The Uniqueness of Water (11.9)

Water is a liquid at room temperature despite its low molar mass. Water forms strong hydrogen bonds, resulting in its high boiling point. Its high polarity also enables it to dissolve many polar and ionic compounds, and even nonpolar gases. Water expands upon freezing, so that ice is less dense than liquid water. Water is critical both to the existence of life and to human health.

Crystalline Structures (11.10–11.13)

In X-ray crystallography, the diffraction pattern of X-rays is used to determine the crystal structure of solids. The crystal lattice is represented by a unit cell, a structure that reproduces the entire lattice when repeated in all three dimensions. Three basic cubic unit cells are the simple cubic, the body-centered cubic, and the face-centered cubic. Some crystal lattices can also be depicted as closest-packed structures, including the hexagonal closest-packing structure (not cubic) and the cubic closest-packing structure (which has a face-centered cubic unit cell). The types of crystal solids are molecular, ionic, and atomic solids. Atomic solids can be divided into three different types: nonbonded, metallic, and covalent. Band theory is a model for bonding in solids in which the atomic orbitals of the atoms are combined and delocalized over the entire crystal solid.

Key Equations and Relationships

Clausius–Clapeyron Equation: Relationship between Vapor Pressure (P_{vap}), the Heat of Vaporization (H_{vap}), and Temperature (T) (11.5)

$$\ln P_{vap} = \frac{-\Delta H_{vap}}{RT} + \ln \beta \; (\beta \text{ is a constant})$$

$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{vap}}{R}\left[\frac{1}{T_2} - \frac{1}{T_1}\right]$$

Bragg's Law: Relationship between Light Wavelength (λ), Angle of Reflection (θ), and Distance (d) between the Atomic Layers (11.10)

$$n\lambda = 2d \sin \theta \; (n = \text{integer})$$

Key Skills

Determining Whether a Molecule Has Dipole–Dipole Forces (11.3)
• Example 11.1 • For Practice 11.1 • Exercises 49–60

Determining Whether a Molecule Displays Hydrogen Bonding (11.3)
• Example 11.2 • For Practice 11.2 • Exercises 49–60

Using the Heat of Vaporization in Calculations (11.5)
• Example 11.3 • For Practice 11.3 • For More Practice 11.3 • Exercises 71–74

Using the Clausius–Clapeyron Equation (11.5)
• Examples 11.4, 11.5 • For Practice 11.4, 11.5 • Exercises 75–78

Using Bragg's Law in X-Ray Diffraction Calculations (11.10)
• Example 11.6 • For Practice 11.6 • Exercises 95, 96

Relating Density to Crystal Structure (11.11)
• Example 11.7 • For Practice 11.7 • Exercises 99–102

EXERCISES

Review Questions

1. Explain why a gecko is able to walk on a polished glass surface.
2. Why are intermolecular forces important?
3. What are the main properties of liquids (in contrast to gases and solids)?
4. What are the main properties of solids (in contrast to liquids and gases)?
5. What is the fundamental difference between an amorphous solid and a crystalline solid?
6. What factors cause changes between the solid and liquid state? The liquid and gas state?
7. Describe the relationship between the state of a substance, its temperature, and the strength of its intermolecular forces.
8. From what kinds of interactions do intermolecular forces originate?
9. Why are intermolecular forces generally much weaker than bonding forces?
10. What is the dispersion force? What does the magnitude of the dispersion force depend on? How can you predict the magnitude of the dispersion force for closely related elements or compounds?

11. What is the dipole–dipole force? How can you predict the presence of dipole–dipole forces in a compound?

12. How is the miscibility of two liquids related to their polarity?

13. What is hydrogen bonding? How can you predict the presence of hydrogen bonding in a compound?

14. What is the ion–dipole force? Why is it important?

15. What is surface tension? How does surface tension result from intermolecular forces? How is it related to the strength of intermolecular forces?

16. What is viscosity? How does viscosity depend on intermolecular forces? What other factors affect viscosity?

17. What is capillary action? How does it depend on the relative strengths of adhesive and cohesive forces?

18. Explain what happens in the processes of vaporization and condensation. Why does the rate of vaporization increase with increasing temperature and surface area?

19. Why is vaporization endothermic? Why is condensation exothermic?

20. How is the volatility of a substance related to the intermolecular forces present within the substance?

21. What is the heat of vaporization for a liquid and why is it useful?

22. Explain the process of dynamic equilibrium. How is dynamic equilibrium related to vapor pressure?

23. What happens to a system in dynamic equilibrium when it is disturbed in some way?

24. How is vapor pressure related to temperature? What happens to the vapor pressure of a substance when the temperature is increased? Decreased?

