Pearson

Exercises

Review Questions

1. What is a solution? What are the solute and solvent?

2. What does it mean when we say that a substance is soluble in another substance? Which units do we use to report solubility?

3. Why do two ideal gases thoroughly mix when combined? What drives the mixing?

4. What is entropy? What role does entropy play in the formation of solutions?

5. What kinds of intermolecular forces are involved in solution formation?

6. Explain how the relative strengths of solute–solute interactions, solvent–solvent interactions, and solvent–solute interactions affect solution formation.

7. What does the statement *like dissolves like* mean with respect to solution formation?

8. List the three steps involved in evaluating the enthalpy changes associated with solution formation.

9. What is the heat of hydration $(\Delta H_{\text{hydration}})$? How does the enthalpy of solution depend on the relative magnitudes of ΔH_{solute} and $\Delta H_{\text{hydration}}$?

10. Explain dynamic equilibrium with respect to solution formation. What is a saturated solution? An unsaturated solution? A supersaturated solution?

11. How does temperature affect the solubility of a solid in a liquid? How is this temperature dependence exploited to purify solids through recrystallization?

12. How does temperature affect the solubility of a gas in a liquid? How does this temperature dependence affect the amount of oxygen available for fish and other aquatic animals?

13. How does pressure affect the solubility of a gas in a liquid? How does this pressure dependence account for the bubbling that occurs upon opening a can of soda?

14. What is Henry's law? For what kinds of calculations is Henry's law useful?

15. What are the common units for expressing solution concentration?

16. How are parts by mass and parts by volume used in calculations?

17. What is the effect of a nonvolatile solute on the vapor pressure of a liquid? Why is the vapor pressure of a solution different from the vapor pressure of the pure liquid solvent?

18. What is Raoult's law? For what kind of calculations is Raoult's law useful?

19. Explain the difference between an ideal and a nonideal solution.

20. What is the effect on vapor pressure of a solution with particularly *strong* solute–solvent interactions? With particularly *weak* solute–solvent interactions?

21. Explain why the lower vapor pressure for a solution containing a nonvolatile solute results in a higher boiling point and lower melting point compared to the pure solvent.

22. What are colligative properties?

23. What is osmosis? What is osmotic pressure?

24. Explain the significance of the van't Hoff factor (i) and its role in determining the colligative properties of solutions containing ionic solutes.

Problems by Topic

Problems by Topic

Note: Answers to all odd-numbered Problems can be found in Appendix III. Exercises in the Problems by Topic section are paired, with each odd-numbered problem followed by a similar even-numbered problem. Exercises in the Cumulative Problems section are also paired but more loosely. Because of their nature, Challenge Problems and Conceptual Problems are unpaired.

Solubility

25. Pick an appropriate solvent from Table 13.3 to dissolve each substance. State the kind of intermolecular forces that would occur between the solute and solvent in each case.

- **a.** motor oil (nonpolar)
- **b.** ethanol (polar, contains an OH group)
- **c.** lard (nonpolar)
- **d.** potassium chloride (ionic)

26. Pick an appropriate solvent from Table 13.3 to dissolve each substance. State the kind of intermolecular forces that would occur between the solute and solvent in each case.

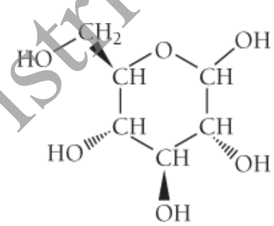
- **a.** isopropyl alcohol (polar, contains an OH group)
- **b.** sodium chloride (ionic)
- **c.** vegetable oil (nonpolar)
- **d.** sodium nitrate (ionic)

27. Which molecule would you expect to be more soluble in water, $CH_3CH_2CH_2CH_2OH$ or $HOCH_2CH_2CH_2OH$?

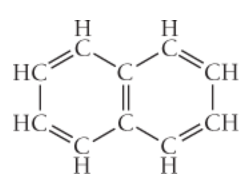
28. Which molecule would you expect to be more soluble in water, CCl_4 or CH_2Cl_2?

29. For each compound, would you expect greater solubility in water or in hexane? Indicate the kinds of intermolecular forces that occur between the solute and the solvent in which the molecule is most soluble.

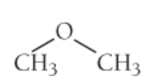
a. glucose



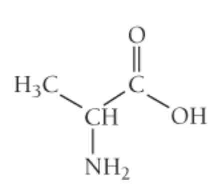
b. naphthalene



c. dimethyl ether

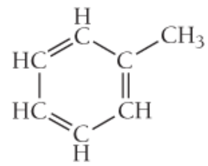


d. alanine (an amino acid)

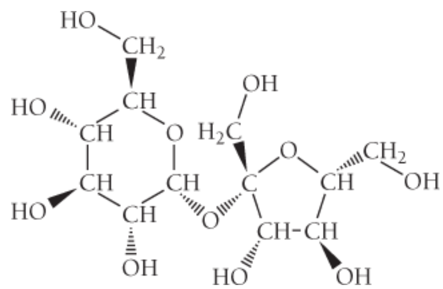


30. For each compound, would you expect greater solubility in water or in hexane? Indicate the kinds of intermolecular forces that would occur between the solute and the solvent in which the molecule is most soluble.

