Some Basic Concepts of Chemistry

Question1

The mass of a mixture containing NaCl and NaBr is 4.0 g . If Na is 30% of the total mixture, the composition of NaCl in the mixture is $(\mathrm{Na}=23\mathrm{u}, \mathrm{Cl}=35.5\mathrm{u}, \mathrm{Br}=80\mathrm{u})$

AP EAPCET 2025 - 26th May Evening Shift

Options:

A.

48%

B.

55%

C.

45%

D.

52%

Answer: C

Solution:

Step 1: Find the mass of sodium (Na) in the mixture.

Na makes up 30% of the total mixture.

So, mass of Na = $0.30 \times 4 = 1.2$ g

Step 2: Set up variables for the unknowns.

Let x = mass of NaCl in grams.
Let y = mass of NaBr in grams.

Step 3: Write the first equation (total mass).

$x + y = 4 \quad \dots (i)$

Step 4: Write the second equation (mass of Na from both salts).

Molar mass of NaCl = $23 + 35.5 = 58.5$ u.
Molar mass of NaBr = $23 + 80 = 103$ u.

Each gram of NaCl contains $\frac{23}{58.5}$ g of Na, and each gram of NaBr contains $\frac{23}{103}$ g of Na.

So, total mass of Na from both salts is:

$x \times \frac{23}{58.5} + y \times \frac{23}{103} = 1.2 \quad \dots (ii)$

Step 5: Solve the two equations.

From (i): $y = 4 - x$

Substitute y into (ii):

$x \times \frac{23}{58.5} + (4 - x) \times \frac{23}{103} = 1.2$

(Solving this equation for x gives $x = 1.8068$ g.)

Step 6: Find the percentage of NaCl in the mixture.

$\%$ of NaCl $= \frac{\text{Mass of NaCl}}{\text{Total mass}} \times 100 = \frac{1.8068}{4} \times 100 \approx 45\%$

Question2

Two acids A and B are titrated separately, 25 mL of $0.5 \mathrm{M} \mathrm{Na}_2\mathrm{CO}_3$ solution requires 10 mL of A and 40 mL of B for complete neutralisation. The volume (in L) of A and B required to produce 1 L of 1 N acid solution respectively are

AP EAPCET 2025 - 24th May Morning Shift

Options:

A.

0.2, 0.8

B.

0.8, 0.2

C.

$0.3, 0.7$

D.

$0.7, 0.3$

Answer: A

Solution:

Normality of Na_2CO_3 = molarity $\times n$-factor

$= 0.5M \times 2 = 1$ N

Normality of acid A,

$$N_{\text{acid } A} \times V_{\text{acid } A} = N_{Na_2CO_3} \times V_{Na_2CO_3}$$
$$N_{\text{acid } A} = \frac{1 \times 25}{10} = 2.5 \text{ N}$$

Similarly normality of acid B

$N_{\text{acid } B} = 0.625$ N

Volume of acid A needed for 1 L of 1 N solution

$$N_{\text{acid } A} \times V_{A \text{ final}} = N_{\text{final}} \times V_{\text{final}}$$
$$V_{A \text{ final}} = \frac{1 \times 1}{2.5} = 0.4 \text{ L} \quad \ldots(i)$$

Volume of acid B needed for 1 L of 1 N solution,

$$V_{B \text{ final}} = \frac{1 \times 1}{0.625} = 1.6 \text{ L} \quad \ldots(ii)$$

Question3

An ideal gas mixture of C_2H_6 and C_2H_4 occupies a volume of 28 L at 1 atm and 273 K . This mixture reacts completely with 128 g of O_2 to produce CO_2 and $H_2O(l)$. What is the mole fraction of C_2H_4 in the mixture ?

AP EAPCET 2025 - 23rd May Evening Shift

Options:

A.

0.4

B.

0.8

C.

0.5

D.

0.6

Answer: D

Solution:

Total moles of gas,

$$\begin{aligned} n_{\text{total}} &= \frac{\text{Volume}}{\text{Molar volume at STP}} \\ &= \frac{28}{224} = 1.25\text{moles} \end{aligned}$$

$$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$$
$$C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$$

Moles of $O_2 = \frac{128}{32} = 4$ mol

Let x be the moles of C_2H_6 and y be moles of C_2H_4

$$x + y = 1.25 \qquad \ldots(i)$$
$$\frac{7}{2}x + 3y = 4 \qquad \ldots(ii)$$

Solve for y, $y = 0.75$ mol

Mole fraction of $C_2H_4 = \frac{0.75}{1.25} = 0.6$

Question4

Complete combustion of ethane gives only gaseous products. In a closed vessel, 15 g of ethane and 112 g of O_2 were allowed to completely react. What is the total number of moles of gaseous substances present in the vessel at the end of the reaction?

Options:

A.

4.25

B.

2.5

C.

1.75

D.

8.50

Answer: A

Solution:

The balanced equation for complete reaction is,

$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

Moles of $C_2H_6 = \frac{15 \text{ g}}{30.07} = 0.499$ mol

Moles of $O_2 = \frac{112}{32} = 3.50$ mol

Moles of O_2 required for 0.499 mol of C_2H_6

$0.499 \times \frac{7}{2} = 1.747 \text{ mol} O_2$

Since, 3.50 mol, O_2 is available and only 1.747 mol .

