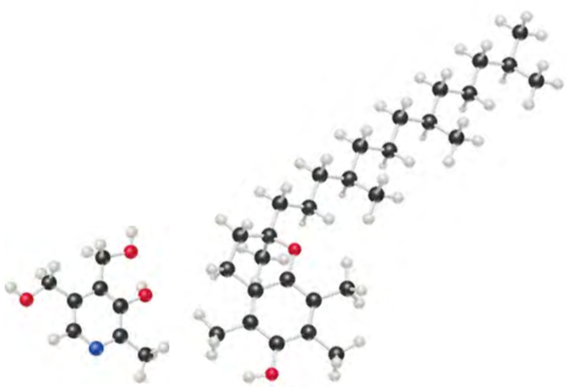
Properties of Solutions
Problem Set

13.3 How does the lattice energy of an ionic solid affect its solubility in water? [Section 13.1]

13.5 Which of the following is the best representation of a saturated solution? Explain your reasoning. [Section 13.2]

13.7 The structures of vitamins E and B_6 are shown below. Predict which is largely water soluble and which is largely fat soluble. Explain. [Section 13.3]



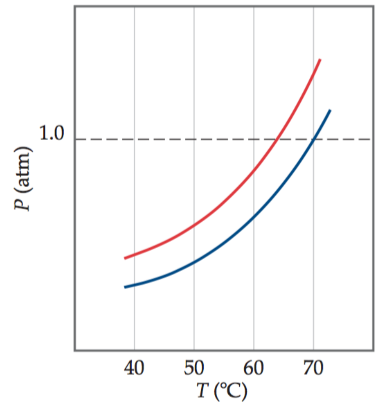
13.8 If you wanted to prepare a solution of CO in water at 25 °C in which the CO concentration was 2.5 m*M*, what pressure of CO would you need to use? (See Figure 13.19.) [Section 13.3]

13.9 The figure shows two identical volumetric flasks containing the same solution at two temperatures.

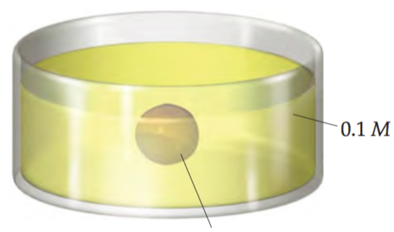
(a) Does the molarity of the solution change with the change in temperature? Explain.

(b) Does the molality of the solution change with the change in temperature? Explain. [Section 13.4]

13.10 The following diagram shows the vapor-pressure curves of a volatile solvent and a solution of that solvent containing a nonvolatile solute. **(a)** Which line represents the solution? **(b)** What are the normal boiling points of the solvent and the solution? [Section 13.5]



13.11 Suppose you had a balloon made of some highly flexible semipermeable membrane. The balloon is filled completely with a 0.2 *M* solution of some solute and is submerged in a 0.1 *M* solution of the same solute:



Initially, the volume of solution in the balloon is 0.25 L. Assuming the volume outside the semipermeable membrane is large, as the illustration shows, what would you expect for the solution volume inside the balloon once the system has come to equilibrium through osmosis? [Section 13.5]

13.13 In general, the attractive intermolecular forces between solvent and solute particles must be comparable or greater than solute–solute interactions for significant solubility to occur. Explain this statement in terms of the overall energetics of solution formation.

13.14 **(a)** Considering the energetics of solute–solute, solvent–solvent, and solute–solvent interactions, explain why NaCl dissolves in water but not in benzene (C_6H_6). **(b)** What factors cause a cation to be strongly hydrated?

13.15 Indicate the type of solute–solvent interaction (Section 11.2) that should be most important in each of the following solutions: **(a)** CCl_4 in benzene (C_6H_6), **(b)** methanol (CH_3OH) in water, **(c)** KBr in water, **(d)** HCl in acetonitrile (CH_3CN).

13.18 When ammonium chloride dissolves in water, the solution becomes colder. **(a)** Is the solution process exothermic or endothermic? **(b)** Why does the solution form?