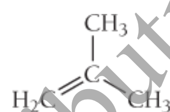
a. toluene



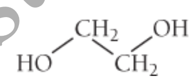
b. sucrose (table sugar)



c. isobutene



d. ethylene glycol



Energetics of Solution Formation

31. When ammonium chloride (NH_4Cl) is dissolved in water, the solution becomes colder.

a. Is the dissolution of ammonium chloride endothermic or exothermic?

b. What can you conclude about the relative magnitudes of the lattice energy of ammonium chloride and its heat of hydration?

c. Sketch a qualitative energy diagram similar to Figure 13.6 for the dissolution of NH_4Cl

d. Why does the solution form? What drives the process?

32. When lithium iodide (LiI) is dissolved in water, the solution becomes hotter.

a. Is the dissolution of lithium iodide endothermic or exothermic?

b. What can you conclude about the relative magnitudes of the lattice energy of lithium iodide and its heat of hydration?

c. Sketch a qualitative energy diagram similar to Figure 13.6 for the dissolution of LiI.

d. Why does the solution form? What drives the process?

33. Silver nitrate has a lattice energy of –820 kJ/mol and a heat of solution of –22.6 kJ/mol. Calculate the heat of hydration for silver nitrate.

34. Use the given data to calculate the heats of hydration of lithium chloride and sodium chloride. Which of the two cations, lithium or sodium, has stronger ion–dipole interactions with water? Why?

	Lattice Energy	

Compound	(kJ/mol)	ΔH_{soln} (kJ/mol)
LiCl	−834	−37.0
NaCl	−769	+3.88

35. Lithium iodide has a lattice energy of -7.3×10^2 kJ/mol and a heat of hydration of −793 kJ/mol. Find the heat of solution for lithium iodide and determine how much heat is evolved or absorbed when 15.0 g of lithium iodide completely dissolves in water.
36. Potassium nitrate has a lattice energy of −163.8 kcal/mol and a heat of hydration of −155.5 kcal/mol. How much potassium nitrate has to dissolve in water to absorb 1.00×10^2 kJ of heat?

Solution Equilibrium and Factors Affecting Solubility

37. A solution contains 25 g of NaCl per 100.0 g of water at 25 °C. Is the solution unsaturated, saturated, or supersaturated? (Refer to Figure 13.10.)
38. A solution contains 32 g of KNO_3 per 100.0 g of water at 25 °C. Is the solution unsaturated, saturated, or supersaturated? (Refer to Figure 13.10.)
39. A KNO_3 solution containing 45 g of KNO_3 per 100.0 g of water is cooled from 40 °C to 0 °C. What happens during cooling? (Refer to Figure 13.10.)
40. A KCl solution containing 42 g of KCl per 100.0 g of water is cooled from 60 °C to 0 °C. What happens during cooling? (Refer to Figure 13.10.)
41. Some laboratory procedures involving oxygen-sensitive reactants or products call for using water that has been boiled (and then cooled). Explain.
42. A person preparing a fish tank fills the tank with water that has been boiled (and then cooled). When the person puts fish into the tank, they die. Explain.
43. Scuba divers breathing air at increased pressure can suffer from nitrogen narcosis—a condition resembling drunkenness—when the partial pressure of nitrogen exceeds about 4 atm. What property of gas/water solutions causes this to happen? How can a diver reverse this effect?
44. Scuba divers breathing air at increased pressure can suffer from oxygen toxicity—too much oxygen in their bloodstream—when the partial pressure of oxygen exceeds about 1.4 atm. What happens to the amount of oxygen in a diver's bloodstream when he or she breathes oxygen at elevated pressures? How can this be reversed?
45. Calculate the mass of nitrogen dissolved at room temperature in an 80.0-L home aquarium. Assume a total pressure of 1.0 atm and a mole fraction for nitrogen of 0.78.
46. Use Henry's law to determine the molar solubility of helium at a pressure of 1.0 atm and 25 °C.

Concentrations of Solutions

47. An aqueous NaCl solution is made using 112 g of NaCl diluted to a total solution volume of 1.00 L. Calculate the molarity, molality, and mass percent of the solution. (Assume a density of 1.08 g/mL for the solution.)
48. An aqueous KNO_3 solution is made using 72.5 g of KNO_3 diluted to a total solution volume of 2.00 L. Calculate the molarity, molality, and mass percent of the solution. (Assume a density of 1.05 g/mL for the solution.)
49. To what volume should you dilute 50.0 mL of a 5.00-M KI solution so that 25.0 mL of the diluted solution contains 3.05 g of KI?
50. To what volume should you dilute 125 mL of an 8.00-M $CuCl_2$ solution so that 50.0 mL of the diluted solution contains 4.67 g $CuCl_2$?
51. Silver nitrate solutions are used to plate silver onto other metals. What is the maximum amount of silver (in grams) that can be plated out of 4.8 L of an $AgNO_3$ solution containing 3.4% Ag by mass? Assume that the density of the solution is 1.01 g/mL.
52. A dioxin-contaminated water source contains 0.085% dioxin by mass. How much dioxin is present in 2.5 L of this water? Assume a density of 1.00 g/mL.

