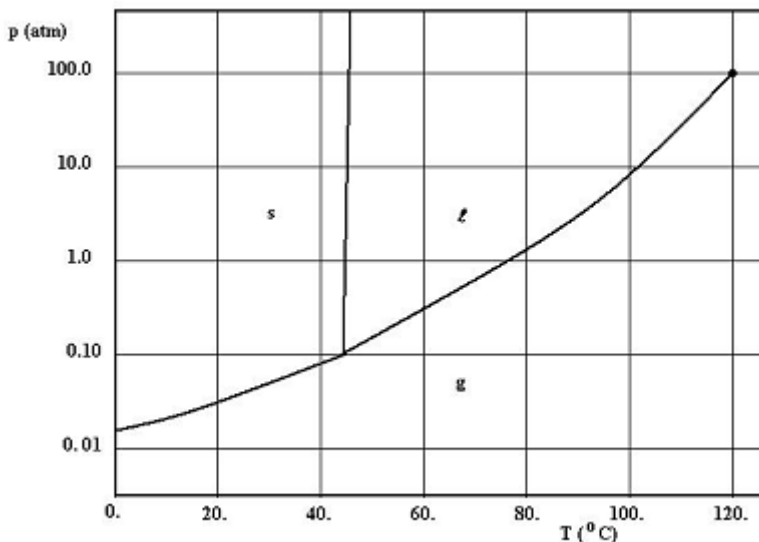


Problems - Chapter 13 (with solutions)

1) The phase diagram for a pure chemical substance is given below, and may be used to answer the following questions.



a) What are the values for p and T at the triple point of the substance?

At the triple point $p = 0.10$ atm, $T = 44.$ $^{\circ}\text{C}$.

b) What are the values for p and T at the critical point of the substance?

At the critical point $p = 100.$ atm and $T = 120.$ $^{\circ}\text{C}$.

c) Give the temperature and pressure corresponding to the normal freezing point, the normal boiling point, and the normal sublimation point for the substance. If a normal point for any of these phase transitions does not exist, explain why the point does not exist.

The word "normal" in the question refers to normal pressure, or $p = 1.00$ atm. A normal phase transition occurs when two phases can exist simultaneously at equilibrium when $p = 1.00$ atm. Therefore

normal freezing point $p = 1.00$ atm, $T = 44.$ $^{\circ}\text{C}$

normal boiling point $p = 1.00$ atm, $T = 77.$ $^{\circ}\text{C}$

normal sublimation point There is no normal sublimation point, because the line representing $p = 1.00$ atm never crosses the solid-gas boundary.

2) Define the following terms: a) solution, b) miscible, c) hydration, d) percent by mass (solute), e) colligative property, f) hypotonic.

a) solution - A homogeneous mixture, that is, a mixture whose composition is the same throughout.

b) miscible - Refers to two liquids that will always form a solution no matter how much of each liquid is used.

c) hydration - The process of surrounding solute particles in aqueous solution with water molecules. Hydration of ions occurs by ion-dipole attractive forces, and is particularly strong.

$$\text{d) percent by mass (solute)} = \frac{\text{mass solute}}{\text{mass solution}} \times 100 \%$$

e) colligative property - A property of a solution of a nonvolatile solute and a volatile solvent which at most (for ideal conditions) depends only on the physical properties of the solvent and the concentration of solute particles.

f) hypotonic - A solution with a lower concentration of dissolved substances than blood plasma. Red blood cells will swell, and can burst open, if placed in a hypotonic solution.

3) The following question concerns mixing of liquids.

a) Consider mixing two liquids where mixing is exothermic ($\Delta H_{\text{soln}} < 0$). Would you expect a solution to form (yes/maybe/no)? Justify your answer.

Yes. Mixing increases randomness (entropy), which favors solution formation. Since the mixing is exothermic, mixing lowers the energy, which also favors solution formation. Since both factors favor solution formation, I would expect a solution to form.

b) Consider mixing two liquids where mixing is endothermic ($\Delta H_{\text{soln}} > 0$). Would you expect a solution to form (yes/maybe/no)? Justify your answer.

Maybe. Mixing increases randomness (entropy), which favors solution formation. However, the mixing is endothermic, and so mixing raises the energy, which favors not forming a solution. Since one factor favors solution formation and one factor does not, a solution might form. Whether a solution forms depends on which of the two opposing factors is more important, and so will mainly depend on how large the value is for ΔH_{soln} . The larger the value for ΔH_{soln} (the more endothermic the mixing process) the less likely a solution will form.

