2.1 DEFINITION OF SOLUTION, SOLVENT AND SOLUTE

When a small amount of sugar (solute) is mixed with water, sugar uniformly dissolves in water and a sugar solution is obtained. In this solution, sugar molecules are uniformly dispersed in molecules of water.

Similarly, a common salt (NaCl) when dissolve in water, if uniformly disperse in water and salt solutions is obtained so a solution of salt in water consist of ions of salt (Na\(^+\), Cl\(^-\)) dispersed in water.

Solutions are homogenous mixtures in which one substance is said to have been dissolved in the other.

The dissolved substance may be present as individual molecules or ions throughout the other substance. Since both the components of a solution are present in the molecular or ionic state, it constitutes a perfectly uniform and transparent system.

Components of Solution

In the study of solution, it is customary to designate the components in solution. The components are:

Solvent

The component present in larger proportion is known as solvent.

Solute

The component present in smaller proportion is known as solute.

Solution = Solvent + Solute

Ex. Sugar solution = Sugar (solute) + Water (solvent)
    Common salt solution = Salt (solute) + Water (solvent)

2.2 TYPES OF SOLUTION

Homogenous Solution

A solution in which two substances are mixed has uniform composition and the components cannot be identified separately.

Ex. Sugar solutions — Two substances

Heterogenous Solution

A solution in which two or more substances are mixed has non-uniform composition and the components can be identified separately.
Ex. Naphthalene solution — Two substances

Naphthalene (solute)  
Water (solvent)

In naphthalene solution, both water and naphthalene can be identified and separated from one another and non-uniform in their properties like density, concentration and viscosity etc.

**Aqueous Solution**

Solution containing water as solvent is called aqueous solution.

**Ammoniacal Solution**

Solution containing ammonia as solvent is called ammoniacal solution.

**Non-aqueous Solution**

Solution containing solvent other than water is called non-aqueous solution.

**Classification of Solutions**

Based on the physical state of components, the solutions are classified into gaseous solutions, liquid solutions and solid solutions.

1. **Gaseous solutions:** In these solutions, gas is solvent and the solute may be solid or liquid or gas.
   
   (a) Gas in Gas
   
   Ex. H₂ and O₂ mixture, air
   
   (b) Liquid in Gas
   
   Ex. Water in air
   
   (c) Solid in Gas
   
   Ex. Camphor in air

2. **Liquid solutions:** In these solutions, liquid is the solvent and the solute may be solid or liquid or gas.
   
   (a) Gas in Liquid
   
   Ex. Soda water (CO₂ in water)
   
   (b) Liquid in Liquid
   
   Ex. Alcohol in water
   
   (c) Solid in Liquid
   
   Ex. Salt in water

3. **Solid solutions:** In these solutions, solid is the solvent and the solute may be solid or liquid or gas.
   
   (a) Gas in Solid
   
   Ex. H₂ palladium
   
   (b) Liquid in Solid
   
   Ex. Hg in Zn
   
   (c) Solid in Solid
   
   Ex. Alloys (Zn in Cu)
Based on the relative amounts of the dissolved solute, the solutions can be classified into three types. They are:

(i) Saturated solutions
(ii) Unsaturated solutions
(iii) Supersaturated solutions.

**Saturated Solution**

A solution containing maximum amount of the dissolved solute at a given temperature is called saturated solution. This solution can’t dissolve any more of solute. A saturated solution contains little amount of undissolved solute. A dynamic equilibrium exists between undissolved solute and dissolved solute of saturated solutions.

\[
\text{Solute (undissolved)} \xrightarrow{\text{dynamic equilibrium}} \text{Solute (dissolved)}
\]

If a little solute is added to saturated solutions, it settles undissolved.

**Unsaturated Solution**

A solution in which the amount of dissolved solute is less than that required for saturation is called unsaturated solution. If a little amount of solute is added to the unsaturated solution, that solute is dissolved.

