Solutions - Part 1

Objectives

After going through this lesson, the learners will be able to understand the following:

- Differentiate among types of solutions.
- Describe the formation of different types of solution.
- Express and calculate concentration of solution in different units.

Contents Outline

- Introduction
- Types of Solution
- Expressing Concentration of Solution
- Relationship Between Different Units of Concentrations

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.

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 a 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	Gas	Chloroform mixed with nitrogen gas
	Solid	Gas	Camphor in nitrogen gas
Liquid Solutions	Gas	Liquid	Oxygen dissolved in water
	Liquid	Liquid	Ethanol dissolved in water
	Solid	Liquid	Glucose dissolved in water
Solid Solutions	Gas	Solid	Solution of hydrogen in palladium
	Liquid	Solid	Amalgam of mercury with sodium
	Solid	Solid	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).

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 descriptions can add to a lot of confusion and thus require 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:

$$Mass\ \%\ of\ a\ component\ = \frac{Mass\ of\ the\ component\ of\ the\ solution}{Total\ mass\ of\ the\ solution} \times 100 \qquad (1)$$

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,

$mass\ percentage\ of\ benzene\ = (22/144) \times 100\ =\ 15.28\ \%\ (w/w)$

Similarly,

$mass\ percentage\ of\ carbon\ tetrachloride\ = (122/144) \times 100\ =\ 84.72\ \%\ (w/w)$

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:

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

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

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

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

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

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

$$Volume\ \%\ of\ a\ component\ = \frac{Volume\ of\ the\ component\ the\ solution}{Total\ volume\ of\ the\ solution} \times 100 \qquad (2)$$

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 255.4 K (–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:

$$Parts\ per\ million\ = \frac{Number\ of\ parts\ of\ the\ component}{Total\ number\ of\ parts\ of\ components\ of\ the\ solution} \times 10^{6} \qquad (3)$$

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 6×10^{-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:

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

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

$$x_A = \frac{n_A}{n_A + n_B} \qquad (5)$$

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

$$x_A = \frac{n_i}{n_1 + n_2 .... + n_i} = \frac{n_i}{\Sigma n_i} \qquad (6)$$

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

$$x_1 + x_2 + ... + x_i = 1 \qquad (7)$$

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.

$$Molar\ mass\ of\ C_2H_6O_2 = 12\ g\ mol^{-1} \times 2 + 1\ g\ mol^{-1} \times 6 + 16\ g\ mol^{-1} \times 2$$

$$= 62\ g\ mol^{-1}$$

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

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

$$x_{glycol} = \frac{moles\ Of\ C_2H_6O_2}{moles\ of\ C_2H_6O_2 + moles\ of\ H_2O}$$

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

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

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,

$$Molarity = \frac{Moles\ of\ solute}{Volume\ of\ solution\ in\ litre} \quad (8)$$

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:

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

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

Using eq (8), we have

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

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

$= 0.278\ mol\ dm^{-3}$

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 = 30 g /270.933 g mol^{-1}

= 0.103 mol

$$Molarity = \frac{moles\ of\ solute}{Volume\ of\ solution\ in\ litre}$$

$$= 0.103\ mol/4.3\ dm^3 = 0.0239\ mol.\ dm^{-3} = 0.0239\ M$$

[Since 1L = 1dm^3]

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

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

$$Molarity = \frac{moles\ of\ solute}{Volume\ of\ solution\ in\ litre} = \frac{\left(\frac{0.5}{1000}\right)\times 30\ mol}{500\frac{mL}{1000}\ mL\ dm^{-3}} = 0.03\ \text{mol.dm}^{-3} = 0.03\ \text{M}$$

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 = 100 g

Mass of nitric acid in 100g solution = 68 g

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

$$Molar\ mass\ of\ HNO_3 = 63\ g\ mol^{-1}$$

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

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

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

$$Molality = \frac{Moles\ of\ solute}{Mass\ of\ solvent\ in\ kg} \quad (9)$$

For example, 1.00 mol kg^{-1} (or 1.00 m) solution of KCl means that 1 mol (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

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

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

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

$$Molality\ of\ C_2H_4O_2 = \frac{Moles\ of\ C_2H_4O_2}{Mass\ of\ benzene\ (kg)} = \frac{0.417\ mol}{75\times10^{-3}\ kg} = 5.56\ \text{mol kg}^{-1}$$

$$= 5.56\ \text{m}$$

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}

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

Relationship Between Different Units of Concentrations

Among 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 $Mass\ of\ solute \ = \ n_2 \times M_2$

$Mass\ of\ solution \ = \ V \times \rho$

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

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

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

Note: The molarity "M" has unit mol.dm^{-3}, the density of solution "ρ" should be expressed in kgdm^{-3} units and the molar mass of solute M_2 should be expressed in kgmol^{-1} unit as this will result in a molality unit i.e. molkg^{-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 the solution, respectively.