25. Define the terms *boiling point* and *normal boiling point*.

26. What is the Clausius–Clapeyron equation and why is it important?

27. Explain what happens to a substance when it is heated in a closed container to its critical temperature.

28. What is sublimation? Give a common example of sublimation.

29. What is fusion? Is fusion exothermic or endothermic? Why?

30. What is the heat of fusion and why is it important?

31. Examine the heating curve for water in Section 11.7 (Figure 11.36). Explain why the curve has two segments in which heat is added to the water but the temperature does not rise.

32. Examine the heating curve for water in Section 11.7 (Figure 11.36). Explain the significance of the slopes of each of the three rising segments. Why are the slopes different?

33. What is a phase diagram? Draw a generic phase diagram and label its important features.

34. What is the significance of crossing a line in a phase diagram?

35. How do the properties of water differ from those of most other substances?

36. Explain the basic principles involved in X-ray crystallography. Include Bragg's law in your explanation.

37. What is a crystalline lattice? How is the lattice represented with the unit cell?

38. Make a drawing of each unit cell: simple cubic, body-centered cubic, and face-centered cubic.

39. For each of the cubic cells in the previous problem, give the coordination number, edge length in terms of *r*, and number of atoms per unit cell.

40. What is the difference between hexagonal closest packing and cubic closest packing? What are the unit cells for each of these structures?

41. What are the three basic types of solids and the composite units of each? What types of forces hold each type of solid together?

42. In an ionic compound, how are the relative sizes of the cation and anion related to the coordination number of the crystal structure?

43. Show how the cesium chloride, sodium chloride, and zinc blende unit cells each contain a cation-to-anion ratio of 1:1.

44. Show how the fluorite structure accommodates a cation-to-anion ratio of 1:2.

45. What are the three basic subtypes of atomic solids? What kinds of forces hold each of these subtypes together?

46. In band theory of bonding for solids, what is a *band*? What is the difference between the *valence band* and the *conduction band*?

47. What is a band gap? How does the band gap differ in metals, semiconductors, and insulators?

48. Explain how doping can increase the conductivity of a semiconductor. What is the difference between an n-type semiconductor and a p-type semiconductor?

Problems by Topic

Intermolecular Forces

49. Determine the kinds of intermolecular forces that are present in each element or compound:

a. N_2 **b.** NH_3 **c.** CO **d.** CCl_4

50. Determine the kinds of intermolecular forces that are present in each element or compound:

a. Kr **b.** NCl_3 **c.** SiH_4 **d.** HF

51. Determine the kinds of intermolecular forces that are present in each element or compound:

a. HCl **b.** H_2O **c.** Br_2 **d.** He

52. Determine the kinds of intermolecular forces that are present in each element or compound:

a. PH_3 **b.** HBr **c.** CH_3OH **d.** I_2

53. Arrange these compounds in order of increasing boiling point. Explain your reasoning.

a. CH_4 **b.** CH_3CH_3
c. CH_3CH_2Cl **d.** CH_3CH_2OH

54. Arrange these compounds in order of increasing boiling point. Explain your reasoning.

a. H_2S **b.** H_2Se **c.** H_2O

55. For each pair of compounds, pick the one with the highest boiling point. Explain your reasoning.

a. CH_3OH or CH_3SH **b.** CH_3OCH_3 or CH_3CH_2OH
c. CH_4 or CH_3CH_3

56. For each pair of compounds, pick the one with the higher boiling point. Explain your reasoning.

a. NH_3 or CH_4 **b.** CS_2 or CO_2 **c.** CO_2 or NO_2

57. For each pair of compounds, pick the one with the higher vapor pressure at a given temperature. Explain your reasoning.
a. Br_2 or I_2 **b.** H_2S or H_2O **c.** NH_3 or PH_3

58. For each pair of compounds, pick the one with the higher vapor pressure at a given temperature. Explain your reasoning.
a. CH_4 or CH_3Cl
b. $CH_3CH_2CH_2OH$ or CH_3OH
c. CH_3OH or H_2CO

59. Which pairs of substances would you expect to form homogeneous solutions when combined? For those that form homogeneous solutions, indicate the type of forces that are involved.
a. CCl_4 and H_2O **b.** KCl and H_2O
c. Br_2 and CCl_4 **d.** CH_3CH_2OH and H_2O

60. Which pairs of compounds would you expect to form homogeneous solutions when combined? For those that form homogeneous solutions, indicate the type of forces that are involved.
a. $CH_3CH_2CH_2CH_2CH_3$ and $CH_3CH_2CH_2CH_2CH_2CH_3$
b. CBr_4 and H_2O
c. $LiNO_3$ and H_2O
d. CH_3OH and $CH_3CH_2CH_2CH_2CH_3$

Surface Tension, Viscosity, and Capillary Action

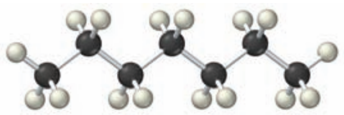
61. Which compound would you expect to have greater surface tension, acetone $[(CH_3)_2CO]$ or water (H_2O)? Explain.

