7.1 pg. 174 #3, 4; pg. 175 #5, 6; pg. 179 #7, 8; pg. 181 #9 to 11, 13, 14 #3) 4.65 mol Si #4) 2.17 x 10<sup>23</sup> molecules H<sub>2</sub>O #5) 2.75 x 10<sup>24</sup> atoms #6) 7.72 mol NO<sub>2</sub> #7) a. 30.0 g/mol b. 137.5 g/mol c. 60.0 g/mol d. 108.0 g/mol #8) a. 71.0 g b. 46.0 g c. 331.6 g d. 60.1 g #9) a. 94.2 g/mol b. 136.2 g/mol c. 317.2 g/mol #10) a. 175.3 g/mol b. 139.6 g/mol c. 84.0 g/mol d. 294.3 g/mol #11) One mole of any substance contains Avogadro's number (6.02 x 10<sup>23</sup>) of representative particles. #13) a. 3 atoms c. 3 atoms b. 4 atoms d. 9 atoms #14) a. 2.49 x 10<sup>-1</sup> mol NH<sub>3</sub> b.  $2 \times 10^{-15} \text{ mol } O_2$ c. 0.100 mol Br<sub>2</sub> d. 7.99 mol Li 7.2 pg. 183 #16 to 19; pg. 184 #20, 21; pg. 185 #22, 23; pg. 186 #24 to 28; #16) a. 1.30 x 10<sup>2</sup> g b. 1.27 g c. 1.55 g #17) a. 355 g b. 225 g #18) a. 3.43 x 10<sup>-2</sup> mol B b. 0.343 mol TiO<sub>2</sub> c. 8.82 mol (NH<sub>4</sub>)<sub>2</sub>CO<sub>3</sub> #19) a. 0.987 mol N<sub>2</sub>O<sub>3</sub> b. 2.68 mol N<sub>2</sub> c. 1.21 mol Na<sub>2</sub>O #20) a. 7.17 x 10<sup>-2</sup> L b. 21.5 L c. 82.9 L #21) a. 3.00 mol SO<sub>2</sub> b. 3.93 x 10<sup>-2</sup> mol He c. 44.6 mol  $C_2H_6$ #22) 80.2 g/mol #23) 3.74 g/L

#24) a. 6.5 g Beb.  $67.2 \text{ g N}_2$ c.  $5.44 \text{ g H}_2\text{O}_2$ d.  $8.34 \times 10^2 \text{ g Ca(NO}_3)_2$ #25) a.  $7.85 \times 10^{23}$  molecules NO2b.  $1.21 \text{ L Cl}_2$ c.  $12.9 \text{ g CH}_4$ 

#26) They would have the same volume but different masses; equal volumes of gases have the same number of molecules at the same temperature and pressure.

#27) a. 2.5 mol H<sub>2</sub> b. 2.40 x  $10^{-6}$  mol Li<sub>2</sub>HPO<sub>4</sub> c. 6.93 mol Al d. 2.13 mol SnF<sub>2</sub> #28) gas A: 28.0 g/mol, nitrogen gas B: 64.1 g/mol, sulfur dioxide gas C: 16.0 g/mol, methane

**7.2 (Part of 12.5) pg. 348 #31 to 33; pg. 349 #34 to 36; pg. 353 #40 and 41** #31) 5.60 L

#32) 1.66 x 10<sup>1</sup> L

#33) 1.38 x 10<sup>23</sup> nitrogen molecules

#34) 1.50 L He (g)

#35) 4.48 L

#36) 7.67 x 10<sup>1</sup> L

#40) by using Avogadro's hypothesis and the molar mass and molar volume of the gas

#41) a. 38 L b. 0.40 L c. 5.6 x 10<sup>3</sup> L

7.3 pg. 189 #29, 30; pg. 191 #31 to 33; pg. 192 #34; pg. 193 # 35, 36; pg. 194 #37, 38; pg. 195 #39 to 43

#29) a. 72.2% Mg, 27.8% N	b. 87.1% Ag, 12.9% S
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#30) a. 93.0% Hg, 7.0% O

#31) a. 80.0% C, 20.0% H b. 19.2% Na, 0.83% H, 26.7% S, 53.3% O c. 26.2% N, 7.5% H, 66.4% Cl

#32) a. 46.7% N b. 82.4% N c. 35.0% N

#33) a. 70 g H b. 0.17 g H c. 0.16 g H

#34) a.	58.4 g N	b. 103 g N	C.	43.8 g N	
#35) a.	ОН	b. CH <sub>3</sub>	C.	HgSO <sub>4</sub> d. C <sub>2</sub> HNO	3
#36) C <sub>3</sub>	H <sub>8</sub> N				
#37) a.	$C_2H_6O_2$		b.	$C_6H_4CI_2$	
#38) a.	$C_{6}H_{12}O_{6}$		b.	Na <sub>2</sub> Cr <sub>2</sub> O <sub>7</sub>	
#39) a.	74.2% N, 25.8% O		b.	39.3% Na, 60.7% Cl	
#40) C	<sub>5</sub> H <sub>10</sub> O <sub>2</sub>				
#41) a.	25.4% Ca, 30.4% C,	3.8% H, 40.5% O	b.	3.7% H, 44.4% C, 51.9% N	
#42) a.	4.7g H		b.	14 g H	
#43) a. c.	Molecular Molecular and Empir	ical	b. d.	Molecular Molecular and Empirical	

#### Oteplet 7 REVIEW

#### Answers

- 44. number, mass, or volume; Examples will vary.
- 45. a. molecule c. molecule b. formula unit d. atom
- 46. a. 3 c. 9
  - **b**. 2 **d**. 10
- 47. All contain  $6.02 \times 10^{23}$  molecules.
- 48. 1.00 mol C<sub>2</sub>H<sub>6</sub>
- **49. a**.  $1.81 \times 10^{24}$  atoms Sn
  - **b**.  $2.41 \times 10^{23}$  formula units KCl
  - c.  $4.52 \times 10^{24}$  molecules SO<sub>2</sub>
  - **d**.  $2.89 \times 10^{21}$  formula units NaI
- **50. a**. 98.0 g **d**. 132.1 g
  - **b.** 76.0 g **e.** 89.0 g
  - **c.** 100.1 g **f.** 159.8 g
- **51. a.** 60.1 g **c.** 106.8 g
- **b.** 28.0 g **d.** 63.5 g
- 52. Answers will vary but should include:(1) Determine the number of
- moles of each atom from the
- formula. (2) Look up the atomic mass of
- each element.
- (3) Multiply the number of moles of each atom by its molar
- mass.

(4) Sum these products.

- **53.**  $71.0 \text{ g } \text{Cl}_2$
- 54. Answers will vary.
- **55. a.** 0.258 mol SiO<sub>2</sub>
  - **b.**  $4.80 \times 10^{-4}$  mol AgCl
  - **c.** 1.12 mol Cl<sub>2</sub>
  - d. 0.106 mol KOH
    e. 5.93 mol Ca(C<sub>2</sub>H<sub>3</sub>O<sub>2</sub>)<sub>2</sub>
  - f.  $2.00 \times 10^{-2}$  mol Ca
  - $1.2.00 \times 10^{-1}$  mol Ca
- **56. a.** 108 g C<sub>5</sub>H<sub>12</sub>
  - **b**. 547 g  $F_2$
  - **c.** 71.8 g Ca(CN)<sub>2</sub>
  - **d**. 238 g H<sub>2</sub>O<sub>2</sub>
  - e. 224 g NaOH
- **f.** 1.88 g Ni **57. a.**  $1.7 \times 10^2$  L Ar
- **57. a.**  $1.7 \times 10^2$  LAr **c.** 26.9 LO<sub>2</sub> **b.** 9.9 LC<sub>2</sub>H<sub>6</sub>
- **58. a.** 1.96 g/L **c.** 2.05 g/L **b.** 0.902 g/L
- 59. a. 234 L SO<sub>3</sub>
   b. 2.99 × 10<sup>-22</sup> g C<sub>9</sub>H<sub>8</sub>O<sub>4</sub>
   c. 3.13 × 10<sup>25</sup> atoms
- 60. a. 5.9% H, 94.1% S b. 22.6% N, 6.5% H, 19.4% C,
  - 51.6% O

# Chapter 7 REVIEW

#### CONCEPT PRACTICE

- 44. List three common ways that matter is measured. Give examples of each. 7.1
- 45. Name the representative particle (atom, molecule, or formula unit) of each substance. 7.1
  a. oxygen
  b. sodium sulfide
  c. sulfur dioxide
  d. potassium
- 46. How many hydrogen atoms are in a representative particle of each substance? 7.1
  a. Al(OH)<sub>3</sub>
  b. H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>
  c. (NH<sub>4</sub>)<sub>2</sub>HPO<sub>4</sub>
  d. C<sub>4</sub>H<sub>10</sub>O
- **47**. Which contains more molecules: 1.00 mol H<sub>2</sub>O<sub>2</sub>, 1.00 mol C<sub>2</sub>H<sub>6</sub>, or 1.00 mol CO? 7.1
- 48. Which contains more atoms: 1.00 mol H<sub>2</sub>O<sub>2</sub>, 1.00 mol C<sub>2</sub>H<sub>6</sub>, or 1.00 mol CO? *7.1*
- 49. Find the number of representative particles in each substance. 7.1
  a. 3.00 mol Sn
  c. 7.50 mol SO<sub>2</sub>
  - **b.** 0.400 mol KCl **d.**  $4.80 \times 10^{-3}$  mol NaI
- **51.** Calculate the mass of 1.00 mol of each of these substances. *7.1* 
  - a. silicon dioxide (SiO<sub>2</sub>)
  - **b**. diatomic nitrogen (N<sub>2</sub>)
  - c. iron(III) hydroxide (Fe(OH)<sub>3</sub>)
     d. copper (Cu)
- 52. List the steps you would take to calculate the molar mass of any compound. 7.1
- 53. What is the gram molecular mass (gmm) of chlorine? 7.1
- **54.** Construct a numerical problem to illustrate the size of Avogadro's number. Exchange problems with a classmate and then compare your answers. *7.1*
- 55. How many moles is each of the following? 7.2

   a. 15.5 g SiO<sub>2</sub>
   d. 5.96 g KOH

   b. 0.0688 g AgCl
   e. 937 g Ca(C<sub>2</sub>H<sub>3</sub>O<sub>2</sub>)<sub>2</sub>
  - **c.** 79.3 g  $Cl_2$  **f.** 0.800 g Ca
- 56. Find the mass of each substance. 7.2
  a. 1.50 mol C<sub>5</sub>H<sub>12</sub>
  b. 14.4 mol F<sub>2</sub>
  c. 0.780 mol Ca(CN)<sub>2</sub>
  - d. 7.00 mol H<sub>2</sub>O<sub>2</sub>
  - e. 5.60 mol NaOH
  - f.  $3.21 \times 10^{-2}$  mol Ni

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c. 41.7% Mg, 54.9% O, 3.4% H
d. 42.1% Na, 18.9% P, 39.0% O
61. a. 3.33 g S
c. 40.6 g Mg
b. 5.65 g N
d. 152 g P
62. d. 77.7% Fe in FeO
63. H<sub>2</sub>O<sub>2</sub>
64. a. molecular
b. molecular
c. empirical
f. empirical

- Calculate the volume of each of the following gases at STP. 7.2
  - a. 7.6 mol Ar
  - **b**. 0.44 mol C<sub>2</sub>H<sub>6</sub> **c**. 1.20 mol O<sub>2</sub>
- What is the density of each of the following gases at STP? 7.2

a. C<sub>3</sub>H<sub>8</sub> b. Ne c. NO<sub>2</sub>

- 59. Find each of the following quantities. 7.2a. the volume, in liters, of 835 g SO<sub>3</sub> at STP
  - **b.** the mass, in grams, of a molecule of  $aspin_n$ (C<sub>9</sub>H<sub>8</sub>O<sub>4</sub>)
  - c. the number of atoms in 5.78 mol NH4NO3
- **60**. Calculate the percent composition of each compound. 7.3
  - a. H<sub>2</sub>S c. Mg(OH)<sub>2</sub>
  - b. (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub> d. Na<sub>3</sub>PO<sub>4</sub>
- 61. Using your answers from Problem 60, calculate the number of grams of these elements. 73
   a. sulfur in 3.54 g H<sub>2</sub>S
  - b. nitrogen in 25.0 g (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub>
  - c. magnesium in 97.4 g Mg(OH)<sub>2</sub>
  - d. phosphorus in 804 g Na<sub>3</sub>PO<sub>4</sub>
- 62. Which of the following compounds has the highest iron content? 7.3
  a. FeCl<sub>2</sub>
  c. Fe(OH)<sub>2</sub>
  - **b.**  $Fe(C_2H_3O_2)_3$  **d.** FeO
- 63. You find that 7.36 g of a compound has decomposed to give 6.93 g of oxygen. The only other element in the compound is hydrogen. If the molar mass of the compound is 34.0 g/mol, what is its molecular formula? 7.3
- 64. Classify each formula as an empirical or a molecular formula. 7.3

a.  $S_2Cl_2$  c.  $Na_2SO_3$  e.  $C_{17}H_{19}NO_3$ b.  $C_6H_{10}O_4$  d.  $C_5H_{10}O_5$  f.  $(NH_4)_2CO_3$ 

- What is the molecular formula for each compound? Each compound's empirical formula and molar mass is given. 7.3
   CH<sub>2</sub>O, 90 g/mol
  - a. HgCl, 472.2 g/mol
  - e. C<sub>3</sub>H<sub>5</sub>O<sub>2</sub>, 146 g/mol
- 66. Determine the molecular formula for each compound. 7.3
  a. 94.1% O and 5.9% H; molar mass = 34 g
  - b. 40.0% C, 6.6% H, and 53.4% O; molar mass = 120 g
- 65. a. C<sub>3</sub>H<sub>6</sub>O<sub>3</sub>
  - **b**.  $Hg_2Cl_2$
  - C. C<sub>6</sub>H<sub>10</sub>O<sub>4</sub>
- 66. a. H<sub>2</sub>O<sub>2</sub>
  - **b**. C<sub>4</sub>H<sub>8</sub>O<sub>4</sub>

#### 9.1 pg. 240-241 #3, 4, 6 to 8

- #3) 2 molecules  $H_2$  + 1 molecule  $O_2 \rightarrow$  2 molecules  $H_2O$ 2 mol  $H_2$  + 1 mol  $O_2 \rightarrow$  2 mol  $H_2O$ 44.8 L  $H_2$  + 22.4 L  $O_2 \rightarrow$  44.8 L  $H_2O$
- #4) 2 mol  $C_2H_2$  + 5 mol  $O_2 \rightarrow$  4 mol  $CO_2$  + 2 mol  $H_2O$ 44.8 L  $C_2H_2$  + 112 L  $O_2 \rightarrow$  89.6 L  $CO_2$  + 44.8 LI  $H_2O$ 212 g reactants  $\rightarrow$  212 g products

$$\begin{array}{cccc} \mbox{#6}) & C_2H_5OH_{(l)} + 3 & O_{2\,(g)} \rightarrow 2 & CO_{2\,(g)} + 3 & H_2O_{(l)} \\ \mbox{a.} & 1 & \mbox{molecule} & C_2H_5OH_{(l)} + 3 & \mbox{molecules} & O_{2\,(g)} \rightarrow 2 & \mbox{molecules} & CO_{2\,(g)} + 3 & \mbox{molecules} & H_2O_{(l)} \\ & 1 & \mbox{mol} & C_2H_5OH_{(l)} + 3 & \mbox{mol} & O_{2\,(g)} \rightarrow 2 & \mbox{molecules} & CO_{2\,(g)} + 3 & \mbox{molecules} & H_2O_{(l)} \\ \mbox{b.} & 46.0 & \mbox{g} & C_2H_5OH_{(l)} + 96.0 & \mbox{g} & O_{2\,(g)} \rightarrow 88.0 & \mbox{g} & CO_{2\,(g)} + 54.0 & \mbox{g} & H_2O_{(l)} \\ & 142 & \mbox{g} & \mbox{reactants} \rightarrow 142.0 & \mbox{g} & \mbox{products} \end{array}$$

#7) The number of moles of reactants and products depends on the chemical reaction. For some reactions, moles of reactants and products are equal, but this is not generally the case.

