
11. STATES OF MATTER; LIQUIDS AND SOLIDS

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

11.1 First, calculate the heat required to vaporize 1.00 kg of ammonia:

$$1.00 \text{ kg NH}_3 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{23.4 \text{ kJ}}{1 \text{ mol}} = 1374.04 \text{ kJ}$$

The amount of water at 0°C that can be frozen to ice at 0°C with this heat is

$$1374.04 \text{ kJ} \times \frac{1 \text{ mol H}_2\text{O}}{6.01 \text{ kJ}} \times \frac{18.01 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 4117.54 \text{ g} = 4.12 \text{ kg H}_2\text{O}$$

11.2 Use the two-point form of the Clausius-Clapeyron equation to calculate P_2 :

$$\begin{aligned} \ln \frac{P_2}{760 \text{ mmHg}} &= \frac{26.8 \times 10^3 \text{ J/mol}}{8.31 \text{ J/(K} \cdot \text{mol)}} \left[\frac{1}{319 \text{ K}} - \frac{1}{308 \text{ K}} \right] \\ &= 3225 \text{ K} \left[\frac{-1.12 \times 10^{-4}}{\text{K}} \right] = -0.36106 \end{aligned}$$

(continued)

Converting to antilogs gives

$$\frac{P_2}{760 \text{ mmHg}} = \text{antilog}(-0.36106) = e^{-0.36106} = 0.69693$$

$$P_2 = 0.69693 \times 760 \text{ mmHg} = 529.6 = 5.30 \times 10^2 \text{ mmHg}$$

11.3 Use the two-point form of the Clausius-Clapeyron equation to solve for ΔH_{vap} :

$$\ln \frac{757 \text{ mmHg}}{760 \text{ mmHg}} = \frac{\Delta H_{\text{vap}}}{8.31 \text{ J/(K} \cdot \text{mol)}} \left[\frac{1}{368 \text{ K}} - \frac{1}{378 \text{ K}} \right]$$

$$0.37169 = \frac{\Delta H_{\text{vap}}}{8.31 \text{ J/(K} \cdot \text{mol)}} \left[\frac{7.188 \times 10^{-5}}{\text{K}} \right]$$

$$\Delta H_{\text{vap}} = 4.296 \times 10^4 \text{ J/mol} \quad (43.0 \text{ kJ/mol})$$

11.4 a. Liquefy methyl chloride by a sufficient increase in pressure below 144°C .

b. Liquefy oxygen by compressing to 50 atm below -119°C .

11.5 a. Propanol has a hydrogen atom bonded to an oxygen atom. Therefore, hydrogen bonding is expected. Because propanol is polar (from the O–H bond), we also expect dipole-dipole forces. Weak London forces exist, too, because such forces exist between all molecules.

b. Linear carbon dioxide is not polar, so only London forces exist among CO_2 molecules.

c. Bent sulfur dioxide is polar, so we expect dipole-dipole forces; we also expect the usual London forces.

11.6 The order of increasing vapor pressure is butane (C_4H_{10}), propane (C_3H_8), and ethane (C_2H_6). Because London forces tend to increase with increasing molecular weight, we would expect the molecule with the highest molecular weight to have the lowest vapor pressure.

11.7 Because ethanol has an H atom bonded to an O atom, strong hydrogen bonding exists in ethanol but not in methyl chloride. Hydrogen bonding explains the lower vapor pressure of ethanol compared to methyl chloride.

11.8 a. Zinc, a metal, is a metallic solid.

b. Sodium iodide, an ionic substance, exists as an ionic solid.

c. Silicon carbide, a compound in which carbon and silicon might be expected to form covalent bonds to other carbon and silicon atoms, exists as a covalent network solid.

d. Methane, at room temperature a gaseous molecular compound with covalent bonds, freezes as a molecular solid.

11.9 Only MgSO_4 is an ionic solid; $\text{C}_2\text{H}_5\text{OH}$, CH_4 , and CH_3Cl form molecular solids; thus, MgSO_4 should have the highest melting point. Of the molecular solids, CH_4 has the lowest molecular weight (16.0 amu) and would be expected to have the lowest melting point. Both $\text{C}_2\text{H}_5\text{OH}$ and CH_3Cl have approximately the same molecular weights (46.0 amu vs. 50.5 amu), but $\text{C}_2\text{H}_5\text{OH}$ exhibits strong hydrogen bonding and, therefore, would be expected to have the higher melting point. The order of increasing melting points is CH_4 , CH_3Cl , $\text{C}_2\text{H}_5\text{OH}$, and MgSO_4 .

11.10 Each of the four corners of the cell contains one atom, which is shared by a total of four unit cells. Therefore, the corners contribute one whole atom. This is:

$$\frac{\text{Atoms}}{\text{Unit cell}} = 4 \text{ corners} \times \frac{1/4 \text{ atom}}{1 \text{ corner}} = 1 \text{ atom}$$

11.11 Use the edge length to calculate the volume of the unit cell. Then, use the density to determine the mass of one atom. Divide the molar mass by the mass of one atom.

$$V = (3.509 \times 10^{-10} \text{ m})^3 = 4.321 \times 10^{-29} \text{ m}^3$$

$$D = \frac{0.534 \text{ g}}{1 \text{ cm}^3} \times \left[\frac{100 \text{ cm}}{1 \text{ m}} \right]^3 = 5.34 \times 10^5 \text{ g/m}^3$$

$$\text{Mass of 1 unit} = d \times V$$

$$= (5.34 \times 10^5 \text{ g/m}^3) \times (4.321 \times 10^{-29} \text{ m}^3) = 2.3074 \times 10^{-23} \text{ g}$$

(continued)

There are two atoms in a body-centered cubic unit cell; thus, the mass of one lithium atom is

$$1/2 \times 2.3074 \times 10^{-23} \text{ g} = 1.1537 \times 10^{-23} \text{ g}$$

The known atomic weight of lithium is 6.941 amu, so Avogadro's number is

$$N_A = \frac{6.941 \text{ g/mol}}{1.1537 \times 10^{-23} \text{ g/atom}} = 6.016 \times 10^{23} = 6.02 \times 10^{23} \text{ atoms/mol}$$

- 11.12 Use Avogadro's number to convert the molar mass of potassium to the mass per one atom.

$$\frac{39.0983 \text{ g K}}{1 \text{ mol K}} \times \frac{1 \text{ mol K}}{6.022 \times 10^{23} \text{ atoms}} = \frac{6.4925 \times 10^{-23} \text{ g K}}{1 \text{ atom}}$$

There are two K atoms per unit cell; therefore, the mass per unit cell is

$$\frac{6.4925 \times 10^{-23} \text{ g K}}{1 \text{ atom}} \times \frac{2 \text{ atoms}}{1 \text{ unit cell}} = \frac{1.2985 \times 10^{-22} \text{ g}}{1 \text{ unit cell}}$$

The density of 0.856 g/cm^3 is equal to the mass of one unit cell divided by its unknown volume, V . After solving for V , determine the edge length from the cube root of the volume.

$$0.856 \text{ g/cm}^3 = \frac{1.2985 \times 10^{-22} \text{ g}}{V}$$

$$V = \frac{1.2985 \times 10^{-22} \text{ g}}{0.856 \text{ g/cm}^3} = 1.517 \times 10^{-22} \text{ cm}^3 \quad (1.517 \times 10^{-28} \text{ m}^3)$$

$$\text{Edge length} = \sqrt[3]{1.517 \times 10^{-28} \text{ m}^3} = 5.333 \times 10^{-10} = 5.33 \times 10^{-10} \text{ m} \quad (533 \text{ pm})$$

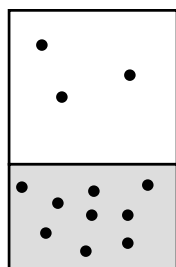
■ Answers to Concept Checks

- 11.1 a. (i) At $t = 0$, since the system is not at equilibrium and there are no H_2O molecules in the gaseous state, you would expect the rate of evaporation to exceed the rate of condensation. At $t = 1$, since evaporation has proceeded at a greater rate than condensation, there must now be fewer molecules in the liquid state, resulting in a lower level of $\text{H}_2\text{O}(l)$.

(continued)

- (ii) At $t = 1$, since some of the H_2O has gone into the vapor state, the vapor pressure must be higher.
- (iii) At $t = 1$, since evaporation has occurred, there must be more molecules in the vapor state.
- (iv) At $t = 0$, since the system is not at equilibrium, and there are no H_2O molecules in the gaseous state, you would expect the rate of evaporation to exceed the rate of condensation.

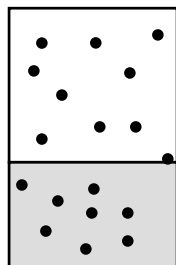
(v)



$t = 1$

- b. (i) Between $t = 1$ and $t = 2$, the system has still not reached equilibrium. Therefore, because the rate of evaporation continues to exceed the rate of condensation, you would expect the water level to be lower.
- (ii) Prior to reaching equilibrium at $t = 2$, you would continue to observe a rate of evaporation greater than the rate of condensation, resulting in a higher vapor pressure than $t = 1$.
- (iii) Since evaporation has been occurring at a greater rate than condensation between points $t = 1$ and $t = 2$, you would expect more molecules in the vapor state at $t = 2$.
- (iv) When the system has reached equilibrium at $t = 2$, the rate of evaporation equals the rate of condensation.

(v)



$t = 2$

- 11.2 You would have to cook the egg for a longer time. The reason is that, since there is lower atmospheric pressure at high altitude, water boils at a lower temperature than near sea level. Since the temperature is lower, it would take longer to transfer an equivalent amount of heat to the egg.
- 11.3 a. If the first reaction occurred, the mixture of hydrogen and oxygen that resulted would form an explosive mixture. At 100°C, 2 moles of water make 61.2 L of gaseous water. In the incorrect reaction, 2 moles of water produce 91.8 L of gaseous products.
- b. Since you would be breaking strong chemical bonds and forming relatively weak bonds, the enthalpy for the first reaction (the wrong reaction) would be many times greater (more positive) than for the second reaction.
- c. Apply Hess's Law. The enthalpy for the wrong reaction would be equal to two times ΔH_f° for $\text{H}_2\text{O}(\text{l})$ plus the heat required to raise the temperature of two moles of water from 25 °C to 100 °C.
- 11.4 a. First, consider the B balls (small). There are four atoms, each completely inside the cell. Thus, there are four B atoms per cell. Next, there are fourteen A atoms (large). Of these, eight are in corners, and contribute 1/8 to the cell. Six atoms are in faces, and contribute 1/2 to the cell. Thus, there are $8 \times (1/8) + 6 \times (1/2) = 4$ A atoms per cell. The ratio of A atoms to B atoms is 4 to 4, or 1 to 1. Thus, the formula of the compound is AB.
- b. Since all of the B atoms are completely within the cell, the shape of the cell is determined by the A atoms only. It is a face-centered cubic unit cell.

