**"Cultivating excellence in every student"**

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## Chapter-2: SOLUTIONS

# Expressing concentration of solutions

**SOLUTIONS are homogeneous mixtures in which the substances are so thoroughly mixed that they cannot be distinguished from one another**.
Most solutions contain a solid **(solute)** dissolved in a liquid **(solvent)**; however, there are solutions of gases as well.
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 the need for a quantitative description of the solution.***
There are several ways by which we can describe the concentration of the solution **quantitatively.**

The **CONCENTRATION** of a substance in solution provides away to find **how much of the substance exists in a given volume of the solution.**
**Chemists use the "mole" to describe the amount of substance in a solution**
A majority of reactions in the laboratories are carried out in solutions. Therefore, it is important to understand as how the amount of substance is expressed when it is present in the form of a solution. **The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the following ways.**
**1. Mass per cent or weight per cent (w/w %)**
**2. Mole fraction**
**3. Molarity (M)**
**4. Molality (m)**
**5. Normality (N)**
**6. Parts per million (ppm)**

**1. <u>Mass per cent or weight per cent</u> (w/w %)**
Weight percent is often used to express the concentration of a solid substance dissolved in a liquid. The weight percent is equal to the weight of the substance (solute) divided by the total weight of the solution and multiplied by 100 to get a percentage.

$$\text{Weight percent} = \frac{\text{weight of solute}}{\text{total weight of solution}} \times 100$$

For example, a 1% sodium chloride solution in water contains 1 gram of sodium chloride (NaC*l*) in a total of 100 grams of solution. The amount of water is 99 grams, the difference between the total weight (100 grams) and the weight of the solute (1 gram).

**<u>Volume Percent</u>**
Volume percent is often used to express the concentration of a liquid solute in a liquid solvent. The volume percent is equal to the volume of the solvent divided by the total volume of the solution and multiplied by 100 to get a percentage.

$$\textbf{Volume percent} = \frac{\textbf{Volume of solute}}{\textbf{total volume of solution}} \textbf{x 100}$$

Common rubbing alcohol is a 70% solution of rubbing alcohol (isopropyl alcohol) in water; that is, it contains 70 mL of isopropyl alcohol in 100 mL of total solution.

**2. <u>Mole fraction</u>**
The mole fraction is moles of target substance divided by total moles involved. The symbol for the mole fraction is the lower-case Greek letter chi, $\chi$. It can be given with a subscript: $\chi_{\text{solute}}$ is an example.

**Problem:** 0.100 mole of NaCl is dissolved into 100.0 grams of pure $H_2O$. What is the mole fraction of NaCl?

Sol.

$$\frac{100.0\ \text{g}}{18.0\ \text{g mol}^{-1}} = \textbf{5.56 mol of } \mathbf{H_2O}$$

Add that to the 0.100 mol of NaCl = 5.56 + 0.100 = 5.66 mol total

$$\text{Mole fraction of NaCl} = \frac{0.100\ \text{mol}}{5.66\ \text{mol}} = 0.018$$

$$\text{The mole fraction of the } H_2O = \frac{5.56\ \text{mol}}{5.66\ \text{mol}} = 0.982$$

**Another way to figure out the last substance is 1.00 minus (the total of all other mole fractions). In this case 1.00 - 0.018 = 0.982.**
**Remember that all the mole fractions in the solution should total up to one.**

**Notice that the mole fraction has no units on it and is written as a decimal value. Do not change it to percent.**

**Problem:** A solution is prepared by mixing 25.0 g of water, $H_2O$, and 25.0 g of ethanol, $C_2H_5OH$. Determine the mole fractions of each substance.

**Solution:**

1) Determine the moles of each substance:

$$H_2O \Rightarrow \frac{\mathbf{25.0\ g}}{\mathbf{18.0\ g/mol}} = 1.34\ \text{mol}$$

$$C_2H_5OH \Rightarrow \frac{\mathbf{25.0\ g}}{\mathbf{46.07\ g/mol}} = 0.543\ \text{mol}$$

2) Determine mole fractions:

$$H_2O \Rightarrow \frac{\mathbf{1.34\ mol}}{\mathbf{(1.34\ mol + 0.543\ mol)}} = 0.71$$

$$C_2H_5OH \Rightarrow \frac{\mathbf{0.543\ mol}}{\mathbf{(1.34\ mol + 0.543\ mol)}} = 0.29$$

