

CHEM 101

Chapter 12 Problems

12.9 Why is naphthalene (C_{10}H_8) more soluble than CsF in benzene?

CsF is an ionic solid; the ion-ion attractions are too strong to be overcome in the dissolving process in benzene. The ion-induced dipole interaction is too weak to stabilize the ion.

Nonpolar naphthalene molecules form a molecular solid in which the only interparticle forces are of the weak dispersion type. The same forces operate in liquid benzene causing naphthalene to dissolve with relative ease. Like dissolves like.

12.10 Explain why ethanol ($\text{C}_2\text{H}_5\text{OH}$) is not soluble in cyclohexane (C_6H_{12})

Cyclohexane cannot form hydrogen bonds with ethanol, and therefore cannot attract ethanol molecules strongly enough to form a solution.

12.15 Calculate the percent by mass of the solute in each of the following aqueous solutions:

(a) 5.50 g of NaBr in 78.2 g of solution,

(b) 31.0 g of KCl in 152 g of water,

(c) 4.5 g of toluene in 29 g of benzene C₆H₆.

$$(a) \quad \frac{5.50 \text{ g NaBr}}{78.2 \text{ g soln}} \times 100\% = \mathbf{7.03\%}$$

$$(b) \quad \frac{31.0 \text{ g KCl}}{(31.0 + 152) \text{ g soln}} \times 100\% = \mathbf{16.9\%}$$

$$(c) \quad \frac{4.5 \text{ g toluene}}{(4.5 + 29) \text{ g soln}} \times 100\% = \mathbf{13\%}$$

- 12.16** Calculate the amount of water (in grams) that must be added to
- (a) 5.00 g of urea $(\text{NH}_2)_2\text{CO}$ in the preparation of a 16.2 percent by mass solution.
- (b) 26.2 g of MgCl_2 in the preparation of a 1.5 percent by mass solution.

(a) The percent by mass is defined as

$$\text{percent by mass of solute} = \frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100\%$$

Substituting in the percent by mass of solute and the mass of solute, we can solve for the mass of solvent (water).

$$16.2\% = \frac{5.00 \text{ g urea}}{5.00 \text{ g urea} + \text{mass of water}} \times 100\%$$

$$(0.162)(\text{mass of water}) = 5.00 \text{ g} - (0.162)(5.00\text{g})$$

$$\text{mass of water} = 25.9 \text{ g}$$

(b) Similar to part (a),

$$1.5\% = \frac{26.2 \text{ g MgCl}_2}{26.2 \text{ g MgCl}_2 + \text{mass of water}} \times 100\%$$

$$\text{mass of water} = 1.72 \times 10^3 \text{ g}$$

12.17 Calculate the molality of each of the following solutions:

- (a) 14.3 g of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) in 676 g of water,
(b) 7.20 moles of ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$) in 3546 g of water

- (a) The molality is the number of moles of sucrose (molar mass 342.3 g/mol) divided by the mass of the solvent (water) in kg.

$$\text{mol sucrose} = 14.3 \cancel{\text{g sucrose}} \times \frac{1 \text{ mol}}{342.3 \cancel{\text{g sucrose}}} = 0.0418 \text{ mol}$$

$$\text{Molality} = \frac{0.0418 \text{ mol sucrose}}{0.676 \text{ kg H}_2\text{O}} = \mathbf{0.0618 \text{ } m}$$

$$\text{(b) Molality} = \frac{7.20 \text{ mol ethylene glycol}}{3.546 \text{ kg H}_2\text{O}} = \mathbf{2.03 \text{ } m}$$

12.18 Calculate the molality of each of the following aqueous solutions:

(a) 2.50 M NaCl solution (density of solution = 1.08 g/mL),

(b) 48.2 percent by mass KBr solution.

$$(a) \quad \text{mass of 1 L soln} = 1000 \cancel{\text{ mL}} \times \frac{1.08 \text{ g}}{1 \cancel{\text{ mL}}} = 1080 \text{ g}$$

$$\text{mass of water} = 1080 \text{ g} - \left(2.50 \cancel{\text{ mol}} \text{ NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \cancel{\text{ mol}} \text{ NaCl}} \right) = 934 \text{ g} = 0.934 \text{ kg}$$

$$m = \frac{2.50 \text{ mol NaCl}}{0.934 \text{ kg H}_2\text{O}} = \mathbf{2.68 \text{ } m}$$

