SOLUTIONS

⇒ Homogeneous mix. of two/ more components.

$$\text{Solutions} \rightarrow \begin{cases} \text{Solute (smaller quantity)} \\ + \\ \text{Solvent (larger quantity)} \end{cases}$$

⇒ Types of solutions :-

		Eg :-
1. Gaseous solutions (Solvent = gas)	Solid in gas	Camphor in N_2
	Liq. in gas	Humidity in air.
	Gas in gas	$O_2 + N_2$
2. Liquid Solutions (Solvent = liquid)	Solid in liq.	Sugar solution
	Liquid in liq.	Ethanol in water
	Gas in liq.	Aerated drinks
3. Solid Solutions (Solvent = Solid)	Solid in solid	Alloys
	Liq. in solid	Amalgam
	Gas in solid	H_2 in Pd

⇒ Concentration terms :-

1. $$\%\ \text{Mass (W/W)} = \frac{\text{Mass of the component}}{\text{Total mass of solution}} \times 100$$

2. $$\%\ \text{Volume (V/V)} = \frac{\text{Volume of the component}}{\text{Total volume of solution}} \times 100$$

3. $$\text{Normality} = \frac{\text{No. of gm equivalent of solute}}{\text{Vol. of solution (in L)}}$$

4. % Mass / volume (w/v) = $\frac{\text{Mass of the component}}{\text{Total volume of solution}} \times 100$

5. ppm = $\frac{\text{No. of parts of the component}}{\text{Total no. of parts of all components of solution}} \times 10^6$.

6. Mole fraction = $\frac{\text{No. of moles of the component}}{\text{Total no. of moles of all components}}$

7. Molarity = $\frac{\text{No. of moles of solute}}{\text{Vol. of solution (in L)}}$

8. Molality = $\frac{\text{No. of moles of solute}}{\text{Mass of solvent (in kg)}}$

*. Molarity → Depends on Temp. (∵ Vol^m depends on Temp)

* (ppm, Molality, Mass %, Mole fraction → Independent of Temp.

∵ Mass is independent of Temp.

9. Strength (in gL^{-1} or $g\,dm^{-3}$) = $\frac{\text{Mass of solute (g)}}{\text{Vol. of solution } (dm^3)}$

Relationship between Volume & Molarity or Normality:

$$N_1V_1 = N_2V_2$$

$$M_1V_1 = M_2V_2$$

* Relation between Molarity and (W/W)%

Molarity = $\frac{(W/W)\% \times 10 \times \text{density}}{\text{Molecular Mass}}$ (d in g/ml or kg/L)

* Relation between Molarity (M) & Molality (m)

$$m = \frac{1000 \times M}{1000 \times d - M \times M_w}$$

(d in g/ml or kg/L)

M_w = M.M of solute.

*. **Solubility** → It is the max. amount that can be dissolved in a specified amount of solvent at a specific temperature.

* Solubility of solid in liquid :-

→ Nature of solute & solvent → "Like dissolves like".

→ Temp. ⇒ $\Delta_{sol} H > 0$; solubility $\propto$ Temp.

⇒ $\Delta_{sol} H < 0$; solubility $\propto \frac{1}{\text{Temp.}}$.

→ Pressure → Unaffected (∵ Solids & liquids are highly incompressible)

* Solubility of gas in liquid :-

⇒ Nature of gas & the solvent ⇒ Chemical similarity enhances solubility of gas.

⇒ Temperature $\propto \frac{1}{\text{Solubility}}$ (∵ Dissolution of gas is exothermic)

⇒ Pressure $\propto$ Solubility.

#. Henry's law

1st form ⇒ $\boxed{p = K_H \cdot x}$

p = partial pressure of gas

x = mole fraction of gas

K_H = Henry's const.

Unit of K_H = (atm or bar)

2nd form

$$\boxed{C = Kp}$$

C = Solubility of gas.
K = Const of proportionality
p = partial pressure.

Units : $C \Rightarrow mol\ L^{-1}$
$K = mol\ L^{-1}\ atm^{-1}$ or $mol\ L^{-1}\ bar^{-1}$.
p = atm or bar

#. solubility $\propto \frac{1}{K_H}$

#. K_H depends on nature of the gas.

#. $K_H \propto$ Temp.

*. Limitations :-

→ Applicable to ideal gas (High pressure & low temp.)
→ No association/ dissociation of gas in solvent.