#8) 2 atoms K + 2 molecules of  $H_2O \rightarrow 2$  formula units KOH + 1 molecule  $H_2$ 2 mol K + 2 mol of  $H_2O \rightarrow 2$  mol KOH + 1 mol  $H_2$ 114.2 reactants  $\rightarrow$  114.2 products

# 9.2 pg. 244 #9b, 10; pg. 245 #11, 12; pg. 248 #13, 14; pg. 249 #15, 16; pg. 250 #17 to 22

- #9) b. 7.4 mol
- #10) a. 11.1 mol b. 0.52 mol
- #11) 2.03 g C<sub>2</sub>H<sub>2</sub>
- #12) 1.36 mol CaC<sub>2</sub>
- #13) 4.82 x 10<sup>22</sup> molecules O<sub>2</sub>
- #14) 11.5 g NO<sub>2</sub>
- #15) 1.93 L O<sub>2</sub>
- #16) 0.28 L PH<sub>3</sub>
- #17) 18.6 mL SO<sub>2</sub>

#18) 1.9 dL CO<sub>2</sub>

#19) a.	15.3 mol O <sub>2</sub>	b.	10.2 mol CO <sub>2</sub> , 13.6 mol H <sub>2</sub>	0	
#20) a.	16.0 g CH <sub>4</sub> 1 mol CH <sub>4</sub>	b.	$\frac{1 \ mol \ CH_4}{22.4 \ L \ CH_4}$ C. $\frac{6.02 \times 10^{-10}}{10^{-10}}$	$\frac{1 \text{ mol } CH_4}{10^{23} \text{ molecules } CH_4}$	
#21) a.	176 g CO <sub>2</sub> , 36.0 g H	20	b. $1.60 \times 10^2 \text{ g O}_2$	с	. 212 g = 212 g
#22) a.	18.8 L HF		b. 2.69 x 10 <sup>23</sup> molecul	es H <sub>2</sub> c	. 966 g $SnF_2$

# 9.3 pg. 254 #23, 24; pg. 255 #25; pg. 256 #26; pg. 258 #27, 28; pg. 259 #29 to 32

#23) a. 8.10 mol O <sub>2</sub>	required; $O_2$ is the limiting reactant (reagent)	b. 4.20 mol H <sub>2</sub> O
#24) a.  5.40 mol O <sub>2</sub>	required; $C_2H_4$ is the limiting reactant	b. 5.40 mol $H_2O$
#25) a. HCl	b. 0.164 g H <sub>2</sub>	
#26) 43.2 g H <sub>2</sub> O		
#27) 91.6%		

#28) 83.5%

#29) Limiting reactant determines the amount of product possible from a reaction; excess reactant is not used up by the reaction.

#30) 70.5%

#31) The actual (or experimental) yield is an experimental determined value; a theoretical yield is a value calculated using the balanced chemical equation. The theoretical yield is normally larger than the actual yield, but that is not always the case. The percent yield is a ratio of the actual yield over the theoretical yield expressed as a percentage.

#32) 21.4 g SO<sub>3</sub>

### Chapter () Review

#### Answers

33. a. Two formula units KClO<sub>3</sub> decompose to form 2 formula units KCl and 3 molecules O<sub>2</sub>.
b. Four molecules NH<sub>3</sub> react with 6 molecules NO to form 5 molecules N<sub>2</sub> and 6 molecules H<sub>2</sub>O.

**c.** Four atoms K react with 1 molecule  $O_2$  to form 2 formula units  $K_2O$ .

34. a. Two mol KClO<sub>3</sub> decompose to form 2 mol KCl and 3 mol O<sub>2</sub>.
b. Four mol NH<sub>3</sub> react with 6 mol NO to form 5 mol N<sub>2</sub> and 6 mol H<sub>2</sub>O.

**c.** Four mol K react with 1 mol O<sub>2</sub> to form 2 mol K<sub>2</sub>O.

- **35. a.** 245.2 g **b.** 248.0 g **c.** 188.4 g All obey the law of conservation of mass.
- 36. Answers will vary but should include the idea of writing a ratio using the coefficients of two substances from a balanced equation as the number of moles of each substance reacting or being formed.
- 37. a. 0.54 mol
  - **b**. 13.6 mol
  - c. 0.984 mol
  - d. 236 mol
- **38.** a. 11.3 mol CO, 22.5 mol H<sub>2</sub>
   b. 112 g CO, 16.0 g H<sub>2</sub>
  - $c. 11.4 g H_2$
- **39**. **a**. 372 g F<sub>2</sub>
  - **b**. 1.32 g NH<sub>3</sub>
  - c.  $123 g N_2 F_4$
- **40.** The coefficients indicate the relative number of moles (or particles) of reactants and products.
- 41. a. 51.2 g H₂O
   b. 5.71 × 10<sup>23</sup> molecules NH<sub>3</sub>
  - **c**. 23.2 g Li<sub>3</sub>N
- **42.** The amount of the limiting reagent determines the maximum amount of product that can be formed. The excess reagent is only partially consumed in the reaction.
- **43**. To identify the limiting reagent, express quantities of reactants as moles; compare to the mole ratios from the balanced equation.

# Chapter 9 **REVIEW**

#### **CONCEPT PRACTICE**

- Interpret each chemical equation in terms of interacting particles. 9.1
  - **a.**  $2\text{KClO}_3(s) \longrightarrow 2\text{KCl}(s) + 3\text{O}_2(g)$
  - **b.**  $4NH_3(g) + 6NO(g) \longrightarrow 5N_2(g) + 6H_2O(g)$
  - **c.**  $4K(s) + O_2(g) \longrightarrow 2K_2O(s)$
- Interpret each equation in Problem 33 in terms of interacting numbers of moles of reactants and products. 9.1
- **35.** Calculate and compare the mass of the reactants with the mass of the products for each equation in Problem 33. Show that each balanced equation obeys the law of conservation of mass. 9.1
- 36. Explain the term mole ratio in your own words. When would you use this term? 9.2
- **37**. Carbon disulfide is an important industrial solvent. It is prepared by the reaction of coke with sulfur dioxide. 9.2
  - $5C(s) + 2SO_2(g) \longrightarrow CS_2(l) + 4CO(g)$
  - a. How many moles of CS<sub>2</sub> form when 2.7 mol C reacts?
  - **b**. How many moles of carbon are needed to react with 5.44 mol SO<sub>2</sub>?
  - c. How many moles of carbon monoxide form at the same time that 0.246 mol CS<sub>2</sub> forms?
  - d. How many mol SO<sub>2</sub> are required to make 118 mol CS<sub>2</sub>?
- Methanol (CH<sub>3</sub>OH) is used in the production of many chemicals. Methanol is made by reacting carbon monoxide and hydrogen at high temperature and pressure. 9.2

 $CO(g) + 2H_2(g) \longrightarrow CH_3OH(g)$ 

- a. How many moles of each reactant are needed to produce  $3.60 \times 10^2 \text{ g CH}_3\text{OH}$ ?
- b. Calculate the number of grams of each reactant needed to produce 4.00 mol CH<sub>3</sub>OH.
- c. How many grams of hydrogen are necessary to react with 2.85 mol CO?
- The reaction of fluorine with ammonia produces dinitrogen tetrafluoride and hydrogen fluoride. 9.2

 $5F_2(g) + 2NH_3(g) \longrightarrow N_2F_4(g) + 6HF(g)$ 

**a.** If you have 66.6 g  $NH_3$ , how many grams of  $F_2$  are required for complete reaction?

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- b. How many grams of NH<sub>3</sub> are required to produce 4.65 g HF?
- c. How many grams of N<sub>2</sub>F<sub>4</sub> can be produced from 225 g F<sub>2</sub>?
- 40. What information about a chemical reaction is derived from the coefficients in a balanced equation? 9.2
- Lithium nitride reacts with water to form ammonia and aqueous lithium hydroxide. 9.2 Li<sub>3</sub>N(s) + 3H<sub>2</sub>O(l) → NH<sub>3</sub>(g) + 3LiOH(an)
  - a. What mass of water is needed to react with 32.9 g Li<sub>3</sub>N?
  - b. When the above reaction takes place, how many molecules of NH<sub>3</sub> are produced?
  - c. Calculate the number of grams of Li<sub>3</sub>N that must be added to an excess of water to produce 15.0 L NH<sub>3</sub> (at STP).
- 42. What is the significance of the limiting reagent in a chemical process? What happens to the amount of any reagent that is present in an excess? 9.3
- 43. How would you identify a limiting reagent in a chemical reaction? 9.3
- 44. For each balanced equation, identify the limiting reagent for the given combination of reactants, 9.3

		~			
a.	2Al 3.6 mol	+	3Cl <sub>2</sub> 5.3 mol	$\rightarrow$	2AICI <sub>3</sub>
b.	2H <sub>2</sub> 6.4 mol	+	O <sub>2</sub> 3.4 mol	$\rightarrow$	2H <sub>2</sub> O
C.	2P <sub>2</sub> O <sub>5</sub> 0.48 mol	+	6H <sub>2</sub> O 1.52 mol	$\rightarrow$	4H <sub>3</sub> PO <sub>4</sub>
d.	4P 14.5 mol	+	5O <sub>2</sub> 18.0 mol	$\rightarrow$	2P <sub>2</sub> O <sub>5</sub>

- 45. For each reaction in Problem 44, calculate the number of moles of product formed. 9.3
- **46.** For each reaction in Problem 44, calculate the number of moles of excess reagent remaining after the reaction. 9.3
- Heating an ore of antimony (Sb<sub>2</sub>S<sub>3</sub>) in the presence of iron gives the element antimony and iron(II) sulfide.

 $Sb_2S_3(s) + 3Fe(s) \longrightarrow 2Sb(s) + 3FeS(s)$ 

When 15.0 g Sb<sub>2</sub>S<sub>3</sub> reacts with an excess of Fe, 9.84 g Sb is produced. What is the percent yield of this reaction? 9.3 In an experiment, varying masses of sodium metal are reacted with a fixed initial mass of chlorine gas. The amounts of sodium used and the amounts of sodium chloride formed are shown on the following graph. 9.3



a. Explain the general shape of the graph. b. Estimate the amount of chlorine gas used in

- this experiment at the point where the curve becomes horizontal.
- What does the percent yield of a chemical reaction measure? 9.3
- **31.** The manufacture of compound F requires five separate chemical reactions. The initial reactant, compound A, is converted to compound B, compound B is converted to compound C, and so forth. The diagram below summarizes the stepwise manufacture of compound F, including the percent yield for each step. Provide the missing quantities or missing percent yields. Assume that the reactant and product in each step react in a one-to-one mole ratio. Set



#### CONCEPT MASTERY

- Calcium carbonate reacts with phosphoric acid to produce calcium phosphate, carbon dioxide, and water.
  - $3CaCO_{3}(s) + 2H_{3}PO_{4}(aq) \longrightarrow$   $Ca_{3}(PO_{4})_{2}(aq) + 3CO_{2}(q) + 3H_{2}O(l)$
  - a. How many grams of phosphoric acid react with excess calcium carbonate to produce 3.74 g Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>?
  - b. Calculate the number of grams of CO<sub>2</sub> formed when 0.773 g H<sub>2</sub>O is produced.
- 52. Nitric acid and zinc react to form zinc nitrate, ammonium nitrate, and water.

 $4Zn(s) + 10HNO_3(aq) \longrightarrow$ 

 $4\text{Zn}(\text{NO}_3)_2(aq) + \text{NH}_4\text{NO}_3(aq) + 3\text{H}_2\text{O}(l)$ a. How many atoms of zinc react with 1.49 g

- HNO<sub>3</sub>?
- b. Calculate the number of grams of zinc that must react with an excess of  $HNO_3$  to form 29.1 g  $NH_4NO_3$ .
- Hydrazine (N<sub>2</sub>H<sub>4</sub>) is used as rocket fuel. It reacts with oxygen to form nitrogen and water.
  - $N_2H_4(l) + O_2(g) \longrightarrow N_2(g) + 2H_2O(g)$ **a**. How many liters of  $N_2$  (at STP) form when
  - 1.0 kg N<sub>2</sub>H<sub>4</sub> reacts with 1.0 kg O<sub>2</sub>?
  - b. How many grams of the excess reagent remain after the reaction?
- 54. When 50.0 g of silicon dioxide is heated with set excess of carbon, 32.2 g of silicon carbide to produced.

 $SiO_{C}(s) \div 3C(s) \longrightarrow SiC(s) + 2CO(g)$  What is the percent yield of this reaction? b. How many grams of CO gas are made?

55. If the reaction below proceeds with a 96.8% yield, how many kilograms of CaSO<sub>4</sub> are formed when 5.24 kg SO<sub>2</sub> reacts with an excess of CaCO<sub>3</sub> and O<sub>2</sub>?

 $2CaCO_3(s) + 2SO_2(g) + O_2(g) \longrightarrow \\ 2CaSO_4(s) + 2CO_2(g)$ 

56. Ammonium nitrate will decompose explosively at high temperatures to form nitrogen, oxygen, and water vapor.

 $2NH_4NO_3(s) \longrightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$ What is the total number of liters of gas formed when 228 g NH<sub>4</sub>NO<sub>3</sub> is decomposed? (Assume STP.)