(b) 100 g of the solution contains 48.2 g KBr and 51.8 g H₂O.

$$\text{mol of KBr} = 48.2 \cancel{\text{ g}} \text{ KBr} \times \frac{1 \text{ mol KBr}}{119.0 \cancel{\text{ g}} \text{ KBr}} = 0.405 \text{ mol KBr}$$

$$\text{mass of H}_2\text{O (in kg)} = 51.8 \cancel{\text{ g}} \text{ H}_2\text{O} \times \frac{1 \text{ kg}}{1000 \cancel{\text{ g}}} = 0.0518 \text{ kg H}_2\text{O}$$

$$m = \frac{0.405 \text{ mol KBr}}{0.0518 \text{ kg H}_2\text{O}} = \mathbf{7.82 \text{ } m}$$

12.19 Calculate the molalities of the following aqueous solution:

b) 0.87 *M* NaOH solution (density of solution = 1.04 g/mL),

$$\text{(b) mass of NaOH} = 0.87 \cancel{\text{ mol NaOH}} \times \frac{40.00 \text{ g NaOH}}{1 \cancel{\text{ mol NaOH}}} = 35 \text{ g NaOH}$$

$$\text{mass solvent (H}_2\text{O)} = 1040 \text{ g} - 35 \text{ g} = 1005 \text{ g} = 1.005 \text{ kg}$$

$$\text{molality} = \frac{0.87 \text{ mol NaOH}}{1.005 \text{ kg H}_2\text{O}} = \mathbf{0.87 \text{ } m}$$

12.23 Calculate the molarity and the molality of an NH_3 solution made up of 30.0 g of NH_3 in 70.0 g of water.

The density of the solution is 0.982 g/mL

$$\text{mol NH}_3 = 30.0 \cancel{\text{g NH}_3} \times \frac{1 \text{ mol NH}_3}{17.03 \cancel{\text{g NH}_3}} = 1.76 \text{ mol NH}_3$$

$$\text{Volume of the solution} = 100.0 \cancel{\text{g soln}} \times \frac{1 \cancel{\text{mL}}}{0.982 \cancel{\text{g}}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 0.102 \text{ L}$$

$$\text{molarity} = \frac{1.76 \text{ mol NH}_3}{0.102 \text{ L soln}} = \mathbf{17.3 \text{ M}}$$

$$\text{kg of solvent (H}_2\text{O)} = 70.0 \cancel{\text{g H}_2\text{O}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} = 0.0700 \text{ kg H}_2\text{O}$$

$$\text{molality} = \frac{1.76 \text{ mol NH}_3}{0.0700 \text{ kg H}_2\text{O}} = \mathbf{25.1 \text{ m}}$$

12.37 The solubility of CO₂ in water at 25°C and 1 atm is 0.034 mol/L. What is its solubility under atmospheric conditions? (The partial pressure of CO₂ in air is 0.0003 atm.) Assume that CO₂ obeys Henry's law

We first find the value of k for Henry's law

$$k = \frac{c}{P} = \frac{0.034 \text{ mol/L}}{1 \text{ atm}} = 0.034 \text{ mol/L} \cdot \text{atm}$$

For atmospheric conditions we write:

$$c = kP = (0.034 \text{ mol/L} \cdot \cancel{\text{atm}})(0.00030 \cancel{\text{ atm}}) = 1.0 \times 10^{-5} \text{ mol/L}$$

12.49 A solution is prepared by dissolving 396 g of sucrose ($C_{12}H_{22}O_{11}$) in 624 g of water. What is the vapor pressure of this solution at 30°C? (The vapor pressure of water is 31.8 mmHg at 30°C.)

$$\text{Moles sucrose} = 396 \cancel{\text{g}} \times \frac{1 \text{ mol}}{342.3 \cancel{\text{g}}} = 1.16 \text{ mol sucrose}$$

$$\text{Moles water} = 624 \cancel{\text{g}} \times \frac{1 \text{ mol}}{18.02 \cancel{\text{g}}} = 34.6 \text{ mol water}$$