## <u>Problems for practice</u>

1. **A solution contains 10.0 g pentane, 10.0 g hexane and 10.0 g benzene. What is the mole fraction of hexane?**
2. **How many grams of water must be used to dissolve 100.0 grams of sucrose ($C_{12}H_{22}O_{11}$) to prepare a 0.020 mole fraction of sucrose in the solution?**
3. **Surprisingly, water (in the form of ice) is slightly soluble in liquid nitrogen. At -196 °C, (the boiling point of liquid nitrogen) the mole fraction of water in a saturated solution is $1.00 \times 10^{-5}$. Compute the mass of water that can dissolve in 1.00 kg of boiling liquid nitrogen.**
4. **What is the mole fraction of cinnamic acid in a mixture that is 50.0% weight urea in cinnamic acid (urea = 60.06 g/mol; cinnamic acid = 148.16 g/mol).**

**3. <u>MOLAR CONCENTRATION or MOLARITY (M)</u>** of a substance is the number of moles of the substance contained in **1 L of solution**.

$$\text{i.e. Molarity} = \frac{\textbf{No.of moles of solute}}{\textbf{Volume of solution in litres}}$$

$$\text{Molarity of the solution} = \frac{W_B}{M_B \times V}$$

If V is taken in ml ($cm^3$), then

$$\text{Molarity of the solution} = \frac{W_B}{M_B \times V} \times 1000$$

**The unit of molarity is mol $litre^{-1}$ or mol $dm^{-3}$**

e.g. If 2.0 L of solution contain 5.0 mol of NaCl, what is the molarity of the NaCl?

$$\text{molar concentration} = \frac{5.0\ \text{mol}}{2.0\ \text{L}} = 2.5\ \frac{\text{mol}}{\text{L}}$$

$$\text{The unit symbol for } \frac{\text{mol}}{\text{L}} \text{ is “M”}$$

When expressed in words, the unit symbol **“M”** is written as **“molar”.**

The short-hand symbol for **“molar concentration of …”** is a set of brackets: **[…]**, so **[NaCl]** means the **“molar concentration of NaCl.”**

**If density’s’ of the solution in gm/ml and mass % (x %) of solute is given then the molarity can be calculated as**

$$\text{As density‘d’ of solution} = \frac{\text{mass of solution}}{\text{volume of solution}}$$

$$\therefore \text{Mass of 1 litre solution} = (1000 \times d)\ \text{g}$$

$$\text{Or, mass of solute in 1 litre} = \frac{x}{100} \times (1000 \times d)\ \text{g} = (x \times d \times 10)\ \text{g}$$

$$\text{Number of moles of solute in 1 litre} = \frac{\text{Mass of solute in grams}}{\text{Gram molecular mass of solute}} = \frac{x \times d \times 10}{M_B}$$

Where $M_B$ is gram molecular mass of solute

$$\therefore \mathbf{M} = \frac{x \times d \times 10}{M_B}$$

**In case of dilution, the molarity can be calculated by using dilution formula which can be given as,**

$$\mathbf{M_1V_1 = M_2V_2}$$

Where $M_1$, $M_2$, $V_1$ and $V_2$ are the molarities and volumes respectively.

**Molarity of mixture:**

When there are two or more samples of solution containing same solute and solvent with molarities $M_1$, $M_2$ ….and volumes $V_1$, $V_2$ …respectively.

**After mixing the molarity of mixed solution can be calculated as**,

$$\mathbf{M_1V_1 + M_2V_2 + \ldots = M_R (V_1 + V_2 + \ldots)}$$

**Or**

$$M_R = \frac{M_1V_1+M_2V_2+\cdots}{(V_1+V_2+\cdots)}$$

**Where, $M_R$ = resultant molarity or molarity of mixture.**

## Problems

| | |
|---|---|
| *Problem:* | What mass of NaOH is contained in 3.50 L of 0.200 M NaOH? |
| *Solution:* | The molarity (M) and volume (V) are given so moles can be found. Moles can then be converted to mass. <br> solving $M = \frac{mol}{V}$ for n gives mol = M x V <br> moles of NaOH = $0.200 \frac{mol}{L} \times 3.50\ L = 0.700$ mol <br> mass of NaOH = $0.700\ mol \times \frac{40.0\ g}{mol}$ = **28.0 g** |

