Unit 1

Solutions

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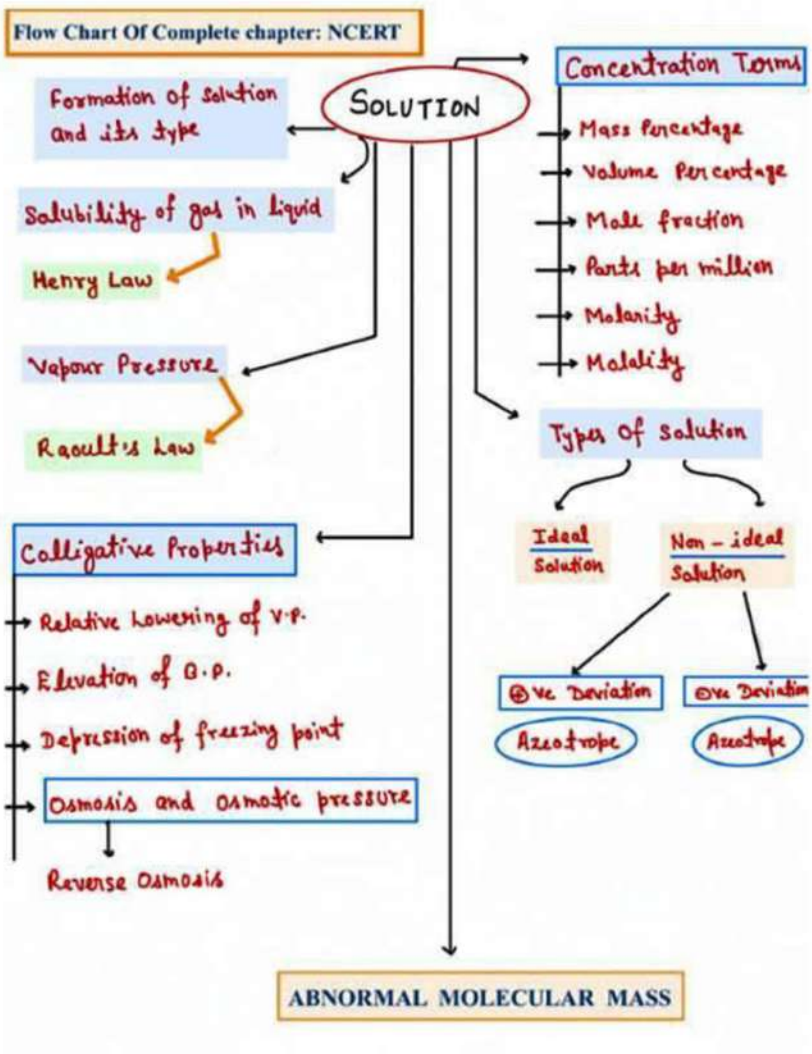
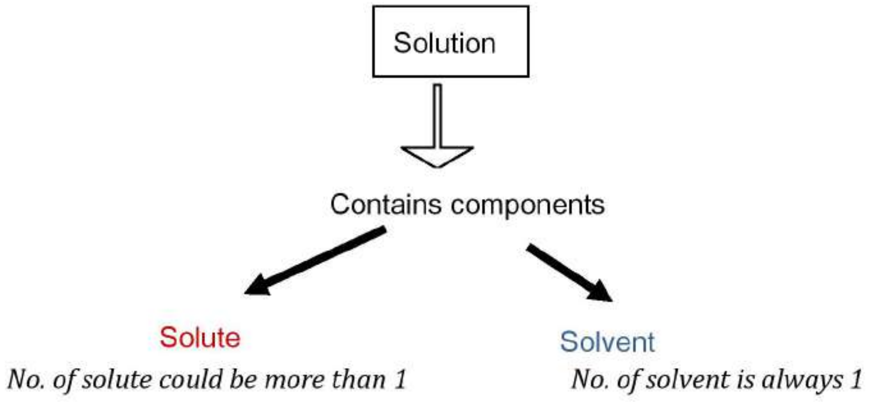


Prerequisites

Solution: A solution is a homogeneous mixture of two or more than two components.



Depending on the number of components, solutions can be termed as binary, ternary, quaternary etc.