The mole fraction of water is:

$$X_{H_2O} = \frac{34.6 \text{ mol}}{34.6 \text{ mol} + 1.16 \text{ mol}} = 0.968$$

The vapor pressure of the solution is found as follows:

$$P_{\text{solution}} = X_{H_2O} \times P_{H_2O}^{\circ} = (0.968)(31.8 \text{ mmHg}) = \mathbf{30.8 \text{ mmHg}}$$

12.50 How many grams of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) must be added to 552 g of water to give a solution with a vapor pressure 2.0 mmHg less than that of pure water at 20°C? (The vapor pressure of water at 20°C is 17.5 mmHg.)

$$\Delta P = X_2 P_1^\circ$$

$$\Delta P = X_{\text{sucrose}} P_{\text{water}}^\circ$$

$$X_{\text{sucrose}} = \frac{\Delta P}{P_{\text{water}}^\circ} = \frac{2.0 \text{ mmHg}}{17.5 \text{ mmHg}} = 0.11$$

From the definition of mole fraction, we can calculate moles of sucrose.

$$X_{\text{sucrose}} = \frac{n_{\text{sucrose}}}{n_{\text{water}} + n_{\text{sucrose}}}$$

$$\text{moles of water} = 552 \text{ g} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = 30.6 \text{ mol H}_2\text{O}$$

$$X_{\text{sucrose}} = 0.11 = \frac{n_{\text{sucrose}}}{30.6 + n_{\text{sucrose}}}$$

$$n_{\text{sucrose}} = 3.8 \text{ mol sucrose}$$

Using the molar mass of sucrose as a conversion factor, we can calculate the mass of sucrose.

$$\text{mass of sucrose} = 3.8 \text{ mol sucrose} \times \frac{342.3 \text{ g sucrose}}{1 \text{ mol sucrose}} = 1.3 \times 10^3 \text{ g sucrose}$$

12.52 The vapor pressures of ethanol ($\text{C}_2\text{H}_5\text{OH}$) and 1-propanol ($\text{C}_3\text{H}_7\text{OH}$) at 35°C are 100 mmHg and 37.6 mmHg, respectively. Assume ideal behavior and calculate the partial pressures of ethanol and 1-propanol at 35°C over a solution of ethanol in 1-propanol, in which the mole fraction of ethanol is 0.300

For any solution the sum of the mole fractions of the components is always 1.00, so the mole fraction of 1-propanol is 0.700. The partial pressures are:

$$P_{\text{ethanol}} = X_{\text{ethanol}} \times P_{\text{ethanol}}^\circ = (0.300)(100 \text{ mmHg}) = \mathbf{30.0 \text{ mmHg}}$$

$$P_{\text{1-propanol}} = X_{\text{1-propanol}} \times P_{\text{1-propanol}}^\circ = (0.700)(37.6 \text{ mmHg}) = \mathbf{26.3 \text{ mmHg}}$$

12.55 What are the boiling point and freezing point of a 2.47 *m* solution of naphthalene in benzene? (The boiling point and freezing point of benzene are 80.1°C and 5.5°C, respectively.)

$$\Delta T_b = K_b m = (2.53^\circ\text{C}/\cancel{m})(2.47 \cancel{m}) = 6.25^\circ\text{C}$$

The new **boiling point** is $80.1^\circ\text{C} + 6.25^\circ\text{C} = 86.4^\circ\text{C}$

$$\Delta T_f = K_f m = (5.12^\circ\text{C}/\cancel{m})(2.47 \cancel{m}) = 12.6^\circ\text{C}$$

The new **freezing point** is $5.5^\circ\text{C} - 12.6^\circ\text{C} = -7.1^\circ\text{C}$

12.56 An aqueous solution contains the amino acid glycine ($\text{NH}_2\text{CH}_2\text{COOH}$). Assuming that the acid does not ionize in water, calculate the molality of the solution if it freezes at -1.1°C .

$$m = \frac{\Delta T_f}{K_f} = \frac{1.1^\circ\cancel{\text{C}}}{1.86^\circ\cancel{\text{C}}/\cancel{m}} = 0.59 \text{ } m$$

12.60 A solution is prepared by condensing 4.00 L of a gas, measured at 27°C and 748 mmHg pressure, into 58.0 g of benzene. Calculate the freezing point of this solution.