**Supersaturated Solution**

A solution containing more amount of solute than required for saturation at a given temperature is called supersaturated solution. A supersaturated solution is metastable. It readily forms saturated solutions on slight disturbance on addition of a small crystal of the solute.

**2.3. MOLE CONCEPT**

A mole is defined as “The amount of substance containing the same number of particles as the number of atoms present in 0.012 kg (or 12 gms) of carbon-12”.

\[\text{Or}\]

Mole is the quantity of substance that contains \(6.023 \times 10^{23}\) particles (Avogadro’s number)

For example, the gram molecular weight of \(\text{H}_2\) is 2 gms.
2 gm of H₂ is equal to one mole of H₂.

The gram molecular weight of Na₂CO₃ is 106 gms.

∴ 106 gms of Na₂CO₃ is equal to 1 mole of Na₂CO₃.

If we know the gram molecular weight of any substance, we can find the number of moles.

\[
\text{Number of moles} = \frac{\text{Weight of the substance}}{\text{Gram molecular weight}}
\]

**Example 1:** Find the number of present in 196 gms of H₂SO₄.

**Solution:**

The gram molecular weight of H₂SO₄ = 98

Weight of H₂SO₄ = 196

\[
\text{Number of moles of H₂SO₄} = \frac{196}{98} = 2
\]

196 gms of H₂SO₄ has 2 moles.

**Example 2:** What is the weight of 0.2 moles of AgNO₃.

**Solution:**

Number of moles of AgNO₃ = 0.2

Gram molecular weight of AgNO₃ = 108

\[
\text{Number of moles of AgNO₃} = \frac{\text{Weight of AgNO₃}}{\text{Gram molecular weight}}
\]

\[
0.2 = \frac{\text{Weight of AgNO₃}}{108}
\]

∴ Weight of AgNO₃ = 0.2 × 108

= 21.6 gms

2.4. **MOLARITY**

It is indicated by ‘M’.

‘It is the number of moles of solute present in one litre of solutions’

\[
\text{Molarity} = \frac{\text{Number of moles}}{1 \text{ litre of solution}}
\]

\[
\text{Molarity} = \frac{\text{Weight of the substance}}{\text{Gram molecular weight}}
\]

∴

\[
\text{Molarity} = \frac{\text{Weight of the substance}}{\text{Gram molecular weight}} \times \frac{1000}{\text{Volume in ml}}
\]

or

\[
M = \frac{W}{\text{gr. mol. wt.}} \times \frac{1000}{V \text{ in ml}}
\]

If the molarity is given we can find out the weight of the substance by using the following equations.
\[ W = M \times \text{gr. mol. wt.} \times \frac{V \text{ in ml}}{1000} \]

For dilution, molarity relation

\[ M_1 V_1 = M_2 V_2 \]

where
- \( M_1 \) = Molarity of concentrated solution
- \( V_1 \) = Volume of concentrated solution
- \( M_2 \) = Molarity of dilute solution
- \( V_2 \) = Volume of dilute solution

For volumetric analysis, molarity relation

\[ \frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2} \]

where
- \( M_1 \) = Molarity of Ist solution
- \( V_1 \) = Volume of Ist solution
- \( n_1 \) = Number of moles of Ist solution
- \( M_2 \) = Molarity of IInd solution
- \( V_2 \) = Volume of IInd solution
- \( n_2 \) = Number of moles of IInd solution

The unit of molarity is moles/lit. Molarity depends on temperature. If temperature increases, molarity decreases.

### 2.5 NORMALITY (N)

The number of gram equivalents of the solute present in 1 litre or 1000 ml of solution at a given temperature is called normality (N).

