## 1. Details of Module and its structure

| Module Detail | |
|---|---|
| Subject Name | Chemistry |
| Course Name | Chemistry 03 (Class XII, Semester 01) |
| Module Name/Title | Solutions – Part 1<br>(Types of Solution and Units of Concentration) |
| Module Id | lech_10201 |
| Pre-requisites | Knowledge of terms; pure substance, mixture, amount of substance (moles), physical state, etc. |
| Objectives | After going through this lesson, the learners will be able to:<br>• Differentiate among type of solutions.<br>• Describe the formation of different types of solution.<br>• Express and calculate concentration of solution in different units. |
| Keywords | Solute, Solvent, Solution, Concentration of Solution |

## 2. Development Team

| Role | Name | Affiliation |
|---|---|---|
| National MOOC Coordinator (NMC) | Prof. Amarendra P. Behera | CIET, NCERT, New Delhi |
| Program Coordinator | Dr. Mohd. Mamur Ali | CIET, NCERT, New Delhi |
| Course Coordinator (CC) / PI | Prof. Alka Mehrotra<br>Prof. R. K. Parashar | DESM, NCERT, New Delhi<br>DESM, NCERT, New Delhi |
| Course Co-Coordinator / Co-PI | Dr. Aerum Khan | CIET, NCERT, New Delhi |
| Subject Matter Expert (SME) | Ms. Karuna Gupta | Ahlcon International School, Mayur Vihar, Phase -1, Delhi-91 |
| Review Team | Dr. Amit Kumar<br>Dr Shampa Bhattacharya<br>Ms Sarojini Sinha<br>Dr. Aerum Khan | Dayal Singh College, University of Delhi<br>Hans Raj College, University of Delhi<br>Vice-Principal, PGT, Seth Anandram Jaipuria School, Ghaziabad, UP<br>CIET, NCERT, New Delhi |

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## 1. Introduction

In normal life we rarely come across pure substances. Most of these are mixtures containing two or more pure substances. Their utility or importance in life depends on their composition. For example, the properties of brass (mixture of copper and zinc) are quite different from those of German silver (mixture of copper, zinc and nickel) or bronze (mixture of copper and tin); 1 part per million (ppm) of fluoride ions in water prevents tooth decay, while 1.5 ppm causes the tooth to become mottled and high concentrations of fluoride ions can be poisonous (for example, sodium fluoride is used in rat poison); intravenous injections are always dissolved in water containing salts at particular ionic concentrations that match with blood plasma concentrations and so on. Thus, usage or importance of solution in everyday life depends upon its composition and concentration. In this Module, we will consider mostly liquid solutions and their formation. We will begin with types of solutions and then we will discuss various alternatives in which concentration of a solute can be expressed in liquid solution.

## 2. Types of Solutions

**Solutions** are homogeneous mixtures of two or more than two components. By homogenous mixture we mean that its composition and properties are uniform throughout the mixture. Generally, the component that is present in the largest quantity is known as **solvent**. One or more components present in the solution other than solvent are called **solutes**. A solution, depending upon the number of constituents present, is classified as binary, tertiary, quaternary and so on. A binary solution consists of two components whereas three components are present in a tertiary solution. Each component of a solution may be solid, liquid or in gaseous state. Solvent determines the physical state in which solution exists.

**Table 1: Types of Solutions**

| Type of Solution | Solute | Solvent | Common Examples |
|---|---|---|---|
| *Gaseous Solutions* | Gas | Gas | Mixture of oxygen and nitrogen gases |

| | | | |
|---|---|---|---|
| | Liquid<br>Solid | Gas<br>Gas | Chloroform mixed with nitrogen gas<br>Camphor in nitrogen gas |
| *Liquid Solutions* | Gas<br>Liquid<br>Solid | Liquid<br>Liquid<br>Liquid | Oxygen dissolved in water<br>Ethanol dissolved in water<br>Glucose dissolved in water |
| *Solid Solutions* | Gas<br>Liquid<br>Solid | Solid<br>Solid<br>Solid | Solution of hydrogen in palladium<br>Amalgam of mercury with sodium<br>Copper dissolved in gold |

Depending upon the physical state of the solvent the solution can be classified in three different types and these are summarised in Table 1. In this module we shall consider only binary solutions (i.e., consisting of two components).

