Chemistry

(Chapter – 1) (Some Basic Concepts of Chemistry)

(Class – XI)

Question 1.1:

Calculate the molecular mass of the following:

(i) H_2O **(ii)** CO_2 **(iii)** CH_4

Answer 1.1:

(i) H_2O:

The molecular mass of water, H_2O

= (2 × Atomic mass of hydrogen) + (1 × Atomic mass of oxygen)

= [2(1.0084) + 1(16.00 u)]

= 2.016 u + 16.00 u

= 18.016

= 18.02 u

(ii) CO_2:

The molecular mass of carbon dioxide, CO_2

= (1 × Atomic mass of carbon) + (2 × Atomic mass of oxygen)

= [1(12.011 u) + 2 (16.00 u)]

= 12.011 u + 32.00 u

= 44.01 u

(iii) CH_4:

The molecular mass of methane, CH_4

= (1 × Atomic mass of carbon) + (4 × Atomic mass of hydrogen)

= [1(12.011 u) + 4 (1.008 u)]

= 12.011 u + 4.032 u

= 16.043 u

Question 1.2:

Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).

Answer 1.2:

The molecular formula of sodium sulphate is Na_2SO_4 .

Molar mass of Na_2SO_4 = [(2 × 23.0) + (32.066) + 4 (16.00)]

= 142.066 g

Mass percent of an element $= \dfrac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$

∴ Mass percent of sodium:

$$= \frac{46.0\,\text{g}}{142.066\,\text{g}} \times 100$$

$$= 32.379$$

$$= 32.4\%$$

Mass percent of sulphur:

$$= \frac{32.066\text{g}}{142.066\text{g}} \times 100$$

$$= 22.57$$

$$= 22.6\%$$

Mass percent of oxygen:

$$= \frac{64.0\text{g}}{142.066\text{g}} \times 100$$

$$= 45.049$$

$$= 45.05\%$$

Question 1.3:

Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass.

Answer 1.3:

% of iron by mass = 69.9 % [Given]

% of oxygen by mass = 30.1 % [Given] Relative moles of iron in iron oxide:

$$= \frac{\text{\% of iron by mass}}{\text{Atomic mass of iron}}$$

$$= \frac{69.9}{55.85}$$

$$= 1.25$$

Relative moles of oxygen in iron oxide:

$$= \frac{\text{\% of oxygen by mass}}{\text{Atomic mass or oxygen}}$$

$$= \frac{30.1}{16.00}$$

$$= 1.88$$

Simplest molar ratio of iron to oxygen:

= 1.25: 1.88

= 1: 1.5

$\simeq$ 2: 3

$\therefore$ The empirical formula of the iron oxide is Fe_2O_3.

Question 1.4:

Calculate the amount of carbon dioxide that could be produced when

(i) 1 mole of carbon is burnt in air.

(ii) 1 mole of carbon is burnt in 16 g of dioxygen.

(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Answer 1.5:

The balanced reaction of combustion of carbon can be written as:

(i) As per the balanced equation, 1 mole of carbon burns in1 mole of dioxygen (air) to produce1 mole of carbon dioxide.

(ii) According to the question, only 16 g of dioxygen is available. Hence, it will react with 0.5 mole of carbon to give 22 g of carbon dioxide. Hence, it is a limiting reactant.

(iii) According to the question, only 16 g of dioxygen is available. It is a limiting reactant. Thus, 16 g of dioxygen can combine with only 0.5 mole of carbon to give 22 g of carbon dioxide.

Question 1.5:

Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}

Answer 1.5:

0.375 M aqueous solution of sodium acetate

≡ 1000 mL of solution containing 0.375 moles of sodium acetate

∴Number of moles of sodium acetate in 500 mL

$$= \frac{0.375}{1000} \times 500$$
$$= 0.1875 \text{ mole}$$

Molar mass of sodium acetate = 82.0245 g mole^{-1} (Given)

∴ Required mass of sodium acetate = (82.0245 g mol^{-1}) (0.1875 mole)

= 15.38 g

Question 1.6:

Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

Answer 1.6:

Mass percent of nitric acid in the sample = 69 % [Given]

Thus, 100 g of nitric acid contains 69 g of nitric acid by mass.