4) Explain the variation in the solubility in water of the listed alcohols:

<u>Compound</u>	<u>Solubility (g/100 g H₂O) at T = 20. °C</u>
CH ₃ OH	∞
CH ₃ CH ₂ OH	∞
CH ₃ CH ₂ CH ₂ OH	∞
CH ₃ CH ₂ CH ₂ CH ₂ OH	9.0
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ OH	2.7

∞ means the alcohol and water are completely miscible

The –OH group of the alcohols can form strong hydrogen bonds with water molecules, but this property diminishes in importance as the chain length increases.

The longer the C–C chain, the less significant the –OH bond, and the more the molecule "looks like" a nonpolar hydrocarbon. Hence, as the C–C chain length increases, the molecule becomes less polar. Since "like dissolves like", as the molecules become less polar, the solubility in water (a polar solvent) decreases.

5) Which of the following liquids is expected to be miscible with water – methanol (CH₃OH), acetone (CH₃COCH₃), cyclohexane (C₆H₁₂)?

Methanol and acetone are both polar molecules, and so expected to homogeneously mix with water, a polar liquid. Cyclohexane is nonpolar, and so is not expected to homogeneously mix with water.

Note that drawing the Lewis structure of methanol and acetone makes it easy to see they are polar. You may assume any molecule that is a hydrocarbon (contains only carbon and hydrogen) is nonpolar.

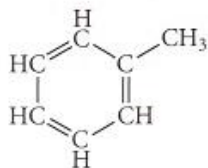
6) What does the statement like dissolves like mean with respect to solution formation?

The statement "like dissolves like" means that polar solutes tend to dissolve well in polar solvents, and that nonpolar solutes tend to dissolve well in nonpolar solvents. Polar solutes usually do not dissolve well in nonpolar solvents, and nonpolar solutes usually do not dissolve well in polar solvents.

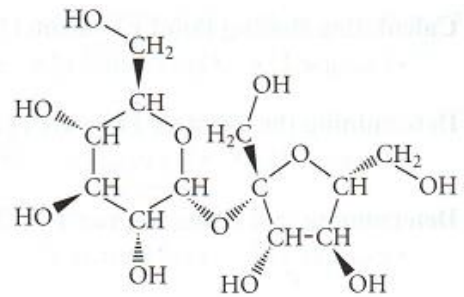
The underlying explanation for this "rule" is that $\Delta H_{\text{soln}} \cong 0$ when the solute and the solvent are both polar or when they are both nonpolar. Solution formation will usually occur in this case because of the increase in randomness (entropy) when a solution forms. When the solvent is polar and the solute is nonpolar, or when the solvent is nonpolar and the solute is polar, it is usually the case that $\Delta H_{\text{soln}} \gg 0$. The increase in enthalpy (essentially an increase in energy) required to form a solution, generally prevents solution formation from taking place.

7) For each compound would you expect greater solubility in water (H₂O) or in hexane (C₆H₁₄)? Indicate the kinds of intermolecular forces that would occur between the solute and the solvent in which the molecule is more soluble.

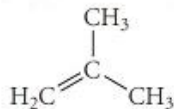
a. toluene



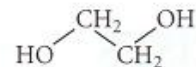
b. sucrose
(table sugar)



c. isobutene



d. ethylene glycol



Molecules a and c (toluene and isobutene) are both hydrocarbons (they contain only carbon and hydrogen atoms). The molecules are nonpolar. Therefore these two substances should dissolve well in hexane (a nonpolar solvent) but not dissolve well in water (a polar solvent). They will interact with the hexane by dispersion forces.

Molecules b and d (sucrose and ethylene glycol) contain several polar -OH bonds, and so the molecules are polar. Therefore these two molecules should dissolve well in water (a polar solvent) but not dissolve well in hexane (a nonpolar solvent). They will interact with water by hydrogen bonding (a type of dipole-dipole interaction) and dispersion forces.

8) Calculate the molality of each of the following aqueous solutions.

a) 2.55 M NaCl solution, density = $D = 1.08 \text{ g/mL}$.