53. A hard-water sample contains 0.0085% Ca by mass (in the form of Ca^{2+} ions). How much water (in grams) contains 1.2 g of Ca? (1.2 g of Ca is the recommended daily allowance of calcium for adults between 19 and 24 years old.)

54. Lead is a toxic metal that affects the central nervous system. A Pb-contaminated water sample contains 0.0011% Pb by mass. How much of the water (in mL) contains 150 mg of Pb? (Assume a density of 1.0 g/mL.)

55. You can purchase nitric acid in a concentrated form that is 70.3% HNO_3 by mass and has a density of 1.41 g/mL. How can you prepare 1.15 L of 0.100 M HNO_3 from the concentrated solution?

56. You can purchase hydrochloric acid in a concentrated form that is 37.0% HCl by mass and that has a density of 1.20 g/mL. How can you prepare 2.85 L of 0.500 M HCl from the concentrated solution?

57. Describe how to prepare each solution from the dry solute and the solvent.

- **a.** 1.00×10^2 mL of 0.500 M KCl
- **b.** 1.00×10^2 g of 0.500 m KCl
- **c.** 1.00×10^2 g of 5.0% KCl solution by mass

58. Describe how to prepare each solution from the dry solute and the solvent.

- **a.** 125 mL of 0.100 M $NaNO_3$
- **b.** 125 g of 0.100 m $NaNO_3$
- **c.** 125 g of 1.0% $NaNO_3$ solution by mass

59. A solution is prepared by dissolving 28.4 g of glucose ($C_6H_{12}O_6$) in 355 g of water. The final volume of the solution is 378 mL. For this solution, calculate the concentration in each unit.

- **a.** molarity
- **b.** molality
- **c.** percent by mass
- **d.** mole fraction
- **e.** mole percent

60. A solution is prepared by dissolving 20.2 mL of methanol (CH_3OH) in 100.0 mL of water at 25 °C. The final volume of the solution is 118 mL. The densities of methanol and water at this temperature are 0.782 g/mL and 1.00 g/mL, respectively. For this solution, calculate the concentration in each unit.

- **a.** molarity
- **b.** molality
- **c.** percent by mass
- **d.** mole fraction
- **e.** mole percent

61. Household hydrogen peroxide is an aqueous solution containing 3.0% hydrogen peroxide by mass. What is the molarity of this solution? (Assume a density of 1.01 g/mL.)

62. One brand of laundry bleach is an aqueous solution containing 4.55% sodium hypochlorite (NaOCl) by mass. What is the molarity of this solution? (Assume a density of 1.02 g/mL.)

63. An aqueous solution contains 36% HCl by mass. Calculate the molality and mole fraction of the solution.

64. An aqueous solution contains 5.0% NaCl by mass. Calculate the molality and mole fraction of the solution.

Vapor Pressure of Solutions

65. A beaker contains 100.0 mL of pure water. A second beaker contains 100.0 mL of seawater. The two beakers are left side by side on a lab bench for one week. At the end of the week, the liquid level in both beakers has decreased. However, the level has decreased more in one of the beakers than in the other. Which one and why?

66. Which solution has the highest vapor pressure?

- **a.** 20.0 g of glucose ($C_6H_{12}O_6$) in 100.0 mL of water

b. 20.0 g of sucrose $(C_{12}H_{22}O_{11})$ in 100.0 mL of water

c. 10.0 g of potassium acetate $KC_2H_3O_2$ in 100.0 mL of water

67. Calculate the vapor pressure of a solution containing 24.5 g of glycerin $(C_3H_8O_3)$ in 135 mL of water at 30.0 °C. The vapor pressure of pure water at this temperature is 31.8 torr. Assume that glycerin is not volatile and dissolves molecularly (that is, it is not ionic), and use a density of 1.00 g/mL for the water.

68. A solution contains naphthalene $(C_{10}H_8)$ dissolved in hexane (C_6H_{14}) at a concentration of 12.35% naphthalene by mass. Calculate the vapor pressure at 25 °C of hexane above the solution. The vapor pressure of pure hexane at 25 °C is 151 torr.

69. A solution contains 50.0 g of heptane (C_7H_{16}) and 50.0 g of octane (C_8H_{18}) at 25 °C. The vapor pressures of pure heptane and pure octane at 25 °C are 45.8 torr and 10.9 torr, respectively. Assuming ideal behavior, answer each question.

a. What is the vapor pressure of each solution component in the mixture?

b. What is the total pressure above the solution?

c. What is the composition of the vapor in mass percent?

d. Why is the composition of the vapor different from the composition of the solution?

70. 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.)

71. A solution contains 4.08 g of chloroform $(CHCl_3)$ and 9.29 g of acetone (CH_3COCH_3). The vapor pressures at 35 °C of pure chloroform and pure acetone are 295 torr and 332 torr, respectively. Assuming ideal behavior, calculate the vapor pressures of each component and the total vapor pressure above the solution. The experimentally measured total vapor pressure of the solution at 35 °C is 312 torr. Is the solution ideal? If not, what can you say about the relative strength of chloroform–acetone interactions compared to the acetone–acetone and chloroform–chloroform interactions?