■ Answers to Review Questions

- 11.1 The six different phase transitions, with examples in parentheses, are melting (snow melting), sublimation (dry ice subliming directly to carbon dioxide gas), freezing (water freezing), vaporization (water evaporating), condensation (dew forming on the ground), and gas-solid condensation or deposition (frost forming on the ground).
- 11.2 Iodine can be purified by heating it in a beaker covered with a dish containing ice or ice water. Only pure iodine should sublime, crystallizing on the cold bottom surface of the dish above the iodine. The common impurities in iodine do not sublime nor do they vaporize significantly.
- 11.3 The vapor pressure of a liquid is the partial pressure of the vapor over the liquid, measured at equilibrium. In molecular terms, vapor pressure involves molecules of a liquid vaporizing from the liquid phase, colliding with any surface above the liquid, and exerting pressure on it. The equilibrium is a dynamic one because molecules of the liquid are continually leaving the liquid phase and returning to it from the vapor phase.

- 11.4 Steam at 100°C will melt more ice than the same weight of water at 100°C because it contains much more energy in the form of its heat of vaporization. It will transfer this energy to the ice and condense in doing so. The condensed steam and the water will both transfer heat to the ice as the temperature then drops.
- 11.5 The heat of fusion is smaller than the heat of vaporization because melting requires only enough energy for molecules to escape from their sites in the crystal lattice, leaving other molecular attractions intact. In vaporization, sufficient energy must be added to break almost all molecular attractions and also to do the work of pushing back the atmosphere.
- 11.6 Evaporation leads to cooling of a liquid because the gaseous molecules require heat to evaporate; as they leave the other liquid molecules, they remove the heat energy required to vaporize them. This leaves less energy in the liquid whose temperature then drops.
- 11.7 As the temperature increases for a liquid and its vapor in the closed vessel, the two, which are separated by a meniscus, gradually become identical. The meniscus first becomes fuzzy and then disappears altogether as the temperature reaches the critical temperature. Above this temperature, only the vapor exists.
- 11.8 A permanent gas can be liquefied only by lowering the temperature below its critical temperature while compressing the gas.
- 11.9 The pressure in the cylinder of nitrogen at room temperature (above its critical temperature of -147°C) decreases continuously as gas is released because the number of molecules in the vapor phase, which governs the pressure, decreases continuously. The pressure in the cylinder of propane at room temperature (below its critical temperature) is constant because liquid propane and gaseous propane exist at equilibrium in the cylinder. The pressure will remain constant at the vapor pressure of propane until only gaseous propane remains. At that point, the pressure will decrease until all of the propane is gone.
- 11.10 The vapor pressure of a liquid depends on the intermolecular forces in the liquid phase since the ease with which a molecule leaves the liquid phase depends on how strongly it is attracted to the other molecules. If such molecules attract each other strongly, the vapor pressure will be relatively low; if they attract each other weakly, the vapor pressure will be relatively high.
- 11.11 Surface tension makes a liquid act as though it had a skin because, for an object to break through the surface, the surface area must increase. This requires energy, so there is some resistance to the object breaking through the surface.
- 11.12 London forces, also known as dispersion forces, originate between any two molecules that are weakly attracted to each other by means of small instantaneous dipoles that occur as a result of the varying positions of the electrons during their movement about their nuclei.

- 11.13 Hydrogen bonding is a weak to moderate attractive force that exists between a hydrogen atom covalently bonded to a very electronegative atom, X (N, O, or F), and a lone pair of electrons on another small, electronegative atom, Y. (X and Y may be the same or different elements.) Hydrogen bonding in water involves a hydrogen atom of one water molecule bonding to a lone pair of electrons on the oxygen atom of another water molecule.
- 11.14 Molecular substances have relatively low melting points because the forces broken by melting are weak intermolecular attractions in the solid state, not strong bonding attractions.
- 11.15 A crystalline solid has a well-defined, orderly structure; an amorphous solid has a random arrangement of structural units.
- 11.16 In a face-centered cubic cell, there are atoms at the center of each face of the unit cell in addition to those at the corners.
- 11.17 The structure of thallium(I) iodide is a simple cubic lattice for both the metal ions and the anions. Thus, the structure consists of two interpenetrating cubic lattices of cation and anion.
- 11.18 The coordination number of Cs^+ in CsCl is eight; the coordination number of Na^+ in NaCl is six; and the coordination number of Zn^{2+} in ZnS is four.
- 11.19 Starting with the edge length of a cubic crystal, we can calculate the volume of a unit cell by cubing the edge length. Then, knowing the density of the crystalline solid, we can calculate the mass of the atoms in the unit cell. Then, the mass of the atoms in the unit cell is divided by the number of atoms in the unit cell, to give the mass of one atom. Dividing the mass of one mole of the crystal by the mass of one atom yields a value for Avogadro's number.
- 11.20 X rays can strike a crystal and be reflected at various angles; at most angles, the reflected waves will be out of phase and will interfere destructively. At certain angles, however, the reflected waves will be in phase and will interfere constructively, giving rise to a diffraction pattern.

■ Answers to Conceptual Problems

- 11.21 a. The number of molecules in the gas phase is directly related to the kinetic energy needed to escape into the gas phase. The molecules with the most kinetic energy will have the most molecules in the gas phase. Since molecules of C have the highest kinetic energy, they will have the majority of molecules in the gas phase.
- b. The molecules with the strongest intermolecular attractions will have the lowest kinetic energy. Since molecules of A have the lowest kinetic energy, they will have the strongest intermolecular attractions.

(continued)

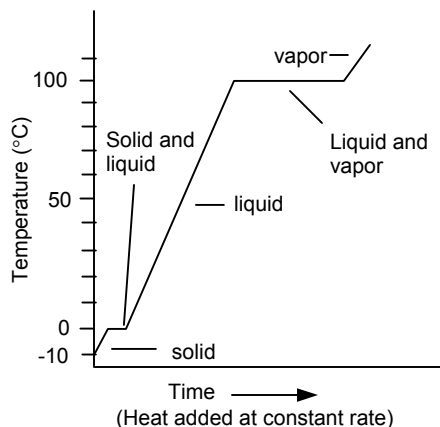
- c. The molecules with the strongest intermolecular attractions will have the lowest vapor pressure. Thus, molecules of A would have the lowest vapor pressure.

11.22 You will need to compare the heats of fusion and vaporization of substance X ($\Delta H_{\text{fus}} = 9.0 \text{ kJ/mol}$ and $\Delta H_{\text{vap}} = 20.3 \text{ kJ/mol}$) with the values for water, which are $\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$, and $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$. Comparing values shows that ΔH_{fus} is 1.5 times larger for substance X, and ΔH_{vap} is 2.0 times larger for H_2O .

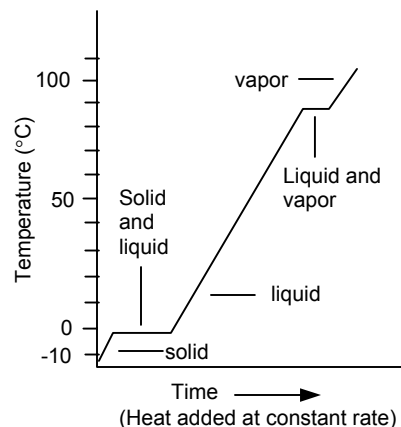
Heating the substance, or water, from -10°C to the boiling point is a three-step process. Step 1 is to heat the solid from -10°C to 0°C , the freezing point. The heat required for this step is equal to mass \times specific heat capacity \times temperature change. Step 2 is to melt the solid to liquid at 0°C . The heat required for this step is equal to moles $\times \Delta H_{\text{fus}}$. Step 3 is to heat the liquid from 0°C to 100°C . The heat required for this step is equal to mass \times specific heat capacity \times temperature change.

- a. Since the masses, heat capacities, and temperature changes for water and for substance X are all equal, the heat required for step 1 and step 3 are the same for both. Since ΔH_{fus} is larger for substance X (per mole), step 2 will require more heat for substance X and thus take longer. Therefore, H_2O will reach the boiling point first.
- b. To completely boil away the substance, an additional step is required. Step 4 is to boil the liquid to vapor at 100°C . The heat required for this step is equal to moles $\times \Delta H_{\text{vap}}$. Since the ΔH_{vap} values are much larger than the ΔH_{fus} values, step 4 will require much more heat than step 2 for both substance X and H_2O . Since ΔH_{vap} is smaller for substance X (per mole), step 4 will require less heat for substance X and thus take less time. The total heat required for the four steps is directly proportional to the time it would take to completely boil away the substance. Steps 1 and 3 are the same for both. Step 2 takes 1.5 times as long for substance X, but step 4 takes 2.0 times as long for water. Since step 4 requires the most heat, water will require more time to complete this step, so substance X will boil away first.
- c. The heating curves for substance X and for water are shown below.