Recall that molality = $\frac{\text{moles solute}}{\text{kg solvent}}$

$MW(\text{NaCl}) = 58.44 \text{ g/mol}$

Assume 1000.0 mL of solution. (Note that you will get the same answer if you assume a different volume of solution, but 1000.0 mL is a convenient amount to work with.)

$$\text{Then mass solution} = 1000.0 \text{ mL soln} \frac{1.08 \text{ g soln}}{\text{mL soln}} = 1080. \text{ g soln} = 1.080 \text{ kg soln}$$

$$\text{moles NaCl} = 1.000 \text{ L soln} \frac{2.55 \text{ mol NaCl}}{\text{L soln}} = 2.55 \text{ mol NaCl}$$

$$\text{mass of water} = \text{mass soln} - \text{mass NaCl}$$

$$= 1080. \text{ g} - [2.55 \text{ mol NaCl} \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}}] = 931. \text{ g} = 0.931 \text{ kg}$$

$$\text{And so molality} = \frac{2.55 \text{ mol NaCl}}{0.931 \text{ kg water}} = 2.74 \text{ mol/kg}$$

b) 45.2 percent by mass KBr.

Here it is convenient to assume we have 100.0 g solution. Since the solution is 45.2 % by mass KBr, 100 g of the solution contains 45.2 g KBr and 54.8 g H₂O. Since MW(KBr) = 119.0 g/mol, then

$$\text{moles KBr} = 45.2 \text{ g KBr} \frac{1 \text{ mol KBr}}{119.0 \text{ g KBr}} = 0.380 \text{ mol KBr}$$

$$\text{kg water} = 54.8 \text{ g water} \frac{1 \text{ kg}}{1000. \text{ g}} = 0.0548 \text{ kg water}$$

$$\text{And so molality} = \frac{0.380 \text{ mol KBr}}{0.0548 \text{ kg water}} = 6.93 \text{ mol/kg}$$

9) The density of an aqueous solution containing 15.0 percent ethyl alcohol (EtOH = CH₃CH₂OH) by mass is D = 0.984 g/mL.

a) Calculate the molality of this solution.

Assume 100.0 g of solution. (Note that you will get the same answer if you assume a different mass of solution, but 100.0 g is an easy amount to work with.)

The mass of ethyl alcohol in 100.0 g solution is 0.150 (100.0 g) = 15.0 g EtOH

The mass of water in 100.0 g solution is 100.0 g - 15.0 g = 85.0 g = 0.0850 kg water

The molecular mass of ethyl alcohol is MW = 46.07 g/mol, and so the moles of ethyl alcohol in 100.0 g solution is

$$\text{moles EtOH} = 15.0 \text{ g EtOH} \frac{1 \text{ mol}}{46.07 \text{ g}} = 0.3256 \text{ mol EtOH}$$

The molality of ethyl alcohol is then $m = \frac{0.3256 \text{ mol EtOH}}{0.850 \text{ kg water}} = 3.83 \text{ mol/kg}$

b) Calculate the molarity of the solution

The volume occupied by 100.0 g of solution is

$$V = 100.0 \text{ g soln} \frac{1 \text{ mL}}{0.984 \text{ g}} = 102. \text{ mL} = 0.102 \text{ L}$$

The moles of EtOH is 0.3256 mol (from part a). Therefore

$$\text{molarity} = \frac{0.3256 \text{ mol EtOH}}{0.102 \text{ L soln}} = 3.19 \text{ mol/L}$$

c) What volume of solution would contain 0.250 moles of ethyl alcohol?

$$V = 0.250 \text{ mol EtOH} \frac{1 \text{ L}}{3.19 \text{ mol}} = 0.0784 \text{ L} = 78.4 \text{ mL soln}$$

10) A solution is 14.8% by mass hexane (C_6H_{14} , MW = 86.2 g/mol) in carbon tetrachloride (CCl_4 , MW = 153.8 g/mol). What are the molality and the mole fraction of hexane in the solution?

Assume 100.0 g of solution. Then there are 14.8 g hexane (H) and 85.2 g carbon tetrachloride (C).

$$\text{moles hexane} = 14.8 \text{ g} \frac{1 \text{ mol H}}{86.2 \text{ g H}} = 0.1717 \text{ mol H}$$

$$\text{moles carbon tetrachloride} = 85.2 \text{ g} \frac{1 \text{ mol C}}{153.8 \text{ g C}} = 0.5540 \text{ mol C}$$

$$m_{\text{H}} = \frac{\text{mol H}}{\text{kg C}} = \frac{0.1717 \text{ mol H}}{0.0852 \text{ kg C}} = 2.02 \text{ mol/kg}$$

$$X_{\text{H}} = \frac{\text{mol H}}{\text{mol H} + \text{mol C}} = \frac{0.1717 \text{ mol}}{0.1717 \text{ mol} + 0.5540 \text{ mol}} = 0.237$$