O_2 is needed, C_2H_6 is limiting reactant. Mol of CO_2 produced

$0.499 \times \frac{4}{2} = 0.998 \text{ mol} CO_2$

mol of H_2O produced

$0.499 \times \frac{6}{2} = 1.497 \text{ mol} H_2O$

Moles of excess $O_2 = 3.50 - 1.747$

$= 1.753$ mol

Total mol of gaseous substance, $0.998 + 1.497 + 1.753 = 4.25$ mol

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Question5

The composition of a sample of wustite is $Fe_{0.93}O_{1.00}$. Percentage of iron in the form of Fe^{3+} ion is nearly

AP EAPCET 2025 - 23rd May Morning Shift

Options:

A.

85

B.

15

C.

93

D.

7

Answer: B

Solution:

Let x and y be number of Fe^{2+} and Fe^{3+} ions.

$x + y = 0.93 \qquad \ldots(i)$

Total negative charge from oxygen is,

$1 \times 2 = 2$

Total positive charge from Fe ions

$= 2x + 3y$

For charge neutrality $= 2x + 3y = 2 \qquad \ldots(ii)$

Solving Eqs. (i) and (ii)

$y = 0.14$

$$Fe^{3+}\% = \frac{\text{Number of } Fe^{3+} \text{ ions}}{\text{Total ions}} \times 100$$

$$\Rightarrow \frac{0.14}{0.93} \times 100 \approx 15\%$$

Question6

Consider the following

(A) 0.0025

(B) 500.0

(C) 2.0034

Number of significant figures in A, B and C respectively, are

AP EAPCET 2025 - 22nd May Evening Shift

Options:

A.

$5, 4, 4$

B.

$2, 4, 2$

C.

$4, 3, 2$

D.

$2, 4, 5$

Answer: D

Solution:

Let's find the **number of significant figures** carefully.

(A) 0.0025

- Leading zeros are **not significant**.
- Only the digits **2** and **5** are significant.

✅ **Significant figures** = **2**

(B) 500.0

- Trailing zeros after the decimal point **are significant**.
- 500.0 → digits 5, 0, 0, and 0 after decimal.

✅ **Significant figures** = **4**

(C) 2.0034

- All nonzero digits are significant.
- Zeros between nonzero digits **are significant**.
- So digits: 2, 0, 0, 3, 4 → all count.

✅ **Significant figures** = **5**

✅ Final Answer

(A, B, C) ⇒ (2, 4, 5)

Correct Option: D

$\boxed{2,\ 4,\ 5}$

Question7

1.84 g of a mixture of $CaCO_3$ and $MgCO_3$ is strongly heated to get a residue of 0.96 g . The percentage of $CaCO_3$ in the mixture is

AP EAPCET 2025 - 22nd May Morning Shift

Options:

A.

50.34

B.

49.66

C.

54.34

D.

45.66

Answer: C

Solution:

Let x be the mass of $CaCO_3$,y be the mass of $MgCO_3$

$x+y=1.84$ g $\quad\quad\ldots(i)$

When $CaCO_3$ deompose

$$\underset{100\text{ g}}{CaCO_3} \longrightarrow \underset{56\text{ g}}{CaO} + CO_2$$

Mass of CaO produced $=\frac{56}{100}x$

$$\underset{84}{MgCO_3} \longrightarrow \underset{40}{MgO} + CO_2$$

Mass of MgO produced $=\frac{40}{84}x$

Total mass of residue

$\frac{56}{100}x+\frac{40}{84}y=0.96$ g $\quad\quad\ldots(ii)$

Solving Eqs. (i) and (ii),

$x=1$

Mass of $CaCO_3=1$ g

$\%$ of $CaCO_3=\frac{1}{1.84}\times 100=54.34\%$

Question8

209 g of an element reacts with chlorine to form 315.5 g of its chloride. What is the weight (in g) of oxygen that reacts with 418 g of same element ?

$(\mathrm{Cl} = 35.5\mathrm{u}; \mathrm{O} = 16\mathrm{u})$

AP EAPCET 2025 - 21st May Evening Shift

Options:

A.

24

B.

48

C.

96

D.

36

Answer: C

Solution:

Mass of chlorine is the difference between the mass of the chloride and mass of the elements.

$315.5 - 209 = 106.59$

Equivalent weight of element

$$= \frac{\text{Mass of element}}{\text{Mass of chlorine}} \times \begin{array}{l}\text{Atomic weight} \\ \text{of chlorine}\end{array}$$

$$= \frac{209 \text{ g}}{106.5 \text{ g}} \times 35.5\mathrm{u} \approx 69.66\mathrm{u}$$

$\therefore$ Mass of oxygen

$$= \frac{\text{Mass of element}}{\text{Equivalent weight}} \times \begin{array}{l}\text{Atomic weight} \\ \text{of oxygen}\end{array}$$

of element

$$= \frac{418 \text{ g}}{69.66\mathrm{u}} \times 16\mathrm{u} = 6 \times 16 \text{ g} = 96 \text{ g}$$

Question9

Observe the following

I. 0.0063

II. 132.00

III. 1004

The number of significant figures in I, II and III is respectively.

