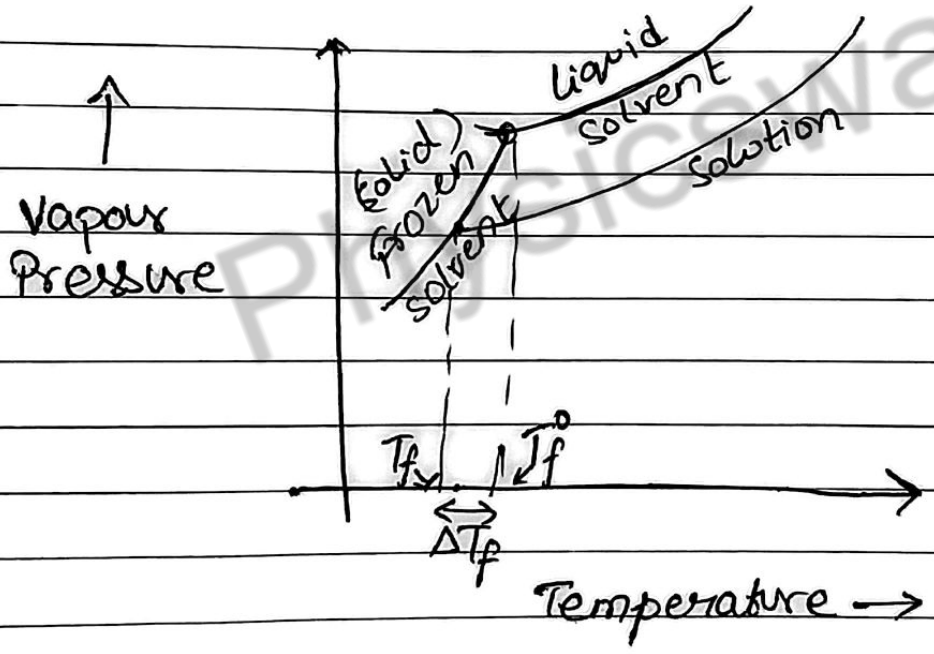
Depression in Freezing Point.

Freezing point is the temperature at which the vapour Pressure of a substance int it's liquid phase becomes equal to the vapour poressure in the solid phase

⇒ we also know, On addition of non volatile Solute the vapour Pressure of a Solution decreases.



from the graph we can see,

i) vapour Pressure of non-volatile Solute conteining Solution is lower than Vapour Pressure of pure liquid Solvent.

ii) Freezing Point of pure Solvent is T_f°
(vapour Pressure of Solvent (Liquid) = V.P. of Frozen Solvent

∞ ∞ ΔT_f is a Colligative property as it depends upon no of solute particles and not on nature of solute particles

iii) Freezing Point of Solution is T_f

(Vapour Pressure of Solution = Vapour Pressure of solid Solvent)

$T_f < T_f^\circ$

Depression in Freezing point

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

experimentally, we find

$\Delta T_f \propto m$ (molality of solution) (colligative ∞ ∞ ∞ property)

$$\Delta T_f = K_f m$$

K_f → Molal Depression Constant / Cryoscopic Constant

K_f → depends on solvent

K_f → unit K kg/mol

ex: K_f for water is 1.86 K kg/mol

$$m \text{ (molality)} = \frac{\text{no of moles of non volatile solute.}}{\text{Mass of Solvent (in kg)}}$$

Q i) How much C_2H_5OH should be added to 1L H_2O so that it will not freeze at ~~-20°C~~ -2°C? ($K_f = 1.86$) (density of water = 10^3 kg/m^3)

Solution: a) 20g b) 50g c) 80g d) 100g

③ 1L water = 1kg water

C_2H_5OH
Molar Mass = 46

$\Delta T_f = K_f \times m$

$.2 = 1.86 \times \frac{\text{no of moles of solute}}{\text{Mass of Solvent (in kg)}}$

$2 = 1.86 \times \frac{\frac{x}{46}}{1kg}$

$x = 49.4g \simeq 50g$

Q2) The freezing point of a solution containing 50cm³ of ethylene glycol in 50g of water is found to be -34°C. Assuming ideal behaviour, calculate the density of ethylene glycol. (K_f for water = 1.86 K kg/mol)

a) 1.02 g/cm³
b) 2.12 g/cm³
c) 1.13 g/cm³

Solution:

$\Delta T_f = K_f \times m$

$34 = 1.86 \times \frac{\text{no of moles of solute}}{\text{Mass of Solvent (in kg)}}$

Let mass of ethylene glycol be x

$34 = 1.86 \times \frac{\frac{x}{62}}{\frac{50}{1000}}$

CH_2-CH_2
$|\quad\quad|$
$OH\quad OH$

Molar Mass = 62

$x = 56.66\,g$

density of ethylene glycol $= \frac{\text{mass}}{\text{volume}} = \frac{56.66g}{50cm^3} = 1.13\ g/cm^3$

Q) A 5% solution (by mass) of cane sugar in water has freezing point of 271K. Calculate the freezing point of 5% glucose in water if freezing point of pure water is 273.15K.

Cane Sugar $C_{12}H_{22}O_{11}$ $(M = 342)$
Glucose $C_6H_{12}O_6$ $(M = 180)$

Solution:

$\Delta T_f = K_f \times m$ (for cane sugar)

$2.15 = K_f \times m$ 5% = 5g solute in 100g solution

$\Rightarrow$ 5g cane sugar + 95g water

$$2.15 = K_f \times \frac{\frac{5}{342}}{\frac{95}{1000}} \quad \text{(i)}$$

$\Delta T_f = K_f \times m$ (for glucose)

$$\Delta T_f = K_f \times \frac{\frac{5}{180}}{\frac{95}{1000}} \quad \text{(ii)} \quad (5g \text{ glucose} + 95g \text{ water})$$

(ii) ÷ (i)

$$\frac{\Delta T_f}{2.15} = \frac{342}{180}$$

$$\Delta T_f = 4.09$$

$$T_f = T_f^\circ - 4.09 = 273.15 - 4.09 = 269.06K$$

If K_f is not given (rare case)

$$K_f = \frac{R \times (T_f^\circ)^2 \times M}{1000 \times \Delta H_{fusion}}$$

(T_f° → in kelvin)

R → Universal Gas Constant (8.314 J/mol K OR 2 calorie/mol K)
(value as per unit of ΔH_{fusion})

T_f° → Freezing Point of Solvent (in kelvin)

M → Molar Mass of Solvent

ΔH_{fusion} → Molar Enthalpy of fusion.

($\Delta H_{fusion} = M \times L_{fusion}$)

M → Molar Mass of Solvent

L_{fusion} → Latent Heat of fusion

Q3) Calculate Molal depression Constant of a solvent which has Freezing Point at 16.6°C and Latent Heat of fusion 180.75 J/g

Solution $T_f^\circ = 16.6 + 273 = 289.6\,K$

Let M be Molar Mass of solvent

$\Delta H_{fusion} = M \times L_f = M \times 180.75\ J/g$

$R = 8.314\ J/mol\,K$

$$K_f = \frac{R \times (T_f^\circ)^2 \times M}{1000 \times \Delta H_{fusion}} = \frac{8.314 \times (289.6)^2 \times M}{1000 \times M \times 180.75}$$

$$\boxed{K_f = 3.857\ K\,kg/mol}$$