62. Water (a) "wets" some surfaces and beads up on others. Mercury (b), in contrast, beads up on almost all surfaces. Explain this difference.

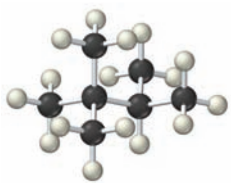
(a) (b)

63. The structures of two isomers of heptane are shown here. Which of these two compounds would you expect to have the greater viscosity?

Compound A

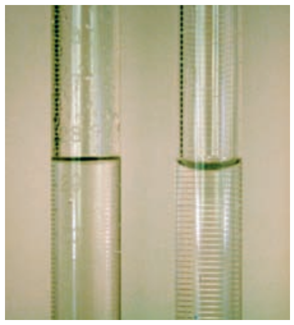


Compound B



64. Explain why the viscosity of multigrade motor oils is less temperature dependent than that of single-grade motor oils.

65. Water in a glass tube that contains grease or oil residue displays a flat meniscus (left); whereas water in a clean glass tube displays a concave meniscus (right). Explain this difference.



66. When a thin glass tube is put into water, the water rises 1.4 cm. When the same tube is put into hexane, the hexane rises only 0.4 cm. Explain the difference.

Vaporization and Vapor Pressure

67. Which will evaporate more quickly: 55 mL of water in a beaker with a diameter of 4.5 cm, or 55 mL of water in a dish with a diameter of 12 cm? Will the vapor pressure of the water be different in the two containers? Explain.

68. Which will evaporate more quickly: 55 mL of water (H_2O) in a beaker or 55 mL of acetone $[(CH_3)_2CO]$ in an identical beaker under identical conditions? Is the vapor pressure of the two substances different? Explain.

69. Spilling room-temperature water over your skin on a hot day will cool you down. Spilling room-temperature vegetable oil over your skin on a hot day will not. Explain the difference.

70. Why is the heat of vaporization of water greater at room temperature than it is at its boiling point?

71. The human body obtains 915 kJ of energy from a candy bar. If this energy were used to vaporize water at 100.0 °C, how much water (in liters) could be vaporized? (Assume the density of water is 1.00 g/mL.)

72. A 100.0-mL sample of water is heated to its boiling point. How much heat (in kJ) is required to vaporize it? (Assume a density of 1.00 g/mL.)

73. Suppose that 0.95 g of water condenses on a 75.0-g block of iron that is initially at 22 °C. If the heat released during condensation goes only to warming the iron block, what is the final temperature (in °C) of the iron block? (Assume a constant enthalpy of vaporization for water of 44.0 kJ/mol.)

74. Suppose that 1.15 g of rubbing alcohol (C_3H_8O) evaporates from a 65.0-g aluminum block. If the aluminum block is initially at 25 °C, what is the final temperature of the block after the evaporation of the alcohol? Assume that the heat required for the vaporization of the alcohol comes only from the aluminum block and that the alcohol vaporizes at 25 °C.

75. This table displays the vapor pressure of ammonia at several different temperatures. Use the data to determine the heat of vaporization and normal boiling point of ammonia.

Temperature (K)	Pressure (torr)
200	65.3
210	134.3
220	255.7
230	456.0
235	597.0

76. This table displays the vapor pressure of nitrogen at several different temperatures. Use the data to determine the heat of vaporization and normal boiling point of nitrogen.

Temperature (K)	Pressure (torr)
65	130.5
70	289.5
75	570.8
80	1028
85	1718

77. Ethanol has a heat of vaporization of 38.56 kJ/mol and a normal boiling point of 78.4 °C. What is the vapor pressure of ethanol at 15 °C?

78. Benzene has a heat of vaporization of 30.72 kJ/mol and a normal boiling point of 80.1 °C. At what temperature does benzene boil when the external pressure is 445 torr?