11) An aqueous KNO_3 solution is made using 72.5 g of KNO_3 dissolved to a total solution volume of 2.00 L. Calculate the molarity, molality, and mass percent KNO_3 in the solution. Assume a density of 1.050 g/mL for the solution.

$$\text{MW}(\text{KNO}_3) = 101.1 \text{ g/mol} \quad \text{MW}(\text{H}_2\text{O}) = 18.02 \text{ g/mol}$$

$$g(\text{KNO}_3) = 72.5 \text{ g} \quad \text{mol}(\text{KNO}_3) = 72.5 \text{ g} \frac{1 \text{ mol}}{101.1 \text{ g}} = 0.7171 \text{ mol}$$

$$g \text{ solution} = 2000 \text{ mL} \frac{1.05 \text{ g}}{1 \text{ mL}} = 2100. \text{ g}$$

Since $g \text{ solution} = g \text{ water} + g \text{ KNO}_3$,

$$g \text{ water} = g \text{ solution} - g \text{ KNO}_3 = 2100. \text{ g} - 72.5 \text{ g} = 2027.5 \text{ g}$$

So

$$\text{molarity (M)} = \frac{0.7171 \text{ mol KNO}_3}{2.00 \text{ L soln}} = 0.359 \text{ mol/L}$$

$$\text{molality (m)} = \frac{0.7171 \text{ mol KNO}_3}{2.0275 \text{ kg water}} = 0.354 \text{ mol/kg}$$

$$\text{mass percent KNO}_3 = \frac{72.5 \text{ g KNO}_3}{2100. \text{ g soln}} \cdot 100\% = 3.45 \% \text{ KNO}_3 \text{ by mass}$$

12) A stock solution is prepared by dissolving 18.21 g of sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$, MW = 82.03 g/mol) in water. The final volume of the solution is $V = 250.0 \text{ mL}$.

a) What is the molarity of sodium acetate in the solution?

$$\text{moles sodium acetate} = 18.21 \text{ g} \frac{1 \text{ mol}}{82.03 \text{ g}} = 0.2220 \text{ mol sodium acetate}$$

$$\text{molarity} = \frac{0.2220 \text{ mol}}{0.2500 \text{ L}} = 0.8880 \text{ mol/L}$$

b) How many mL of the stock solution of sodium acetate are needed to prepare 100.0 mL of a 0.1000 mol/L sodium acetate solution?

$$M_i V_i = M_f V_f$$

$$\text{and so } V_i = \frac{M_f}{M_i} V_f = \frac{(0.1000 \text{ M})}{(0.8880 \text{ M})} (100.0 \text{ mL}) = 11.36 \text{ mL stock solution}$$

13) The Henry's law constant for methyl bromide (CH_3Br , MW = 94.9 g/mol) in water is 0.159 mol/L·atm at $T = 25. \text{ }^\circ\text{C}$. How many grams of methyl bromide will dissolve in 1.000 L of water at this temperature when the pressure of methyl bromide above the water is $p = 320. \text{ torr}$?

Let MeBr = methyl bromide

From Henry's law $[MBr] = k p_{MBr}$

$$p_{MeBr} = 320. \text{ torr} \frac{1 \text{ atm}}{760 \text{ torr}} = 0.421 \text{ atm}$$

$$[MeBr] = (0.159 \text{ mol/L} \cdot \text{atm}) (0.421 \text{ atm}) = 0.0669 \text{ mol/L}$$

So 1.000 L of water contains 0.0669 moles of methyl bromide.

$$\text{grams MeBr} = (0.0669 \text{ mol}) (94.9 \text{ g/mol}) = 6.36 \text{ g}$$

14) The following question concerns ideal solutions.

a) How do we define an ideal solution of two gases?

An ideal solution of two gases is a homogeneous mixture of the gases where each partial pressure is given by the ideal gas law

$$p_A = \frac{n_A RT}{V} \quad p_B = \frac{n_B RT}{V}$$

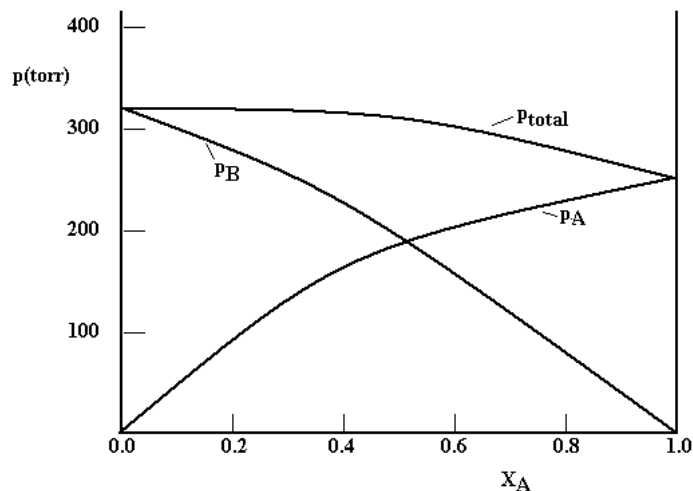
b) How do we define an ideal solution of two volatile liquids?