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#### Oterter 9

# REVIEW

- 44. a. Cl<sub>2</sub> c. P<sub>2</sub>O<sub>5</sub>
  - **b**. H<sub>2</sub> **d**. O<sub>2</sub>
- 45. a. 3.5 mol AlCl<sub>3</sub>
  - **b**. 6.4 mol H<sub>2</sub>O
  - **c.** 0.96 mol  $H_3PO_4$
  - **d**. 7.20 mol  $P_2O_5$
- 46. a. 0.1 mol Al
  - b. 0.2 mol O<sub>2</sub>
    c. 0.08 mol H<sub>2</sub>O
  - **d**. 0.1 mol P
- 47. 91.5%
- **48. a.** Initially, the amount of NaCl formed increases as the amount of Na used increases. For this part of the curve sodium is the limiting reagent. Beyond a mass of about 2.5 g of Na, the amount of product formed remains constant because chlorine is now the limiting reagent.

**b.** Chlorine becomes the limiting reagent when the mass of sodium exceeds 2.5 g. This corresponds to a mass of about 3.9 g chlorine.

- 49. The efficiency of a reaction; the actual yield/theoretical yield expressed as a percent.
- **50.** 50.0% yield; 0.500 mol; 0.0500 mol 20.0% yield
- **51**. **a**. 2.36 g H<sub>3</sub>PO<sub>4</sub>
  - **b.** 1.89 g CO<sub>2</sub>
- **52. a.**  $5.70 \times 10^{21}$  atoms Zn **b.** 95. 2 g Zn
- **53. a.**  $7.0 \times 10^2 L N_2$ 
  - b. no reagent in excess
- **54. a.** 96.4% **b.** 45.0 g
- 55. 10.7 kg CaSO4
- 56. 224 L gas

#### 11.1 and 11.3 Homework Answer Key

#### **11.1** pg. 299 #1 to 10

#1) 2.0 J/(g □ °C)

#2) 0.511 J/(g □ °C)

#3) 1.8 kJ

#4) Energy is the capacity for doing work or supplying heat. Heat is energy that transfers between objects across a temperature gradient.

#5) The specific heat of a substance independent of its mass. The heat capacity of an object is proportional to its mass.

#6) It will be the same.

#7) Water has a higher heat capacity than concrete. The sun's heat raises the temperature of the concrete more than that of the water.

#8)  $1.76 \times 10^3$  cal (1.76 kcal = 1.76 Cal); 7.36 x  $10^3$  J (7.36 kJ)

#9) 2.36 x 10<sup>-1</sup> J/(g □ °C)

#10) 2.8 x 10<sup>2</sup> kJ

#### **11.3** pg. 311 #22 and 23; pg. 313 #24 to 29

- #22) 144 kJ
- #23) 0.19kJ

#24) 301 kJ

#25) 3.42 mol NH<sub>4</sub>NO<sub>3 (s)</sub>

- #26) a. molar heat of vaporization, endothermic
  - c. molar heat of fusion, endothermic
- b. molar heat of solution, endothermic
- e. molar heat of solidification, exothermic
- d. molar heat of condensation, exothermic

#27) 2.68 kJ

#28) When a mole of steam condenses, it releases a substantial amount of heat, its molar heat of condensation.

#29) The ice absorbs sufficient heat from the surroundings to change from the solid to the liquid state.

#### **11.2 Homework Answer Key**

**11.2 pg. 302 #11 and 12; pg. 304 #13 and 14; pg. 306 #15 to 19** #11) 1.5 kJ

- #12) 146 J
- #13) 6.63 kJ
- #14) 89.4 kJ
- #15) 2 Mg  $_{(s)}$  + O<sub>2 (g)</sub>  $\rightarrow$  2 MgO  $_{(s)}$  + 1204 kJ
- #16) 3.36 x 10<sup>2</sup> kJ
- #17) the heat released or absorbed in a chemical change.
- #18) -3.98 x 10<sup>3</sup> kJ
- #19) because a phase change involves an energy change

#### **11.4 Homework Answer Key**

#### **11.4** pg. 317 #30 and 31; pg. 318 #32 to 35

#30) a. -30.91 kJ b. 178.4 kJ c. -113.0 kJ

- #31) CO is a compound.
- #32) -8.456 x 10<sup>2</sup> kJ
- #33) -1.960 x 10<sup>2</sup> kJ

#34) Answers will vary, but should include the idea that chemical equations can be added algebraically along with their enthalpies to obtain the enthalpy of a different chemical reaction.

#35) The sign of  $\Delta H$  must be changed.

# CHAPTER 11

# KEVIEW

#### Answers

- **36.** Answers will vary, but should include the idea that energy is conserved in every physical and chemical process.
- **37.** Heat flows from the object at the higher temperature to the object at the lower temperature.
- **38.** Potential energy is energy stored in a substance because of its chemical composition.
- **39.** the chemical composition of the substance and its mass
- **40.** 1 Cal = 1000 cal = 1 kcal
- **41. a**.  $8.50 \times 10^{-1}$  Calorie
  - **b.**  $1.86 \times 10^3$  J
    - **c.**  $1.8 \times 10^3$  J
    - **d**.  $1.1 \times 10^2$  cal
- **42.** Answers will vary, but should mention that thermochemistry measures heat flow across the boundary between the system and the surroundings.
- **43.** A negative sign is given to heat flow from the system to the surroundings. A positive sign is given to heat flow to the system from the surroundings.
- 44. a. exothermic

**b**. The immediate surroundings are the glass beaker and the air. If one or more of the substances is an aqueous solution, the water is also considered part of the surroundings.

- 45. a. exothermic
  - b. endothermic
  - **c**. exothermic
  - d. endothermic
- 46. enthalpy
- **47.** A calorimeter is an instrument used to measure heat changes in physical or chemical processes.
- 48. The foam cup will absorb heat. Some heat will be lost to the air. If the reactants are not completely mixed, temperature measurements will be inaccurate.
- 49. bomb calorimeter
- **50.** one atmosphere pressure (101.3 kPa); all reactants and products in their normal physical state

# Chapter 11 REVIEW

#### CONCEPT PRACTICE

- **36**. Explain in your own words the law of conservation of energy. 11.1
- **37**. What always happens when two objects of different temperatures come in contact? Give an example from your own experience. *11.1*
- Define potential energy in terms of chemistry. 11.1
- **39**. What factors determine the heat capacity of an object? *11.1*
- **40**. What is the relationship between a calorie and a Calorie? *11.1*
- 41. Make the following conversions. 11.1
  a. 8.50 × 10<sup>2</sup> cal to Calories
  b. 444 cal to joules
  - c. 1.8 kJ to joules
  - **d**.  $4.5 \times 10^{-1}$  kJ to calories
- **42**. Why do you think it is important to define the system and the surroundings? *11.1*
- **43**. Describe the sign convention that is used in thermochemical calculations. *11.1*
- **44**. Two substances in a glass beaker chemically react, and the glass beaker becomes too hot to touch. *11.1* 
  - a. Is this an exothermic or endothermic reaction?
  - b. If the two substances are defined as the system, what constitutes the surroundings?
- **45**. Classify these processes as exothermic or endothermic. *11.1* 
  - a. condensing steam
  - b. evaporating shorted
  - c. burning alcohol
  - d. baking a pointo
- 46. What special name is given to a heat change at constant pressure? 11.2
- 47. What is the function of a calorimeter? 11.2
- **48**. There are some obvious sources of error in experiments that use foam cups as calorimeters. Name at least three. *11.2*
- **49**. What special device would you use to measure the heat released at constant volume? *11.2*
- **50**. Give the standard conditions for heat of combustion. *11.2*
- **51**. What information is given in a thermochemical equation? *11.2*
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- **51.** amount of heat released or absorbed in the chemical change at constant pressure
- 52. Heat is being used to melt the ice.
- **53. a.**  $-2.10 \times 10^{1}$  kJ **b.**  $-1.8 \times 10^{1}$  kJ
  - **c.**  $-5.56 \times 10^2 \, \text{kJ}$
  - **d.** 6.5 kJ
- 54. increase
- **55.** It allows the calculation of the enthalpy of a reaction from the known enthalpies of two or more other reactions.

- 52. Explain why ice melts at 0 °C without an increase of temperature, even though heat is flowing from the surroundings into the system (the ice). 11.3
- 53. Calculate the quantity of heat gained or log in the following changes. 11.3
  a. 3.50 mol of water freezes at 0 °C.
  b. 0.44 mol of steam condenses at 100 °C
  - c. 1.25 mol NaOH(s) dissolves in water.
  - d. 0.15 mol C<sub>2</sub>H<sub>5</sub>OH(l) vaporizes at 78.3 °C.
- Sodium acetate dissolves readily in water according to the following equation. 11.3

$$NaC_2H_3O_2(s) \longrightarrow NaC_2H_3O_2(aq)$$

$$\Delta H = -17.3 \text{ kJ/mol}$$

Would this process increase or decrease the temperature of the water?

- 55. Explain the usefulness of Hess's law of heat summation in thermochemistry. 11.4
- 56. A considerable amount of heat is required for the decomposition of aluminum oxide. 11.4

$$2Al_2O_3(s) \longrightarrow 4Al(s) + 3O_2(g)$$
$$AH = 3352 \text{ kI}$$

- b. Is the reaction exothermic or endothermic?
- **57.** Calculate the heat change for the formation lead(IV) chloride by the reaction of lead(II) chloride with chlorine. *11.4*

$$PbCl_{2}(s) + Cl_{2}(g) \longrightarrow PbCl_{4}(l)$$
$$\Delta H = ?$$

Use the following thermochemical equations

$$Pb(s) + 2Cl_2(g) \longrightarrow PbCl_4(l)$$

$$\Delta H = -329.2 \text{ kJ}$$

$$Ph(s) + Cl_{2}(g) \longrightarrow PhCl_{2}(s)$$

$$\Delta H = -359.4 \text{ kJ}$$

58. From the following reactions: 11.4

$$\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \longrightarrow NO(g)$$

$$\Delta H = 90.4 \text{ kJ/mol}$$

 $\frac{1}{2}\mathrm{N}_{2}(g) + \mathrm{O}_{2}(g) \longrightarrow \mathrm{NO}_{2}(g)$ 

 $\Delta H = 33.6 \text{ kJ/mol}$ 

determine the heat of reaction for:  $NO(g) + \frac{1}{2}O_2(g) \longrightarrow NO_2(g)$  $\Delta H = ?$ 

- **56. a.**  $-1.676 \times 10^3 \text{ kJ}$
- **b**. exothermic
- **57.**  $3.02 \times 10^1 \text{ kJ}$
- **58.**  $-5.68 \times 10^1 \, \text{kJ}$

59. Calculate the heat change for the formation of copper(I) oxide from its elements. 11.4

$$\operatorname{Cu}(s) + \frac{1}{2}\operatorname{O}_2(g) \longrightarrow \operatorname{CuO}(s)$$

Use the following thermochemical equations to make the calculation.

> $CuO(s) + Cu(s) \longrightarrow Cu_2O(s)$  $\Delta H = -11.3 \text{ kJ}$  $Cu_2O(s) + \frac{1}{2}O_2(g) \longrightarrow 2CuO(s)$  $\Delta H = -114.6 \, \text{kJ}$

- 60. What is the standard heat of formation of a free element in its standard state? 11.4
- 61. Consider the statement, "the more negative the value of  $\Delta H_{\rm f}^0$ , the more stable the compound." Is this statement true or false? Explain. 11.4
- 62. Calculate the change in enthalpy (in kJ) for the following reactions. 11.4

**a**.  $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$  **b**.  $2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$ 

63. What is the standard heat of formation of a compound? 11.4

CONCEPT MASTERY

- 64. Equal masses of two substances absorb the same amount of heat. The temperature of substance A increases twice as much as the temperature of substance B. Which substance has the higher specific heat? Explain.
- 65. If 3.20 kcal of heat is added to 1.00 kg of ice at 0 °C, how much water at 0 °C is produced, and how much ice remains?
- 66. The amounts of heat required to change different quantities of carbon tetrachloride  $(CCl_4)(l)$  into vapor are given in the table.

Mass of CCI	He	at
(g)	(J)	(cal)
2.90	652	156
7.50	1689	404
17.0	3825	915
26.2	5894	1410
39.8	8945	2140
51.0	11453	2740

a. Graph the data, using heat as the dependent variable.

b. What is the slope of the line?

c. The heat of vaporization of  $CCl_4(l)$  is 53.8 cal/g. How does this value compare with the slope of the line?

- 67. Calculate the heat change in calories when 45.2 g of steam at 100 °C condenses to water at the same temperature. What is the heat change in joules?
- 68. Find the enthalpy change for the formation of phosphorus pentachloride from its elements.

 $2P(s) + 5Cl_2(g) \longrightarrow 2PCl_5(s)$ 

Use the following thermochemical equations.

$$PCl_{5}(s) \longrightarrow PCl_{3}(g) + Cl_{2}(g)$$
$$\Delta H = 87.9 \text{ kJ}$$
$$2P(s) + 3Cl_{2}(g) \longrightarrow 2PCl_{3}(g)$$

g)

$$\Delta H = -574 \,\mathrm{kJ}$$

- 69. Use standard heats of formation  $(\Delta H_f^0)$  to calculate the change in enthalpy for these reactions. a.  $2C(s, \text{graphite}) + O_2(g) \longrightarrow 2CO(g)$ 
  - **b.**  $2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$

c. 
$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

- 70. An ice cube with a mass of 40.0 g melts in water originally at 25.0 °C.
  - a. How much heat does the ice cube absorb from the water when it melts? Report your answer in calories, kilocalories, and joules.
  - b. Calculate the number of grams of water that can be cooled to 0 °C by the melting ice cube.
- 71. The molar heat of vaporization of ethanol (C<sub>2</sub>H<sub>5</sub>OH(l)) is 43.5 kJ/mol. Calculate the heat required to vaporize 25.0 g of ethanol at its boiling point.
- 72. An orange contains 445 kJ of energy. What mass of water could this same amount of energy raise from 25.0 °C to the boiling point?
- 72. The combustion of ethane (C2H4) is an exothermicreaction.

$$C_2H_{4,2} + 2C_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$$
  
=  $-1.39 \times 10^3 \text{ kJ}$ 

Calculate the amount of heat liberated when 4.79 g C.H. reacts with excess oxygen.