Molar mass of nitric acid (HNO_3)

= {1 + 14 + 3(16)} g mol^{-1}

= 1 + 14 + 48

= 63 g mol^{-1}

∴ Number of moles in 69 g of HNO_3

$$= \frac{69 \text{g}}{63 \text{g mol}^{-1}}$$
$$= 1.095 \text{mol}$$

Volume of 100g of nitric acid solution

$$= \frac{\text{Mass of solution}}{\text{density of solution}}$$
$$= \frac{100 \text{g}}{1.41 \text{g mL}^{-1}}$$
$$= 70.92 \text{mL} \equiv 70.92 \times 10^{-3} \text{ L}$$

Concentration of nitric acid

$$= \frac{1.095 \text{ mole}}{70.92 \times 10^{-3} \text{ L}}$$
$$= 15.44 \text{ mol/L}$$

∴Concentration of nitric acid = 15.44 mol/L

Question 1.7:

How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

Answer 1.7:

1 mole of $CuSO_4$ contains 1 mole of copper.

Molar mass of $CuSO_4$ = (63.5) + (32.00) + 4(16.00)

= 63.5 + 32.00 + 64.00

= 159.5 g

159.5 g of $CuSO_4$ contains 63.5 g of copper.

⇒ 100 g of $CuSO_4$ will contain $\frac{63.5 \times 100 \text{ g}}{159.5}$ of copper.

∴ Amount of copper that can be obtained from 100 g $CuSO_4$ $= \frac{63.5 \times 100}{159.5}$

= 39.81 g

Question 1.8:

Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molar mass of the oxide is 159.69 g mol^{-1}.

Answer 1.8:

Mass percent of iron (Fe) = 69.9% (Given)

Mass percent of oxygen (O) = 30.1% (Given)

Number of moles of iron present in the oxide $= \frac{69.90}{55.85}$

= 1.25

Number of moles of oxygen present in the oxide $= \frac{30.1}{16.0}$

= 1.88

Ratio of iron to oxygen in the oxide,

$$= 1.25 : 1.88$$

$$= \frac{1.25}{1.25} : \frac{1.88}{1.25}$$

= 1 : 1.5

= 2 : 3

∴ The empirical formula of the oxide is Fe_2O_3.

Empirical formula mass of Fe_2O_3 = [2(55.85) + 3(16.00)] g Molar mass of Fe_2O_3 = 159.69 g

$$\therefore n = \frac{\text{Molar mass}}{\text{Emprical formula mass}} = \frac{159.69 \text{ g}}{159.7 \text{ g}}$$
$$= 0.999$$
$$= 1 \text{(approx)}$$

Molecular formula of a compound is obtained by multiplying the empirical formula with *n*.

Thus, the empirical formula of the given oxide is Fe_2O_3 and *n* is 1.

Hence, the molecular formula of the oxide is Fe_2O_3.

Question 1.9:

Calculate the atomic mass (average) of chlorine using the following data:

	% Natural Abundance	Molar Mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Answer 1.9:

The average atomic mass of chlorine

$$= \left[\begin{pmatrix} \text{Fractional abundance} \\ \text{of } {}^{35}\text{Cl} \end{pmatrix} \begin{pmatrix} \text{Molar mass} \\ \text{of } {}^{35}\text{Cl} \end{pmatrix} + \begin{pmatrix} \text{Fractional} \\ \text{abundance} \\ \text{of } {}^{37}\text{Cl} \end{pmatrix} \begin{pmatrix} \text{Molar mass} \\ \text{of } {}^{37}\text{Cl} \end{pmatrix} \right]$$

$$= \left[\left\{ \left(\frac{75.77}{100} \right) (34.9689\,\text{u}) \right\} + \left\{ \left(\frac{24.23}{100} \right) (36.9659\,\text{u}) \right\} \right]$$

= 26.4959 + 8.9568

= 35.4527 u

∴ The average atomic mass of chlorine = 35.4527 u

Question 1.10:

In three moles of ethane (C_2H_6), calculate the following:

(i) Number of moles of carbon atoms.