Sublimation and Fusion

79. How much energy is released when 65.8 g of water freezes?

80. Calculate the amount of heat required to completely sublime 50.0 g of solid dry ice (CO_2) at its sublimation temperature. The heat of sublimation for carbon dioxide is 32.3 kJ/mol.

81. An 8.5-g ice cube is placed into 255 g of water. Calculate the temperature change in the water upon the complete melting of the ice. Assume that all of the energy required to melt the ice comes from the water.

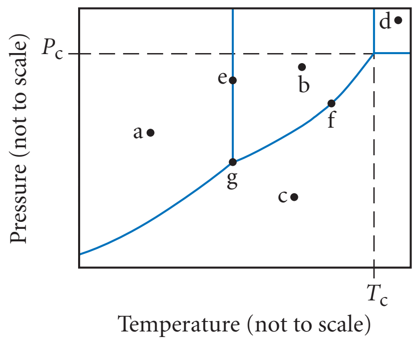
82. How much ice (in grams) would have to melt to lower the temperature of 352 mL of water from 25 °C to 5 °C? (Assume the density of water is 1.0 g/mL.)

83. How much heat (in kJ) is required to warm 10.0 g of ice, initially at −10.0 °C, to steam at 110.0 °C? The heat capacity of ice is 2.09 J/g · °C and that of steam is 2.01 J/g · °C.

84. How much heat (in kJ) is evolved in converting 1.00 mol of steam at 145.0 °C to ice at −50.0 °C? The heat capacity of steam is 2.01 J/g · °C and of ice is 2.09 J/g · °C.

Phase Diagrams

85. Consider the phase diagram shown here. Identify the states present at points *a* through *g*.



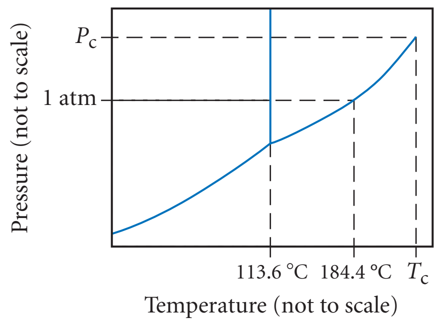
86. Consider the phase diagram for iodine shown here and answer each of the following questions.

a. What is the normal boiling point for iodine?

b. What is the melting point for iodine at 1 atm?

c. What state is present at room temperature and normal atmospheric pressure?

d. What state is present at 186 °C and 1.0 atm?



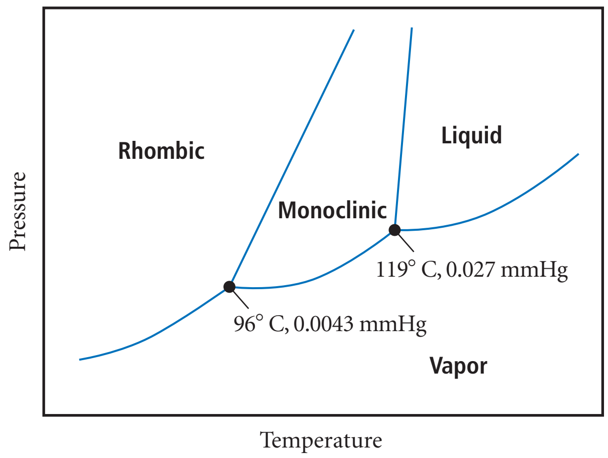
87. Nitrogen has a normal boiling point of 77.3 K and a melting point (at 1 atm) of 63.1 K. Its critical temperature is 126.2 K and critical pressure is 2.55×10^4 torr. It has a triple point at 63.1 K and 94.0 torr. Sketch the phase diagram for nitrogen. Does nitrogen have a stable liquid state at 1 atm?

88. Argon has a normal boiling point of 87.2 K and a melting point (at 1 atm) of 84.1 K. Its critical temperature is 150.8 K and critical pressure is 48.3 atm. It has a triple point at 83.7 K and 0.68 atm. Sketch the phase diagram for argon. Which has the greater density, solid argon or liquid argon?

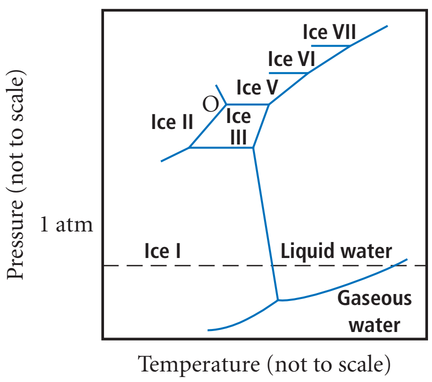
89. The phase diagram for sulfur is shown below. The rhombic and monoclinic states are two solid states with different structures.

a. Below what pressure will solid sulfur sublime?

b. Which of the two solid states of sulfur is most dense?