We first find the number of moles of gas using the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{\left(748 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right)(4.00 \text{ L})}{(27 + 273) \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.0821 \text{ L} \cdot \text{atm}} = 0.160 \text{ mol}$$

$$\text{molality} = \frac{0.160 \text{ mol}}{0.0580 \text{ kg benzene}} = 2.76 \text{ m}$$

$$\Delta T_f = K_f m = (5.12^\circ\text{C}/m)(2.76 \text{ m}) = 14.1^\circ\text{C}$$

$$\text{freezing point} = 5.5^\circ\text{C} - 14.1^\circ\text{C} = -8.6^\circ\text{C}$$

2.62 A solution of 2.50 g of a compound having the empirical formula $\text{C}_6\text{H}_5\text{P}$ in 25.0 g of benzene is observed to freeze at 4.3°C . Calculate the molar mass of the solute and its molecular formula.

First, from the freezing point depression we can calculate the molality of the solution. See Table 12.2 of the text for the normal freezing point and K_f value for benzene.

$$\Delta T_f = (5.5 - 4.3)^\circ\text{C} = 1.2^\circ\text{C}$$

$$m = \frac{\Delta T_f}{K_f} = \frac{1.2^\circ\cancel{\text{C}}}{5.12^\circ\cancel{\text{C}}/\text{m}} = 0.23 \text{ m}$$

Multiplying the molality by the mass of solvent (in kg) gives moles of unknown solute. Then, dividing the mass of solute (in g) by the moles of solute, gives the molar mass of the unknown solute.

$$\begin{aligned} ? \text{ mol of unknown solute} &= \frac{0.23 \text{ mol solute}}{1 \cancel{\text{kg}} \text{ benzene}} \times 0.0250 \cancel{\text{kg}} \text{ benzene} \\ &= 0.0058 \text{ mol solute} \end{aligned}$$

$$\text{molar mass of unknown} = \frac{2.50 \text{ g}}{0.0058 \text{ mol}} = 4.3 \times 10^2 \text{ g/mol}$$

The empirical molar mass of $\text{C}_6\text{H}_5\text{P}$ is 108.1 g/mol. Therefore, the molecular formula is $(\text{C}_6\text{H}_5\text{P})_4$ or $\text{C}_{24}\text{H}_{20}\text{P}_4$.

12.63 What is the osmotic pressure (in atm) of a 1.36 M aqueous solution of urea $[(\text{NH}_2)_2\text{CO}]$ at 22.0°C?

$$\pi = MRT = (1.36 \text{ mol/L})(0.0821 \text{ L}\cdot\text{atm/K}\cdot\text{mol})(22.0 + 273)\text{K} = 32.9 \text{ atm}$$

12.64 A solution containing 0.8330 g of a polymer of unknown structure in 170.0 mL of an organic solvent was found to have an osmotic pressure of 5.20 mmHg at 25°C. Determine the molar mass of the polymer?

$$\pi = MRT$$

$$M = \frac{\pi}{RT} = \frac{\left(5.20 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right)}{298 \text{ K}} \times \frac{\text{mol}\cdot\text{K}}{0.0821 \text{ L}\cdot\text{atm}} = 2.80 \times 10^{-4} M$$

Multiplying the molarity by the volume of solution (in L) gives moles of solute (polymer).

$$? \text{ mol of polymer} = (2.80 \times 10^{-4} \text{ mol/L})(0.170 \text{ L}) = 4.76 \times 10^{-5} \text{ mol polymer}$$

Lastly, dividing the mass of polymer (in g) by the moles of polymer, gives the molar mass of the polymer.

$$\text{molar mass of polymer} = \frac{0.8330 \text{ g polymer}}{4.76 \times 10^{-5} \text{ mol polymer}} = 1.75 \times 10^4 \text{ g/mol}$$