13.24 The solubility of $MnSO_4 \cdot H_2O$ in water at 20 °C is 70 g per 100 mL of water. **(a)** Is a 1.22 *M* solution of $MnSO_4 \cdot H_2O$ in water at 20 °C saturated, supersaturated, or unsaturated? **(b)** Given a solution of $MnSO_4 \cdot H_2O$ of unknown concentration, what experiment could you perform to determine whether the new solution is saturated, supersaturated, or unsaturated?

13.27 Water and glycerol, $CH_2(OH)CH(OH)CH_2OH$, are miscible in all proportions. What does this mean? How do the OH groups of the alcohol molecule contribute to this miscibility?

13.28 Oil and water are immiscible. What does this mean? Explain in terms of the structural features of their respective molecules and the forces between them.

13.29 Common laboratory solvents include acetone (CH_3COCH_3), methanol (CH_3OH), toluene ($C_6H_5CH_3$), and water. Which of these is the best solvent for nonpolar solutes? Explain.

13.33 Which of the following in each pair is likely to be more soluble in hexane, C_6H_{14}: **(a)** CCl_4 or $CaCl_2$; **(b)** benzene (C_6H_6) or glycerol, $CH_2(OH)CH(OH)CH_2OH$; **(c)** octanoic acid, $CH_3CH_2CH_2CH_2CH_2CH_2CH_2COOH$, or acetic acid, CH_3COOH? Explain your answer in each case.

13.35 **(a)** Explain why carbonated beverages must be stored in sealed containers. **(b)** Once the beverage has been opened, why does it maintain more carbonation when refrigerated than at room temperature?

13.36 Explain why pressure substantially affects the solubility of O_2 in water but has little effect on the solubility of NaCl in water.

13.37 The Henry's law constant for helium gas in water at 30 °C is 3.7×10^{-4} *M*/atm and the constant for N_2 at 30 °C is 6.0×10^{-4} *M*/atm. If the two gases are each present at 1.5 atm pressure, calculate the solubility of each gas.

13.40 **(a)** What is the mass percentage of iodine (I_2) in a solution containing 0.035 mol I_2 in 125 g of CCl_4? **(b)** Seawater contains 0.0079 g Sr^{2+} per kilogram of water. What is the concentration of Sr^{2+} measured in ppm?

13.43 Calculate the molarity of the following aqueous solutions: **(a)** 0.540 g $Mg(NO_3)_2$ in 250.0 mL of solution, **(b)** 22.4 g $LiClO_4 \cdot 3\,H_2O$ in 125 mL of solution, **(c)** 25.0 mL of 3.50 *M* HNO_3 diluted to 0.250 L.

13.46 **(a)** What is the molality of a solution formed by dissolving 1.12 mol of KCl in 16.0 mol of water? **(b)** How many grams of sulfur (S_8) must be dissolved in 100.0 g naphthalene ($C_{10}H_8$) to make a 0.12 *m* solution?

13.47 A sulfuric acid solution containing 571.6 g of H_2SO_4 per liter of solution has a density of 1.329 g/cm^3. Calculate **(a)** the mass percentage, **(b)** the mole fraction, **(c)** the molality, **(d)** the molarity of H_2SO_4 in this solution.

13.49 The density of acetonitrile (CH_3CN) is 0.786 g/mL and the density of methanol (CH_3OH) is 0.791 g/mL. A solution is made by dissolving 22.5 mL CH_3OH in 98.7 mL CH_3CN. **(a)** What is the mole fraction of methanol in the solution? **(b)** What is the molality of the solution? **(c)** Assuming that the volumes are additive, what is the molarity of CH_3OH in the solution?

13.51 Calculate the number of moles of solute present in each of the following aqueous solutions: **(a)** 600 mL of 0.250 *M* $SrBr_2$, **(b)** 86.4 g of 0.180 *m* KCl, **(c)** 124.0 g of a solution that is 6.45% glucose ($C_6H_{12}O_6$) by mass.

