19** If Avogadro number  $N_A$ , is changed from  $6.022 \times 10^{23} \text{ mol}^{-1}$  to  $6.022 \times 10^{20} \text{ mol}^{-1}$  this would change [CBSE AIPMT 2015]

- (a) the definition of mass in units of grams  
(b) the mass of one mole of carbon  
(c) the ratio of chemical species to each other in a balanced equation  
(d) the ratio of elements to each other in a compound

**Ans. (b)**

If Avogadro number  $N_A$ , is changed from  $6.022 \times 10^{23} \text{ mol}^{-1}$  to  $6.022 \times 10^{20} \text{ mol}^{-1}$ , this would change the mass of one mole of carbon.

$\therefore$  1 mole of carbon has mass = 12 g

or  $6.022 \times 10^{23}$  atoms of carbon have mass = 12 g

$\therefore 6.022 \times 10^{20}$  atoms of carbon have mass

$$= \frac{12}{6.022 \times 10^{23}} \times 6.022 \times 10^{20} = 0.012 \text{ g}$$

- 20** How many grams of concentrated nitric acid solution should be used to prepare 250 mL of 2.0 M  $\text{HNO}_3$ ? The concentrated acid is 70%  $\text{HNO}_3$ . [NEET 2013]

- (a) 45.0 g conc.  $\text{HNO}_3$   
(b) 90.0 g conc.  $\text{HNO}_3$   
(c) 70.0 g conc.  $\text{HNO}_3$   
(d) 54.0 g conc.  $\text{HNO}_3$

**Ans. (a)**

Given, molarity of solution = 2

Volume of solution = 250 mL

$$= \frac{250}{1000} = \frac{1}{4} \text{ L}$$

Molar mass of

$$\text{HNO}_3 = 1 + 14 + 3 \times 16 = 63 \text{ g mol}^{-1}$$

$\therefore$  Molarity

$$= \frac{\text{Weight of HNO}_3}{\text{Molecular mass of HNO}_3}$$

$$\times \text{volume of solution (L)}$$

$\therefore$  Weight of

$$\text{HNO}_3 = \text{molarity} \times \text{molecular mass}$$

$$\times \text{volume (L)}$$

$$= 2 \times 63 \times \frac{1}{4} \text{ g}$$

$$= 31.5 \text{ g}$$

It is the weight of 100%  $\text{HNO}_3$

But the given acid is 70%  $\text{HNO}_3$

$$\therefore \text{Its weight} = 31.5 \times \frac{100}{70} \text{ g}$$

$$= 45 \text{ g}$$

- 21**  $6.02 \times 10^{20}$  molecules of urea are present in 100 mL of its solution.

The concentration of solution is

[NEET 2013]

(a) 0.02 M

(b) 0.01 M

(c) 0.001 M

(d) 0.1 M

**Ans. (b)**

Given, number of molecules of urea  
 $= 6.02 \times 10^{20}$

$$\therefore \text{Number of moles} = \frac{6.02 \times 10^{20}}{N_A}$$

$$= \frac{6.02 \times 10^{20}}{6.023 \times 10^{23}}$$

$$= 0.999 \times 10^{-3}$$

$$\approx 1 \times 10^{-3} \text{ mol}$$

Volume of the solution

$$= 100 \text{ mL} = \frac{100}{1000} \text{ L} = 0.1 \text{ L}$$

Concentration of urea solution (in  $\text{mol L}^{-1}$ )

$$= \frac{1 \times 10^{-3}}{0.1} \text{ mol L}^{-1}$$

$$= 1 \times 10^{-2} \text{ mol L}^{-1}$$

$$= 0.01 \text{ mol L}^{-1}$$

- 22** The number of atoms in 0.1 mole of a triatomic gas is

$$(N_A = 6.023 \times 10^{23} \text{ mol}^{-1})$$

[CBSE AIPMT 2010]

(a)  $6.026 \times 10^{22}$

(b)  $1.806 \times 10^{23}$

(c)  $3.600 \times 10^{23}$

(d)  $1.800 \times 10^{22}$

**Ans. (b)**

Number of atoms = number of moles

$$\times N_A \times \text{atomicity}$$

$$= 0.1 \times 6.023 \times 10^{23} \times 3$$

$$= 1.806 \times 10^{23} \text{ atoms}$$

- 23** Volume occupied by one molecule of water (density =  $1 \text{ g cm}^{-3}$ ) is

[CBSE AIPMT 2008]

