⇒ Homogeneous mix. of two / more components.

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

⇒ Types of solutions :-

1. Gaseous solutions (Solvent = gas)

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

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 = (W/W)% × 10 × 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 ∝ $\frac{1}{\text{Temp.}}$

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

* Solubility of gas in liquid :—

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

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

⇒ Pressure ∝ Solubility.

Henry's law

1st form ⇒ $p = K_H \cdot x$

p = partial pressure of gas

x = Mole fraction of gas

$C = K_H p$

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

Units : C ⇒ 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)

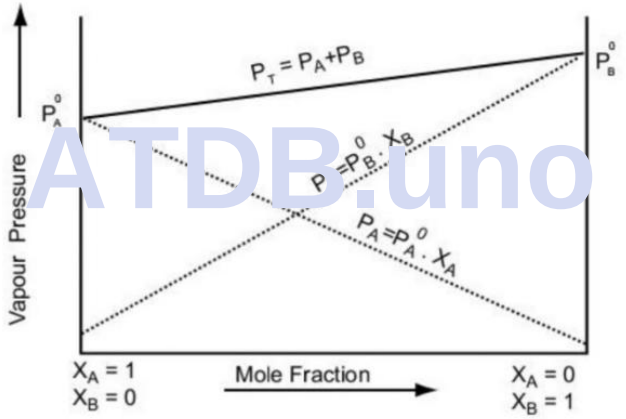
⇒ Solute (solid or liq.)

⇒ 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 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.

$$P_i = y_i P_T$$

(y_i = Mole fraction in vapour phase

(P_i = partial pressure

x = Mole fraction (in liq. phase)

y = Mole fraction (in vap. phase)

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

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

* 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

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

($\because$ 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)

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

$1 - \frac{p}{p_0} = 1 - x_1$

$\frac{p^\circ - p}{p^\circ} = x_2$

RLVP = 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)

→ $\Delta V_{mix} = 0$

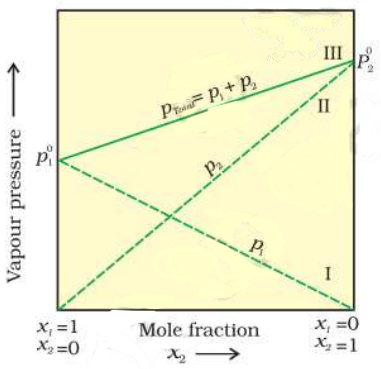
→ $\Delta H_{mix} = 0$

→ Interactions remains same after mixing. No azeotrope formation

→ $P_A = P^\circ_A \chi_A$

→ $P_B = P^\circ_B \chi_B$

→ $P_S = P^\circ_A \chi_A + P^\circ_B \chi_B$



Ideal solution

Non-ideal sol^n.

* (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

→ $P_A > \chi_A P^\circ_A$

→ $P_B > \chi_B P^\circ_B$

→ $P_S > \chi_A P^\circ_A + \chi_B P^\circ_B$

-ve deviation

→ $\Delta V_{mix} = -ve$

→ $\Delta H_{mix} = -ve$

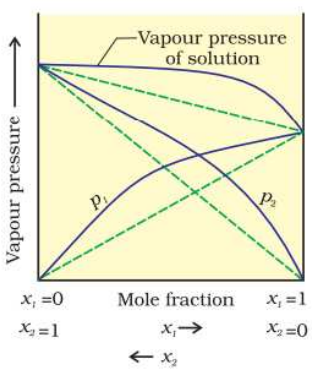
→ Interactions are stronger after mixing

→ Forms maximum boiling azeotropes

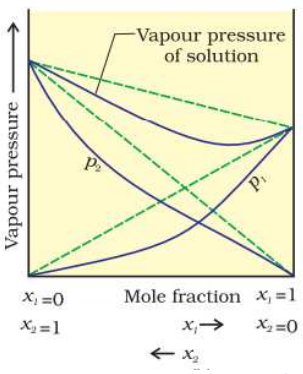
→ $P_A < P^\circ_A \chi_A$

→ $P_B < P^\circ_B \chi_B$

→ $P_S < P^\circ_A \chi_A + P^\circ_B \chi_B$



+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.

$$\frac{P^\circ - P_s}{P^\circ} = \chi_{solute} = \frac{n_2}{in_2 + n_1}$$

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)

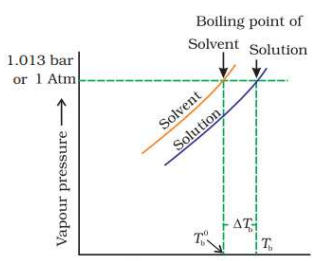
m = molal conc.

K_b = Ebullioscopic const.

R = gas const.

M = M.M of solvent

$\Delta_{vap}H$ = Enthalpy of vapouri-

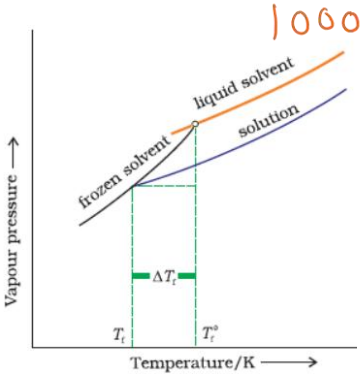


3. Depression in freezing point

$$\Delta T_f = i K_f \cdot m$$

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

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



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)

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

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

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

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