Student Solutions Manual

Introductory Chemistry
A Foundation
SIXTH EDITION

Zumdahl

James F. Hall
University of Massachusetts Lowell
# Contents

<table>
<thead>
<tr>
<th>CHAPTER</th>
<th>TITLE</th>
<th>PAGE</th>
</tr>
</thead>
<tbody>
<tr>
<td>PREFACE</td>
<td></td>
<td>V</td>
</tr>
<tr>
<td>CHAPTER 1</td>
<td>CHEMISTRY: AN INTRODUCTION</td>
<td>1</td>
</tr>
<tr>
<td>CHAPTER 2</td>
<td>MEASUREMENTS AND CALCULATIONS</td>
<td>2</td>
</tr>
<tr>
<td>CHAPTER 3</td>
<td>MATTER</td>
<td>14</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 1–3</td>
<td></td>
<td>17</td>
</tr>
<tr>
<td>CHAPTER 4</td>
<td>CHEMICAL FOUNDATIONS: ELEMENTS, ATOMS, AND IONS</td>
<td>21</td>
</tr>
<tr>
<td>CHAPTER 5</td>
<td>NOMENCLATURE</td>
<td>29</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 4 &amp; 5</td>
<td></td>
<td>37</td>
</tr>
<tr>
<td>CHAPTER 6</td>
<td>CHEMICAL REACTIONS: AN INTRODUCTION</td>
<td>41</td>
</tr>
<tr>
<td>CHAPTER 7</td>
<td>REACTIONS IN AQUEOUS SOLUTION</td>
<td>45</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 6 &amp; 7</td>
<td></td>
<td>55</td>
</tr>
<tr>
<td>CHAPTER 8</td>
<td>CHEMICAL COMPOSITION</td>
<td>59</td>
</tr>
<tr>
<td>CHAPTER 9</td>
<td>CHEMICAL QUANTITIES</td>
<td>84</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 8 &amp; 9</td>
<td></td>
<td>105</td>
</tr>
<tr>
<td>CHAPTER 10</td>
<td>ENERGY</td>
<td>110</td>
</tr>
<tr>
<td>CHAPTER 11</td>
<td>MODERN ATOMIC THEORY</td>
<td>116</td>
</tr>
<tr>
<td>CHAPTER 12</td>
<td>CHEMICAL BONDING</td>
<td>123</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 10–12</td>
<td></td>
<td>135</td>
</tr>
<tr>
<td>CHAPTER 13</td>
<td>GASES</td>
<td>141</td>
</tr>
<tr>
<td>CHAPTER 14</td>
<td>LIQUIDS AND SOLIDS</td>
<td>157</td>
</tr>
<tr>
<td>CHAPTER 15</td>
<td>SOLUTIONS</td>
<td>163</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 13–15</td>
<td></td>
<td>180</td>
</tr>
<tr>
<td>CHAPTER 16</td>
<td>ACIDS AND BASES</td>
<td>187</td>
</tr>
<tr>
<td>CHAPTER 17</td>
<td>EQUILIBRIUM</td>
<td>195</td>
</tr>
<tr>
<td>CUMULATIVE REVIEW CHAPTERS 16 &amp; 17</td>
<td></td>
<td>203</td>
</tr>
<tr>
<td>CHAPTER 18</td>
<td>OXIDATION–REDUCTION REACTIONS/ELECTROCHEMISTRY</td>
<td>209</td>
</tr>
<tr>
<td>CHAPTER 19</td>
<td>RADIOACTIVITY AND NUCLEAR ENERGY</td>
<td>219</td>
</tr>
<tr>
<td>CHAPTER 20</td>
<td>ORGANIC CHEMISTRY</td>
<td>224</td>
</tr>
<tr>
<td>CHAPTER 21</td>
<td>BIOCHEMISTRY</td>
<td>235</td>
</tr>
</tbody>
</table>
Preface

This guide contains the even-numbered solutions for the end-of-chapter problems in the sixth editions of *Introductory Chemistry*, *Introductory Chemistry: A Foundation*, and *Basic Chemistry* by Steven S. Zumdahl. Several hundred new problems and questions have been prepared for the new editions of the text, which we hope will be of even greater help to your students in gaining an understanding of the fundamental principles of chemistry.

We have tried to give the most detailed solutions possible to all the problems even though some problems give repeat drill practice on the same subject. Our chief attempt at brevity is to give molar masses for compounds without showing the calculation (after the subject of molar mass itself has been discussed). We have also made a conscious effort in this guide to solve each problem in the manner discussed in the textbook. The instructor, of course, may wish to discuss alternative methods of solution with his or her students.

One topic that causes many students concern is the matter of significant figures and the determination of the number of digits to which a solution to a problem should be reported. To avoid truncation errors in the solutions contained in this guide, the solutions typically report intermediate answers to one more digit than appropriate for the final answer. The final answer to each problem is then given to the correct number of significant figures based on the data provided in the problem.

Many very dedicated people have worked long hours to prepare this guide. Particular thanks go to Bess Deck, Liz Hogan and Sean McGann at Houghton Mifflin for their patience, competence, and good cheer.

James F. Hall
University of Massachusetts Lowell

James_Hall@uml.edu
CHAPTER 1

Chemistry: An Introduction

CHAPTER ANSWERS

2. The answer will depend on student examples.

4. Answer depends on student response.

6. This answer depends on your own experience, but consider the following examples: oven cleaner (the label says it contains sodium hydroxide; it converts the burned-on grease in the oven to a soapy material that washes away); drain cleaner (the label says it contains sodium hydroxide; it dissolves the clog of hair in the drain); stomach antacid (the label says it contains calcium carbonate; it makes me belch and makes my stomach feel better); hydrogen peroxide (the label says it is a 3% solution of hydrogen peroxide; when applied to a wound, it bubbles); depilatory cream (the label says it contains sodium hydroxide; it removes unwanted hair from skin).

8. The scientist must recognize the problem and state it clearly, propose possible solutions or explanations, and then decide through experimentation which solution or explanation is best.

10. Answer depends on student response. A quantitative observation must include a number. For example, “There are two windows in this room.” represents a quantitative observation, but, “The walls of this room are yellow.” is a qualitative observation.

12. A natural law is a summary of observed, measurable behavior that occurs repeatedly and consistently. A theory is our attempt to explain such behavior. The conservation of mass observed during chemical reactions is an example of a natural law. The idea that the universe began with a “big bang” is an example of a theory.

14. Chemistry is not just a set of facts that have to be memorized. To be successful in chemistry, you have to be able to apply what you have learned to new situations, new phenomena, and new experiments. Rather than just learning a list of facts or studying someone else’s solution to a problem, your instructor hopes you will learn how to solve problems yourself, so that you will be able to apply what you have learned in future circumstances.

16. In real-life situations, the problems and applications likely to be encountered are not simple textbook examples. One must be able to observe an event, hypothesize a cause, and then test the hypothesis. One must be able to carry what has been learned in class forward to new, different situations.
CHAPTER 2
Measurements and Calculations

CHAPTER ANSWERS

2. “Scientific notation” means we have to put the decimal point after the first significant figure and then express the order of magnitude of the number as a power of ten. So we want to put the decimal point after the first two:

\[
\frac{63.5 \text{ g}}{125 \text{ cm}^3}
\]

To be able to move the decimal point three places to the left in going from 2,421 to 2.421, means I will need a power of \(10^3\) after the number, where the exponent three shows that I moved the decimal point three places to the left.

\[2,421 \rightarrow 2.421 \times 10^{\text{some power}} = 2.421 \times 10^3\]

4. If the number is greater than one, the exponent is positive; if the number is less than one, the exponent is negative.
   a. negative
   b. positive
   c. negative
   d. positive

6.
   a. already a decimal number
   b. The decimal point must be moved three places to the right: 1995
   c. The decimal point must be moved two places to the right: 199.5
   d. The decimal point must be moved four places to the left: 0.0001995

8.
   a. The decimal point has to be moved three places to the right, so the exponent will be 3.
   b. The decimal point has to be moved four places to the left, so the exponent will be –4.
   c. The decimal point has to be moved two places to the right, so the exponent will be 2.
   d. This number is already written as a power of ten but is not in standard scientific notation; realizing that \(0.994 = 9.94 \times 10^1\) and combining exponents gives the final exponent as 1.

10.
   a. five spaces to the right
   b. three spaces to the left
   c. eight spaces to the left
12.  
   a. The decimal point must be moved four places to the left: 0.0004915.
   b. The decimal point must be moved three spaces to the right: 994.
   c. The decimal point must be moved three spaces to the left: 0.02495.
   d. already a decimal number
   e. The decimal point must be moved two spaces to the right: 693.4.
   f. The decimal point must be moved one space to the right: 693.4.

14.  
   a. \( \frac{1}{0.00032} = 3.1 \times 10^3 \)
   b. \( \frac{10^3}{10^{-3}} = 1 \times 10^6 \)
   c. \( \frac{10^3}{10^3} = 1 \times (1 \times 10^0) \); any number divided by itself is unity.
   d. \( \frac{1}{55,000} = 1.8 \times 10^{-5} \)
   e. \( \frac{(10^5)(10^4)(10^{-4})}{10^{-2}} = 1 \times 10^7 \)
   f. \( \frac{43.2}{(4.32 \times 10^{-5})} = \frac{4.32 \times 10^4}{4.32 \times 10^{-5}} = 1.00 \times 10^6 \)
   g. \( \frac{(4.32 \times 10^{-5})}{432} = \frac{4.32 \times 10^{-5}}{4.32 \times 10^2} = 1.00 \times 10^{-7} \)
   h. \( \frac{1}{(10^5)(10^{-6})} = 1/(10^{-1}) = 1 \times 10^1 \)

16. time  

18.  
   a. \( 10^{-2} \)
   b. \( 10^6 \)
   c. \( 10^9 \)
   d. \( 10^{-1} \)
   e. \( 10^{-3} \)
   f. \( 10^{-6} \)

20. Since a pound is 453.6 g, the 125-g can will be slightly less than ¼ pound.

22. Since 1 in = 2.54 cm, the nail is approximately an inch long.

24. Since a liter is slightly more than a quart, the 2–L bottle is larger.

26. 1.62 m is approximately 5 ft, 4 in. The woman is slightly taller.

28.  
   a. centimeter
   b. meter
c. kilometer

30. d (The other units would give very large numbers for the distance.)

32. Table 2.6 indicates that a nickel coin weighs about 5 g; 1 kg = 1000 g: so
\[
\frac{1000 \text{ g}}{5 \text{ g/nickel}} = \text{about 200 nickels.}
\]

34. uncertainty

36. The scale of the ruler shown is marked only to the nearest tenth of a centimeter; writing 2.850 would imply that the scale was marked to the nearest hundredth of a centimeter (and that the zero in the thousandths place had been estimated).

38.

a. one
b. infinite (definition)
c. infinite (fixed number)
d. two
e. probably two

40. It is better to round off only the final answer and to carry through extra digits in intermediate calculations. If there are enough steps to the calculation, rounding off in each step may lead to a cumulative error in the final answer.

42.

a. \(9.96 \times 10^{-1}\)
b. \(4.40 \times 10^{3}\)
c. \(8.22 \times 10^{-1}\)
d. \(4.00 \times 10^{-9}\)
e. \(8.42 \times 10^{-2}\)

44.

a. \(8.8 \times 10^{-4}\)
b. \(9.375 \times 10^{4}\)
c. \(8.97 \times 10^{-1}\)
d. \(1.00 \times 10^{3}\)

46. The total mass would be determined by the number of decimal places available on the readout of the scale/balance. For example, if a balance whose readout is to the nearest 0.01 g were used, the total mass would be reported to the second decimal place. For example, 32.05 g + 29.15 g + 31.09 g would be reported as 92.29 g to the second decimal place. For the calculation 44.05 g + 33.91 g + 48.38 g, the sum would be reported as 126.34 g (a total of five significant figures, but given to the second decimal place).

48. Most calculators would display 0.66666666. If the 2 and 3 were experimentally determined numbers, this quotient would imply far too many significant figures.

50. none
52.  
   a.  5.16 (The answer can be given only to the second decimal place because 3.04 is known only to the second decimal place.)  
   b.  2423 (2.423 × 10³; the numbers have to be expressed in the same power of 10 before adding 2119 + 304 = 2423)  
   c.  0.516 (5.159 × 10⁻¹; both numbers are the same power of 10)  
   d.  2423 (This is the same problem as part b only the numbers were not expressed in scientific notation.)  

54.  
   a. one (the factor of two has only one significant figure)  
   b. four (the sum within the parentheses will contain four significant figures)  
   c. two (based on the factor 4.7 × 10⁻⁶ having only two significant figures)  
   d. three (based on the factor 63.9 having only three significant figures)  

56.  
   a.  \[ \frac{2.0944 + 0.0003233 + 12.22}{7.001} = \frac{14.3147233}{7.001} = 2.045 \]  
   b.  \[ \frac{(1.42 \times 10^2 + 1.021 \times 10^3)}{(3.1 \times 10^{-1})} = \frac{(142 + 1021)}{(3.1 \times 10^{-1})} = \frac{1163}{3.1 \times 10^{-1}} = 3751 = 3.8 \times 10^3 \]  
   c.  \[ \frac{(9.762 \times 10^{-3})(1.43 \times 10^2 + 4.51 \times 10^1)}{(143 + 45.1)} = \frac{(9.762 \times 10^{-3})(188.1)}{188.1} = 5.19 \times 10^{-5} \]  
   d.  \[ (6.1982 \times 10^{-4})^2 = (6.1982 \times 10^{-4})(6.1982 \times 10^{-4}) = 3.8418 \times 10^{-7} \]  

58. an infinite number (a definition)  

60. \[ \frac{1000 \text{ mL}}{1 \text{ L}} \text{ and } \frac{1 \text{ L}}{1000 \text{ mL}} \]  

62. \[ \frac{1 \text{ lb}}{0.79} \]  

64.  
   a.  \[ 2.23 \text{ m} \times \frac{1.094 \text{ yd}}{1 \text{ m}} = 2.44 \text{ yd} \]  
   b.  \[ 46.2 \text{ yd} \times \frac{1 \text{ m}}{1.094 \text{ yd}} = 42.2 \text{ m} \]  
   c.  \[ 292 \text{ cm} \times \frac{1 \text{ in}}{2.54 \text{ cm}} = 115 \text{ in} \]  
   d.  \[ 881.2 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 2238 \text{ cm} \]  
   e.  \[ 1043 \text{ km} \times \frac{1 \text{ mi}}{1.6093 \text{ km}} = 648.1 \text{ mi} \]
Chapter 2: Measurements and Calculations

f. \[ 445.5 \text{ mi} \times \frac{1.6093 \text{ km}}{1 \text{ mi}} = 716.9 \text{ km} \]

g. \[ 36.2 \text{ m} \times \frac{1 \text{ km}}{1000 \text{ m}} = 0.0362 \text{ km} \]

h. \[ 0.501 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{100 \text{ cm}}{1 \text{ m}} = 5.01 \times 10^4 \text{ cm} \]

66.

a. \[ 5.25 \text{ oz} \times \frac{1 \text{ lb}}{16 \text{ oz}} = 0.328 \text{ lb} \]

b. \[ 125 \text{ g} \times \frac{1 \text{ lb}}{453.59 \text{ g}} = 0.276 \text{ lb} \]

c. \[ 125 \text{ g} \times \frac{1 \text{ lb}}{453.59 \text{ g}} \times \frac{16 \text{ oz}}{1 \text{ lb}} = 4.41 \text{ oz} \]

d. \[ 125 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.125 \text{ L} \]

e. \[ 125 \text{ mL} \times \frac{1.057 \text{ qt}}{1000 \text{ mL}} \times \frac{2 \text{ pt}}{1 \text{ qt}} = 0.264 \text{ pt} \]

f. \[ 2.5 \text{ mi} \times \frac{1.6093 \text{ km}}{1 \text{ mi}} = 4.0 \text{ km} \]

g. \[ 2.5 \text{ mi} \times \frac{1.6093 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} = 4.0 \times 10^3 \text{ m} \]

h. \[ 2.5 \text{ mi} \times \frac{1.6093 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{100 \text{ cm}}{1 \text{ m}} = 4.0 \times 10^5 \text{ cm} \]

68. \[ 190 \text{ mi} = 1.9 \times 10^2 \text{ mi} \text{ to two significant figures} \]

\[ 1.9 \times 10^2 \text{ mi} \times \frac{1 \text{ km}}{0.62137 \text{ mi}} = 3.1 \times 10^5 \text{ km} \]

\[ 3.1 \times 10^5 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} = 3.1 \times 10^8 \text{ m} \]

\[ 1.9 \times 10^2 \text{ mi} \times \frac{5280 \text{ ft}}{1 \text{ mi}} = 1.0 \times 10^6 \text{ ft} \]

70. \[ 1 \times 10^{-10} \text{ m} \times \frac{100 \text{ cm}}{1 \text{ m}} = 1 \times 10^{-8} \text{ cm} \]

\[ 1 \times 10^{-8} \text{ cm} \times \frac{1 \text{ in}}{2.54 \text{ cm}} = 4 \times 10^{-9} \text{ in.} \]

\[ 1 \times 10^{-8} \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = 0.1 \text{ nm} \]
72. freezing
74. 273
76. Fahrenheit (F)
78. \( T_c = T_k - 273 \)
   a. \(-210 + 273 = 63 \text{ K}\)
   b. \(275 \text{ K} - 273 = 2^\circ \text{C}\)
   c. \(778 \text{ K} - 273 = 505^\circ \text{C}\)
   d. \(778 + 273 = 1051 \text{ K}\)
80. \( T_f = 1.80(T_c) + 32 \)
   a. \(1.80(78.1) + 32 = 173^\circ \text{F}\)
   b. \(1.80(40.) + 32 = 104^\circ \text{F}\)
   c. \(1.80(-273) + 32 = -459^\circ \text{F}\)
   d. \(1.80(32) + 32 = 90.0^\circ \text{F}\)
82. \( T_f = 1.80(T_c) + 32 \quad T_c = (T_f - 32)/1.80 \)
   a. \(275 - 273 = 2^\circ \text{C}\)
   b. \((82 - 32)/1.80 = 28^\circ \text{C}\)
   c. \(1.80(-21) + 32 = -5.8^\circ \text{F} (-6^\circ \text{F})\)
   d. \((-40 - 32)/1.80 = -40^\circ \text{C}\) (the Celsius and Fahrenheit temperatures are the same).
84. \( \text{g/cm}^3 \) (g/mL)
86. 100 in\(^3\)
88. Density is a characteristic property that is always the same for a pure substance.
90. copper
92. density \(= \frac{\text{mass}}{\text{volume}}\)
   a. \(d = \frac{122.4 \text{ g}}{5.5 \text{ cm}^3} = 22 \text{ g/cm}^3\)
   b. \(v = 0.57 \text{ m}^3 \times \left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 = 5.7 \times 10^5 \text{ cm}^3\)

   \[ d = \frac{1.9302 \times 10^3 \text{ g}}{5.7 \times 10^5 \text{ cm}^3} = 0.034 \text{ g/cm}^3 \]
   c. \(m = 0.0175 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 17.5 \text{ g}\)
\[d = \frac{17.5 \text{ g}}{18.2 \text{ mL}} = 0.962 \text{ g/mL} = 0.962 \text{ g/cm}^3\]

d. \[v = 0.12 \text{ m}^3 \times \left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 = 1.2 \times 10^5 \text{ cm}^3\]

\[d = \frac{2.49 \text{ g}}{1.2 \times 10^5 \text{ cm}^3} = 2.1 \times 10^{-5} \text{ g/cm}^3\]

94. \[1.0 \text{ kg} = 1.0 \times 10^3 \text{ g}\]

\[1.0 \times 10^3 \text{ g} \times \frac{1 \text{ mL}}{1.097 \text{ g}} = 0.91 \text{ L} \text{ (two significant figures)}\]

96. \[m = 3.5 \text{ lb} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 1.59 \times 10^3 \text{ g}\]

\[v = 1.2 \times 10^4 \text{ in.}^3 \times \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^3 = 1.97 \times 10^5 \text{ cm}^3\]

\[d = \frac{1.59 \times 10^3 \text{ g}}{1.97 \times 10^5 \text{ cm}^3} = 8.1 \times 10^{-3} \text{ g/cm}^3\]

The material will float.