## 3. *Expressing Concentration of Solutions*

Composition of a solution can be described by expressing its concentration. The latter can be expressed either qualitatively or quantitatively. For example, qualitatively we can say that the solution is dilute (i.e., relatively very small quantity of solute) or it is concentrated (i.e., relatively very large quantity of solute). But in real life these kinds of description can add to lot of confusion and thus requires a quantitative description of the solution.

There are several ways by which we can describe the concentration of the solution quantitatively.

(i) **Mass percentage (w/w):** The mass percentage of a component of a solution is defined as:

$$\text{Mass \% of a component} \quad \frac{Mass\, ofthe\, component \in the\, solution}{Total\, mass\, of\, the\, solution} \times 100 \qquad (1)$$

For example, if a solution is described by 10% glucose in water by mass, it means that 10 g of glucose is dissolved in 90 g of water resulting in a 100 g solution. Concentration described by mass percentage is commonly used in industrial chemical applications. For example, commercial bleaching solution contains 3.62 mass percentage (3.62 % w/w), of sodium hypochlorite in water.

***Example 1:*** Calculate the mass percentage of benzene ($C_6H_6$) and carbon tetrachloride ($CCl_4$) if 22 g of benzene is dissolved in 122 g of carbon tetrachloride.

***Solution:***

Total mass of the solution = Mass of $C_6H_6$ + Mass of $CCl_4$= (22+122) g =144 g

In 144 g of solution, mass of benzene = 122 g

Hence in 100 g of solution, mass percentage of **benzene** = (22/144) × 100 = **15.28 % (w/w)**

Similarly, mass percentage of **carbon tetrachloride** = (122/144) × 100 = **84.72 % (w/w)**

***Example 2:*** A solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass. Calculate the mass percentage of the resulting solution.

***Solution:***

Total mass of solution after mixing = (300 + 400) g = 700 g

Amount of solute present in 300 g of 25% solution = 300 g × (25/100) = 75 g

Amount of solute present in 400 g of 40% solution = 400 g × (40/100) = 160 g

Total mass of solute = (160 + 75) g = 235 g

Hence mass% of resulting solution = (235/700) × 100 = **33.57% (w/w)**

(ii) **Volume percentage (v/v)**: The volume percentage is defined as:

$$\text{Volume \% of a component} \quad \frac{Volumeofthecomponent \in thesolution}{Totalvolumeofthesolution} \times 100 \qquad (2)$$

For example, 10% ethanol solution in water means that 10 ml of ethanol is dissolved in water such that the total volume of the solution is 100 ml. Solutions containing liquids are commonly expressed in this unit. For example, a 35% (v/v) solution of ethylene glycol, antifreeze, is used in cars for cooling the engine. At this concentration the antifreeze lowers the freezing point of water to 255.4K (–17.6°C).

(iii) **Mass by volume percentage (w/v)**: Another unit which is commonly used in medicine and pharmacy is mass by volume percentage. It is the mass of solute dissolved in 100 ml of the solution.

(iv) **Parts per million:** When a solute is present in trace quantities, it is convenient to express concentration in parts per million (ppm) and is defined as:

$$\text{Parts per million} \quad \frac{Numberofpartsofthecomponent}{Totalnumberofpartsofa \ll componentsof\ thesolution} \times 10^6 \qquad (3)$$

As in the case of percentage, concentration in parts per million can also be expressed as mass to mass, volume to volume and mass to volume. A litre of sea water (which weighs 1030 g) contains about 6 $\times 10^{-3}$ g of dissolved oxygen ($O_2$). Such a small concentration is also expressed as 5.8 g per $10^6$ g (5.8 ppm) of sea water. The concentration of pollutants in water or atmosphere is often expressed in terms of μg $mL^{-1}$ or ppm.

(v) **Mole fraction:** Commonly used symbol for mole fraction is x and subscript used on the right hand side of x denotes the component. It is defined as:

Mole fraction of a component $\dfrac{Moles of the component}{Total moles of all the components}$ (4)

For example, in a binary mixture, if the amount (number of moles) of A and B are $n_A$ and $n_B$ respectively, the mole fraction of A will be

$$x_A = \frac{n_A}{n_A + n_B} \tag{5}$$

For a solution containing i number of components, we have:

$$x_A = \frac{n_i}{n_1 + n_2 \ldots\ldots\ldots + n_i} = \frac{n_i}{\sum n_i} \tag{6}$$

It can be shown that in a given solution sum of all the mole fractions is unity, i.e.

$$x_1 + x_2 + \ldots\ldots\ldots\ldots\ldots\ldots + x_i = 1 \tag{7}$$

Mole fraction unit is very useful in relating some physical properties of solutions, say vapour pressure with the concentration of the solution and quite useful in describing the calculations involving gas mixtures.