(ii) Number of moles of hydrogen atoms.

(iii) Number of molecules of ethane.

Answer 1.10:

(i) 1 mole of C_2H_6 contains 2 moles of carbon atoms.

∴ Number of moles of carbon atoms in 3 moles of C_2H_6

= 2 × 3 = 6

(ii) 1 mole of C_2H_6 contains 6 moles of hydrogen atoms.

∴ Number of moles of carbon atoms in 3 moles of C_2H_6

= 3 × 6 = 18

(iii) 1 mole of C_2H_6 contains 6.023×10^{23} molecules of ethane.

∴ Number of molecules in 3 moles of C_2H_6

$= 3 \times 6.023 \times 10^{23} = 18.069 \times 10^{23}$

Question 1.11:

What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L?

Answer 1.11:

Molarity (M) of a solution is given by,

$$= \frac{\text{Number of moles of solute}}{\text{Volume of solution in Litres}}$$

$$= \frac{\text{Mass of sugar/molar mass of sugar}}{2\text{ L}}$$

$$= \frac{20\text{g} / \left[(12\times12)+(1\times22)+(11\times16)\right]\text{g}}{2\text{ L}}$$

$$= \frac{20\text{g} / 342\text{ g}}{2\text{ L}}$$

$$= \frac{0.0585\text{ mol}}{2\text{ L}}$$

= 0.02925 mol L^{-1}

$\therefore$ Molar concentration of sugar = 0.02925 mol L^{-1}

Question 1.12:

If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution?

Answer 1.12:

Molar mass of methanol (CH_3OH) = (1 × 12) + (4 × 1) + (1 × 16)

= 32 g mol^{-1}

= 0.032 kg mol^{-1}

Molarity of methanol solution $= \frac{0.793\text{ kg L}^{-1}}{0.032\text{ kg mol}^{-1}}$

= 24.78 mol L^{-1}

(Since density is mass per unit volume)

Applying,

$M_1V_1 = M_2V_2$

(Given solution) (Solution to be prepared)

$(24.78 \text{ mol L}^{-1})\ V_1 = (2.5 \text{ L})\ (0.25 \text{ mol L}^{-1})$

$V_1 = 0.0252$ L

$V_1 = 25.22$ mL

Question 1.13:

Pressure is determined as force per unit area of the surface. The SI unit of pressure, Pascal is as shown below:

$1\text{Pa} = 1\text{N m}^{-2}$

If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in Pascal.

Answer 1.13:

Pressure is defined as force acting per unit area of the surface.

$$P = \frac{F}{A}$$

$$= \frac{1034\,\text{g} \times 9.8\ \text{ms}^{-2}}{\text{cm}^2} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{(100)^2\ \text{cm}^2}{1\ \text{m}^2}$$

$= 1.01332 \times 10^5 \text{ kg m}^{-1}\text{s}^{-2}$

We know,

$1 \text{ N} = 1 \text{ kg ms}^{-2}$

Then,

$1 \text{ Pa} = 1 \text{ Nm}^{-2} = 1 \text{ kg m}^{-2}\text{s}^{-2}$ 1

$\text{Pa} = 1 \text{ kg m}^{-1}\text{s}^{-2}$

$\therefore$ Pressure $= 1.01332 \times 10^5$ Pa

Question 1.14:

What is the SI unit of mass? How is it defined?

Answer 1.14:

The SI unit of mass is kilogram (kg). 1 Kilogram is defined as ***the mass equal to the mass of the international prototype of kilogram***.