An ideal solution of two volatile liquids is a homogeneous mixture of liquids where the equilibrium partial pressure above the solution is given by Raoult's law

$$p_A = X_A p_A^\circ \quad p_B = X_B p_B^\circ$$

where p_A° and p_B° are the vapor pressures for the pure liquids.

15) The diagram below is for a solution of two volatile liquids A and B, and shows the partial pressure of A, the partial pressure of B, and the total pressure above a solution as a function of X_A , the mole fraction of A in the solution.



a) Do A and B form an ideal solution? Justify your answer.

No, A and B do not form an ideal solution. For an ideal solution the partial pressures of A and B are given by Raoult's law, and so a plot of partial pressure or total pressure vs X_A should give straight lines. Since that is not observed here, A and B do not obey Raoult's law and so do not form an ideal solution.

b) If A and B do not form an ideal solution, are the A--B interactions stronger than, approximately equal to, or weaker than the average of the A---A and B---B interactions? Justify your answer.

The partial pressures of A and B are higher than we would get if Raoult's law were obeyed. That means that the molecules of A and B are less strongly held in the liquid than if Raoult's law were obeyed. Therefore, the A---B interactions are weaker than the average of the A---A and B---B interactions.

c) What is p_B° , the vapor pressure of pure B, at the temperature corresponding to the diagram above?

Pure B corresponds to $X_A = 0$. For this value of X_A , the pressure, which corresponds to the vapor pressure of pure B, is $p_B^\circ = 315$ torr.

16) A solution contains a mixture of pentane and hexane at room temperature. The solution has a vapor pressure of 258 torr. Pure pentane and hexane have vapor pressures of 425 torr and 151 torr, respectively, at room temperature. What is the mole fraction composition of the mixture? Assume ideal behavior.

Let P = pentane and H = hexane, and assume Raoult's law applies.

$$p_{\text{total}} = p_P + p_H = X_P p_P^\circ + X_H p_H^\circ$$

We have one equation and two unknowns. But $X_P + X_H = 1$, so $X_P = 1 - X_H$, so

$$\begin{aligned} p_{\text{total}} &= (1 - X_H) p_P^\circ + X_H p_H^\circ = p_P^\circ - X_H p_P^\circ + X_H p_H^\circ \\ &= p_P^\circ + X_H (p_H^\circ - p_P^\circ) \end{aligned}$$

$$\text{Or, finally, } X_H = \frac{(p_{\text{total}} - p_P^\circ)}{(p_H^\circ - p_P^\circ)} = \frac{(258 - 425) \text{ torr}}{(151 - 425) \text{ torr}} = 0.609$$

And so $X_P = 1 - 0.609 = 0.391$

17) How many grams of glucose ($C_6H_{12}O_6$, MW = 180.2 g/mol) must be added to 552. g of water (H_2O , MW = 18.02 g/mol) to give a solution with a vapor pressure 0.20 torr less than that of pure water at 20.0 °C? The vapor pressure of pure water at this temperature is $p^\circ_{H_2O} = 17.5$ torr.

For vapor pressure lowering $\Delta p = X_B p^\circ$

$$\text{So } X_{\text{glu}} = \frac{\Delta p}{p^\circ} = \frac{0.20 \text{ torr}}{17.5 \text{ torr}} = 0.0114$$

$$\text{And so } X_{H_2O} = 1.000 - 0.0114 = 0.9886$$

The number of moles of H_2O is

$$n(H_2O) = \frac{552. \text{ g}}{(18.02 \text{ g/mol})} = 30.63 \text{ mol } H_2O$$

$$\text{Since } \frac{n(\text{glucose})}{n(H_2O)} = \frac{X_{\text{glu}}}{X_{H_2O}}$$

$$n(\text{glucose}) = n(H_2O) \frac{X_{\text{glu}}}{X_{H_2O}} = (30.63 \text{ mol}) \frac{0.0114}{0.9886} = 0.353 \text{ mol glucose}$$

And so

$$g(\text{glucose}) = 0.353 \text{ mol } \frac{180.2 \text{ g}}{1 \text{ mol}} = 63.6 \text{ g glucose}$$

18) An ethylene glycol solution contains 21.2 g of ethylene glycol ($C_2H_6O_2$, MW = 62.07 g/mol) in 85.4 mL of water, (H_2O , MW = 18.02 g/mol). Determine the freezing point and boiling point of the solution. Assume a density of 1.000 g/mL for water. Note that for water $K_f = 1.86$ °C•kg/mol and $K_b = 0.512$ °C•kg/mol, and that ethylene glycol is a nonvolatile and nonionizing solute.