Heating curve for water



Heating curve for X



- 11.23 The water the farmers spray above and on their fruit is warmer than the temperature of the fruit on the trees. Therefore, as the temperature of the air drops, it absorbs heat from the water, converting it into ice, before absorbing any heat from the fruit. The heat released when the liquid-to-solid phase change occurs prevents the fruit from freezing.
- 11.24a. Bottle A is most likely mislabeled. If it is an ionic compound, the boiling point should be higher than 35°C . Most ionic compounds are solids with high melting points.
- b. The substance with the highest boiling point will have the strongest intermolecular attractions. Thus, the compound in bottle C has the strongest intermolecular attractions.
- c. The substance with the lowest boiling point will have the highest vapor pressure. Thus, the substance in bottle B will have the highest vapor pressure.
- 11.25 a. Considering the A atoms, there are nine A atoms per cell. Of these, eight are in corners and contribute $1/8$ per cell, and one is completely inside the cell. Thus there are $8 \times (1/8) + 1 \times (1) = 2$ A atoms per cell. Next, considering the B atoms, there are six per cell. Of these, four are in faces and contribute $1/2$ per cell, and two are completely inside the cell. Thus, there are $4 \times (1/2) + 2 \times (1) = 4$ B atoms per unit cell. The ratio of A atoms to B atoms is two to four, or one to two. Thus, the formula of the compound is AB_2 .
- b. The A atoms are in the arrangement of a body-centered cell.
- 11.26 Consulting the phase diagram for water (Figure 11.11), you see that, by increasing the pressure on solid ice at constant temperature, you will convert solid ice into liquid water. The film of liquid that forms allow the skates and sleds to slide so well.
- 11.27 As the water evaporates, molecules with higher kinetic energy escape the liquid and leave behind molecules with lower energy. The result is a drop in temperature of the liquid. Since the cup is well insulated, the energy lost with the evaporated molecules is not rapidly replaced.
- 11.28 The heat released when the vapor of a substance condenses to liquid is equal to the negative (opposite) of the heat of vaporization for the substance. Due to its strong intermolecular attractions water has a larger heat of vaporization, so it releases more heat when condensing on the skin, causing a more severe burn.

- 11.29 This question can be answered by looking at the packing efficiencies of the three types of crystals, which are: simple cubic, 52.4 percent; body-centered cubic, 68 percent; and face-centered cubic, 74 percent.
- The highest density would correspond to the type of crystal with the highest packing efficiency. Of the three types of crystals, the face-centered cubic has the highest packing efficiency (74 percent) and thus has the highest density.
 - The most empty space corresponds to the type of crystal with the lowest packing efficiency. Of the three types of crystals, the simple cubic has the lowest packing efficiency (52.4 percent) and, therefore, the most empty space.
- 11.30 Since the liquids have comparable molar masses, differences in properties can be attributed to differences in intermolecular forces. The flask on the left has more molecules in the vapor phase than the flask on the right, and, therefore, the substance in this flask has a higher vapor pressure. Since vapor pressure decreases as intermolecular forces increase, the intermolecular forces for the substance in the flask on the right are stronger than for the substance in the flask on the left.
- Substance A has hydrogen bonding while substance B does not. Hydrogen bonding is a stronger intermolecular force than a dipole-dipole force. This implies that substance A with the lower vapor pressure is in the flask on the right.
 - The amount of vapor present at 35 °C (an increase of 15 °C) would be greater in both flasks compared to the amount present at 20 °C, but since the intermolecular forces of substance B are weaker, substance B will experience a larger increase.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

- 11.31 a. Vaporization
- Freezing of eggs and sublimation of ice
 - Condensation
 - Gas-solid condensation, deposition
 - Freezing

11.32 a. Sublimation

b. Vaporization

c. Sublimation of the filament and gas-solid condensation of the vapor

d. Freezing

e. Melting, fusion

11.33 Dropping a line from the intersection of a 350-mmHg line with the diethyl ether curve in Figure 11.7 intersects the temperature axis at about 10°C.

11.34 Dropping a line from the intersection of a 250-mmHg line with the carbon tetrachloride curve in Figure 11.7 intersects the temperature axis at about 40°C.

11.35 The total amount of energy provided by the heater in 4.54 min is

$$4.54 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{3.48 \text{ J}}{\text{s}} = 947.95 \text{ J} \text{ (0.94795 kJ)}$$

The heat of fusion per mole of I_2 is

$$\frac{0.9479 \text{ kJ}}{15.5 \text{ g I}_2} \times \frac{2 \times 126.9 \text{ g I}_2}{\text{mol I}_2} = 15.52 = \frac{15.5 \text{ kJ}}{\text{mol I}_2}$$

11.36 The total amount of energy provided by the heater in 6.92 min is

$$6.92 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{4.66 \text{ J}}{1 \text{ s}} = 1934.8 \text{ J} \text{ (1.9348 kJ)}$$

The heat of fusion per mole of Cd is

$$\frac{1.9348 \text{ kJ}}{15.5 \text{ g Cd}} \times \frac{112.4 \text{ g Cd}}{\text{mol Cd}} = 6.074 = \frac{6.07 \text{ kJ}}{\text{mol Cd}}$$

11.37 The heat absorbed per 2.25 g of isopropyl alcohol, $\text{C}_3\text{H}_8\text{O}$, is

$$2.25 \text{ g C}_3\text{H}_8\text{O} \times \frac{1 \text{ mol C}_3\text{H}_8\text{O}}{60.09 \text{ g C}_3\text{H}_8\text{O}} \times \frac{42.1 \text{ kJ}}{1 \text{ mol C}_3\text{H}_8\text{O}} = 1.576 = 1.58 \text{ kJ}$$

- 11.38 For 39.3 g of butane, C_4H_{10} , the heat needed is

$$39.3 \text{ g } C_4H_{10} \times \frac{1 \text{ mol } C_4H_{10}}{58.12 \text{ g } C_4H_{10}} \times \frac{21.3 \text{ kJ}}{1 \text{ mol } C_4H_{10}} = 14.\underline{40} = 14.4 \text{ kJ}$$

- 11.39 Because all the heat released by freezing the water is used to evaporate the remaining water, you must first calculate the amount of heat released in the freezing:

$$9.31 \text{ g } H_2O \times \frac{\text{mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{6.01 \text{ kJ}}{\text{mol } H_2O} = 3.10505 \text{ kJ}$$

Finally, calculate the mass of H_2O that was vaporized by the 3.10505 kJ of heat:

$$3.10505 \text{ kJ} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} \times \frac{1 \text{ mol } H_2O}{44.9 \text{ kJ}} = 1.246 = 1.25 \text{ g } H_2O$$

- 11.40 Enough ice must have been added so the heat consumed in melting the ice is equal to the heat released in cooling the water from 21.0°C to 0.0°C .

$$\begin{aligned} \text{Heat released by cooling} &= [(33.6 \text{ g})(0.0^\circ\text{C} - 21.0^\circ\text{C})(4.18 \text{ J/g}\cdot^\circ\text{C})] \\ &= -2.9494 \times 10^3 \text{ J} = -2.9494 \text{ kJ} \end{aligned}$$

Convert this heat (2.9494 kJ) to the mass of ice melted:

$$2.9494 \text{ kJ} \times \frac{1 \text{ mol } H_2O}{6.01 \text{ kJ}} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 8.8433 = 8.84 \text{ g } H_2O$$

- 11.41 Calculate how much heat is released by cooling 64.3 g of H_2O from 55°C to 15°C .

$$\begin{aligned} \text{Heat rel'd} &= (64.3 \text{ g})(15^\circ\text{C} - 55^\circ\text{C})\left(\frac{4.18 \text{ J}}{1 \text{ g}\cdot^\circ\text{C}}\right) \\ &= -1.07509 \times 10^4 \text{ J} = -10.7509 \text{ kJ} \end{aligned}$$

The heat released is used first to melt the ice and then to warm the liquid from 0°C to 15°C . Let the mass of ice equal y grams. Then, for fusion, and for warming, we have

$$\text{Fusion: } (y \text{ g } H_2O) \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{6.01 \text{ kJ}}{1 \text{ mol } H_2O} = 0.3335 y \text{ kJ}$$

(continued)

$$\text{Warming: } (y \text{ g H}_2\text{O})(15^\circ\text{C} - 0^\circ\text{C}) \left(\frac{4.18 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \right) = 62.70 y \text{ J } (0.06270 y \text{ kJ})$$

Because the total heat required for melting and warming must equal the heat released by cooling, equate the two, and solve for y .

$$10.7509 \text{ kJ} = 0.3335y \text{ kJ} + 0.0627y \text{ kJ} = y(0.3335 + 0.0627) \text{ kJ}$$

$$y = 10.7509 \text{ kJ} \div 0.3962 \text{ kJ} = 27.13 \text{ (grams)}$$

Thus, 27 g of ice were added.

- 11.42 If all the steam condensed, the quantity of heat used to warm the water in the flask must equal the heat released by condensation and cooling of the steam. Thus, first find the quantity of heat used to warm the water in the flask.

$$\begin{aligned} \text{Heat required} &= (275 \text{ g})(83^\circ\text{C} - 21^\circ\text{C})(4.18 \text{ J/(g} \cdot ^\circ\text{C)}) \\ &= 71.26 \times 10^3 \text{ J } (71.26 \text{ kJ}) \end{aligned}$$

Now let y = the mass of the steam. Then, find the quantity of heat released by condensation:

$$y \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{-40.7 \text{ kJ}}{1 \text{ mol H}_2\text{O}} = -2.259 y \text{ kJ}$$

Now find the quantity of heat released by cooling y grams of H_2O from 100°C to 83°C :

$$\text{Heat rel'd.} = y (83^\circ\text{C} - 100^\circ\text{C}) (4.18 \text{ J/(g} \cdot ^\circ\text{C)}) = -71.06y \text{ J } (-0.07106y \text{ kJ})$$

The total quantity of heat released in condensing and cooling the steam is equal in magnitude, but opposite in sign, to the quantity of heat required to warm the water in the flask:

$$-71.26 \text{ kJ} = -2.259y \text{ kJ} + (-0.07106y) \text{ kJ}$$

$$y \text{ g steam} = \frac{71.26 \text{ kJ}}{(2.259 + 0.07106) \text{ kJ}} = 30.58 = 31 \text{ g}$$

Thus, 31 g of steam condensed.