| | |
|---|---|
| *Problem:* | What is the molarity of pure sulphuric acid, $H_2SO_4$, having a density of 1.839 g/mL? |
| *Solution:* | Since both density and molarity both have units of amount/volume, <br> $\text{density} = \frac{\text{amount (as mass)}}{\text{volume}}$ <br> and <br> $\text{molarity} = \frac{\text{amount (as moles)}}{\text{volume}}$ <br> therefore all we need to do is convert from mass to moles using molar mass. <br> $[H_2SO_4] = \frac{1.839\ g}{0.001\ L} \times \frac{1\ mol}{98.1\ g}$ = **18.7 M** |

| | |
|---|---|
| *Problem:* | If 200.0 mL of 0.500 M NaCl is added to 300.0 mL of water, what is the resulting [NaCl] in the mixture. |
| *Solution:* | $$M_2 = M_1 \times \frac{V_1}{V_2}$$ $$[NaCl] = 0.500\ M \times \frac{200.0\ mL}{(200.0 + 300.0)\ mL} = \mathbf{0.200\ M}$$ |
| *Problem:* | What volume of 6.00 M HCl is needed to make 2.00 L of 0.125 M HCl? |
| *Solution:* | Since $$M_1 \times V_1 = M_2 \times V_2$$ then, $$V_1 = \frac{M_2 \times V_2}{M_1}$$ $$V_{HCl} = \frac{0.125\ M \times 2.00\ L}{6.00\ M} = \mathbf{0.0417\ L}$$ |
| *Problem:* | If 300.0 mL of 0.250 M NaCl is added to 500.0 mL of 0.100 M NaCl, what is the resulting [NaCl] in the mixture? |
| *Solution:* | Calculate the moles of solute from each solution: $$M = \frac{moles}{volume}$$ $$moles = M \times V$$ $$moles_1 = 0.250\ M \times 0.300\ L = 0.075\ moles$$ $$moles_2 = 0.100\ M \times 0.500\ L = 0.050\ moles$$ $$Total\ moles = 0.075\ mol + 0.050\ mol = 0.125\ mol$$ $$M = \frac{0.125\ mol}{0.800\ L} = 0.156\ M$$ |

**4. Molality**

**It is defined as the number of moles of solute dissolved per kilogram of solvent. It is designated by the symbol 'm'.**

The label 2.0 m $H_2SO_4$ is read "2 molal sulphuric acid" and is prepared by adding 2.0 mol of $H_2SO_4$ to 1 kg of solvent. Molality is expressed as:

$$m = \frac{1000\, n_B}{W_A}$$

Where $n_B$ is the number of moles of the solute and $W_A$ is the mass in grams of solvent. The Molality of a solution does not change with temperature.

**Example 1: Find out the molarity of the solution which contains 32.0 g of methyl alcohol ($CH_3OH$) in 200 mL solution.**

**Solution :** Molar mass of $CH_3OH = 12 + 1 \times 3 + 16 + 1 = 32\ g\ mol^{-1}$

$$\text{Number of moles of } CH_3OH = \frac{32\ g}{32\ g\ mol^{-1}} = 1\ mol$$

Volume of the solution = 200 mL = 0.2 litre

$$\therefore \quad \text{Molarity} = \frac{No.\ of\ moles\ of\ solute}{Volume\ of\ solution\ in\ litres} = \frac{1}{0.2} = 5\ M$$

**Example 2: What is the Molality of a sulphuric acid solution of density 1.20 g/cm3 containing 50% sulphuric acid by mass?**

**Solution :** Mass of $1 cm^3$ of $H_2SO_4$ solution = 1.20 g

Mass of 1 litre ($1000\ cm^3$) of $H_2SO_4$ solution = $1.20 \times 1000 = 1200$ g

Mass of $H_2SO_4$ in 100 g solution of $H_2SO_4$ = 50 g

$$\text{Mass of } H_2SO_4 \text{ in 1200 g solution of } H_2SO_4 = \frac{50}{100} \times 1200 = 600\ g$$