Normality (N) = \( \frac{\text{No. of gram equivalents of solute}}{\text{Volume of solution in litres}} \)

or

\[ N = \frac{x}{V \text{ in litres}} \]

where \( x \) = No. of gram equivalents of solute

It is calculated by the following relation

\[ x = \frac{\text{weight of solute in gram (W)}}{\text{gram equivalent weight of solute (GEW)}} \]

Then

\[ N = \frac{W}{\text{GEW}} \times \frac{1}{V \text{ in litres}} \]

or

\[ N = \frac{W}{\text{GEW}} \times \frac{1000}{V \text{ in litres}} \]

Normality relation for dilution, \( N_1 V_1 = N_2 V_2 \)

- \( N_1 \) = Normality of concentrated solution
- \( V_1 \) = Volume of concentrated solution
Solutions

\[ N_2 = \text{Normality of dilute solution} \]
\[ V_2 = \text{Volume of dilute solution} \]

Normality relation in volumetric analysis

\[ N_1V_1 = N_2V_2 \]

where

\[ N_1 = \text{Normality of Ist solution} \]
\[ V_1 = \text{Volume of Ist solution} \]
\[ N_2 = \text{Normality of IInd solution} \]
\[ V_2 = \text{Volume of IInd solution} \]

* The unit of normality is gram equivalents/litre
* Normality depends on temperature.

If 1 gram equivalent of solute dissolved in 1 litre or 1000 ml of water, the normality of solution is 1. This solution is known as 1 normal solution or 1 N solution. Similarly, if 0.1 gram equivalent weight of solute is dissolved in 1 litre or 1000 ml the resulting solution is known as decinormal or 0.1 N solution.

**Example 1:** If 49 grams of \( \text{H}_2\text{SO}_4 \) (GEW of \( \text{H}_2\text{SO}_4 = 49 \) g) dissolved in 1 litre or 1000 ml of water, the normality of solution is 1 N.

**Example 2:** If 5.3 grams of \( \text{Na}_2\text{CO}_3 \) (GEW of \( \text{Na}_2\text{CO}_3 = 53 \) g) dissolved in 1 litre or 1000 ml of water, the normality of solution is 0.1 N.

### 2.6 CALCULATION OF EQUIVALENT WEIGHT OF SUBSTANCES

The equivalent weight of a substance is the weight of it that reacts with 1g of hydrogen or 8 g of oxygen.

1. **Equivalent weight of an acid:** It is the ratio of molecular weight to its basicity.

   \[ E_{\text{Acid}} = \frac{\text{Molecular weight}}{\text{Basicity}} \]

   Basicity is the number of displaceable hydrogen atoms present in an acid.

   **Table 2.2: Equivalent weight of some Acids**

<table>
<thead>
<tr>
<th>Name of acid</th>
<th>Formula of acid</th>
<th>Basicity</th>
<th>Equivalent weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>1</td>
<td>36.5/1 = 36.5</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>HNO₃</td>
<td>1</td>
<td>63/1 = 63</td>
</tr>
<tr>
<td>Sulphuric acid</td>
<td>H₂SO₄</td>
<td>2</td>
<td>98/2 = 49</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH₃COOH</td>
<td>1</td>
<td>60/1 = 60</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>3</td>
<td>98/3 = 32.66</td>
</tr>
<tr>
<td>Oxalic acid</td>
<td>H₂C₂O₄</td>
<td>2</td>
<td>90/2 = 45</td>
</tr>
</tbody>
</table>

2. **Equivalent weight of a base:** It is the ratio of molecular weight to its acidity.

   \[ E_{\text{Base}} = \frac{\text{Molecular weight}}{\text{Acidity}} \]

   Acidity is the number of replaceable hydroxyl groups of the base is called acidity.
Table 2.3: Equivalent weight of some Bases