90. The high-pressure phase diagram of ice is shown here. Notice that, under high pressure, ice can exist in several different solid forms. What three forms of ice are present at the triple point marked O? What is the density of ice II compared to ice I (the familiar form of ice). Would ice III sink or float in liquid water?



The Uniqueness of Water

91. Water has a high boiling point for its relatively low molar mass. Why?

92. Water is a good solvent for many substances. What is the molecular basis for this property and why is it significant?

93. Explain the role of water in moderating Earth's climate.

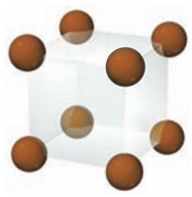
94. How is the density of solid water compared to that of liquid water atypical among substances? Why is this significant?

Types of Solids and Their Structures

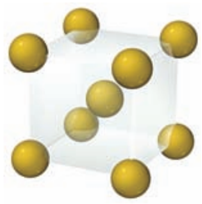
95. An X-ray beam with $\lambda = 154$ pm incident on the surface of a crystal produced a maximum reflection at an angle of $\theta = 28.3°$. Assuming $n = 1$, calculate the separation between layers of atoms in the crystal.

96. An X-ray beam of unknown wavelength is diffracted from a NaCl surface. If the interplanar distance in the crystal is 286 pm, and the angle of maximum reflection is found to be 7.23°, what is the wavelength of the X-ray beam? (Assume $n = 1$.)

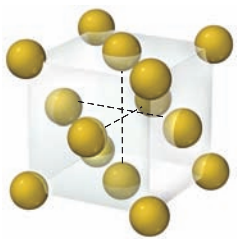
97. Determine the number of atoms per unit cell for each metal.



(a) Polonium

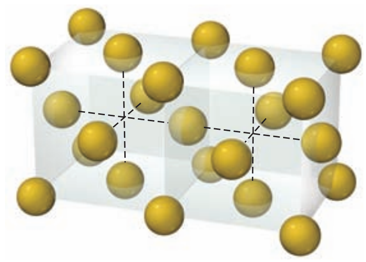


(b) Tungsten

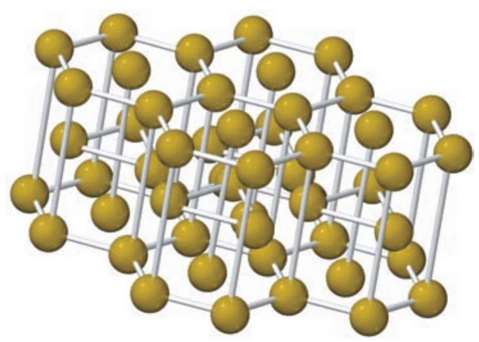


(c) Nickel

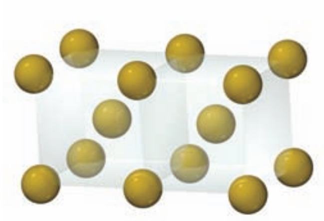
98. Determine the coordination number for each structure.



(a) Gold



(b) Ruthenium



(c) Chromium

99. Platinum crystallizes with the face-centered cubic unit cell. The radius of a platinum atom is 139 pm. Calculate the edge length of the unit cell and the density of platinum in g/cm^3.

100. Molybdenum crystallizes with the body-centered unit cell. The radius of a molybdenum atom is 136 pm. Calculate the edge length of the unit cell and the density of molybdenum.

101. Rhodium has a density of 12.41 g/cm^3 and crystallizes with the face-centered cubic unit cell. Calculate the radius of a rhodium atom.

102. Barium has a density of 3.59 g/cm^3 and crystallizes with the body-centered cubic unit cell. Calculate the radius of a barium atom.

103. Polonium crystallizes with a simple cubic structure. It has a density of 9.3 g/cm^3, a radius of 167 pm, and a molar mass of 209 g/mol. Use this data to estimate Avogadro's number (the number of atoms in one mole).

104. Palladium crystallizes with a face-centered cubic structure. It has a density of 12.0 g/cm^3, a radius of 138 pm, and a molar mass of 106.42 g/mol. Use this data to estimate Avogadro's number.