To use the equations for boiling point elevation and freezing point depression we need to find the molality of ethylene glycol in the solution. Let ET = ethylene glycol.

So the moles of ethylene glycol is

$$n_{\text{ET}} = \frac{21.2 \text{ g}}{62.07 \text{ g/mol}} = 0.3415 \text{ mol ET}$$

There are 85.4 mL of water, and since the density of water is 1.00 g/mL, there are 85.4 g (or 0.0854 kg) of water. So the molality of the solution is

$$m = \frac{0.3415 \text{ mol ET}}{0.0854 \text{ kg water}} = 4.00 \text{ mol/kg}$$

The formulas for boiling point elevation and freezing point depression are

$$\Delta T_b = K_b m \quad K_b = 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

$$\Delta T_f = K_f m \quad K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

So $\Delta T_b = K_b m = (0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}) (4.00 \text{ mol/kg}) = 2.05 \text{ }^\circ\text{C}$

$$\Delta T_f = K_f m = (1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}) (4.00 \text{ mol/kg}) = 7.44 \text{ }^\circ\text{C}$$

Since the normal boiling point and normal freezing point for water are 100.00 °C and 0.00 °C, respectively, the boiling point and freezing point for the solution are

$$T_b = 100.00 \text{ }^\circ\text{C} + 2.05 \text{ }^\circ\text{C} = 102.05 \text{ }^\circ\text{C}$$

$$T_f = 0.00 \text{ }^\circ\text{C} - 7.44 \text{ }^\circ\text{C} = -7.44 \text{ }^\circ\text{C}$$

Note that the boiling point of the solution is higher than that for pure water and the freezing point of the solution is lower than that for pure water.

19) An aqueous solution containing 35.9 g of an unknown nonvolatile and nonionizing molecular compound in 150.0 g of water was found to have a freezing point of - 1.30 °C. Calculate the molar mass of the unknown compound. For water, $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$.

For freezing point depression $\Delta T_f = K_f m$, where m = molality of solute particles

$$\text{So } m = \frac{\Delta T_f}{K_f} \quad \Delta T_f = T_f^\circ - T_f = 0.0^\circ\text{C} - (-1.3 \text{ }^\circ\text{C}) = 1.3 \text{ }^\circ\text{C}$$

$$\text{So } m = \frac{1.3 \text{ }^\circ\text{C}}{(1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol})} = 0.699 \text{ mol/kg}$$

There are 150.0 g of solvent (water). The number of moles of solute is

$$n_{\text{solute}} = 0.1500 \text{ kg solvent} \frac{0.699 \text{ mol}}{1 \text{ kg}} = 0.105 \text{ mol solute}$$

$$\text{So MW} = \frac{\text{g solute}}{\text{mol solute}} = \frac{35.9 \text{ g}}{0.105 \text{ mol}} = 342. \text{ g/mol}$$

20) The density of a 0.900 % by mass aqueous solution of sodium chloride (NaCl, MW = 58.44 g/mol) is $D = 1.0046 \text{ g/mL}$ at $T = 25.0 \text{ }^\circ\text{C}$. What is the osmotic pressure of the above solution (in atm), relative to pure water, at this temperature?

Assume 1.000 L of solution. Then the mass of the solution is 1004.6 g. Since the solution is 0.900 % by mass NaCl, the mass of NaCl is

$$\text{mass} = (0.00900) (1004.6 \text{ g}) = 9.04 \text{ g NaCl}$$

$$\text{The moles of NaCl is } n = 9.04 \text{ g NaCl} \frac{1 \text{ mol}}{58.44 \text{ g}} = 0.1547 \text{ moles NaCl}$$

NaCl ionizes in water by the process $\text{NaCl(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

The moles of particles is

$$n = (0.1547 \text{ moles NaCl}) \frac{2 \text{ moles particles}}{1 \text{ mole NaCl}} = 0.3094 \text{ moles particles}$$

$$\text{So } [B] = \frac{0.3094 \text{ moles particles}}{1 \text{ L solution}} = 0.3094 \text{ mol/L}$$

Since $\Pi = [B]RT$

$$\Pi = (0.3094 \text{ mol/L}) (0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K}) (298.2 \text{ K}) = 7.57 \text{ atm}$$

Note that since the colligative properties are in terms of particle concentration, ionization of an ionic compound in water must be accounted for.