(a)  $9.0 \times 10^{-23} \text{ cm}^3$

(b)  $6.023 \times 10^{-23} \text{ cm}^3$

(c)  $3.0 \times 10^{-23} \text{ cm}^3$

(d)  $5.5 \times 10^{-23} \text{ cm}^3$

**Ans. (c)**

$$1 \text{ mole} = 6.023 \times 10^{23} \text{ molecule}$$

$$18 \text{ g} = 6.02 \times 10^{23} \text{ molecule}$$

$$18 \text{ g} = \text{mass of } 6.02 \times 10^{23}$$

water molecules

Mass of one water molecule

$$= \frac{18}{6.023 \times 10^{23}} \text{ g}$$

$$\text{Density} = 1 \text{ g cm}^{-3}$$

$$\text{Volume} = \frac{\text{Mass of one water molecule}}{\text{Density}}$$

$$= \frac{18}{6.023 \times 10^{23} \times 1} \text{ cm}^3$$

$$\approx 3.0 \times 10^{-23} \text{ cm}^3$$

- 24** The maximum number of molecules are present in

[CBSE AIPMT 2004]

(a) 15 L of  $\text{H}_2$  gas at STP

(b) 5 L of  $\text{N}_2$  gas at STP

(c) 0.5 g of  $\text{H}_2$  gas

(d) 10 g of  $\text{O}_2$  gas

**Ans. (a)**

In 15 L of  $\text{H}_2$  gas at STP,

the number of molecules

$$= \frac{6.023 \times 10^{23}}{22.4} \times 15$$

$$= 4.033 \times 10^{23}$$

In 5 L of  $\text{N}_2$  gas at STP,

the number of molecules

$$= \frac{6.023 \times 10^{23} \times 5}{22.4}$$

$$= 1.344 \times 10^{23}$$

In 0.5 g of  $\text{H}_2$  gas,

the number of molecules

$$= \frac{6.023 \times 10^{23} \times 0.5}{2}$$

$$= 1.505 \times 10^{23}$$

In 10 g of  $\text{O}_2$  gas,

the number of molecules

$$= \frac{6.023 \times 10^{23} \times 10}{32}$$

$$= 1.882 \times 10^{23}$$

Hence, maximum number of molecules are present in 15 L of  $\text{H}_2$  at STP.

- 25** Percentage of Se in peroxidase anhydrase enzyme is 0.5% by weight (at. weight = 78.4), then minimum molecular weight of peroxidase anhydrase enzyme is

[CBSE AIPMT 2001]

(a)  $1.568 \times 10^3$

(b) 15.68

(c)  $2.168 \times 10^4$

(d)  $1.568 \times 10^4$

**Ans. (d)**

Suppose the molecular weight of enzyme = x

0.5% by weight means in 100 g of enzyme weight of Se = 0.5 g

$$\therefore \text{In } x \text{ g of enzyme weight of Se} = \frac{0.5}{100} \times x$$

$$\text{Hence, } 78.4 = \frac{0.5 \times x}{100}$$

$$\therefore x = 15680$$

$$= 1.568 \times 10^4$$

- 26** The number of atoms in 4.25 g of  $\text{NH}_3$  is approximately

[CBSE AIPMT 1999]

(a)  $4 \times 10^{23}$

(b)  $2 \times 10^{23}$

(c)  $1 \times 10^{23}$

(d)  $6 \times 10^{23}$

**Ans. (d)**

Weight of  $\text{NH}_3 = 4.25 \text{ g}$

Number of moles of

$$\text{NH}_3 = \frac{\text{Weight}}{\text{Molecular weight}}$$

$$= \frac{4.25}{17} = 0.25 \text{ mol}$$

Number of molecules in 0.25 mole of  $\text{NH}_3$   
 $= 0.25 \times 6.023 \times 10^{23}$

So, number of atoms

$$= 4 \times 0.25 \times 6.023 \times 10^{23}$$

$$= 6.0 \times 10^{23}$$

- 27** Haemoglobin contains 0.33% of iron by weight. The molecular weight of haemoglobin is approximately 67200 g. The number of iron atoms (at. weight of Fe is 56) present in one molecule of haemoglobin are

[CBSE AIPMT 1998]

(a) 1

(b) 6

(c) 4

(d) 2

**Ans. (c)**

$\therefore$  0.33 % of iron by weight means 100 g of haemoglobin has 0.33 g of iron

100 g of haemoglobin contains iron  
 $= 0.33 \text{ g}$

$\therefore$  67200 g of haemoglobin contains iron

$$= \frac{0.33 \times 67200}{100} \text{ g}$$

$$= 221.76 \text{ g of Fe}$$

$$\begin{aligned}\text{Number of Fe-atoms} &= \frac{221.76}{56} \\ &= 3.96 \approx 4\end{aligned}$$

