

Chapter 11- Properties of Solutions

11.1 Solution Composition

molarity- $\frac{\text{\# of moles}}{\text{L of solution}}$

mass percent- $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$

mole fraction- $\frac{\text{moles of solute}}{\text{moles of solution}}$

molality- $\frac{\text{moles of solute}}{\text{kg of solvent}}$

* sample exercises 11.1 and 11.2 may be helpful

11.2 The Energies of Solution Formation

ΔH of solution- enthalpy of solution, which may be defined as + or –

$$= \Delta H_1 + \Delta H_2 + \Delta H_3$$

ΔH_1 - expanding solute

ΔH_2 - expanding solvent

ΔH_3 - combining expanded solute and solvent to form solution

Δ hydration- enthalpy of hydration (interactions between solvent and solute)

Ex. H_2O (ΔH_2) and ionic solids (ΔH_3)

*sample exercise 11.3

Question: Decide whether liquid hexane (C_6H_{14}) or liquid methanol (CH_3OH) is the more appropriate solvent for the substances grease ($\text{C}_{20}\text{H}_{42}$) and potassium iodide (KI).

Answer: Hexane is a nonpolar solvent because it contains C-H bonds. Thus hexane will work best for the nonpolar solute grease. Methanol has an O-H group that makes it significantly polar. Thus it will serve as the better solvent for the ionic solid KI.

11.3 Factors Affecting Solubility

A) structure

B) pressure (Henry's Law) - $P = KC$

P = partial pressure, k = constant characteristic of gas, C = concentration of dissolved Gas

-Henry's law only works if there is no reaction between the solute and solvent

*sample exercise 11.4

Question: A certain soft drink is bottled so that a bottle at 25°C contains CO₂ gas at a pressure of 5.0 atm over the liquid. Assuming that the partial pressure of CO₂ in the atmosphere is 4.0 x 10⁻⁴ atm, calculate the equilibrium concentrations of CO₂ in the soda both before and after the bottle is opened. Henry's law constant for CO₂ in aqueous solution is 32 Lxatm/mol at 25°C.

Answer:

We can write Henry's law for CO₂ as

$$P_{\text{CO}_2} = k_{\text{CO}_2} C_{\text{CO}_2}$$

Where $k_{\text{CO}_2} = 32 \text{ L} \times \text{atm/mol}$. In the unopened bottle, $P_{\text{CO}_2} = 5.0 \text{ atm}$ and

$$C_{\text{CO}_2} = (P_{\text{CO}_2})/(k_{\text{CO}_2}) = 5.0\text{atm}/(32 \text{ L} \times \text{atm/mol}) = .16 \text{ mol/L}$$

In the opened bottle, the CO₂ in the soda eventually reaches equilibrium with the atmospheric CO₂, so $P_{\text{CO}_2} = 4.0 \times 10^{-4} \text{ atm}$ and

$$C_{\text{CO}_2} = P_{\text{CO}_2}/k_{\text{CO}_2} = (4.0 \times 10^{-4})/(32 \text{ L} \times \text{atm/mol}) = 1.2 \times 10^{-5} \text{ mol/L}$$

Note the large change in concentration of CO₂. This is why soda goes "flat" after being open for a while.

C) temperature- most solids become more soluble with increased temperature
-gases typically become less soluble with increased temperature

11.4 The Vapor Pressure of Solutions

-dissolved nonvolatile solute decreases the vapor pressure of the solvent by inhibiting its escape

$$- P_{\text{solution}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}}$$

P- obscured vapor pressure

X- mole fraction of solvent

P^o_{solvent} – vapor pressure of pure solvent

*sample exercise 11.5

Question: Calculate the expected vapor pressure at 25°C for a solution prepared by dissolving 158.0g of common table sugar (sucrose, molar mass= 342.5 g/mol) in 643.5 cm³ of water. At 25°C the density of water is 0.9971 g/cm³ and the vapor pressure is 23.76 torr.