*. Applications :-

⇒ Carbonation of soft drinks / soda (Sealed under high pressur
⇒ Scuba diving (Dilution with He & O_2 to prevent bends)
⇒ Altitude sickness (Anoxia) ⇒ (Low conc. of O_2 in high altitudes)

*. $$\ln \frac{C_2}{C_1} = \frac{\Delta H}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

(C_1, C_2 = Conc of sol.)
(T_1, T_2 = Temp)
(ΔH = heat of solution)

*. Vapour pressure of liquid solutions

→ Solvent (liquid)

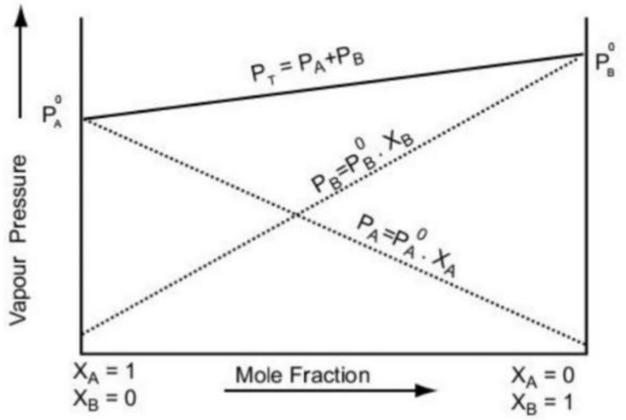
→ Solute (solid or liq.)

⇒ For a solution of volatile liquids (Raoult's law)

$P_s = p_1^\circ x_1 + p_2^\circ x_2$ (P_s = Total pressure)

$= p_1^\circ (1 - x_2) + p_2^\circ x_2$ (p° = v.p in pure state)

$= p_1^\circ + x_2 (p_2^\circ - p_1^\circ)$ (x = mole fraction)



For a sol^n of volatile liquids, vapour pressure of each component ∝ mole fraction of components in liq. phase.

Composition of vapour phase in equilibrium with solution.

$$\boxed{p_i = y_i P_T}$$

(y_i = Mole fraction in vapour phase)

(p_i = partial pressure of components)

x = Mole fraction (in liq phase)
y = Mole fraction (in vap. phase)

⇒ $p_i = x_i p^\circ$; $p_i = y_i P_T$

*. $\boxed{x_i p^\circ = y_i P_T}$

* Raoult's law as a special case of Henry's law :-

$p_i = x_i p^\circ$; $p_i = K_H x_i$

when, $p^\circ = K_H$.

*. Raoult's law for non-volatile solutes

→ Escaping tendency of solvent molecules(↓) ∴ V.P ↓
(∵ solute particles occupy the surface of solvent)

⇒ Hence, V.P of solution = V.P of solvent in sol^n.

$$P = x_1 P^\circ$$

(x_1 = Mole fraction of solvent)

$$\frac{P}{P^\circ} = x_1$$

$$1 - \frac{P}{P_0} = 1 - x_1$$

$$\underbrace{\frac{P^\circ - P}{P^\circ}}_{\text{RLVP}} = \underbrace{x_2}_{\text{Mole fraction of solute}}$$

#. Relative lowering of V.P of solution having a non-volatile solute = Mole fraction of solute.

Liq - liq. solutions

Ideal sol^n

* (Obeys Raoult's law)

→ $\Delta V_{mix} = 0$

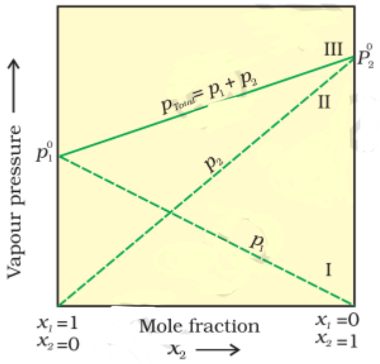
→ $\Delta H_{mix} = 0$

→ Interactions remains same after mixing.
No azeotrope formation

→ $p_A = p^0_A x_A$

→ $p_B = p^0_B x_B$

→ $p_s = p^0_A x_A + p^0_B x_B$

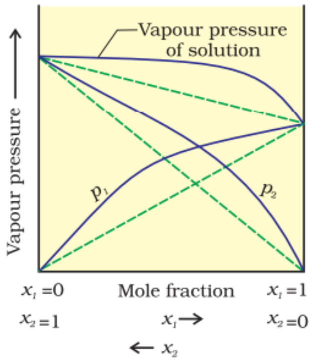


Ideal solution

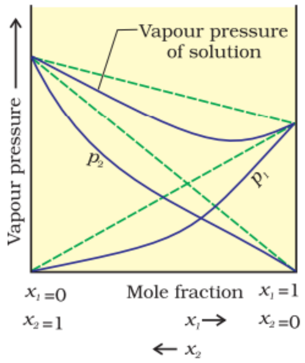
Non-ideal sol^n.