No of solute	No of solven	Solution
1	1	Binary
2	1	Ternary
3	1	Quaternary

A few definitions▶

Dissolution: When a solid solute is added to the solvent, some solute dissolves and its concentration increases in solution. This process is known as dissolution.

Crystallization: Some solute particles in solution collide with the solid solute particles and get separated out of solution. This process is known as crystallization.

A stage is reached when the two processes occur at the same rate.

$$Solute + Solvent \xleftrightarrow{At\ dynamic\ equilibrium} \text{Solution}$$

Vapour Pressure: The pressure exhibited by vapor present above a liquid surface is known as vapor pressure.

Volatile substance: Volatile substances have the capability to go into the vapour phase even at low temperature. Eg. Acetone, ethanol, naphthalene.

2.SOLUTIONS.pdf - Google Drive

Concentration: It is the amount of solute in given amount of solution.

Expressing Concentration of Solutions

Composition of a solution can be described by expressing its concentration. There are several ways by which we can describe the concentration of the solution.

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

$$\text{Mass \% of a component} = \frac{\text{Mass of the component in the solution}}{\text{Total mass of the solution}} \times 100$$

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

$$\text{Volume \% of a component} = \frac{\text{Volume of the component}}{\text{Total volume of solution}} \times 100$$

(iii) ***Mass by volume percentage (w/V)***: 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:

$$\text{Parts per million} = \frac{\text{Number of parts of the component}}{\text{Total number of parts of all components of the solution}} \times 10^6$$

The concentration of pollutants in water or atmosphere is often expressed in terms of μg mL^{-1} or ppm

(v) ***Mole fraction***: It is defined as:

$$\text{Mole fraction of a component} = \frac{\text{Number of moles of the component}}{\text{Total number of moles of all the components}}$$

Mole fraction is denoted by $'x'$.
For example, in a binary mixture having components A and B.

No. of moles of A, $n_A = \frac{mass\ of\ A\ taken}{Molar\ mass\ of\ A}$

No. of moles of B, $n_B = \frac{mass\ of\ B\ taken}{Molar\ mass\ of\ B}$

Mole fraction of A, $x_A = \frac{n_A}{n_A+n_B}$

Mole fraction of B, $x_B = \frac{n_B}{n_A+n_B}$

In a given solution, sum of all the mole fractions is unity, i.e. $x_1 + x_2 + x_3 + \cdots\cdots\cdots\cdots + x_n = 1$

(vi) ***Molarity***: Molarity (M) is defined as number of moles of solute dissolved in one litre (or one cubic decimetre) of solution.

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

Or,

$$\text{Molarity} = \frac{\text{Mass of solute taken}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in mL}}$$

(vii) ***Normailty***: Normality (N) is defined as the number of gram-equivalents of solute dissolved in one litre (or one cubic decimetre) of solution.

$$\text{Normality} = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution in litre}}$$

Or,

$$\text{Normality} = \frac{\text{Mass of solute taken}}{\text{Equivalent mass of solute}} \times \frac{1000}{\text{Volume of solution in mL}}$$

Equivalent mass=$\frac{molar\ mass}{n}$

Where *n*= valency or
total positive or negavtive
charge of an ionic compound

(viii) ***Molality*** (m) is the number of moles of solute present in 1kg of solvent.

$$\text{Molality} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kilograms}}$$

Or,

$$\text{Molality} = \frac{\text{Mass of solute taken}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Mass of solvent in gram}}$$

(ix) ***Formality***: Ionic solutes do not exist in the form of molecules. Their molecular mass is expressed Gram formula mass. Formality is defined as Number of gram formula masses of solute/ volume of solution in litre.

- **Mass %, ppm, mole fraction** and molality are **independent of temperature**, whereas **molarity, normality, formality** depend on temperature. This is because volume depends on temperature and the mass does not.

Solubility

It is the maximum amount that can be dissolved in a specified amount of solvent at a specified temperature.

Various terms related to solution

Saturated solution: It is a solution in which **no more solute can be dissolved** at the same temperature and pressure. The solution which is in dynamic equilibrium with undissolved solute is the saturated solution and contains the maximum amount of solute dissolved in a given amount of solvent.

Unsaturated solution: An unsaturated solution is one in which **more solute can be dissolved** at the same temperature.

Supersaturated Solution: A supersaturated solution is **one in which more solute is dissolved** than is necessary to make a saturated solution.