98. \[5.25 \text{ g} \times \frac{1 \text{ cm}^3}{10.5 \text{ g}} = 0.500 \text{ cm}^3 = 0.500 \text{ mL}\]

\[11.2 \text{ mL} + 0.500 \text{ mL} = 11.7 \text{ mL}\]

100.

a. \[50.0 \text{ cm}^3 \times \frac{19.32 \text{ g}}{1 \text{ cm}^3} = 966 \text{ g}\]

b. \[50.0 \text{ cm}^3 \times \frac{7.87 \text{ g}}{1 \text{ cm}^3} = 394 \text{ g}\]

c. \[50.0 \text{ cm}^3 \times \frac{11.34 \text{ g}}{1 \text{ cm}^3} = 567 \text{ g}\]

d. \[50.0 \text{ cm}^3 \times \frac{2.70 \text{ g}}{1 \text{ cm}^3} = 135 \text{ g}\]

102.

a. \[3.011 \times 10^{23} = 301,100,000,000,000,000,000,000\]

b. \[5.091 \times 10^9 = 5,091,000,000\]

c. \[7.2 \times 10^2 = 720\]

d. \[1.234 \times 10^5 = 123,400\]

e. \[4.32002 \times 10^{-4} = 0.000432002\]

f. \[3.001 \times 10^{-2} = 0.03001\]
g. \[ 2.9901 \times 10^{-7} = 0.00000029901 \]

h. \[ 4.2 \times 10^{-1} = 0.42 \]

104.

a. centimeters
b. meters
c. kilometers
d. centimeters
e. millimeters

106.

a. \[ 36.2 \text{ blim} \times \frac{1400 \text{ kryll}}{1 \text{ blim}} = 5.07 \times 10^4 \text{ kryll} \]

b. \[ 170 \text{ kryll} \times \frac{1 \text{ blim}}{1400 \text{ kryll}} = 0.12 \text{ blim} \]

c. \[ 72.5 \text{ kryll}^2 \times \left( \frac{1 \text{ blim}}{1400 \text{ kryll}} \right)^2 = 3.70 \times 10^{-5} \text{ blim}^2 \]

108. \[ 52 \text{ cm} \times \frac{1 \text{ in}}{2.54 \text{ cm}} = 20. \text{ in} \]

110. \[ 1 \text{ lb} \times \frac{1 \text{ kg}}{2.2 \text{ lb}} \times \frac{1 \text{ euro}}{1 \text{ kg}} = \frac{\$1.20}{2.45 \text{ euro}} = \$1.33 \text{ per pound} \]

112. \[ ^\circ X = 1.26^\circ C + 14 \]

114. \[ d = \frac{36.8 \text{ g}}{10.5 \text{ L}} = 3.50 \text{ g/L} \quad (3.50 \times 10^{-3} \text{ g/cm}^3) \]

116. For ethanol, \[ 100. \text{ mL} \times \frac{0.785 \text{ g}}{1 \text{ mL}} = 78.5 \text{ g} \]

For benzene, \[ 1000 \text{ mL} \times \frac{0.880 \text{ g}}{1 \text{ mL}} = 880. \text{ g} \]

total mass, \[ 78.5 + 880. = 959 \text{ g} \]

118.

a. negative
b. negative
c. positive
d. zero
e. negative
120.
   a. 2; positive
   b. 11; negative
   c. 3; positive
   d. 5; negative
   e. 5; positive
   f. 0; zero
   g. 1; negative
   h. 7; negative

122.
   a. 1; positive
   b. 3; negative
   c. 0; zero
   d. 3; positive
   e. 9; negative

124.
   a. The decimal point must be moved five places to the left; \(2.98 \times 10^{-5} = 0.0000298\).
   b. The decimal point must be moved nine places to the right; \(4.358 \times 10^{9} = 4,358,000,000\).
   c. The decimal point must be moved six places to the left; \(1.9928 \times 10^{-6} = 0.0000019928\).
   d. The decimal point must be moved 23 places to the right; \(6.02 \times 10^{23} = 602,000,000,000,000,000,000,000\).
   e. The decimal point must be moved one place to the left; \(1.01 \times 10^{-1} = 0.101\).
   f. The decimal point must be moved three places to the left; \(7.87 \times 10^{-3} = 0.00787\).
   g. The decimal point must be moved seven places to the right; \(9.87 \times 10^{7} = 98,700,000\).
   h. The decimal point must be moved two places to the right; \(3.7899 \times 10^{2} = 378.99\).
   i. The decimal point must be moved one place to the left; \(1.093 \times 10^{-1} = 0.1093\).
   j. The decimal point must be moved zero places; \(2.9004 \times 10^{0} = 2.9004\).
   k. The decimal point must be moved four places to the left; \(3.9 \times 10^{-4} = 0.00039\).
   l. The decimal point must be moved eight places to the left; \(1.904 \times 10^{-8} = 0.00000001904\).

126.
   a. \(1/10^{2} = 1 \times 10^{-2}\)
   b. \(1/10^{-2} = 1 \times 10^{2}\)
   c. \(55/10^{3} = \frac{5.5 \times 10^{1}}{1 \times 10^{3}} = 5.5 \times 10^{-2}\)
d. \( \frac{3.1 \times 10^6}{10^{-3}} = \frac{3.1 \times 10^6}{1 \times 10^{-3}} = 3.1 \times 10^9 \)

e. \( (10^6)^{1/2} = 1 \times 10^3 \)

f. \( \frac{(10^6)(10^4)}{(10^7)} = \frac{(1 \times 10^6)(1 \times 10^4)}{(1 \times 10^7)} = 1 \times 10^8 \)

g. \( \frac{1}{0.0034} = \frac{1}{3.4 \times 10^3} = 2.9 \times 10^2 \)

h. \( \frac{3.453}{10^{-4}} = \frac{3.453}{1 \times 10^{-4}} = 3.453 \times 10^4 \)

128. Kelvin, K

130. centimeter

132. 0.105 m

134. 1 kg (100 g = 0.1 kg)

136. 10 cm (1 cm = 10 mm)

138. 2.8 (the hundredths place is estimated)

140.

a. 0.000426

b. \( 4.02 \times 10^{-5} \)

c. \( 5.99 \times 10^6 \)

d. 400.

e. 0.00600

142.

a. 2149.6 (The answer can be given only to the first decimal place because 149.2 is known only to the first decimal place)

b. \( 5.37 \times 10^3 \) (The answer can be given only to two decimal places because 4.34 is known only to two decimal places. Moreover, since the power of ten is the same for each number, the calculation can be performed directly.)

c. Before performing the calculation, the numbers have to be converted so that they contain the same power of ten.
\[ 4.03 \times 10^{-2} - 2.044 \times 10^{-3} = 4.03 \times 10^{-2} - 0.2044 \times 10^{-2} = 3.83 \times 10^{-2} \] (The answer can be given only to the second decimal place because \( 4.03 \times 10^{-2} \) is known only to the second decimal place.)

d. Before performing the calculation, the numbers have to be converted so that they contain the same power of ten.
\[ 2.094 \times 10^5 - 1.073 \times 10^6 = 2.094 \times 10^5 - 10.73 \times 10^5 = -8.64 \times 10^5 \]
Chapter 2: Measurements and Calculations

144.

a. \((2.9932 \times 10^4)(2.4443 \times 10^2 + 1.0032 \times 10^1) = (2.9932 \times 10^4)(2.4443 \times 10^3 + 1.0032 \times 10^1) = (2.9932 \times 10^4)(25.446 \times 10^1) = 7.6166 \times 10^6\)

b. \((2.34 \times 10^2 + 2.443 \times 10^{-1})/(0.0323) = (2.34 \times 10^2 + 0.002443 \times 10^2)/(0.0323) = (2.34 \times 10^2)/(0.0323) = 7.24 \times 10^3\)

c. \((4.38 \times 10^{-3})^2 = 1.92 \times 10^{-5}\)

d. \((5.9938 \times 10^{-6})^{1/2} = 2.4482 \times 10^{-3}\)

146. \[
\frac{1 \text{ year}}{12 \text{ months}} = \frac{12 \text{ months}}{1 \text{ year}}
\]

148.

a. \[908 \text{ oz} \times \frac{1 \text{ lb}}{16 \text{ oz}} \times \frac{1 \text{ kg}}{2.2046 \text{ lb}} = 25.7 \text{ kg}\]

b. \[12.8 \text{ L} \times \frac{1 \text{ qt}}{0.94633 \text{ L}} \times \frac{1 \text{ gal}}{4 \text{ qt}} = 3.38 \text{ gal}\]

c. \[125 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ qt}}{0.94633 \text{ L}} = 0.132 \text{ qt}\]

d. \[2.89 \text{ gal} \times \frac{4 \text{ qt}}{1 \text{ gal}} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.09 \times 10^4 \text{ mL}\]

e. \[4.48 \text{ lb} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 2.03 \times 10^3 \text{ g}\]

f. \[550 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.0567 \text{ qt}}{1 \text{ L}} = 0.58 \text{ qt}\]

150. Assuming exactly 6 gross, 864 pencils

152.

a. Celsius temperature = \((175 - 32)/1.80 = 79.4^\circ \text{C}\)  
   Kelvin temperature = \(79.4 + 273 = 352 \text{ K}\)

b. \(255 - 273 = -18 \, ^\circ \text{C}\)

c. \((-45 - 32)/1.80 = -43^\circ \text{C}\)

d. \(1.80(125) + 32 = 257^\circ \text{F}\)

154. \[85.5 \text{ mL} \times \frac{0.915 \text{ g}}{1 \text{ mL}} = 78.2 \text{ g}\]
156. \[ m = 155 \text{ lb} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 7.031 \times 10^4 \text{ g} \]

\[ \nu = 4.2 \text{ ft}^3 \times \left( \frac{12 \text{ in}}{1 \text{ ft}} \right)^3 \times \left( \frac{2.54 \text{ cm}}{1 \text{ in}} \right)^3 = 1.189 \times 10^5 \text{ cm}^3 \]

\[ d = \frac{7.031 \times 10^4 \text{ g}}{1.189 \times 10^5 \text{ cm}^3} = 0.59 \text{ g/cm}^3 \]

158. \[ T_F = 1.80(T_C) + 32 \]

a. 23 °F
b. 32 °F
c. –321 °F
d. –459 °F
e. 187 °F
f. –459 °F
CHAPTER 3

Matter

CHAPTER ANSWERS

2. forces
4. liquids
6. gaseous
8. The stronger the interparticle forces, the more rigid is the sample overall.
10. Gases are easily compressed into smaller volumes whereas solids and liquids are not. Because a gaseous sample consists mostly of empty space, it is the empty space that is compressed when pressure is applied to a gas.
12. This is a chemical change; the mercury disappears and is replaced by an orange solid.
14. Magnesium is malleable and ductile.
16. The most common physical changes are changes in state: solid to liquid, liquid to gaseous, solid to gaseous (and their opposites).
18.
   a. physical: The iron is only being heated.
   b. chemical: The sugars in the marshmallow are being reduced to carbon.
   c. chemical: Most strips contain a peroxide that decomposes.
   d. chemical: The bleach oxidizes dyes in the fabric.
   e. physical: Evaporation is only a change of state.
   f. physical: The salt is modifying only the physical properties of the solution, not undergoing a chemical reaction.
   g. chemical: The drain cleaner breaks bonds in the hair.
   h. physical: Students will most likely reply that this is a physical change since the perfume is evaporating; the sensation of smell, however, depends on chemical processes.
   i. physical: The sublimation is only a change of state.
   j. physical: The wood is only being physically divided into smaller pieces.
   k. chemical: The cellulose in the wood is reacting with oxygen gas.
20. Compounds consist of two or more elements combined together chemically in a fixed composition, no matter what their source may be. For example, water on earth consists of molecules containing one oxygen atom and two hydrogen atoms. Water on Mars (or any other planet) has the same composition.
22. compounds
24. Typically, the properties of a compound and the elements that constitute it are very different. Consider the properties of liquid water and the hydrogen and oxygen gases from which the water was prepared. Consider the properties of sodium chloride (table salt) and the sodium metal and chlorine gas from which it might have been prepared.

26. Assuming the magnesium and sulfur had been measured out in exactly the correct ratio for complete reaction, what would remain after heating would be a pure compound. If there were an excess of either magnesium or sulfur, however, the material left after reaction would be a mixture of the compound and the excess reagent.

28. solutions: window cleaner, shampoo, rubbing alcohol
   mixtures: salad dressing, jelly beans, the change in my pocket

30.
   a. primarily a pure compound, but fillers and anti-caking agents may have been added
   b. mixture
   c. mixture
   d. pure substance

32.
   a. homogeneous
   b. heterogeneous
   c. heterogeneous
   d. heterogeneous
   e. homogeneous

34. Consider a mixture of salt (sodium chloride) and sand. Salt is soluble in water, sand is not. The mixture is added to water, stirred to dissolve the salt and then filtered. The salt solution passes through the filter; the sand remains on the filter. The water can then be evaporated from the salt.

36. The solution is heated to vaporize (boil) the water. The water vapor is then cooled so that it condenses back to the liquid state, and the liquid is collected. After all the water is vaporized from the original sample, pure sodium chloride will remain. The process consists of physical changes.

38. Since \( X \) is a pure substance, the fact that two different solids form when electrical current is passed indicates that \( X \) must be a compound.

40. Because vaporized water is still the same substance as solid water, no chemical reaction has occurred. Sublimation is a physical change.

42. far apart
44. chemical
46. chemical
48. electrolysis

50.
   a. heterogeneous
   b. heterogeneous
c. heterogeneous (Unless you work hard to get all the lumps out!)
d. Although strictly heterogeneous, it may appear homogeneous.
e. heterogeneous

52. Answer depends on student’s response

54. physical, chemical

56. O₂ and P₄ are both still elements, even though the ordinary forms of these elements consist of molecules containing more than one atom (but all atoms in each respective molecule are the same). P₂O₅ is a compound because it is made up of two or more different elements (not all the atoms in the P₂O₅ molecule are the same).

58. Assuming there is enough water present in the mixture to have dissolved all the salt, filter the mixture to separate out the sand from the mixture. Then distill the filtrate (consisting of salt and water), which will boil off the water, leaving the salt.

60. The most obvious difference is the physical states: Water is a liquid under room conditions; hydrogen and oxygen are both gases. Hydrogen is flammable. Oxygen supports combustion. Water does neither.
Cumulative Review Chapters 1–3

ANSWERS

2. By now, after having covered three chapters in this book, it is hoped that you have adopted an “active” approach to your study of chemistry. You may have discovered, perhaps through a disappointing grade on a quiz (though we hope not), that you have to get really involved with chemistry. You can’t just sit and take notes or just look over the solved examples in the textbook. You have to learn to solve problems. You have to learn how to interpret problems and how to reduce them to the simple mathematical relationships you have studied. Whereas in some courses you might get by on just giving back on exams the facts or ideas presented in class, in chemistry you have to be able to extend and synthesize what has been discussed and to apply the material to new situations. Don’t get discouraged if this is difficult at first: it’s difficult for everyone at first.

4. It is difficult sometimes for students (especially beginning students) to understand why certain subjects are required for a given college major. The faculty of your major department, however, have collectively many years of experience in the subject in which you have chosen to specialize. They really do know which courses will be helpful to you in the future. They may have had trouble with the same courses that now give you trouble, but they realize that all the work will be worth it in the end. Some courses you take, particularly in your major field itself, have obvious and immediate utility. Other courses, oftentimes chemistry included, are provided to give you a general background knowledge that may prove useful in understanding your own major or other subjects related to your major. In perhaps a burst of bravado, chemistry has been called “the central science” by one team of textbook authors. This moniker is very true; however, in order to understand biology, physics, nutrition, farming, home economics, or whatever, it helps to have a general background in chemistry.