Answer:

We will use Raoult's law in the form

$$P_{\text{soln}} = X_{\text{H}_2\text{O}} P^{\circ}_{\text{H}_2\text{O}}$$

To calculate the mole fraction of water in the solution, we must first determine the number of moles of sucrose:

$$\begin{aligned}\text{Moles of sucrose} &= 158.0 \text{ g sucrose} \times (1 \text{ mol sucrose}) / (342.3 \text{ g sucrose}) \\ &= 0.4616 \text{ mol sucrose}\end{aligned}$$

To determine the moles of water present, we first convert volume to mass using the density:

$$643.5 \text{ cm}^3 \text{ H}_2\text{O} \times (1 \text{ mol H}_2\text{O}) / (18.01 \text{ g H}_2\text{O}) = 35.63 \text{ mol H}_2\text{O}$$

The number of moles of water is therefore

$$641.6 \text{ g H}_2\text{O} \times (1 \text{ mole H}_2\text{O}) / (18.01 \text{ g H}_2\text{O}) = 35.63 \text{ mol H}_2\text{O}$$

The mole fraction of water in the solution is

$$\begin{aligned}X_{\text{H}_2\text{O}} &= (\text{mol H}_2\text{O}) / (\text{mol H}_2\text{O} + \text{mol sucrose}) = (35.63 \text{ mol}) / (35.63 \text{ mol} + .4616 \\ \text{mol}) \\ &= (35.63 \text{ mol}) / (36.09 \text{ mol}) = .9873\end{aligned}$$

Then $P_{\text{soln}} = X_{\text{solvent}} P_{\text{solvent}}^{\circ} = (.9873)(23.76 \text{ torr}) = 23.46 \text{ torr}$

Thus the vapor pressure of water has been lowered from 23.76 torr in the pure state to 23.46 torr in the solution. The vapor pressure has been lowered by .30 torr.

11.5 Boiling-Point Elevation and Freezing-Point Depression (Colligative Properties)

boiling-point elevation- nonvolatile solute elevates the boiling point of a solvent by lowering its vapor pressure.

$$-\Delta T = K_b M_{\text{solute}}$$

ΔT - difference in boiling point

K_b - molality boiling-point elevation constant

M_{solute} - molality of solute

*sample exercise 11.8

Question: A solution was prepared by dissolving 18.00 g glucose in 150.0 g water. The resulting solution was found to have a boiling point of 100.34°C. Calculate the molar mass of glucose. Glucose is a molecular solid that is present as individual molecules in solution.

Solution:

We make use of the equation

$$\begin{aligned}\Delta T &= K_b m_{\text{solute}} \\ \text{where } \Delta T &= 100.34^{\circ}\text{C} - 100.00^{\circ}\text{C} = .34^{\circ}\text{C}\end{aligned}$$

From Table 11.5 on P. 528, for water $K_b = .51$. The molality of this solution then can be calculated by rearranging the boiling-point elevation equation to give

$$m_{\text{solute}} = (\Delta T)/(K_b) = (.34^\circ\text{C})/(.51^\circ\text{C} \times \text{kg/mol}) = .67 \text{ mol/kg}$$

The solution was prepared using .1500kg water. Using the definition of molality we can find the number of moles of glucose in the solution.

$$m_{\text{solute}} = .67 \text{ mol/kg} = (\text{mol solute})/(\text{kg solvent}) = (n_{\text{glucose}})/(.1500 \text{ kg})$$

$$n_{\text{glucose}} = (.67 \text{ mol/kg})(.1500 \text{ kg}) = .10 \text{ mol}$$

Thuse .10 mol glucose has a mass of 18.00 g, and 1.0 mol glucose has a mass of 180 g (10 x 18.00 g). The molar mass of glucose is 180 g/mol.

Freezing-point depression- solute depresses the freezing- point of the solvent by lowering the vapor pressure of the liquid (have to reach new liquid/solid equilibrium)

$$\Delta T = K_f m_{\text{solute}} \text{ (similar to boiling-point elevation)}$$

* sample exercise 11.9

Question: What mass of ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$, molar mass = 62.1 g/mol), the main component of antifreeze, must be added to 10.0 L water to produce a solution for use in a car's radiator that freezes at -10.0°F (-23.3°C)? Assume the density of water is exactly 1g/mL.