72. A solution of methanol and water has a mole fraction of water of 0.312 and a total vapor pressure of 211 torr at 39.9 °C. The vapor pressures of pure methanol and pure water at this temperature are 256 torr and 55.3 torr, respectively. Is the solution ideal? If not, what can you say about the relative strengths of the solute–solvent interactions compared to the solute–solute and solvent–solvent interactions?

Freezing Point Depression, Boiling Point Elevation, and Osmosis

73. A glucose solution contains 55.8 g of glucose $(C_6H_{12}O_6)$ in 455 g of water. Determine the freezing point and boiling point of the solution. (Assume a density of 1.00 g/mL for water.)

74. An ethylene glycol solution contains 21.2 g of ethylene glycol $(C_2H_6O_2)$ in 85.4 mL of water. Determine the freezing point and boiling point of the solution. (Assume a density of 1.00 g/mL for water.)

75. Calculate the freezing point and melting point of a solution containing 10.0 g of naphthalene $(C_{10}H_8)$ in 100.0 mL of benzene. Benzene has a density of 0.877 g/cm^3.

76. Calculate the freezing point and melting point of a solution containing 7.55 g of ethylene glycol $(C_2H_6O_2)$ in 85.7 mL of ethanol. Ethanol has a density of 0.789 g/cm^3.

77. An aqueous solution containing 17.5 g of an unknown molecular (nonelectrolyte) compound in 100.0 g of water has a freezing point of −1.8 °C. Calculate the molar mass of the unknown compound.

78. An aqueous solution containing 35.9 g of an unknown molecular (nonelectrolyte) compound in 150.0 g of water has a freezing point of −1.3 °C. Calculate the molar mass of the unknown compound.

79. Calculate the osmotic pressure of a solution containing 24.6 g of glycerin $(C_3H_8O_3)$ in 250.0 mL of solution at 298 K.

80. What mass of sucrose ($C_{12}H_{22}O_{11}$)would you combine with $5.00 \times 10^2$of water to make a solution with an osmotic pressure of 8.55 atm at 298 K? (Assume a density of 1.0 g/mL for the solution.)

81. A solution containing 27.55 mg of an unknown protein per 25.0 mL solution was found to have an osmotic pressure of 3.22 torr at 25 °C. What is the molar mass of the protein?

82. Calculate the osmotic pressure of a solution containing 18.75 mg of hemoglobin in 15.0 mL of solution at 25 °C. The molar mass of hemoglobin is 6.5×10^4 g/mol

83. Calculate the freezing point and boiling point of each aqueous solution, assuming complete dissociation of the solute.

- **a.** $0.100\ m\ K_2S$
- **b.** 21.5 g of $CuCl_2$ in 4.50×10^2 g water
- **c.** 5.5% $NaNO_3$ by mass (in water)

84. Calculate the freezing point and boiling point in each solution, assuming complete dissociation of the solute.

- **a.** 10.5 g $FeCl_3$ in 1.50×10^2 g water
- **b.** 3.5% KCl by mass (in water)
- **c.** $0.150\ m\ MgF_2$

85. What mass of salt (NaCl) should you add to 1.00 L of water in an ice-cream maker to make a solution that freezes at –10.0 °C? Assume complete dissociation of the NaCl and a density of 1.00 g/mL for water.

86. Determine the required concentration (in percent by mass) for an aqueous ethylene glycol ($C_2H_6O_2$)solution to have a boiling point of 104.0 °C.

87. Use the van't Hoff factors in Table 13.7 to calculate each colligative property.

- **a.** the melting point of a 0.100-m iron(III) chloride solution
- **b.** the osmotic pressure of a 0.085-M potassium sulfate solution at 298 K
- **c.** the boiling point of a 1.22% by mass magnesium chloride solution

88. Referring to the van't Hoff factors in Table 13.7, calculate the mass of solute required to make each aqueous solution.

- **a.** a sodium chloride solution containing 1.50×10^2 gof water that has a melting point of –1.0 °C
- **b.** 2.50×10^2 mLof a magnesium sulfate solution that has an osmotic pressure of 3.82 atm at 298 K
- **c.** an iron(III) chloride solution containing 2.50×10^2 gof water that has a boiling point of 102 °C

89. A 1.2-m aqueous solution of an ionic compound with the formula MX_2 has a boiling point of 101.4 °C. Calculate the van't Hoff factor (i) for MX_2 at this concentration.

90. A 0.95-m aqueous solution of an ionic compound with the formula MX has a freezing point of –3.0 °C. Calculate the van't Hoff factor (i) for MX at this concentration.

91. A 0.100-M ionic solution has an osmotic pressure of 8.3 atm at 25 °C. Calculate the van't Hoff factor (i) for this solution.

92. A solution contains 8.92 g of KBr in 500.0 mL of solution and has an osmotic pressure of 6.97 atm at 25 °C. Calculate the van't Hoff factor (i) for KBr at this concentration.