<table>
<thead>
<tr>
<th>Commonly Used Prefixes in the Metric System</th>
</tr>
</thead>
<tbody>
<tr>
<td>Prefix</td>
</tr>
<tr>
<td>mega-</td>
</tr>
<tr>
<td>kilo-</td>
</tr>
<tr>
<td>deci-</td>
</tr>
<tr>
<td>centi-</td>
</tr>
<tr>
<td>milli-</td>
</tr>
<tr>
<td>nano-</td>
</tr>
</tbody>
</table>

The metric system is in use in most of the world because its system of units and multiples is simple to remember, and the system permits easy conversion between units. The various multiples and subdivisions of the basic units are based on factors of ten, which are also the basis for our number system. The United States uses a historical system in which there is no simple relationship among most units. Although several attempts have been made to gradually change the United States over to the metric system, no widespread support for the program has been achieved. Although 2-liter soda bottles were accepted without too much complaint (since they
replaced a similar-sized 2-quart bottle), the thought of having a size 85 waist (in cm) may be repugnant to too many Americans! There obviously also would be a great cost to industry to retool all machinery and measurement devices in metric units.

6. Whenever a scientific measurement is made, we always employ the instrument or measuring device we are using to the limits of its precision. On a practical basis, this usually means that we estimate our reading of the last significant figure of the measurement. An example of the uncertainty in the last significant figure is given for measuring the length of a pin in the text in Figure 2.5. Scientists appreciate the limits of experimental techniques and instruments and always assume that the last digit in a number representing a measurement has been estimated. Because the last significant figure in every measurement is assumed to be estimated, it is never possible to exclude uncertainty from measurements. The best we can do is to try to improve our techniques and instruments so that we get more significant figures for our measurements.

8. Dimensional analysis is a method of problem solving that pays particular attention to the units of measurement and uses these units as if they were algebraic symbols that multiply, divide, and cancel. Consider the following example. A dozen eggs costs $1.25. Suppose we want to know how much one egg costs and also how much three dozen of eggs will cost. To solve these problems, we need to make use of two equivalence statements:

$$1 \text{ dozen eggs} = 12 \text{ eggs}$$
$$1 \text{ dozen eggs} = $1.25$$

The first of these equivalence statements is obvious; everyone knows that 12 eggs is “equivalent” to one dozen. The second statement also expresses an equivalence; if you give the grocer $1.25, he or she will give you a dozen eggs. From these equivalence statements we can construct the conversion factors we need to answer the two questions. For the first question (what does one egg cost) we can set up the calculation as follows

$$\frac{$1.25}{12 \text{ eggs}} = $0.104 = $0.10$$

as the cost of one egg. Similarly, for the second question (the cost of 3 dozen eggs), we can set up the conversion as follows

$$3 \text{ dozen} \times \frac{$1.25}{1 \text{ dozen}} = $3.75$$

as the cost of three dozen eggs. See Section 2.6 of the text for how we construct conversion factors from equivalence statements.

10. Defining what scientists mean by “matter” often seems circular to students. Scientists say that matter is something that “has mass and occupies space” without ever really explaining what it means to “have mass” or to “occupy space”! The concept of matter is so basic and fundamental that it becomes difficult to give a good textbook definition other than to say that matter is the “stuff” of which everything is made. Matter can be classified and subdivided in many ways depending on what we are trying to demonstrate.

On the most fundamental basis, all matter is composed of tiny particles (such as protons, electrons, neutrons, and the other subatomic particles). On one higher level, these tiny particles are combined in a systematic manner into units called atoms. Atoms, in turn, may be combined to constitute molecules. And finally, large groups of molecules may be placed together to form a bulk sample of substance that we can see.
Matter can also be classified as to the physical state a particular substance happens to take. Some substances are solids, some are liquids, and some are gases. Matter can also be classified as to whether it is a pure substance (one type of molecule) or a mixture (more than one type of molecule) and furthermore whether a mixture is homogeneous or heterogeneous.

12. Chemists tend to give a functional definition of what they mean by an “element,” defining an element as a fundamental substance that cannot be broken down into any simpler substances by chemical methods. Compounds, on the other hand, can be broken down into simpler substances (the elements of which the compound is composed). For example, sulfur and oxygen are both elements (sulfur occurs as S₈ molecules and oxygen as O₂ molecules). When sulfur and oxygen are placed together and heated, the compound sulfur dioxide (SO₂) forms. When we analyze the sulfur dioxide produced, we notice that each and every molecule consists of one sulfur atom and two oxygen atoms and on a mass basis, consists of 50% each of sulfur and oxygen. We describe this by saying that sulfur dioxide has a constant composition. The fact that a given compound has constant composition is usually expressed in terms of the mass percentages of the elements present in the compound. The reason the mass percentages are constant is because of a constant number of atoms of each type present in the compound’s molecules. If a scientist anywhere in the universe analyzed sulfur dioxide, he or she would find the same composition. If a scientist finds something that does not have the same composition, then the substance cannot be sulfur dioxide.

14. a. The decimal point must be moved five places to the left; 2.29 \times 10^5
b. The decimal point must be moved two places to the right; 421
c. The decimal point must be moved five places to the left; 0.0000593
d. The decimal point must be moved one place to the left, and the exponents must be combined; 1.93 \times 10^{-3}
e. The decimal point must be moved seven places to the right; 9.3 \times 10^7
f. The decimal point must be moved three places to the right and the exponents must be combined; 3.18 \times 10^1

16. a. four (two before the decimal point, two after the decimal point)
b. two (based on 1.9)
c. two
d. two (one before the decimal point, one after the decimal point)
e. three
f. two (based on two significant figures in the denominator)
g. four
h. five (three before the decimal point, two after the decimal point)

18. density = mass/volume  
   mass = volume \times density  
   volume = mass/density

   a. density = \frac{4.21 \text{ g}}{4.31 \text{ mL}} = 0.977 \text{ g/mL}

   b. mass = 1210 \text{ mL} \times \frac{0.891 \text{ g}}{1 \text{ mL}} = 1.08 \times 10^3 \text{ g}
c. volume = \( \frac{225 \text{ g}}{9.21 \text{ g/mL}} \) = 24.4 mL

d. mass = 24.5 mL \times \frac{1.31 \text{ g}}{1 \text{ mL}} = 32.1 \text{ g}

e. volume = \( \frac{5280 \text{ g}}{1.81 \text{ g/mL}} \) = 2920 mL = 2.92 L

f. density = \( \frac{72.4 \text{ g}}{82.4 \text{ mL}} \) = 0.879 g/mL
CHAPTER ANSWERS

2. Robert Boyle

4. There are at least 115 elements presently known: of these 88 occur naturally and the remainders are manmade. Table 4.1 lists the most common elements on the Earth.

6. Trace elements are those elements that are present in only tiny amounts in the body, but are critical for many bodily processes and functions.

8. The symbols for these elements are based upon their names in other languages.

10.

a. 8
b. 5
c. 2
d. 9
e. 13
f. 12
g. 6
h. 11
i. 7
j. 1

12. praeseodymium Pr
lawrencium Lr
californium Cf
nobelium No
hafnium Hf
14. **B**: barium, Ba; berkelium, Bk; beryllium, Be; bismuth, Bi; bohrium, Bh; boron, B; bromine, Br

**N**: neodymium, Nd; neon, Ne; neptunium, Np; nickel, Ni; niobium, Nb; nitrogen, N; nobelium, No

**P**: palladium, Pd; phosphorus, P; platinum, Pt; plutonium, Pu; polonium, Po; potassium, K; praseodymium, Pr; promethium, Pm; protactinium, Pa

**S**: samarium, Sm; scandium, Sc; seaborgium, Sg; selenium, Se; silicon, Si; silver, Ag; sodium, Na; strontium, Sr; sulfur, S

16.

a. Elements are made of tiny particles called atoms.

b. All the atoms of a given element are identical.

c. The atoms of a given element are different from those of any other element.

d. A given compound always has the same relative numbers and types of atoms.

e. Atoms are neither created nor destroyed in chemical processes. A chemical reaction simply changes the way the atoms are grouped together.

18. According to Dalton all atoms of the same element are identical; in particular, every atom of a given element has the same mass as every other atom of that element. If a given compound always contains the same relative numbers of atoms of each kind, and those atoms always have the same masses, then it follows that the compound made from those elements would always contain the same relative masses of its elements.

20.

a. PbO₂

b. CoCl₃

c. C₆H₁₂O₆

d. Al₂O₃

e. Na₂CO₃

f. CaH₂

22.

a. False; Rutherford’s bombardment experiments with metal foil suggested that the alpha particles were being deflected by coming near a dense, positively charged atomic nucleus.

b. False; The proton and the electron have opposite charges, but the mass of the electron is much smaller than the mass of the proton.

c. True

24. The protons and neutrons are found in the nucleus. The protons are positively charged; the neutrons have no electrical charge. Protons and neutrons each have approximately the same mass.

26. neutron; electron

28. Because they are located in the exterior regions of the atom, it is the electrons of an atom that most interact with other atoms and are therefore most responsible for the atom’s chemical behavior.
30. The atomic number represents the number of protons in the nucleus of the atom and makes the atom a particular element. The mass number represents the total number of protons and neutrons in the nucleus of an atom and distinguishes one isotope of an element from another.

32. mass

34. Atoms of the same element (i.e., atoms with the same number of protons in the nucleus) may have different numbers of neutrons, and so will have different masses.

36.

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Symbol</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>14</td>
<td>Si</td>
<td>silicon</td>
</tr>
<tr>
<td>54</td>
<td>Xe</td>
<td>xenon</td>
</tr>
<tr>
<td>79</td>
<td>Au</td>
<td>gold</td>
</tr>
<tr>
<td>56</td>
<td>Ba</td>
<td>barium</td>
</tr>
<tr>
<td>53</td>
<td>I</td>
<td>iodine</td>
</tr>
<tr>
<td>50</td>
<td>Sn</td>
<td>tin</td>
</tr>
<tr>
<td>48</td>
<td>Cd</td>
<td>cadmium</td>
</tr>
</tbody>
</table>

38.

a. $^{26}_{14}$Si
b. $^{30}_{15}$P
c. $^{47}_{24}$Cr
d. $^{60}_{27}$Co
e. $^{62}_{30}$Zn
f. $^{39}_{19}$K

40.

a. 19 protons, 20 neutrons, 19 electrons
b. 24 protons, 29 neutrons, 24 electrons
c. 34 protons, 50 neutrons, 34 electrons
d. 33 protons, 43 neutrons, 33 electrons
e. 36 protons, 55 neutrons, 36 protons
f. 27 protons, 32 neutrons, 27 electrons

42.

<table>
<thead>
<tr>
<th>Element</th>
<th>Number of Neutrons</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>nitrogen</td>
<td>6</td>
<td>7</td>
<td>13</td>
<td>$^{13}_{7}$N</td>
</tr>
<tr>
<td>nitrogen</td>
<td>7</td>
<td>7</td>
<td>14</td>
<td>$^{14}_{7}$N</td>
</tr>
<tr>
<td>lead</td>
<td>124</td>
<td>82</td>
<td>206</td>
<td>$^{206}_{82}$Pb</td>
</tr>
<tr>
<td>iron</td>
<td>31</td>
<td>26</td>
<td>57</td>
<td>$^{57}_{26}$Fe</td>
</tr>
<tr>
<td>krypton</td>
<td>48</td>
<td>36</td>
<td>84</td>
<td>$^{84}_{36}$Kr</td>
</tr>
</tbody>
</table>
44. Elements with similar chemical properties are aligned vertically in families known as groups.

46. Metallic elements are found towards the left and bottom of the periodic table; there are far more metallic elements than there are nonmetals.

48. The gaseous nonmetallic elements are hydrogen, nitrogen, oxygen, fluorine, chlorine, plus all the Group 8 elements (noble gases). There are no gaseous metallic elements under room conditions.

50. The metalloids are the elements found on either side of the “stairstep” region that is marked on most periodic tables. The metalloid elements show some properties of both metals and nonmetals.

52. 
   a. fluorine, chlorine, bromine, iodine, astatine
   b. lithium, sodium, potassium, rubidium, cesium, francium
   c. beryllium, magnesium, calcium, strontium, barium, radium
   d. helium, neon, argon, krypton, xenon, radon

54. 

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Number</th>
<th>Group Number</th>
<th>Metal/Nonmetal</th>
</tr>
</thead>
<tbody>
<tr>
<td>calcium</td>
<td>Ca</td>
<td>20</td>
<td>2</td>
<td>metal</td>
</tr>
<tr>
<td>radon</td>
<td>Rn</td>
<td>86</td>
<td>8</td>
<td>nonmetal</td>
</tr>
<tr>
<td>rubidium</td>
<td>Rb</td>
<td>37</td>
<td>1</td>
<td>metal</td>
</tr>
<tr>
<td>phosphorus</td>
<td>P</td>
<td>15</td>
<td>5</td>
<td>nonmetal</td>
</tr>
<tr>
<td>germanium</td>
<td>Ge</td>
<td>32</td>
<td>4</td>
<td>metalloid</td>
</tr>
</tbody>
</table>

56. Most of the elements are too reactive to be found in the uncombined form in nature and are found only in compounds.

58. These elements are found uncombined in nature and do not readily react with other elements. For many years it was thought that these elements formed no compounds at all, but this has now been shown to be untrue.

60. diatomic gases: $\text{H}_2$, $\text{N}_2$, $\text{O}_2$, $\text{Cl}_2$, and $\text{F}_2$

   monatomic gases: $\text{He}$, $\text{Ne}$, $\text{Kr}$, $\text{Xe}$, $\text{Rn}$, and $\text{Ar}$

62. chlorine

64. diamond

66. electrons

68. $3^+$

70. -ide

72. nonmetallic

74.
   a. 10
   b. 22
   c. 10
   d. 10
   e. 23
Chapter 4: Chemical Foundations: Elements, Atoms, and Ions

76.

a. two electrons gained
b. three electrons gained
c. three electrons lost
d. two electrons lost
e. one electron lost
f. two electrons lost.

78.

a. P\(^{3-}\)
b. Ra\(^{2+}\)
c. At\(^-\)
d. no ion
e. Cs\(^+\)
f. Se\(^{2-}\)

80. Sodium chloride is an ionic compound consisting of Na\(^+\) and Cl\(^-\) ions. When NaCl is dissolved in water, these ions are set free and can move independently to conduct the electrical current. Sugar crystals, although they may visually appear similar, contain no ions. When sugar is dissolved in water, it dissolves as uncharged molecules. There are no electrically charged species present in a sugar solution to carry the electrical current.

82. The total number of positive charges must equal the total number of negative charges so that there will be no net charge on the crystals of an ionic compound. A macroscopic sample of compound must ordinarily not have any net charge.

84.

a. The smallest common multiple of two and three is six; two 3+ ions are balanced by three 2– ions; Cr\(_2\)S\(_3\)
b. One 2+ ion balances one 2– ion; CrO
c. Three 1– ions are needed to balance one 3+ ion; AlF\(_3\)
d. The smallest common multiple of two and three is six; two 3+ ions are balanced by three 2– ions; Al\(_2\)O\(_3\)
e. One 3+ ion balances one 3– ion; AlP
f. Three 1+ ions are needed to balance one 3– ion; Li\(_3\)N

86.

a. 7; halogens
b. 8; noble gases
c. 2; alkaline earth elements

d. 2; alkaline earth elements

e. 4

f. 6; (The members of Group 6 are sometimes called the chalcogens.)

g. 8; noble gases

h. 1; alkali metals

88.

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Atomic Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group 3</td>
<td></td>
<td></td>
</tr>
<tr>
<td>boron</td>
<td>B</td>
<td>5</td>
</tr>
<tr>
<td>aluminum</td>
<td>Al</td>
<td>13</td>
</tr>
<tr>
<td>gallium</td>
<td>Ga</td>
<td>31</td>
</tr>
<tr>
<td>indium</td>
<td>In</td>
<td>49</td>
</tr>
<tr>
<td>Group 5</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrogen</td>
<td>N</td>
<td>7</td>
</tr>
<tr>
<td>phosphorus</td>
<td>P</td>
<td>15</td>
</tr>
<tr>
<td>arsenic</td>
<td>As</td>
<td>33</td>
</tr>
<tr>
<td>antimony</td>
<td>Sb</td>
<td>51</td>
</tr>
<tr>
<td>Group 6</td>
<td></td>
<td></td>
</tr>
<tr>
<td>oxygen</td>
<td>O</td>
<td>8</td>
</tr>
<tr>
<td>sulfur</td>
<td>S</td>
<td>16</td>
</tr>
<tr>
<td>selenium</td>
<td>Se</td>
<td>34</td>
</tr>
<tr>
<td>tellurium</td>
<td>Te</td>
<td>52</td>
</tr>
<tr>
<td>Group 8</td>
<td></td>
<td></td>
</tr>
<tr>
<td>helium</td>
<td>He</td>
<td>2</td>
</tr>
<tr>
<td>neon</td>
<td>Ne</td>
<td>10</td>
</tr>
<tr>
<td>argon</td>
<td>Ar</td>
<td>18</td>
</tr>
<tr>
<td>krypton</td>
<td>Kr</td>
<td>36</td>
</tr>
</tbody>
</table>

90. Most of the mass of an atom is concentrated in the nucleus; the protons and neutrons that constitute the nucleus have similar masses, and these particles are nearly two thousand times heavier than electrons. The chemical properties of an atom depend on the number and location of the electrons it possesses. Electrons are found in the outer regions of the atom and are the particles most likely to be involved in interactions among atoms.