Answer: The freezing point must be lowered from 0°C to -23.3°C . To determine the molality of ethylene glycol needed to accomplish this, we can use the equation

$$\Delta T = K_f m_{\text{solute}}$$

where $\Delta T = 23.3^\circ\text{C}$ and $K_f = 1.86$ (from Table 11.5). Solving for the molality gives

$$m_{\text{solute}} = (\Delta T)/(K_f) = (23.3^\circ\text{C})/(1.86^\circ\text{C} \times \text{kg/mol}) = 12.5 \text{ mol/kg}$$

This means that 12.5 mol ethylene glycol must be added per kilogram of water. We have 10.0 L, or 10.0 kg, of water. Therefore, the total number of moles of ethylene glycol needed is

$$(12.5 \text{ mol})/(\text{kg}) \times 10.0 \text{ kg} = 1.25 \times 10^2 \text{ mol}$$

The mass of ethylene glycol needed is

$$1.25 \times 10^2 \text{ mol} \times 62.1 \text{ g/mol} = 7.76 \times 10^3 \text{ g (or 7.76 kg)}$$

11.6 Osmotic Pressure

$$-\pi = MRT$$

semipermeable membrane- allow solvent but not solute molecules to pass through

osmosis- the flow of solvent into a solution through a semipermeable membrane

osmotic pressure- excess hydrostatic pressure on the solution

isotonic solutions- solutions that have identical osmotic pressures

reverse osmosis- opposite flow to that of osmosis (external pressure greater than osmotic pressure)

sample exercise 11.12

Question: What concentration of sodium chloride in water is needed to produce an aqueous solution isotonic with blood ($\pi = 7.70$ atm at 25°C)?

Solution:

We can calculate the molarity of the solute from the equation

$$\pi = MRT \text{ or } M = \pi/(RT)$$

$$M = (7.70 \text{ atm}) / \{(.08206 \text{ L} \times \text{atm}/\text{K} \times \text{mol})(298\text{K})\} = .315 \text{ mol/L}$$

This represents the total molarity of solute particles. But NaCl gives two ions per formula unit. Therefore, the concentration of NaCl needed is $.315 \text{ M}/2$, or $.158 \text{ M}$. That is,



11.7 Colligative Properties of Electrolyte Solution

-Vant Hoff factor (i) is an expected value used to determine the number of solute particles to affect solutions (ion pairing causes the solution to deviate from I) and is affected more by increased ion charges

-modified freezing-point and boiling-point equation for electrolyte solutions:

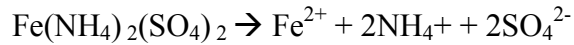
$$\Delta T = imk$$

$$\pi = iMRT \text{ for osmotic}$$

sample exercise 11.13

Question: The observed osmotic pressure for a $.10 \text{ M}$ solution of $\text{Fe}(\text{NH}_4)_2(\text{SO}_4)_2$ at 25°C is 10.8 atm. Compare the expected and experimental values for i .

Answer: The ionic solid $\text{Fe}(\text{NH}_4)_2(\text{SO}_4)_2$ dissociates in water to produce 5 ions:



Thus the expected value for i is 5. We can obtain the experimental value for i by using the equation for osmotic pressure:

$$\pi = iMRT \text{ or } i = \pi/(MRT)$$

where $\pi = 10.8 \text{ atm}$, $M = .10 \text{ mol/L}$, $R = (.08206 \text{ L x atm})/(\text{k x mol})$, and $T = 25 + 273 = 298 \text{ K}$. Substituting these values into the equation gives

$$i = \pi/(MRT) = (10.8 \text{ atm})/\{(.10 \text{ mol/L})(.08206 \text{ L x atm/K x mol})(298 \text{ K})\} = 4.4$$

The experimental value for i is less than the expected value, presumably because of ion pairing.

11.8 Colloids

tyndall effect- the scattering of light by particles within a suspension (used to distinguish between a suspension and a true solution)

colloid (colloidal dispersion)- suspension of tiny particles

- classified according to the states of matter
- the dispersed phase and dispersing medium (table 11.7)

coagulation- colloidal particles aggregate