- 11.43 At the normal boiling point, the vapor pressure of a liquid is 760.0 mmHg. Use the Clausius-Clapeyron equation to find P_2 when $P_1 = 760.0$ mmHg, $T_1 = 334.85$ K (61.7°C), and $T_2 = 309.35$ K (36.2°C). Also use $\Delta H_{\text{vap}} = 31.4 \times 10^3$ J/mol.

$$\ln \frac{P_2}{P_1} = \frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln \frac{P_2}{760 \text{ mmHg}} = \frac{31.4 \times 10^3 \text{ J/mol}}{8.31 \text{ J/K} \cdot \text{mol}} \left(\frac{1}{334.85 \text{ K}} - \frac{1}{309.35 \text{ K}} \right) = -0.93018$$

Taking antilogs of both sides gives

$$\frac{P_2}{760 \text{ mmHg}} = e^{-0.93018} = 0.3944$$

$$P_2 = 0.3944 \times 760 \text{ mmHg} = 299.8 = 3.00 \times 10^2 \text{ mmHg (300. mmHg)}$$

- 11.44 At methanol's normal boiling point, its vapor pressure is 760.0 mmHg. Use the Clausius-Clapeyron equation to calculate P_2 when $P_1 = 760$ mmHg, $T_1 = 338.2$ K, and $T_2 = 295.2$ K. Also use $\Delta H_{\text{vap}} = 37.4 \times 10^3$ J/mol

$$\ln \frac{P_2}{P_1} = \frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln \frac{P_2}{760 \text{ mmHg}} = \frac{37.4 \times 10^3 \text{ J/mol}}{8.31 \text{ J/K} \cdot \text{mol}} \left(\frac{1}{338.2 \text{ K}} - \frac{1}{295.2 \text{ K}} \right) = -1.9384$$

Taking antilogs of both sides gives

$$\frac{P_2}{760 \text{ mmHg}} = e^{-1.9384} = 0.14393$$

$$P_2 = 0.14393 \times 760 \text{ mmHg} = 109.38 = 109 \text{ mmHg}$$

- 11.45 From the Clausius-Clapeyron equation,

$$\begin{aligned} \Delta H_{\text{vap}} &= R \left(\frac{T_2 T_1}{T_2 - T_1} \right) \left[\ln \frac{P_2}{P_1} \right] \\ &= [8.31 \text{ J/(K} \cdot \text{mol)}] \left[\frac{(553.2 \text{ K})(524.2 \text{ K})}{(553.2 - 524.2) \text{ K}} \right] \left[\ln \frac{760.0 \text{ mmHg}}{400.0 \text{ mmHg}} \right] \\ &= 5.3336 \times 10^4 \text{ J/mol} = 53.3 \text{ kJ/mol} \end{aligned}$$

11.46 From the Clausius-Clapeyron equation,

$$\begin{aligned}\Delta H_{\text{vap}} &= R \left(\frac{T_2 T_1}{T_2 - T_1} \right) \left[\ln \frac{P_2}{P_1} \right] \\ &= [8.31 \text{ J/(K}\cdot\text{mol)}] \left[\frac{(319.7 \text{ K})(301.2 \text{ K})}{(319.7 - 301.2) \text{ K}} \right] \left[\ln \frac{760.0 \text{ mmHg}}{400.0 \text{ mmHg}} \right] \\ &= 2.7762 \times 10^4 \text{ J/mol} = 27.8 \text{ kJ/mol}\end{aligned}$$

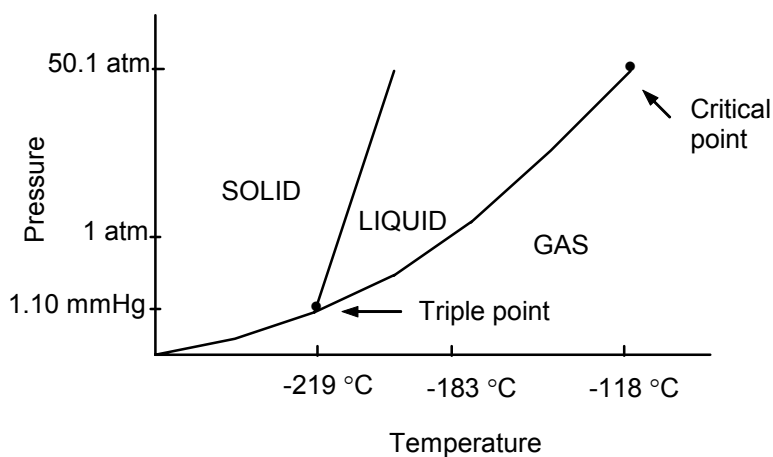
11.47 a. At point A, the substance will be a gas.

- b. The substance will be a gas.
- c. The substance will be a liquid.
- d. no

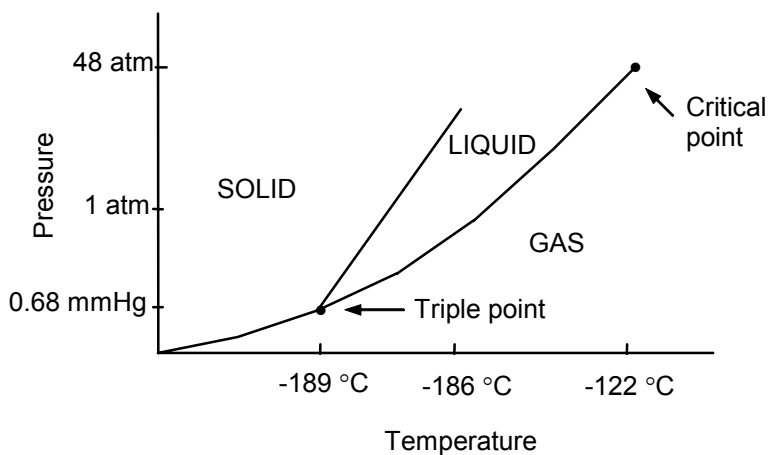
11.48 a. At point A, the substance will be a gas.

- b. The substance will be a solid.
- c. The substance will be a gas.
- d. yes

11.49 The phase diagram for oxygen is shown below. It is plotted from these points: triple point = -219°C , boiling point = -183°C , and critical point = -118°C .



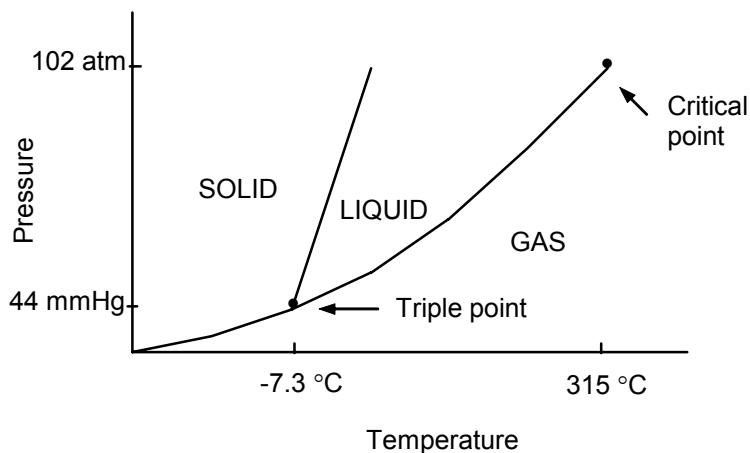
- 11.50 The phase diagram for argon below is plotted from these points: triple point = -189°C , boiling point = -186°C , and critical point = -122°C .



- 11.51 Liquefied at 25°C : SO_2 and C_2H_2 . To liquefy CH_4 , lower its temperature below -82°C , and then compress it. To liquefy CO , lower its temperature below -140°C , and then compress it.

- 11.52 a. If CF_4 is in the tank, it's not in liquid form because the liquid phase cannot exist above -46°C .
- b. If C_4H_{10} is in the tank, it's not in liquid form because 21°C is above its boiling point (1.0 atm).

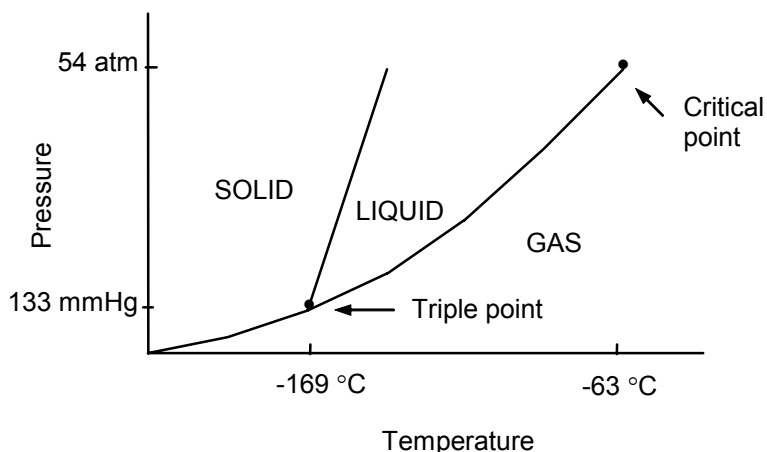
- 11.53 Br_2 phase diagram:



(continued)

- a. Circle "solid." The pressure of 40 mmHg is lower than the pressure at the triple point so the liquid phase cannot exist.
- b. Circle "liquid." The pressure of 400 mmHg is above the triple point so the gas will condense to a liquid.

11.54 Kr phase diagram:



- a. Circle "sublimes." The pressure of 130 mmHg is lower than the pressure at the triple point so the liquid phase cannot exist.
- b. Circle "melts." The pressure of 760 mmHg is higher than the pressure at the triple point so the solid melts to the liquid phase.

11.55 Yes, the heats of vaporization of 0.9, 5.6, and 20.4 kJ/mol (for H_2 , N_2 , and Cl_2 , respectively) increase in the order of the respective molecular weights of 2.016, 28.02, and 71.0. (London forces increase in order of increasing molecular weight.)

11.56 The heats of vaporization of 1.8 kJ/mol for Ne and 6.8 kJ/mol for O_2 increase in the order of the respective molecular weights of 20.1 and 32.0 because only London forces are involved. The heat of vaporization of 34.5 kJ/mol for methanol (molecular weight = 32.0) is higher than that of oxygen because of strong hydrogen bonding between H and O.