$\therefore$ Mass of water in the solution = 1200 – 600 = 600 g

Molar mass of $H_2SO_4 = 98\ g\ mol^{-1}$

$$\text{No. of moles of } H_2SO_4 = \frac{Mass\ in\ grams}{Molar\ mass} = \frac{600\ g}{98\ g\ mol^{-1}}$$

$$\therefore \quad \text{Molarity} = \frac{No.\ of\ moles\ of\ H_2SO_4}{Mass\ of\ water\ in\ grams} \times 1000$$

$$= \frac{600}{98} \times \frac{1}{600} \times 1000 = 6.8\ m$$

## 5. Normality (N)

**It is defined as the number of gram equivalents (equivalent weight in grams) of a solute present per liter of the solution.**
**Unit of normality is gram equivalents litre$^{-1}$.**
**Normality changes with temperature since it involves volume.**
**When a solution is diluted 'x' times, its normality also decreases by 'x' times.**

Solutions in term of normality generally expressed as,

**N= Normal solution; 5N= Penta normal,**

**10N= Deca normal; N/2 = semi normal**

**N/10 = Deci normal; N/5 = Penti normal**

**N/100 or 0.01N= centinormal,**

**M/1000 or 0.001= millinormal**

**Mathematically normality can be calculated by following formulas,**

**(i)** $$\text{Normality} = \frac{number of\, gm.equiv.of\, solute}{volume of\, solution\ (l)}$$

**(ii)** $$N = \frac{weight\ of\ solute\ in\ gms.}{gm.equiv.wt.of\ solute \times volume\ of\ solution\ (l)}$$

**(iii)** $$N = \frac{wt.of\ solute\ per\ liter\ of\ solution}{gm.eq.wt.of\ solute}$$

**(iv)** $$N = \frac{wt.of\ solute}{gm.eq.wt.of\ solute} \times \frac{1000}{volume\ of\ solution\ in\ ml}$$

**(v)** $$N = \frac{percent\ of\ solute \times 10}{gm.eq.wt.of\ solute}$$

**(vi)** $$N = \frac{strength\ in\ gl^{-1} of\ solution}{gm.eq.wt.of\ solute}$$

**(vii)** $$N = \frac{wt\ \% \times density \times 10}{eq.wt.}$$

**(vii) If volume $V_1$ and normality $N_1$ is so changed that new normality and volume $N_2$ and $V_2$ then,**

$$N_1V_1 = N_2V_2 \text{ (Normality equation)}$$

**(ix) When two solutions of the same solute are mixed then normality of mixture (N) is**

$$N = \frac{N_1V_1 + N_2V_2}{V_1+V_2}$$

**(x) Vol. of water to be added i.e., $V_2–V_1$ to get a solution of normality $N_2$ from $V_1$ ml of normality $N_1$**

$$V_2 - V_1 = \left(\frac{N_1-N_2}{N_2}\right)V_1$$

**(xi) Normality of the acidic mixture** $= \frac{V_1N_1 + V_2N_2}{V_1+V_2}$

**Solved Example**

**Ex.1 The molarity of 20% (W/W) solution of sulphuric acid is 2.55 M. The density of the solution is:**
**(A) 1.25 g $cm^{-3}$ (B) 0.125 g $L^{-1}$ (C) 2.55 g $cm^{-3}$ (D) unpredictable**

**(Ans. A)**

**Sol.** Volume of 100 g of solution $= \frac{100}{d}$ ml

$$M = \frac{20 \times d \times 1000}{100 \times 98}$$

$$\text{Or } d = \frac{2.55 \times 100 \times 98}{20 \times 1000} = 1.249 » 1.25$$

**Ex.2 The density of a solution containing 13% by mass of sulphuric acid is 1.09 g/mL. Calculate the molarity and normality of the solution-**
**(A) 1.445 M (B) 14.45 M (C) 144.5 M (D) 0.1445 M**