74. Calculate the heat change  $(\Delta H)$  for the formade from its elements.

 $N_2(g) + O_2(g) \longrightarrow 2NO(g)$ 

Use these thermochemical equations.

$$4NH_3(g) + 3O_2(g) \longrightarrow 2N_2(g) + 6H_2O(l)$$

$$\Delta H = -1.53 \times 10^3 \,\mathrm{kJ}$$

 $4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)$  $\Delta H = -1.17 \times 10^3 \, \text{kJ}$ 

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#### HHAPTER

## 231/B//

- **59**.  $-1.259 \times 10^2$  kJ
- 60. zero
- 61. This statement is true, because stability implies lower energy. The greater the release of heat, the more stable is the compound relative to its elements (all of which have  $\Delta H_f^0 = 0$ ).
- 62. a.  $-8.902 \times 10^2$  kJ **b**.  $-5.660 \times 10^2$  kJ
- **63.**  $\Delta H_f^0$  for the formation of one mole of a compound from its elements
- 64. substance B; For equal masses, the substance with the greater heat capacity undergoes the smaller temperature change.
- **65.**  $4.00 \times 10^1$  g water;  $9.60 \times 10^2$  g ice
- 66. a. See students' graphs. b. about 54 cal/g c. The two values are essentially
- the same.
- **67.**  $2.44 \times 10^4$  cal;  $1.02 \times 10^5$  J
- **68.**  $-7.50 \times 10^2$  kJ
- **69. a.**  $-2.21 \times 10^2$  kJ
  - **b**.  $-1.96 \times 10^2$  kJ
  - **c.**  $-9.046 \times 10^2$  kJ
- **70.** a.  $3.19 \times 10^3$  cal, 3.19 kcal, 1.34 $\times 10^4 \text{ J}$
- **b.**  $1.28 \times 10^2 \, \text{g H}_2\text{O}$
- **71.**  $2.36 \times 10^1 \text{ kJ}$
- **72.**  $1.42 \times 10^3$  g
- **73.**  $2.38 \times 10^2$  kJ
- 74.  $1.8 \times 10^2$  kJ

#### opuruer iii

# REVIEW

#### Answers

- **75.**  $6.72 \times 10^1 \text{ kJ}$
- 76. a. (2) volume
  b. (3) endothermic
  c. (4) heat
- 77. The region denoted by  $\Delta H_{\text{fus}}$ represents the coexistence of solid and liquid. The region denoted by  $\Delta h_{\text{vap}}$  represents the coexistence of liquid and vapor.
- **78**. 0.40
- **79.**  $1.20 \times 10^{24}$  H<sub>2</sub> molecules
- **80.**  $\operatorname{Ag}^+(aq) + \operatorname{Cl}^-(aq) \to \operatorname{AgCl}(s)$
- **81.**  $1.18 \times 10^1 \, \text{g} \, \text{O}_2$
- **82. a.**  $3.24 \times 10^{1}$  kcal,  $1.36 \times 10^{2}$  kJ **b.** 8.13 kg
- **83.**  $-1.37 \times 10^2 \, \text{kJ}$
- 84. 9.6 g
- 85. 45.4°C
- **86.**  $\Delta H_{vap}$  for water at 70°C is approximately 42 kJ/mol. 1 L of water (.1000 mL) has a mass of 1000 g and contains 55.6 mol water. Therefore, the amount of heat required is 42 kJ/mole × 55.6 mol =  $2.34 \times 10^3$  kJ.

75. How much heat must be removed from a 45.0 -g sample of naphthalene (C<sub>10</sub>H<sub>8</sub>) at its freezing point to bring about solidification? The heat of fusion of naphthalene is 191.2 kJ/mol.

#### **CRITICAL THINKING**

- Choose the term that best completes the second relationship.
  - a. kilojoules:heat

cm <sup>3</sup> :	
(1) mass	(3) energy
(2) volume	(4) weight
b. right: left	
exothermic:	<u> </u>
(1) combustion	(3) endothermic
(2) heat	(4) joule

- 77. Refer to Figure 11.15. Which region of the graph represents the coexistence of solid and liquid? Liquid and vapor?

#### **CUMULATIVE REVIEW**

- 78. What fraction of the average kinetic energy of hydrogen gas at 100 K does hydrogen gas have at 40 K?
- **79.** How many hydrogen molecules are in 66.8 L H<sub>2</sub>(g) at STP?
- Write the net ionic equation for the teaction of aqueous solutions of sodium chloride and alver acetate.
- 81. How many grams of oxygen are formed by the decomposition of 25.0 g of hydrogen peroxide?
   2H<sub>2</sub>O<sub>2</sub>(l) → 2H<sub>2</sub>O(l) + O<sub>2</sub>(g)

#### **CONCEPT** CHALLENGE

- **82**. The temperature of a person with an extremely high fever can be lowered with a sponge bath of 2-propanol (isopropyl alcohol). The heat of vaporization of 2-propanol is 11.1 kcal/mol.
  - a. How many kilocalories of heat are removed from the skin of a person when 175 g of 2-propanol evaporates? How many kilojoules?
  - **b**. How many kilograms of water would this energy loss cool in lowering the temperature from 40.0 °C to 36.0 °C?

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**83.** Ethane  $(C_2H_6(g))$  can be formed by the reaction of ethene  $(C_2H_4(g))$  with hydrogen gas.

 $C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$ 

Use the heats of combustion for the following reactions to calculate the heat change for the formation of ethane from ethene and hydrogen

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$
$$\Delta H = -5.72 \times 10^2 \,\text{kJ}$$

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2H_2O(l) + 2CO_{l(g)}$$
$$\Delta H = -1.401 \times 10^3 \text{ kJ}$$

$$2C_2H_6(g) + 7O_2(g) \longrightarrow 6H_2O(l) + 4C_{0_2(g)}$$
$$\Delta H = -3.100 \times 10^3 \text{ kJ}$$

- 84. An ice cube at 0 °C was dropped into 30.0 goi water in a cup at 45.0 °C. At the instant that all of the ice was melted, the temperature of the water in the cup was 19.5 °C. What was the mass of the ice cube?
- **85.** 41.0 g of glass at 95 °C is placed in 175 gof water at 21 °C in an insulated container. They are allowed to come to the same temperature. What is the final temperature of the glass -water mixture? The specific heat of glass is 2.1 cal/(g × °C).
- 86. The enthalpy of vaporization of water at various temperatures is given in the graph. From this graph, estimate the amount of heat required convert 1 L of water to steam on the summited Mount Everest (29 002 ft), where the boiling temperature of water is 70 °C.



#### 10.1, 10.2 and 10.4 Homework Answer Key

#### 10.1 pg. 271 #1, 2; pg. 272 #3 to 7

#1) 51.3 kPa, 0.507 atm

#2) 33.7 kPa is greater than 0.25 atm

#3) Answers will vary but should include the idea that particles exhibit rapid, random motion with elastic collisions.

#4) The collision of gas particles with an object causes gas pressure.

- #5) 72.6 kPa
- #6) Heat the water.
- #7) by one-half.

#### 10.2 pg. 279 #8 to 13

#8) Liquid molecules slide past one another despite intermolecular forces.

#9) Liquid particles are more attracted to one another and have little space between them.

#10) Evaporation, which happens only at the surface, can occur below the boiling point.

#11) a. About 76°C b. About 50°C c. About 62°C

#12) At boiling, vapor pressure equals external pressure. As external pressure increases, the temperature needed to produce an equivalent vapor pressure - the boiling point - must increase.

#13) When the molecules with the highest kinetic energy escape, average kinetic energy of the liquid is lowered.

#### 10.4 pg. 286 #17 to 19

#17) the state of water at given temperatures and pressures and the conditions at which phase changes occur.

#18) During sublimation, a substance changes directly from solid to vapor; drying without extensive heating.

#19) The triple point is the temperature and pressure at which an equilibrium exists among all three phases of a substance.

# chapter 10 REVIEW

# CONCEPT PRACTICE

191.20

- 20. What is meant by elastic collision? 10.1
- 21. List the various units used to measure pressure, and identify the SI unit. 10.1
- 22. Change 1656 kPa to atm. 10.1
- 23. Convert 190 mm Hg to the following. 10.1 a. kilopascals
  - b. atmospheres of pressure
- 24. How much pressure (in mm Hg) does a gas exert at 3.1 atm? 10.1
- 25. Explain the relationship between the absolute temperature of a substance and the kinetic energy of its particles. 10.1
- 26. How is the average kinetic energy of water molecules affected when you pour hot water from a kettle into cups at the same temperature as the water? 10.1
- 27. What does the abbreviation STP represent? 10.1
- Express standard temperature in kelvins and standard pressure in kilopascals and in millimeters of mercury. 10.1
- 29. What is significant about the temperature absolute zero? 10.1
- 30. By what factor does the average kinetic energy of gas molecules in an aerosol container increase when the temperature is raised from 27 °C (300 K) to 627 °C (900 K)? 10.1
- 31. A liquid is a condensed state of matter. Explain. 10.2
- 32. Explain why liquids and gases differ 10.2a. in physical state.b. in compressibility.
- Compare the evaporation of a contained liquid with that of an uncontained liquid. 10.2
- 34. Explain vapor pressure and dynamic equilibrium. 10.2
- 35. Explain why increasing the temperature of a liquid increases its rate of evaporation. 10.2
- **36.** Would you expect an equilibrium vapor pressure to be reached above a liquid in an open container? Why? *10.2*
- **37**. Describe the effect that increasing temperature has on the vapor pressure of a liquid. *10.2*
- Distinguish between the boiling point and the normal boiling point of a liquid. 10.2

39. Use the graph to answer each question. 10.2



- a. What is the vapor pressure of water at 40 °C?
- b. At what temperature is the vapor pressure of water 600 mm Hg?
- c. What is the significance of the vapor pressure of water at 100 °C?
- **40.** Use Figure 10.11 to determine the temperature at which water will boil in an open vessel when the atmospheric pressure is 400 mm Hg. *10.2*
- 41. At the top of Mount Everest, water boils at only 69 °C. Use Figure 10.11 to estimate the atmospheric pressure at the top of this mountain. 10.2
- 42. Explain how boiling is a cooling process. 10.2
- 43. Name at least one physical property that would permit you to distinguish a molecular solid from an ionic solid. 10.3
- Describe what happens when a solid is heated to its melting point. 10.3
- Molecular solids usually have lower melting points than ionic solids. Why? 10.3
- 46. When you remove the lid from a food container that has been left in a freezer for several months, you discover a large collection of ice crystals on the underside of the lid. Explain what has happened. 10.4
- 47. Any liquid stays at a constant temperature while it is boiling. Why? 10.4

#### CONCEPT MASTERY

- **48**. Describe evaporation, vapor pressure, and boiling point.
- 49. Mount McKinley (6194 m) in Alaska is the tallest peak in North America. The atmospheric pressure at its peak is 330 mm Hg. Find the boiling point of water there. Use Figure 10.11.

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### Answers

- An elastic collision transfers energy from one particle to another. There is no change in the total kinetic energy.
- 21. pascal (Pa), millimeter of mercury (mm Hg), atmosphere (atm); Pa is the SI unit.
- 22. 16.35 atm
- 23. a. 25 kPa
- b. 0.25 atm
- **24.**  $2.4 \times 10^3$  mm Hg

 Kinetic energy is directly proportional to the absolute temperature.

- **26**. The average kinetic energy is unaffected.
- 27. standard temperature and pressure, 0 °C and 1 atm
- 28. 273 K; 101.3 kPa, 760 mm Hg
- 29. Average kinetic energy of particles is zero.
- 30. The average kinetic energy triples.
- Its volume is minimally affected by an increase in pressure.
- 32. a. Because of attractive forces between parti-

#### Chapter 10

# REVIEW

cles, liquids are denser than gases.

- **b**. The molecules of liquids are in close contact and cannot be squeezed together. The molecules of gases are far apart and, thus, compressible.
- 33. In both cases, particles with sufficient kinetic energy move from the liquid to the vapor phase. In a container, a dynamic equilibrium is set up between the liquid and its vapor.
- 34. Vapor pressure results from collisions of vapor particles with the container's walls. A dynamic equilibrium exists when the rate of evaporation of the liquid equals the rate of condensation of the vapor.
- **35.** More molecules have enough energy to escape attractions within the liquid.
- 36. No; vapor continuously leaves the surface of the liquid preventing dynamic equilibrium from being established.
- **37**. It increases the kinetic energy, which increases the vapor pressure.
- 38. The boiling point is the temperature at which the vapor pressure equals the external pressure. At the normal boiling point, external pressure is 1 atm.
- **39.** a. about 55 mm Hg
  b. about 93 °C
  c. 760 mm Hg is standard pressure.
- 40. about 82 °C
- 41. about 250 mm Hg
- **42.** Escaping molecules have more kinetic energy than remaining molecules, which lowers the average kinetic energy.
- **43.** Ionic compounds have crystalline structures with relatively high melting points.
- Its molecules gain sufficient kinetic energy to overcome attractive forces.
- **45.** The forces between the molecules of molecular solids are weaker.

### Chapter 10

## REVIEW

#### Answers

- **46.** Moisture in the food has sublimed and then resolidified on the container lid.
- 47. Molecules use the added heat to escape the liquid; average kinetic energy remains the same.
- 48. Evaporation is the conversion of a liquid to a vapor at temperatures below the boiling point.
  Vapor pressure is the force per unit area exerted by the vapor particles on the container walls.
  Boiling point is the temperature at which vapor pressure equals external pressure.
- 49. about 77 °C
- **50.** The mass of air pressing on molecules below it; the mass of air above a mountain is less than the mass at sea level.
- **51.** Although there is an ongoing exchange of particles, net amounts of vapor molecules and liquid molecules remain constant.
- 52. At −196 °C, the kinetic energy of the air particles decreases drastically as does the pressure the particles exert, thus reducing the balloon's flexible volume. The kinetic energy of the air particles increases as the balloon warms back to room temperature.
- 53. It increases.
- 54. 819 K
- **55.** There is sufficient average kinetic energy to disrupt the forces holding the solid crystal together.
- 56. a. 2 b. 1
- 57. Possible answers include: odors travel through a room; ink will move throughout a beaker of water.
- **58**. They are the same because the temperature is the same.
- **59.** No; collisions between nonatomic objects involve the conversion of kinetic energy into heat.

- **50.** What causes atmospheric pressure, and why is it much lower on the top of a mountain than it is at sea level?
- 51. Why is the equilibrium that exists between a liquid and its vapor in a closed container called a dynamic equilibrium?
- 52. Pouring liquid nitrogen onto a balloon decreases the volume of the balloon dramatically, as shown in the photograph. Afterward, the balloon reinflates. Use kinetic theory to explain this sequence of events. The temperature of liquid nitrogen is −196 °C.



- 53. What happens to the average kinetic energy of the water molecules in your body when you get a fever?
- 54. The temperature of the gas in an aerosol container is 0 °C (273 K). To what temperature must this gas be raised to increase the average kinetic energy of the gas molecules by a factor of three?
- **55.** Explain what happens at the melting point of a substance.

#### CRITICAL THEFTING

- Choose the term that best completes the second relationship.
  - a. temperature : thermometer atmospheric pressure : \_\_\_\_\_\_ (1) valve (3) volume (2) barometer (4) gauge
  - b. evolution : organisms
  - kinetic theory :
  - (1) particles (3) heat
  - (2) temperature (4) pressure
- 57. What everyday evidence suggests that all matter is in constant motion?
- 58. Is the average kinetic energy of the particles in a block of ice at 0 °C the same as or different from the average kinetic energy of the particles in a gas-filled weather balloon at 0 °C? Explain.
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- **60.** The energy needed to evaporate perspiration comes from skin cells, which helps lower body temperature.
- **61.** Some compounds have stronger intermolecular forces than others.
- **62. a**. 1, 2, 1, 2 **c**. 4, 5, 4, 6
  - **b**. 1, 1, 1, 4 **d**. 2, 13, 12, 14
- **63.**  $3.74 \text{ g/L} = 0.003 74 \text{ g/cm}^3$
- **64.** C<sub>2</sub>H<sub>4</sub>
- **65. a**. 13.9 mol SO<sub>2</sub> **c**. 0.021 mol CO<sub>2</sub> **b**. 0.0472 mol NH<sub>3</sub>

- 59. Are perfectly elastic collisions possible between objects that you can see?
- 60. How does perspiration help cool your body on a hot day?
- 61. Why do different liquids have different boiling points?