11.57 a. London forces

- b. London and dipole-dipole forces, H-bonding
- c. London and dipole-dipole forces
- d. London forces

- 11.58 Both a and d exhibit only London forces. The substance depicted in b has a dipole moment as well as London forces. Compound c is linear but is an ionic species. Thus, it has other forces present in addition to London forces.
- 11.59 The order is $\text{CCl}_4 < \text{SiCl}_4 < \text{GeCl}_4$ (in order of increasing molecular weight).
- 11.60 The order is $\text{He} < \text{Ar} < \text{Kr}$ (in order of increasing atomic weight).
- 11.61 CCl_4 has the lowest vapor pressure because it has the largest molecular weight and the greatest London forces even though HCCl_3 and H_3CCl have dipole-dipole interactions.
- 11.62 ClF has the smallest molecular weight and hence the smallest intermolecular forces and the highest vapor pressure. Thus, it should have the lowest boiling point.
- 11.63 The order of increasing vapor pressure is $\text{HOCH}_2\text{CH}_2\text{OH}$, $\text{FCH}_2\text{CH}_2\text{OH}$, $\text{FCH}_2\text{CH}_2\text{F}$. There is no hydrogen bonding in the third molecule; the second molecule can hydrogen-bond at only one end; and the first molecule can hydrogen-bond at both ends for the strongest interaction.
- 11.64 The order is $\text{HOCH}_2\text{CH}_2\text{CH}_2\text{OH} < \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH} < \text{CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3$. There is no hydrogen bonding in the third molecule; the second molecule can hydrogen-bond at only one end; and the first molecule can hydrogen-bond at both ends for the strongest interaction and lowest vapor pressure.
- 11.65 The order is $\text{CH}_4 < \text{C}_2\text{H}_6 < \text{CH}_3\text{OH} < \text{CH}_2\text{OHCH}_2\text{OH}$. The weakest forces are the London forces in CH_4 and C_2H_6 , which increase with molecular weight. The next strongest interaction is in CH_3OH , which can hydrogen-bond at only one end of the molecule. The strongest interaction is in the last molecule, which can hydrogen-bond at both ends.
- 11.66 The order is $\text{C}_2\text{H}_6 < \text{C}_3\text{H}_8 < (\text{CH}_3)_3\text{N} < \text{C}_4\text{H}_9\text{OH}$. The weakest forces are the London forces in C_2H_6 , C_3H_8 , and $(\text{CH}_3)_3\text{N}$, which increase with molecular weight. The strongest interaction is in $\text{C}_4\text{H}_9\text{OH}$, which can hydrogen-bond from oxygen to hydrogen, a stronger force than in the other three molecules.
- 11.67 a. Metallic b. Metallic c. Covalent network
 d. Molecular e. Ionic

- 11.68 a. Ionic
d. Molecular
- b. Ionic
e. Molecular
- c. Molecular
- 11.69 a. Metallic
c. Molecular
- b. Covalent network (like diamond)
d. Molecular
- 11.70 a. Not molecular (ionic)
c. Not molecular (metallic)
- b. Molecular
d. Molecular
- 11.71 The order is $(\text{C}_2\text{H}_5)_2\text{O} < \text{C}_4\text{H}_9\text{OH} < \text{KCl} < \text{CaO}$. Melting points increase in the order of attraction between molecules or ions in the solid state. Hydrogen bonding in $\text{C}_4\text{H}_9\text{OH}$ causes it to melt at a higher temperature than $(\text{C}_2\text{H}_5)_2\text{O}$. Both KCl and CaO are ionic solids with much stronger attraction than the organic molecules. In CaO, the higher charges cause the lattice energy to be higher than in KCl.
- 11.72 The order is $\text{C}_2\text{H}_6 < \text{CH}_3\text{OH} < \text{NaCl} < \text{Si}$. Melting points increase in the order of attraction between molecules or atoms. Hydrogen bonding in CH_3OH causes it to melt at a higher temperature than C_2H_6 . Because NaCl is an ionic solid, it melts at a higher temperature than either of the previous two. Silicon is a covalent network solid with the highest melting point.
- 11.73 a. Low-melting and brittle
c. Malleable and electrically conducting
- b. High-melting, hard, and brittle
d. Hard and high-melting
- 11.74 a. Metallic (from conductivity and luster)
b. Covalent network (from high-melting, hard, and nonconducting liquid)
c. Ionic (from high melting point and conducting liquid)
d. Molecular (from low melting point and odor or vapor pressure at room temperature)
- 11.75 a. LiCl
b. SiC
c. CH_3I
d. Co
- 11.76 a. Pb
b. CaCl_2
c. P_4S_3
d. BN
- 11.77 In a simple cubic lattice with one atom at each lattice point, there are atoms only at the corners of unit cells. Each corner is shared by eight unit cells, and there are eight corners per unit cell. Therefore, there is one atom per unit cell.

- 11.78 There are two atoms per unit cell, one from the corners and one atom at the center of the unit cell.

- 11.79 Calculate the volume of the unit cell, change density to g/m^3 , and then convert volume to mass, using density:

$$\text{Volume} = (2.866 \times 10^{-10} \text{ m})^3 = 2.354 \times 10^{-29} \text{ m}^3$$

$$\frac{7.87 \text{ g}}{1 \text{ cm}^3} \times \left(\frac{100 \text{ cm}}{1 \text{ m}} \right)^3 = 7.87 \times 10^6 \text{ g}/\text{m}^3$$

$$\text{Mass of one cell} = (7.87 \times 10^6 \text{ g}/\text{m}^3) \times (2.354 \times 10^{-29} \text{ m}^3) = 1.8526 \times 10^{-22} \text{ g}$$

Because Fe is a body-centered cubic cell, there are two Fe atoms in the cell, and

$$\text{Mass of one Fe atom} = (1.8526 \times 10^{-22} \text{ g}) \div 2 = 9.263 \times 10^{-23} \text{ g}$$

Using the molar mass to calculate the mass of one Fe atom, you find the agreement is good:

$$\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \times \frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}} = 9.2743 \times 10^{-23} \text{ g/Fe atom}$$

- 11.80 Calculate the volume of the unit cell, change density to g/m^3 , and then convert volume to mass, using density:

$$\text{Volume} = (3.524 \times 10^{-10} \text{ m})^3 = 4.376 \times 10^{-29} \text{ m}^3$$

$$\frac{8.91 \text{ g}}{1 \text{ cm}^3} \times \left(\frac{100 \text{ cm}}{1 \text{ m}} \right)^3 = 8.91 \times 10^6 \text{ g}/\text{m}^3$$

$$\text{Mass of one cell} = (8.91 \times 10^6 \text{ g}/\text{m}^3) \times (4.376 \times 10^{-29} \text{ m}^3) = 3.899 \times 10^{-22} \text{ g}$$

Because Ni is a face-centered unit cell, there are four Ni atoms per cell, and

$$\text{Mass of one Ni atom} = (3.899 \times 10^{-22} \text{ g}) \div 4 = 9.747 \times 10^{-23} \text{ g}$$

Using the molar mass to calculate Avogadro's number, you obtain

$$\frac{58.70 \text{ g Ni}}{1 \text{ mol Ni}} \times \frac{1 \text{ atom Ni}}{9.747 \times 10^{-23} \text{ g Ni}} = 6.022 \times 10^{23} \text{ atoms/mol}$$

11.81 There are four Cu atoms in the face-centered cubic structure, so the mass of one cell is

$$4 \text{ Cu atoms} \times \frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} \times \frac{63.5 \text{ g Cu}}{1 \text{ mol Cu}} = 4.218 \times 10^{-22} \text{ g}$$

$$\text{Cell volume} = \frac{4.218 \times 10^{-22} \text{ g}}{8.93 \text{ g/cm}^3} = 4.723 \times 10^{-23} \text{ cm}^3$$

All edges are the same length in a cubic cell, so the edge length, l , is

$$l = \sqrt[3]{V} = \sqrt[3]{4.723 \times 10^{-23} \text{ cm}^3} = 3.614 \times 10^{-8} \text{ cm} \\ = 3.61 \times 10^{-8} \text{ cm} \quad (361 \text{ pm})$$

11.82 There are two Ba atoms in the face-centered cubic structure, so the mass of one cell is

$$2 \text{ Ba atoms} \times \frac{1 \text{ mol Ba}}{6.022 \times 10^{23} \text{ Ba atoms}} \times \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 4.561 \times 10^{-22} \text{ g}$$

$$\text{Cell volume} = \frac{4.561 \times 10^{-22} \text{ g}}{3.51 \text{ g/cm}^3} = 1.2994 \times 10^{-22} \text{ cm}^3$$

All edges are the same length in a cubic cell, so the edge length, l , is

$$l = \sqrt[3]{V} = \sqrt[3]{1.2994 \times 10^{-22} \text{ cm}^3} = 5.0651 \times 10^{-8} \text{ cm} \\ = 5.07 \times 10^{-8} \text{ cm} \quad (507 \text{ pm})$$

11.83 Calculate the volume from the edge length of 407.9 pm ($4.079 \times 10^{-8} \text{ cm}$), and then use it to calculate the mass of the unit cell:

$$\text{Cell volume} = (4.079 \times 10^{-8} \text{ cm})^3 = 6.7869 \times 10^{-23} \text{ cm}^3$$

$$\text{Cell mass} = (19.3 \text{ g/cm}^3)(6.7869 \times 10^{-23} \text{ cm}^3) = 1.3098 \times 10^{-21} \text{ g}$$

Calculate the mass of one gold atom:

$$1 \text{ Au atom} \times \frac{1 \text{ mol Au}}{6.022 \times 10^{23} \text{ Au atoms}} \times \frac{196.97 \text{ g Au}}{1 \text{ mol Au}} = 3.2708 \times 10^{-22} \text{ g}$$

(continued)

$$\frac{1.3098 \times 10^{-21} \text{ g}}{1 \text{ unit cell}} \times \frac{1 \text{ Au atom}}{3.2708 \times 10^{-22} \text{ g Au}} = \frac{4.004 \text{ Au atoms}}{\text{unit cell}}$$

Since there are four atoms per unit cell, it is a face-centered cubic.