- 28** The number of moles of oxygen in 1 L of air containing 21% oxygen by volume, under standard conditions, is

[CBSE AIPMT 1995]

- (a) 0.0093 mole (b) 2.10 moles  
(c) 0.186 mole (d) 0.21 mole

**Ans. (a)**

$$\begin{aligned}\text{Volume of oxygen in 1 L of air} \\ &= \frac{21}{100} \times 1000 = 210 \text{ mL}\end{aligned}$$

$\therefore$  22400 mL volume at STP is occupied by oxygen = 1 mole

Therefore, number of moles occupied by 210 mL

$$= \frac{210}{22400} = 0.0093 \text{ mol}$$

- 29** The percentage weight of Zn in white vitriol [ $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ ] is approximately equal to (at. mass of Zn = 65, S = 32, O = 16 and H = 1)

[CBSE AIPMT 1995]

- (a) 33.65% (b) 32.56%  
(c) 23.65% (d) 22.65%

**Ans. (d)**

$$\begin{aligned}\text{Molecular weight of} \\ \text{ZnSO}_4 \cdot 7\text{H}_2\text{O} &= 65 + 32 + (4 \times 16) + 7(18) \\ &= 287\end{aligned}$$

$$\begin{aligned}\therefore \text{Percentage weight of Zn} &= \frac{65}{287} \times 100 \\ &= 22.65\%\end{aligned}$$

- 30** The total number of valence electrons in 4.2 g of  $\text{N}_3^-$  ion is ( $N_A$  is the Avogadro's number)

[CBSE AIPMT 1994]

- (a)  $2.1 N_A$  (b)  $4.2 N_A$  (c)  $1.6 N_A$  (d)  $3.2 N_A$

**Ans. (c)**

$$\text{Moles of } \text{N}_3^- \text{ ion} = \frac{4.2}{42} = 0.1$$

Each nitrogen atom has 5 valence electrons. Therefore, total number of electrons in  $\text{N}_3^-$  ion = 16

$$\begin{aligned}\text{Total number of electrons in 0.1 mole or} \\ 4.2 \text{ g of } \text{N}_3^- \text{ ion} &= 0.1 \times 16 \times N_A = 1.6 N_A\end{aligned}$$

- 31** The number of gram molecules of oxygen in  $6.02 \times 10^{24}$  CO molecules is

[CBSE AIPMT 1990]

- (a) 10 g molecules (b) 5 g molecules  
(c) 1 g molecule (d) 0.5 g molecule

**Ans. (b)**

$$\begin{aligned}6.023 \times 10^{23} \text{ molecules of CO} \\ &= 1 \text{ mole of CO}\end{aligned}$$

$$\begin{aligned}6.02 \times 10^{24} \text{ molecules of CO} \\ &= 10 \text{ moles of CO}\end{aligned}$$

$$= 10 \text{ g atoms of O} = 5 \text{ g molecules of O}_2$$

- 32** The number of oxygen atoms in 4.4 g of  $\text{CO}_2$  is

[CBSE AIPMT 1990]

- (a)  $1.2 \times 10^{23}$  (b)  $6 \times 10^{22}$   
(c)  $6 \times 10^{23}$  (d)  $12 \times 10^{23}$

**Ans. (a)**

$$\begin{aligned}1 \text{ mole of CO}_2 &= 44 \text{ g of CO}_2 \\ &= 6.023 \times 10^{23} \text{ molecules}\end{aligned}$$

$$\begin{aligned}\therefore 4.4 \text{ g of CO}_2 &= 0.1 \text{ mole of CO}_2 \\ &= 6.023 \times 0.1 \times 10^{23} \text{ molecules} \\ &= 6.023 \times 10^{22} \text{ molecules} \\ &= 6.023 \times 10^{22} \text{ molecules of O}_2 \\ &= 2 \times 6.023 \times 10^{22} \text{ atoms of O} \\ &\approx 1.2 \times 10^{23} \text{ atoms of O}\end{aligned}$$

- 33** Ratio of  $C_p$  and  $C_v$  of a gas 'X' is

1:4. The number of atoms of the gas 'X' present in 11.2 L of it at NTP will be

[CBSE AIPMT 1989]

- (a)  $6.02 \times 10^{23}$  (b)  $1.2 \times 10^{23}$   
(c)  $3.01 \times 10^{23}$  (d)  $2.01 \times 10^{23}$

**Ans. (a)**

For the gas X ratio of  $C_p/C_v = 1:4$

So, the gas X is diatomic.