AP EAPCET 2025 - 21st May Morning Shift

Options:

A.

$4, 3, 5$

B.

$4, 5, 4$

C.

$4, 3, 4$

D.

$2, 5, 4$

Answer: D

Solution:

The number of significant figures are

$0.0063 = 2$

[Zero before number is non-significant]

$132.00 = 5$

[Zero after decimals are significant]

$1004 = 4$

[Zero between two non-zero are significant]

Question10

Identify the correct statements from the following.

I. Reaction of hydrogen with fluorine occurs even in dark.

II. Manufacture of ammonia by Haber process is an endothermic reaction.

III. HF is electron rich hydride.

AP EAPCET 2024 - 22th May Evening Shift

Options:

A. I, III only

B. I, II, III

C. II, III only

D. I,II only

Answer: A

Solution:

Statements given in (I) and (III) are correct, while statement given in (ii) is incorrect. The correct from is :

Manufacture of ammonia by Haber process is an exothermic reaction.

Question11

10 g of a metal (M) reacts with oxygen to form 11.6 got oxide. What is the equivalent weight of M ?

AP EAPCET 2024 - 22th May Morning Shift

Options:

A. 50 g

B. 0.02

C. 0.02 g

D. 50

Answer: D

Solution:

Determine Oxygen Mass:

The metal's mass is $10\,\mathrm{g}$. The resulting oxide weighs $11.6\,\mathrm{g}$. Thus, the mass of oxygen can be calculated as:

Mass of oxygen $= 11.6\,\mathrm{g} - 10\,\mathrm{g} = 1.6\,\mathrm{g}$

Equivalent Mass of Oxygen:

It is given that the equivalent mass of oxygen is $8\,\mathrm{g}$.

Calculate Equivalent Mass of Metal (M):

Using the relationship for equivalent mass, which involves the mass of the substance and the mass of the reactive component (in this case, oxygen), we can set up the following equation:

$$\frac{\text{Equivalent mass of metal}}{\text{Equivalent mass of oxygen}} = \frac{\text{Mass of metal}}{\text{Mass of oxygen}}$$

Plugging in the values:

$$\frac{x}{8} = \frac{10}{1.6}$$

Solving for x:

$$x = \frac{80}{1.6} = 50$$

Thus, the equivalent weight of the metal M is 50.

Question12

A flask contains 98 mg of H_2SO_4. If 3.01×10^{20} molecules of H_2SO_4 are removed from the flask. The number of moles of H_2SO_4 remained in flask is

$$\left(N_A = 6.02 \times 10^{23}\right)$$

Options:

A. 1×10^{-4}

B. 5×10^{-4}

C. 1.66×10^{-3}

D. 9.95×10^{-3}

Answer: B

Solution:

To find the remaining moles of H_2SO_4 in the flask, follow these steps:

Step 1: Calculate the Molar Mass

The molar mass of H_2SO_4 is determined by adding the atomic masses of its constituent elements:

Hydrogen (H): $2 \times 1 = 2$

Sulfur (S): 32

Oxygen (O): $4 \times 16 = 64$

Thus, the total molar mass of H_2SO_4 is:

$2 + 32 + 64 = 98\,\text{g/mol}$

Step 2: Convert Mass from mg to g

Given:

98 mg of H_2SO_4.

Since 1 g = 1000 mg, convert milligrams to grams:

$98\,\text{mg} = \frac{98}{1000}\,\text{g} = 0.098\,\text{g}$

Step 3: Determine Initial Molecules

The number of molecules in 98 mg of H_2SO_4 can be found using Avogadro's number ($N_A = 6.02 \times 10^{23}$ molecules/mol):

98 grams of H_2SO_4 contains 6.02×10^{23} molecules. Therefore, 0.098 grams will have:

$\frac{6.02 \times 10^{23}}{1000} = 6.02 \times 10^{20}$ molecules

Step 4: Calculate Remaining Molecules

After removing 3.01×10^{20} molecules, the remaining number of molecules is:

$6.02 \times 10^{20} - 3.01 \times 10^{20} = 3.01 \times 10^{20}$ molecules

Step 5: Convert Remaining Molecules to Moles

Using the proportion of molecules to moles:

6.02×10^{23} molecules $= 1$ mol

Consequently, 3.01×10^{20} molecules are:

$\frac{3.01 \times 10^{20}}{6.02 \times 10^{23}} = 5 \times 10^{-4}$ moles

Thus, the remaining H_2SO_4 in the flask is 5×10^{-4} moles.