- 11.84 Calculate the volume from the edge ($288.5 \text{ pm} = 2.885 \times 10^{-8} \text{ cm}$). Use it to calculate the mass:

$$\text{Cell volume} = (2.885 \times 10^{-8} \text{ cm})^3 = 2.401 \times 10^{-23} \text{ cm}^3$$

$$\text{Cell mass} = (7.20 \text{ g/cm}^3)(2.401 \times 10^{-23} \text{ cm}^3) = 1.729 \times 10^{-22} \text{ g}$$

Calculate the mass of one chromium atom:

$$1 \text{ Cr atom} \times \frac{1 \text{ mol Cr}}{6.022 \times 10^{23} \text{ Au atoms}} \times \frac{51.996 \text{ g Cr}}{1 \text{ mol Cr}} = 8.634 \times 10^{-23} \text{ g Cr}$$

$$\frac{1.729 \times 10^{-22} \text{ g}}{1 \text{ unit cell}} \times \frac{1 \text{ Cr atom}}{8.634 \times 10^{-23} \text{ g Cr}} \times \frac{2.002 \text{ Cr atoms}}{1 \text{ unit cell}}$$

Since there are two atoms per unit cell, it is a body-centered cubic.

- 11.85 Calculate the volume from the edge ($316.5 \text{ pm} = 3.165 \times 10^{-8} \text{ cm}$). Use it to calculate the mass:

$$\text{Cell volume} = (3.165 \times 10^{-8} \text{ cm})^3 = 3.1705 \times 10^{-23} \text{ cm}^3$$

For a body-centered cubic lattice, there are two atoms per cell, so their mass is

$$2 \text{ W atoms} \times \frac{1 \text{ mol W}}{6.022 \times 10^{23} \text{ W atoms}} \times \frac{183.8 \text{ g W}}{1 \text{ mol W}} = 6.1043 \times 10^{-22} \text{ g W}$$

$$\text{Density} = \frac{6.1043 \times 10^{-22} \text{ g W}}{3.1705 \times 10^{-23} \text{ cm}^3} = 19.253 = 19.25 \text{ g/cm}^3$$

- 11.86 Calculate the volume from the edge ($495.0 \text{ pm} = 4.950 \times 10^{-8} \text{ cm}$). Use it to calculate the mass:

$$\text{Cell volume} = (4.950 \times 10^{-8} \text{ cm})^3 = 1.21\bar{2}9 \times 10^{-22} \text{ cm}^3$$

For a face-centered cubic lattice, there are four atoms per cell, so their mass is

$$4 \text{ Pb atoms} \times \frac{1 \text{ mol Pb}}{6.022 \times 10^{23} \text{ Pb atoms}} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 1.3763 \times 10^{-21} \text{ g Pb}$$

$$\text{Density} = \frac{1.3763 \times 10^{-21} \text{ g Pb}}{1.2129 \times 10^{-22} \text{ cm}^3} = 11.3\bar{4}7 = 11.35 \text{ g/cm}^3$$

- 11.87 Use Avogadro's number to calculate the number of atoms in 1.74 g ($= d \times 1.000 \text{ cm}^3$):

$$\begin{aligned} 1.74 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.305 \text{ g Mg}} \times \frac{6.022 \times 10^{23} \text{ Mg atoms}}{1 \text{ mol Mg}} \\ = 4.311 \times 10^{22} \text{ Mg atoms} \end{aligned}$$

Because the space occupied by the Mg atoms $= 0.741 \text{ cm}^3$, each atom's volume is

$$\text{Volume 1 Mg atom} = \frac{0.741 \text{ cm}^3}{4.311 \times 10^{22} \text{ Mg atoms}} = 1.7\bar{1}9 \times 10^{-23} \text{ cm}^3$$

$$\text{Volume} = \frac{4\pi r^3}{3}$$

so

$$r = \sqrt[3]{\frac{3V}{4\pi}}$$

$$\begin{aligned} r &= \sqrt[3]{\frac{3}{4\pi}(1.719 \times 10^{-23} \text{ cm}^3)} = 1.6\bar{0}1 \times 10^{-8} = 1.60 \times 10^{-8} \text{ cm} \\ &= 1.60 \times 10^2 \text{ pm} \end{aligned}$$

- 11.88 Use Avogadro's number to calculate the number of atoms in 3.51 g ($= d \times 1.000 \text{ cm}^3$):

$$3.51 \text{ g Ba} \times \frac{1 \text{ mol Ba}}{137.33 \text{ g Ba}} \times \frac{6.022 \times 10^{23} \text{ Ba atoms}}{1 \text{ mol Ba}}$$

$$= 1.539 \times 10^{22} \text{ Ba atoms}$$

Because the space occupied by the Ba atoms $= 0.680 \text{ cm}^3$, each atom's volume is

$$\text{Volume one Ba atom} = \frac{0.680 \text{ cm}^3}{1.539 \times 10^{22} \text{ Ba atoms}} = 4.418 \times 10^{-23} \text{ cm}^3$$

$$\text{Volume} = \frac{4\pi r^3}{3}$$

so

$$r = \sqrt[3]{\frac{3V}{4\pi}}$$

$$r = \sqrt[3]{\frac{3}{4\pi}(4.418 \times 10^{-23} \text{ cm}^3)} = 2.193 \times 10^{-8} = 2.19 \times 10^{-8} \text{ cm} = 219 \text{ pm}$$

■ Solutions to General Problems

- 11.89 Water vapor deposits directly to solid water (frost) without forming liquid water. After heating, most of the frost melted to liquid water, which then vaporized to water vapor. Some of the frost may have sublimed directly to water vapor.
- 11.90 Water vapor condenses directly to solid water (snow) in the upper atmosphere. After falling through the warm air mass, the snow melts to liquid water (rain). After falling on a sunny spot, the rain is vaporized to water vapor.
- 11.91 From Table 5.5, the vapor pressures are 18.7 mmHg at 21 °C and 12.8 mmHg at 15 °C. If the moisture did not begin to condense until the air had been cooled to 15 °C, then the partial pressure of water in the air at 21 °C must have been 12.8 mmHg. The relative humidity is

$$\text{Percent relative humidity} = \frac{12.8 \text{ mmHg}}{18.7 \text{ mmHg}} \times 100\% = 68.44 = 68.4 \text{ percent}$$

11.92 The vapor pressure of water at 21 °C is 18.7 mmHg (Table 5.5). Therefore,

$$\text{Percent relative humidity} = 58 \text{ percent} = \frac{x \text{ mmHg}}{18.7 \text{ mmHg}} \times 100\%;$$

$$x = 10.8 \text{ mmHg}$$

The partial pressure of water in the air at 21 °C is thus 10.8 mmHg. The water will begin to condense when the temperature drops to the temperature at which the vapor pressure of water is 10.8 mmHg. This temperature is between 12 °C and 13 °C (Table 5.5).

11.93 After labeling the problem data as below, use the Clausius-Clapeyron equation to obtain ΔH_{vap} , which can then be used to calculate the boiling point.

At $T_1 = 299.3 \text{ K}$, $P_1 = 100.0 \text{ mmHg}$; at $T_2 = 333.8 \text{ K}$, $P_2 = 400.0 \text{ mmHg}$

$$\ln \frac{400.0 \text{ mmHg}}{100.0 \text{ mmHg}} = \frac{\Delta H_{\text{vap}}}{8.31 \text{ J/(K} \cdot \text{mol)}} \left[\frac{333.8 \text{ K} - 299.3 \text{ K}}{333.8 \text{ K} \times 299.3 \text{ K}} \right]$$

$$1.3862 = \Delta H_{\text{vap}} (4.1535 \times 10^{-5} \text{ mol/J})$$

$$\Delta H_{\text{vap}} = 33.37 \times 10^3 \text{ J/mol (33.4 kJ/mol)}$$

Now, use this value of ΔH_{vap} and the following data to calculate the boiling point:

At $T_1 = 299.3 \text{ K}$, $P_1 = 100.0 \text{ mmHg}$; at T_2 (boiling pt.), $P_2 = 760 \text{ mmHg}$

$$\ln \frac{760.0 \text{ mmHg}}{100.0 \text{ mmHg}} = \frac{33.4 \times 10^3 \text{ J/mol}}{8.314 \text{ J/(K} \cdot \text{mol)}} \left[\frac{1}{299.3 \text{ K}} - \frac{1}{T_2} \right]$$

$$2.0281 = 4.0144 \times 10^3 \text{ K} \left[\frac{1}{299.3 \text{ K}} - \frac{1}{T_2} \right]$$

$$\frac{1}{T_2} = \frac{1}{299.3 \text{ K}} - \frac{2.0281}{4.0144 \times 10^3 \text{ K}} = 2.8359 \times 10^{-3} / \text{K}$$

$$T_2 = 352.6 = 353 \text{ K (80 °C)}$$

- 11.94 After labeling the problem data as below, use the Clausius-Clapeyron equation to obtain ΔH_{vap} , which can then be used to calculate the boiling point.

At $T_1 = 293.2 \text{ K}$, $P_1 = 17.5 \text{ mmHg}$; at $T_2 = 353.2 \text{ K}$, $P_2 = 355.1 \text{ mmHg}$

$$\ln \frac{355.1 \text{ mmHg}}{17.5 \text{ mmHg}} = \frac{\Delta H_{\text{vap}}}{8.31 \text{ J/(K} \cdot \text{mol)}} \left[\frac{353.2 \text{ K} - 293.2 \text{ K}}{353.2 \text{ K} \times 293.2 \text{ K}} \right]$$

$$3.0101 = \Delta H_{\text{vap}} (6.9687 \times 10^{-5} \text{ mol/J})$$

$$\Delta H_{\text{vap}} = 43.19 \times 10^3 \text{ J/mol (43.2 kJ/mol)}$$

Now, use this value of ΔH_{vap} and the following data to calculate the boiling point:

At $T_1 = 293.32 \text{ K}$, $P_1 = 17.5 \text{ mmHg}$; at T_2 (boiling pt.), $P_2 = 760 \text{ mmHg}$

$$\ln \frac{760.0 \text{ mmHg}}{17.5 \text{ mmHg}} = \frac{43.2 \times 10^3 \text{ J/mol}}{8.314 \text{ J/(K} \cdot \text{mol)}} \left[\frac{1}{293.2 \text{ K}} - \frac{1}{T_2} \right]$$

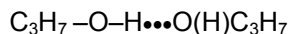
$$3.7711 = 5.1955 \times 10^3 \text{ K} \left[\frac{1}{293.2 \text{ K}} - \frac{1}{T_2} \right]$$

$$\frac{1}{T_2} = \frac{1}{293.2 \text{ K}} - \frac{3.7711}{5.1955 \times 10^3 \text{ K}} = 2.6848 \times 10^{-3} / \text{K}$$

$$T_2 = 372.46 = 372 \text{ K (99 } ^\circ\text{C)}$$

- 11.95 a. As this gas is compressed at 20°C , it will condense into a liquid because 20°C is above the triple point but below the critical point.
- b. As this gas is compressed at -70°C , it will condense directly to the solid phase because the temperature of -70°C is below the triple point.
- c. As this gas is compressed at 40°C , it will not condense because 40°C is above the critical point.
- 11.96 a. As I_2 vapor is cooled at 120 atm, no change to a distinct liquid will be observed, but the I_2 will condense to a solid phase at some definite temperature.
- b. As I_2 vapor is cooled at one atm, the vapor will condense to a liquid and then freeze to the solid phase.
- c. As I_2 vapor is cooled at 50 mmHg (below the triple point), the vapor will condense directly to the solid phase without going through the liquid phase.