92. \(C_6H_{12}O_6\)

94.

a. 29 protons; 34 neutrons; 29 electrons

b. 35 protons; 45 neutrons; 35 electrons

c. 12 protons; 12 neutrons; 12 electrons

96. The chief use of gold in ancient times was as ornamentation, whether in statuary or in jewelry. Gold possesses an especially beautiful luster, and because it is relatively soft and malleable, it could be worked finely by artisans. Among the metals gold is particularly inert to attack by most substances in the environment.
98.
   a. I
   b. Si
   c. W
   d. Fe
   e. Cu
   f. Co

100.
   a. Br
   b. Bi
   c. Hg
   d. V
   e. F
   f. Ca

102.
   a. osmium
   b. zirconium
   c. rubidium
   d. radon
   e. uranium
   f. manganese
   g. nickel
   h. bromine

104.
   a. CO$_2$
   b. AlCl$_3$
   c. HClO$_4$
   d. SCl$_6$

106.
   a. $^{13}_{6}$C
   b. $^{13}_{6}$C
   c. $^{13}_{6}$C
   d. $^{44}_{19}$K
e. $^{41}_{20}\text{Ca}$

f. $^{35}_{19}\text{K}$

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Mass Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{41}_{20}\text{Ca}$</td>
<td>20</td>
<td>21</td>
<td>41</td>
</tr>
<tr>
<td>$^{55}_{25}\text{Mn}$</td>
<td>25</td>
<td>30</td>
<td>55</td>
</tr>
<tr>
<td>$^{109}_{47}\text{Ag}$</td>
<td>47</td>
<td>62</td>
<td>109</td>
</tr>
<tr>
<td>$^{43}_{21}\text{Sc}$</td>
<td>21</td>
<td>24</td>
<td>45</td>
</tr>
</tbody>
</table>
CHAPTER 5
Nomenclature

CHAPTER ANSWERS

2. A binary compound contains only two elements. The major types of binary compounds are ionic (compounds that contain a metal and a nonmetal) and nonionic (compounds containing two nonmetals).

4. cation

6. Some substances do not contain molecules. For example, the substance sodium chloride consists of an extended lattice array of sodium ions, Na⁺, and chloride ions, Cl⁻. Each sodium ion is surrounded by several chloride ions, and each chloride ion is surrounded by several sodium ions. We write the formula as NaCl to indicate the relative number of each ion in the substance, not to indicate that there are “molecules” of sodium chloride.

8. Roman numeral

10.
   a. potassium bromide
   b. zinc chloride
   c. cesium oxide
   d. magnesium sulfide
   e. aluminum iodide
   f. magnesium bromide
   g. beryllium fluoride
   h. barium hydride

12.
   a. Ag₂S
   b. BaH₂
   c. Al₂O₃
   d. MgF₂
   e. correct

14.
   a. copper(II) chloride
   b. copper(I) iodide
   c. manganese(II) bromide
   d. chromium(II) iodide
e. chromium(III) chloride
f. mercury(II) oxide

16.

a. cupric iodide
b. mercurous bromide
c. chromous bromide
d. cobaltous oxide
e. cobaltic oxide
f. stannous chloride

18.

a. xenon difluoride
b. diboron trisulfide
c. dichlorine hept(a)oxide
d. silicon tetrabromide
e. nitrogen monoxide
f. sulfur trioxide

20.

a. barium nitride
b. aluminum sulfide
c. diphosphorus trisulfide
d. calcium phosphide
e. krypton pentafluoride
f. copper(I) selenide/cuprous selenide

22.

a. barium fluoride – ionic
b. radium oxide – ionic
c. dinitrogen oxide – nonionic
d. rubidium oxide – ionic
e. diarsenic pent(a)oxide – nonionic
f. calcium nitride – ionic
24. An oxyanion is a polyatomic ion containing a given element and one or more oxygen atoms. The oxyanions of chlorine and bromine are given below:

<table>
<thead>
<tr>
<th>Oxyanion</th>
<th>Name</th>
<th>Oxyanion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO⁻</td>
<td>hypochlorite</td>
<td>BrO⁻</td>
<td>hypobromite</td>
</tr>
<tr>
<td>ClO₂⁻</td>
<td>chlorite</td>
<td>BrO₂⁻</td>
<td>bromite</td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>chlorate</td>
<td>BrO₃⁻</td>
<td>bromate</td>
</tr>
<tr>
<td>ClO₄⁻</td>
<td>perchlorate</td>
<td>BrO₄⁻</td>
<td>perbromate</td>
</tr>
</tbody>
</table>

26. For a series of oxyanions, the prefix hypo– is used for the anion with the fewest oxygen atoms, and the prefix per– is used for the anion with the most oxygen atoms.

28. IO⁻     hypoiodite
   IO₂⁻    iodite
   IO₃⁻    iodate
   IO₄⁻    periodate

30. 
   a. NO₃⁻
   b. NO₂⁻
   c. NH₄⁺
   d. CN⁻

32. CN⁻     cyanide
   CO₃²⁻    carbonate
   HCO₃⁻    hydrogen carbonate
   C₂H₃O₂⁻  acetate

34. 
   a. ammonium
   b. dihydrogen phosphate
   c. sulfate
   d. hydrogen sulfite (also called bisulfite)
   e. perchlorate
   f. iodate

36. 
   a. ammonium acetate
   b. lithium perchlorate
   c. sodium hydrogen sulfate
   d. gold(III) carbonate
   e. calcium chlorate
f. hydrogen peroxide

38. oxygen (commonly referred to as \textit{oxy}acids)

40.

a. hypochlorous acid
b. sulfurous acid
c. bromic acid
d. hypoiodous acid
e. perbromic acid
f. hydrosulfuric acid
g. hydroselenic acid
h. phosphorous acid

42.

a. lithium nitrate
b. chromium(III) carbonate/chromic carbonate
c. copper(II) carbonate/cupric carbonate
d. copper(I) selenide/cuprous selenide
e. manganese(IV) sulfate
f. magnesium nitrite

44.

a. \( \text{N}_2\text{O} \)
b. \( \text{NO}_2 \)
c. \( \text{N}_2\text{O}_4 \)
d. \( \text{SF}_6 \)
e. \( \text{PBr}_3 \)
f. \( \text{Cl}_4 \)
g. \( \text{OCl}_2 \)

46.

a. \( \text{BaSO}_3 \)
b. \( \text{Ca(H}_2\text{PO}_4)_2 \)
c. \( \text{NH}_4\text{ClO}_4 \)
d. \( \text{NaMnO}_4 \)
e. \( \text{Fe}_2(\text{SO}_4)_3 \)
f. \( \text{CoCO}_3 \)
g. \( \text{Ni(OH)}_2 \)
h. \( \text{ZnCrO}_4 \)
48. 
   a. HCN 
   b. HNO₃ 
   c. H₂SO₄ 
   d. H₃PO₄ 
   e. HClO or HOCl 
   f. HBr 
   g. HBrO₂ 
   h. HF 

50. 
   a. Mg(HSO₄)₂ 
   b. CsClO₄ 
   c. FeO 
   d. H₂Te 
   e. Sr(NO₃)₂ 
   f. Sn(C₂H₃O₂)₄ 
   g. MnSO₄ 
   h. N₂O₄ 
   i. Na₂HPO₄ 
   j. Li₂O₂ 
   k. HNO₂ 
   l. Co(NO₃)₃ 

52. A moist paste of NaCl would contain Na⁺ and Cl⁻ ions in solution and would serve as a conductor of electrical impulses. 

54. H → H⁺ (hydrogen ion: a cation) + e⁻ 
   H + e⁻ → H⁻ (hydrıde ion: an anion) 

56. missing oxyanions: IO₃⁻; ClO₂⁻ 
   missing oxyacids: HClO₄; HClO; HBrO₂ 

58. 
   a. gold(III) bromide, auric bromide 
   b. cobalt(III) cyanide, cobaltic cyanide 
   c. magnesium hydrogen phosphate 
   d. diboron hexahydride (diborane is its common name) 
   e. ammonia 
   f. silver(I) sulfate (usually called silver sulfate)
60. 
   a. ammonium carbonate 
   b. ammonium hydrogen carbonate, ammonium bicarbonate 
   c. calcium phosphate 
   d. sulfurous acid 
   e. manganese(IV) oxide 
   f. iodic acid 
   g. potassium hydride 

62. 
   a. $\text{M(C}_2\text{H}_3\text{O}_2\text{)}_4$ 
   b. $\text{M(MnO}_4\text{)}_4$ 
   c. $\text{MO}_2$ 
   d. $\text{M(HPO}_4\text{)}_2$ 
   e. $\text{M(OH}_4$ 
   f. $\text{M(NO}_2\text{)}_4$ 

64. $\text{M}^+$ compounds: $\text{MD, M}_2\text{E, M}_3\text{F}$ 
   $\text{M}^{2+}$ compounds: $\text{MD}_2, \text{ME}, \text{M}_3\text{F}_2$ 
   $\text{M}^{3+}$ compounds: $\text{MD}_3, \text{M}_2\text{E}_3, \text{MF}$ 

66.  

<table>
<thead>
<tr>
<th></th>
<th>Ca(NO$_3$)$_2$</th>
<th>CaSO$_4$</th>
<th>Ca(HSO$_4$)$_2$</th>
<th>Ca(H$_2$PO$_4$)$_2$</th>
<th>CaO</th>
<th>CaCl$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sr(NO$_3$)$_2$</td>
<td>SrSO$_4$</td>
<td>Sr(HSO$_4$)$_2$</td>
<td>Sr(H$_2$PO$_4$)$_2$</td>
<td>SrO</td>
<td>SrCl$_2$</td>
<td></td>
</tr>
<tr>
<td>NH$_4$NO$_3$</td>
<td>(NH$_4$)$_2$SO$_4$</td>
<td>NH$_4$HSO$_4$</td>
<td>NH$_4$H$_2$PO$_4$</td>
<td>(NH$_4$)$_2$O</td>
<td>NH$_4$Cl</td>
<td></td>
</tr>
<tr>
<td>Al(NO$_3$)$_3$</td>
<td>Al$_2$(SO$_4$)$_3$</td>
<td>Al(HSO$_4$)$_2$</td>
<td>Al(H$_2$PO$_4$)$_2$</td>
<td>Al$_2$O$_3$</td>
<td>AlCl$_3$</td>
<td></td>
</tr>
<tr>
<td>Fe(NO$_3$)$_3$</td>
<td>Fe$_2$(SO$_4$)$_3$</td>
<td>Fe(HSO$_4$)$_2$</td>
<td>Fe(H$_2$PO$_4$)$_2$</td>
<td>Fe$_2$O$_3$</td>
<td>FeCl$_3$</td>
<td></td>
</tr>
<tr>
<td>Ni(NO$_3$)$_2$</td>
<td>NiSO$_4$</td>
<td>Ni(HSO$_4$)$_2$</td>
<td>Ni(H$_2$PO$_4$)$_2$</td>
<td>NiO</td>
<td>NiCl$_2$</td>
<td></td>
</tr>
<tr>
<td>AgNO$_3$</td>
<td>Ag$_2$SO$_4$</td>
<td>AgHSO$_4$</td>
<td>AgH$_2$PO$_4$</td>
<td>Ag$_2$O</td>
<td>AgCl</td>
<td></td>
</tr>
<tr>
<td>Au(NO$_3$)$_3$</td>
<td>Au$_2$(SO$_4$)$_3$</td>
<td>Au(HSO$_4$)$_2$</td>
<td>Au(H$_2$PO$_4$)$_2$</td>
<td>Au$_2$O$_3$</td>
<td>AuCl$_3$</td>
<td></td>
</tr>
<tr>
<td>KNO$_3$</td>
<td>K$_2$SO$_4$</td>
<td>KHSO$_4$</td>
<td>KH$_2$PO$_4$</td>
<td>K$_2$O</td>
<td>KCl</td>
<td></td>
</tr>
<tr>
<td>Hg(NO$_3$)$_2$</td>
<td>HgSO$_4$</td>
<td>Hg(HSO$_4$)$_2$</td>
<td>Hg(H$_2$PO$_4$)$_2$</td>
<td>HgO</td>
<td>HgCl$_2$</td>
<td></td>
</tr>
<tr>
<td>Ba(NO$_3$)$_2$</td>
<td>BaSO$_4$</td>
<td>Ba(HSO$_4$)$_2$</td>
<td>Ba(H$_2$PO$_4$)$_2$</td>
<td>BaO</td>
<td>BaCl$_2$</td>
<td></td>
</tr>
</tbody>
</table>

68. helium 

70. iodine (solid), bromine (liquid), fluorine and chlorine (gases) 

72. 1– 

74. 1–
76.  
   a. \( \text{Al}(13e^-) \rightarrow \text{Al}^{3+}(10e^-) + 3e^- \)
   b. \( \text{S}(16e^-) + 2e^- \rightarrow \text{S}^2-(18e^-) \)
   c. \( \text{Cu}(29e^-) \rightarrow \text{Cu}^+(28e^-) + e^- \)
   d. \( \text{F}(9e^-) + e^- \rightarrow \text{F}^-(10e^-) \)
   e. \( \text{Zn}(30e^-) \rightarrow \text{Zn}^{2+}(28e^-) + 2e^- \)
   f. \( \text{P}(15e^-) + 3e^- \rightarrow \text{P}^3-(18e^-) \)

78.  
   a. Two 1+ ions are needed to balance a 2– ion, so the formula must have two \( \text{Na}^+ \) ions for each \( \text{S}^2- \) ion; \( \text{Na}_2\text{S} \).
   b. One 1+ ion exactly balances a 1– ion, so the formula should have an equal number of \( \text{K}^+ \) and \( \text{Cl}^- \) ions; \( \text{KCl} \).
   c. One 2+ ion exactly balances a 2– ion, so the formula must have an equal number of \( \text{Ba}^{2+} \) and \( \text{O}^{2-} \) ions; \( \text{BaO} \).
   d. One 2+ ion exactly balances a 2– ion, so the formula must have an equal number of \( \text{Mg}^{2+} \) and \( \text{Se}^{2-} \) ions; \( \text{MgSe} \).
   e. One 2+ ion requires two 1– ions to balance charge, so the formula must have twice as many \( \text{Br}^- \) ions as \( \text{Cu}^{3+} \) ions; \( \text{CuBr}_2 \).
   f. One 3+ ion requires three 1– ions to balance charge, so the formula must have three times as many \( \text{I}^- \) ions as \( \text{Al}^{3+} \) ions; \( \text{Al}_3\text{I}_3 \).
   g. Two 3+ ions give a total of 6+, whereas three 2– ions will give a total of 6–. The formula then should contain two \( \text{Al}^{3+} \) ions and three \( \text{O}^{2-} \) ions; \( \text{Al}_2\text{O}_3 \).
   h. Three 2+ ions are required to balance two 3– ions, so the formula must contain three \( \text{Ca}^{2+} \) ions for every two \( \text{N}^{3-} \) ions; \( \text{Ca}_3\text{N}_2 \).

80.  
   a. silver(I) oxide or just silver oxide
   b. correct
   c. iron(III) oxide
   d. plumbic oxide
   e. correct

82.  
   a. As bromide ions always have a 1– charge, the cobaltion must have a 3+ charge; the name is cobaltic bromide.
   b. As iodide ions always have a 1– charge, the lead ion must have a 4+ charge; the name is plumbic iodide.
   c. As oxide ions always have a 2– charge, and as there are three oxide ions, each iron ion must have a 3+ charge; the name is ferric oxide.
d. As sulfide ions always have a 2– charge, the iron ion must have a 2+ charge; the name is ferrous sulfide.

e. As chloride ions always have a 1– charge, the tin ion must have a 4+ charge; the name is stannic chloride.

f. As oxide ions always have a 2– charge, the tin ion must have a 2+ charge; the name is stannous oxide.

84.

a. iron(III) acetate, ferric acetate
b. bromine monofluoride
c. potassium peroxide
d. silicon tetrabromide
e. copper(II) permanganate, cupric permanganate
f. calcium chromate

86.

a. CO\textsubscript{3}\textsuperscript{2–}
b. HCO\textsubscript{3}–
c. C\textsubscript{2}H\textsubscript{3}O\textsubscript{2}–
d. CN–

88.

a. carbonate
b. chlorate
c. sulfate
d. phosphate
e. perchlorate
f. permanganate

90. Answer depends on student choices.

92.

a. NaH\textsubscript{2}PO\textsubscript{4}
b. LiClO\textsubscript{4}
c. Cu(HCO\textsubscript{3})\textsubscript{2}
d. KС\textsubscript{2}H\textsubscript{3}O\textsubscript{2}
e. BaO\textsubscript{2}
f. Cs\textsubscript{2}SO\textsubscript{3}
Cumulative Review Chapters 4 & 5

ANSWERS

2. How many elements could you name? Although you certainly don’t have to memorize all the elements, you should be able to give at least the symbol or name for the most common elements (listed in Table 4.3).

4. Dalton’s atomic theory as presented in this text consists of five main postulates. Although Dalton’s theory was exceptional scientific thinking for its time, some of the postulates have been modified as our scientific instruments and calculation methods have become increasingly more sophisticated. The main postulates of Dalton’s theory are as follows: (1) Elements are made up of tiny particles called atoms; (2) all atoms of a given element are identical; (3) although all atoms of a given element are identical, these atoms are different from the atoms of all other elements; (4) atoms of one element can combine with atoms of another element to form a compound, and such a compound will always have the same relative numbers and types of atoms for its composition; and (5) atoms are merely rearranged into new groupings during an ordinary chemical reaction, and no atom is ever destroyed and no new atom is ever created during such a reaction.

6. The expression nuclear atom indicates that we view the atom as having a dense center of positive charge (called the nucleus) around which the electrons move through primarily empty space. Rutherford’s experiment involved shooting a beam of particles at a thin sheet of metal foil. According to the then-current “plum pudding” model of the atom most of these positively-charged particles should have passed right through the foil. However, Rutherford detected that a significant number of particles effectively bounced off something and were deflected backwards to the source of particles and that other particles were deflected from the foil at large angles. Rutherford realized that his observations could be explained if the atoms of the metal foil had a small, dense, positively-charged nucleus with a significant amount of empty space between nuclei. The empty space between nuclei would allow most of the particles to pass through the atom. However, if a particle hit a nucleus head-on, it would be deflected backwards at the source. If a positively charged particle passed near a positively charged nucleus (but did not hit the nucleus head-on), then the particle would be deflected by the repulsive forces between the positive charges. Rutherford’s experiment conclusively disproved the “plum pudding” model for the atom, which envisioned the atom as a uniform sphere of positive charge with enough negatively charged electrons scattered throughout the atom to balance out the positive charge.