***Example 3:*** Calculate the mole fraction of ethylene glycol ($C_2H_6O_2$) in a solution containing 20% of $C_2H_6O_2$ by mass.

**Solution**

Assume that we have 100 g of solution (one can start with any amount of solution because the results obtained will be the same). Solution will contain 20 g of ethylene glycol and 80 g of water.

Molar mass of $C_2H_6O_2$ = 12 $gmol^{-1}$ × 2 + 1 $gmol^{-1}$ × 6 + 16 $gmol^{-1}$ × 2 = 62 g $mol^{-1}$.

Moles of $C_2H_6O_2$ = $\dfrac{20\,g}{62\,g\,mol^{-1}} = 0.322\,mol$

Moles of water = $\dfrac{80\,g}{18\,g\,mol^{-1}} = 4.444\,mol$

$$x_{glycol} = \frac{Moles of\, C_2H_6O_2}{Moles of\, C_2H_6O_2 + Moles of\, H_2O}$$

$$x_{glycol} = \frac{0.322\,mol}{0.322\,mol + 4.444\,mol} = 0.068$$

Similarly, $x_{water} = \dfrac{4.444\,mol}{0.322\,mol + 4.444\,mol} = 0.932$

Mole fraction of water can also be calculated as: 1 – 0.068 = 0.932

(vi) **Molarity:** Molarity (M) is defined as amount (or moles) of solute dissolved in one litre (or one cubic decimetre) of solution, Molarity $\dfrac{Molesofsolute}{Volumeofsolution \in litre}$ (8)

For example, 0.25 mol $L^{-1}$ (or 0.25 M) solution of NaOH means that 0.25 mol of NaOH has been dissolved in one litre (or one cubic decimetre).

***Example 4:*** Calculate the molarity of a solution containing 5 g of NaOH in 450 ml solution.

***Solution***

Moles of NaOH = $\dfrac{5\,g}{40\,g\,mol^{-1}} = 0.125\,mol$

Volume of the solution in litres = 450 mL/1000mL $L^{-1}$

Using eq (8), we have

Molarity = $\dfrac{0.125\,mol \times 1000\,mL\,L^{-1}}{450\,mL} = 0.278\,M$

$0.278\,mol\,L^{-1}$

$0.278\,moldm^{-3}$

***Example 5:*** Calculate the molarity of each of the following solutions: (a) 30 g of $Co(NO_3)_2.6H_2O$ in 4.3 L of solution (b) 30 mL of 0.5 M $H_2SO_4$ diluted to 500 mL.

***Solution***

a) Molar mass of $Co(NO_3)_2.6H_2O$ = 270.933 g $mol^{-1}$: moles of cobalt nitrate = 30g /270.933 g $mol^{-1}$

= 0.103 mol

Molarity = $\dfrac{molesofsolute}{Volumeofsolution \in litre}$ = 0.103 mol/4.3 $dm^3$ = 0.0239 mol.$dm^3$ = 0.0239 M

[Since 1L = 1$dm^3$]

b) moles of $H_2SO_4$ present in30 mL of 0.5 M= (0.5/ 1000) × 30 mol

After dilution to 500 mL, the molarity of the solution is given by

Molarity = $\dfrac{molesofsolute}{Volumeofsolution \in litre}$ = $\dfrac{(0.5/1000) \times 30\,mol}{500\,mL/1000\,mL\,dm^{-3}}$ = 0.03 mol.$dm^{-3}$ = 0.03 M

***Example 6:*** Concentrated nitric acid used in laboratory work is 68% nitric acid by mass in aqueous solution. What should be the molarity of such a sample of the acid if the density of the solution is 1.504 g $mL^{-1}$.