At NTP, volume of 1 mole of a gas = 22.4 L

$$1 \text{ mole of a gas} = 6.023 \times 10^{23} \text{ molecules}$$

Thus, at NTP 22.4 L contains

$$= 6.023 \times 10^{23} \text{ molecules}$$

So, at NTP 11.2 L contains

$$= \frac{6.023 \times 10^{23} \times 11.2}{22.4} \text{ molecules}$$

$$= 3.01 \times 10^{23} \text{ molecules}$$

Hence, number of atoms of gas 'X' (diatomic)

$$= 3.01 \times 10^{23} \times 2 \text{ atoms}$$

$$= 6.02 \times 10^{23} \text{ atoms}$$

- 34** 1 cc  $\text{N}_2\text{O}$  at NTP contains

[CBSE AIPMT 1988]

- (a)  $\frac{1.8}{224} \times 10^{22}$  atoms  
(b)  $\frac{6.02}{22400} \times 10^{23}$  molecules  
(c)  $\frac{1.32}{224} \times 10^{23}$  electrons  
(d) All of the above

**Ans. (d)**

$$\begin{aligned}\text{At NTP } 22400 \text{ cc of } \text{N}_2\text{O} &\text{ contains} \\ &= 6.02 \times 10^{23} \text{ molecules}\end{aligned}$$

$$\begin{aligned}\therefore 1 \text{ cc } \text{N}_2\text{O} &\text{ will contain} \\ &= \frac{6.02 \times 10^{23}}{22400} \text{ molecules}\end{aligned}$$

$$\begin{aligned}\text{In } \text{N}_2\text{O} \text{ molecule, number of atoms} \\ &= 2 + 1 = 3\end{aligned}$$

Thus, number of atoms

$$= \frac{3 \times 6.02 \times 10^{23}}{22400} \text{ atoms}$$

$$= \frac{1.8 \times 10^{22}}{224} \text{ atoms}$$

$$\begin{aligned}\text{In } \text{N}_2\text{O} \text{ molecule, number of electrons} \\ &= 7 + 7 + 8 = 22\end{aligned}$$

Hence, number of electrons

$$= \frac{6.02 \times 10^{23}}{22400} \times 22 \text{ electrons}$$

$$= \frac{1.32 \times 10^{23}}{224} \text{ electrons}$$

- 35** At STP, the density of  $\text{CCl}_4$  vapour in g/L will be nearest to

[CBSE AIPMT 1988]

- (a) 6.87 (b) 3.42 (c) 10.26 (d) 4.57

**Ans. (a)**

$$\begin{aligned}1 \text{ mole CCl}_4 \text{ vapours} \\ &= 12 + 4 \times 35.5 = 154 \text{ g}\end{aligned}$$

$$\begin{aligned}\text{At STP, volume of 1 mole of a gas} \\ &= 22.4 \text{ L}\end{aligned}$$

$$\text{Thus, } 154 \text{ g} = 22.4 \text{ L}$$

$$\begin{aligned}\therefore \text{Density of CCl}_4 \text{ vapours} &= \frac{154}{22.4} \text{ g L}^{-1} \\ &= 6.87 \text{ g L}^{-1}\end{aligned}$$

## TOPIC 4

### Stoichiometric and Volumetric Calculations

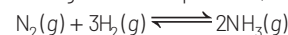
- 36** The number of moles of hydrogen molecules required to produce 20 moles of ammonia through Haber's process is

[NEET (National) 2019]

- (a) 20 (b) 30 (c) 40 (d) 10

**Ans. (b)**

According to Haber's process,



Now, according to above equations

2 moles of ammonia ( $\text{NH}_3$ ) require = 3 moles of  $\text{H}_2$

$$\therefore 1 \text{ mole of } \text{NH}_3 \text{ require} = 3/2 \text{ moles of } \text{H}_2$$

or, 20 moles of  $\text{NH}_3$  require  $= \frac{3}{2} \times 20$

moles of  $\text{H}_2 = 30$  moles of  $\text{H}_2$ .