- 11.97 In propanol, hydrogen bonding exists between the hydrogen of the OH group and the lone pair of electrons of oxygen of the OH group of an adjacent propanol molecule. For two adjacent propanol molecules, the hydrogen bond may be represented as follows:



- 11.98 In hydrogen peroxide, hydrogen bonding exists between any hydrogen and the lone pair of electrons of oxygen of an adjacent hydrogen peroxide. For two adjacent hydrogen peroxide molecules, the hydrogen bond may be represented as follows:



- 11.99 Ethylene glycol molecules are capable of hydrogen bonding to each other, whereas pentane molecules are not. The greater intermolecular forces in ethylene glycol are reflected in greater resistance to flow (viscosity) and high boiling point.
- 11.100 Pentylamine molecules are capable of hydrogen bonding to each other, but triethylamine molecules are not. The greater intermolecular forces in pentylamine cause a higher boiling point and greater resistance to flow.
- 11.101 Aluminum (Group IIIA) forms a metallic solid. Silicon (Group IVA) forms a covalent network solid. Phosphorus (Group VA) forms a molecular solid. Sulfur (Group VIA) forms a molecular (amorphous) solid.
- 11.102 AlF_3 forms an ionic solid. SiF_4 , PF_3 , and SF_4 form molecular solids.
- 11.103a. Lower: KCl. The lattice energy should be lower for ions with a lower charge. A lower lattice energy implies a lower melting point.
- b. Lower: CCl_4 . Both are molecular solids, so the compound with the lower molecular weight should have weaker London forces and, therefore, the lower melting point.
- c. Lower: Zn. Melting points for Group IIB metals are lower than for metals near the middle of the transition-metal series.
- d. Lower: $\text{C}_2\text{H}_5\text{Cl}$. Ethyl chloride cannot hydrogen-bond, but acetic acid can. The compound with the weaker intermolecular forces has the lower melting point.
- 11.104a. Lower: C_6H_{14} . A molecular solid has a lower melting point than an ionic solid.
- b. Lower: 1-propanol. The 1-propanol can hydrogen-bond at only one end; ethylene glycol can hydrogen-bond at both ends so its intermolecular forces are stronger.

- c. Lower: Na. Sodium is a metallic solid, but Si is a covalent network solid (high melting point).
- d. Lower: CH₄. Both form molecular solids, but CH₄ has a lower molecular weight with weaker London forces and, therefore, the lower melting point.

11.105 The face-centered cubic structure means one atom is at each lattice point. All edges are the same length in such a structure, so the volume is

$$\text{Volume} = l^3 = (3.839 \times 10^{-8} \text{ cm})^3 = 5.6579 \times 10^{-23} \text{ cm}^3$$

$$\text{Mass of unit cell} = dV = (22.42 \text{ g/cm}^3)(5.6579 \times 10^{-23} \text{ cm}^3) = 1.2685 \times 10^{-21} \text{ g}$$

There are four atoms in a face-centered cubic cell, so

$$\begin{aligned} \text{Mass of one Ir atom} &= \text{mass of unit cell} \div 4 = (1.2685 \times 10^{-21} \text{ g}) \div 4 \\ &= 3.1712 \times 10^{-22} \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Molar mass of Ir} &= (3.1712 \times 10^{-22} \text{ g/Ir atom}) \times (6.022 \times 10^{23} \text{ Ir atoms/mol}) \\ &= 190.96 = 191.0 \text{ g/mol} \quad (\text{The atomic weight} = 191.0 \text{ amu.}) \end{aligned}$$

11.106 The body-centered cubic structure means that one atom is at each lattice point. All edges are the same length in such a structure, so the volume is

$$\text{Volume} = l^3 = (3.306 \times 10^{-8} \text{ cm})^3 = 3.6133 \times 10^{-23} \text{ cm}^3$$

$$\text{Mass of unit cell} = dV = (16.69 \text{ g/cm}^3)(3.6133 \times 10^{-23} \text{ cm}^3) = 6.0306 \times 10^{-22} \text{ g}$$

There are two atoms in a body-centered cubic cell, so

$$\begin{aligned} \text{Mass of one Ta atom} &= \text{mass of unit cell} \div 2 = (6.0306 \times 10^{-22} \text{ g}) \div 2 \\ &= 3.0153 \times 10^{-22} \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Molar mass of Ta} &= (3.0153 \times 10^{-22} \text{ g/Ta atom}) \times (6.022 \times 10^{23} \text{ Ta atoms/mol}) \\ &= 181.58 = 181.6 \text{ g/mol} \quad (\text{The atomic weight} = 181.6 \text{ amu.}) \end{aligned}$$

- 11.107 From Problem 11.81, the cell edge length (l) is 361.4 pm. There are four copper atom radii along the diagonal of a unit-cell face. Because the diagonal square = $l^2 + l^2$ (Pythagorean theorem),

$$4r = \sqrt{2} l^2 = \sqrt{2} l, \text{ or } r = \frac{\sqrt{2}}{4} (361.4 \text{ pm}) = 127.7 = 128 \text{ pm}$$

- 11.108 A body-centered cubic cell has two atoms per cell. The mass of the unit cell is

$$2 \text{ Rb atoms} \times \frac{1 \text{ mol Rb}}{6.022 \times 10^{23} \text{ Rb atoms}} \times \frac{85.468 \text{ g Rb}}{1 \text{ mol Rb}} = 2.8385 \times 10^{-22} \text{ g}$$

The volume of the unit cell is

$$\text{Volume} = \frac{2.8385 \times 10^{-22} \text{ g}}{1.532 \text{ g/cm}^3} = 1.8528 \times 10^{-22} \text{ cm}^3$$

$$l = \sqrt[3]{\text{Volume}} = \sqrt[3]{1.8528 \times 10^{-22} \text{ cm}^3} = 5.7009 \times 10^{-8} \text{ cm}$$

Because the corner spheres touch the body-centered sphere, the length of the body diagonal (diagonal passing through the center of the cell) must be four times the radius of the Rb atom. Also, from the geometry of a cube and the Pythagorean theorem, the square of the body diagonal equals $l^2 + d^2$ (l is the unit-cell edge length, and d is the diagonal along a face of the unit cell).

Because $d^2 = l^2 + l^2$, or $d = \sqrt{2} l$, you can write

$$(\text{Body diagonal})^2 = l^2 + d^2 = l^2 + 2l^2 = 3l^2$$

$$\text{Body diagonal} = \sqrt{3} l = \sqrt{3} (5.7009 \times 10^{-8} \text{ cm}) = 9.8742 \times 10^{-8} \text{ cm}$$

$$\text{Radius of Rb atom} = (9.8742 \times 10^{-8} \text{ cm}) \div 4 = 2.4686 \times 10^{-8} \text{ cm} (246.9 \text{ pm})$$

- 11.109 The body diagonal (diagonal passing through the center of the cell) is four times the radius, r , of a sphere. Also, from the geometry of a cube and the Pythagorean theorem, the body diagonal equals $\sqrt{3} l$, where l is the edge length of the unit cell. Thus

$$4r = \sqrt{3} (l)$$

or

$$l = \frac{4r}{\sqrt{3}}$$

(continued)

Because the unit cell contains two spheres, the volume occupied by the spheres is

$$V_{\text{spheres}} = 2 \times \frac{4}{3} \pi r^3$$

and

$$V_{\text{cell}} = l^3 = \left[\frac{4r}{\sqrt{3}} \right]^3 = \frac{64r^3}{3\sqrt{3}}$$

Finally, to obtain the percent volume of the cell occupied, divide V_{spheres} by V_{cell} :

$$\begin{aligned} \text{Percent V} &= \frac{V_{\text{spheres}}}{V_{\text{cell}}} \times 100\% = \frac{2 \left[\frac{4\pi r^3}{3} \right]}{\frac{64r^3}{3\sqrt{3}}} \times 100\% = \frac{\pi\sqrt{3}}{8} \times 100\% \\ &= 68.01 = 68 \text{ percent} \end{aligned}$$

11.110 Because the spheres touch along the diagonal of a face, d , the radius, r , of the spheres is

$$r = d/4 = l(\sqrt{2})/4$$

or

$$l = \frac{4r}{\sqrt{2}}$$

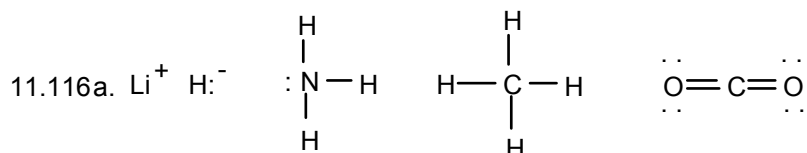
$$V_{\text{cell}} = l^3 = \left[\frac{4r}{\sqrt{2}} \right]^3$$

For a face-centered cubic structure, there are four spheres per cell, so the volume, V_{spheres} , occupied by the spheres is

$$V_{\text{spheres}} = 4 \left[\frac{4\pi r^3}{3} \right]$$

$$\begin{aligned} \text{Percent V} &= \frac{V_{\text{spheres}}}{V_{\text{cell}}} \times 100\% = \frac{4 \left[\frac{4\pi r^3}{3} \right]}{\left[\frac{4r}{\sqrt{2}} \right]^3} \times 100\% = \frac{\pi\sqrt{2}}{6} \times 100\% \\ &= 74.04 = 74 \text{ percent} \end{aligned}$$