Solubility can be three different

- Solubility of a Gas in a Liquid
- Solubility of a Liquid in a Liquid
- Solubility of a Solid in a Liquid

❖ Solubility of one substance into another depends on the followings

(i) *Nature of the substances*

Like dissolves like

Polar or ionic solute $\xrightarrow{\text{Dissolves in}}$ **Polar solvents**

Non-polar or covalent solute $\xrightarrow{\text{Dissolves in}}$ **Non-Polar solvents**

(ii) *Effect of Temperature*

The solubility of a **solid in a liquid** or **liquid in liquid** is significantly affected by temperature changes. In general, if in a *nearly saturated solution*, the dissolution process is endothermic ($\Delta_{sol} H > 0$), the solubility should increase with rise in temperature and if it is exothermic ($\Delta_{sol} H < 0$) the solubility should decrease.

Solubility of **gases in liquids** decreases with rise in temperature. When dissolved, the gas molecules are present in liquid phase and the process of dissolution can be considered similar to condensation and heat is evolved in this process.

(iii) *Effect of Pressure*

Pressure does not have any significant effect on solubility of **solids in liquids** and **liquid in liquid**. It is so because solids and liquids are highly incompressible and practically remain unaffected by changes in pressure.

The solubility of gases in liquid increases with increase of pressure.

I. Solubility of a Solid in a Liquid

A quantitative relation between pressure and solubility of a gas in a solvent is given by Henry's law.

Henry's law states that at a constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas present above the surface of liquid or solution.

The most commonly used form of Henry's law states that "**the partial pressure of the gas in vapour phase (p) is proportional to the mole fraction of the gas (x) in the solution**"

$$p \ \alpha \ x$$

$$\boxed{p = K_H . x}$$

Henry's law

Where K_H – Henry's law constant

Value of K_H depends on the nature of the gas.

II. Solubility of a Liquid in a Liquid

Let us consider a binary solution of two volatile liquids and denote the two components as 1 and 2. When taken in a closed vessel, both the components would evaporate and eventually an equilibrium would be established between vapour phase and the liquid phase.

Let, Partial vapour pressure of component 1 = p_1
Partial vapour pressure of component 2 = p_2
Total vapour pressure = p_{total}
Mole fraction of component 1 = x_1
Mole fraction of component 2 = x_2

Then according to **Raoult's law** which states that **for a solution of volatile liquids, the partial vapour pressure of each component of the solution is directly proportional to its mole fraction present in solution**.

Thus, for component 1

$$p_1 \ \alpha \ x_1$$

And $p_1 = p_1^0 x_1$ where p_1^0 is the vapour pressure of pure component 1 at the same temperature.

Similarly, for component 2

$$p_2 \ \alpha \ x_2$$

And $p_2 = p_2^0 x_2$ where p_2^0 is the vapour pressure of pure component 1 at the same temperature.

According to Dalton's law of partial pressures, the total pressure over the solution phase in the container will be the sum of the partial pressures of the components of the solution and is given as:

$$p_{total} = p_1 + p_2 \ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots \text{(1)}$$

Substituting the values of p_1 and p_2, we get

$$p_{total} = p_1^0 x_1 + p_2^0 x_2 \ldots\ldots\ldots\ldots\text{(2)}$$

And we know

$$x_1 + x_2 = 1$$
$$x_1 = 1 - x_2 \ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots \text{(3)}$$

Substituting equation 3 in equation 2, we have

$$\begin{aligned} p_{total} &= p_1^0(1 - x_2) + p_2^0 x_2 \\ &= p_1^0 - p_1^0 x_2 + p_2^0 x_2 \\ &= p_1^0 + p_2^0 x_2 - p_1^0 x_2 \end{aligned}$$

$$\boxed{p_{total} = p_1^0 + (p_2^0 - p_1^0)x_2} \ldots\ldots\ldots\ldots\ldots \text{(4)}$$

Following conclusions can be drawn from equation (4).

- Total vapour pressure over the solution can be related to the mole fraction of any one component.
- Total vapour pressure over the solution varies linearly with the mole fraction of component 2.
- Depending on the vapour pressures of the pure components 1 and 2, total vapour pressure over the solution decreases or increases with the increase of the mole fraction of component 1.

A plot of p_1 or p_2 versus the mole fractions x_1 and x_2 a solution gives a linear plot as shown