**Note** Involvement of any limiting reagent is not mentioned in question.

- 37** 20.0 g of a magnesium carbonate sample decomposes on heating to give carbon dioxide and 8.0 g magnesium oxide. What will be the percentage purity of magnesium carbonate in the sample? (Atomic weight of Mg = 24)

[CBSE AIPMT 2015]

(a) 75 (b) 96 (c) 60 (d) 84

**Ans. (d)**

**Key Concept** In the given problem we have provided practical yield of MgO. For calculation of percentage yield of MgO, we need theoretical yield of MgO. For this we shall use mole concept.



$$\begin{aligned} \text{Moles of MgCO}_3 &= \frac{\text{Weight in gram}}{\text{Molecular weight}} \\ &= \frac{20}{84} = 0.238 \text{ mol} \end{aligned}$$

From Eq. (i)

1 mole of  $\text{MgCO}_3$  gives = 1 mol MgO

$\therefore$  0.238 mole  $\text{MgCO}_3$  will give

$$= 0.238 \text{ mol MgO}$$

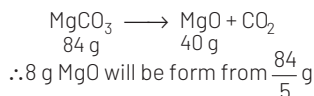
$$= 0.238 \times 40 \text{ g}$$

$$= 9.52 \text{ g MgO}$$

Now, practical yield of MgO = 8 g

$$\therefore \% \text{ purity} = \frac{8}{9.52} \times 100 = 84\%$$

**Alternate Method**



$$\therefore \% \text{ purity} = \frac{84}{5} \times \frac{100}{20} = 84\%$$

- 38** What is the mass of precipitate formed when 50 mL of 16.9% solution of  $\text{AgNO}_3$  is mixed with 50 mL of 5.8% NaCl solution? (Ag = 107.8, N = 14, O = 16, Na = 23, Cl = 35.5)

[CBSE AIPMT 2015]

(a) 28 g (b) 3.5 g  
(c) 7 g (d) 14 g

**Ans. (c)**

**Plan** For the calculation of mass of AgCl precipitated, we find mass of  $\text{AgNO}_3$  and

NaCl in equal volume with the help of mole concept.

16.9% solution of  $\text{AgNO}_3$  means 16.9 g  $\text{AgNO}_3$  is present in 100 mL solution.

$\therefore$  8.45 g  $\text{AgNO}_3$  will be present in 50 mL solution.

Similarly,

5.8 g NaCl is present in 100 mL solution

$\therefore$  2.9 g NaCl is present in 50 mL solution



Initial mole				
8.45	2.9	0	0	
169.8	58.5			
$= 0.049 = 0.049$				

**After reaction**

0 0 0.049 0.049

$\therefore$  Mass of AgCl precipitated

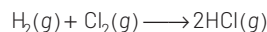
$$= 0.049 \times 143.5 = 7 \text{ g}$$

- 39** When 22.4 L of  $\text{H}_2(\text{g})$  is mixed with 11.2 L of  $\text{Cl}_2(\text{g})$ , each at STP, the moles of  $\text{HCl}(\text{g})$  formed is equal to [CBSE AIPMT 2014]

(a) 1 mole of HCl (g)  
(b) 2 moles of HCl (g)  
(c) 0.5 mole of HCl (g)  
(d) 1.5 moles of HCl (g)

**Ans. (a)**

The given problem is related to the concept of stoichiometry of chemical equations. Thus, we have to convert the given volumes into their moles and then, identify the limiting reagent [possessing minimum number of moles and gets completely used up in the reaction]. The limiting reagent gives the moles of product formed in the reaction.



Initial vol. 22.4 L 11.2 L 2 mol

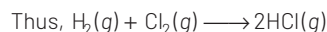
$\therefore$  22.4 L volume at STP is occupied by

$$\text{Cl}_2 = 1 \text{ mole}$$

$\therefore$  11.2 L volume will be occupied by

$$\text{Cl}_2 = \frac{1 \times 11.2}{22.4} \text{ mol} = 0.5 \text{ mol}$$

22.4 L volume at STP is occupied by  $\text{H}_2 = 1 \text{ mol}$



1 mol 0.5 mol

Since,  $\text{Cl}_2$  possesses minimum number of moles, thus it is the limiting reagent.

As per equation,

$$1 \text{ mole of } \text{Cl}_2 = 2 \text{ moles of HCl}$$

$$\therefore 0.5 \text{ mole of } \text{Cl}_2 = 2 \times 0.5 \text{ mole of HCl}$$

$$= 1.0 \text{ mole of HCl}$$

Hence, 1.0 mole of  $\text{HCl}(\text{g})$  is produced by 0.5 mole of  $\text{Cl}_2$  [or 11.2 L].