Question13

0.1 mole of potassium permanganate was heated at $300^{\circ}C$. What is weight (ing) of the residue?$(Mn = 55u,\ K = 39u, O = 16u)$

AP EAPCET 2024 - 21th May Morning Shift

Options:

A. 14.2

B. 1.6

C. 15.8

D. 7.1

Answer: A

Solution:

The decomposition reaction is,

$$\underset{2\text{ moles}}{2MMOO_4} \longrightarrow \underset{1\text{ moles}}{K_2MnO_4} + \underset{1\text{ moles}}{MnO_2} + \underset{1\text{ moles}}{O_2 \uparrow}$$

For 0.1 moles of $KMnO_4$

$= 0.1/2 = 0.05$moles

(same for K_2MnO_4 and MnO_2)

Molar mass of products,

$K_2MnO_4 = 197 \text{ g/mol}$
$MnO_2 = 87 \text{ g/mol}$

Weight of $K_2MnO_4 = 0.05 \times 197$

$= 9.85 \text{ g}$

Weight of $MnO_2 = 0.05 \times 87$

$= 4.35 \text{ g}$

Weight of residue = Weight of K_2MnO_4 and MnO_2

$9.85 + 4.35 = 14.2 \text{ g}$

Question14

100 mL of $\frac{M}{10}$ $Ca(NO_3)_2$ and 200 mL of $\frac{M}{10}$ KNO_3 solutions are mixed. What is the normality of resulted solution with respect to NO_3^-?

AP EAPCET 2024 - 21th May Morning Shift

Options:

A. 0.1 N

B. 0.2 N

C. 0.13 N

D. 0.066 N

Answer: C

Solution:

To determine the normality of the resulting solution with respect to NO_3^-, we begin by examining the individual solutions and their normalities:

For $Ca(NO_3)_2$**:**

Molarity (M): $\frac{M}{10}$

n**-factor**: 2 (because $Ca(NO_3)_2$ provides two NO_3^- ions per formula unit)

Normality (N): Calculated as $N = M \times n\text{-factor} = \frac{1}{10} \times 2 = 0.2\,\text{N}$

For KNO_3:

Molarity (M): $\frac{M}{10}$

n**-factor**: 1 (because KNO_3 provides one NO_3^- ion per formula unit)

Normality (N): Calculated as $N = M \times n\text{-factor} = \frac{1}{10} \times 1 = 0.1\,\text{N}$

To find the normality of the mixed solution with respect to NO_3^-, use the formula for mixing solutions:

$$N_{\text{result}} = \frac{V_1N_1 + V_2N_2}{V_1 + V_2}$$

$V_1 = 100\,\text{mL}$, $N_1 = 0.2\,\text{N}$ for $Ca(NO_3)_2$

$V_2 = 200\,\text{mL}$, $N_2 = 0.1\,\text{N}$ for KNO_3

Plugging in the values:

$$N_{\text{result}} = \frac{100 \times 0.2 + 200 \times 0.1}{100 + 200} = \frac{20 + 20}{300} = \frac{40}{300} = 0.13\,\text{N}$$

Thus, the normality of the resulting solution with respect to NO_3^- is 0.13 N.

Question15

At STP ' x ' g of a metal hydrogen carbonate $(M\text{HCO}_3)$ (molar mass 84 g mol^{-1}) on heating gives CO_2, which can completely react with 0.2 moles of MOH (molar mass 40 g mol^{-1}) to give $MHCO_3$. The value of ' x' is

AP EAPCET 2024 - 20th May Evening Shift

Options:

A. 67.2

B. 33.6

C. 11.2

D. 22.4

Answer: B

Solution:

To determine the value of 'x', we need to analyze two reactions involving metal hydrogen carbonate ($MHCO_3$).

Reaction I: Decomposition of $MHCO_3$

$$2\,MHCO_3 \xrightarrow{\Delta} CO_2 + M_2CO_3 + H_2O$$

In this reaction, 2 moles of $MHCO_3$ decompose to give 1 mole of CO_2.

For x g of $MHCO_3$, the number of moles is $\frac{x}{84}$.

This results in $\frac{x}{2\times84}$ moles of CO_2.

Reaction II: Reaction of CO_2 with MOH

$$CO_2 + MOH \longrightarrow MHCO_3$$

In this reaction, 1 mole of CO_2 reacts with 1 mole of MOH to form 1 mole of $MHCO_3$.

From Reaction I, we have $\frac{x}{2\times84}$ moles of CO_2.

According to the problem, this amount of CO_2 completely reacts with 0.2 moles of MOH.

Set the moles of CO_2, $\frac{x}{2\times84}$, equal to 0.2:

$$\frac{x}{2\times84} = 0.2$$

Solving for x:

$$x = 0.2 \times 2 \times 84 = 33.6$$

Thus, the value of x is 33.6.

Question16

The mass % of urea solution is 6 . The total weight of the solution is 1000 g . What is its concentration in $\mathrm{mol L^{-1}}$? (Density of water $= 1.0\ \mathrm{g\ mL^{-1}}$)

$(\mathrm{C} = 12\mathrm{u}, \mathrm{N} = 14\mathrm{u}, \mathrm{O} = 16\mathrm{u}, \mathrm{H} = 1\mathrm{u})$

AP EAPCET 2024 - 20th May Evening Shift

Options:

A. 1.5

B. 1.064

C. 1.12

D. 0.80

Answer: B

Solution:

To determine the concentration of a 6% urea solution with a total weight of 1000 g, follow these steps:

Calculate the amount of urea:

The solution contains 6% urea by mass. Thus, the mass of urea is:

$\text{Urea} = 1000\,\text{g} \times \frac{6}{100} = 60\,\text{g}$

The molar mass of urea (CH_4N_2O) is calculated using the atomic masses, which are:

Carbon (C): 12 u

Nitrogen (N): 14 u

Oxygen (O): 16 u

Hydrogen (H): 1 u

Therefore, the molar mass of urea is:

$\text{Molar mass of urea} = 12 + (4 \times 1) + (2 \times 14) + 16 = 60\,\text{g/mol}$

Convert the mass of urea to moles:

$\text{Moles of urea} = \frac{60\text{ g}}{60\text{ g/mol}} = 1\,\text{mol}$

Calculate the volume of water:

The remaining mass of the solution is water:

$\text{Water} = 1000\,\text{g} - 60\,\text{g} = 940\,\text{g}$

Using the density of water (1 g/mL), convert this mass to volume:

$\text{Volume of water} = \frac{940\text{ g}}{1\text{ g/mL}} = 940\,\text{mL} = 0.94\,\text{L}$

Calculate the concentration of the urea solution:

The concentration in molarity (mol/L) is given by:

$\text{Concentration} = \frac{\text{Moles of urea}}{\text{Volume of solution in liters}} = \frac{1\text{ mol}}{0.94\text{ L}} = 1.0638\,\text{M}$

Rounding to three significant figures gives:

$\text{Concentration} \approx 1.064\,\text{mol/L}$

Therefore, the concentration of the urea solution is approximately 1.064 mol/L.

Question17

$$\mathrm{Xe}(g)+2\ \mathrm{F}_2(g) \xrightarrow[7\ \text{bar}]{873\ \mathrm{K}} \mathrm{XeF}_4(\mathrm{~s})$$

The ratio of $\mathrm{Xe}:\mathrm{F}_2$ required in the above reaction is

AP EAPCET 2024 - 20th May Evening Shift

Options:

A. 1 : 2

B. 1 : 5

C. 1 : 20

D. 1 : 12

Answer: B

Solution:

Given reaction is

$$\underset{1\ \text{mole}}{\mathrm{Xe}(g)}\quad+\underset{5\ \text{moles}}{2\ \mathrm{F}_2(g)} \xrightarrow[7\ \text{bar}]{873\ \mathrm{K}} \mathrm{XeF}_4(s)$$

The ratio of $\mathrm{Xe}:\mathrm{Fe}_2$ is $1:5$.

Question18

The density of nitric acid solution is $1.5\ \mathrm{g\ mL}^{-1}$. Its weight percentage is 68 . What is the approximate concentration (in molL^{-1}) of nitric acid ?

$$(\mathrm{N}=14\mathrm{u};\mathrm{O}=16\mathrm{u};\mathrm{H}=1\mathrm{u})$$

Options:

A. 14.2

B. 11.6

C. 18.2

D. 16.2

Answer: D

Solution:

To find the concentration of nitric acid in mol/L, we'll use the provided data about its density and weight percentage.

Given:

Density of the solution = 1.5 g/mL

Weight percentage of nitric acid = 68%

Calculate the mass of the solution in 1L:

Since the density is given as 1.5 g/mL, the mass of 1 liter (1000 mL) of the solution would be:

Mass of solution $= 1.5 \text{ g/mL} \times 1000 \text{ mL} = 1500 \text{ g}$

Determine the mass of nitric acid in the solution:

The weight percentage indicates that 68% of the solution's mass is nitric acid. Therefore, the mass of nitric acid is:

Mass of $HNO_3 = \frac{68}{100} \times 1500 \text{ g} = 1020 \text{ g}$

Calculate the molar mass of nitric acid (HNO_3):

Using the atomic masses provided:

Molar mass of $HNO_3 = 1\,(H) + 14\,(N) + 16 \times 3\,(O) = 63 \text{ g/mol}$

Find the concentration in mol/L:

To find the number of moles of HNO_3 in 1L:

Number of moles $= \frac{\text{Mass of } HNO_3}{\text{Molar mass } HNO_3} = \frac{1020 \text{ g}}{63 \text{ g/mol}} = 16.20 \text{ mol/L}$

Therefore, the concentration of nitric acid is approximately 16.20 mol/L.

Question19

A hydrocarbon containing C and H has $92.3\%C$ When 39 g of hydrocarbon was completely burnt in 0 X moles of water and Y moles of CO_2 were formed H_2 with Na metal, What is the weight (in g) of ounp $^\circ$ consumed $)(C = 12u : H = 1f)$

AP EAPCET 2024 - 18th May Morning Shift

Options:

A. 120

B. 240

C. 360

D. 489

Answer: A

Solution:

$\text{Carbon} = 92.3/12 = 7.7$

$\text{Hydrogen} = 7.7/1 = 7.7$

$\text{Ratio C : H} = 1 : 1$

So, emperical formula $= CH$

In question, 39 g of hydrocarbon is given, it means 3 moies of CH is needed.

$$\text{So, } 3CH + \frac{15}{4}O_2 \longrightarrow \underset{\substack{\downarrow \\ Y}}{3CO_2} + \underset{\substack{\downarrow \\ X}}{\frac{3}{2}H_2O}$$

Also, $2Na + 2H_2O \longrightarrow \mathbf{2NaOH + H_2}$

$2H_2O$ gives 1 $molH_2$

so, $\frac{3}{2}$ mol of H_2O gives $\frac{1}{2} \times 1.5$ mol of H_2

$= 0.75\ molH_2$

So, all H_2O is consumed to liberate 0.75 mole of H_2.