8. Isotopes represent atoms of the same element that have different atomic masses. Isotopes are a result of the fact that atoms of a given element may have different numbers of neutrons in their nuclei. Isotopes have the same atomic number (number of protons in the nucleus) but have different mass numbers (total number of protons and neutrons in the nucleus). The different isotopes of an atom are indicated by symbolism of the form \(^\text{A}^\text{Z}\) in which \(Z\) represents the atomic number and \(A\) the mass number of element \(X\). For example, \(^{13}\text{C}\) represents a nuclide of carbon with atomic number six (six protons in the nucleus) and mass number 13 (reflecting six protons plus seven neutrons in the nucleus). The various isotopes of an element have identical chemical properties because the chemical properties of an atom are a function of the electrons in the atom.
(not the nucleus). The physical properties of the isotopes of an element (and compounds containing those isotopes) may differ because of the difference in mass of the isotopes.

10. Most elements are too reactive to be found in nature in other than the combined form. Aside from the noble metals gold, silver, and platinum, the only other elements commonly found in nature in the uncombined state are some of the gaseous elements (such as O₂, N₂, He, Ar, etc.) and the solid nonmetals carbon and sulfur.

12. Ionic compounds typically are hard, crystalline solids with high melting and boiling points. Ionic substances like sodium chloride, when dissolved in water or when melted, conduct electrical currents. Chemists have taken this evidence to mean that ionic substances consist of positively and negatively charged particles (ions). Although an ionic substance is made up of positively and negatively charged particles, there is no net electrical charge on a sample of such a substance because the total number of positive charges is balanced by an equal number of negative charges. An ionic compound of just cations or just anions could not possibly exist; there must be a balance of charge or the compound will be very unstable (like charges repel each other).

14. When naming ionic compounds, we name the positive ion (cation) first. For simple binary Type I ionic compounds, the ending –ide is added to the root name of the element that is the negative ion (anion). For example, for the Type I ionic compound formed between potassium and sulfur, K₂S, the name would be potassium sulfide; potassium is the cation, sulfur is the anion (with the suffix –ide added). Type II compounds are named by either of two systems, the “ous–ic” system (which is falling out of use) and the “Roman numeral” system, which is preferred by most chemists. Type II compounds involve elements that form more than one stable ion. It is therefore necessary to specify which ion is present in a given compound. For example, iron forms two types of stable ion; Fe²⁺ and Fe³⁺. Iron can react with oxygen to form either of two stable oxides, FeO or Fe₂O₃, depending on which cation is involved. Under the Roman numeral naming system, FeO would be named iron(II) oxide to show that it contains Fe²⁺ ions and Fe₂O₃ would be named iron(III) oxide to indicate that it contains Fe³⁺ ions. The Roman numeral used in a name corresponds to the charge of the specific ion present in the compound. Under the less-favored “ous–ic” system, for an element that forms two stable ions, the ending –ous is used to indicate the lower-charged ion whereas the ending –ic is used to indicate the higher-charged ion. FeO and Fe₂O₃ would thus be named ferrous oxide and ferric oxide, respectively. The “ous–ic” system has fallen out of favor because it does not indicate the actual charge on the ion, but only that it is the lower- or higher-charged of the two. This can lead to confusion: For example Fe²⁺ is called ferrous ion in this system, but Cu²⁺ is called cupric ion (since there is also a Cu⁺ stable ion).

16. A polyatomic ion is an ion containing more than one atom. Some common polyatomic ions you should be familiar with are listed in Table 5.4. Parentheses are used in writing formulas containing polyatomic ions to indicate unambiguously how many of the polyatomic ions are present in the formula to make certain there is no mistake as to what is meant by the formula. For example, consider the substance calcium phosphate. The correct formula for this substance is Ca₃(PO₄)₂, which indicates that three calcium ions are combined for every two phosphate ions (check the total number of positive and negative charges to see why this is so). If we did not write the parenthesis around the formula for the phosphate ion; that is, if we had written Ca₃PO₄, people reading this formula might think that there were 42 oxygen atoms present!

18. Acids, in general, are substances that produce protons (H⁺ ions) when dissolved in water. For acids that do not contain oxygen, the prefix hydro– and the suffix –ic are used with the root name of the element present in the acid (for example, HCl, hydrochloric acid; H₂S, hydrosulfuric acid; HF, hydrofluoric acid). The nomenclature of acids whose anions contain oxygen is more complicated. A series of prefixes and suffixes is used with the name of the non-oxygen atom in the anion of the acid. These prefixes and suffixes indicate the relative (not actual) number of
oxygen atoms present in the anion. Most of the elements that form oxyanions form two such anions; for example, sulfur forms sulfite ion \((\text{SO}_3^{2–})\) and sulfate ion \((\text{SO}_4^{2–})\), and nitrogen forms nitrite ion \((\text{NO}_2^{–})\) and nitrate ion \((\text{NO}_3^{–})\). For an element that forms two oxyanions, the acid containing the anions will have the ending \(-ous\) if the anion is the \(-ite\) anion and the ending \(-ic\) if the anion is the \(-ate\) anion. For example, \(\text{HNO}_2\) is nitrous acid, and \(\text{HNO}_3\) is nitric acid; \(\text{H}_2\text{SO}_3\) is sulfurous acid, and \(\text{H}_2\text{SO}_4\) is sulfuric acid. The halogen elements (Group 7) each form four oxyanions and consequently four oxyacids. The prefix \(\text{hypo–}\) is used for the oxyacid that contains fewer oxygen atoms than the \(-ite\) anion, and the prefix \(\text{per–}\) is used for the oxyacid that contains more oxygen atoms than the \(-ate\) anion. For example,

<table>
<thead>
<tr>
<th>Acid</th>
<th>Name</th>
<th>Anion</th>
<th>Anion Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\text{HBrO})</td>
<td>hypobromous acid</td>
<td>(\text{BrO}^–)</td>
<td>hypobromite</td>
</tr>
<tr>
<td>(\text{HBrO}_2)</td>
<td>bromous acid</td>
<td>(\text{BrO}_2^{–})</td>
<td>bromite</td>
</tr>
<tr>
<td>(\text{HBrO}_3)</td>
<td>bromic acid</td>
<td>(\text{BrO}_3^{–})</td>
<td>bromate</td>
</tr>
<tr>
<td>(\text{HBrO}_4)</td>
<td>perbromic acid</td>
<td>(\text{BrO}_4^{–})</td>
<td>perbromate</td>
</tr>
</tbody>
</table>

20. How many elements in each family could you name? Elements in the same family have the same type of electronic configuration and tend to undergo similar chemical reactions with other groups. For example, \(\text{Li, Na, K, Rb, and Cs}\) all react with elemental chlorine gas, \(\text{Cl}_2\), to form an ionic compound of general formula \(\text{M}^+\text{Cl}^–\).

22.

a. 1 proton, 1 neutron, 1 electron
b. 1 proton, no neutrons, 1 electron
c. 1 proton, 2 neutrons, 1 electron
d. 31 protons, 40 neutrons, 31 electrons
e. 82 protons, 125 neutrons, 82 electrons
f. 82 protons, 130 neutrons, 82 electrons
g. 28 protons, 31 neutrons, 28 electrons
h. 12 protons, 13 neutrons, 12 electrons

24.

a. 12 protons, 10 electrons
b. 26 protons, 24 electrons
c. 26 protons, 23 electrons
d. 9 protons, 10 electrons
e. 28 protons, 26 electrons
f. 30 protons, 28 electrons
g. 27 protons, 24 electrons
h. 7 protons, 10 electrons
i. 16 protons, 18 electrons
j. 37 protons, 36 electrons
k. 34 protons, 36 electrons
l. 19 protons, 18 electrons

26.
   a. copper(II) chloride, cupric chloride
   b. cobalt(III) chloride, cobaltic chloride
   c. iron(II) oxide, ferrous oxide
   d. manganese(IV) sulfide
   e. manganese(II) sulfide, manganous sulfide
   f. copper(I) oxide, cuprous oxide
   g. tin(IV) chloride, stannic chloride
   h. magnesium bromide
   i. hydrogen peroxide

28.
   a. \( \text{NH}_4^+ \), ammonium ion
   b. \( \text{SO}_3^{2-} \), sulfite ion
   c. \( \text{NO}_3^- \), nitrate ion
   d. \( \text{SO}_4^{2-} \), sulfate ion
   e. \( \text{NO}_2^- \), nitrite ion
   f. \( \text{CN}^- \), cyanide ion
   g. \( \text{OH}^- \), hydroxide ion
   h. \( \text{ClO}_4^- \), perchlorate ion
   i. \( \text{ClO}^- \), hypochlorite ion
   j. \( \text{PO}_4^{3-} \), phosphate ion

30.
   a. \( \text{B}_2\text{O}_3 \), diboron trioxide
   b. \( \text{NO}_2 \), nitrogen dioxide
   c. \( \text{PCl}_5 \), phosphorus pentachloride
   d. \( \text{N}_2\text{O}_4 \), dinitrogen tetroxide
   e. \( \text{P}_2\text{O}_5 \), diphosphorus pentoxide
   f. \( \text{ICl} \), iodine monochloride
   g. \( \text{SF}_6 \), sulfur hexafluoride
   h. \( \text{N}_2\text{O}_3 \), dinitrogen trioxide
CHAPTER 6

Chemical Reactions: An Introduction

CHAPTER ANSWERS

2. Most of these products contain a peroxide, which decomposes releasing oxygen gas.

4. Bubbling takes place as the hydrogen peroxide chemically decomposes into water and oxygen gas.

6. The two components are both liquids, but harden to a solid when combined. There is also heat evolved during the reaction.

8. atoms

10. the same

12. water

14. \( \text{H}_2\text{O}_2(aq) \rightarrow \text{H}_2(g) + \text{O}_2(g) \)

16. \( \text{AgNO}_3(aq) + \text{HCl}(aq) \rightarrow \text{AgCl}(s) + \text{HNO}_3(aq) \)

\( \text{Pb(NO}_3)_2(aq) + \text{HCl}(aq) \rightarrow \text{PbCl}_2(s) + \text{HNO}_3(aq) \)

18. \( \text{C}_3\text{H}_8(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \)

\( \text{C}_3\text{H}_3(g) + \text{O}_2(g) \rightarrow \text{CO}(g) + \text{H}_2\text{O}(g) \)

20. \( \text{CaCO}_3(s) + \text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \)

22. \( \text{SiO}_2(s) + \text{C}(s) \rightarrow \text{Si}(s) + \text{CO}(g) \)

24. \( \text{H}_2\text{S}(g) + \text{O}_2(g) \rightarrow \text{SO}_2(g) + \text{H}_2\text{O}(g) \)

26. \( \text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3(aq) \)

\( \text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) \)

28. \( \text{NO}(g) + \text{O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g) \)

30. \( \text{NH}_3(g) + \text{HNO}_3(aq) \rightarrow \text{NH}_4\text{NO}_3(s) \)

32. \( \text{Xe}(g) + \text{F}_2(g) \rightarrow \text{XeF}_4(s) \)

34. \( \text{NH}_4\text{Cl}(s) + \text{NaOH}(s) \xrightarrow{\text{heat}} \text{NH}_3(g) + \text{H}_2\text{O}(g) + \text{NaCl}(s) \)

36. whole numbers
38.

a. \( \text{Al}(s) + \text{CuO}(s) \rightarrow \text{Al}_2\text{O}_3(s) + \text{Cu}(l) \)

balance Al: \(2\text{Al}(s) + \text{CuO}(s) \rightarrow \text{Al}_2\text{O}_3(s) + \text{Cu}(l)\)
balance O: \(2\text{Al}(s) + 3\text{CuO}(s) \rightarrow \text{Al}_2\text{O}_3(s) + \text{Cu}(l)\)
balance Cu: \(2\text{Al}(s) + 3\text{CuO}(s) \rightarrow \text{Al}_2\text{O}_3(s) + 3\text{Cu}(l)\)
balanced equation: \(2\text{Al}(s) + 3\text{CuO}(s) \rightarrow \text{Al}_2\text{O}_3(s) + 3\text{Cu}(l)\)

b. \( \text{S}_8(s) + \text{F}_2(g) \rightarrow \text{SF}_6(g) \)

balance sulfur: \(\text{S}_8(s) + \text{F}_2(g) \rightarrow 8\text{SF}_6(g)\)
balance fluorine: \(\text{S}_8(s) + 24\text{F}_2(g) \rightarrow 8\text{SF}_6(g)\)
balanced equation: \(\text{S}_8(s) + 24\text{F}_2(g) \rightarrow 8\text{SF}_6(g)\)

c. \( \text{Xe}(g) + \text{F}_2(g) \rightarrow \text{XeF}_6(s) \)

balance fluorine: \(\text{Xe}(g) + 3\text{F}_2(g) \rightarrow \text{XeF}_6(s)\)
balanced equation: \(\text{Xe}(g) + 3\text{F}_2(g) \rightarrow \text{XeF}_6(s)\)

d. \( \text{NH}_4\text{Cl}(g) + \text{KOH}(s) \rightarrow \text{NH}_3(g) + \text{H}_2\text{O}(g) + \text{KCl}(s) \)

The equation is already balanced.

e. \( \text{SiC}(s) + \text{Cl}_2(g) \rightarrow \text{SiCl}_4(l) + \text{C}(s) \)

balance chlorine: \(\text{SiC}(s) + 2\text{Cl}_2(g) \rightarrow \text{SiCl}_4(l) + \text{C}(s)\)
balanced equation: \(\text{SiC}(s) + 2\text{Cl}_2(g) \rightarrow \text{SiCl}_4(l) + \text{C}(s)\)

f. \( \text{K}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow \text{KOH}(aq) \)

balance potassium: \(\text{K}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow 2\text{KOH}(aq)\)
balanced equation: \(\text{K}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow 2\text{KOH}(aq)\)

g. \( \text{N}_2\text{O}_5(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(aq) \)

balance nitrogen: \(\text{N}_2\text{O}_5(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq)\)
balanced equation: \(\text{N}_2\text{O}_5(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq)\)

h. \( \text{H}_2\text{S}(g) + \text{Cl}_2(g) \rightarrow \text{S}_8(s) + \text{HCl}(g) \)

balance sulfur: \(8\text{H}_2\text{S}(g) + \text{Cl}_2(g) \rightarrow \text{S}_8(s) + \text{HCl}(g)\)
balance hydrogen: \(8\text{H}_2\text{S}(g) + 8\text{Cl}_2(g) \rightarrow \text{S}_8(s) + 16\text{HCl}(g)\)
balanced equation: \(8\text{H}_2\text{S}(g) + 8\text{Cl}_2(g) \rightarrow \text{S}_8(s) + 16\text{HCl}(g)\)

40.

a. \( \text{Na}_2\text{SO}_4(aq) + \text{CaCl}_2(aq) \rightarrow \text{CaSO}_4(s) + 2\text{NaCl}(aq) \)

b. \( 3\text{Fe}(s) + 4\text{H}_2\text{O}(g) \rightarrow \text{Fe}_3\text{O}_4(s) + 4\text{H}_2(g) \)

c. \( \text{Ca(OH)}_2(aq) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + 2\text{H}_2\text{O}(l) \)

d. \( \text{Br}_2(g) + 2\text{H}_2\text{O}(l) + \text{SO}_2(g) \rightarrow 2\text{HBr}(aq) + \text{H}_2\text{SO}_4(aq) \)
42.

a. \(4\text{NaCl}(s) + 2\text{SO}_2(g) + 2\text{H}_2\text{O}(g) + \text{O}_2(g) \rightarrow 2\text{Na}_2\text{SO}_4(s) + 4\text{HCl}(g)\)

b. \(3\text{Br}_2(l) + \text{I}_2(s) \rightarrow 2\text{IBr}_2(s)\)

c. \(\text{Ca}(s) + 2\text{H}_2\text{O}(g) \rightarrow \text{Ca(OH)}_2(aq) + \text{H}_2(g)\)

d. \(2\text{BF}_3(g) + 3\text{H}_2\text{O}(g) \rightarrow \text{B}_2\text{O}_3(s) + 6\text{HF}(g)\)

e. \(\text{SO}_2(g) + 2\text{Cl}_2(g) \rightarrow \text{SOCl}_2(l) + \text{Cl}_2\text{O}(g)\)

f. \(\text{Li}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow 2\text{LiOH}(aq)\)

g. \(\text{Mg}(s) + \text{CuO}(s) \rightarrow \text{MgO}(s) + \text{Cu}(l)\)

h. \(\text{Fe}_3\text{O}_4(s) + 4\text{H}_2(g) \rightarrow 3\text{Fe}(l) + 4\text{H}_2\text{O}(g)\)

44.

a. \(\text{Ba(NO}_3)_2(aq) + \text{Na}_2\text{CrO}_4(aq) \rightarrow \text{BaCrO}_4(s) + 2\text{NaNO}_3(aq)\)

b. \(\text{PbCl}_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{PbSO}_4(s) + 2\text{KCl}(aq)\)

c. \(\text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l)\)

d. \(\text{SrC}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Sr(OH)}_2(s) + \text{C}_2\text{H}_2(g)\)

e. \(\text{Ba}_2\text{O}_5(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + \text{H}_2\text{O}_2(aq)\)

g. \(2\text{AsI}_3(s) \rightarrow 2\text{As}(s) + 3\text{I}_2(s)\)

h. \(2\text{CuSO}_4(aq) + 4\text{KI}(s) \rightarrow 2\text{CuI}(s) + \text{I}_2(s) + 2\text{K}_2\text{SO}_4(aq)\)

46. \(\text{Na}(s) + \text{O}_2(g) \rightarrow \text{Na}_2\text{O}_2(s)\)

\(\text{Na}_2\text{O}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{NaOH}(aq) + \text{O}_2(g)\)

48. \(\text{C}_12\text{H}_22\text{O}_{11}(aq) + \text{H}_2\text{O}(l) \rightarrow 4\text{C}_2\text{H}_5\text{OH}(aq) + 4\text{CO}_2(g)\)

50. \(2\text{Al}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 4\text{Al}(s) + 3\text{CO}_2(g)\)

52. \(2\text{Li}(s) + \text{S}(s) \rightarrow \text{Li}_2\text{S}(s)\)

\(2\text{Na}(s) + \text{S}(s) \rightarrow \text{Na}_2\text{S}(s)\)

\(2\text{K}(s) + \text{S}(s) \rightarrow \text{K}_2\text{S}(s)\)

\(2\text{Rb}(s) + \text{S}(s) \rightarrow \text{Rb}_2\text{S}(s)\)

\(2\text{Cs}(s) + \text{S}(s) \rightarrow \text{Cs}_2\text{S}(s)\)

\(2\text{Fr}(s) + \text{S}(s) \rightarrow \text{Fr}_2\text{S}(s)\)

54. \(\text{BaO}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{BaO}(s) + \text{H}_2\text{O}_2(aq)\)

56. \(2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)\)
58. \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \)

60. The senses we call “odor” and “taste” are really chemical reactions of the receptors in our body with molecules in the food we are eating. The fact that the receptors no longer detect the “fishy” odor or taste suggests that adding the lemon juice or vinegar has changed the nature of the amines in the fish.

