Molarity vs. molality

$$\frac{\text{mol solute}}{\text{L of solution}}$$

solute

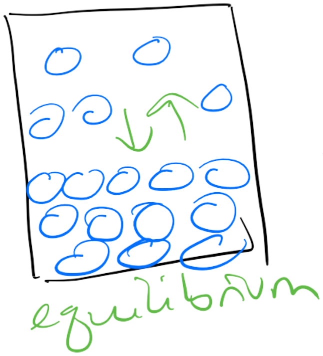
Solvent

$$\frac{\text{mol Solute}}{\text{Kg of solvent}}$$

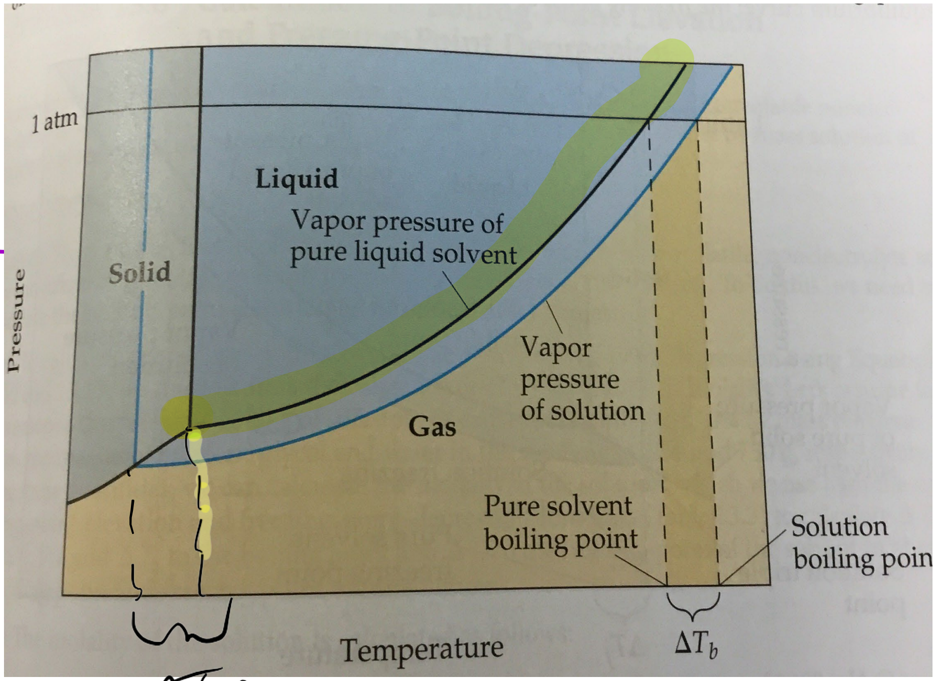
(H_2O)

Colligative Properties

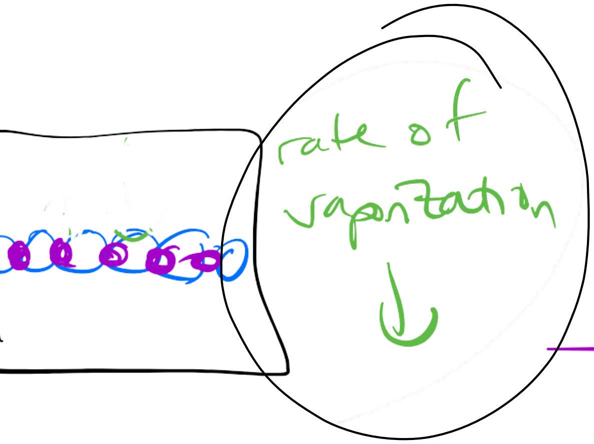
→ depend on amount of solu
not identity



add → solute



ΔT_f



→ equilibrium re-established
less molecules in gas phase

☆ vapor pressure is lowered

volatile solvent has a vapor pressure

non-volatile solute no measurable vapor pressure

colligative properties

- ☆ vapor pressure lowering
- boiling pt elevation
- freezing pt depression
- osmotic pressure