21) What is the osmotic pressure, in atmospheres, of a 1.57 M aqueous solution of urea ($(\text{NH}_2)_2\text{CO}$, MW = 60.06 g/mol), a nonvolatile and nonionizing solute, at $T = 27. \text{ }^\circ\text{C}$.

$$\Pi = MRT = (1.57 \text{ mol/L})(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(27. + 273.) \text{ K} = 38.7 \text{ atm}$$

22) An isotonic saline solution is an aqueous solution that has approximately the same value for osmotic pressure as human body cells. Isotonic saline contains 9.0 g of sodium chloride (NaCl, MW = 58.44 g/mol) per liter of solution.

a) What is the molarity of sodium chloride in an isotonic saline solution?

$$M = \frac{\text{mol NaCl}}{\text{L soln}} \quad \text{mol NaCl} = 9.00 \text{ g} \frac{1 \text{ mol}}{58.44 \text{ g}} = 0.1540 \text{ mol}$$

$$\text{So } M = \frac{0.1540 \text{ mol NaCl}}{1 \text{ L soln}} = 0.1540 \text{ mol/L NaCl}$$

b) What is the osmotic pressure (relative to pure water) of an isotonic saline solution? Assume $T = 37.^\circ\text{C}$. Give your final answer in units of atm.

We may use our answer in a to find the molarity of solute particles

$$M(\text{particles}) = \frac{0.1540 \text{ mol NaCl}}{\text{L soln}} \frac{2 \text{ mol particles}}{1 \text{ mol NaCl}} = 0.3080 \text{ mol/L particles}$$

$$\Pi = [B]RT = (0.3080 \text{ mol/L})(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(310. \text{ K}) = 7.84 \text{ atm}$$

c) Based on your answer to b, explain why red blood cells rupture when placed in deionized water.

The osmotic pressure for the fluid inside a red blood cell, relative to that for pure water, should be about the same as for an isotonic saline solution, and so will be about 7.8 atm. This is such a large value for pressure that the membrane of the red blood cell will rupture.

23) Arrange the following aqueous solutions in order of decreasing freezing point, and explain your reasoning: 0.50 m HCl, 0.50 m glucose, 0.50 m CH_3COOH .

The equation for freezing point depression, including the van't Hoff factor (i), is

$$\Delta T_f = i K_f m$$

The freezing point will be depressed most by the solution that contains the most solute particles. Since all of the solutions have the same concentration of solute, the freezing point will be depressed most by the solute with the largest value for i , the van't Hoff factor.

We can classify each solute as a strong electrolyte, a weak electrolyte, or a nonelectrolyte. HCl is a strong electrolyte, so under ideal conditions it will completely dissociate into two particles per molecule ($i = 2$). The concentration of particles will be 1.00 m. Acetic acid is a weak electrolyte, so it will only dissociate to a small extent ($1 < i < 2$). The concentration of particles will be greater than 0.50 m but less than 1.00 m. Glucose is a nonelectrolyte ($i = 1$), so glucose molecules remain as glucose molecules in solution. The concentration of particles will be 0.50 m. For these solutions, the order of freezing points is:

$$0.50 \text{ m glucose} > 0.50 \text{ m acetic acid} > 0.50 \text{ m HCl}$$

The HCl solution will have the lowest freezing point (greatest freezing point depression), and the glucose solution will have the highest freezing point (smallest freezing point depression).

24) A quantity of 7.480 g of an organic compound is dissolved in water to make 300.0 mL of solution. The solution has an osmotic pressure of 1.43 atm at $T = 27. \text{ }^\circ\text{C}$. The analysis of this compound shows that it contains 41.8 percent C, 4.7 percent H, 37.3 percent O, and 16.3 percent N. Find the molecular formula for the compound.