62. \( \text{Fe}(s) + \text{S}(s) \rightarrow \text{FeS}(s) \)

64. \( \text{K}_2\text{CrO}_4(aq) + \text{BaCl}_2(aq) \rightarrow \text{BaCrO}_4(s) + 2\text{KCl}(aq) \)

66. \( 2\text{NaCl}(aq) + \text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g) + \text{Cl}_2(g) \)

70. \( \text{CuO}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O}(l) \)

72. \( \text{Na}_2\text{SO}_3(aq) + \text{S}(s) \rightarrow \text{Na}_2\text{S}_2\text{O}_3(aq) \)

74.

a. \( \text{Cl}_2(g) + 2\text{KI}(aq) \rightarrow 2\text{KCl}(aq) + \text{I}_2(s) \)

b. \( \text{CaC}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(s) + \text{C}_2\text{H}_2(g) \)

c. \( 2\text{NaCl}(s) + \text{H}_2\text{SO}_4(l) \rightarrow \text{Na}_2\text{SO}_4(s) + 2\text{HCl}(g) \)

d. \( \text{CaF}_2(s) + \text{H}_2\text{SO}_4(l) \rightarrow \text{CaSO}_4(s) + 2\text{HF}(g) \)

e. \( \text{K}_2\text{CO}_3(s) \rightarrow \text{K}_2\text{O}(s) + \text{CO}_2(g) \)

f. \( 3\text{BaO}(s) + 2\text{Al}(s) \rightarrow \text{Al}_2\text{O}_3(s) + 3\text{Ba} \)

g. \( 2\text{Al}(s) + 3\text{F}_2(g) \rightarrow 2\text{AlF}_3(s) \)

h. \( \text{CS}_2(g) + 3\text{Cl}_2(g) \rightarrow \text{CCl}_4(l) + \text{S}_2\text{Cl}_2(g) \)

76.

a. \( \text{Pb(NO}_3)_2(aq) + \text{K}_2\text{CrO}_4(aq) \rightarrow \text{PbCrO}_4(s) + 2\text{KNO}_3(aq) \)

b. \( \text{BaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{NaCl}(aq) \)

c. \( 2\text{CH}_3\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)

d. \( \text{Na}_2\text{CO}_3(aq) + \text{S}(s) + \text{SO}_2(g) \rightarrow \text{CO}_2(g) + \text{Na}_2\text{S}_2\text{O}_3(aq) \)

e. \( \text{Cu}(s) + 2\text{H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + \text{SO}_2(g) + 2\text{H}_2\text{O}(l) \)

f. \( \text{MnO}_2(s) + 4\text{HCl}(aq) \rightarrow \text{MnCl}_2(aq) + \text{Cl}_2(g) + 2\text{H}_2\text{O}(l) \)

g. \( \text{As}_2\text{O}_3(s) + 6\text{KI}(aq) + 6\text{HCl}(aq) \rightarrow 2\text{AsI}_3(s) + 6\text{KCl}(aq) + 3\text{H}_2\text{O}(l) \)

h. \( 2\text{Na}_2\text{S}_2\text{O}_3(aq) + \text{I}_2(aq) \rightarrow \text{Na}_2\text{S}_4\text{O}_6(aq) + 2\text{NaI}(aq) \)
CHAPTER 7
Reactions in Aqueous Solution

CHAPTER ANSWERS
2. Driving forces are types of changes in a system that pull a reaction in the direction of product formation; driving forces discussed in Chapter 7 are formation of a solid, formation of water, formation of a gas, and transfer of electrons.

4. The net charge of a precipitate must be zero. The total number of positive charges equals the total number of negative charges.

6. ions

8. Chemists know when a solution contains independent ions because such a solution will readily allow an electrical current to pass through it. The simplest experiment that demonstrates this uses the sort of light–bulb conductivity apparatus described in the text; if the light bulb glows strongly, then the solution must contain a lot of ions that conduct the electricity well.

10. Answer depends on student choices.

12.
   a. soluble; Rule 3
   b. Rule 6: Most sulfide salts are only slightly soluble.
   c. Rule 5: Most hydroxides are only slightly soluble.
   d. soluble; Rule 2
   e. soluble; Rule 4
   f. Rule 6: Most sulfide salts are only slightly soluble.
   g. soluble; Rule 2
   h. Rule 6: Most carbonate salts are only slightly soluble

14.
   a. Rule 5: Most hydroxides are only slightly soluble.
   b. Rule 6: Most phosphate salts are only slightly soluble.
   c. Rule 6: Most carbonate salts are only slightly soluble.
   d. Rule 6: Most sulfide salts are only slightly soluble.

16.
   a. MnCO₃; Rule 6: Most carbonates are only slightly soluble.
   b. CaSO₄; Rule 4, exception for sulfates
   c. Hg₂Cl₂; Rule 3, exception for chlorides
   d. soluble
e. Ni(OH)$_2$; Rule 5: Most hydroxides are only slightly soluble.

f. BaSO$_4$; Rule 4, exception for sulfates

18. The precipitates are marked in boldface type.

a. Rule 6: Most sulfide salts are insoluble. Na$_2$S(aq) + CuCl$_2$(aq) → CuS(s) + 2NaCl(aq)

b. Rule 6: Most phosphate salts are insoluble. K$_3$PO$_4$(aq) + AlCl$_3$(aq) → 3KCl(aq) + AlPO$_4$(s)

c. Rule 4: barium sulfate is a listed exception. H$_2$SO$_4$(aq) + BaCl$_2$(aq) → BaSO$_4$(s) + 2HCl(aq)

d. Rule 5: Most hydroxide compounds are insoluble. 3NaOH(aq) + FeCl$_3$(aq) → 3NaCl(aq) + Fe(OH)$_3$(s)

e. Rule 3: a listed exception for chlorides. 2NaCl(aq) + Hg$_2$(NO$_3$)$_2$(aq) → 2NaNO$_3$(aq) + Hg$_2$Cl$_2$(s)

f. Rule 6: Most carbonate salts are insoluble. 3K$_2$CO$_3$(aq) + 2Cr(C$_2$H$_3$O$_2$)$_3$(aq) → 6KC$_2$H$_3$O$_2$(aq) + Cr$_2$(CO$_3$)$_3$(s)

20. Hint: When balancing equations involving polyatomic ions, especially in precipitation reactions, balance the polyatomic ions as a unit, not in terms of the atoms the polyatomic ions contain (e.g., treat nitrate ion, NO$_3^-$, as a single entity, not as one nitrogen and three oxygen atoms). When finished balancing, however, be sure to count the individual number of atoms of each type on each side of the equation.

a. CaCl$_2$(aq) + AgNO$_3$(aq) → Ca(NO$_3$)$_2$(aq) + AgCl(s)
   balance chlorine: CaCl$_2$(aq) + AgNO$_3$(aq) → Ca(NO$_3$)$_2$(aq) + 2AgCl(s)
   balance silver: CaCl$_2$(aq) + 2AgNO$_3$(aq) → Ca(NO$_3$)$_2$(aq) + 2AgCl(s)
   balanced equation: CaCl$_2$(aq) + 2AgNO$_3$(aq) → Ca(NO$_3$)$_2$(aq) + 2AgCl(s)

b. AgNO$_3$(aq) + K$_2$CrO$_4$(aq) → Ag$_2$CrO$_4$(s) + KNO$_3$(aq)
   balance silver: 2AgNO$_3$(aq) + K$_2$CrO$_4$(aq) → Ag$_2$CrO$_4$(s) + KNO$_3$(aq)
   balance nitrate ion: 2AgNO$_3$(aq) + K$_2$CrO$_4$(aq) → Ag$_2$CrO$_4$(s) + 2KNO$_3$(aq)
   balanced equation: 2AgNO$_3$(aq) + K$_2$CrO$_4$(aq) → Ag$_2$CrO$_4$(s) + 2KNO$_3$(aq)

c. BaCl$_2$(aq) + K$_2$SO$_4$(aq) → BaSO$_4$(s) + 2KCl(aq)
   balance potassium: BaCl$_2$(aq) + K$_2$SO$_4$(aq) → BaSO$_4$(s) + 2KCl(aq)
   balanced equation: BaCl$_2$(aq) + K$_2$SO$_4$(aq) → BaSO$_4$(s) + 2KCl(aq)

22. The products are determined by having the ions “switch partners.” For example, for a general reaction AB + CD →, the possible products are AD and CB if the ions switch partners. If either AD or CB is insoluble, then a precipitation reaction has occurred. In the following reaction, the formula of the precipitate is given in boldface type.

a. CaCl$_2$(aq) + 2AgC$_2$H$_3$O$_2$(aq) → 2AgCl(s) + Ca(C$_2$H$_3$O$_2$)$_2$(aq)
   Rule 3; exception for chloride

b. Ba(NO$_3$)$_2$(aq) + 2NH$_4$OH(aq) → Ba(OH)$_2$(s) + 2NH$_4$NO$_3$(aq)
   Rule 5: Most hydroxides are only slightly soluble.
c. \[ \text{NiCl}_2(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{NiCO}_3(s) + 2\text{NaCl}(aq) \]

Rule 6: Most carbonates are only slightly soluble.

24. Spectator ions are ions that \textit{remain in solution} during a precipitation/double displacement reaction. For example in the reaction

\[ \text{BaCl}_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{KCl}(aq) \]

the \( \text{K}^+ \) and \( \text{Cl}^- \) ions are the spectator ions.

26. The net ionic equation for a reaction indicates \textit{only those ions that go to form the precipitate} and does not show the spectator ions present in the solutes mixed. The identity of the precipitate is determined from the Solubility Rules (Table 7.1).

a. \[ \text{Ca(NO}_3)_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CaSO}_4(s) + 2\text{HNO}_3(aq) \]
\[ \text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{CaSO}_4(s) \]

b. \[ \text{Ni(NO}_3)_2(aq) + 2\text{NaOH}(aq) \rightarrow \text{Ni(OH)}_2(s) + 2\text{NaNO}_3(aq) \]
\[ \text{Ni}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Ni(OH)}_2(s) \]

c. \[ 3\text{(NH}_4)_2\text{S}(aq) + 2\text{FeCl}_3(aq) \rightarrow \text{Fe}_2\text{S}_3(s) + 6\text{NH}_4\text{Cl}(aq) \]
\[ 2\text{Fe}^{3+}(aq) + 3\text{S}^{2-}(aq) \rightarrow \text{Fe}_2\text{S}_3(s) \]

28. \( \text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s) \)

\( \text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{PbCl}_2(s) \)

\( \text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{Hg}_2\text{Cl}_2(s) \)

30. \( \text{Co}^{3+}(aq) + \text{S}^{2-}(aq) \rightarrow \text{CoS}(s) \)

\( 2\text{Co}^{3+}(aq) + 3\text{S}^{2-}(aq) \rightarrow \text{Co}_2\text{S}_3(s) \)

\( \text{Fe}^{2+}(aq) + \text{S}^{2-}(aq) \rightarrow \text{FeS}(s) \)

\( 2\text{Fe}^{3+}(aq) + 3\text{S}^{2-}(aq) \rightarrow \text{Fe}_2\text{S}_3(s) \)

32. Strong bases fully produce hydroxide ions when dissolved in water. The strong bases are also strong electrolytes.

34. acids: \( \text{HCl}, \text{H}_2\text{SO}_4, \text{HNO}_3, \text{HClO}_4, \text{HBr} \)

bases: \( \text{NaOH}, \text{KOH}, \text{RbOH}, \text{CsOH} \)

36. A salt is the ionic product remaining in solution when an acid neutralizes a base. For example, in the reaction \( \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) \), sodium chloride is the salt produced by the neutralization reaction.

38. \( \text{RbOH}(s) \rightarrow \text{Rb}^+(aq) + \text{OH}^-(aq) \)

\( \text{CsOH}(s) \rightarrow \text{Cs}^+(aq) + \text{OH}^-(aq) \)

40. In general, the salt formed in an aqueous acid–base reaction consists of the \textit{positive ion of the base} involved in the reaction combined with the \textit{negative ion of the acid}. The hydrogen ion of the strong acid combines with the hydroxide ion of the strong base to produce water, which is the other product of the acid–base reactions.

a. \[ \text{KOH}(aq) + \text{HCl}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KCl}(aq) \]

b. \[ \text{NaOH}(aq) + \text{HClO}_4(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaClO}_4(aq) \]
c. \( \text{CsOH}(aq) + \text{HNO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CsNO}_3(aq) \)

d. \( 2\text{KOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq) \)

42. Answer depends on student choice of example: \( \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s) \) is an example.

44. The metallic element loses electrons, and the nonmetallic element gains electrons.

46. Each magnesium atom would lose two electrons. Each oxygen atom would gain two electrons (so the \( \text{O}_2 \) molecule would gain four electrons). Two magnesium atoms would be required to react with each oxygen, \( \text{O}_2 \), molecule. Magnesium ions are charged \( 2^+ \); oxide ions are charged \( 2^- \).

48. Each potassium atom loses one electron. The sulfur atom gains two electrons. So, two potassium atoms are required to react with one sulfur atom.

\[
2 \times (\text{K} \rightarrow \text{K}^+ + e^-) \\
\text{S} + 2e^- \rightarrow \text{S}^{2-}
\]

50.

a. \( \text{P}_4(s) + \text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(s) \)

balance oxygen: \( \text{P}_4(s) + 5\text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(s) \)

balanced equation: \( \text{P}_4(s) + 5\text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(s) \)

b. \( \text{MgO}(s) + \text{C}(s) \rightarrow \text{Mg}(s) + \text{CO}(g) \)

This equation is already balanced.

c. \( \text{Sr}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Sr(OH)}_2(aq) + \text{H}_2(g) \)

balance oxygen: \( \text{Sr}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Sr(OH)}_2(aq) + \text{H}_2(g) \)

balanced equation: \( \text{Sr}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Sr(OH)}_2(aq) + \text{H}_2(g) \)

d. \( \text{Co}(s) + \text{HCl}(aq) \rightarrow \text{CoCl}_2(aq) + \text{H}_2(g) \)

balance hydrogen: \( \text{Co}(s) + 2\text{HCl}(aq) \rightarrow \text{CoCl}_2(aq) + \text{H}_2(g) \)

balanced equation: \( \text{Co}(s) + 2\text{HCl}(aq) \rightarrow \text{CoCl}_2(aq) + \text{H}_2(g) \)

52. Examples of formation of water:

\( \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \)

\( \text{H}_2\text{SO}_4(aq) + 2\text{KOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq) \)

Examples of formation of a gaseous product:

\( \text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g) \)

\( 2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g) \)

54. For each reaction, the type of reaction is first identified, followed by some of the reasoning that leads to this choice (There may be more than one way in which you can recognize a particular type of reaction.).

a. oxidation–reduction (Oxygen changes from the combined state to the elemental state.)

b. oxidation–reduction (Zinc changes from the elemental to the combined state; hydrogen changes from the combined to the elemental state.)

c. acid–base (\( \text{H}_2\text{SO}_4 \) is a strong acid, and \( \text{NaOH} \) is a strong base; water and a salt are formed.)
d. acid–base, precipitation [H₂SO₄ is a strong acid, and Ba(OH)₂ is a base; water and a salt are formed; an insoluble product forms.]

e. precipitation (From the Solubility Rules of Table 7.1, AgCl is only slightly soluble.)

f. precipitation [From the Solubility Rules of Table 7.1, Cu(OH)₂ is only slightly soluble.]

g. oxidation–reduction (Chlorine and fluorine change from the elemental to the combined state.)

h. oxidation–reduction (Oxygen changes from the elemental to the combined state.)

i. acid–base (HNO₃ is a strong acid and Ca(OH)₂ is a strong base; a salt and water are formed.)

56. oxidation–reduction

58. A decomposition reaction is one in which a given compound is broken down into simpler compounds or constituent elements. The reactions

\[ \text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \]
\[ 2\text{HgO}(\text{s}) \rightarrow 2\text{Hg}(\text{l}) + \text{O}_2(\text{g}) \]

both represent decomposition reactions. Such reactions often (but not necessarily always) may be classified in other ways. For example, the reaction of HgO(s) is also an oxidation–reduction reaction.