* (Doesn't obey Raoult's law)

+ve deviation	−ve deviation
→ $\Delta V_{mix} = +ve$	→ $\Delta V_{mix} = -ve$
→ $\Delta H_{mix} = +ve$	→ $\Delta H_{mix} = -ve$
→ Interactions are lower after mixing	→ Interactions are stronger after mixing
→ Forms minimum boiling azeotropes	→ Forms maximum boiling azeotropes
→ $p_A > x_A p^0_A$	→ $p_A < p^0_A x_A$
→ $p_B > x_B p^0_B$	→ $p_B < p^0_B x_B$
→ $p_s > x_A p^0_A + x_B p^0_B$	→ $p_s < p^0_A x_A + p^0_B x_B$



+ve deviation



−ve deviation

Azeotropes

- → Mix. of two/more liq. that boils at const temp.
- → Composition of liq phase = Composition of vapour phase.
- → Remains unseparated by simple distillation
- → Min. boiling azeotropes → Boils at lower temp than any of the components (+ve deviation)
- → Max boiling azeotropes → Boils at higher temp than any of the components (-ve deviation).

Colligative properties → (Depends on no. of solute particles irrespective of nature of solute).

1. Relative lowering of vapour pressure.

$$\frac{P^{o} - P_s}{P^{o}} = \chi_{solute} = \frac{i n_2}{i n_2 + n_1}$$

P^{o} = v.p of pure solvent

P_s = v.p of solution.

n_2 = no. of moles of solute

n_1 = no. of moles of solvent

i = Van't Hoff factor

2. Elevation in B.P

* $\Delta T_b = i \cdot K_b \cdot m$

* $\Delta T_b = T_b - T_b^{o}$

* $$K_b = \frac{R \times M_1 \times T_b^2}{1000 \times \Delta_{vap} H}$$

T_b = B.P of Solution (in K)

T_b^{o} = B.P of solvent (in K)

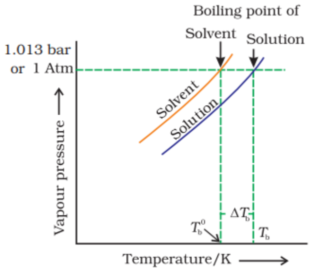
m = molal conc.

K_b = Ebullioscopic const.

R = gas const.

M = M.M of solvent

$\Delta_{vap} H$ = Enthalpy of vapourisation

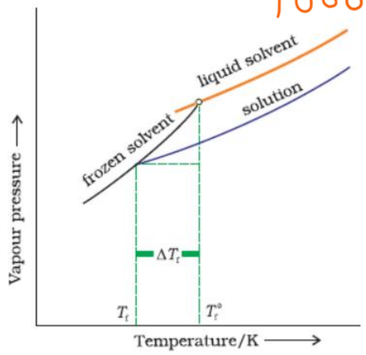


3. Depression in freezing point.

$\Delta T_f = iK_f \cdot m$

$\Delta T_f = T_f^{\circ} - T_f$

$$K_f = \frac{R \times M_1 \times T_f^2}{1000 \times \Delta_{fus} H}$$



i = Vant Hoff factor

K_f = Cyroscopic const.

m = molal conc.

R = gas const

T_f° = F.P of solvent (in K)

T_f = F.P of Solution (K)

M_1 = M.M of solvent

$\Delta_{fus} H$ = Enthalpy of fusion.

4. Osmotic pressure (π)

$\pi = i \cdot CRT$

$\pi = i \cdot \frac{n}{V} RT$

C = Molar concentration

T = temperature

n = no. of moles of solute

V = vol. of solution

i = Van't Hoff factor

* Van't Hoff factor (i)

$$i = \frac{\text{No. of moles after association/dissociation}}{\text{No. of moles before association/dissociation}}$$

$$= \frac{\text{Observed value of the colligative property}}{\text{Normal value of the same property}}$$

$$= \frac{\text{Normal M.M}}{\text{Abnormal M.M}}$$

* $i = \frac{1+(n-1)\alpha}{1}$ (For dissociation)

$i = \frac{1-\alpha+\frac{\alpha}{n}}{1}$ (For association)

(α = degree of dissociation or association)