<table>
<thead>
<tr>
<th>Name of the base</th>
<th>Formula of base</th>
<th>Acidity</th>
<th>Equivalent weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium hydroxide</td>
<td>NaOH</td>
<td>1</td>
<td>40/1 = 40</td>
</tr>
<tr>
<td>Potassium hydroxide</td>
<td>KOH</td>
<td>1</td>
<td>56/1 = 56</td>
</tr>
<tr>
<td>Ammonium hydroxide</td>
<td>NH₄OH</td>
<td>1</td>
<td>35/1 = 35</td>
</tr>
<tr>
<td>Magnesium hydroxide</td>
<td>Mg(OH)₂</td>
<td>2</td>
<td>58/2 = 29</td>
</tr>
<tr>
<td>Calcium hydroxide</td>
<td>Ca(OH)₂</td>
<td>2</td>
<td>74/2 = 37</td>
</tr>
</tbody>
</table>

3. **Equivalent weight of a salt**: It is the ratio of molecular weight to the total valency of cations or anions.

\[
E_{\text{Salt}} = \frac{\text{Molecular weight}}{\text{Total valency of cations or anions}}
\]

Table 2.4: Equivalent weight of some Salts

<table>
<thead>
<tr>
<th>Name of the salt</th>
<th>Formula of salt</th>
<th>Total valency of cations or anions</th>
<th>Equivalent weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium chloride</td>
<td>NaCl</td>
<td>1</td>
<td>58.5/1 = 58.5</td>
</tr>
<tr>
<td>Sodium carbonate</td>
<td>Na₂CO₃</td>
<td>2</td>
<td>106/2 = 53</td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>MgCl₂</td>
<td>2</td>
<td>95/2 = 47.5</td>
</tr>
<tr>
<td>Magnesium sulphate</td>
<td>MgSO₄</td>
<td>2</td>
<td>120/2 = 60</td>
</tr>
<tr>
<td>Calcium carbonate</td>
<td>CaCO₃</td>
<td>2</td>
<td>100/2 = 50</td>
</tr>
<tr>
<td>Silver nitrate</td>
<td>AgNO₃</td>
<td>1</td>
<td>170/1 = 170</td>
</tr>
<tr>
<td>Copper sulphate</td>
<td>CuSO₄</td>
<td>2</td>
<td>159.5/2 = 79.75</td>
</tr>
</tbody>
</table>

4. **Equivalent weight of an oxidising agent**: It is the ratio of molecular weight to the number of electrons gained.

\[
E_{\text{OA}} = \frac{\text{Molecular weight}}{\text{No. of electrons gained}}
\]

Table 2.5: Equivalent weight of some oxidising agents

<table>
<thead>
<tr>
<th>Name of the compound</th>
<th>Formula of compound</th>
<th>No. of electrons gained</th>
<th>Equivalent weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium permanganate</td>
<td>KMnO₄</td>
<td>5 (in acidic medium)</td>
<td>158/5 = 31.6</td>
</tr>
<tr>
<td>Potassium permanganate</td>
<td>KMnO₄</td>
<td>3 (in neutral)</td>
<td>158/3 = 52.6</td>
</tr>
<tr>
<td>Potassium dichromate</td>
<td>K₂Cr₂O₇</td>
<td>6</td>
<td>294/6 = 49</td>
</tr>
</tbody>
</table>

5. **Equivalent weight of a reducing agent**: It is the ratio of molecular weight to the number of electrons lost.

\[
E_{\text{RA}} = \frac{\text{Molecular weight}}{\text{No. of electrons lost}}
\]
Table 2.6: Equivalent weight of some reducing agents

<table>
<thead>
<tr>
<th>Name of the compound</th>
<th>Formula of compound</th>
<th>No. of electrons lost</th>
<th>Equivalent weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mohr’s salt</td>
<td>FeSO₄(NH₄)₂SO₄ · 6H₂O</td>
<td>1</td>
<td>392/1 = 392</td>
</tr>
<tr>
<td>Hypo</td>
<td>Na₂S₂O₃</td>
<td>1</td>
<td>158/1 = 158</td>
</tr>
<tr>
<td>Oxalic acid</td>
<td>H₂C₂O₄ · 2H₂O</td>
<td>2</td>
<td>126/2 = 63</td>
</tr>
<tr>
<td>Ferrous sulphate</td>
<td>FeSO₄ · 7H₂O</td>
<td>1</td>
<td>278/1 = 278</td>
</tr>
</tbody>
</table>