60. Compounds like those in this problem, which contain only carbon and hydrogen, are called hydrocarbons. When a hydrocarbon is reacted with oxygen (O₂), the hydrocarbon is almost always converted to carbon dioxide and water vapor. Because water molecules contain an odd number of oxygen atoms, and O₂ contains an even number of oxygen atoms, it is often difficult to balance such equations. For this reason, it is simpler to balance the equation using fractional coefficients if necessary and then multiply by a factor that will give whole number coefficients for the final balanced equation.

a. \[ \text{C}_3\text{H}_8(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \]
   balance carbon: \[ \text{C}_3\text{H}_8(\text{g}) + \text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \]
   balance hydrogen: \[ \text{C}_3\text{H}_8(\text{g}) + \text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g}) \]
   balance oxygen: \[ \text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g}) \]
   balanced equation: \[ 2\text{C}_3\text{H}_8(\text{g}) + 13\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 10\text{H}_2\text{O}(\text{g}) \]

b. \[ \text{C}_2\text{H}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \]
   balance carbon: \[ \text{C}_2\text{H}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \]
   balance hydrogen: \[ \text{C}_2\text{H}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \]
   balance oxygen: \[ \text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \]
   balanced equation: \[ \text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \]

c. \[ \text{C}_4\text{H}_{10}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \]
   balance carbon: \[ \text{C}_4\text{H}_{10}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \]
   balance hydrogen: \[ \text{C}_4\text{H}_{10}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O}(\text{g}) \]
   balance oxygen: \[ \text{C}_4\text{H}_{10}(\text{g}) + (13/2)\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O}(\text{g}) \]
   balanced equation: \[ 2\text{C}_4\text{H}_{10}(\text{g}) + 13\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 10\text{H}_2\text{O}(\text{g}) \]
62. A reaction in which small molecules or atoms combine to make a larger molecule is called a synthesis reaction. An example would be the synthesis of sodium chloride from the elements,

\[ 2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s). \]

A reaction in which a molecule is broken down into simpler molecules or atoms is called a decomposition reaction. An example would be the decomposition of sodium hydrogen carbonate when heated,

\[ 2\text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g). \]

Specific examples will depend on the students’ input.

64.

a. \( 8\text{Fe}(s) + \text{S}_8(s) \rightarrow 8\text{FeS}(s) \)

b. \( 4\text{Co}(s) + 3\text{O}_2(g) \rightarrow 2\text{Co}_2\text{O}_3(s) \)

c. \( \text{Cl}_2\text{O}_7(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HClO}_4(aq) \)

66.

a. \( 2\text{NI}_3(s) \rightarrow \text{N}_2(g) + 3\text{I}_2(s) \)

b. \( \text{BaCO}_3(s) \rightarrow \text{BaO}(s) + \text{CO}_2(g) \)

c. \( \text{Ca}_3\text{H}_12\text{O}_6(s) \rightarrow 6\text{C}(s) + 6\text{H}_2\text{O}(g) \)

d. \( \text{Cu(NH}_3)_4\text{SO}_4(s) \rightarrow \text{CuSO}_4(s) + 4\text{NH}_3(g) \)

e. \( 3\text{NaN}_3(s) \rightarrow \text{Na}_3\text{N}(s) + 4\text{N}_2(g) \)

68. In several cases, the given ion may be precipitated by many reactants. The following are only three of the possible examples.

a. Chloride ion would precipitate when treated with solutions containing silver ion, lead(II) ion, or mercury(I) ion.

\[
\begin{align*}
\text{Ag}^+(aq) + \text{Cl}^-(aq) & \rightarrow \text{AgCl}(s) \\
\text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) & \rightarrow \text{PbCl}_2(s) \\
\text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) & \rightarrow \text{Hg}_2\text{Cl}_2(s)
\end{align*}
\]

b. Calcium ion would precipitate when treated with solutions containing sulfate ion, carbonate ion, and phosphate ion.

\[
\begin{align*}
\text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) & \rightarrow \text{CaSO}_4(s) \\
\text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) & \rightarrow \text{CaCO}_3(s) \\
3\text{Ca}^{2+}(aq) + 2\text{PO}_4^{3-}(aq) & \rightarrow \text{Ca}_3(\text{PO}_4)_2(s)
\end{align*}
\]

c. Iron(III) ion would precipitate when treated with solutions containing hydroxide, sulfide, or carbonate ions.

\[
\begin{align*}
\text{Fe}^{3+}(aq) + 3\text{OH}^-(aq) & \rightarrow \text{Fe(OH)}_3(s) \\
2\text{Fe}^{3+}(aq) + 3\text{S}^{2-}(aq) & \rightarrow \text{Fe}_2\text{S}_3(s) \\
2\text{Fe}^{3+}(aq) + 3\text{CO}_3^{2-}(aq) & \rightarrow \text{Fe}_2(\text{CO}_3)_3(s)
\end{align*}
\]
d. Sulfate ion would precipitate when treated with solutions containing barium ion, calcium ion, or lead(II) ion.

\[
\begin{align*}
\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) & \rightarrow \text{BaSO}_4(s) \\
\text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) & \rightarrow \text{CaSO}_4(s) \\
\text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) & \rightarrow \text{PbSO}_4(s)
\end{align*}
\]

e. Mercury(I) ion would precipitate when treated with solutions containing chloride ion, sulfide ion, or carbonate ion.

\[
\begin{align*}
\text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) & \rightarrow \text{Hg}_2\text{Cl}_2(s) \\
\text{Hg}_2^{2+}(aq) + \text{S}^2-(aq) & \rightarrow \text{Hg}_2\text{S}(s) \\
\text{Hg}_2^{2+}(aq) + \text{CO}_3^{2-}(aq) & \rightarrow \text{Hg}_2\text{CO}_3(s)
\end{align*}
\]

e. Silver ion would precipitate when treated with solutions containing chloride ion, sulfide ion, or carbonate ion.

\[
\begin{align*}
\text{Ag}^+(aq) + \text{Cl}^-(aq) & \rightarrow \text{AgCl}(s) \\
2\text{Ag}^+(aq) + \text{S}^2-(aq) & \rightarrow \text{Ag}_2\text{S}(s) \\
2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) & \rightarrow \text{Ag}_2\text{CO}_3(s)
\end{align*}
\]

70. The formulas of the salts are indicated in boldface type.

a. \( \text{HNO}_3(aq) + \text{KOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KNO}_3(aq) \)

b. \( \text{H}_2\text{SO}_4(aq) + \text{Ba(OH)}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{BaSO}_4(s) \)

c. \( \text{HClO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaClO}_4(aq) \)

d. \( 2\text{HCl}(aq) + \text{Ca(OH)}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{CaCl}_2(aq) \)

72.

a. soluble (Rule 2: Most potassium salts are soluble.)

b. soluble (Rule 2: Most ammonium salts are soluble.)

c. insoluble (Rule 6: Most carbonate salts are only slightly soluble.)

d. insoluble (Rule 6: Most phosphate salts are only slightly soluble.)

e. soluble (Rule 2: Most sodium salts are soluble.)

f. insoluble (Rule 6: Most carbonate salts are only slightly soluble.)

g. soluble (Rule 3: Most chloride salts are soluble.)

74. The precipitates are marked in boldface type.

a. Rule 3: AgCl is listed as an exception.

\( \text{AgNO}_3(aq) + \text{HCl}(aq) \rightarrow \text{AgCl}(s) + \text{HNO}_3(aq) \)

b. Rule 6: Most carbonate salts are only slightly soluble.

\( \text{CuSO}_4(aq) + (\text{NH}_4)_2\text{CO}_3(aq) \rightarrow \text{CuCO}_3(s) + (\text{NH}_4)_2\text{SO}_4(aq) \)

c. Rule 6: Most carbonate salts are only slightly soluble.

\( \text{FeSO}_4(aq) + \text{K}_2\text{CO}_3(aq) \rightarrow \text{FeCO}_3(s) + \text{K}_2\text{SO}_4(aq) \)
d. no reaction
e. Rule 6: Most carbonate salts are only slightly soluble.
\[ \text{Pb(NO}_3\text{)}_2(\text{aq}) + \text{Li}_2\text{CO}_3(\text{aq}) \rightarrow \text{PbCO}_3(s) + 2\text{LiNO}_3(\text{aq}) \]
f. Rule 5: Most hydroxide compounds are only slightly soluble.
\[ \text{SnCl}_4(\text{aq}) + 4\text{NaOH}(\text{aq}) \rightarrow \text{Sn(OH})_4(s) + 4\text{NaCl}(\text{aq}) \]

76. \( \text{Fe}^{2+}(\text{aq}) + \text{S}^2-(\text{aq}) \rightarrow \text{FeS}(s) \)
\[ 2\text{Cr}^{3+}(\text{aq}) + 3\text{S}^2-(\text{aq}) \rightarrow \text{Cr}_2\text{S}_3(s) \]
\[ \text{Ni}^{2+}(\text{aq}) + \text{S}^2-(\text{aq}) \rightarrow \text{NiS}(s) \]

78. These anions tend to form insoluble precipitates with many metal ions. The following are illustrative for cobalt(II) chloride, tin(II) chloride, and copper(II) nitrate reacting with the sodium salts of the given anions.

a. \[ \text{CoCl}_2(\text{aq}) + \text{Na}_2\text{S}(\text{aq}) \rightarrow \text{CoS}(s) + 2\text{NaCl}(\text{aq}) \]
\[ \text{SnCl}_2(\text{aq}) + \text{Na}_2\text{S}(\text{aq}) \rightarrow \text{SnS}(s) + 2\text{NaCl}(\text{aq}) \]
\[ \text{Cu(NO}_3\text{)}_2(\text{aq}) + \text{Na}_2\text{S}(\text{aq}) \rightarrow \text{CuS}(s) + 2\text{NaNO}_3(\text{aq}) \]

b. \[ \text{CoCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{CoCO}_3(s) + 2\text{NaCl}(\text{aq}) \]
\[ \text{SnCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{SnCO}_3(s) + 2\text{NaCl}(\text{aq}) \]
\[ \text{Cu(NO}_3\text{)}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{CuCO}_3(s) + 2\text{NaNO}_3(\text{aq}) \]

c. \[ \text{CoCl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Co(OH)}_2(s) + 2\text{NaCl}(\text{aq}) \]
\[ \text{SnCl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Sn(OH)}_2(s) + 2\text{NaCl}(\text{aq}) \]
\[ \text{Cu(NO}_3\text{)}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Cu(OH)}_2(s) + 2\text{NaNO}_3(\text{aq}) \]

d. \[ 3\text{CoCl}_2(\text{aq}) + 2\text{Na}_3\text{PO}_4(\text{aq}) \rightarrow \text{Co}_3(\text{PO}_4)_2(s) + 6\text{NaCl}(\text{aq}) \]
\[ 3\text{SnCl}_2(\text{aq}) + 2\text{Na}_3\text{PO}_4(\text{aq}) \rightarrow \text{Sn}_3(\text{PO}_4)_2(s) + 6\text{NaCl}(\text{aq}) \]
\[ 3\text{Cu(NO}_3\text{)}_2(\text{aq}) + 2\text{Na}_3\text{PO}_4(\text{aq}) \rightarrow \text{Cu}_3(\text{PO}_4)_2(s) + 6\text{NaNO}_3(\text{aq}) \]

80. 

a. \( \text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}_2 \)
Balance sodium: \( 2\text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}_2 \)
Balanced equation: \( 2\text{Na}(s) + \text{O}_2(g) \rightarrow \text{Na}_2\text{O}_2(s) \)

b. \( \text{Fe}(s) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{FeSO}_4(\text{aq}) + \text{H}_2(g) \)
Equation is already balanced!

c. \( \text{Al}_2\text{O}_3 \rightarrow \text{Al} + \text{O}_2 \)
Balance oxygen: \( 2\text{Al}_2\text{O}_3 \rightarrow 4\text{Al} + 3\text{O}_2 \)
Balance aluminum: \( 2\text{Al}_2\text{O}_3 \rightarrow 4\text{Al} + 3\text{O}_2 \)
Balanced equation: \( 2\text{Al}_2\text{O}_3(s) \rightarrow 4\text{Al}(s) + 3\text{O}_2(g) \)
d. \( \text{Fe} + \text{Br}_2 \rightarrow \text{FeBr}_3 \)
   Balance bromine: \( \text{Fe} + 3\text{Br}_2 \rightarrow 2\text{FeBr}_3 \)
   Balance iron: \( 2\text{Fe} + 3\text{Br}_2 \rightarrow 2\text{FeBr}_3 \)
   Balanced equation: \( 2\text{Fe}(s) + 3\text{Br}_2(l) \rightarrow 2\text{FeBr}_3(s) \)

e. \( \text{Zn} + \text{HNO}_3 \rightarrow \text{Zn(NO}_3)_2 + \text{H}_2 \)
   Balance nitrate ions: \( \text{Zn} + 2\text{HNO}_3 \rightarrow \text{Zn(NO}_3)_2 + \text{H}_2 \)
   Balanced equation: \( \text{Zn}(s) + 2\text{HNO}_3(aq) \rightarrow \text{Zn(NO}_3)_2(aq) + \text{H}_2(g) \)

82.
   a. \( 2\text{C}_4\text{H}_{10}(l) + 13\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g) \)
   b. \( \text{C}_4\text{H}_{10}\text{O}(l) + 6\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 5\text{H}_2\text{O}(g) \)
   c. \( 2\text{C}_4\text{H}_{10}\text{O}_2(l) + 11\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g) \)

84.
   a. \( 2\text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(g) + \text{CO}_2(g) \)
   b. \( 2\text{NaClO}_3(s) \rightarrow 2\text{NaCl}(s) + 3\text{O}_2(g) \)
   c. \( 2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g) \)
   d. \( \text{C}_12\text{H}_22\text{O}_11(s) \rightarrow 12\text{C}(s) + 11\text{H}_2\text{O}(g) \)
   e. \( 2\text{H}_2\text{O}_2(l) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) \)

86. \( \text{Al}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + \text{H}_2(g) \)
   \( \text{Zn}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{H}_2(g) \)
   \( \text{Mg}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{MgSO}_4(aq) + \text{H}_2(g) \)
   \( \text{Co}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CoSO}_4(aq) + \text{H}_2(g) \)
   \( \text{Ni}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{NiSO}_4(aq) + \text{H}_2(g) \)

88.
   a. one
   b. one
   c. two
   d. two
   e. three

90. A very simple example that fits the bill is \( \text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) \).

92.
   a. \( 2\text{C}_3\text{H}_8\text{O}(l) + 9\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 8\text{H}_2\text{O}(g) \)
      oxidation–reduction, combustion
b. \( \text{HCl}(aq) + \text{AgC}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{AgCl}(s) + \text{HC}_2\text{H}_3\text{O}_2(aq) \)
   precipitation, double–displacement

c. \( 3\text{HCl}(aq) + \text{Al(OH)}_3(s) \rightarrow \text{AlCl}_3(aq) + 3\text{H}_2\text{O}(l) \)
   acid–base, double–displacement

d. \( 2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) \)
   oxidation–reduction, decomposition

e. \( \text{N}_2\text{H}_4(l) + \text{O}_2(g) \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(g) \)
   oxidation–reduction, combustion

94. \( 2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s) \)
\( 2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s) \)
\( \text{Zn}(s) + \text{Cl}_2(g) \rightarrow \text{ZnCl}_2(s) \)
\( \text{Ca}(s) + \text{Cl}_2(g) \rightarrow \text{CaCl}_2(s) \)
\( 2\text{Fe}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{FeCl}_3(s); \text{Fe}(s) + \text{Cl}_2(g) \rightarrow \text{FeCl}_2(s) \)
Cumulative Review Chapters 6 & 7

ANSWERS

2. A chemical equation indicates the substances necessary for a chemical reaction to take place as well as what is produced by that chemical reaction. The substances to the left of the arrow in a chemical equation are called the reactants; those to the right of the arrow are referred to as the products. In addition, if a chemical equation has been balanced, then the equation indicates the relative proportions in which the reactant molecules combine to form the product molecules.

4. It is never permissible to change the subscripts of a formula when balancing a chemical equation. Changing the subscripts changes the identity of a substance from one chemical to another. For example, consider the unbalanced chemical equation

\[ \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \]

If you changed the formula of the product from \( \text{H}_2\text{O}(\text{l}) \) to \( \text{H}_2\text{O}_2(\text{l}) \), the equation would appear to be “balanced.” However, \( \text{H}_2\text{O} \) is water whereas \( \text{H}_2\text{O}_2 \) is hydrogen peroxide—a completely different chemical substance that is not prepared by reaction of the elements hydrogen and oxygen.

When we balance a chemical equation, it is permitted to adjust only the coefficients of a formula because changing a coefficient merely changes the number of molecules of a substance being used in the reaction without changing the identity of the substance. For the example above, we can balance the equation by putting coefficients of 2 in front of the formulas of \( \text{H}_2 \) and \( \text{H}_2\text{O} \); these coefficients do not change the nature of what is reacting and what product is formed.

\[ 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) \]

6. A precipitation reaction is one in which a solid forms when the reactants are combined; the solid is called a precipitate. If you were to perform such a reaction, the mixture would turn cloudy as the reactants were combined, and a solid would eventually settle from the mixture on standing. There are many examples of such precipitation reactions; consult the solubility rules in Table 7.1 if you need help. One example would be to combine barium nitrate and sodium carbonate solutions. A precipitate of barium carbonate would form.

\[ \text{Ba(NO}_3\text{)}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{BaCO}_3(\text{s}) + 2\text{NaNO}_3(\text{aq}) \]

8. In summary, nearly all compounds containing the nitrate, sodium, potassium, and ammonium ions are soluble in water. Most salts containing the chloride and sulfate ions are soluble in water, with specific exceptions (See Table 7.1 for these exceptions.). Most compounds containing the hydroxide, sulfide, carbonate, and phosphate ions are not soluble in water unless the compound also contains one of the cations mentioned above (\( \text{Na}^+, \text{K}^+, \text{NH}_4^+ \)).