**(Ans. A)**

**Sol.** Volume of 100 gram of the solution $= \frac{100}{d}$

$$= \frac{100}{1.09} \text{ mL} = \frac{100}{1.09 \times 1000} \text{ litre}$$

$$= \frac{1}{1.09 \times 10} \text{ litre}$$

Number of moles of $H_2SO_4$ in 100 gram of the solution $= \frac{13}{98}$

$$\text{Molarity} = \frac{\text{No. of moles of } H_2SO_4}{\text{Volume of solution in litre}}$$

$$= \frac{13}{98} \times \frac{1.09 \times 10}{1} = \mathbf{1.445\ M}$$

**Ex.3 Calculate the molarity of pure water (d = 1g/L)**
**(A) 555 M (B) 5.55 M (C) 55.5 M (D) None**
**(Ans. C)**

**Sol.** Consider 1000 mL of water
Mass of 1000 mL of water

$= 1000 \times 1 = 1000$ gram

Number of moles of water $= \frac{1000}{18} = 55.5$

Molarity $= \frac{\text{No. of moles of water}}{\text{Volume in litre}}$

$= \frac{55.5}{1}$ **= 55.5 M**

**Ex.4 Calculate the quantity of sodium carbonate (anhydrous) required to prepare 250 ml of**
**0.1 M solution-**
**(A) 2.65 gram (B) 4.95 gram (C) 6.25 gram (D) None**
**(Ans. A)**

**Sol.** We know that

Molarity $= \frac{W}{M \times V}$

Where;
W = Mass of $Na_2CO_3$ in gram
M = Molecular mass of $Na_2CO_3$ in grams =106
V =Volume of solution in litres $= \frac{250}{1000} = 0.25$

Molarity $= \frac{1}{10}$

Hence, $= \frac{1}{10} = \frac{W}{106 \times 0.25}$

Or W $= \frac{106 \times 0.25}{10}$ **= 2.65 gram**

**Ex.5 Find the Molality of $H_2SO_4$ solution whose specific gravity is 1.98 g $ml^{-1}$ and 95% by volume $H_2SO_4$**
**(A) 7.412 (B) 8.412 (C) 9.412 (D) 10.412**
**(Ans. C)**

**Sol.** $H_2SO_4$ is 95% by volume
wt. of $H_2SO_4$ = 95g
Vol of solution = 100ml
Moles of $H_2SO_4$ = , and weight of solution = 100 × 1.98 = 198 g
Weight of water = 198 _ 95 = 103 g

Molality $= \frac{95 \times 1000}{98 \times 103} = 9.412$, Hence molality of $H_2SO_4$ solution is **9.412**

**Ex.6.** Calculate **molality of 1 litre solution of 93% $H_2SO_4$ by volume. The density of solution is 1.84 gm ml$^{-1}$**
**(A) 9.42 (B) 10.42 (C) 11.42 (D) 12.42**

**(Ans. B)**
**Sol.** Given $H_2SO_4$ is 93% by volume
wt. of $H_2SO_4$ = 93g
Volume of solution = 100ml
weight of solution = 100 × 1.84 gm
= 184 gm
wt. of water = 184 _ 93 = 91 gm
Molality = $\frac{\text{Moles}}{\text{wt. of water in kg}}$
= $\frac{93 \times 1000}{98 \times 91}$ = 10.42

**Ex.7 Suppose 5gm of $CH_3COOH$ is dissolved in one litre of Ethanol. Assume no reaction between them. Calculate molality of resulting solution if density of Ethanol is 0.789 gm/ml.**
**(A) 0.0856 (B) 0.0956 (C) 0.1056 (D) 0.1156**

**(Ans. C)**

**Sol.** Wt .of $CH_3COOH$ dissolved = 5g, Eq. of $CH_3COOH$ dissolved = $\frac{5}{60}$
Volume of ethanol = 1 litre = 1000ml, Weight of ethanol = 1000 × 0.789 = 789g
Molality of solution = $\frac{\text{Moles of solute}}{\text{wt of solvent in kg}}$ = $\frac{\frac{5}{60 \times 789}}{1000}$ = 0.1056

**Ex.8 Calculate the molarity and normality of a solution containing 0.5 gm of NaOH dissolved in 500 ml. solution-**
**(A) 0.0025 M, 0.025 N (B) 0.025 M, 0.025 N**
**(C) 0.25 M, 0.25 N (D) 0.025 M, 0.0025 N**

**(Ans. B)**
**Sol.** Wt. of NaOH dissolved = 0.5 gm
Vol. of NaOH solution = 500 ml
**Calculation of molarity**
0.5 g of NaOH = $\frac{0.5}{40}$ moles of NaOH
[Q Mol. wt of NaOH = 40]
= 0.0125 moles
Thus 500 ml of the solution contain NaOH = 0.0125 moles