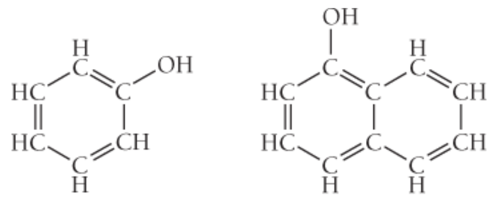
93. Calculate the vapor pressure at 25 °C of an aqueous solution that is 5.50% NaCl by mass. (Assume complete dissociation of the solute.)

94. An aqueous $CaCl_2$ solution has a vapor pressure of 81.6 mmHg at 50 °C. The vapor pressure of pure water at this temperature is 92.6 mmHg. What is the concentration of $CaCl_2$ in mass percent? (Assume complete dissociation of the solute.)

Cumulative Problems

95. The solubility of carbon tetrachloride (CCl_4)in water at 25 °C is 1.2 g/L. The solubility of chloroform ($CHCl_3$)at the same temperature is 10.1 g/L. Why is chloroform almost ten times more soluble in water than carbon tetrachloride?

96. The solubility of phenol in water at 25 °C is 8.7 g/L. The solubility of naphthol at the same temperature is only 0.074 g/L. Examine the structures of phenol and naphthol shown here and explain why phenol is so much more soluble than naphthol.



Phenol Naphthol

97. Potassium perchlorate ($KCIO_4$)has a lattice energy of −599 kJ/mol and a heat of hydration of −548 kJ/mol. Find the heat of solution for potassium perchlorate and determine the temperature change that occurs when 10.0 g of potassium perchlorate is dissolved with enough water to make 100.0 mL of solution. (Assume a heat capacity of 4.05 J/g· °C for the solution and a density of 1.05 g/mL.)

98. Sodium hydroxide (NaOH) has a lattice energy of −887 kJ/mol and a heat of hydration of −932 kJ/mol. How much solution could be heated to boiling by the heat evolved by the dissolution of 25.0 g of NaOH? (For the solution, assume a heat capacity of 4.0 J/g · °C, an initial temperature of 25.0 °C, a boiling point of 100.0 °C, and a density of 1.05 g/mL.)

99. A saturated solution forms when 0.0537 L of argon, at a pressure of 1.0 atm and temperature of 25 °C, is dissolved in 1.0 L of water. Calculate the Henry's law constant for argon.

100. A gas has a Henry's law constant of 0.112 M/atm. What total volume of solution is needed to completely dissolve 1.65 L of the gas at a pressure of 725 torr and a temperature of 25 °C?

101. The Safe Drinking Water Act (SDWA) sets a limit for mercury—a toxin to the central nervous system—at 0.0020 ppm by mass. Water suppliers must periodically test their water to ensure that mercury levels do not exceed this limit. Suppose water becomes contaminated with mercury at twice the legal limit (0.0040 ppm). How much of this water would a person have to consume to ingest 50.0 mg of mercury?

102. Water softeners often replace calcium ions in hard water with sodium ions. Because sodium compounds are soluble, the presence of sodium ions in water does not cause the white, scaly residues caused by calcium ions. However, calcium is more beneficial to human health than sodium because calcium is a necessary part of the human diet, while high levels of sodium intake are linked to increases in blood pressure. The U.S. Food and Drug Administration (FDA) recommends that adults ingest less than 2.4 g of sodium per day. How many liters of softened water, containing a sodium concentration of 0.050% sodium by mass, does a person have to consume to exceed the FDA recommendation? (Assume a water density of 1.0 g/mL.)

103. An aqueous solution contains 12.5% NaCl by mass. What mass of water (in grams) is contained in 2.5 L of the vapor above this solution at 55 °C? The vapor pressure of pure water at 55 °C is 118 torr. (Assume complete dissociation of NaCl.)

104. The vapor above an aqueous solution contains 19.5 mg water per liter at 25 °C. Assuming ideal behavior, what is the concentration of the solute within the solution in mole percent?

105. What is the freezing point of an aqueous solution that boils at 106.5 °C?

106. What is the boiling point of an aqueous solution that has a vapor pressure of 20.5 torr at 25 °C? (Assume a nonvolatile solute.)

107. An isotonic solution contains 0.90% NaCl mass to volume. Calculate the percent mass to volume for isotonic solutions containing each solute at 25 °C. Assume a van't Hoff factor of 1.9 for all *ionic* solutes.

- **a.** KCl
- **b.** NaBr
- **c.** glucose($C_6H_{12}O_6$)

108. Magnesium citrate, $Mg_3(C_6H_5O_7)_2$, belongs to a class of laxatives called *hyperosmotics*, which cause rapid emptying of the bowel. When a person consumes a concentrated solution of magnesium citrate, it passes through the intestines, drawing water and promoting diarrhea, usually within 6 hours. Calculate the osmotic pressure of a magnesium citrate laxative solution containing 28.5 g of magnesium citrate in 235 mL of solution at 37 °C (approximate body temperature). Assume complete dissociation of the ionic compound.

109. A solution is prepared from 4.5701 g of magnesium chloride and 43.238 g of water. The vapor pressure of water above this solution is 0.3624 atm at 348.0 K. The vapor pressure of pure water at this temperature is 0.3804 atm. Find the value of the van't Hoff factor (i) for magnesium chloride in this solution.