6. Equivalent weight of an element: It is the ratio of atomic weight to the valency

\[ E_{\text{ele}} = \frac{\text{Atomic weight}}{\text{Valency}} \]

Table 2.7: Equivalent weight of some elements

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Valency</th>
<th>Equivalent weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>1</td>
<td>23/1 = 23</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg</td>
<td>2</td>
<td>24/2 = 12</td>
</tr>
<tr>
<td>Aluminium</td>
<td>Al</td>
<td>3</td>
<td>27/3 = 9</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag</td>
<td>1</td>
<td>108/1 = 108</td>
</tr>
<tr>
<td>Ferrous</td>
<td>Fe⁰²</td>
<td>2</td>
<td>56/2 = 28</td>
</tr>
<tr>
<td>Ferric</td>
<td>Fe⁰³</td>
<td>3</td>
<td>56/3 = 18.6</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn</td>
<td>2</td>
<td>65.4/2 = 32.7</td>
</tr>
<tr>
<td>Cupric</td>
<td>Cu⁰²</td>
<td>2</td>
<td>63.5/2 = 31.75</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
<td>1</td>
<td>39/1 = 39</td>
</tr>
</tbody>
</table>

2.7 NUMERICAL PROBLEMS ON MOLARITY, NORMALITY

Problem 1: 2 moles of a solute is dissolved in 5 litres of solution. What is its molarity?

Solution:

Number of moles of solute \((n) = 2\)

Volume of solution \((V) = 5\) litres

\[ M = ? \]

\[ M = \frac{n}{V} = \frac{2}{5} = 0.4 \text{ M} \]

Problem 2: Find the number of moles of solute present in 500 ml of 0.2 M solution.

Solution:

Number of moles of solute \((n) = ?\)

Volume of solution \((V) = 500\) ml = 0.5 lit.

Molarity of solution \((M) = 0.2\)

\[ M = \frac{n}{V} \]

\[ n = M \times V = 0.2 \times 0.5 = 0.1 \]
Problem 3: Find the molarity of solution containing 171 g sugar (sucrose) in 2 litres? (Molecular formula of sucrose = C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}, \ MW = 342)

Solution:

\[
\text{Wt. of sucrose (W) = 171 g} \\
\text{GMW of solute = 342 g} \\
\text{Volume of solution = 2 lit.} \\
M = ?
\]

\[
M = \frac{W}{\text{GMW}} \times \frac{1}{V \text{ in litres}} \\
= \frac{171}{342} \times \frac{1}{2} = \frac{1}{4} \text{ or } 0.25 \text{ M}
\]

Problem 4: 21.2 grams of Na\textsubscript{2}CO\textsubscript{3} are dissolved in 500 ml of solution. Find the molarity of solution.

Solution:

\[
W = 21.2 \text{ g} \\
\text{GMW} = 106 \text{ g} \\
V = 500 \text{ ml}
\]

\[
M = \frac{W}{\text{GMW}} \times \frac{1000}{V \text{ in ml}} \\
= \frac{21.2}{106} \times \frac{1000}{500} = 0.4 \text{ M}
\]

Problem 5: Find the weight of H\textsubscript{2}SO\textsubscript{4} required to prepare 400 ml of 0.5 M solution.

Solution:

\[
W = ? \\
\text{GMW} = 98 \text{ g} \\
V = 400 \text{ ml} \\
M = 0.5
\]

\[
M = \frac{W}{\text{GMW}} \times \frac{1000}{V} \Rightarrow 0.5 \\
W = \frac{M \times \text{GMW} \times V}{1000} \text{ g} \\
= \frac{0.5 \times 98 \times 400}{1000} = 19.6 \text{ g}
\]

Problem 6: What weight of urea (NH\textsubscript{2}CONH\textsubscript{2}) is required to prepare 2 lit. of 0.2 M solution?