Question 1.15:

Match the following prefixes with their multiples:

	Prefixes	**Multiples**
(i)	micro	10^6
(ii)	deca	10^9
(iii)	mega	10^{-6}
(iv)	giga	10^{-15}
(v)	femto	10

Answer 1.15:

	Prefix	Multiples
(i)	micro	10^{-6}
(ii)	deca	10
(iii)	mega	10^6
(iv)	giga	10^9
(v)	femto	10^{-15}

Question 1.16:

What do you mean by significant figures?

Answer 1.16:

Significant figures are those meaningful digits that are known with certainty.

They indicate uncertainty in an experiment or calculated value. For example, if 15.6 mL is the result of an experiment, then 15 is certain while 6 is uncertain, and the total number of significant figures are 3.

Hence, significant figures are defined as the total number of digits in a number including the last digit that represents the uncertainty of the result.

Question 1.17:

A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(i) Express this in percent by mass.

(ii) Determine the molality of chloroform in the water sample.

Answer 1.17:

(i) 1 ppm is equivalent to 1 part out of 1 million (10^6) parts.

Mass percent of 15 ppm chloroform in water $= \frac{15}{10^6} \times 100$

$\simeq 1.5 \times 10^{-3}\ \%$

(ii) 100 g of the sample contains 1.5×10^{-3} g of $CHCl_3$.

⇒ 1000 g of the sample contains 1.5×10^{-2} g of $CHCl_3$.

$\therefore$ Molality of chloroform in water

$$= \frac{1.5 \times 10^{-2}\ \text{g}}{\text{Molar mass of } CHCl_3}$$

Molar mass of $CHCl_3$ = 12.00 + 1.00 + 3(35.5)

= 119.5 g mol^{-1}

∴ Molality of chloroform in water = 0.0125×10^{-2} m

= 1.25×10^{-4} m

Question 1.18:

Express the following in the scientific notation:

(i) 0.0048

(ii) 234,000

(iii) 8008

(iv) 500.0

(v) 6.0012

Answer 1.18:

(i) 0.0048 = 4.8×10^{-3}

(ii) 234, 000 = 2.34×10^{5}

(iii) 8008 = 8.008×10^{3}

(iv) 500.0 = 5.000×10^{2}

(v) 6.0012 = 6.0012

Question 1.19:

How many significant figures are present in the following?

(i) 0.0025

(ii) 208

(iii) 5005

(iv) 126,000

(v) 500.0

(vi) 2.0034

Answer 1.19:

(i) 0.0025

There are 2 significant figures.

(ii) 208

There are 3 significant figures.

(iii) 5005

There are 4 significant figures.

(iv) 126,000

There are 3 significant figures.

(v) 500.0

There are 4 significant figures.

(vi) 2.0034

There are 5 significant figures.

Question 1.20:

Round up the following upto three significant figures:

(i) 34.216

(ii) 10.4107

(iii) 0.04597

(iv) 2808

Answer 1.20:

(i) 34.2

(ii) 10.4

(iii) 0.0460

(iv) 2810

Question 1.21:

The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

	Mass of dinitrogen	Mass of dioxygen
(i)	14 g	16 g
(ii)	14 g	32 g
(iii)	28 g	32 g
(iv)	28 g	80 g

(a) Which law of chemical combination is obeyed by the above experimental data? Give its statement.

(b) Fill in the blanks in the following conversions:

(i) 1 km = ... mm = ... pm

(ii) 1 mg = ... kg = ... ng

(iii) 1 mL =..................... L = ... dm^3

Answer 1.21:

(a)

If we fix the mass of dinitrogen at 28 g, then the masses of dioxygen that will combine with the fixed mass of dinitrogen are 32 g, 64 g, 32 g, and 80 g.

The masses of dioxygen bear a whole number ratio of 1:2:2:5. Hence, the given experimental data obeys the law of multiple proportions. The law states that if two elements combine to form more than one compound, then the masses of one element that combines with the fixed mass of another element are in the ratio of small whole numbers.