13.54 Describe how you would prepare each of the following aqueous solutions: **(a)** 1.50 L of 0.110 *M* $(NH_4)_2SO_4$ solution, starting with solid $(NH_4)_2SO_4$; **(b)** 225 g of a solution that is 0.65 *m* in Na_2CO_3, starting with the solid solute; **(c)** 1.20 L of a solution that is 15.0% $Pb(NO_3)_2$ by mass (the density of the solution is 1.16 g/mL), starting with solid solute; **(d)** a 0.50 *M* solution of HCl that would just neutralize 5.5 g of $Ba(OH)_2$ starting with 6.0 *M* HCl.

13.56 Commercial concentrated aqueous ammonia is 28% NH_3 by mass and has a density of 0.90 g/mL. What is the molarity of this solution?

13.58 Caffeine ($C_8H_{10}N_4O_2$) is a stimulant found in coffee and tea. If a solution of caffeine in chloroform ($CHCl_3$) as a solvent has a concentration of 0.0500 *m*, calculate **(a)** the percent caffeine by mass, **(b)** the mole fraction of caffeine.

13.61 List four properties of a solution that depend on the total concentration but not the type of particle or particles present as solute. Write the mathematical expression that describes how each of these properties depends on concentration.

13.62 How does increasing the concentration of a nonvolatile solute in water affect the following properties: **(a)** vapor pressure, **(b)** freezing point, **(c)** boiling point; **(d)** osmotic pressure?

13.63 Consider two solutions, one formed by adding 10 g of glucose ($C_6H_{12}O_6$) to 1 L of water and the other formed by adding 10 g of sucrose ($C_{12}H_{22}O_{11}$) to 1 L of water. Are the vapor pressures over the two solutions the same? Why or why not?

13.64 **(a)** What is an *ideal solution*? **(b)** The vapor pressure of pure water at 60 °C is 149 torr. The vapor pressure of water over a solution at 60 °C containing equal numbers of moles of water and ethylene glycol (a nonvolatile solute) is 67 torr. Is the solution ideal according to Raoult's law? Explain.

[13.67] At 63.5 °C the vapor pressure of H_2O is 175 torr, and that of ethanol (C_2H_5OH) is 400 torr. A solution is made by mixing equal masses of H_2O and C_2H_5OH. **(a)** What is the mole fraction of ethanol in the solution? **(b)** Assuming ideal-solution behavior, what is the vapor pressure of the solution at 63.5 °C? **(c)** What is the mole fraction of ethanol in the vapor above the solution?

13.69 **(a)** Why does a 0.10 *m* aqueous solution of NaCl have a higher boiling point than a 0.10 *m* aqueous solution of $C_6H_{12}O_6$? **(b)** Calculate the boiling point of each solution. **(c)** The experimental boiling point of the NaCl solution is lower than that calculated, assuming that NaCl is completely dissociated in solution. Why is this the case?

13.71 List the following aqueous solutions in order of increasing boiling point: 0.120 *m* glucose, 0.050 *m* LiBr, 0.050 *m* $Zn(NO_3)_2$.

13.75 How many grams of ethylene glycol ($C_2H_6O_2$) must be added to 1.00 kg of water to produce a solution that freezes at −5.00 °C?

13.76 What is the freezing point of an aqueous solution that boils at 105.0 °C?

13.78 Seawater contains 3.4 g of salts for every liter of solution. Assuming that the solute consists entirely of NaCl (over 90% is), calculate the osmotic pressure of seawater at 20 °C.

13.82 A dilute aqueous solution of an organic compound soluble in water is formed by dissolving 2.35 g of the compound in water to form 0.250 L of solution. The resulting solution has an osmotic pressure of 0.605 atm at 25 °C. Assuming that the organic compound is a nonelectrolyte, what is its molar mass?

[13.83] The osmotic pressure of a 0.010 *M* aqueous solution of $CaCl_2$ is found to be 0.674 atm at 25 °C. **(a)** Calculate the van't Hoff factor, *i*, for the solution. **(b)** How would you expect the value of *i* to change as the solution becomes more concentrated? Explain.