1000 ml of the solution contain = $\frac{0.0125}{500} \times 1000 = 0.025$ M
Hence molarity of the solution = **0.025 M**
**Calculation of normality**
Since NaOH is mono acidic ;
Eq. wt. of NaOH = Mol. wt. of NaOH = 40
0.5 gm of NaOH = $\frac{0.5}{40}$ gm equivalents = 0.0125 gm equivalents
Thus 500 ml of the solution contain NaOH = 0.0125 gmequ.
1000 ml of the solution contain
$= \frac{0.0125}{500} \times 1000 = 0.025$
Hence normality of the solution = **0.025 N**

**Ex.9 Calculate the molality and mole fraction of the solute in aqueous solution containing 3.0 gm of urea per 250 gm of water (Mol. wt. of urea = 60).**
**(A) 0.2 m, 0.00357 (B) 0.4 m, 0.00357 (C) 0.5 m, 0.00357 (D) 0.7m, 0.00357**

**(Ans. A)**

**Sol.** Wt. of solute (urea) dissolved = 3.0 gm
Wt. of the solvent (water) = 250 gm
Mol. wt. of the solute = 60
3.0 gm of the solute = $\frac{3.0}{60}$ moles = 0.05 moles
Thus 250 gm of the solvent contain = 0.05 moles of solute
1000 gm of the solvent contain = $\frac{0.05 \times 1000}{250}$ = 0.2 moles
Hence molality of the solution = 0.2 m

**In short,**
Molality = No. of moles of solute/1000 g of solvent
Molality = $\frac{3/60}{250} \times 1000 = \mathbf{0.2\ m}$
**Calculation of mole fraction**
3.0 gm of solute = 3/60 moles = 0.05 moles
250 gm of water = $\frac{250}{18}$ moles = 13.94 moles
Mole fraction of the solute
$= \frac{0.05}{0.05 + 13.94} = \frac{0.05}{13.99} = \mathbf{0.00357}$

**Ex.10 A solution has 25% of water, 25% ethanol and 50% acetic acid by mass. Calculate the mole fraction of each component.**
**(A) 0.50, 0.3, 0.19 (B) 0.19, 0.3, 0.50 (C) 0.3, 0.19, 0.50 (D) 0.50, 0.19, 0.3**

**(Ans. D)**

**Sol.**
Since 18 g of water = 1 mole

25 g of water = $\frac{25}{18}$ = 1.38 mole
Similarly, 46 g of ethanol = 1 mole

25 g of ethanol = $\frac{25}{46}$ = 0.55 moles
Again, 60 g of acetic acid = 1 mole

50 g of acetic acid = $\frac{50}{60}$ = 0.83 mole

Mole fraction of water = $\frac{1.38}{1.38+0.55+0.83}$ = **0.50**

Similarly, Mole fraction of ethanol = $\frac{0.55}{1.38+0.55+0.83}$ = **0.19**

Mole fraction of acetic acid = $\frac{0.83}{1.38+0.55+0.83}$ = **0.3**

**Ex.11. 15 gram of methyl alcohol is dissolved in 35 gram of water. What is the mass percentage of methyl alcohol in solution?**
**(A) 30% (B) 50% (C) 70% (D) 75%**

**(Ans. A)**

**Sol.** Total mass of solution = (15 35) gram = 50 gram
mass percentage of methyl alcohol

$= \frac{\text{Mass of methyl alcohol}}{\text{Mass of solution}} \times 100$

$= \frac{15}{50} \times 100 =$ **30%**

**Ex.12 Calculate the masses of cane sugar and water required to prepare 250 gram of 25% cane sugar solution-**
**(A) 187.5 gram, 62.5 gram (B) 62.5 gram, 187.5 gram**
**(C) 162.5 gram, 87.5 gram (D) None of these**
**(Ans. B)**
**Sol.** Mass percentage of sugar = 25
We know that

Mass percentage = $\frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$,    So, 25 = $\frac{\text{Mass of cane sugar}}{250} \times 100$

or Mass of cane sugar = $\frac{25 \times 250}{100}$ = **62.5 gram**
Mass of water = (250 - 62.5) = **187.5 gram**