Now the molality (m) of solution is given by

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

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

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 *kgmol*$^{-1}$ unit consequently the RHS expression will yield the unit of molality.

(iii) Relationship between molarity and mole fraction

Let M be the molarity of a solution having n_2 moles of solute present in 1 dm^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 ρ of the solution can be calculated as

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

Therefore, the molarity of solution is given by

$$Molarity\ \text{“}M\text{”} = \frac{moles\ of\ solute}{Volume\ of\ solution\ in\ liters}$$

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

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

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

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

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

Note: Here x_1 and x_2 are dimensionless quantity and it must be noted that the density of solution "ρ" 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;ρ = 0.994 g cm^{-3}

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

$$= 46\ g\ mol^{-1}$$

We know that molar mass of water (M_2) is 18 g mol^{-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

$$x_2 = \frac{mM_1}{1+MM_1}$$

Given data

Molality (m) of methanol in solution = 3.6 mol kg^{-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 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 15×10^{-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. $Molality\ of\ solution = \frac{15 \times 10^{-3}\ g\ kg^{-1}}{119.5\ g\ mol^{-1}}$

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

= 1.25×10^{-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

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

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

$$Molality\ = \frac{moles\ of\ solute}{mass\ of\ solvent\ in\ kg} = \frac{0.1205\ mol}{80g\times10^{-3}kg\ g^{-1}} = 1.51\ mol\ kg^{-1}\ =\ 1.51\ m$$

$$Molarity\ = \frac{moles\ of\ solute}{volume\ of\ solution\ in\ litre} = \frac{0.1205\ mol}{83.1946\ mL\times10^{-3}dm^{3}mL^{-1}}$$

$$= 1.45\ mol.dm^{-3}\ =\ 1.45\ M$$

$Mole\ fraction\ of\ KI\ =\ 0.1205\ mol\ /\ (0.1205\ mol\ +\ 4.44\ mol)$

$=\ 0.1205\ mol/4.5605\ mol\ =\ 0.026$

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

$Moles\ of\ Na2CO_3\ =\ x\ g/106\ g\ mol^{-1}\ =\ x/106\ mol$

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

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

Therefore,

$x/106\ mol\ =\ (1\ -\ x)/84\ mol$

$84\ x\ =\ 106\ (1 - x)$

$84\ x\ =\ 106 - 106x$

$190x\ =\ 106$

$x\ =\ 0.558\ g$

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 $Na_2CO_3\ NaHCO_3$

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

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

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

$0.00526\, mol\, of\, Na2CO_3\, requires\, 2 \times 0.00526\, mol = 0.01052\, mol\, of\, HCl.$

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 for 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}

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

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

$$molality = \frac{moles\, of\, solute}{mass\, of\, solvent\, in\, kg} = \frac{0.055\, mol}{90g\times10^{-3}kg\, g^{-1}} = 0.617 mol.kg^{-1} = 0.617\, m\mathrm{L}$$

$$molarity = \frac{moles\, of\, solute}{volume\, of\, solution\, in\, litre} = \frac{0.055\, mol}{83.33mL\times10^{-3}dm^3mL^{-1}} = 0.66\, mol.dm^{-3} = 0.66\, M$$

$$Total\, moles = moles\, of\, glucose + moles\, of\, water = 0.055 + 5 = 5.055$$

$$mole\, fraction\, of\, glucose = \frac{moles\, of\, component}{Total\, moles\, of\, all\, the\, components} = \frac{0.055\, mol}{5.055\, mol} = 0.01$$

$$mole\, fraction\, of\, water = \frac{moles\, of\, component}{Total\, moles\, of\, all\, the\, components} = 5\, mol\, /5.055\, mol = 0.99$$

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}

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

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

$$molality\ = \frac{moles\ of\ solute}{mass\ of\ solvent\ in\ kg} = \frac{3.59\ mol}{200g \times 10^{-3} Kk\ g^{-1}} = 17.95\ \text{mol.kg}^{-1} = \mathbf{17.95\ m}$$

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

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

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

= **9.107 M**