12.66 A solution of 6.85 g of a carbohydrate in 100.0 g of water has a density of 1.024 g/mL and an osmotic pressure of 4.61 atm at 20.0°C. Calculate the molar mass of the carbohydrate

$$M = \frac{\pi}{RT} = \frac{4.61 \cancel{\text{atm}}}{(20 + 273)\cancel{\text{K}}} \times \frac{\cancel{\text{mol}} \cdot \cancel{\text{K}}}{0.0821 \text{ L} \cdot \cancel{\text{atm}}} = 0.192 \text{ mol/L}$$

Next we use the density and the solution mass to find the volume of the solution.

$$\text{mass of soln} = 6.85 \text{ g} + 100.0 \text{ g} = 106.9 \text{ g soln}$$

$$\text{volume of soln} = 106.9 \cancel{\text{g}} \text{ soln} \times \frac{1 \text{ mL}}{1.024 \cancel{\text{g}}} = 104.4 \text{ mL} = 0.1044 \text{ L}$$

Multiplying the molarity by the volume (in L) gives moles of solute (carbohydrate).

$$\text{mol of solute} = M \times L = (0.192 \cancel{\text{mol/L}})(0.1044 \cancel{\text{L}}) = 0.0200 \text{ mol solute}$$

Finally, dividing mass of carbohydrate by moles of carbohydrate gives the molar mass of the carbohydrate.

$$\text{molar mass} = \frac{6.85 \text{ g carbohydrate}}{0.0200 \text{ mol carbohydrate}} = \mathbf{343 \text{ g/mol}}$$

12.81 Lysozyme is an enzyme that cleaves bacterial cell walls. A sample of lysozyme extracted from egg white has a molar mass of 13,930 g. A quantity of 0.100 g of this enzyme is dissolved in 150 g of water at 25°C. Calculate the vapor-pressure lowering, the depression in freezing point, the elevation in boiling point, and the osmotic pressure of this solution. (The vapor pressure of water at 25°C is 23.76 mmHg.)

$$n_{\text{lysozyme}} = 0.100 \cancel{\text{g}} \times \frac{1 \text{ mol}}{13930 \cancel{\text{g}}} = 7.18 \times 10^{-6} \text{ mol}$$

$$n_{\text{water}} = 150 \cancel{\text{g}} \times \frac{1 \text{ mol}}{18.02 \cancel{\text{g}}} = 8.32 \text{ mol}$$

Vapor pressure lowering: $\Delta P = X_{\text{lysozyme}} P_{\text{water}}^{\circ} = \frac{n_{\text{lysozyme}}}{n_{\text{lysozyme}} + n_{\text{water}}} (23.76 \text{ mmHg})$

$$\Delta P = \frac{7.18 \times 10^{-6} \cancel{\text{mol}}}{[(7.18 \times 10^{-6}) + 8.32] \cancel{\text{mol}}} (23.76 \text{ mmHg}) = 2.05 \times 10^{-5} \text{ mmHg}$$

Freezing point depression: $\Delta T_f = K_f m = (1.86^{\circ}\text{C}/\cancel{m}) \left(\frac{7.18 \times 10^{-6} \cancel{\text{mol}}}{0.150 \cancel{\text{kg}}} \right) = 8.90 \times 10^{-5} ^{\circ}\text{C}$

Boiling point elevation: $\Delta T_b = K_b m = (0.52^{\circ}\text{C}/\cancel{m}) \left(\frac{7.18 \times 10^{-6} \cancel{\text{mol}}}{0.150 \cancel{\text{kg}}} \right) = 2.5 \times 10^{-5} ^{\circ}\text{C}$

Osmotic pressure: As stated above, we assume the density of the solution is 1.00 g/mL. The volume of the solution will be 150 mL.

$$\pi = MRT = \left(\frac{7.18 \times 10^{-6} \cancel{\text{mol}}}{0.150 \cancel{\text{L}}} \right) (0.0821 \cancel{\text{L}} \cdot \cancel{\text{atm}} / \cancel{\text{mol}} \cdot \cancel{\text{K}}) (298 \cancel{\text{K}}) = 1.17 \times 10^{-3} \text{ atm} = 0.889 \text{ mmHg}$$

Note that only the osmotic pressure is large enough to measure.