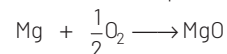
- 40** 1.0 g of magnesium is burnt with 0.56 g of oxygen in a closed vessel. Which reactant is left in excess and how much? [CBSE AIPMT 2014]

(At. weight of Mg = 24, O = 16)

(a) Mg, 0.16 g  
(b)  $\text{O}_2$ , 0.16 g  
(c) Mg, 0.44 g  
(d)  $\text{O}_2$ , 0.28 g

**Ans. (a)**

The balanced chemical equation is



24 g 16 g 40 g

From the above equation, it is clear that, 24 g of Mg reacts with 16 g of  $\text{O}_2$ .

Thus, 1.0 g of Mg reacts with

$$\frac{16}{24} \text{ g of } \text{O}_2 = 0.67 \text{ g of } \text{O}_2.$$

But only 0.56 g of  $\text{O}_2$  is available which is less than 0.67 g. Thus,  $\text{O}_2$  is the limiting reagent.

Further, 16 g of  $\text{O}_2$  reacts with 24 g of Mg.

$\therefore$  0.56 g of  $\text{O}_2$  will react with Mg

$$= \frac{24}{16} \times 0.56$$

$$= 0.84 \text{ g}$$

$\therefore$  Amount of Mg left unreacted

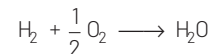
$$= (1.0 - 0.84) \text{ g Mg}$$

$$= 0.16 \text{ g Mg}$$

- 41** 10 g of hydrogen and 64 g of oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be [CBSE AIPMT 2009]

(a) 2 moles (b) 3 moles  
(c) 4 moles (d) 1 mole

**Ans. (c)**



1 mol  $\frac{1}{2}$  mol 1 mol

$\frac{10}{2}$  mol  $\frac{64}{32}$  mol ?

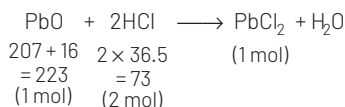
5 mol 2 mol

$\therefore \frac{1}{2}$  mole of  $\text{O}_2$  gives = 1 mole of  $\text{H}_2\text{O}$

$\therefore$  2 moles of  $\text{O}_2$  will give =  $1 \times 2 \times 2$   
= 4 moles of water

- 42** How many moles of lead (II) chloride will be formed from a reaction between 6.5 g of PbO and 3.2 g of HCl? [CBSE AIPMT 2008]  
(a) 0.044 (b) 0.333 (c) 0.011 (d) 0.029

**Ans. (d)**



$$\text{Mole of PbO} = \frac{6.5}{223} = 0.029$$

$$\text{Mole of HCl} = \frac{3.2}{36.5} = 0.087$$

Since, 1 mole of PbO reacts with 2 moles of HCl, thus in this reaction PbO is the limiting reagent.

Hence, 1 mole of PbO forms  
= 1 mole of PbCl<sub>2</sub>

0.029 mole of PbO will form = 0.029 mole of PbCl<sub>2</sub>

**43** What volume of oxygen gas (O<sub>2</sub>)

measured at 0°C and 1 atm, is needed to burn completely 1L of propane gas (C<sub>3</sub>H<sub>8</sub>) measured under the same conditions?

[CBSE AIPMT 2008]

- (a) 7 L (b) 6 L  
(c) 5 L (d) 10 L

**Ans. (c)**



$$22.4 \text{ L} \quad 5 \times 22.4 \text{ L}$$

For the combustion of 22.4 L propane, oxygen required = 5 × 22.4 L

For the combustion of 1 L of propane oxygen required

$$= \frac{5 \times 22.4}{22.4} \text{ L} = 5 \text{ L}$$

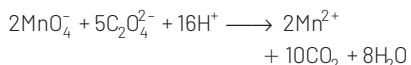
**44** Number of moles of MnO<sub>4</sub><sup>-</sup> required

to oxidise one mole of ferrous oxalate completely in acidic medium will be [CBSE AIPMT 2008]

- (a) 0.6 mole (b) 0.4 mole  
(c) 7.5 moles (d) 0.2 mole

**Ans. (b)**

In acidic medium MnO<sub>4</sub><sup>-</sup> oxidises ferrous oxalate as follows:



∴ 5 moles of oxalate ions are oxidised by 2 moles of MnO<sub>4</sub><sup>-</sup>.