(b)

(i) 1 km = 1 km × $\frac{1000\text{ m}}{1\text{ km}} \times \frac{100\text{ cm}}{1\text{ m}} \times \frac{10\text{ mm}}{1\text{ cm}}$

$\therefore$ 1 km = 10^6 mm

1 km = 1 km × $\frac{1000\text{ m}}{1\text{ km}} \times \frac{1\text{ pm}}{10^{-12}\text{ m}}$

$\therefore$ 1 km = 10^{15} pm

Hence, 1 km = 10^6 mm = 10^{15} pm

(ii) 1 mg = 1 mg $\frac{1\ \text{g}}{1000\ \text{mg}} \times \frac{1\ \text{kg}}{1000\ \text{g}}$ ×

⇒ 1 mg = 10^{-6} kg

1 mg = 1 mg × $\frac{1\ \text{g}}{1000\ \text{mg}} \times \frac{1\ \text{ng}}{10^{-9}\text{g}}$

⇒ 1 mg = 10^6 ng

$\therefore$ 1 mg = 10^{-6} kg = 10^6 ng

(iii) 1 mL = 1 mL × $\frac{1\ \text{L}}{1000\ \text{mL}}$

⇒ 1 mL = 10^{-3} L

1 mL = 1 cm^3 = 1 $\frac{1\ \text{dm} \times 1\ \text{dm} \times 1\ \text{dm}}{10\ \text{cm} \times 10\ \text{cm} \times 10\ \text{cm}}$ cm^3

⇒ 1 mL = 10^{-3} dm^3

$\therefore$ 1 mL = 10^{-3} L = 10^{-3} dm^3

Question 1.22:

If the speed of light is 3.0×10^8 m s^{-1}, calculate the distance covered by light in 2.00 ns.

Answer 1.22:

According to the question:

Time taken to cover the distance = 2.00 ns

= 2.00×10^{-9} s

Speed of light = 3.0×10^8 ms^{-1}

Distance travelled by light in 2.00 ns

= Speed of light × Time taken

$= (3.0 \times 10^8 \text{ ms}^{-1}) (2.00 \times 10^{-9} \text{ s})$

$= 6.00 \times 10^{-1}$ m

= 0.600 m

Question 1.23:

In a reaction

$A + B_2 \rightarrow AB_2$

Identify the limiting reagent, if any, in the following reaction mixtures.

(i) 300 atoms of A + 200 molecules of B

(ii) 2 mol A + 3 mol B

(iii) 100 atoms of A + 100 molecules of B

(iv) 5 mol A + 2.5 mol B

(v) 2.5 mol A + 5 mol B

Answer 1.23:

A limiting reagent determines the extent of a reaction. It is the reactant which is the first to get consumed during a reaction, thereby causing the reaction to stop and limiting the amount of products formed.

(i) According to the given reaction, 1 atom of A reacts with 1 molecule of B. Thus, 200 molecules of B will react with 200 atoms of A, thereby leaving 100 atoms of A unused. Hence, B is the limiting reagent.

(ii) According to the reaction, 1 mole of A reacts with 1 mole of B. Thus, 2 mole of A will react with only 2 mole of B. As a result, 1 mole of A will not be consumed. Hence, A is the limiting reagent.

(iii) According to the given reaction, 1 atom of A combines with 1 molecule of B. Thus, all 100 atoms of A will combine with all 100 molecules of B. Hence, the mixture is stoichiometric where no limiting reagent is present.

(iv) 1 mole of atom A combines with 1 mole of molecule B. Thus, 2.5 mole of B will combine with only 2.5 mole of A. As a result, 2.5 mole of A will be left as such. Hence, B is the limiting reagent.

(v) According to the reaction, 1 mole of atom A combines with 1 mole of molecule B. Thus, 2.5 mole of A will combine with only 2.5 mole of B and the remaining 2.5 mole of B will be left as such. Hence, A is the limiting reagent.

Question 1.24:

Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:

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

(i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen.

(ii) Will any of the two reactants remain unreacted?

(iii) If yes, which one and what would be its mass?

Answer 1.24:

(i) Balancing the given chemical equation,

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

From the equation, 1 mole (28 g) of dinitrogen reacts with 3 mole (6 g) of dihydrogen to give 2 mole (34 g) of ammonia.