Raoult's Law

$$\chi \, P_{\text{silvent}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}}$$

$$X_{\text{solvent}} = \frac{n_{\text{solvent}}}{n_{\text{solvent}} + n_{\text{solute}}}$$

solutes

⊗ ideal solution
solutes do not react w/ water

A, B & C

Raoults Law continued

$$X_A = \frac{n_A}{n_A + n_B + n_C}$$

Glycerol is a non-volatile, non-electrolyte with a density of 1.26 g/ml at 25.0° C. Calculate the vapor pressure at 25.0 °C of a solution made by adding 50.0 ml of glycerol to 500.0 ml of water. The vapor pressure of pure water at 25 C is 23.8 torr and its density is 1.00 g/ml. The molar mass of glycerol is 92.09 g/mol.

$$P_{new} = X_{solvent} P^{\circ}_{solvent}$$

$$P_{new} =$$

(solvent)

$$n_{H_2O} = 500.0\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.0\,g} = 27.78\,mol\,H_2O$$

(solute)

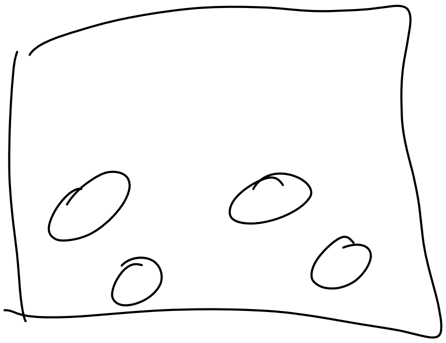
$$n_{glycerol} = \underset{3SF}{\underline{50.0}}\,\cancel{ml} \times \frac{1.26\,g}{\cancel{ml}} \times \frac{1\,mol}{92.09\,g} = 0.6841\,mol\ glycerol$$

$$X_{H_2O} = \frac{27.78\,mol}{27.78\,mol + 0.6841\,mol}$$

$$X = 0.976$$

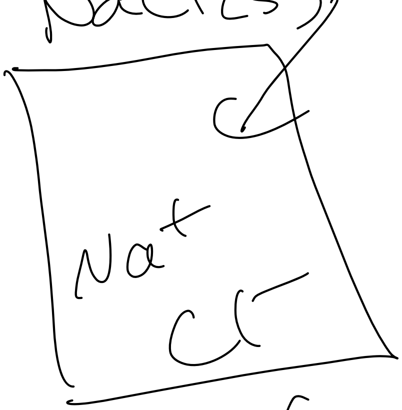
$$P = (23.8\,torr) \times (0.976) = 23.2\,torr$$

non-electrolyte vs. electrolyte



mol of glycerol "1"

NaCl(s)



moles of ions is "2"

$$X_{solvent} = \frac{n_{solvent}}{n_{solvent} + 2 \times n_{NaCl}}$$

Example: Raoult's law problem with a strong electrolyte

predict the vapor pressure of a solution prepared by mixing 35.8 g of sodium sulfate with 175.0 g of water. Vapor pressure of pure water @ 25.0°C is 23.8 torr. What is the new vapor pressure?

$Na_2SO_{4(s)} \rightarrow Na^+$ SO_4^{2-} $\boxed{Na_2SO_4}$

$2Na^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

moles ion = 3

Na^+ + Na^+

SO_4^{2-}

$$P_{new} = X_{solvent} P^\circ_{solvent}$$

$$X_{solvent}: \frac{n_{H_2O}}{n_{H_2O} + 3(n_{Na_2SO_4})}$$

$$n_{H_2O}: 175\,g \times \frac{1\,mol}{18\,g} = 9.72\,mol\ H_2O$$

$$n_{Na_2SO_4}: 35.0\,g\ Na_2SO_4 \times \frac{1\,mol}{142\,g} = 0.246\,mol\ Na_2SO_4$$

$$X_{solvent}: \frac{9.72\,mol}{9.72\,mol + 3(0.246\,mol)}$$

$$X_{H_2O}: 0.929$$

$$P_{new} = 0.929 \times 23.8\,torr$$

$$\boxed{P = 22.1\,torr}$$

Boiling Point Elevation

change in temp.