***Solution***

Let us consider

Mass of solution = 100g:

Mass of nitric acid in 100g solution = 68g

Volume of 100g of solution = 100g / 1.504 g $mL^{-1}$ = 66.49 ml

Molar mass of $HNO_3$ = 63 g $mol^{-1}$: Moles of $HNO_3$ = 68 g /63 g $mol^{-1}$ = 1.08 mol

Molarity = $\frac{moles of solute}{volume of solution \in litre} = \frac{1.08\,mol \times 1000\,ml\,dm^{-3}}{66.49\,ml}$ = **16.24 M**

(vii) **Molality:** Molality (m) is defined as the amount (moles) of the solute per kilogram (kg) of the solvent and is expressed as:

$$\text{Molality}\quad \frac{Moles of solute}{Mass of solvent \in kg} \tag{9}$$

For example, 1.00 mol $kg^{-1}$ (or 1.00 m) solution of KCl means that 1mol (or 74.5 g) of KCl is dissolved in 1 kg of water.

***Example 7***: Calculate molality of 2.5 g of ethanoic acid ($CH_3COOH$) in 75 g of benzene.

***Solution***

Molar mass of $C_2H_4O_2$ = 12 $gmol^{-1}$ × 2 + 1 $gmol^{-1}$ × 4 + 16 $gmol^{-1}$ × 2 = 60 g $mol^{-1}$.

Moles of $C_2H_4O_2$ = $\frac{2.5\,g}{60\,g\,mol^{-1}} = 0.417\,mol$

Mass of benzene in kg = 75g × $10^{-3}$ kg.$g^{-1}$ = 75 × $10^{-3}$ kg

Molality of $C_2H_4O_2$ = $\frac{Moles of\ C_2H_4O_2}{Mass of benzene\,(kg)} = \frac{0.417\,mol}{75 \times 10^{-3}\,kg}$ = 5.56 mol $kg^{-1}$

= 5.56 m

Each method of expressing concentration of the solutions has its own merits and demerits. Mass %, ppm, mole fraction and molality are independent of temperature, whereas molarity is a function of temperature. This is because volume depends on temperature and the mass does not.

***Example 8***: Calculate the mass of urea ($NH_2CONH_2$) required in making 2.5 kg of 0.25 molal aqueous solution.

***Solution***

Molar mass of urea = 60 g $mol^{-1}$

Mass of urea= (60 g $mol^{-1}$ × 0.25 mol $kg^{-1}$ × 2.5 kg) = 37.5 g

## 4. Relationship between different units of concentrations

Of the various units of concentration discussed above molality, molarity and mole fraction are commonly used to express the composition of solution. At times, we require inter conversion of these units so as to be able to solve a numeric problem at hand. In this section we try to establish the relationship between commonly used concentration units.

***(i)*** *Relationship between molarity and molality of solution*

Let us consider

Moles of solute in solution = $n_2$

Volume of solution = V

Molar mass of solute = $M_2$

Density of solution = ρ

Now Mass of solute = $n_2 \times M_2$

Mass of solution = V × ρ

Therefore mass of solvent = Mass of solution – Mass of solute

$= V \times \rho - n_2 \times M_2$

The molality of solution is given by

$$molality = \frac{moles of solute}{mass of solvent \in kg unit}$$

i.e.

$$m = \frac{n_2}{V\rho - n_2 M_2}$$

Dividing numerator and denominator of the above equation by V, we get

$$m = \frac{n_2 / V \,(mol)}{\frac{V}{V} \times \rho - \frac{n_2}{V} M_2}$$

$$m = \frac{M}{\rho - M M_2}$$

**Note:** The molarity "*M*" has unit *mol.dm*$^{-3}$, the density of solution " $\rho$ " should be expressed in *Kg.dm*$^{-3}$ unit and the molar mass of solute $M_2$ should expressed in *Kg.mol*$^{-1}$ unit as this will result in molality unit i.e. *mol.Kg*$^{-1}$ of the right hand side (rhs) expression.

***(ii)*** *Relationship between molality and mole fraction*

Let $x_1$ and $x_2$ be the mole fraction of solvent and solute in the solution, respectively.

i.e.

$$x_1 = \frac{n_1}{n_1 + n_2}$$

$$x_2 = \frac{n_2}{n_1 + n_2}$$

Here $n_1$ and $n_2$ are no. of moles of solvent and solute in he solution, respectively.

