⇒ Homogeneous mix. of two / more components.

⇒ Solutions → Solute (smaller quantity)
+
→ Solvent (larger quantity)

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

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

3\. Normality = No. of gm equivalent of solute

$\times 100$

Total volume of solution

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

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

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

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

*. Molarity → Depends on Temp. ($\because$ Vol^m depends on Temp)

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

$\because$ Mass is independent of Temp.

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

Relationship between Volume & Molarity or Normality:

$$N_1V_1 = N_2V_2$$

$$M_1V_1 = M_2V_2$$

* Relation between Molarity and (w/w)%

$\text{Molarity} = \dfrac{(w/w)\% \times 10 \times \text{density}}{}$ (d in g/ml)

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

(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 ∝ 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 ∝ Solubility.

Henry's law

1st form ⇒ $\boxed{p = K_H \cdot x}$

p = partial pressure of gas

x = mole fraction of gas

$\Rightarrow$ $\boxed{C = Kp}$

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 pressure)

⇒ 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)

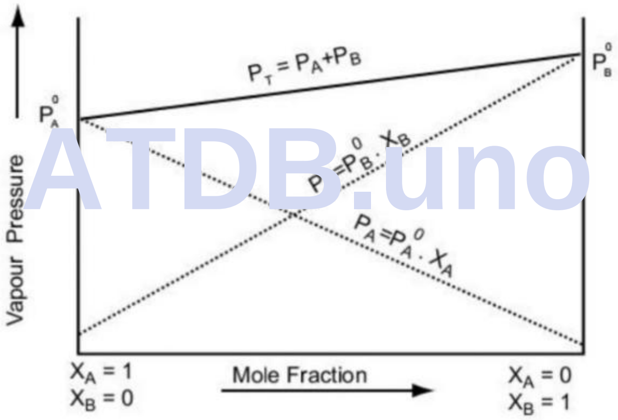
→ Solute (solid or liq.)

→ 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 solⁿ 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

x = Mole fraction (in liq phase)

y = Mole fraction (in vap. phase)

$\Rightarrow \quad p_i = x_i p^\circ \quad ; \quad p_i = y_i P_T$

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

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

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

when, $p^\circ = K_H$.

*. **Raoult's law for non-volatile solutes**

$\Rightarrow$ Escaping tendency of solvent molecules (↓) ∴ V.P ↓

(∵ solute particles occupy the surface of solvent)

$\Rightarrow$ 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 solⁿ

* (Obeys Raoult's law)

→ $\Delta V_{mix} = 0$

→ $\Delta H_{mix} = 0$

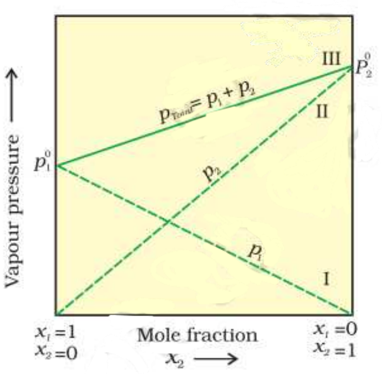
→ Interactions remains same after mixing.

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 solⁿ.

* (Doesn't obey Raoult's law)

+ve deviation

→ $\Delta V_{mix} = +ve$

→ $\Delta H_{mix} = +ve$

→ Interactions are lower after mixing

→ Forms minimum boiling azeotropes

→ $P_A > X_A P^0_A$

→ $P_B > X_B P^0_B$

→ $P_S > X_A P^0_A + X_B P^0_B$

-ve deviation

→ $\Delta V_{mix} = -ve$

→ $\Delta H_{mix} = -ve$

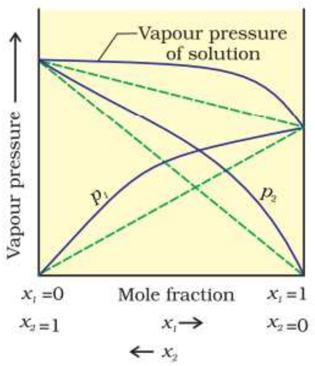
→ Interactions are stronger after mixing

→ Forms maximum boiling azeotropes

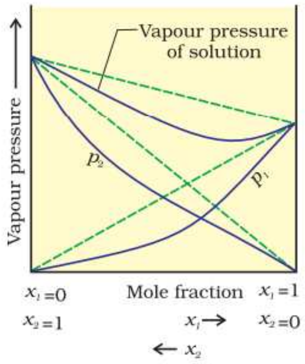
→ $P_A < P^0_A X_A$

→ $P_B < P^0_B X_B$

→ $P_S < P^0_A X_A + P^0_B X_B$



+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.

$$\frac{P^\circ - P_s}{P^\circ} = x_{solute} = \frac{in_2}{in_2+n_1}$$

P° = 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

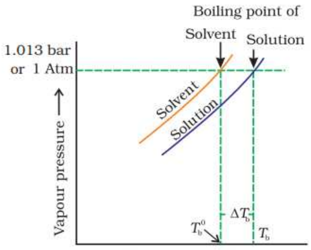
2. Elevation in B.P

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

* $\Delta T_b = T_b - T_b^\circ$

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

T_b = B.P of Solution (in K)
T_b° = 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 vapouri-



3. Depression in freezing point :

$\Delta T_f = i K_f \cdot m$

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

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

i = Vant Hoff factor

K_f = Cyroscopic const.

m = molal conc.

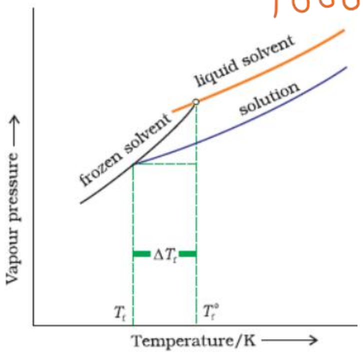
R = gas const

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)

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

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

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

* $i = \frac{1 + (n-1)\alpha}{1}$ $\quad i = \frac{1 - \alpha + \frac{\alpha}{n}}{1}$ (α = degree of dissociation