$$\Delta T_b = K_b m$$

constant unique to each solvent

solute

What is the boiling point elevation when 11.4 g of ammonia (NH_3) is dissolved in 200. g of water? K_b for water is 0.520°C/m.

$$\Delta T_b = K_b m$$

$$\Delta T_b = 0.520\ °C/m$$

Freezing Point depression

$$m = \frac{\text{mol solute}}{\text{Kg solvent}}$$

$\Delta Tf = k_f\ m$

$$11.4\,g\ NH_3 \times \frac{1\,mol}{17.0\,g} = 0.67\,mol\ NH_3$$

$$m = \frac{0.67\,mol}{0.2\,kg}$$

$$3.36\ m$$

$$\Delta T_b = (0.520\ \tfrac{°C}{\cancel{m}})(3.36\ \cancel{m})$$

$$\Delta T_b = 1.74°$$

new b.p. $100^{\circ}.0\,^{\circ}C + 1.74\,^{\circ}C =$

$\boxed{101.7\,^{\circ}C}$

Electrolytes vs. non electrolytes

$\Delta T_b = \textcircled{i}$

"i" - "Van ~~ffff~~ factor"

glucose

$\textcircled{C_6H_{12}O_6}$ vs. NaCl

$$NaCl \rightarrow Na^+ \quad Cl^-$$

if have 0.100m solution of each which solution will have a lower freezing point?

what solution has the highest boiling point?

a. 0.400 m glucose, non-electrolyte

b. 0.100 m $CaCl_2$

c. 0.200 m NaCl

d. 0.150 m $AlCl_3$

Van Hoff factor "i"

$\Delta T_f = i K_f m$

For NaCl $i = 2$

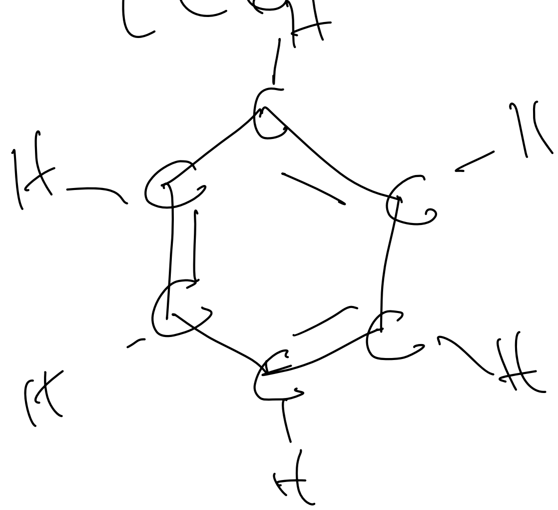
For glucose

$i = 1$

$\Delta T_f = k_f \cdot m$ "freezing pt depression"

Colligative Properties: finding molar mass from freezing pt. depression

When 0.258g of a non-electrolyte was dissolved in 40.0g of benzene (C_6H_6), $K_f = 5.12\ °C/m$. the ΔT_f was 0.27°C.