**Ex.13 Calculate normality of the mixture obtained by mixing 100ml of 0.1N HCl and 50ml of 0.25N NaOH solution.**
**(A) 0.0467 N (B) 0.0367 N (C) 0.0267 N (D) 0.0167 N**
**(Ans. D)**
**Sol.** Meq. ofHCl =100 × 0.1 = 10
Meq.ofNaOH = 50 × 0.25 = 12.5
Q HCl and NaOH neutralize each other with equal eq.
Meq.ofNaOH left = 12.5 _ 10 = 2.5
Volume of new solution = 100 50 = 150 ml.
$N_{NaOH}$ left = $\frac{2.5}{150}$ = 0.0167 N
Hence normality of the mixture obtained is **0.0167 N**

**Ex.14. 300 ml 0.1 M HCl and 200 ml of 0.03M $H_2SO_4$ are mixed. Calculate the normality of the resulting mixture-**
**(A) 0.084 N (B) 0.84 N (C) 2.04 N (D) 2.84 N**
**(Ans.A)**
**Sol.** For HC*l*F or $H_2SO_4$
$V_1$ = 300 ml $V_2$ = 200 ml
$N_1$ = M x Basicity $N_2$ = M x Basicity
= 0.1 x 1 = 0.1 = 0.03 × 2 = 0.06
Normality of the mixture, N = $\frac{V_1N_1+V_2N_2}{V_1+V_2}$ = $\frac{300 \times 0.1+200 \times 0.06}{500}$ = $\frac{30+12}{500}$ = **0.084 N**

**Ex.15 Calculate the amount of each in the following solutions –**

B.. **150 ml of $\frac{N}{7}$ $H_2SO_4$ (ii) 250 ml of 0.2M $NaHCO_3$**
**(iii) 400 ml of $\frac{N}{10}$ $Na_2CO_3$ (iv) 1052 g of 1 m KOH.**
**(A) 52g, 2.12g, 4.2g, 1.05g (B) 1.05g, 4.2g, 2.12g, 52g**
**© 1.05g, 2.12g, 52g, 4.2g (D) 4.2g, 2.12g, 1.05g, 52g**

**(Ans. B)**

**Sol.** (i) Eq. wt. of $H_2SO_4$ = $\frac{\text{Mol. wt}}{\text{Basicity}}$ = $\frac{98}{2}$ = 49
Amount of $H_2SO_4$ per litre (strength) = Normality × Eq. wt. = × 49 = 7 g/litre
Amount in 150 ml = $\frac{7 \times 150}{1000}$ = 1.05 g

(ii) Molecular wt. of

$NaHCO_3$ = 23+ 1+ 12 +48 = 84

Amount of $NaHCO_3$ required to produce 1000 c.c. of one molar solution = 84 g

Amount present per litre in 0.2 M solution = 84 × 0.2 = 16.8 g

Amount present in 250 c.c. = $\frac{16.8 \times 250}{1000}$ = 4.2 g

Equivalent weight of $Na_2CO_3$ = $\frac{\text{Mol.wt}}{\text{No. of positive valencies}}$ = 53

Amount of $Na_2CO_3$ = Normality × Eq. wt. = × 53 = 5.3 g/litre

Amount present in 400 c.c. = $\frac{5.3 \times 400}{1000}$ = 2.12 g

We know that 1 molal solution of a substance contains 1000 g of solvent.

Wt. of KOH in 1052 g of 1 m KOH solution = 1052 _ 1000 = 52 g

**Ex.16 How many kilograms of wet NaOH containing 12% water are required to prepare 60 litres of 0.50 N solution? (A) 1.36 kg (B) 1.50 kg (C) 2.40 gm (D) 3.16 kg**

**(Ans. A)**

**Sol.** One litre of 0.50 N NaOH contains = 0.50 × 40g = 20 g = 0.020 kg

60 litres of 0.50 N NaOH contain

= 0.020 × 60 kg = 1.20 kg NaOH

Since the given NaOH contains 12% water, the amount of pure NaOH in 100 kg of the given NaOH = 100 - 12 = 88 kg

Thus 88 kg of pure NaOH is present in 100 kg wet NaOH

B.P. kg of pure NaOH is present in = $\frac{100}{88}$ × 1.20 = **1.36 kg wet NaOH**