Thus, weight of O_2 consumed

$= \frac{15}{4} \times 32 = 120$ g

Question20

The concentration of 1 L of $CaCO_3$ solution is 1000 ppm . What is its concentration in mol L^{-1} ?

$(\mathrm{Ca} = 40\mathrm{u}, \mathrm{O} = 16\mathrm{u}, \mathrm{C} = 12\mathrm{u})$

AP EAPCET 2024 - 18th May Morning Shift

Options:

A. 10^{-3}

B. 10^{-1}

C. 10^{-4}

D. 10^{-2}

Answer: D

Solution:

To determine the concentration of a $CaCO_3$ solution in molarity (mol/L), given its concentration is 1000 ppm, we proceed as follows:

Mass of $CaCO_3$ in 1 Liter:

Since 1000 ppm means 1000 mg of $CaCO_3$ per liter, it converts to grams as:

$\frac{1000 \times 1000 \text{ mg}}{1000000} = 1 \text{ g}$

Calculation of Moles:

The molar mass of $CaCO_3$ is calculated using the atomic masses:

$\mathrm{Ca} = 40 \mathrm{~u}, \mathrm{C} = 12 \mathrm{~u}, \mathrm{O} = 16 \mathrm{~u}$.

Therefore, the molar mass of $CaCO_3 = 40 + 12 + (16 \times 3) = 100$ g/mol.

Now, calculate the number of moles (n):

$$n = \frac{\text{Mass of} \quad CaCO_3}{\text{Molar mass}} = \frac{1}{100} = 0.01 \text{ mol}$$

Determine Molarity:

The molarity (concentration in mol/L) is given by:

$$\text{Molarity} = \frac{\text{Number of moles}}{\text{Volume of solution in L}} = \frac{0.01}{1} = 0.01 \text{ mol/L} = 10^{-2} \text{ mol/L}$$

Thus, the concentration of the $CaCO_3$ solution is 10^{-2} mol/L.

Question21

50 g of a substance is dissolved in 1 kg of water at $+90^{\circ}C$. The temperature is reduced to $+10^{\circ}C$. The density is increased from 1.1 to 1.15 g cc^{-1}. What is the % change of molarity of the solution?

AP EAPCET 2022 - 5th July Morning Shift

Options:

A. 10

B. 4.5

C. 5

D. 7.3

Answer: B

Solution:

To determine the percentage change in the molarity of the solution, we need to follow these steps:

1. Calculate the initial volume of the solution at $+90^{\circ}C$.
2. Calculate the final volume of the solution at $+10^{\circ}C$.
3. Determine the initial and final molarities of the solution.
4. Compute the percentage change in the molarity.

Given information:

- Mass of the solute (substance) = 50 g
- Mass of the solvent (water) = 1 kg = 1000 g

- Initial density = 1.1 g cc^{-1}
- Final density = 1.15 g cc^{-1}

Step 1: Calculate the initial volume of the solution at $+90°C$.

The total mass of the solution is:

$$\text{Total mass} = 50\,g + 1000\,g = 1050\,g$$

Using the initial density, the initial volume is:

$$V_{\text{initial}} = \frac{\text{Total mass}}{\text{Initial density}} = \frac{1050\,g}{1.1\text{ g/cc}} \approx 954.545\text{ cc}$$

Step 2: Calculate the final volume of the solution at $+10°C$.

Using the final density, the final volume is:

$$V_{\text{final}} = \frac{\text{Total mass}}{\text{Final density}} = \frac{1050\,g}{1.15\text{ g/cc}} \approx 913.043\text{ cc}$$

Step 3: Determine the initial and final molarities of the solution.

Molarity (M) is defined as the number of moles of solute per liter of solution. Let's assume the molar mass (M_s) of the solute (substance) is required (denote it as M_s).

Number of moles of the solute:

$$\text{Number of moles} = \frac{50\,g}{M_s\text{ g/mol}} = \frac{50}{M_s}\text{ mol}$$

Initial molarity:

$$M_{\text{initial}} = \frac{\text{Number of moles}}{\text{Initial volume in liters}} = \frac{\frac{50}{M_s}}{\frac{954.545}{1000}} = \frac{50 \times 1000}{M_s \times 954.545} \approx \frac{50 \times 1.047}{M_s}$$

Final molarity:

$$M_{\text{final}} = \frac{\text{Number of moles}}{\text{Final volume in liters}} = \frac{\frac{50}{M_s}}{\frac{913.043}{1000}} = \frac{50 \times 1000}{M_s \times 913.043} \approx \frac{50 \times 1.095}{M_s}$$

Step 4: Compute the percentage change in molarity.