① $\Delta T_f = K_f \cdot \boxed{m}$

$$m = \frac{\Delta T_f}{K_f}$$

$$m = \frac{0.27\ \cancel{°C}}{5.12\ \frac{\cancel{°C}}{m}} = 0.0527\ m$$

②

$$\checkmark m = \frac{\boxed{\text{mol solute}}}{\checkmark \text{ kg solvent}}$$

$$\text{mol solute} = (0.0400\,\cancel{\text{kg}}) \times \left(0.0527\,\frac{\text{mol}}{\cancel{\text{kg}}}\right)$$

$$= 0.00211 \text{ mol solute}$$

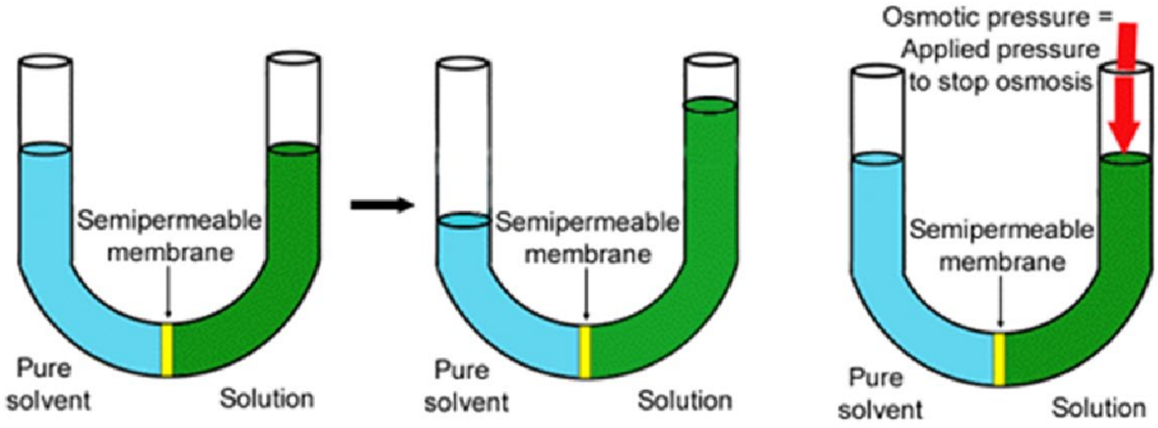
③ $\frac{\text{g}}{\text{mol}}$

$$\text{molar mass} : \frac{0.258\text{g}}{0.00211\text{mol}}$$

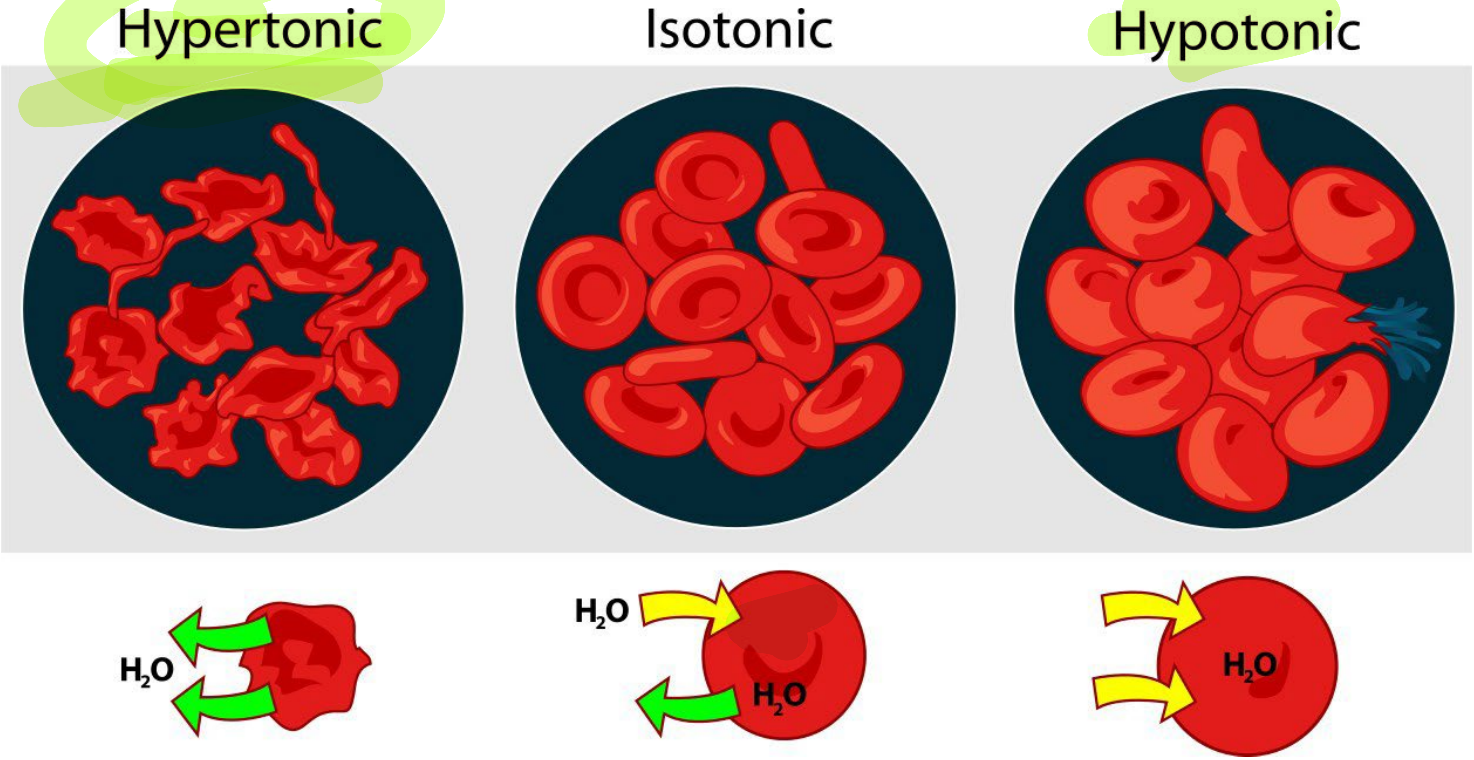
$$\text{Molar mass} = 122 \frac{\text{g}}{\text{mol}}$$