110. When HNO_2 dissolves in water, it partially dissociates according to the equation $HNO_2\,(aq) \rightleftharpoons H^+\,(aq) + NO_2^-\,(aq)$ A solution contains 7.050 g of HNO_2 in 1.000 kg of water. Its freezing point is –0.2929 °C. Calculate the fraction of HNO_2 that has dissociated.

111. A solution of a nonvolatile solute in water has a boiling point of 375.3 K. Calculate the vapor pressure of water above this solution at 338 K. The vapor pressure of pure water at this temperature is 0.2467 atm.

112. The density of a 0.438-M solution of potassium chromate (K_2CrCO_4)at 298 K is 1.063 g/mL. Calculate the vapor pressure of water above the solution. The vapor pressure of pure water at this temperature is 0.0313 atm. (Assume complete dissociation of the solute.)

113. The vapor pressure of carbon tetrachloride, CCl_4, is 0.354 atm, and the vapor pressure of chloroform, $CHCl_3$, is 0.526 atm at 316 K. A solution is prepared from equal masses of these two compounds at this temperature. Calculate the mole fraction of the chloroform in the vapor above the solution. If the vapor above the original solution is condensed and isolated into a separate flask, what would the vapor pressure of chloroform be above this new solution?

114. Distillation is a method of purification based on successive separations and recondensations of vapor above a solution. Use the result of the previous problem to calculate the mole fraction of chloroform in the vapor above a solution obtained by three successive separations and condensations of the vapors above the original solution of carbon tetrachloride and chloroform. Show how this result supports the use of distillation as a separation method.

115. A solution of 49.0% H_2SO_4 by mass has a density of 1.39 g/cm^3 at 293 K. A 25.0 cm^3 sample of this solution is mixed with enough water to increase the volume of the solution to 99.8 cm^3. Find the molarity of sulfuric acid in this solution.

116. Find the mass of urea (CH_4N_2O)needed to prepare 50.0 g of a solution in water in which the mole fraction of urea is 0.0770.

117. A solution contains 10.05 g of unknown compound dissolved in 50.0 mL of water. (Assume a density of 1.00 g/mL for water.) The freezing point of the solution is –3.16 °C. The mass percent composition of the compound is 60.97% C, 11.94% H, and the rest is O. What is the molecular formula of the compound?

118. The osmotic pressure of a solution containing 2.10 g of an unknown compound dissolved in 175.0 mL of solution at 25 °C is 1.93 atm. The combustion of 24.02 g of the unknown compound produced 28.26 g $CO_2$and 8.64 g H_2O What is the molecular formula of the compound (which contains only carbon, hydrogen, and oxygen)?

119. A 100.0-mL aqueous sodium chloride solution is 13.5% NaCl by mass and has a density of 1.12 g/mL. What would you add (solute or solvent) and what mass of it to make the boiling point of the solution $104.4\,^{\circ}C$? (Use $i = 1.8$ for NaCl.)

120. A 50.0-mL solution is initially 1.55% $MgCl_2$by mass and has a density of 1.05 g/mL. What is the freezing point of the solution after you add an additional 1.35 g $MgCl_2$? (Use $i = 2.5$ for $MgCl_2$.)

Challenge Problems

121. The small bubbles that form on the bottom of a water pot that is being heated (before boiling) are due to dissolved air coming out of solution. Use Henry's law and the solubilities given to calculate the total volume of nitrogen and oxygen gas that should bubble out of 1.5 L of water upon warming from 25 °C to 50 °C. Assume that the water is initially saturated with nitrogen and oxygen gas at 25 °C and a total pressure of 1.0 atm. Assume that the gas bubbles out at a temperature of 50 °C. The solubility of oxygen gas at 50 °C is 27.8 mg/L at an oxygen pressure of 1.00 atm. The solubility of nitrogen gas at 50 °C is 14.6 mg/L at a nitrogen pressure of 1.00 atm. Assume that the air above the water contains an oxygen partial pressure of 0.21 atm and a nitrogen partial pressure of 0.78 atm.

122. The vapor above a mixture of pentane and hexane at room temperature contains 35.5% pentane by mass. What is the mass percent composition of the solution? Pure pentane and hexane have vapor pressures of 425 torr and 151 torr, respectively, at room temperature.

123. A 1.10-g sample contains only glucose ($C_6H_{12}O_6$) and sucrose ($C_{12}H_{22}O_{11}$). When the sample is dissolved in water to a total solution volume of 25.0 mL, the osmotic pressure of the solution is 3.78 atm at 298 K. What is the mass percent composition of glucose and sucrose in the sample?

124. A solution is prepared by mixing 631 mL of methanol with 501 mL of water. The molarity of methanol in the resulting solution is 14.29 M. The density of methanol at this temperature is 0.792 g/mL. Calculate the difference in volume between this solution and the total volume of water and methanol that were mixed to prepare the solution.