Percentage change in molarity:

$$\text{Percentage change} = \frac{M_{\text{final}} - M_{\text{initial}}}{M_{\text{initial}}} \times 100$$

Substituting the values:

$$\text{Percentage change} = \frac{\frac{50 \times 1.095}{M_s} - \frac{50 \times 1.047}{M_s}}{\frac{50 \times 1.047}{M_s}} \times 100 = \frac{50 \times (1.095 - 1.047)}{50 \times 1.047} \times 100$$

$$\text{Percentage change} = \frac{(1.095 - 1.047)}{1.047} \times 100 = \frac{0.048}{1.047} \times 100 \approx 4.58\%$$

Since the closest provided option is 4.5, the correct answer is:

Option B: 4.5

Question22

The empirical formula of calgon is

AP EAPCET 2022 - 4th July Evening Shift

Options:

A. Na_2PO_3

B. NaP_2O_3

C. $NaPO_3$

D. Na_3PO_3

Answer: C

Solution:

Calgon is a water softener. The molecular formula of calgon is $Na_6P_6O_{18}$. Thus, the empirical formula of calgon is $NaPO_3$.

Question23

The statement related to law of definite proportions is

AP EAPCET 2022 - 4th July Morning Shift

Options:

A. the ratio of oxygen in H_2O and H_2O_2 with respect to fixed mass of hydrogen atom is a whole number.

B. % of oxygen in H_2O is constant irrespective of the source.

C. equal volume of all gases at the same temperature and pressure should contain equal number of molecules.

D. matter can neither be created nor destroyed.

Answer: B

Solution:

Law of definite proportion states that the individual elements that constitute a chemical compound are always present in a fixed mass ratio.

(a) The ratio of oxygen in H_2O and H_2O_2 with respect to fixed mass of hydrogen atom is a whole number. This is the statement of "law of multiple proportion."

(b) As the ratio of elements is fixed, so percentage of oxygen in H_2O is constant or fixed irrespective of the source. This statement is with reference to law of definite proportion. Thus, this is the correct option.

(c) Avogadro's law states that the "equal volume of all gases at the same temperature and pressure should contain equal number of molecules."

(d) "Mass can neither be created nor be destroyed" is the statement of "law of conservation of mass."

Question24

What are x and y in the following reaction?

$$x\,Pb_3O_4 \longrightarrow yPbO + O_2$$

AP EAPCET 2022 - 4th July Morning Shift

Options:

A. $x = 3, y = 6$

B. $x = 2, y = 4$

C. $x = 2, y = 5$

D. $x = 2, y = 6$

Answer: D

Solution:

The balanced equation is

$2Pb_3O_4 \longrightarrow 6PbO + O_2$

So, value of x and y in the given reactions are 2 and 6 respectively.

Question25

If the volume of $15.9 \mathrm{~g}$ of carbon tetrachloride is $10 \mathrm{~mL}$, calculate its density.

AP EAPCET 2021 - 20th August Evening Shift

Options:

A. $31.8 \mathrm{~g} \mathrm{~mL}^{-1}$

B. $1.59 \mathrm{~g} \mathrm{~mL}^{-1}$

C. $0.159 \mathrm{~g} \mathrm{~mL}^{-1}$

D. $15.9 \mathrm{~g} \mathrm{~mL}^{-1}$

Answer: B

Solution:

$$M = 15.9 \mathrm{~g} \text{ (Given)}$$
$$V = 10 \mathrm{~mL}$$

$$\text{Density } = \frac{\text{Mass}}{\text{Volume}} = \frac{15.9 \mathrm{~g}}{10 \mathrm{~mL}}$$
$$\text{Density } = 1.59 \mathrm{~g} \mathrm{~mL}^{-1}.$$

Question26

$0.63 \mathrm{~g}$ of oxalic acid is dissolved in order to obtain $250 \mathrm{~cm}^3$ of its solution. Find the normality of this solution. [oxalic acid $(\mathrm{COOH})_2 \cdot 2\mathrm{H}_2\mathrm{O}$]

Options:

A. 0.05 N

B. 0.01 N

C. 0.04 N

D. 0.02 N

Answer: C

Solution:

Volume of solution

$$
\begin{aligned}
& =250 \mathrm{~cm}^3=250 \mathrm{~mL} \\
& =250 \times 10^{-3} \mathrm{~L}
\end{aligned}
$$

Mass of oxalic acid $=0.63 \mathrm{~g}$

Molar mass of oxalic acid $=126$

Moles $=0.63 / 126=5 \times 10^{-3}$ moles

$$
\begin{aligned}
\text{Molarity } & =\frac{\text{Number of moles of solution}}{\text{Volume of solution in litre}} \\
M & =\frac{5 \times 10^{-3}}{250 \times 10^{-3}}=0.02 \mathrm{M}
\end{aligned}
$$

Normality = Basicity $\times$ Molarity

Basicity of oxalic acid = 2

$N=2 \times 0.02=0.04$ N.