#### CUMULATIVE REVIEW

- 62. Balance these equations.
  - $\mathbf{a}. V_2 O_5 + H_2 \longrightarrow V_2 O_3 + H_2 O$
  - **b**.  $(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + N_2 + H_2O$
  - **c**.  $NH_3 + O_2 \rightarrow NO + H_2O$
  - $\mathbf{d}. \mathrm{C}_{6}\mathrm{H}_{14} + \mathrm{O}_{2} \rightarrow \mathrm{CO} + \mathrm{H}_{2}\mathrm{O}$
- 63. What is the density of krypton gas at STP?
- **64.** Hydrogen reacts with ethene  $(C_2H_4)$  to form ethane  $(C_2H_6)$ .

 $C_2H_4 + H_2 \longrightarrow C_2H_6$ 

What is the limiting reagent when 40.0 gC\_H, reacts with 3.0 g  $H_2?$ 

- 65. How many moles is each substance?a. 888 g of sulfur dioxide
  - **b.**  $2.84 \times 10^{22}$  molecules of ammonia **c.** 0.47 L of carbon dioxide (at STP)

with dichlorine heptoxide.

66. Perchloric acid forms by the reaction of water

$$Cl_2O_7 + H_2O \longrightarrow HClO_4$$

- a. How many grams of Cl<sub>2</sub>O<sub>7</sub> must react with an excess of H<sub>2</sub>O to form 56.2 g HCl0<sub>i</sub>?
- b. How many mL of water are needed to form 3.40 mol HClO<sub>4</sub>?

#### CEPT CHALLENGE

- 67. If the volume of the container in which there is a liquid-vapor equilibrium is changed, the vapor pressure is not affected. Why?
- 68. The ions in sodium chloride are arranged ina face-centered cubic pattern. Sketch a layer of the ions in a crystal of sodium chloride.
- Using Figure 10.14, identify the crystal system described by these characteristics.
  - a. three unequal axes mutually perpendicular
  - b. three equal axes making equal angles with each other
  - c. two equal axes and one unequal axis mutually perpendicular
  - d. three unequal axes intersecting obliquely
  - e. three axes equal and mutually perpendicular
- **66. a.** 51.2 g Cl<sub>2</sub>O<sub>7</sub> **b.** 30.6 mL H<sub>2</sub>O
- 67. Vapor pressure depends only on the kinetic energy of the escaping molecules.
- 68. Na<sup>+</sup>Cl<sup>-</sup> Na<sup>+</sup>Cl<sup>-</sup> Cl<sup>-</sup> Na<sup>+</sup>Cl<sup>-</sup> Na<sup>+</sup> Na<sup>+</sup>Cl<sup>-</sup> Na<sup>+</sup>Cl<sup>-</sup> Cl<sup>-</sup> Na<sup>+</sup>Cl<sup>-</sup> Na<sup>+</sup>
- 69. a. orthorhombic
  b. rhombohedral
  c. tetragonal
  d. triclinic
  e. cubic

#### 12.1 to 12.4 Homework Answer Key

#### 12.1 pg. 328 #1 to 4

#1) The volume of gas particles is negligible; gas particles do not interact; gas particles move rapidly; the average kinetic energy of the gas particles is directly proportional to the Kelvin temperature.

#2) Kinetic energy is transferred from one gas particle to another without loss of energy when gas particles collide. As Kelvin temperature increases, the kinetic energy increases in direct proportion.

#3) According to kinetic theory, there is considerable space between gas particles, which allows gases to be compressed.

#4) Pressure in kilopascals (kPa), Kelvin temperature (K), volume in Litre (L), and amount in moles (mol).

#### 12.2 pg. 332 #5 to 9

#5) As the number of gas molecules increases in a constant-volume container, the pressure increases. As the volume of the container decreases, pressure increases, if other variables are constant.

#6) As the temperature of a contained gas increases, the pressure increases, and vice versa.

#7) Increase the volume by a factor of four.

#8) Add 100 times as much gas or decrease the volume by a factor of 100.

#9) The pressure will double.

**12.3 pg. 335 #10, 11; pg. 337 #12, 13; pg. 338 #14; pg. 339 #15; pg. 340 #16 to 21** #10) 6.48 L

- #11) 68.3 kPa
- #12) 3.39 L
- #13) 8.36 L
- #14) 2.58 kPa

#15) 341 K, 68°C

#16) 0.342 L

#17) 129 kPa

#18) Boyle's Law:  $P_1V_1 = P_2V_2$  Charles's Law:  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$  Gay-Lussac's Law:  $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ 

#19) When one condition (P, V, or T) is held constant, its variable cancels out, and one of the other three laws is left

#20)  $P_1V_1 = P_2V_2$  P = pressure; V = volume. The subscript 1 represents the starting conditions; the subscript 2 represents the final conditions. Temperature is constant, so it is not in the expression.

#21) 24.2 L

#### 12.4 pg. 342 #22, 23; pg. 343 #24, 25; pg. 346 #26 to 30

- #22) 251 mol He (g)
- #23) 1.71 x 10<sup>3</sup> kPa
- #24) 2.5 g air
- #25) 17.6 L of O<sub>2 (g)</sub>

#26) by using the ideal gas law, PV = nRT

#27) An ideal gas obeys the assumptions of the kinetic molecular theory of gases. A real gas will deviate from ideal behavior except within a small margin of *P*, *V*, *T* conditions.

#28) Real gases have attractions between molecules, and their molecules have volume. At low temperatures, the attractions between molecules pull them together and reduce the gas volume. At high pressures, the volume occupied by the molecules is a significant part of the total volume because the molecules are closer together.

#29) 17.0 L

#30) 2.23 x 10<sup>3</sup> kPa

#### 12.5 Homework Answer Key

**12.5 pg. 348 #31 to 33; pg. 349 #35, 36; pg. 351 #37, 38; pg. 353 #39 to 44** #31) 5.60 L

#32) 16.6 L

- #33) 1.38 x 10<sup>23</sup> nitrogen molecules
- #35) 4.48 L
- #36) 76.7 L
- #37) 93.4 kPa
- #38) 3.3 kPa

#39) **Avogadro's Hypothesis Law**: Equal volumes of gases at the same temperature and pressure contains equal numbers of particles. **Dalton's Law**: At constant volume and temperature, the total pressure exerted by a mixture of gases equals the sum of the partial pressures of the component gases. **Graham's Law**: The rate of effusion of a gas is inversely proportional to the square root of the molar mass of the gas.

#40) by using Avogadro's Law and the Molar Mass and Molar Volume of the gas.

#41) a. 38 L b. 0.40 L c. 5600 L

#42) Rearrange the equation  $P_{\text{total}} = P_1 + P_2 + P_3 + \dots + P_x$  to isolate the desired pressure term. The rate of effusion of a gas can be calculated by using the equation:  $\frac{Rate_1}{Rate_2} = \frac{\sqrt{Molar Mass_B}}{\sqrt{Molar Mass_A}}$ 

#43) This is the volume of 1 mol of any gas at STP.

#44) Carbon monoxide and nitrogen have identical molar masses of 28.0 g/mol.

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## REVIEW

### Answers

- **45**. The increased kinetic energy of the gas particles causes collisions to occur with more force.
- **46.** The gas particles become closer together.
- 47. The pressure doubles.
- 48. The pressure quadruples.
- **49.** Temperatures measured on the Kelvin scale are always positive and directly proportional to the average kinetic energy of the gaseous particles.
- **50.** The volume decreases. The molecules have less kinetic energy, which causes less pressure on the inside of the balloon.
- **51.**  $1.00 \times 10^2$  kPa
- 52. 1.80 L
- **53.**  $1.8 \times 10^{1}$  L
- 54. 846 K (573 °C)
- **55.** High temperatures can sufficiently increase the pressure of the gas remaining in the container to cause it to explode.

**56.**  $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$ 

- **57.**  $1.10 \times 10^3$  kPa
- **58.**  $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$

When the pressure is constant,  $P_1 = P_2$ , so the pressure terms cancel, leaving the equation for Charles's law.

- **59**. Gas particles have a finite volume and are attracted to one another, especially at low temperatures.
- **60.** Its particles have no volume, no forces between them, and elastic collisions. It follows the gas laws at all temperatures and pressures.
- **61.** At low temperatures, gas particles are attracted to one another; the finite volume of gas particles is significant at high pressures.
- 62.  $3.56 \times 10^2$  kPa
- **63. a**. 5.6 × 10<sup>1</sup> L **b**. 6.72 L **c**. 7.84 L
- 64. equal numbers of particles

# Chapter 12 REVIEW

#### CONCEPT PRACTICE

- **45**. Heating a contained gas that is held at a constant volume increases its pressure. Why? *12.1*
- 46. What happens to gas particles when a gas is compressed? 12.1
- **47.** A metal cylinder contains 1 mol of nitrogen gas at STP. What will happen to the pressure if another mole of gas is added to the cylinder, but the temperature and volume do not change? *12.2*
- **48**. If a gas is compressed from 4 L to 1 L and the temperature remains constant, what happens to the pressure? *12.2*
- **49**. Why is Kelvin temperature specified in calculations that involve gases? *12.2*
- **50**. Describe what happens to the volume of a balloon when it is taken outside on a cold winter day. Explain why this happens. *12.2*
- 51. The gas in a closed container has a pressure of 3.00 × 10<sup>2</sup> kPa at 30 ℃ (303 K). What will the pressure be if the temperature is lowered to -172 ℃ (101 K)? 12.3
- 52. Calculate the volume of a gas (in L) at a pressure of  $1.00 \times 10^2$  kPa if its volume at  $1.20 \times 10^2$  kPa is  $1.50 \times 10^3$  mL. *12.3*
- 53. A gas with a volume of 4.0 L at 90.0 kPa expands until the pressure drops to 20.0 kPa. What is the new volume if the temperature remains constant? 12.3
- 54. A gas with a volume of  $3.00 \times 10^2$  into 10500 °C is heated until its volume is  $6.09 \times 10^2$  cm. What is the new temperature of the gas the pressure remains constant durations wheat-ing process? 12.2
- Why do aerosol containers display the warning, "Do not incinerate"? 12.3
- 56. State the combined gas law. 12.3
- 57. A sealed cylinder of gas contains nitrogen gas at  $1.00 \times 10^3$  kPa pressure and a temperature of 20 °C. The cylinder is left in the sun, and the temperature of the gas increases to 50 °C. What is the new pressure in the cylinder? *12.3*
- 58. Show how Charles's law can be derived from the combined gas law. 12.3

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- Explain why it is impossible for an ideal gas to exist. 12.4
- 60. Describe an ideal gas. 12.4
- 61. Explain the reasons why real gases deviate from ideal behavior. 12.4
- 62. If 4.50 g of methane gas (CH<sub>4</sub>) is introduced into an evacuated 2.00-L container at 35 °C, what is the pressure in the container? 12.4
- 63. Calculate the number of liters occupied at STP. 12.5
  - a. 2.5 mol N2(g)
  - **b**. 0.600 g H<sub>2</sub>(g)
  - c. 0.350 mol O<sub>2</sub>(g)
- 64. How would the number of particles of two gases compare if their partial pressures in a container were identical? 12.5
- 65. Which gas effuses faster: hydrogen or chlorine? How much faster? 12.5
- 66. Which gas effuses faster at the same temperature: molecular oxygen or atomic argon? 12.5
- 67. Calculate the ratio of the velocity of helium atoms to the velocity of neon atoms at the same temperature. 12.5
- Calculate the ratio of the velocity of helium atoms to fluorine molecules at the same temperature. 12.5

#### CONCEPT MASTERY

- 69. What can you conclude about the nature of the relationship between two variables with a quotient that is a constant?
- 70. A certain gas effuses four times as fast as onygen (O<sub>2</sub>). What is the molar mass of the gas?
- **71.** A 3.50-L gas sample at 20 °C and a pressure of 86.7 kPa expands to a volume of 8.00 L. The final pressure of the gas is 56.7 kPa. What is the final temperature of the gas, in degrees Celsius!
- 72. During an effusion experiment, it took 75 seconds for a certain number of moles of an unknown gas to pass through a tiny hole. Under the same conditions, the same number of moles of oxygen gas passed through the hole in 30 seconds. What is the molar mass of the unknown gas?

### (RITICAL THINKING

- 73. Choose the term that best completes the second relationship.
  - a. ideal gas: real gas
  - fiction:\_
  - (1) biography
  - (2) novel
  - (3) movie (4) nonfiction
  - b. Charles's law:temperature
    - Boyle's law:\_
    - (1) pressure
    - (2) volume
    - (3) ideal mass
    - (4) mass
  - c. volume: Charles's law
    - pressure:
    - (1) Boyle's law
    - (2) combined gas law
    - (3) Gay-Lussac's law
    - (4) temperature
  - d. inverse relationship: Boyle's law
    - direct relationship:
    - (1) absolute zero
    - (2) Avagadro's hypothesis
    - (3) Charles's law
    - (4) ideal gas law
  - e. kelvins: degrees Celsius
    - kilopascals:
    - (1) atmospheric pressure
    - (2) atmospheres
    - (3) pressure
    - (4) absolute zero
- 74. Gases will expand to fill a vacuum. Why Go Earth's atmospheric gases not escape into the near-vacuum of
  - space?
- 75. How does the vacuum used in Thermos<sup>™</sup> bottles prevent heat transfer?

- 76. What real gas
  - comes closest to having the characteristics of an ideal gas? Why?

#### **CUMULATIVE REVIEW**

- 77. Calculate the molar mass of each substance.
  - a. Ca(CH<sub>3</sub>CO<sub>2</sub>)<sub>2</sub>
  - b. H<sub>3</sub>PO<sub>4</sub>
  - c. C12H22O11
  - d.  $Pb(NO_3)_2$
- 78. Name each compound.
  - a. SnBr<sub>2</sub>
  - b. BaSO4
  - c. Mg(OH)<sub>2</sub> d. IF5
- 79. An atom of lead-206 weighs 17.16 times as much as an atom of carbon.
  - a. What is the molar mass of this isotope of lead? b. How many protons, electrons, and neutrons
  - are in this atom of lead?
- 80. How many kilojoules and kilocalories of heat are required to raise 40.0 g of water from -12 °C to 130 °C?
- 81. Write a balanced equation for each chemical reaction.
  - a. Calcium reacts with water to form calcium hydroxide and hydrogen gas.
  - b. Tetraphosphorus decoxide reacts with water to form phosphoric acid.
  - c. Mercury and oxygen are prepared by heating mercury(II) oxide.
  - d. Aluminum hydroxide and hydrogen sulfide form when aluminum sulfide reacts with water
- 82. Classify each of the reactions in Problem 81 based on type.
- 83. Calculate the molecular formula of each of the following compounds.
  - a. The empirical formula is C2H4O and the molar mass = 88 g/mol.
  - b. The empirical formula is CH and the molar mass = 104 g/mol.
  - c. The molar mass = 90 g/mol. The percent composition is 26.7% C, 71.1% O, and 2.2% H.
- 84. A piece of metal has a mass of 9.92 g and measures 4.5 cm × 1.3 cm × 1.6 mm. What is the density of the metal?
- 85. Calculate the percent composition of 2-propanol (C3H7OH).