Solution:

\[
W = ? \\
V = 2 \text{ lit.} \\
M = 0.5 \\
\text{GMW of urea = 60 g}
\]
M = \frac{W}{\text{GMW} \times \frac{1}{V \text{ in litres}}}

W = M \times \text{GMW} \times V

= 0.2 \times 60 \times 2

= 24 \text{ g}

**Problem 7:** Find the volume of water required to prepare 0.1 M H₂SO₄ from 200 ml of 0.5 M solution.

**Solution:**

M₁ = 0.5

V₁ = 200 ml

M₂ = 0.1

V₂ = ?

M₁V₁ = M₂V₂

0.5 \times 200 = 0.1 \times V₂

V₂ = 1000 ml

Volume of water required V₂ – V₁ = 1000 – 200

= 800 ml

**Problem 8:** 300 ml of water is added to 200 ml of 0.5 M HCl solution. Calculate the molarity of dilute solution.

**Solution:**

M₁ = 0.5

V₁ = 200 ml

V₂ = 300 + 200 = 500 ml

M₂ = ?

M₁V₁ = M₂V₂

M₂ = \frac{M₁V₁}{V₂} = \frac{0.5 \times 200}{500} = 0.2

**Problem 9:** 9.8 grams of H₂SO₄ dissolved in 2 litres of water calculate the normality of solution.

**Solution:**

Wt. of solute (W) = 9.8 g

Volume of solution (V) = 2 lit.

GMW of H₂SO₄ = 98 g

Basicity of H₂SO₄ = 2

GEW of H₂SO₄ = \frac{98}{2} = 49 g

Normality (N) = ?
Engineering Chemistry and Environmental Studies

\[ N = \frac{W}{\text{GEW}} \times \frac{1}{V \text{ in litres}} \]

\[ = \frac{9.8}{49} \times \frac{1}{2} \]

\[ = \frac{1}{10} \text{ or } 0.1 \text{ N} \]

**Problem 10:** Find the normality of solution prepared by dissolving 10 grams of NaOH in 500 ml of water.

**Solution:**

\[ W = 10 \text{ g} \]

\[ V = 500 \text{ ml} \]

GMW of NaOH = 40 g

Acidity of NaOH = 1

GEW of NaOH = \( \frac{40}{1} = 40 \text{ g} \)

\[ N = ? \]

\[ N = \frac{W}{\text{GEW}} \times \frac{1000}{V \text{ in ml}} \]

\[ = \frac{10}{40} \times \frac{1000}{500} = \frac{1}{2} \text{ or } 0.5 \text{ N} \]

**Problem 11:** Calculate the weight of Na₂CO₃ present in 100 ml of 0.5 N solution.

**Solution:**

\[ W = ? \]

\[ V = 100 \text{ ml} \]

\[ N = 0.5 \]

GMW of Na₂CO₃ = 106 g

GEW of Na₂CO₃ = \( \frac{106}{2} = 53 \text{ g} \)

\[ N = \frac{W}{\text{GEW}} \times \frac{1000}{V \text{ in ml}} \]

\[ W = \frac{N \times \text{GEW} \times V}{1000} = \frac{0.5 \times 53 \times 100}{1000} = 2.65 \text{ g} \]

**Problem 12:** Calculate the weight of oxalic acid (H₂C₂O₄ · 2H₂O) required to prepare 0.05 normal solution in 2 litres.