Question27

7.8 g of a compound having molecular formula C_6H_6, on reacting with $CH_3COCl/AlCl_3$ gives 8.4 g of a product which has molecular formula C_8H_8O. Calculate the percentage yield of the product C_8H_8O. (Given, atomic weights of H, C and O respectively are 1, 12 and 16)

Options:

A. 70%

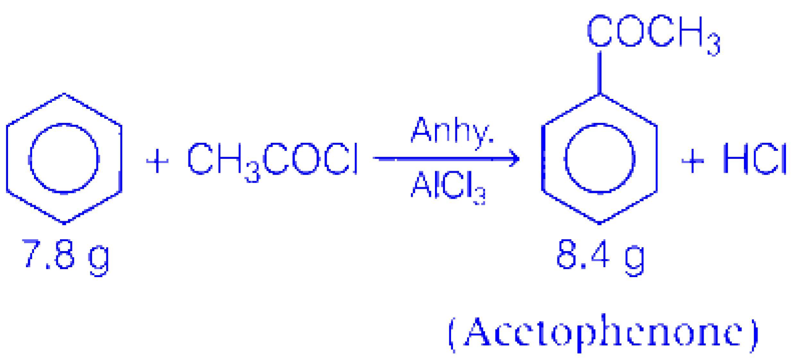
B. 60%

C. 80%

D. 75%

Answer: A

Solution:



1 mole of benzene produces 1 mole of $C_6H_5COCH_3$ 78 g of benzene produces 120 g of $C_6H_5COCH_3$

$\therefore$ 7.8 g of benzene produce

$$= \frac{120}{78} \times 7.8 = 12 \text{ g}$$

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

$$= \frac{8.4}{12} \times 100 = 70\%$$

Question28

The complete combustion of one mole ofbenzene produces .... grams of carbondioxide.

AP EAPCET 2021 - 20th August Evening Shift

Options:

A. 164

B. 220

C. 264

D. 308

Answer: C

Solution:

$$\underset{\text{Benzene}}{2C_6H_6} + 15O_2 \longrightarrow 12CO_2 + 6H_2O$$

2 moles of benzene produce 12 moles of CO_2.

$\therefore$ 1 mol of benzene produce 6 moles of CO_2

$$\text{Mol} = \frac{\text{Given mass}}{\text{Molecular mass}}$$
$$6 = x/44 \Rightarrow x = 264 \text{ g}$$

1 mol of benzene produces 264 g of CO_2.

Question29

An alloy of metals X and Y weighs 12 g andcontains atoms X and Y in the ratio of 2 : 5.The percentage of metal X in the alloy is 20by mass. If the atomic mass of X is 40 amuwhat is the atomic mass of metal Y ?

AP EAPCET 2021 - 20th August Morning Shift

Options:

A. 64 amu

B. 32 amu

C. 60 amu

D. 50 amu

Answer: A

Solution:

Mass of sample of alloy = 12 g

% of metal X in sample $= 20$

$\therefore$ If x is the mass of metal X in sample

$$\frac{x}{12} \times 100 = 20$$
$$x = 2.4 \text{ g}$$

If y is the mass of metal Y in sample, then

$y = 12 - 2.4 = 9.6$ g

Number of atoms of $X = \frac{6.022 \times 10^{23} \times 2.4}{40}$

$= 3.61 \times 10^{22}$

Ratio of atoms of X and $Y = 2 : 5$

Number of atoms of $Y = \frac{3.61 \times 10^{22} \times 5}{2}$

$= 9.025 \times 10^{22}$ atoms.

9.025×10^{22} atoms of Y are present in 9.6 g.

6.022×10^{23} atoms of Y are present in $\frac{9.5}{9.025 \times 10^{22}} \times 6.022 \times 10^{23} = 64$ g

Atomic mass of $Y = 64$ amu.

Question30

If 0.2 moles of sulphuric acid is poured into250 mL of water, calculate the concentrationof the solution.

AP EAPCET 2021 - 19th August Evening Shift

Options:

A. 0.8 N

B. 0.8 M

C. 8 M

D. 0.2 N

Answer: B

Solution:

Moles of sulphuric acid, $n_A = 0.2$

Volume of water, $V = 250$ mL

Concentration of the solution/molarity

$= \frac{n_A}{V(L)} \times 1000 = \frac{0.2}{250} \times 1000 = 0.8$ M

Molarity (molar concentration) : the number ofmoles of solute (sulphuric acid) per litre ofsolution (water).

Question31

When 20 g of $CaCO_3$ is treated with 20 g of HCl, the mass of CO_2 formed would be

AP EAPCET 2021 - 19th August Morning Shift

Options:

A. 10 g

B. 8.8 g

C. 22.2 g

D. 20 g

Answer: B

Solution:

Let $CaCO_3(s)$ be consumed completely.

$CaCO_3(s) + 2HCl(aq.) \longrightarrow CaCl_2(aq)$
Carcium Hydrochloric Calcium
chloride
$+ H_2O(l) + CO_2(g)$
$\therefore$ 100 g of $CaCO_3$ gives $= 44$ g of CO_2 gas
$\therefore$ 20 g of $CaCO_3$ will give $= \frac{44}{100} \times 20$ g of CO_2
$= 8.8$ g of CO_2

Now, HCl, be completely consumed.

73 g of HCl give $= 44$ g of CO_2 gas

$\therefore$ 20 g of HCl gives $= \frac{44}{73} \times 20$ g of $CO_2 = 12.05$ g of CO_2

Hence, $CaCO_3$ give least amount of CO_2 gas. So, $CaCO_3$ is limiting reactant

$\therefore$ Amount of CO_2 formed will be 8.8 g.