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#### Hierier 12

- 65. At any temperature, hydrogen gas diffuses faster than chlorine gas by an approximate factor of six.
- 66. oxygen
- 67. 2.25:1
- 68. 3.08
- 69. The variables are directly proportional.
- 70. 2.0 g
- 71. 165 °C
- 72.  $2.0 \times 10^2$  g
- 73. a. (4)
  - **b**. (1)
  - C. (1)
  - d. (3)
  - e. (2)
- 74. The gases that make up the atmosphere, just like any other form of matter, are held near Earth by the force of gravity.
- 75. A vacuum contains no matter to allow the transfer of kinetic energy between molecules.
- 76. Helium gas; it is composed of small, monatomic atoms with little attraction for each other.
- **77. a**.  $1.58 \times 10^2$  g
  - **b**.  $9.80 \times 10^{1}$  g
  - **c.**  $3.42 \times 10^2$  g
- **d**.  $3.31 \times 10^2$  g 78. a. tin(II) bromide
  - b. barium sulfate
  - c. magnesium hydroxide
  - d. iodine pentafluoride
- **79. a.**  $2.06 \times 10^2$  g b. 82 protons, 82 electrons, 124 neutrons
- **80.**  $1.24 \times 10^2$  kJ;  $2.96 \times 10^1$  kcal
- **81. a.**  $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$ **b**.  $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$ **c.**  $2HgO \rightarrow 2Hg + O_2$ **d**.  $Al_2S_3 + 6H_2O \rightarrow 2Al(OH)_3 +$ 3H<sub>2</sub>S
- 82. a. single-replacement
  - **b**. combination
  - c. decomposition
  - d. double-replacement

85. 60.0% C, 13.3% H, 26.7% O

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- 83. a. C<sub>4</sub>H<sub>8</sub>O<sub>2</sub>
  - b. C<sub>8</sub>H<sub>8</sub> C. C2H2O4

84.  $11 \text{ g/cm}^3$ 

#### 16.1 and 16.2 Homework Answer Key



#7) A single bond is indicated by two dots or one line between two atoms. A double bond is indicated by four dots or two lines between atoms. A triple bond is indicated by six dots or three lines between atoms.

#8) a. Carbon monoxdide (CO) (or similar answers)
b. Ozone (O<sub>3</sub>) (or similar answers)
c. Nitrogen dioxide (NO<sub>2</sub>) (or similar answers)

#9) The structural formula identifies the atoms in the compound and their respective number and arrangement in the molecule.



#### 16.2 pg. 459 #14

#14) In a methane molecule, the four valence electron pairs repel each other, forming the corners of a tetrahedron in which the pairs are equidistant from each other. The angle between the bonds is 109.5°.

#### **16.3 Homework Answer Key**

#### 16.3 pg. 466 #21 to 24, 26

#21) Find the difference in electronegativity values for the two atoms of interest. If the difference is 0.0 to 0.5, 0.5 to 1.5, greater than 1.5, the bond is non-polar, polar covalent, or ionic, respectively.

#22) Dispersion forces (the weakest of the three) are caused by the motion of electrons. Dipole interactions are the attractions between the oppositely charged ends of polar molecules. Hydrogen bonding (the strongest of the three) occurs when a hydrogen atom bonded to a more electronegative atom is attracted to another highly electronegative atom.

#23)  $CCI_4$  is a nonpolar molecule because the atoms are oriented so that the bond polarities cancel.



#26) b., c., and d.

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#### Answers

- 27. Neon has an octet of valence electrons. A chlorine atom achieves an octet by sharing an electron with another chlorine atom.
- 28. a. ionic c. covalent b. ionic d. covalent
- 29. Ionic bonds depend on electrostatic attraction between ions. Covalent bonds depend on electrostatic attraction between shared electrons and nuclei of combining atoms.
- 30. A double covalent bond has four shared electrons; a triple covalent bond has six shared electrons.
- 31. A single atom of an element is usually the central atom.
- C. H:S:H 32. a. II: d. :I:N:I: b. :F:O:F:
- 33. a. :F:F: C. H:C::C:H b. H:Cl: d. H:C::N:
- 34. One atom contributes both electrons to a coordinate covalent bond as in CO.
- 35. An unshared pair of electrons is needed for a coordinate covalent bond. There are no unshared pairs in C-H or C-C bonds.
- 36. Molecules of each compound can be described by more than one electron dot structure.

**37.** 
$$[:\ddot{O}::\ddot{C}:\ddot{O}:]^{2^{-}} \longleftrightarrow [:\ddot{O}:C::\ddot{O}:]^{2^{-}}$$
  
 $:\ddot{O}: :\ddot{O}: :\ddot{O}: :\ddot{O}:$   
 $\longleftrightarrow :[:\ddot{O}:C:\ddot{O}:]^{2^{-}}$   
 $:\ddot{O}: :\ddot{O}:$ 

38. [:Ö:N::Ö:] ↔ [:Ö::N:Ö:]

- 39. The measured mass of a paramagnetic substance appears greater when measured in the presence of a magnetic field than when measured in the absence of a magnetic field.
- 40. a. diamagnetic
  - b. paramagnetic
  - c. paramagnetic
  - d. diamagnetic
- 41. Phosphorus and sulfur are larger atoms and have d orbital electrons.

# Chapter 16 REVIEW

#### CONCEPT PRACTICE

- 27. Explain why neon is monatomic but chlorine is diatomic. 16.1
- 28. Classify the following compounds as ionic or covalent. 16.1
  - a. MgCl<sub>2</sub> b. Na<sub>2</sub>S c. H<sub>2</sub>O d. H<sub>2</sub>S
- 29. Describe the difference between an ionic and a covalent bond. 16.1
- 30. How many electrons do two atoms in a double covalent bond share? How many in a triple covalent bond? 16.1
- 31. Based on the examples given in Section 16.1, state a general rule for determining which atom is the central one in a binary molecular compound. 16.1
- 32. Write plausible electron dot structures for the following substances. Each substance contains only single covalent bonds. 16.1 a. I<sub>2</sub> b. OF<sub>2</sub> c. H<sub>2</sub>S d. NI<sub>3</sub>
- 33. Draw the electron dot structure of each of the following molecules. 16.1 a. F2 b. HCl c. HCCH d. HCN
- 34. Characterize a coordinate covalent bond and give an example. 16.1
- 35. Explain why compounds containing C-N and C-O single bonds can form coordinate covalent bonds with H<sup>+</sup> but compounds containing only C-H and C-C single bonds cannot. 16.1
- 36. What is true for the electron dot structures of all compounds that exhibit resonance? 16.1
- 37. Draw resonance structures for the carbonate ion (CO<sub>3</sub><sup>2-</sup>). Each oxygen is attached to the car bon. 16.1
- 38. Using electron dot structures, draw at least two resonance structures for the nitrite ion (NO2-). The oxygens in NO2<sup>-</sup> are attached to the nitrogen. 16.1
- 39. How can you experimentally determine whether a substance is paramagnetic? 16.1
- 40. Predict whether the following species are diamagnetic or paramagnetic. 16.1 **b**. O<sub>2</sub><sup>-</sup> a. BF<sub>3</sub> c. NO<sub>2</sub> d. F2
- 41. Suggest a reason why phosphorus and sulfur have more than an octet in many of their compounds. Explain why compounds of nitrogen and oxygen never do? 16.1

43. Increasing bond dissociation energy is linked

44. Bond dissociation energy is defined as the en-

ergy needed to break one covalent bond.

2H:C 2 × 393 kJ/mol = 786 kJ/mol

to lower chemical reactivity.

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42. a. diamagnetic

b. diamagnetic

c. diamagnetic

d. diamagnetic

45. H:C:::C:H 1694 kJ/mol

C::::C 908 kJ/mol

- 42. Which of the following species would you predict to be diamagnetic? Paramagnetic? 16.1 b. OHc. H<sub>2</sub>O a. NO3 d. SO
- 43. What is the relationship between the magnitude of a molecule's bond dissociation energy and its expected chemical reactivity? 16,1
- 44. Explain what is meant by bond dissociation energy. 16.1
- 45. Assume the total bond energy in a molecule is the sum of the individual bond energies. Calculate the total bond energy in a mole of ethyne (C2H2). (Hint: Write the electron dot structure to determine the kinds of bonds. Then refer to Table 16.3.) 16.1
- 46. Draw molecular orbital diagrams for the possible diatomic molecule Li2. Would you expect Li2 to exist as a stable molecule? 16.2
- 47. What is the total number of sigma bonds and pi bonds in each molecule? 16.2

**a**. 
$$H-C \equiv N$$
 **b**.  $H-C-N=C=0$ 

- 48. Use VSEPR theory to predict the shapes of the following species. 16.2
  - a. CO2 c. SO<sub>3</sub> e. CO b. SiCl4 d. SCl<sub>2</sub> f. I3
- 49. The molecule CO2 has two carbon-oxygen double bonds. Describe the bonding in the CO2 molecule, which involves hybridized orbitals for carbon and oxygen. 16.2
- 63. What types of hybrid orbitals are involved in the bonding of the carbon atoms in the following molecules? 16.2 a. CH. c.  $HC \equiv CH$ 
  - h.  $H_2C = CH_2$  d.  $N \equiv C C \equiv N$
- 51. What must always be true if a covalent bond is to be polar? 16.3
- 52. The bonds between the following pairs of elements are covalent. Arrange them according to polarity, naming the most polar bond first. 16.3 a. H-Cl c. H-F e.H-H b. H-C d. H-O f. S-Cl
- 53. Arrange the following bonds in order of increasing ionic character. 16.3 a. Cl-F c. K-O e.S-0 b. N-N f. Li-F d.C-H
- 46.
  - The lithium molecule has two electrons in a bonding molecular orbital. Lithium theoretically should exist as a diatomic molecule. The Li2 molecule is moderately stable in the gaseous state.
- 47. a. 2 sigma, 2 pi

b. 6 sigma, 2 pi

- 54. Based on the information about molecular shapes in Section 16.2, which of these molecules would you expect to be polar? 16.3 c. CO<sub>2</sub> a. SO2 d. BF<sub>3</sub> h. H2S
- 5. Depict the hydrogen bonding between two ammonia molecules and between one ammonia molecule and one water molecule. 16.3
- 56. Which compound in each pair exhibits the stronger intermolecular hydrogen bonding? 16.3 a. H<sub>2</sub>S, H<sub>2</sub>O c. HBr, HCl

b. HCl, HF d. NH<sub>3</sub>, H<sub>2</sub>O

57. What is a hydrogen bond? 16.3

58. Why do compounds with strong intermolecular attractive forces have higher boiling points than compounds with weak intermolecular attractive forces? 16.3

#### CONCEPT MASTERY

- 59. Devise a hybridization scheme for PCl3 and predict the molecular shape based on this scheme.
- 60. Write two electron dot structures for the molecule N20. Predict the magnetic properties and shape of this molecule.
- fl. The chlorine and oxygen atoms in thionyl chloride (SOCl<sub>2</sub>) are bonded directly to the sulfur. Write an acceptable electron dot structure for thionyl chloride.
- 2. Explain why each electron dot structure is incorrect. Replace each structure with one that is more acceptable.

:Cl: c. H:C::0 a. C:N

8. Use VSEPR theory to predict the geometry of each of the following.

a. SiCl4 C. CCl4 b. CO32d. SCl<sub>2</sub>

- M. The following graph shows how the percent ionic character of a single bond varies according to the difference in electronegativity between the two elements forming the bond. Answer the following questions, using this graph and Table 14.2.

#### 48. a. linear

- b. tetrahedral
- e. linear

d. bent

- C. trigonal planar f. bent
- 49. The 2s and the 2p orbitals form two  $sp^2$  hybrid orbitals in the carbon atom. One  $sp^2$  hybrid orbital from each oxygen atom forms a sigma bond with the carbon atom. Pi bonds between each oxygen atom and the carbon are formed by the unhybridized 2p orbitals.



- a. What is the relationship between the percent ionic character of single bonds and the electronegativity difference of their elements?
- b. What electronegativity difference will result in a bond with a 50% ionic character?
- c. Estimate the percent ionic character of the bonds formed between (1) lithium and oxygen, (2) nitrogen and oxygen, (3) magnesium and chlorine, and (4) nitrogen and fluorine.
- 65. Using bond dissociation energies, estimate  $\Delta H$ for the following reaction.

 $CO(g) + 2H_2(g) \longrightarrow CH_3OH(g)$ 

- 66. Give the angles between the orbitals of each hybrid.
- **b**.  $sp^2$  hybrids **a**.  $sp^3$  hybrids c. sp hybrids
- 67. Describe the difference between a bonding molecular orbital and an antibonding molecular orbital. How do the energies of these orbitals compare?

#### (RITICAL THINKING

- 68. Choose the term that best completes the second relationship. a. house : lumber
  - bond :\_ (1) atoms
  - (3) valence electrons (4) octet rule (2) molecules
  - b. globe : Earth electron dots : (1) atoms (3) valence electrons
    - (2) molecules (4) octet rule
- 69. Make a list of the elements found in Table 16.2. What do the elements that form covalent bonds have in common?
- 70. Is there a clear difference between a very polar covalent bond and an ionic bond? Explain.

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- 50. a. sp<sup>3</sup> C. sp **b**.  $sp^2$ d. sp
- 51. The two atoms involved must have different electronegativities.
- 52. c., d., a., f., b., e.
- 53. b., d., e., a., c., f.

54.	a. polar	c. nonpolar			
	b. polar	d. nonpolar			

55. H-N:....H-56. a. H<sub>2</sub>O c. HCl b. HF d. H<sub>2</sub>O

**Miniferife** 

- 57. A hydrogen bond is formed by an electrostatic attraction between a hydrogen atom covalently bonded to a more electronegative atom and an unshared electron pair of a nearby atom.
- 58. They require more energy to separate the molecules.
- 59. The 3s and three 3p orbitals of phosphorus hybridize to form four sp<sup>3</sup> hybrid orbitals. The resulting shape is pyramidal with a bond angle of 107° between the sigma bonds.
- 60. :N::N::Ö: :N:::N::Ö: The molecule is diamagnetic and linear.
- 61. :CI:S:CI: :0:
- 62. a. Carbon does not have an ocet.

C::N:

b. A fluorine atom has 10 electrons.