105. Identify each solid as molecular, ionic, or atomic.
a. Ar(s) **b.** H_2O(s) **c.** K_2O(s) **d.** Fe(s)

106. Identify each solid as molecular, ionic, or atomic.
a. $CaCl_2$(s) **b.** CO_2(s) **c.** Ni(s) **d.** I_2(s)

107. Which solid has the highest melting point? Why?

Ar(s), CCl_4(s), LiCl(s), CH_3OH(s)

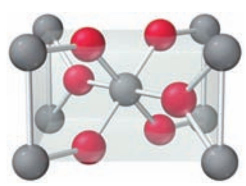
108. Which solid has the highest melting point? Why?

C(s, diamond), Kr(s), NaCl(s), H_2O(s)

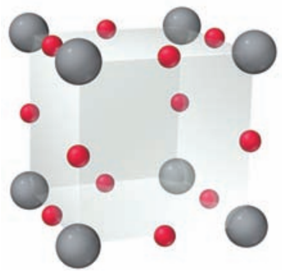
109. In each pair of solids, which one has the higher melting point and why?
a. TiO_2(s) or HOOH(s) **b.** CCl_4(s) or $SiCl_4$(s)
c. Kr(s) or Xe(s) **d.** NaCl(s) or CaO(s)

110. In each pair of solids, which one has the higher melting point and why?
a. Fe(s) or CCl_4(s) **b.** KCl(s) or HCl(s)
c. Ti(s) or Ne(s) **d.** H_2O(s) or H_2S(s)

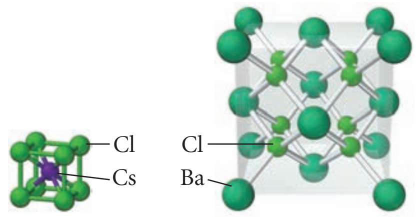
111. An oxide of titanium crystallizes with the unit cell shown here (titanium = gray; oxygen = red). What is the formula of the oxide?



112. An oxide of rhenium crystallizes with the unit cell shown here (rhenium = gray; oxygen = red). What is the formula of the oxide?

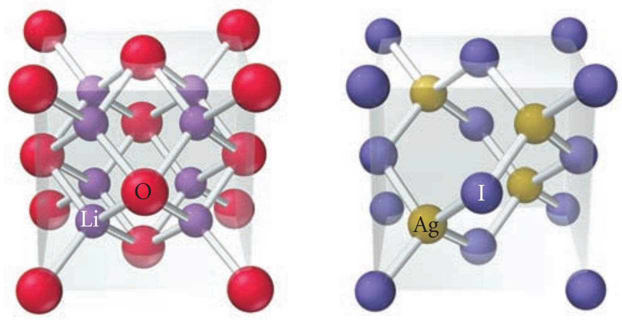


113. The unit cells for cesium chloride and barium(II) chloride are shown below. Show that the ratio of cations to anions in each unit cell corresponds to the ratio of cations to anions in the formula of each compound.



Cesium chloride   Barium(II) chloride

114. The unit cells for lithium oxide and silver iodide are shown here. Show that the ratio of cations to anions in each unit cell corresponds to the ratio of cations to anions in the formula of each compound.



Lithium oxide   Silver iodide

Band Theory

115. Which solid would you expect to have little or no band gap?
a. Zn(*s*) **b.** Si(*s*) **c.** As(*s*)

116. How many molecular orbitals are present in the valence band of a sodium crystal with a mass of 5.45 g?

117. Indicate whether each solid would form an n-type or a p-type semiconductor.
a. germanium doped with gallium
b. silicon doped with arsenic

118. Indicate whether each solid would form an n-type or a p-type semiconductor.
a. silicon doped with gallium
b. germanium doped with antimony

Cumulative Problems

119. Explain the observed trend in the melting points of the hydrogen halides.

HI	−50.8 °C
HBr	−88.5 °C
HCl	−114.8 °C
HF	−83.1 °C

120. Explain the observed trend in the boiling points of these compounds.

H_2Te	−2 °C
H_2Se	−41.5 °C
H_2S	−60.7 °C
H_2O	−100 °C

121. The vapor pressure of water at 25 °C is 23.76 torr. If 1.25 g of water is enclosed in a 1.5-L container, will any liquid be present? If so, what mass of liquid?

122. The vapor pressure of CCl_3F at 300 K is 856 torr. If 11.5 g of CCl_3F is enclosed in a 1.0-L container, will any liquid be present? If so, what mass of liquid?

123. Examine the phase diagram for iodine shown in Figure 11.39(a). What state transitions occur as you uniformly increase the pressure on a gaseous sample of iodine from 0.010 atm at 185 °C to 100 atm at 185 °C? Make a graph, analogous to the heating curve for water shown in Figure 11.36, in which you plot pressure versus time during the pressure increase.