Now the molality (m) of solution is given by

$$m = \frac{mlesofsolute}{massofsolvent \in kgunit}$$

$$m = \frac{n_2}{n_1 \times M_1}$$

dividing the numerator and denominator of rhs of above equation, we get

$$m = \frac{\frac{n_2}{n_1 + n_2}}{\frac{n_1}{n_1 + n_2} \times M_1}$$

$$m = \frac{x_2}{x_1 \times M_1}$$

**Note:** The unit of molar mass of solvent "$M_1$" should be expressed in *Kg.mol*$^{-1}$ unit consequently the rhs expression will yield the unit of molality.

***(iii)*** *Relationship between molarity and mole fraction*

Let M be the molarity of solution having $n_2$ moles of solute present in $1dm^3$ of solution. The mole fraction $x_1$ and $x_2$ of solvent and solute, respectively are given by

$$x_1 = \frac{n_1}{n_1 + n_2}$$

$$x_2 = \frac{n_2}{n_1 + n_2}$$

Here $n_1$ and $n_2$ are moles of solvent and solute, respectively

The mass of solution can be calculated by the following expression

Mass of solution = $n_1 \times M_1 + n_2 \times M_2$

Here $M_1$ and $M_2$ are the molar masses of solvent and solute, respectively

Now using the density $\rho$ of the solution can be calculated as

Volume of solution = $\dfrac{n_1 \times M_1 + n_2 \times M_2}{\rho}$

Therefore, the molarity of solution is given by

Molarity "M" = $\dfrac{molesofsolute}{Volumeofsolution \in liters}$

Molarity = $\dfrac{n_2}{\dfrac{n_1 \times M_1 + n_2 \times M_2}{\rho}}$

dividing numerator and denominator of the above equation by $n_1 + n_2$ we get

Molarity = $\dfrac{\dfrac{n_2}{n_1 + n_2}}{\dfrac{n_1 \times M_1 + n_2 \times M_2}{\rho(n_1 + n_2)}}$

Molarity = $\dfrac{\dfrac{n_2}{n_1 + n_2}\rho}{\dfrac{n_1}{n_1 + n_2} \times M_1 + \dfrac{n_2}{n_1 + n_2} \times M_2}$

**Molarity =** $\frac{x_2\rho}{x_1\times M_1+x_2\times M_2}$

**Note:** Here $x_1$ and $x_2$ are dimensionless quantity and it must be noted that the density of solution " $\rho$ " should be expressed in $g.dm^{-3}$ unit and molar mass of solvent and solute ($M_1$ and $M_2$) should be expressed in $g.mol^{-1}$ unit such that the rhs of the above expression has the unit of molarity i.e. $mol.dm^{-3}$.

***Example 9***: What is the molality of an aqueous solution of NaCl in which amount fraction of NaCl is 0.03

***Solution***

For calculation of molality of the solution we may consider the relation; $m=\frac{x_2}{x_1 M_1}$

Here

$x_2$ is the mole fraction of solute

$x_1$ is the mole fraction of solvent

$M_1$ is the molar mass of the solvent

Thus, we have

$x_2 = 0.03$ and $x_1 = 0.97$

$M_1 = 18\ g\ mol^{-1}$

Using the given data, we have

$$m=\frac{x_2}{x_1 M_1}=\frac{0.03}{0.97\times 18\,g\,mol^{-1}}=1.718\times 10^{-3}\,mol\,g^{-1}$$

$$1.718\times 10^{-3}\,mol\left(10^{-3}\,Kg\right)^{-1}$$

$$1.718\,mol\,Kg^{-1}$$

***Example 10***: In a liquid solution of ethanol in water, the mole fraction of ethanol is 0.03, and the density of solution is 0.994 g $cm^{-3}$. Determine the molarity of the solution.

***Solution***

Mole fraction of solute; $x_2 = 0.03$

Density of solution; $\rho$ = 0.994 g $cm^{-3}$

Molar mass of ethanol; $C_2H_5OH$ = 12 g $mol^{-1}\times 2$ + 1 g $mol^{-1}\times 6$ + 16 g $mol^{-1}\times 1$ = 46 $gmol^{-1}$

We know that molar mass of water ($M_2$) is 18 $gmol^{-1}$

Now the molarity (M) of the solution may be calculated by the formulae

$$M = \frac{x_2 \rho}{x_1 M_1 + x_2 M_2}$$

Thus, we have

$$M = \frac{0.03 \times 0.994\, g\, cm^{-3}}{0.97 \times (18\, g\, mol^{-1}) + 0.03 \times (46\, g\, mol^{-1})}$$

$$M = 1.582 \times 10^{-3}\, mol\, (10^{-1}\, dm)^{-3} = 1.582 \times mol\, dm^{-3}$$