:F:P:F: :F:

c. Carbon has too many  $e^{-}$  (10) and oxygen too few (4).

H:C::Ö: :Cl:

d. The atoms in boron trifluoride have 24 electrons to contribute, not 26. :F:B:F:

- :F:
- 63. a. tetrahedral, 109.5°
  - b. trigonal planar, 120°
- c. tetrahedral, 109.5°
  - d. bent, 105°
- 64. a. The percent ionic character increases as the difference in electronegativities increases. **b**. 1.6
  - C. (1) 85% (2) 10% (3) 62% (4) 23%
- **65.**  $\Delta H = -55 \text{ kJ/mol}$
- 66. a. 109.5° b. 120° c. 180°

#### 17.1 to 17.3 Homework Answer Key

#### 17.1 pg. 478 #1 to 4

#1) A slightly positive hydrogen atom on one water molecule is attracted to a pair of unshared electrons (lone pair) on the oxygen atom of another.

#2) Water molecules are hydrogen-bonded to each other, but not to air molecules. Net attraction is inward, minimizing water surface. Hydrogen bonding prevents escape of water molecules.

#3) A wetting agent that reduces surface tension; it interferes with hydrogen bonding of water.

#4) 4.18 J/(g □ °C)

#### 17.2 pg. 481 #5 and 6

#5) Hydrogen bonds hold liquid water molecules together, making it harder for them to escape to the gas state.

#6) Ice has a honeycomb-like structure of hydrogen-bonded water molecules, in which water molecules are closer together. The density of ice is less than that of water because of ice's open crystal structure. As a result, ice floats on water.

#### 17.3 pg. 488 #8 to 10 and 12

#8) 36.1%

#9) a.  $CaCl_2 + 6 H_2O \rightarrow CaCl_2 \square 6 H_2O$  b. 49.3%

#10) Polar solvents dissolve polar compounds, and nonpolar solvents dissolve nonpolar compounds. *Like* refers to the similar polarity of the substances.

#12) Acetic acid is the solute; water is the solvent.

### Clatter 17 REVIEW

# Answers

- **19.** Oxygen is more electronegative than hydrogen. Because the water molecules are bent, the bond polarities do not cancel.
- **20.** Surface molecules are attracted to the liquid molecules below but not to the air. Molecules inside the liquid are attracted in all directions.
- **21.** Some water molecules have sufficient energy to leave the surface of the liquid. Vapor pressure is the pressure of the water vapor above the liquid.
- **22.** Drops are spherical; objects denser than a liquid can float on the surface.
- It physically interferes with hydrogen bonding and reduces surface tension.
- 24. No; specific heat is expressed per gram.
- **25.** water,  $2.0 \times 10^4$  cal; iron,  $2.2 \times 10^3$  cal
- 26. 11 J
- 27. Ammonia molecules form hydrogen bonds, whereas methane molecules do not. Substances whose molecules form hydrogen bonds tend to have higher boiling points.
- **28.**  $1.3 \times 10^4$  cal
- 29. Ice floats in liquid water. Solids generally sink in their liquids because a substance is generally more dense in the solid state.
- 30. 8.0 kJ
- **31.** Bodies of water would freeze from the bottom up. This would kill many forms of aquatic life.
- **32.** Solutions are homogeneous mixtures in which a solute is dissolved in a solvent. Aqueous solutions are solutions that have water as the solvent.
- 33. The polar water molecules electrostatically attract ions and polar covalent molecules. Polar compounds will dissolve but nonpolar compounds are unaffected as they have no charges.
- **34.** No; the molecules and ions are smaller than the pores of the filter.

#### CONCEPT PRACTICE

- 19. Explain why water molecules are polar. 17.1
- **20.** Why do the particles at the surface of a liquid behave differently from those in the bulk of the liquid? *17.1*
- 21. Describe the origin of the vapor pressure of water. 17.1
- **22.** Describe some observable effects that can be produced by the surface tension of a liquid. *17.1*
- 23. What is a surfactant? Explain how it works. 17.1
- 24. Does the specific heat capacity of water vary depending on the quantity of water? Explain. 17.1
- 25. How many calories are required to heat 256 g of water from 20 °C to 99 °C? How many calories are required to heat the same mass of iron through the same range of temperature? 17.1
- **26**. How many joules are required to vaporize 5.0 mg of water at its boiling point? *17.2*
- 27. Why does ammonia have a much higher boiling point (-33 °C) than methane (-164 °C) even though their molar masses are almost the same? 17.2
- 28. How many calories are liberated when 24 g of steam at 100 °C condenses to liquid water at 100 °C? 17.2
- 29. What characteristic of ice distinguishes it from other solid substances? 17.2
- How many kilojoules are required to melt 24 x of ice at 0 °C? 17.2
- What would be some of the consequences differences di
- **32.** Distinguish between a solution and an aqueous solution. *17.3*
- 33. Why is water an excellent solvent for most ionic compounds and polar covalent molecules but not for nonpolar compounds? 17.3
- 34. Suppose an aqueous solution contains both sugar and salt. Can you separate either of these solutes from the water by filtration? Explain your reasoning. 17.3
- Describe how an ionic compound dissolves in water. 17.3

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- 36. Which of the following substances dissolve appreciably in water? Give reasons for your choice. 17.3
  - a. HCl d. MgSO<sub>4</sub> f. CaCO<sub>3</sub> b. NaI e. CH<sub>4</sub> g. gasoline c. NH<sub>3</sub>
- Why does molten sodium chloride conduct electricity? 17.3
- 38. What is the main distinction between an aqueous solution of a strong electrolyte and an aqueous solution of a weak electrolyte? 17.3
- Describe the water of hydration of a substance. 17.3
- 40. Write formulas for these hydrates. 17.3a. sodium sulfate decahydrate
  - b. magnesium sulfate heptahydrate
     c. barium hydroxide octahydrate
- 41. Name each hydrate. 17.3
  - a.  $SnCl_4 \cdot 5H_2O$ b.  $FeSO_4 \cdot 7H_2O$ c.  $BaBr_2 \cdot 4H_2O$ d.  $FePO_4 \cdot 4H_2O$
- Epsom salt (MgSO<sub>4</sub>·7H<sub>2</sub>O) changes to the monohydrate form at 150 °C. Write an equation for this change. 17.3
- Name two ways to distinguish a suspension from a colloid. 17.4
- 44. Solutions do not demonstrate the Tyndall effect. Why? 17.4
- 45. What makes a colloidal dispersion stable? 17.4
- 48. Define Brownian motion. 17.4

#### **EPT MASTERY**

- 47. Water has its maximum density at 4 °C. Explain why this is so, and discuss the consequences of this fact.
- 48. From your knowledge of intermolecular forces, arrange these liquids in order of increasing surface tension: water ( $H_2O$ ), hexane ( $C_6H_{14}$ ), ethanol ( $C_2H_5OH$ ).
- 49. Name these hydrates and determine the percent by mass of water in each.
  a. Na<sub>2</sub>CO<sub>3</sub>·H<sub>2</sub>O
  b. MgSO<sub>4</sub>·7H<sub>2</sub>O
- 50. If 5.00 g of steam (H<sub>2</sub>O)(g) at 100.0 °C condenses to liquid water (H<sub>2</sub>O)(l) that is cooled w 50.0 °C, how much heat is liberated in
  a. calories?
  b. kilocalories?

**35**. The positive ions are attracted by the negatively charged end of the polar water molecule; the negative ions are attracted by the positively charged end. As the ions are pulled away from the crystal, they are surrounded by the water molecules.

- 36. a. HCl (polar) will dissolve.
  - b. NaI (ionic) will dissolve.
  - c. NH<sub>3</sub> (polar) will dissolve.
  - d. MgSO<sub>4</sub> (ionic) will dissolve.
  - e. CH4 (nonpolar) will not dissolve.

f. CaCO<sub>3</sub> (strong ionic forces) will not dissolve.

- g. Gasoline (nonpolar) will not dissolve.37. Its ions are free to move toward an electrode.
- An aqueous solution of a strong electrolyte conducts electricity much better.
- 39. water in the crystal structure of a substance
- 40. a. Na<sub>2</sub>SO<sub>4</sub>·10H<sub>2</sub>O
  - b. MgSO4·7H<sub>2</sub>O
  - C. Ba(OH)<sub>2</sub>·8H<sub>2</sub>O
- 41. a. tin(IV) chloride pentahydrateb. iron(II) sulfate heptahydrate

- 51. Explain why ions become solvated in aqueous solution.
- 52. A block of ice at 0 °C has a mass of 176.0 g. How much heat must be added to change 25% of this mass of ice to liquid water at 0 °C? Express your answer in calories, kilocalories, joules, and kilojoules. What is the mass of ice remaining?
- 53. You have a solution containing either sugar or salt dissolved in water.
  - a. Can you tell which it is by visual inspection?
  - Give two ways by which you could easily tell which it is.
- 54. Water is a polar solvent; gasoline is a nonpolar solvent. Decide which compounds are more likely to dissolve in water and which are more likely to dissolve in gasoline.

a. sucrose  $(C_{12}H_{22}O_{11})$  c. methane  $(CH_4)$ b. Na<sub>2</sub>SO<sub>4</sub> d. KCl

- Match each term with the following descriptions. A description may apply to more than one term.
  - a. true solution c. suspension
  - b. colloidal
    - (1) does not settle out on standing
  - (2) heterogeneous mixture
  - (3) particle size less than 1.0 nm
  - (4) particles can be filtered out
  - (5) demonstrates Tyndall effect
  - (6) particles are invisible to the unaided eye
  - (7) homogenized milk
  - (8) salt water
  - (9) jelly
- Explain which properties of water are responsible for these occurrences.
  - a. Water in tiny cracks in rocks helps break up the rocks when it freezes.
  - b. Water beads up on a newly waxed car.
  - t. As you exercise and your body temperature increases, your body cools itself by producing sweat.
  - Temperatures below 28 °F damage grapevines. When severe frost is predicted, grape growers spray a mist of water on their vines.
  - e. An efficient way of heating a large building is to generate steam in a boiler and circulate it through pipes to radiators throughout the building.

- **57.** Explain why ethanol ( $C_2H_5OH$ ) will dissolve in both gasoline and water.
- 58. A 25.0 m  $\times$  10.0 m swimming pool is filled with fresh water to a depth of 1.7 m. The water temperature is initially 25 °C. How much heat must be removed from the water to change it all to ice at 0 °C? Express your answer in kilocalories and kilojoules.
- 59. Are all liquids soluble in each other? Explain.
- **60**. Write equations to show how these substances ionize or dissociate in water.
  - a.  $NH_4Cl$  c.  $HC_2H_3O_2$ b.  $Cu(NO_3)_2$  d.  $HgCl_2$
- 61. The following graph shows the density of water over the temperature range 0 °C to 20 °C.

a. What is the maximum density of water?

- b. At what temperature does the maximum density of water occur?
- c. Would it be correct to extend the smooth curve of the graph below 0 °C?



#### (RITICAL THINKING

**62**. Choose the term that best completes the second relationship.

a. plant : green	water molecule :
(1) polar	(3) frozen
(2) ionic	(4) vapor
b. east : west	condensation :
(1) boiling	(3) vaporization
(2) vapor pressure	(4) freezing
c. colloid : emulsion	cat :
(1) dog	(3) poodle
(2) Siamese	(4) fox

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- 46. the random movement of colloidal particles
- 47. Water molecules at 4°C are tightly packed and have maximum density. Below 4°C the water molecules arrange in a regular network because of the attractions between them. As a result, ice has a lower density than water and floats.
- 48. hexane, ethanol, water
- 49. a. sodium carbonate monohydrate, 14.5% H<sub>2</sub>O
  b. magnesium sulfate heptahydrate, 51.1% H<sub>2</sub>O

## Chapter 17 REVIEW

- **50. a**.  $2.95 \times 10^3$  cal **b**. 2.95 kcal
- 51. Ions in solution are surrounded by water molecules. Negative ions are attracted to the hydrogen atoms and positive ions are attracted to the oxygen atoms.
- **52.** 3.52 × 10<sup>3</sup> cal, 3.52 kcal, 1.47 × 10<sup>4</sup> J, 14.7 kJ; 132.0 g H<sub>2</sub>O
- 53. a. no

**b**. drying to examine the crystals, testing for electrical conductivity, doing a flame test

- 54. a. water
   c. gasoline

   b. water
   d. water

   55. 1. a, b
   4. c
   7. b

   2. b, c
   5. b, c
   8. a
- **3**. a **6**. a, b **9**. b **56**. **a**. Water expands when it
  - freezes to ice. **b**. Water is polar and wax is nonpolar, and water has a high surface tension.
  - c. Sweat evaporates to carry away the heat of vaporization.
    d. Freezing damages plants because water inside them expands as it freezes. Water on the surface gives off the heat of fusion to the plants when it freezes.
  - **e**. Steam carries the heat of vaporization.
- 57. Ethanol has a polar hydroxyl end (—OH) that dissolves in water, and a nonpolar hydrocarbon end (C<sub>2</sub>H<sub>5</sub>—) that dissolves in gasoline.
- **58.**  $4.5 \times 10^7$  kcal,  $1.9 \times 10^8$  kJ
- **59.** No; nonpolar molecules do not dissolve in polar molecules.
- 60. a.  $NH_4Cl \rightarrow NH_4^+ + Cl^$ b.  $Cu(NO_3)_2 \rightarrow Cu^{2+} + 2NO_3^$ c.  $HC_2H_3O_2 \rightarrow H^+ + C_2H_3O_2^$ d.  $HgCl_2 \rightarrow Hg^{2+} + 2Cl^-$
- 61. a. 1.00 g/mL b. 4°C
  c. No; the density of ice is 0.917 g/mL at 0°C. There would be a break in the curve at 0°C as liquid water at 0°C changes to solid water (ice) at 0°C.
- 62. a. 1
  - **b**. 3
  - **c.** 2

48. A suspension is a mixture with large particles

that settle on standing. A colloid is a mixture

4. The molecules or ions are too small to have reflective surfaces.

45. the random motion of the dispersion medium <sup>molecules</sup>

¢. barium bromide tetrahydrate ¢. iron(III) phosphate tetrahydrate

Q. MgSO<sub>4</sub>·7H<sub>2</sub>O → MgSO<sub>4</sub>·H<sub>2</sub>O + 6H<sub>2</sub>O

#### 18.1 and 18.2 Homework Answer Key

**18.1 pg. 507 #1 and 2; pg. 508 #3 to 7** #1) 0.44 g/L

#2) 2.6 atm

#3) agitation, temperature, particle size

#4) by using Henry's Law

#5) 272 g NaCl

#6) a. Add solvent. b. Add solute until no more will dissolve.

#7) Solubility usually increases with temperature.