Osmotic Pressure

The minimum pressure that stops the osmosis is equal to the osmotic pressure of the solution



Osmosis in Blood Cells



The average osmotic pressure of blood is 7.7 atm at 25°C. What concentration of glucose will be isotonic with blood?

molarity

$$\Pi = MRT$$

temperature in kelvin

pressure

gas law constant

$0.08206 \frac{L \cdot atm}{mol \cdot K}$

$$P = \frac{n}{V} RT$$

$$\Pi = MRT$$

$$\frac{\Pi}{R \cdot T} = M$$

K = °C + 273

$$M = \frac{7.7 \cancel{atm}}{(0.08206 \frac{L \cdot \cancel{atm}}{mol \cdot \cancel{K}})(298 \cancel{K})}$$

$$M = 0.315 M$$

Osmotic Pressure and Molar Mass Example

A solution is prepared by dissolving 35.0 g of hemoglobin in enough water to make up 1.00 L in volume. The osmotic pressure of the solution is found to be 10.0 mmHg at 25.0 °C. Calculate the molar mass of hemoglobin.

$25.0° + 273$

$10.0\ \text{mmg}$

$\boxed{298K}$

$$\pi = \boxed{M}RT$$

$$R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$$

$$M = \frac{\pi}{RT}$$

$$\pi = \frac{\frac{\text{mass of solut}}{\text{MMsolute}}}{\text{L of soln}} \times (R$$

$$M = \frac{0.0132\ \cancel{atm}}{\left(0.0826 \frac{L \cdot \cancel{atm}}{mol \cdot \cancel{K}}\right)(298\cancel{K})}$$

$$M = 5.4 \times 10^{-5} \frac{mol}{L}$$

$$mol = 5.4 \times 10^{-5} \frac{mol}{\cancel{L}} \times 1\cancel{L}$$

$$mM = \frac{g}{mol} \qquad = 5.4 \times 10^{-5}$$

$$MM = \frac{35.0g}{5.4 \times 10^{-5} mol}$$

$$6.48 \times 10^{4} \frac{g}{mol}$$