***Example 11***: What is the mole fraction of ethanol in an aqueous solution of ethanol, whose molal concentration is 3.6 m

***Solution***

Mole fraction of methanol can be calculated using the expression

$$x_2 = \frac{m M_1}{1 + M M_1}$$

Given data

Molality ( m ) of methanol in solution = 3.6 molKg$^{-1}$

Molar mass of $H_2O$ ($M_1$) = 18 g mol$^{-1}$

Using given data in eq (i), we have

$$x_2 = \frac{3.6\, mol\, Kg^{-1} \times 18\, g\, mol^{-1}}{1 + (3.6\, mol\, Kg^{-1} \times 18\, g\, mol^{-1})}$$

$$1 + (3.6\, mol\, Kg^{-1}) \times (18 \times 10^{-3}\, Kg\, mol^{-1})$$

$$x_2 = \frac{3.6\, mol\, Kg^{-1} \times (18 \times 10^{-3}\, Kg\, mol^{-1})}{1.064}$$

$$x_2 = \frac{0.064}{1.064} = 0.06$$

***Example 12***: A sample of drinking water was found to be severely contaminated with chloroform ($CHCl_3$) which is supposed to be a carcinogenic, the level of contamination was 15 ppm (by mass), determine the molality of chloroform in the water sample.

***Solution***

Molar mass of chloroform = 119.5 g mol$^{-1}$

The concentration of the solution is 15 ppm which means that $15 \times 10^{-3}$ g of chloroform is present in 1 kg of solvent i.e. water

Thus, moles of $CHCl_3$ present in 1 kg of solvent i.e. Molality of solution =

$$\frac{15 \times 10^{-3} g\, kg^{-1}}{119.5\, g\, mol^{-1}} = 1.25 \times 10^{-4} \text{ mol kg}^{-1}$$

$= 1.25 \times 10^{-4}$ m

***Example 13***: Calculate (a) molality (b) molarity and (c) mole fraction of KI if the density of 20% w/w aqueous KI solution is 1.202 g $mL^{-1}$.

***Solution***

Let us consider, mass of solution = 100 g

Mass of KI = 20 g, mass of water = 80 g : moles of water = 80g/18 $g\, mol^{-1}$ = 4.444 mol

Molar mass of KI = 165.9 g $mol^{-1}$: moles of KI = 20g/165.9g$mol^{-1}$ = 0.1205 mol

volume of solution = 100g/1.202 g $mL^{-1}$ = 83.1946 ml

$$\text{Molality} = \frac{moleso\, fsolute}{massofsolvent \in kg} = \frac{0.1205\, mol}{80\, g \times 10^{-3} kg\, g^{-1}} = 1.51 \text{ mol kg}^{-1} = 1.51 \text{ m}$$

$$\text{Molarity} = \frac{molesofsolute}{volumeofsolution \in litre} = \frac{0.1205\, mol}{83.1946\, mL \times 10^{-3} dm^3 mL^{-1}} = \mathbf{1.45\ mol\ dm^3 = 1.45\ M}$$

Mole fraction of KI = 0.1205 mol / (0.1205mol + 4.44mol) = 0.1205 mol/4.5605 mol = **0.026**

***Example 14***: How many mL of 0.1 M HCl are required to react completely with 1 g mixture of $Na_2CO_3$ and $NaHCO_3$ containing equimolar amounts of both?

***Solution***

Molar mass of $Na_2CO_3$= 106 g $mol^{-1}$ : Molar mass of $NaHCO_3$ = 84 g $mol^{-1}$