**18.2 pg. 511 #8 to 11; pg. 513 #12, 13; pg. 514 #14, 15; pg. 515 #16 to 23** #8) 0.10 M

#9) 2.8 M

#10) 0.142 mol

#11) 0.50 mol CaCl<sub>2</sub>; 56 g CaCl<sub>2</sub>

#12) 47.5 mL

#13) Use a pipet to transfer 50 mL of the 1.0 M solution to a 250-mL volumetric flask. Then add distilled water up to the mark.

#14) 5% (v/v)

#15) 12 mL

#16) 0.25 g MgSO<sub>4</sub>

#17) 3.6% CuSO<sub>4</sub> (m/v)

#18) Molarity is found by dividing the number of moles of solute by the number of liters of solution.

#19) Solvent is added to the concentrated solution until the desired molarity is achieved.

#20) Percent by volume equals the volume of solute per volume of solution. Percent (mass/volume) equals mass of solute (in g) per volume of solution (in mL).

#21) a. 0.627 M CuSO<sub>4</sub> b. 0.040 M NaHSO<sub>4</sub>

#22) a. 125 mL b. 0.80 L c. 2.5 mL

#23) 5.0% K<sub>2</sub>SO<sub>4</sub> (m/v)

### 18.3 and 18.4 Homework Answer Key 18.3 pg. 519 #24 to 27

#24) The introduction of solute molecules reduces the number of solvent molecules with enough kinetic energy to escape.

#25) Because vapor pressure has been reduced, more kinetic energy is needed to reach the boiling point. For a solution to freeze, it must lose more kinetic energy than the pure solvent does.

#26) Concentrated; the concentrated solution has more dissolved particles. Boiling point elevation is proportional to the number of dissolved particles.

#27) a.	$MgF_{2(aq)}$	b. KI <sub>(aq)</sub>	c. Kl <sub>(aq)</sub>
	(~~~)	(44)	(0.97)

**18.4 pg. 521 #28, 29; pg. 522 #30, 31; pg. 524 #32, 33; pg. 525 #34 to 39** #28) 12.6 g

#29) 0.285 *m* or 0.285 mol/kg

- #30)  $C_2H_5OH = 0.190$   $H_2O = 0.810$
- #31)  $CCI_4 = 0.437$   $CHCI_3 = 0.563$
- #32) 101.37°C
- #33) 114 g NaCl
- #34) 39.2 g/mol
- #35) 169 g/mol
- #36) 4.0 kg H<sub>2</sub>O
- #37) FeCl<sub>3</sub>

#38) The molality of a solution is directly proportional to its boiling point elevation and freezing point depression.  $\Delta T_f = i K_f$  (molality) and  $\Delta T_b = i K_b$  (molality)

#39) 4.95°C

### Depunde

### REVIEW

#### Answers

- **40**. The solvent is the substance in which the solute is dissolved.
- **41.** Random collisions of the solvent molecules with the solute particles provide enough force to overcome gravity.
- **42.** Miscible liquids dissolve in each other; immiscible liquids do not.
- **43.** Solubility is the amount of solute dissolved in a given amount of solvent to form a saturated solution at a given temperature. A saturated solution contains the maximum possible amount of solute at that temperature. An unsaturated solution contains less dissolved solute than a saturated solution.
- 44. 555 g AgNO<sub>3</sub>
- 45. Particles of solute crystallize.
- **46.** No; if there were undissolved solute, the excess solute would come out of solution.
- **47.** Solubility increases with pressure.
- 48. a. 0.016 g/L b. 0.047 g/L
- **49.** Molarity provides the exact number of moles of solute per liter of solution. Dilute and concentrated are relative terms and are not quantitative.
- **50. a**. 1.3*M* KCl **b**. 0.33*M* MgCl<sub>2</sub>
- **51.** the number of moles of solute dissolved in one liter of solution
- **53. a.** 23 g NaCl **b.** 2.0 g MgCl<sub>2</sub>
- **54. a.** 3.3% KCl **b.** 1.6% NaNO<sub>3</sub>
- **55. a**. 100.26 °C **b**. 101.54 °C
- **56.** Add 27.0 g H<sub>2</sub>O to 32.0 g CH<sub>3</sub>OH.
- 57. 1*M* solution: 1 mol of solute in 1 L of solution; 1*m* solution: 1 mol of solute in 1000 g of solvent

# Chapter 18 REVIEW

#### CONCEPT PRACTICE

- **40**. Name and distinguish between the two components of a solution. *18.1*
- **41**. Explain why the dissolved component does not settle out of a solution. *18.1*
- 42. Explain miscible and immiscible. 18.1
- Define solubility, saturated solution, and unsaturated solution. 18.1
- 44. What mass of AgNO<sub>3</sub> can be dissolved in 250 g of water at 20 °C? 18.1
- **45.** If a saturated solution of sodium nitrate is cooled, what change might you observe? *18.1*
- **46**. Can a solution with undissolved solute be supersaturated? Explain. *18.1*
- 47. What is the effect of pressure on the solubility of gases in liquids? 18.1
- 48. The solubility of methane, the major component of natural gas, in water at 20 °C and 1.00 atm pressure is 0.026 g/L. If the temperature remains constant, what will be the solubility of this gas at the following pressures? 18.1
  a. 0.60 atm
  - **b**. 1.80 atm
- 49. Having a measure of the molarity of a solution is more meaningful than knowing whether a solution is dilute or concentrated. Explain. 18.2
- 50. Calculate the molarity of each relation. 18.2
  a. 1.0 mol KCl in 750 set. of a factories
  b. 0.50 mol MgCl<sub>2</sub> in 1.5 transmitters
- 51. Define molarity. 18.2
- 52. Calculate the moles and grams of solute in each solution. 18.2a. 1.0 L of 0.50M NaCl
  - **b.**  $5.0 \times 10^2$  mL of 2.0*M* KNO<sub>3</sub>
  - **c.** 250 mL of 0.10M CaCl<sub>2</sub>
  - d. 2.0 L of 0.30M Na<sub>2</sub>SO<sub>4</sub>
- 53. Calculate the grams of solute required to make the following solutions. 18.2a. 2.5 L of normal saline solution
  - (0.90% NaCl (m/v))
  - **b**. 50.0 mL of 4.0% MgCl<sub>2</sub> (m/v)
- 54. What is the concentration (in % (m/v)) of the following solutions? 18.2a. 20.0 g KCl in 0.60 L of solution
  - **b**. 32 g NaNO<sub>3</sub> in 2.0 L of solution
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- 58. a. −4.46 °C b. −22.3 °C
- **59. a**. −1.12 °C **b**. −0.74 °C **c**. −1.49 °C
- **60.** mass of solute, mass of solvent, boiling point of solvent, boiling point of solution
- 61. a. The freezing-point depression is twice as great for solute B; solute B must provide twice as many particles in solution.
  b. Solute A forms a saturated solution.
- **62.**  $\Delta T_{\rm f} = -9.60 \,^{\circ}{\rm C}; \, \Delta T_{\rm b} = +4.74 \,^{\circ}{\rm C}$

- 55. What is the boiling point of each solution? Ib:
  a. 0.50 mol glucose in 1000 g H₂O
  b. 1.50 mol NaCl in 1000 g H₂O
- 56. Describe how you would make an aqueous solution of methanol (CH<sub>3</sub>OH) in which the mole fraction of methanol is 0.40. 18.4
- Distinguish between a 1M solution and a lm solution. 18.4
- 58. What is the freezing point of each solution? 144
   a. 1.40 mol Na<sub>2</sub>SO<sub>4</sub> in 1750 g H<sub>2</sub>O
   b. 0.60 mol MgSO<sub>4</sub> in 100 g H<sub>2</sub>O
- **59.** Determine the freezing points of each 0.20m aqueous solution. 18.4
  - a. K<sub>2</sub>SO<sub>4</sub>
  - **b**. CsNO<sub>3</sub>
  - **c**. Al(NO<sub>3</sub>)<sub>3</sub>
- 60. What laboratory measurements must be make to find the molar mass of a solute in a boling point elevation experiment? 18.4

### CONCEPT MASTERY

- **61.** Varying numbers of moles of two different solutes, A and B, were added to identical quantities of water. The graph shows the freezing point of each of the solutions formed with various amounts of solutes.
  - a. Explain the relative slopes of the two lines between 0 and 2 mol of solute added.
  - b. Why does the freezing point for solutionA not continue to drop as amounts of soluteA are added beyond 2.4 mol?



62. Calculate the freezing- and boiling-point changes for a solution containing 12.0 gcl naphthalene (C<sub>10</sub>H<sub>8</sub>) in 50.0 g of benzene.

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#### 20.1, 20.2 and 21.1 Homework Answer Key

#### 20.1: pg. 579 #1 to 5; pg. 609 #34

#1) а. с.	) a. Hydrofluoric acid c. Potassium Hydroxide			<ul><li>b. Nitric acid</li><li>d. Sulfuric acid</li></ul>			
#2) a.	$H_2CrO_{4\ (aq)}$	b. Fe(OH) <sub>3</sub>	C.	$HI_{(aq)}$	d.	LiOH	
#3) a.	Base	b. Both	C.	Both	d.	Acid	
#4) a.	Ba(OH) <sub>2</sub>	b. HBr <sub>(aq)</sub>	C.	RbOH	d.	H₂Se	
#5) а. с.	Hydrofluoric ac Carbonic Acid	bid	b. d.	Chloric acid Aluminum hydroxid	de		

#34) a.	HNO <sub>2 (aq)</sub>	b.	AI(OH) <sub>3 (aq)</sub>	C.	$H_2 Se_{(aq)}$
d.	Sr(OH) <sub>2 (aq)</sub>	e.	H <sub>3</sub> PO <sub>4 (aq)</sub>	f.	CH <sub>3</sub> COOH (aq)

# **20.2:** pg. 582 #7; pg. 586 #8 and 9; pg. 587 #10 and 11; pg. 588 #12 and 13; pg. 589 #14 and 15; pg. 593 #16 to 18; pg. 609 #36 to 45

#7) a. Bas	sic b.	. Basic	c. Acidic	d. Neutral
#8)a.4.0	0 b.	. 3.00	c. 9.00	
#9) a. 12.	00 b.	. 2.00	c. 4.00	
#10) a. 1.	0 x 10 <sup>-5</sup> M b.	. 1.0 x 10 <sup>-7</sup> M	c. 1.0 x 10 <sup>-6</sup> M	d. 1.0 x 10 <sup>-3</sup> M
#11) a. 1.	0 x 10 <sup>-4</sup> M b.	. 1.0 x 10 <sup>-11</sup> M	c. 1.0 x 10 <sup>-8</sup> M	
#12) a. 5.	30 b.	. 9.08	c. 6.57	
#13) a. 9.	63 b.	. 9.30	c. 3.65	
#14) a. 5.	0 x 10 <sup>-8</sup> M b.	. 1.6 x 10 <sup>-2</sup> M	c. 8.9 x 10 <sup>-8</sup> M	d. 2.0 x 10 <sup>-7</sup> M
#15) a. 2.	0 x 10 <sup>-7</sup> M b.	. 6.3 x 10 <sup>-13</sup> M	c. 1.1 x 10 <sup>-7</sup> M	d. 5.0 x 10 <sup>-8</sup> M

#16) a. Hydroxide-ion concentration is greater.

b. Hydrogen-ion concentration is greater.

c. The concentrations are equal.

#17)a. 6.00	b. 4.00	c. 12.00	d. 3.00

#18) a. 1.0 x 10<sup>-8</sup> M b. 1.0 x 10<sup>-5</sup> M c. 1.0 x 10<sup>-2</sup> M

#36) 1.0 x  $10^{-7}$  M for both [H<sup>+</sup>] and [OH<sup>-</sup>] at 25°C

#37) The negative logarithm of the hydrogen-ion concentration.

#38) The hydrogen-ion concentration of pure water at 25°C is  $1.0 \times 10^{-7}$  M. The negative logarithm, or pH, of this concentration is 7.00

#39) a. pH = 2.00, acidic c. pH = 6.00, acidic		b. pH = 12.00, basic d. pH = 6.00, acidic	2
#40) a. 1.0 x 10 <sup>−10</sup> M	b. 1.0 x 10 <sup>-6</sup> M	c. 1.0 x 10 <sup>-2</sup> M	
#41) a. 5.62	b. 8.04	c. 6.3 x 10 <sup>-14</sup> M	d. 2.0 x 10 <sup>-7</sup> M
#42) 1.6 x 10⁻⁴ M			
#43) a. KOH $\rightarrow$ K <sup>+</sup> + OH <sup>-</sup>	b. Mg(OH) <sub>2</sub>	$\rightarrow Mg^{2+} + 2 OH^{-}$	
#44) Acids ionize to give hy	drogen ions in aqueou	us solution. Bases ionize to giv	ve hvdroxide ions

#44) Acids ionize to give hydrogen ions in aqueous solution. Bases ionize to give hydroxide ions in aqueous solution.

#45)	) a. Base	b. Acid	c. base	d. Acid	e. Acid	f. Acid

### 21.1 pg. 616 #1 and 2; pg. 618 #3 and 4; pg. 624 #17 to 19; pg. 640 #36 to 39; pg. 641 # 62

#1) 4.68 mol KOH

#2) 0.20 mol NaOH

#3) 56 mL HCI

#4) 0.128 M

#17) To a measured quantity of an acid (or base) of unknown concentration, drops of a base (or acid) of known concentration are added until the equivalents of base equal the equivalents of acid. The unknown concentration is calculated from setting a stoichiometry calculation line with mole ratio.

#18) a.  $H_2SO_4 + 2 \text{ KOH} \rightarrow 2 \text{ HOH} + \text{K}_2SO_4$ Salt = Potassium Sulfate b. HCl + LiOH  $\rightarrow$  HOH + LiCl Salt = Lithium Chloride c.  $2 H_3 PO_4 + 3 Ca(OH)_2 \rightarrow 6 HOH + Ca_3(PO_4)_2$ Salt = Calcium Phosphate d. 2 HNO<sub>3</sub> + Mg(OH)<sub>2</sub>  $\rightarrow$  2 HOH + Mg(NO<sub>3</sub>)<sub>2</sub> Salt = Magnesium Nitrate #19) a. 0.2 mol b. 2 mol c. 0.2 mol #36) Acid + Base  $\rightarrow$  Water + Salt #37) a.  $HNO_3$  + KOH  $\rightarrow$  HOH + KNO<sub>3</sub> b. 2 HCl + Ca(OH)<sub>2</sub>  $\rightarrow$  2 HOH + CaCl<sub>2</sub> c.  $H_2SO_4$  + 2 NaOH  $\rightarrow$  2 HOH + Na<sub>2</sub>SO<sub>4</sub> #38) Neutralization occurs. #39) a. 1.40 M b. 2.61 M #62) a. 2 HCl + Mg(OH)<sub>2</sub>  $\rightarrow$  2 HOH + MgCl<sub>2</sub> b. 2 HCl + CaCO<sub>3</sub>  $\rightarrow$  H<sub>2</sub>O + CO<sub>2</sub> + CaCl<sub>2</sub> c. AI(OH)\_3 + 3 HCI  $\rightarrow$  3 HOH + AICI\_3