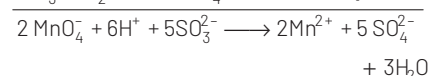
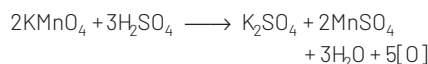
∴ 1 mole of oxalate ion is oxidised by

$$= \frac{2}{5} \text{ mole of MnO}_4^- = 0.4 \text{ mole of MnO}_4^-$$

**45** The number of moles of KMnO<sub>4</sub> that will be needed to react with one mole of sulphite ion in acidic solution is [CBSE AIPMT 2007]

- (a) 4/5 (b) 2/5  
(c) 1 (d) 3/5

**Ans. (b)**



5 moles of sulphite ions react with

= 2 moles of MnO<sub>4</sub><sup>-</sup>

So, 1 mole of sulphite ions react with

$$= \frac{2}{5} \text{ moles of MnO}_4^-.$$

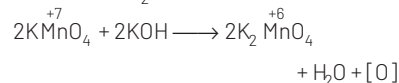
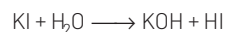
**46** The number of moles of KMnO<sub>4</sub> reduced by one mole of KI in alkaline medium is

[CBSE AIPMT 2005]

- (a) one fifth (b) five  
(c) one (d) two

**Ans. (c)**

In alkaline medium, KMnO<sub>4</sub> is reduced to K<sub>2</sub>MnO<sub>4</sub>



Hence, one mole of KMnO<sub>4</sub> is reduced by one mole of KI.

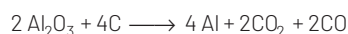
**47** The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of aluminium metal from bauxite by the Hall process is (at. mass of Al = 27)

[CBSE AIPMT 2005]

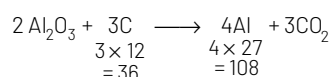
- (a) 180 kg (b) 270 kg (c) 540 kg (d) 90 kg

**Ans. (d)**

In Hall and Heroult process,



but for the removal of only CO<sub>2</sub>, following equation is possible.



∴ For 108 g of Al, 36 g of C is required in above reaction.

∴ For 270 × 10<sup>3</sup> g of Al required amount of C

$$= \frac{36}{108} \times 270 \times 10^3$$

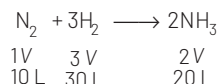
$$= 90 \times 10^3 \text{ g} = 90 \text{ kg}$$

**48** In Haber process 30L of dihydrogen and 30L of dinitrogen were taken for reaction which yielded only 50% of the expected product. What will be the composition of gaseous mixture under the aforesaid condition in the end?

[CBSE AIPMT 2003]

- (a) 20 L ammonia, 10 L nitrogen, 30 L hydrogen  
(b) 20 L ammonia, 25 L nitrogen, 15 L hydrogen  
(c) 20 L ammonia, 20 L nitrogen, 20 L hydrogen  
(d) 10 L ammonia, 25 L nitrogen, 15 L hydrogen

**Ans. (d)**



As only 50% of the expected product is formed, hence only 10 L of NH<sub>3</sub> is formed.

Thus, for the production of 10 L of NH<sub>3</sub>, 5 L of N<sub>2</sub> and 15 L of H<sub>2</sub> are used and composition of gaseous mixture under the aforesaid condition in the end is

$$\text{H}_2 = 30 - 15 = 15 \text{ L}$$

$$\text{N}_2 = 30 - 5 = 25 \text{ L}$$

$$\text{NH}_3 = 10 \text{ L}$$

**49** Which has maximum number of molecules? [CBSE AIPMT 2002]

- (a) 7 g N<sub>2</sub> (b) 2 g H<sub>2</sub>  
(c) 16 g NO<sub>2</sub> (d) 16 g O<sub>2</sub>

**Ans. (b)**

In 7 g nitrogen, number of molecules

$$= \frac{7.0}{28} \text{ mol}$$

$$= 0.25 \times N_A \text{ molecules}$$

where, N<sub>A</sub> = Avogadro number

$$= 6.023 \times 10^{23}$$

$$\text{In 2 g of H}_2 = \frac{2.0}{2} \text{ mol}$$

$$= 1 \times N_A \text{ molecules}$$

$$\text{In } 16 \text{ g of NO}_2 = \frac{16.0}{46} \text{ mol} \\ = 0.348 \times N_A \text{ molecules}$$

$$\text{In } 16 \text{ g of } \text{O}_2 = \frac{16}{32} \text{ mol} = 0.5 \times N_A \text{ molecules}$$

Hence, maximum number of molecules are present in 2 g of  $\text{H}_2$ .