Let the mass of $Na_2CO_3$ = x g, therefore mass of $NaHCO_3$= (1-x) g

Moles of $Na_2CO_3$ = x g/106 g $mol^{-1}$ = x/106 mol

Moles of$NaHCO_3$= (1-x) g /84 g $mol^{-1}$ = (1-x) /84 mol

For an equimolar mixture, we have **Moles of $Na_2CO_3$ = Moles of$NaHCO_3$**

Therefore,

$$x/106 \text{ mol} = (1-x)/84 \text{ mol}$$
$$84\, x = 106\,(1 - x)$$
$$84\, x = 106 - 106\, x$$
$$190x = 106$$
$$x = 0.558 \text{ g}$$

mass of $Na_2CO_3$ = 0.558 g and mass of $NaHCO_3$= 0.442 g

moles of $Na_2CO_3$ = 0.00526 mol

moles of $NaHCO_3$ = 0.00526 mol

Let us consider the stoichiometric reaction between HCl and $Na_2CO_3\, NaHCO_3$

$$Na_2CO_3 + 2\,HCl \rightarrow 2\,NaCl + H_2CO_3$$

$$NaHCO_3 + HCl \rightarrow NaCl + H_2CO_3$$

Thus, for complete neutralization 1 mol of $Na_2CO_3$ requires 2 mol of HCl and 1 mole of $NaHCO_3$ requires 1 mol of HCl.

It therefore means that

0.00526 mol of $Na_2CO_3$ requires 2 × 0.00526 mol = 0.01052 mol of HCl.

0.00526 mol of $NaHCO_3$ requires 0.00526 mol of HCl.

Therefore, for complete neutralization of the mixture of 0.00526 moles of $NaHCO_3$ and 0.00526 moles of $Na_2CO_3$, HCl required is 0.01052 mol + 0.00526 mol = 0.01578 moles.

Thus fr complete neutralisation, volume of 0.1 M HCl required = $\frac{0.01578\,mol}{0.1\,mol\,dm^{-3}}$ = 0.1578 $dm^3$

= 0.1578 $dm^3 \times 10^3$ $mL.dm^{-3}$

= **157.8 ml**

***Example 15***: A solution of glucose in water is labelled as 10% w/w, what would be the molality and mole fraction of each component in the solution? If the density of solution is 1.2 g $mL^{-1}$, then what shall be the molarity of the solution?

***Solution***

Let the mass of solution is 100 g

Therefore mass of glucose in 100 g of solution = 10 g

Mass of water in 100g of solution = 90 g

Molar mass of glucose = 180 g $mol^{-1}$

Moles of water = 90g /18 $gmol^{-1}$ = 5 mol

Moles of glucose = 10g/180$gmol^{-1}$ = 0.055 mol

Molality = $\frac{molesofsolute}{massofsolvent \in kg}$ = $\frac{0.055\,mol}{90\,g \times 10^{-3}\,kg\,g^{-1}}$ = 0.617$mol.kg^{-1}$= 0.617 **m**

Volume of solution = 100g / 1.2 g $mL^{-1}$ = 83.33 ml

Molarity = $\frac{molesofsolute}{volumeofsolution \in litre}$ = $\frac{0.055\,mol}{83.33\,mL \times 10^{-3}\,dm^3\,mL^{-1}}$ = 0.66 $mol.dm^{-3}$ = **0.66 M**

Total moles = moles of glucose + moles of water = 0.055 + 5 = 5.055

Mole fraction of glucose = $\dfrac{moles of component}{Total moles of all the components}$ = $\dfrac{0.055\,mol}{5.055\,mol}$ = **0.01**

Mole fraction of water = $\dfrac{moles of component}{Total moles of all the components}$ = 5 mol /5.055 mol = **0.99**

***Example 16***: An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) and 200 g of water. Calculate the molality of the solution. If the density of the solution is 1.072 g $mL^{-1}$, then what shall be the molarity of the solution?

***Solution***

Molar mass of ethylene glycol = 62 g $mol^{-1}$

Molar mass of water = 18 g $mol^{-1}$

Moles of ethylene glycol = 222.6 g / 62 g $mol^{-1}$ = 3.59 mol

Moles of water = 200 g /18 g $mol^{-1}$ = 11.11 mol

Molality = $\dfrac{moles of solute}{mass of solvent \in kg}$ = $\dfrac{3.59\,mol}{200\,g \times 10^{-3}\,kg\,g^{-1}}$ = 17.95 mol.$kg^{-1}$ = **17.95 m**

Total mass of solution = 222.6 g + 200 g = 422.6 g

Volume of solution = 422.6 g / 1.072 g.$mL^{-1}$ = 394.216 ml

Molarity = $\dfrac{moles of solute}{volume of solution \in litre}$ = $\dfrac{3.59\,mol}{394.216\,mL \times 10^{-3}\,dm^3\,mL^{-1}}$ = 9.107 mol.$dm^{-3}$

= **9.107 M**