125. Two alcohols, isopropyl alcohol and propyl alcohol, have the same molecular formula, C_3H_8O. A solution of the two that is two-thirds by mass isopropyl alcohol has a vapor pressure of 0.110 atm at 313 K. A solution that is one-third by mass isopropyl alcohol has a vapor pressure of 0.089 atm at 313 K. Calculate the vapor pressure of each pure alcohol at this temperature. Explain the difference given that the formula of propyl alcohol is $CH_3CH_2CH_2OH$ and that of isopropyl alcohol is $(CH_3)_2CHOH$

126. A metal, M, of atomic mass 96 amu reacts with fluorine to form a salt that can be represented as MF_x. In order to determine *x* and therefore the formula of the salt, a boiling point elevation experiment is performed. A 9.18-g sample of the salt is dissolved in 100.0 g of water, and the boiling point of the solution is found to be 374.38 K. Find the formula of the salt. (Assume complete dissociation of the salt in solution.)

127. Sulfuric acid in water dissociates completely into H^+ and HSO_4^- ions. The HSO_4^- ion dissociates to a limited extent into H^+ and SO_4^{2-}. The freezing point of a 0.1000-*m* solution of sulfuric acid in water is 272.76 K. Calculate the molality of SO_4^{2-} in the solution, assuming ideal solution behavior.

128. A solution of 75.0 g of benzene (C_6H_6) and 75.0 g of toluene (C_7H_8) has a total vapor pressure of 80.9 mmHg at 303 K. Another solution of 100.0-g benzene and 50.0-g toluene has a total vapor pressure of 93.9 mmHg at this temperature. Find the vapor pressure of pure benzene and pure toluene at 303 K. (Assume ideal solutions.)

129. A solution is prepared by dissolving 11.60 g of a mixture of sodium carbonate and sodium bicarbonate in 1.00 L of water. A 300.0 cm^3 sample of the solution is treated with excess HNO_3 and boiled to remove all the dissolved gas. A total of 0.940 L of dry CO_2 is collected at 298 K and 0.972 atm. Find the molarity of the carbonate and bicarbonate in the solution.

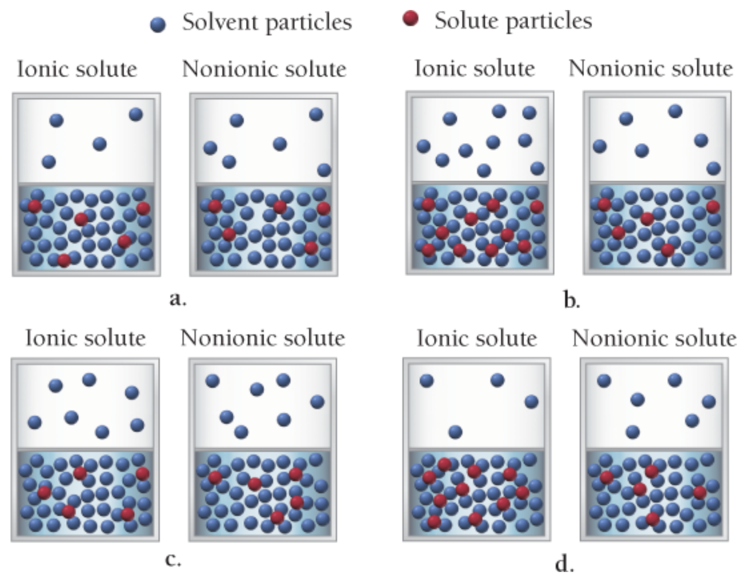
Conceptual Problems

130. Substance A is a nonpolar liquid and has only dispersion forces among its constituent particles. Substance B is also a nonpolar liquid and has about the same magnitude of dispersion forces among its constituent particles as substance A. When substance A and B are combined, they spontaneously mix.

- **a.** Why do the two substances mix?
- **b.** Predict the sign and magnitude of ΔH_{soln}.
- **c.** Determine the signs and relative magnitudes of ΔH_{solute}, $\Delta H_{solvent}$, and ΔH_{mix}.

131. A power plant built on a river uses river water as a coolant. The water is warmed as it is used in heat exchangers within the plant. Should the warm water be immediately cycled back into the river? Why or why not?

132. The vapor pressure of a 1-M ionic solution is different from the vapor pressure of a 1-M nonelectrolyte solution. In both cases, the solute is nonvolatile. Which set of diagrams best represents the differences between the two solutions and their vapors?



133. If each substance listed here costs the same amount per kilogram, which would be most cost-effective as a way to lower the freezing point of water? (Assume complete dissociation for all ionic compounds.) Explain.

- **a.** $HOCH_2CH_2OH$
- **b.** NaCl
- **c.** KCl
- **d.** $MgCl_2$
- **e.** $SrCl_2$

134. A helium balloon inflated on one day falls to the ground the next day. The volume of the balloon decreases somewhat overnight but not by enough to explain why it no longer floats. (If you inflate a new balloon with helium to the same size as the balloon that fell to the ground, the newly inflated balloon floats.) Explain.

Questions for Group Work

Active Classroom Learning

Discuss these questions with the group and record your consensus answer.