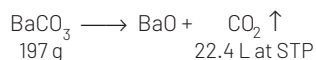
- 50** Assuming fully decomposed, the volume of  $\text{CO}_2$  released at STP on heating 9.85 g of  $\text{BaCO}_3$  (at. mass of Ba = 137) will be

[CBSE AIPMT 2000]

- (a) 1.12 L (b) 0.84 L  
(c) 2.24 L (d) 4.96 L

**Ans. (a)**

On decomposition,  $\text{BaCO}_3$  liberates  $\text{CO}_2$  as



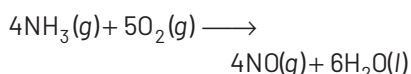
$\therefore$  197 g of  $\text{BaCO}_3$  gives

$$= 22.4 \text{ L of CO}_2 \text{ at STP}$$

$\therefore$  9.85 g of  $\text{BaCO}_3$  will give

$$= \frac{22.4 \times 9.85}{197} = 1.12 \text{ L}$$

- 51** In the reaction,

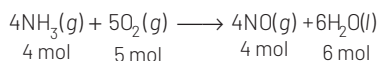


When 1 mole of ammonia and 1 mole of  $\text{O}_2$  are made to react to completion, then

[CBSE AIPMT 1998]

- (a) 1.0 mole of  $\text{H}_2\text{O}$  is produced  
(b) 1.0 mole of NO will be produced  
(c) all the oxygen will be consumed  
(d) all the ammonia will be consumed

**Ans. (c)**



According to equation,

5 moles of  $\text{O}_2$  required = 4 moles of  $\text{NH}_3$

1 mole of  $\text{O}_2$  requires

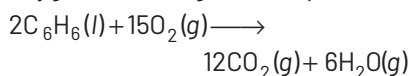
$$= \frac{4}{5} = 0.8 \text{ mole of NH}_3$$

$$\text{While 1 mole of NH}_3 \text{ requires} = \frac{5}{4}$$

$$= 1.25 \text{ moles of O}_2$$

As there is 1 mole of  $\text{NH}_3$  and 1 mole of  $\text{O}_2$ , so all the oxygen will be consumed.

- 52** Liquid benzene ( $\text{C}_6\text{H}_6$ ) burns in oxygen according to the equation,



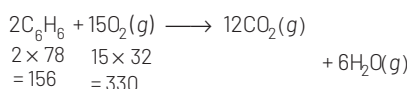
How many litres of  $\text{O}_2$  at STP are needed to complete the

combustion of 39 g of liquid benzene? (Mol. weight of  $\text{O}_2 = 32$ ,  $\text{C}_6\text{H}_6 = 78$ )

[CBSE AIPMT 1996]

- (a) 74 L (b) 11.2 L (c) 22.4 L (d) 84 L

**Ans. (d)**



$\therefore$  156 g of benzene required oxygen

$$= 15 \times 22.4 \text{ L}$$

$\therefore$  1 g of benzene required oxygen

$$= \frac{15 \times 22.4}{156} \text{ L}$$

$\therefore$  39 g of benzene required oxygen

$$= \frac{15 \times 22.4 \times 39}{156}$$

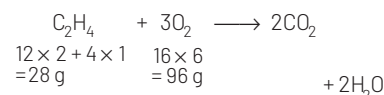
$$= 84.0 \text{ L}$$

- 53** What is the weight of oxygen required for the complete combustion of 2.8 kg of ethylene?

[CBSE AIPMT 1989]

- (a) 2.8 kg (b) 6.4 kg  
(c) 9.6 kg (d) 96 kg

**Ans. (c)**



$\therefore$  For the combustion of  $28 \times 10^{-3}$  kg of ethylene oxygen required =  $96 \times 10^{-3}$  kg

$\therefore$  For the combustion of 2.8 kg of ethylene oxygen required

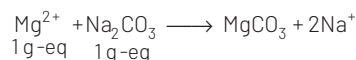
$$= \frac{96 \times 10^{-3} \times 2.8}{28 \times 10^{-3}}$$

$$= 9.6 \text{ kg}$$

- 54** One litre hard water contains 12.00 mg  $\text{Mg}^{2+}$ . Milliequivalents of washing soda required to remove its hardness is [CBSE AIPMT 1988]

- (a) 1  
(b) 12.16  
(c)  $1 \times 10^{-3}$   
(d)  $12.16 \times 10^{-3}$

**Ans. (a)**



$$1 \text{ g-equivalent of Mg}^{2+} = 12 \text{ g of Mg}^{2+}$$

$$= 12000 \text{ mg of Mg}^{2+}$$

Now, 12000 mg of  $\text{Mg}^{2+} \equiv 1000$

milliequivalent of  $\text{Na}_2\text{CO}_3$

12 mg of  $\text{Mg}^{2+} \equiv 1$  milliequivalent of  $\text{Na}_2\text{CO}_3$