- 11.111 a. The boiling point increases as the size (number of electrons in the atom or molecule) increases. The London forces or dispersion forces increase.
- b. Hydrogen bonding occurs between the H–F molecules, and is much stronger than the London forces.
- c. In addition to the dispersion forces, the hydrogen halides are polar (have dipole moments) so there are dipole-dipole interactions.
- 11.112 a. The boiling point increases as the size (number of electrons in the molecule) increases. The London forces or dispersion forces increase.
- b. Hydrogen bonding occurs among the NH_3 molecules.
- c. The nitrogen family compounds consist of polar molecules whereas the carbon family molecules are nonpolar. So, the nitrogen family has dipole-dipole interactions as well as the contributions from the dispersion forces.
- 11.113 a. Diamond and silicon carbide are giant molecules with strong covalent bonds between all the atoms. Graphite is a layered structure, and the forces holding the layers together are weak dispersion forces.
- b. Silicon dioxide is a giant molecule with an infinite array of O–Si–O bonds. Each silicon is bonded to four oxygen atoms in a covalent network solid. Carbon dioxide is a discrete, nonpolar, molecule.
- 11.114. The first member atoms are very small and highly electronegative. Oxygen readily forms double bonds, whereas the larger atoms have difficulty getting close enough together to form multiple bonds. So, we have O_2 vs. S_8 .
- Because oxygen has a higher attraction for hydrogen, the dipole moment is so large that it is possible to have hydrogen bonding in H_2O , but there is no hydrogen bonding in H_2S .
- 11.115 a. CO_2 consists of discrete non-polar molecules that are held together in the solid by weak dispersion forces. SiO_2 is a giant molecule with all the atoms held together by strong covalent bonds.
- b. $\text{HF}(\text{l})$ has extensive hydrogen bonding among the molecules. $\text{HCl}(\text{l})$ boils much lower because it doesn't have H-bonding.
- c. SiF_4 is a larger molecule (it has more electrons than CF_4), so it has stronger dispersion forces and a higher boiling point than CF_4 . Both molecules are tetrahedrally symmetrical and, therefore, non-polar.



- b. LiH has the highest boiling point because of the ionic bonding and crystalline lattice of LiH.
- c. CH_4 has the lowest boiling point because it is a small nonpolar molecule. The only intermolecular forces are dispersion forces.
- d. NH_3 is a polar molecule which also has H-bonding between the molecules.

■ Solutions to Cumulative-Skills Problems

11.117 Use the ideal gas law to calculate n , the number of moles of N_2 :

$$\text{N}_2: n = \frac{PV}{RT} = \frac{(745/760 \text{ atm})(5.40 \text{ L})}{(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(293 \text{ K})} = 0.2201 \text{ mol}$$

$$\text{C}_3\text{H}_8\text{O}: n = 0.6149 \text{ g C}_3\text{H}_8\text{O} \times \frac{1 \text{ mol C}_3\text{H}_8\text{O}}{60.094 \text{ g C}_3\text{H}_8\text{O}} = 0.010232 \text{ mol}$$

$$X_{\text{C}_3\text{H}_8\text{O}} = \frac{0.010232 \text{ mol C}_3\text{H}_8\text{O}}{(0.010232 \text{ mol} + 0.2201 \text{ mol})} = 0.04442 \text{ mole fraction}$$

$$\text{Partial } P = 0.04442 \times 745 \text{ mmHg} = 33.093 \text{ mmHg} = 33.1 \text{ mmHg}$$

$$\text{Vapor pressure of C}_3\text{H}_8\text{O} = 33.1 \text{ mmHg}$$

11.118 Use the ideal gas law to calculate n , the number of moles of N_2 :

$$\text{N}_2: n = \frac{PV}{RT} = \frac{(768/760 \text{ atm})(6.35 \text{ L})}{(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})} = 0.2624 \text{ mol}$$

$$\text{C}_3\text{H}_6\text{O}: n = 6.550 \text{ g C}_3\text{H}_6\text{O} \times \frac{1 \text{ mol C}_3\text{H}_6\text{O}}{58.05 \text{ g C}_3\text{H}_6\text{O}} = 0.11283 \text{ mol}$$

(continued)

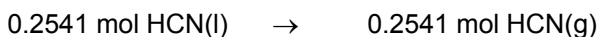
$$X_{\text{C}_3\text{H}_6\text{O}} = \frac{0.11283 \text{ mol C}_3\text{H}_6\text{O}}{(0.11283 \text{ mol} + 0.2624 \text{ mol})} = 0.3007 \text{ mole fraction}$$

$$\text{Partial P} = 0.3007 \times 768 \text{ mmHg} = 230.9 \text{ mmHg} = 231 \text{ mmHg}$$

$$\text{Vapor pressure of C}_3\text{H}_6\text{O} = 231 \text{ mmHg}$$

11.119 Calculate the moles of HCN in 10.0 mL of the solution (density = 0.687 g HCN/mL HCN):

$$10.0 \text{ mL HCN} \times \frac{0.687 \text{ g HCN}}{1 \text{ mL HCN}} \times \frac{1 \text{ mol HCN}}{27.03 \text{ g HCN}} = 0.2541 \text{ mol HCN}$$

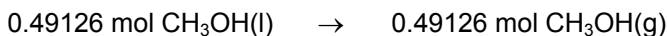


$$(\Delta H_f^\circ = 105 \text{ kJ/mol}) \quad (\Delta H_f^\circ = 135 \text{ kJ/mol})$$

$$\Delta H^\circ = 0.2541 \text{ mol} \times [135 \text{ kJ/mol} - 105 \text{ kJ/mol}] = 7.623 = 7.6 \text{ kJ}$$

11.120 Calculate moles of CH₃OH in the 20.0-mL solution (density = 0.787 g CH₃OH/mL CH₃OH):

$$20.0 \text{ mL CH}_3\text{OH} \times \frac{0.787 \text{ g CH}_3\text{OH}}{1 \text{ mL CH}_3\text{OH}} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \\ = 0.49126 \text{ mol CH}_3\text{OH}$$



$$(\Delta H_f^\circ = -238.6 \text{ kJ/mol}) \quad (\Delta H_f^\circ = -201.2 \text{ kJ/mol})$$

$$\Delta H^\circ = 0.49126 \text{ mol} \times [-201.2 \text{ kJ/mol} - (-238.6 \text{ kJ/mol})] = 18.37 = 18.4 \text{ kJ}$$

11.121 First, convert the mass to moles; then, multiply by the standard heat of formation to obtain the heat absorbed in vaporizing this mass:

$$12.5 \text{ g P}_4 \times \frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} = 0.1009 \text{ mol P}_4$$

$$0.1009 \text{ mol P}_4 \times \frac{95.4 \text{ J}}{^\circ\text{C} \cdot \text{mol P}_4} \times (44.1 ^\circ\text{C} - 25.0 ^\circ\text{C})$$

$$= 183.86 \text{ J} = 0.18386 \text{ kJ}$$

$$2.63 \text{ kJ/mol P}_4 \times 0.1009 \text{ mol P}_4 = 0.26537 \text{ kJ}$$

$$\text{Total heat} = 0.26537 \text{ kJ} + 0.18386 \text{ kJ} = 0.44923 = 0.449 \text{ kJ} = 449 \text{ J}$$

- 11.122 First, convert the mass to moles; then, multiply by the standard heat of formation to obtain the heat absorbed in vaporizing this mass:

$$25.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 1.0874 \text{ mol Na}$$

$$1.0874 \text{ mol Na} \times \frac{28.2 \text{ J}}{^\circ\text{C} \cdot \text{mol Na}} \times (97.8^\circ\text{C} - 25.0^\circ\text{C}) = 2232 \text{ J} \text{ (2232 kJ)}$$

$$2.60 \text{ kJ/mol Na} \times 1.0874 \text{ mol Na} = 2.8273 \text{ kJ}$$

$$\text{Total heat} = 2.8273 \text{ kJ} + 2.232 \text{ kJ} = 5.0597 = 5.06 \text{ kJ}$$

- 11.123 Use the ideal gas law to calculate the total number of moles of monomer and dimer:

$$n = \frac{PV}{RT} = \frac{(436/760)\text{atm} \times 1.000 \text{ L}}{[0.082057 \text{ L} \cdot \text{atm}/(\text{K} \cdot \text{mol})] \times 373.75 \text{ K}}$$

$$= 0.0187057 \text{ mol monomer} + \text{dimer}$$

$$(0.0187057 \text{ mol monomer} + \text{dimer}) \times \frac{0.630 \text{ mol dimer}}{1 \text{ mol dimer} + \text{monomer}}$$

$$= 0.01178 \text{ mol dimer}$$

$$0.0187057 \text{ mol both} - 0.01178 \text{ mol dimer} = 0.00692 \text{ mol monomer}$$

$$\text{Mass dimer} = 0.01178 \text{ mol dimer} \times \frac{120.1 \text{ g dimer}}{1 \text{ mol dimer}} = 1.414 \text{ g dimer}$$

$$\text{Mass monomer} = 0.00692 \text{ mol monomer} \times \frac{60.05 \text{ g monomer}}{1 \text{ mol monomer}}$$

$$= 0.4155 \text{ g monomer}$$

$$\text{Density} = \frac{1.414 \text{ g} + 0.4155 \text{ g}}{1.000 \text{ L}} = 1.829 = 1.83 \text{ g/L vapor}$$

11.124 Use the ideal gas law to calculate the total number of moles of monomer and dimer:

$$n = \frac{PV}{RT} = \frac{(146/760)\text{atm} \times 1.000 \text{ L}}{[0.082057 \text{ L} \cdot \text{atm}/(\text{K} \cdot \text{mol})] \times 344.45 \text{ K}}$$

$$= 0.006796 \text{ mol monomer} + \text{dimer}$$

Let X_d = mole fraction of dimer; then, write one equation in one unknown based on 1.000 L of gas:

$$0.702 \text{ g both} = 0.006796 \text{ mol } X_d (120.1 \text{ g/mol}) + 0.006796 (1 - X_d) (60.05 \text{ g/mol})$$

$$0.702 \text{ g} = 0.8162 X_d + 0.4081 - 0.4081 X_d$$

$$X_d = \frac{(0.702 - 0.4081)}{0.4081} = 0.7201 = 0.720$$

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