124. Carbon tetrachloride displays a triple point at 249.0 K and a melting point (at 1 atm) of 250.3 K. Which state of carbon tetrachloride is more dense, the solid or the liquid? Explain.

125. Four ice cubes at exactly 0 °C with a total mass of 53.5 g are combined with 115 g of water at 75 °C in an insulated container. If no heat is lost to the surroundings, what will be the final temperature of the mixture?

126. A sample of steam with a mass of 0.552 g and at a temperature of 100 °C condenses into an insulated container holding 4.25 g of water at 5.0 °C. Assuming that no heat is lost to the surroundings, what will be the final temperature of the mixture?

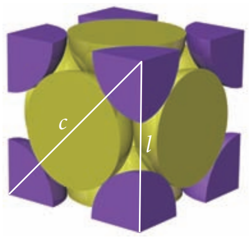
127. Air conditioners not only cool air, but dry it as well. Suppose that a room in a home measures 6.0 m × 10.0 m × 2.2 m. If the outdoor temperature is 30 °C and the vapor pressure of water in the air is 85% of the vapor pressure of water at this temperature, what mass of water must be removed from the air each time the volume of air in the room is cycled through the air conditioner? The vapor pressure for water at 30 °C is 31.8 torr.

128. A sealed flask contains 0.55 g of water at 28 °C. The vapor pressure of water at this temperature is 28.36 mmHg. What is the minimum volume of the flask in order that no liquid water be present in the flask?

129. Silver iodide crystallizes in the zinc blende structure. The separation between nearest neighbor cations and anions is approximately 325 pm and the melting point is 558 °C. Cesium chloride, by contrast, crystallizes in the cesium chloride structure shown in Figure 11.51. Even though the separation between nearest neighbor cations and anions is greater (348 pm), the melting point is higher (645 °C). Explain.

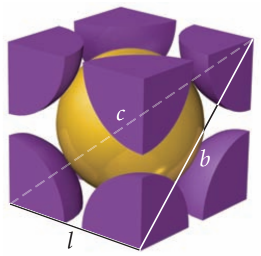
130. Copper iodide crystallizes in the zinc blende structure. The separation between nearest neighbor cations and anions is approximately 311 pm and the melting point is 606 °C. Potassium chloride, by contrast, crystallizes in the rock salt structure. Even though the separation between nearest neighbor cations and anions is greater (319 pm), the melting point is higher (776 °C). Explain.

131. Consider the face-centered cubic structure shown here:



- **a.** What is the length of the line (labeled *c*) that runs diagonally across one of the faces of the cube in terms of *r* (the atomic radius)?
- **b.** Use the answer to part a and the Pythagorean theorem to derive the expression for the edge length (*l*) in terms of *r*.

132. Consider the body-centered cubic structure shown here:



- **a.** What is the length of the line (labeled *c*) that runs from one corner of the cube diagonally through the center of the cube to the other corner in terms of *r* (the atomic radius)?
- **b.** Use the Pythagorean theorem to derive an expression for the length of the line (labeled *b*) that runs diagonally across one of the faces of the cube in terms of the edge length (*l*).
- **c.** Use the answer to parts (a) and (b) along with the Pythagorean theorem to derive the expression for the edge length (*l*) in terms of *r*.

133. The unit cell in a crystal of diamond belongs to a crystal system different from any we have discussed. The volume of a unit cell of diamond is $0.0454\ nm^3$ and the density of diamond is $3.52\ g/cm^3$. Find the number of carbon atoms in a unit cell of diamond.

134. The density of an unknown metal is $12.3\ g/cm^3$ and its atomic radius is 0.134 nm. It has a face-centered cubic lattice. Find the atomic mass of this metal.

135. Based on the phase diagram of CO_2 shown in Figure 11.39(b), describe the state changes that occur when the temperature of CO_2 is increased from 190 K to 350 K at a constant pressure of (a) 1 atm, (b) 5.1 atm, (c) 10 atm, (d) 100 atm.

136. Consider a planet where the pressure of the atmosphere at sea level is 2500 mmHg. Will water behave in a way that can sustain life on the planet?

137. An unknown metal is found to have a density of $7.8748\ g/cm^3$ and to crystallize in a body-centered cubic lattice. The edge of the unit cell is found to be 0.28664 nm. Calculate the atomic mass of the metal.

138. When spheres of radius *r* are packed in a body-centered cubic arrangement, they occupy 68.0% of the available volume. Use the fraction of occupied volume to calculate the value of *a*, the length of the edge of the cube in terms of *r*.