**Solution:**

\[ W = ? \]

\[ N = 0.05 \]

\[ V = 2 \text{ litres} \]
GMW of oxalic acid = 126 g

Basicity = 2

GEW of oxalic acid = 126/2 = 63 g

\[ N = \frac{W}{\text{GEW}} \times \frac{1}{V \text{ in litres}} \]

or

\[ W = N \times \text{GEW} \times V \]

= 0.05 \times 63 \times 2

= 6.3 g

**Problem 13:** Find the volume of water to be added to 250 ml of 0.05 N, Na₂CO₃ to get 0.01 N solution.

**Solution:** For dilution, the equation is 

\[ N_1 V_1 = N_2 V_2 \]

\[ N_1 = 0.05 \]

\[ V_1 = 250 \text{ ml} \]

\[ N_2 = 0.01 \]

\[ V_2 = ? \]

\[ N_1 V_1 = N_2 V_2 \]

or

\[ V_2 = \frac{N_1 V_1}{N_2} = \frac{0.05 \times 250}{0.01} = 1250 \text{ ml} \]

Water to be added = 1250 – 250 = 1000 ml

**Problem 14:** Calculate the normality of 20 ml of NaOH that exactly neutralises the 50 ml of 0.02 N, H₂SO₄.

**Solution:**

Normality of H₂SO₄ (N₁) = 0.02

Volume of H₂SO₄ (V₁) = 50 ml

Normality of NaOH (N₂) = ?

Volume of NaOH = 20 ml

\[ N_1 V_1 = N_2 V_2 \]

or

\[ N_2 = \frac{N_1 V_1}{V_2} = \frac{0.02 \times 50}{20} = 0.05 \]
Questions and Answers

Q. 1. Explain the terms solvent and solute with suitable examples.

Ans. The larger component of solution is called solvent and the smaller component is called solute.

Example: If 5 grams of a salt dissolved in 100 ml of water, salt solution is formed. In this solution, salt is solute and water is solvent.

Q. 2. What are the saturated, unsaturated and supersaturated solutions?

Ans. A solution containing maximum amount of the dissolved solute at a given temperature is called saturated solution.

A solution in which the amount of dissolved solute is less than that required for saturation is called unsaturated solution.

A solution containing more amount of dissolved solute than that required for saturation is called supersaturated solution.

Q. 3. What is molarity? Calculate molarity of a solution prepared by dissolving 5.85 g of NaCl in 500 ml of water.

Ans. The number of moles of solute present in 1 litre of solution at a given temperature is called molarity.

\[
M = \frac{W}{GMW} \times \frac{1000}{V \text{ in ml}}
\]

\[
= \frac{5.85}{58.5} \times \frac{1000}{500} = 0.2
\]

Q. 4. What is normality? Find the normality of a solution prepared by dissolving 1.58 g of KMnO₄ is 200 ml of water. (GMW of KMnO₄ is 158 g)

Ans. The number of gram equivalents of solute dissolved in 1 litre of solution is called normality.

\[
N = \frac{W}{GEW} \times \frac{1000}{V \text{ in ml}}
\]

\[
= \frac{1.58}{31.6} \times \frac{1000}{200} = \frac{1}{4} \text{ or } 0.25
\]

(KMnO₄ is an oxidising agent. It takes five electrons)
Q. 5. Calculate molarity and normality of a solution prepared by dissolving 10.6 g of \( \text{Na}_2\text{CO}_3 \) in 2 litres of water.

Ans.

\[
\begin{align*}
W &= 10.6 \text{ g} \\
\text{GMW} &= 106 \text{ g} \\
V &= 2 \text{ lit.} \\
\text{GEW} &= \frac{106}{2} = 53 \text{ g} \\
M &= \frac{W}{\text{GMW}} \times \frac{1}{V \text{ in litres}} \\
    &= \frac{10.6}{106} \times \frac{1}{2} \\
    &= \frac{1}{20} \text{ or } 0.05 \text{ M} \\
N &= \frac{W}{\text{GEW}} \times \frac{1}{V \text{ in litres}} \\
    &= \frac{10.6}{53} \times \frac{1}{2} \\
    &= \frac{1}{10} \text{ or } 0.1 \text{ N}
\end{align*}
\]