$\Rightarrow$ 2.00×10^3 g of dinitrogen will react with $\frac{6\text{ g}}{28\text{ g}} \times 2.00 \times 10^3\text{ g}$ dihydrogen

i.e., 2.00×10^3 g of dinitrogen will react with 428.6 g of dihydrogen.

Given,

Amount of dihydrogen = 1.00×10^3 g Hence,

N_2 is the limiting reagent.

$\therefore$ 28 g of N_2 produces 34 g of NH_3.

Hence, mass of ammonia produced by 2000 g of N_2 $= \frac{34\text{ g}}{28\text{ g}} \times 2000\text{ g}$

= 2428.57 g

(ii) N_2 is the limiting reagent and H_2 is the excess reagent. Hence, H_2 will remain unreacted.

(iii) Mass of dihydrogen left unreacted = 1.00×10^3 g – 428.6 g
= 571.4 g

Question 1.25:

How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

Answer 1.25:

Molar mass of Na_2CO_3 = (2 × 23) + 12.00 + (3× 16) = 106 g mol^{-1}

Now, 1 mole of Na_2CO_3 means 106 g of Na_2CO_3.

$\therefore$ 0.5 mol of Na_2CO_3 $= \dfrac{106\,g}{1\ mole} \times 0.5\ mol\ Na_2CO_3$

= 53 g Na_2CO_3

⇒ 0.50 M of Na_2CO_3 = 0.50 mol/L Na_2CO_3

Hence, 0.50 mol of Na_2CO_3 is present in 1 L of water or 53 g of Na_2CO_3 is present in 1 L of water.

Question 1.26:

If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Answer 1.26:

Reaction of dihydrogen with dioxygen can be written as:

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

Now, two volumes of dihydrogen react with one volume of dihydrogen to produce two volumes of water vapour.

Hence, ten volumes of dihydrogen will react with five volumes of dioxygen to produce ten volumes of water vapour.

Question 1.27:

Convert the following into basic units:

(i) 28.7 pm

(ii) 15.15 pm

(iii) 25365 mg

Answer 1.27:

(i) 28.7 pm:

1 pm = 10^{-12} m

$\therefore$ 28.7 pm = 28.7×10^{-12} m

= 2.87×10^{-11} m

(ii) 15.15 pm:

1 pm = 10^{-12} m

$\therefore$ 15.15 pm = 15.15×10^{-12} m

= 1.515×10^{-12} m

(iii) 25365 mg:

1 mg = 10^{-3} g

25365 mg = $2.5365 \times 10^{4} \times 10^{-3}$ g

Since,

1 g = 10^{-3} kg

2.5365×10^{1} g = $2.5365 \times 10^{-1} \times 10^{-3}$ kg

$\therefore$ 25365 mg = 2.5365×10^{-2} kg

Question 1.28:

Which one of the following will have largest number of atoms?

(i) 1 g Au (s)

(ii) 1 g Na (s)

(iii) 1 g Li (s)

(iv) 1 g of Cl_2(g)

Answer 1.28:

(i). 1 g of Au (s) $= \frac{1}{197}$ mol of Au (s)

$= \frac{6.022 \times 10^{23}}{197}$ atoms of Au (s)

$= 3.06 \times 10^{21}$ atoms of Au (s)

(ii). 1 g of Na (s) $= \frac{1}{23}$ mol of Na (s)

$= \frac{6.022 \times 10^{23}}{23}$ atoms of Na (s)

$= 0.262 \times 10^{23}$ atoms of Na (s)

$= 26.2 \times 10^{21}$ atoms of Na (s)

(iii). 1 g of Li (s) $= \frac{1}{7}$ mol of Li (s)

$= \frac{6.022 \times 10^{23}}{7}$ atoms of Li (s)

$= 0.86 \times 10^{23}$ atoms of Li (s)

$= 86.0 \times 10^{21}$ atoms of Li (s)

(iv). 1 g of Cl_2 (g) $= \frac{1}{71}$ mol of Cl_2 (g)

(Molar mass of Cl_2 molecule $= 35.5 \times 2 = 71$ g mol^{-1})

$= \frac{6.022 \times 10^{23}}{71}$ atoms of Cl_2 (g)

$= 0.0848 \times 10^{23}$ atoms of Cl_2 (g)

$= 8.48 \times 10^{21}$ atoms of Cl_2 (g)

Hence, 1 g of Li (s) will have the largest number of atoms.