Challenge Problems

139. Potassium chloride crystallizes in the rock salt structure. Estimate the density of potassium chloride using the ionic radii given in Chapter 8.

140. Butane (C_4H_{10}) has a heat of vaporization of 22.44 kJ/mol and a normal boiling point of −0.4 °C. A 250-mL sealed flask contains 0.55 g of butane at −22 °C. How much butane is present as a liquid? If the butane is warmed to 25 °C, how much is present as a liquid?

141. Liquid nitrogen can be used as a cryogenic substance to obtain low temperatures. Under atmospheric pressure, liquid nitrogen boils at 77 K, allowing low temperatures to be reached. However, if the nitrogen is placed in a sealed, insulated container connected to a vacuum pump, even lower temperatures can be reached. Why? If the vacuum pump has sufficient capacity, and is left on for an extended period of time, the liquid nitrogen will start to freeze. Explain.

142. Calculate the fraction of empty space in cubic closest packing to five significant figures.

143. A tetrahedral site in a closest-packed lattice is formed by four spheres at the corners of a regular tetrahedron. This is equivalent to placing the spheres at alternate corners of a cube. In such a closest-packed arrangement the spheres are in contact and if the spheres have a radius *r*, the diagonal of the face of the cube is 2*r*. The tetrahedral hole is inside the middle of the cube. Find the length of the body diagonal of this cube and then find the radius of the tetrahedral hole.

144. Given that the heat of fusion of water is −6.02 kJ/mol, that the heat capacity of $H_2O(l)$ is 75.2 J/mol · K and that the heat capacity of $H_2O(s)$ is 37.7 J/mol · K, calculate the heat of fusion of water at −10 °C.

145. The heat of combustion of CH_4 is 890.4 kJ/mol and the heat capacity of H_2O is 75.2 J/mol · K. Find the volume of methane measured at 298 K and 1.00 atm required to convert 1.00 L of water at 298 K to water vapor at 373 K.

146. Two liquids, A and B, have vapor pressures at a given temperature of 24 mmHg and 36 mmHg, respectively. We prepare solutions of A and B at a given temperature and measure the total pressures above the solutions. We obtain the following data:

Solution	Amt A (mol)	Amt B (mol)	P (mmHg)
1	1	1	30
2	2	1	28
3	1	2	32
4	1	3	33

Predict the total pressure above a solution of 5 mol A and 1 mol B.

147. Three 1.0-L flasks, maintained at 308 K, are connected to each other with stopcocks. Initially the stopcocks are closed. One of the flasks contains 1.0 atm of N_2, the second 2.0 g of H_2O, and the third, 0.50 g of ethanol, C_2H_6O. The vapor pressure of H_2O at 308 K is 42 mmHg and that of ethanol is 102 mmHg. The stopcocks are then opened and the contents mix freely. What is the pressure?

Conceptual Problems

148. One prediction of global warming is the melting of global ice, which may result in coastal flooding. A criticism of this prediction is that the melting of icebergs does not increase ocean levels any more than the melting of ice in a glass of water increases the level of liquid in the glass. Is this a valid criticism? Does the melting of an ice cube in a cup of water raise the level of the liquid in the cup? Why or why not? In response to this criticism, scientists have asserted that they are not worried about melting icebergs, but rather the melting of ice sheets that sit on the continent of Antarctica. Would the melting of this ice increase ocean levels? Why or why not?

149. The rate of vaporization depends on the surface area of the liquid. However, the vapor pressure of a liquid does not depend on the surface area. Explain.

150. Substance A has a smaller heat of vaporization than substance B. Which of the two substances will undergo a larger change in vapor pressure for a given change in temperature?

151. The density of a substance is greater in its solid state than in its liquid state. If the triple point in the phase diagram of the substance is below 1.0 atm, then which will necessarily be at a lower temperature, the triple point or the normal melting point?

152. A substance has a heat of vaporization of ΔH_{vap} and heat of fusion of ΔH_{fus}. Express the heat of sublimation in terms of ΔH_{vap} and ΔH_{fus}.

153. Examine the heating curve for water in Section 11.7 (Figure 11.36). If heat is added to the water at a constant rate, which of the three segments in which temperature is rising will have the least steep slope? Why?

154. A root cellar is an underground chamber used to store fruits, vegetables, and even meats. In extreme cold, farmers put large vats of water into the root cellar to prevent the fruits and vegetables from freezing. Explain why this works.

155. Suggest an explanation for the observation that the heat of fusion of a substance is always smaller than its heat of vaporization.