135. Explain why 1-propanol ($CH_3CH_2CH_2OH$) is miscible in both water (H_2O) and hexane (C_6H_6) when hexane and water are barely soluble in each other.

136. Have each group member make a flashcard with one of the following on the front: ΔH_{soln}, $\Delta H_{lattice}$, ΔH_{solven}, ΔH_{mix}, and $\Delta H_{hydration}$. On the back of the card, each group member should describe (in words) the ΔH process his or her card lists and how that ΔH relates to other ΔH values mathematically. Each member then presents his or her ΔH to the group. After everyone has presented, members should trade cards and quiz each other.

137. Complete the table by adding *increases*, *decreases*, or *no effect*:

	Increasing Temperature	Increasing Pressure
Solubility of gas in water		
Solubility of a solid in water		

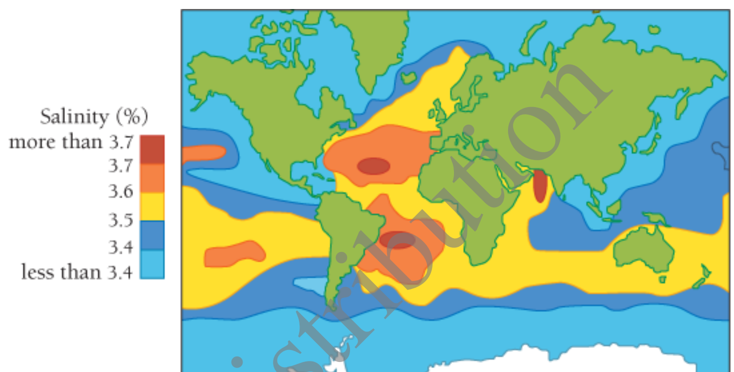
138. When 13.62 g (about one tablespoon) of table sugar (sucrose, $C_{12}H_{22}O_{11}$) is dissolved in 241.5 mL of water (density 0.997 g/mL), the final volume is 250.0 mL (about one cup). Have each group member calculate one of the following for the solution, and present his or her answer to the group:

- **a.** mass percent
- **b.** molarity
- **c.** molality

139. Calculate the expected boiling and freezing points for the solution in the previous problem. If you had to bring this syrup to the boiling point for a recipe, would you expect it to take much more time than it takes to boil the same amount of pure water? Why or why not? Would the syrup freeze in a typical freezer (−18 °C)? Why or why not?

Data Interpretation and Analysis

140. The salinity of seawater can vary in the world's oceans as shown in the map, which indicates salinity in units of percent by mass NaCl. Examine the image and answer the questions.



Salinity of Seawater

- **a.** Which regions of the globe generally have higher salinity? Lower salinity? State your answer as a general trend in the salinity of seawater.
- **b.** Speculate on possible reasons for the trend you observed in part a.
- **c.** Calculate the freezing point of a sample of seawater taken from the middle of the Atlantic Ocean. Use a van't Hoff factor of 1.9 for your calculation.
- **d.** Make a graph of the freezing point of the seawater versus salinity for the range of salinities in the world's oceans.

Answers to Conceptual Connections

Cc 13.1 The first alcohol on the list is methanol, which is highly polar and forms hydrogen bonds with water. It is miscible in water and has only limited solubility in hexane, which is nonpolar. However, as the carbon chain gets longer in the series of alcohols, the OH group becomes less important relative to the growing nonpolar carbon chain. Therefore, the alcohols become progressively less soluble in water and more soluble in hexane. This table demonstrates the rule of thumb *like dissolves like.* Methanol is like water and therefore dissolves in water. It is unlike hexane and therefore has limited solubility in hexane. As you move down the list, the alcohols become increasingly like hexane and increasingly unlike water and therefore become increasingly soluble in hexane and increasingly insoluble in water.

Cc 13.2 You can conclude that $|\Delta H_{\text{solute}}| < |\Delta H_{\text{hydration}}|$. Because ΔH_{soln} is negative, the absolute value of the negative term ($\Delta H_{\text{hydration}}$) must be greater than the absolute value of the positive term (ΔH_{solute})

Cc 13.3 **(b)** Some potassium bromide precipitates out of solution. The solubility of most solids decreases with decreasing temperature. However, the solubility of gases increases with decreasing temperature. Therefore, the nitrogen becomes more soluble and will not bubble out of solution.

Cc 13.4 Ammonia is the only compound on the list that is polar, so we would expect its solubility in water to be greater than those of the other gases (which are all nonpolar).

Cc 13.5 The solution has a molality of 10.0 *m*. You combined 1.00 mol of solute with 0.100 kg of solvent, so the molality is 1.00 mol/0.100 kg = 10.0 *m*.

Cc 13.6 The solute–solvent interactions must be stronger than the solute–solute and solvent–solvent interactions. The stronger interactions lower the vapor pressure from the expected ideal value of 150 mmHg.

Cc 13.7 Solution B because K_b for ethanol is greater than K_b for water.

Cc 13.8 **(c)** The 0.50-M $MgCl_2$ solution has the highest boiling point because it has the highest concentration of particles. We expect 1 mol of $MgCl_2$ to form 3 mol of particles in solution (although it effectively forms slightly fewer).