Question 1.29:

Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

Answer 1.29:

Mole fraction of C_2H_5OH $= \dfrac{\text{Number of moles of } C_2H_5OH}{\text{Number of moles of solution}}$

$$0.040 = \frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + n_{H_2O}} \text{...........(1)}$$

Number of moles present in 1 L water:

$$n_{H_2O} = \frac{1000 \text{ g}}{18 \text{ g mol}^{-1}}$$
$$n_{H_2O} = 55.55 \text{ mol}$$

Substituting the value of n_{H_2O} in equation (1),

$$\frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + 55.55} = 0.040$$

$$n_{C_2H_5OH} = 0.040 n_{C_2H_5OH} + (0.040)(55.55)$$
$$0.96 n_{C_2H_5OH} = 2.222 \text{ mol}$$
$$n_{C_2H_5OH} = \frac{2.222}{0.96} \text{ mol}$$
$$n_{C_2H_5OH} = 2.314 \text{ mol}$$

$\therefore$ Molarity of solution $= \dfrac{2.314 \text{ mol}}{1 \text{ L}}$

= 2.314 M

Question 1.30:

What will be the mass of one ^{12}C atom in g?

Answer 1.30:

1 mole of carbon atoms = 6.023×10^{23} atoms of carbon

= 12 g of carbon

$\therefore$ Mass of one ^{12}C atom $= \dfrac{12 \text{ g}}{6.022 \times 10^{23}}$

$= 1.993 \times 10^{-23}$ g

Question 1.31:

How many significant figures should be present in the answer of the following calculations?

(i) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$

(ii) 5×5.364

(iii) $0.0125 + 0.7864 + 0.0215$

Answer 1.31:

(i) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$

Least precise number of calculation = 0.112

$\therefore$ Number of significant figures in the answer

= Number of significant figures in the least precise number = 3

(ii) 5×5.364

Least precise number of calculation = 5.364

$\therefore$ Number of significant figures in the answer = Number of significant figures in 5.364

= 4

(iii) $0.0125 + 0.7864 + 0.0215$

Since the least number of decimal places in each term is four, the number of significant figures in the answer is also 4.

Question 1.32:

Use the data given in the following table to calculate the molar mass of naturally occurring *argon* isotopes:

Isotope	Isotopic molar mass	Abundance
^{36}Ar	35.96755 $gmol^{-1}$	0.337%
^{38}Ar	37.96272 gmol–1	0.063%
^{40}Ar	39.9624 $gmol^{-1}$	99.600%

Answer 1.32:

Molar mass of argon

$$= \left[\left(35.96755 \times \frac{0.337}{100}\right) + \left(37.96272 \times \frac{0.063}{100}\right) + \left(39.9624 \times \frac{90.60}{100}\right)\right] \text{gmol}^{-1}$$

$$= [0.121 + 0.024 + 39.802] \text{ gmol}^{-1}$$

= 39.947 $gmol^{-1}$

Question 1.33:

Calculate the number of atoms in each of the following

(i) 52 moles of Ar **(ii)** 52 u of He **(iii)** 52 g of He.

Answer 1.33:

(i) 1 mole of Ar = 6.022×10^{23} atoms of Ar

$\therefore$ 52 mol of Ar = $52 \times 6.022 \times 10^{23}$ atoms of Ar

= 3.131×10^{25} atoms of Ar

(ii) 1 atom of He = 4 u of He

Or,

4 u of He = 1 atom of He

1 u of He $= \frac{1}{4}$ atom of He

52u of He $= \frac{52}{4}$ atom of He

= 13 atoms of He

(iii) 4 g of He = 6.022×10^{23} atoms of He

$\therefore$ 52 g of He $= \dfrac{6.022 \times 10^{23} \times 52}{4}$ atoms of He

= 7.8286×10^{24} atoms of He

Question 1.34:

A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g.