Find the concentration of the solution, then work out the molar mass. Note that $27. \text{ }^\circ\text{C} = 300. \text{ K}$. The concentration is:

$$M = \frac{\Pi}{RT} = \frac{(1.43 \text{ atm})}{(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K}) (300. \text{ K})} = 0.0581 \text{ mol/L}$$

The solution volume is 0.3000 L so the number of moles of solute is:

$$n = 0.3000 \text{ L} \frac{0.0581 \text{ mol}}{\text{L}} = 0.0174 \text{ mol}$$

The molar mass is then

$$\text{MW} = \frac{7.480 \text{ g}}{0.0174 \text{ mol}} = 430. \text{ g/mol}$$

The empirical formula can be found most easily by assuming a 100.0 g sample of the substance.

$$\text{moles C} = 41.8 \text{ g} \frac{1 \text{ mol}}{12.01 \text{ g}} = 3.48 \text{ mol C} / 1.16 = 3.00$$

$$\text{moles H} = 4.7 \text{ g} \frac{1 \text{ mol}}{1.008 \text{ g}} = 4.7 \text{ mol H} / 1.16 = 4.05$$

$$\text{moles O} = 37.3 \text{ g} \frac{1 \text{ mol}}{16.00 \text{ g}} = 2.33 \text{ mol O} / 1.16 = 2.01$$

$$\text{moles N} = 16.3 \text{ g} \frac{1 \text{ mol}}{14.01 \text{ g}} = 1.16 \text{ mol N} / 1.16 = 1.00$$

This gives the empirical formula, $\text{C}_3\text{H}_4\text{O}_2\text{N}$, which has a mass of 86.0 g per formula unit. The molar mass is five times this amount ($430 \div 86.0 = 5.0$), so the molecular formula is $(\text{C}_3\text{H}_4\text{O}_2\text{N})_5$ or $\text{C}_{15}\text{H}_{20}\text{O}_{10}\text{N}_5$.

25) A solution is prepared by dissolving 2.18 g of potassium nitrate (KNO_3 , $\text{MW} = 101.1 \text{ g/mol}$) in water. The final volume of the solution is $V = 400.0 \text{ mL}$. What is the osmotic pressure of the solution at $T = 20.0 \text{ }^\circ\text{C}$?

The number of moles of KNO_3 is

$$n = 2.18 \text{ g} \frac{1 \text{ mol}}{101.1 \text{ g}} = 0.0216 \text{ mol}$$

and so the molarity of the solution is $M = \frac{0.0216 \text{ mol KNO}_3}{0.4000 \text{ L}} = 0.0539 \text{ mol/L KNO}_3$

However, KNO_3 is an ionic compound $\text{KNO}_3(\text{s}) \rightarrow \text{K}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$

and so the concentration of solute particles is

$$\frac{0.0539 \text{ mol KNO}_3}{\text{L soln}} \frac{2 \text{ mol particle}}{1 \text{ mol KNO}_3} = 0.1078 \text{ mol particles}$$

The osmotic pressure is therefore

$$\Pi = [B]RT = (0.1078 \text{ mol/L})(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(293.1 \text{ K}) = 2.59 \text{ atm}$$

26) Consider 0.0100 mol/kg aqueous solutions of iron III chloride (FeCl_3), calcium nitrate ($\text{Ca}(\text{NO}_3)_2$), and copper II sulfate (CuSO_4). Which of these solutions will have the highest freezing point, and which of these solutions will have the highest melting point? Justify your answer.

The formulas for boiling point elevation and freezing point depression

$$\Delta T_b = K_b m_B$$

$$\Delta T_f = K_f m_B$$

The solutions all have the same molality. Since boiling point increases, the compound forming the largest number of particles per formula unit will have the highest boiling point. Since freezing point decreases (and recall that freezing point and melting point occur at the same temperature, and are just inverse processes), the compound forming the smallest number of particles per formula unit will have the highest melting point (think about why this is true).

FeCl_3 – 4 particles per formula unit

$\text{Ca}(\text{NO}_3)_2$ – 3 particles per formula unit

CuSO_4 – 2 particles per formula unit

The FeCl_3 solution will have the highest concentration of solute particles, and therefore the highest boiling point. The CuSO_4 solution will have the lowest concentration of solute particles, and therefore the highest melting point.

27) What is a colloid? How are colloids similar to solutions, and how do they differ from solutions?

A colloid is a homogeneous mixture of particles within a solvent, where each particle is composed of a number of individual molecules. Colloids are similar to solutions in that both are homogeneous (the same composition throughout). Colloids are different from solutions in that different particles making up a colloid will contain different numbers of molecules, and so will not all be exactly the same.

For example, a cloud is a colloid consisting of water droplets suspended in air. The water droplets contain a large number of water molecules, and different droplets contain a different number of molecules of water.