Calculate

(i) empirical formula,

(ii) molar mass of the gas, and

(iii) molecular formula.

Answer 1.34:

(i) 1 mole (44 g) of CO_2 contains 12 g of carbon.

$\therefore$ 3.38 g of CO_2 will contain carbon $= \dfrac{12 \text{ g}}{44 \text{ g}} \times 3.38 \text{ g}$

= 0.9217 g

18 g of water contains 2 g of hydrogen.

$\therefore$ 0.690 g of water will contain hydrogen $= \dfrac{2 \text{ g}}{18 \text{ g}} \times 0.690$

= 0.0767 g

Since carbon and hydrogen are the only constituents of the compound, the total mass of the compound is:

= 0.9217 g + 0.0767 g

= 0.9984 g

$\therefore$ Percent of C in the compound $= \dfrac{0.9217 \text{ g}}{0.9984 \text{ g}} \times 100$

= 92.32%

Percent of H in the compound $= \dfrac{0.0767 \text{ g}}{0.9984 \text{ g}} \times 100$

= 7.68%

Moles of carbon in the compound $= \dfrac{92.32}{12.00}$

= 7.69

Moles of hydrogen in the compound = $\dfrac{7.68}{1}$

= 7.68

∴ Ratio of carbon to hydrogen in the compound = 7.69: 7.68 = 1: 1

Hence, the empirical formula of the gas is CH.

(ii) Given,

Weight of 10.0L of the gas (at S.T.P) = 11.6 g

∴ Weight of 22.4 L of gas at STP $= \dfrac{11.6 \text{ g}}{10.0 \text{L}} \times 22.4 \text{ L}$

= 25.984 g

≈ 26 g

Hence, the molar mass of the gas is 26 g.

(iii) Empirical formula mass of CH = 12 + 1 = 13 g

$$n = \frac{\text{Molar mass of gas}}{\text{Empirical formula mass of gas}}$$

$$= \frac{26 \text{g}}{13 \text{g}}$$

n = 2

∴ Molecular formula of gas = $(CH)_n$

= C_2H_2

Question 1.35:

Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction, $CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Answer 1.35:

0.75 M of HCl ≡ 0.75 mol of HCl are present in 1 L of water

≡ [(0.75 mol) × (36.5 g mol^{-1})] HCl is present in 1 L of water

≡ 27.375 g of HCl is present in 1 L of water

Thus, 1000 mL of solution contains 27.375 g of HCl.

∴ Amount of HCl present in 25 mL of solution

$$= \frac{27.375\ g}{1000\ mL} \times 25\ mL$$

= 0.6844 g

From the given chemical equation,

$$CaCO_{3(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

2 mol of HCl (2 × 36.5 = 71 g) react with 1 mol of $CaCO_3$ (100 g).

∴ Amount of $CaCO_3$ that will react with 0.6844 g $= \frac{100}{71} \times 0.6844\ g$

= 0.9639 g

Question 1.36:

Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction

$$4HCl_{(aq)} + MnO_{2(s)} \rightarrow 2H_2O_{(l)} + MnCl_{2(aq)} + Cl_{2(g)}$$

How many grams of HCl react with 5.0 g of manganese dioxide?

Answer 1.36:

1 mol [55 + 2 × 16 = 87 g] MnO_2 reacts completely with 4 mol [4 × 36.5 = 146 g] of HCl.

5.0 g of MnO_2 will react with

$= \frac{146\text{ g}}{87\text{ g}} \times 5.0\text{ g}$ of HCl

= 8.4 g of HCl

Hence, 8.4 g of HCl will react completely with 5.0 g of manganese dioxide.