The solubility rules are phrased as if you had a sample of a given solute and wanted to see if you could dissolve it in water. These rules can also be applied, however, to predict the identity of the solid produced in a precipitation reaction. A given combination of ions will not be soluble in water whether you take a pure compound out of a reagent bottle or whether you generate the insoluble combination of ions during a chemical reaction. For example, the solubility rules say that \( \text{BaSO}_4 \) is not soluble in water. This not only means that a pure sample of \( \text{BaSO}_4 \) taken from a reagent bottle will not dissolve in water, but also that if \( \text{Ba}^{2+} \) ion and \( \text{SO}_4^{2-} \) ion end up together in
the same solution, they will precipitate as BaSO₄. If we were to combine barium chloride and sulfuric acid solutions

\[ \text{BaCl}_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{HCl}(\text{aq}) \]

then, because barium sulfate is not soluble in water, a precipitate of BaSO₄(s) would form. Because a precipitate of BaSO₄(s) would form no matter what barium compound or what sulfate compound were mixed, we can write the net ionic equation for the reaction as

\[ \text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) \]

Thus if, for example, barium nitrate solution were combined with sodium sulfate solution, a precipitate of BaSO₄ would form. Barium sulfate is insoluble in water regardless of its source.

10. Acids (such as the citric acid found in citrus fruits and the acetic acid found in vinegar) were first noted primarily because of their sour taste. The first bases noted were characterized by their bitter taste and slippery feel on the skin. Acids and bases chemically react with (neutralize) each other forming water. The net ionic equation is

\[ \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) \]

The strong acids and bases fully ionize when they dissolve in water. Because these substances fully ionize, they are strong electrolytes. The common strong acids are HCl (hydrochloric), HNO₃ (nitric), H₂SO₄ (sulfuric), and HClO₄ (perchloric). The most common strong bases are the alkali metal hydroxides, particularly NaOH (sodium hydroxide) and KOH (potassium hydroxide).

12. Oxidation–reduction reactions are electron-transfer reactions. Oxidation represents a loss of electrons by an atom, molecule, or ion whereas reduction is the gain of electrons by such a species. Because an oxidation–reduction process represents the transfer of electrons between species, you can’t have one without the other also taking place. The electrons lost by one species must be gained by some other species. An example of a simple oxidation reduction reaction between a metal and a nonmetal could be the following

\[ \text{Mg}(\text{s}) + \text{F}_2(\text{g}) \rightarrow \text{MgF}_2(\text{s}) \]

In this process, Mg atoms lose two electrons each to become Mg²⁺ ions in MgF₂; Mg is oxidized. Each F atom of F₂ gains one electron to become an F⁻ ion for a total of two electrons gained for each F₂ molecule; F₂ is reduced.

\[ \text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^- \]

\[ 2(\text{F} + \text{e}^- \rightarrow \text{F}^-) \]

14. In general, a synthesis reaction represents the reaction of elements or simple compounds to produce more complex substances. There are many examples of synthesis reactions, for example

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) \]

\[ \text{NaOH}(\text{aq}) + \text{CO}_2(\text{g}) \rightarrow \text{NaHCO}_3(\text{s}) \]

Decomposition reactions represent the breakdown of a more complex substance into simpler substances. There are many examples of decomposition reactions, for example

\[ 2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) \]

Synthesis and decomposition reactions are very often also oxidation–reduction reactions, especially if an elemental substance reacts or is generated. It is not necessary, however, for synthesis and decomposition reactions to always involve oxidation–reduction. The reaction between NaOH and CO₂ given as an example of a synthesis reaction does not represent oxidation–reduction.
16.  
   a.  \( \text{C(s)} + \text{O}_2(g) \rightarrow \text{CO}_2(g) \)  
   b.  \( 2\text{C(s)} + \text{O}_2(g) \rightarrow 2\text{CO(g)} \)  
   c.  \( 2\text{Li(l)} + 2\text{C(s)} \rightarrow \text{Li}_2\text{C}_2(s) \)  
   d.  \( \text{FeO(s)} + \text{C(s)} \rightarrow \text{Fe(l)} + \text{CO(g)} \)  
   e.  \( \text{C(s)} + 2\text{F}_2(g) \rightarrow \text{CF}_4(g) \)  

18.  
   a.  \( \text{Ba(NO}_3)_2(aq) + \text{K}_2\text{CrO}_4(aq) \rightarrow \text{BaCrO}_4(s) + 2\text{KNO}_3(aq) \)  
   b.  \( \text{NaOH(aq)} + \text{CH}_3\text{COOH(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{NaCH}_3\text{COO(aq)} \) (Then evaporate the water from the solution.)  
   c.  \( \text{AgNO}_3(aq) + \text{NaCl(aq)} \rightarrow \text{AgCl(s)} + \text{NaNO}_3(aq) \)  
   d.  \( \text{Pb(NO}_3)_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{PbSO}_4(s) + 2\text{HNO}_3(aq) \)  
   e.  \( 2\text{NaOH(aq)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)} \) (Then evaporate the water from the solution.)  
   f.  \( \text{Ba(NO}_3)_2(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{BaCO}_3(s) + 2\text{NaNO}_3(aq) \)  

20.  
   a.  \( \text{FeO(s)} + 2\text{HNO}_3(aq) \rightarrow \text{Fe(NO}_3)_2(aq) + \text{H}_2\text{O(l)} \)  
      acid–base, double-displacement  
   b.  \( 2\text{Mg(s)} + 2\text{CO}_2(g) + \text{O}_2(g) \rightarrow 2\text{MgCO}_3(s) \)  
      synthesis; oxidation–reduction  
   c.  \( 2\text{NaOH(s)} + \text{CuSO}_4(aq) \rightarrow \text{Cu(OH)_2(s)} + \text{Na}_2\text{SO}_4(aq) \)  
      precipitation, double-displacement  
   d.  \( \text{HI(aq)} + \text{KOH(aq)} \rightarrow \text{KI(aq)} + \text{H}_2\text{O(l)} \)  
      acid–base, double-displacement  
   e.  \( \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O(g)} \)  
      combustion; oxidation–reduction  
   f.  \( \text{Co(NH}_3)_6\text{Cl}_3(s) \rightarrow \text{CoCl}_2(s) + 6\text{NH}_3(g) \)  
      decomposition  
   g.  \( 2\text{HCl(aq)} + \text{Pb(C}_2\text{H}_3\text{O}_2)_2(aq) \rightarrow 2\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{PbCl}_2(s) \)  
      precipitation, double-displacement  
   h.  \( \text{C}_12\text{H}_22\text{O}_11(s) \rightarrow 12\text{C(s)} + 11\text{H}_2\text{O(g)} \)  
      decomposition; oxidation–reduction  
   i.  \( 2\text{Al(s)} + 6\text{HNO}_3(aq) \rightarrow 2\text{Al(NO}_3)_3(aq) + 3\text{H}_2(g) \)  
      oxidation–reduction; single-displacement
j. \[ 4B(s) + 3O_2(g) \rightarrow 2B_2O_3(s) \]
synthesis; oxidation–reduction

22. Specific examples will depend on students’ responses. The following are general equations that illustrate each type of reaction:

Precipitation: typical when solutions of two ionic solutes are mixed, and one of the new combinations of ions is insoluble.

\[ A^+B^-(aq) + C^+D^-(aq) \rightarrow AD(s) + C^+D^-(aq) \]

Single-displacement: one element replaces a less reactive element from a compound.

\[ A(s) + B^+(aq) \rightarrow A^+C^-(aq) + B(s) \]

Combustion: a rapid oxidation reaction most commonly involving \( O_2(g) \). Most examples in the text involve the combustion of hydrocarbons or hydrocarbon derivatives.

\[(\text{hydrocarbon or derivative}) + O_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)\]

Synthesis: elements or simple compounds combine to make more complicated molecules.

\[ A(s) + B(s) \rightarrow AB(s) \]

Oxidation–reduction: reactions in which electrons are transferred from one species to another. Examples of oxidation–reduction reactions include single-displacement, combustion, synthesis, and decomposition reactions.

Decomposition: a compound breaks down into elements and/or simpler compounds.

\[ AB(s) \rightarrow A(s) + B(s) \]

Acid–base neutralization: a neutralization takes place when a proton from an acid combines with a hydroxide ion from a base to make a water molecule.

24.

a. no reaction
b. \[ \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) \]
c. \[ \text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s) \]
d. \[ \text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) \]
e. \[ \text{Fe}^{3+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_2(s) \]
f. \[ \text{Ni}^{2+}(aq) + S^{2-}(aq) \rightarrow \text{NiS}(s) \]
g. \[ \text{Mg}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{MgCO}_3(s) \]
h. no reaction
CHAPTER 8
Chemical Composition

CHAPTER ANSWERS
2. 500. g \times \frac{1 \text{ cork}}{1.63 \text{ g}} = 306.7 = 307 \text{ corks}

500. g \times \frac{1 \text{ stopper}}{4.31 \text{ g}} = 116 \text{ stoppers}

1 \text{ kg of corks contains } \left( 1000 \text{ g} \times \frac{1 \text{ cork}}{1.63 \text{ g}} \right) = 613.49 = 613 \text{ corks}

613 \text{ stoppers would weigh } \left( 613 \text{ stoppers} \times \frac{4.31 \text{ g}}{1 \text{ stopper}} \right) = 2644 \text{ g} = 2640 \text{ g}

The ratio of the mass of a stopper to the mass of a cork is (4.31 g/1.63 g). So the mass of stoppers that contains the same number of stoppers as there are corks in 1000 g of corks is

1000 \text{ g} \times \frac{4.31 \text{ g}}{1.63 \text{ g}} = 2644 \text{ g} = 2640 \text{ g}.

4. The average atomic mass takes into account the various isotopes of an element and the relative abundances in which those isotopes are found.

6.

a. 40.08 \text{ amu Ca} \times \frac{1 \text{ Ca atom}}{40.08 \text{ amu}} = 1 \text{ Ca atom}

b. 919.5 \text{ amu W} \times \frac{1 \text{ W atom}}{183.9 \text{ amu}} = 5 \text{ W atoms}

c. 549.4 \text{ amu Mn} \times \frac{1 \text{ Mn atom}}{54.94 \text{ amu}} = 10 \text{ Mn atoms}

d. 6345 \text{ amu I} \times \frac{1 \text{ I atom}}{126.9 \text{ amu}} = 50 \text{ I atoms}

e. 2072 \text{ amu Pb} \times \frac{1 \text{ Pb atom}}{207.2 \text{ amu}} = 10 \text{ Pb atoms}
8. One tin atom has a mass of 118.7 amu.

A sample containing 35 tin atoms would weigh: 
\[ 35 \times \frac{118.7 \text{ amu}}{1 \text{ atom}} = 4155 \text{ amu} \]

2967.5 amu of tin would represent: 
\[ 2967.5 \text{ amu} \times \frac{1 \text{ tin atom}}{118.7 \text{ amu}} = 25 \text{ tin atoms}. \]

10. \(3 \times \) Avogadro’s number \((3 \times 6.022 \times 10^{23} = 1.807 \times 10^{24}, 3.00 \text{ mol})\)

12. 32.00 g of O\(_2\) (the molar mass of O\(_2\)) contains the same number of atoms as 28.02 g of N\(_2\) (the molar mass of N\(_2\)). Each quantity represents Avogadro’s number of its respective molecules.

14. The ratio of the atomic mass of Co to the atomic mass of F is \((58.93 \text{ amu}/19.00 \text{ amu})\), and the mass of cobalt is given by

\[ 57.0 \text{ g} \times \frac{58.93 \text{ amu}}{19.00 \text{ amu}} = 177 \text{ g Co}. \]

16. 14.01 g of nitrogen atoms contains \(6.022 \times 10^{23}\) N atoms; therefore the mass of one nitrogen atom is given by

\[ 1 \text{ N atom} \times \frac{14.01 \text{ g}}{6.022 \times 10^{23} \text{ atoms}} = 2.32 \times 10^{-23} \text{ g} \]

18. 0.50 mol O atoms \(\times \frac{16.00 \text{ g}}{1 \text{ mol}} = 8.0 \text{ g O} \)

\[ 4 \text{ mol H atoms} \times \frac{1.008 \text{ g H}}{1 \text{ mol}} = 4 \text{ g H} \]

Half a mole of O atoms weighs more than 4 moles of H atoms.

20.

a. 66.50 g F \(\times \frac{1 \text{ mol}}{19.00 \text{ g}} = 3.500 \text{ mol of F atoms} \)

b. 401.2 mg Hg \(\times \frac{1 \text{ mmol}}{200.6 \text{ mg}} = 2.000 \text{ mmol Hg} (1 \text{ mmol} = 1/1000 \text{ mol}) \)

c. 84.27 g Si \(\times \frac{1 \text{ mol}}{28.09 \text{ g}} = 3.000 \text{ mol Si} \)

d. 48.78 g Pt \(\times \frac{1 \text{ mol}}{195.1 \text{ g}} = 0.2500 \text{ mol Pt} \)

e. 2431 g Mg \(\times \frac{1 \text{ mol}}{24.31 \text{ g}} = 100.0 \text{ mol Mg} \)

f. 47.97 g \(\times \frac{1 \text{ mol}}{95.94 \text{ g}} = 0.5000 \text{ mol Mo} \)
22.  

a. \(1.76 \times 10^{-3} \text{ mol Cs} \times \frac{132.9 \text{ g}}{1 \text{ mol}} = 2.34 \times 10^{-1} \text{ g Cs}\)

b. \(0.0125 \text{ mol Ne} \times \frac{20.18 \text{ g}}{1 \text{ mol}} = 0.252 \text{ g Ne}\)

c. \(5.29 \times 10^3 \text{ mol Pb} \times \frac{207.2 \text{ g}}{1 \text{ mol}} = 1.10 \times 10^6 \text{ g Pb}\)

d. \(0.00000122 \text{ mol Na} \times \frac{22.99 \text{ g}}{1 \text{ mol}} = 2.80 \times 10^{-5} \text{ g Na}\)

e. \(5.51 \text{ millimol As} \times \frac{74.92 \text{ g}}{1 \text{ mol}} = 413 \text{ mg} = 0.413 \text{ g As}\)

f. \(8.72 \text{ mol C} \times \frac{12.01 \text{ g}}{1 \text{ mol}} = 105 \text{ g C}\)

24.  

a. \(425 \text{ Na atoms} \times \frac{22.99 \text{ amu}}{1 \text{ Na atom}} = 9.77 \times 10^3 \text{ amu}\)

b. \(425 \text{ Na atoms} \times \frac{22.99 \text{ g Na}}{6.022 \times 10^{23} \text{ Na atoms}} = 1.62 \times 10^{-20} \text{ g}\)

c. \(425 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 9.77 \times 10^3 \text{ g}\)

d. \(425 \text{ mol Na} \times \frac{6.022 \times 10^{23} \text{ Na atoms}}{1 \text{ mol Na}} = 2.56 \times 10^{26} \text{ sodium atoms}\)

e. \(425 \text{ g Na} \times \frac{6.022 \times 10^{23} \text{ Na atoms}}{22.99 \text{ g Na}} = 1.11 \times 10^{25} \text{ sodium atoms}\)

f. \(425 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 18.5 \text{ mol Na}\)

g. \(425 \text{ g Na} \times \frac{24.31 \text{ g Mg}}{22.99 \text{ g Na}} = 449 \text{ g Mg}\)

26. The molar mass is calculated by summing the individual atomic masses of the atoms in the formula.

28.  

a. carbon monoxide

mass of 1 mol C = 12.01 g

mass of 1 mol O = 16.00 g

molar mass of CO = (12.01 g + 16.00 g) = 28.01 g
b. sodium carbonate

mass of 2 mol Na = 2(22.99 g) = 45.98 g
mass of 1 mol C = 12.01 g
mass of 3 mol O = 3(16.00 g) = 48.00 g
molar mass of Na₂CO₃ = (45.98 g + 12.01 g + 48.00 g) = 105.99 g

c. iron(III) nitrate/ferric nitrate

mass of 1 mol Fe = 55.85 g
mass of 3 mol N = 3(14.01 g) = 42.03 g
mass of 9 mol O = 9(16.00 g) = 96.00 g
molar mass of Fe(NO₃)₃ = (55.85 g + 42.03 g + 96.00 g) = 241.88 g

d. hydrogen iodide

mass of 1 mol H = 1.008 g
mass of 1 mol I = 126.9 g
molar mass of HI = 127.9 g

e. sulfur trioxide

mass of 1 mol S = 32.07 g
mass of 3 mol O = 3(16.00 g) = 48.00 g
molar mass of SO₃ = (32.07 g + 48.00 g) = 80.07 g

30. 

a. aluminum fluoride

mass of 1 mol Al = 26.98 g
mass of 3 mol F = 3(19.00 g) = 57.00 g
molar mass of AlF₃ = (26.98 g + 57.00 g) = 83.98 g

b. sodium phosphate

mass of 3 mol Na = 3(22.99 g) = 68.97 g
mass of 1 mol P = 30.97 g
mass of 4 mol O = 4(16.00 g) = 64.00 g
molar mass of Na₃PO₄ = (68.97 g + 30.97 g + 64.00 g) = 163.94 g

c. magnesium carbonate

mass of 1 mol Mg = 24.31 g
mass of 1 mol C = 12.01 g
mass of 3 mol O = 3(16.00 g) = 48.00 g
molar mass of MgCO₃ = (24.31 g + 12.01 g + 48.00 g) = 84.32 g
d. lithium hydrogen carbonate/lithium bicarbonate

mass of 1 mol Li = 6.941 g

mass of 1 mol H = 1.008 g

mass of 1 mol C = 12.01 g

mass of 3 mol O = 3(16.00 g) = 48.00 g

molar mass of LiHCO₃ = (6.941 g + 1.008 g + 12.01 g + 48.00 g) = 67.96 g

e. chromium(III) oxide/chromic oxide

mass of 2 mol Cr = 2(52.00 g) = 104.0 g

mass of 3 mol O = 3(16.00 g) = 48.00 g

molar mass of Cr₂O₄ = 152.0 g

32.

a. molar mass NaCl = 58.44 g; 52.1 mg = 0.0521 g

\[ 0.0521 \text{ g} \times \frac{1 \text{ mol}}{58.44 \text{ g}} = 8.92 \times 10^{-4} \text{ mol} \]

b. molar mass MgCO₃ = 84.32 g

\[ 10.5 \text{ g} \times \frac{1 \text{ mol}}{84.32 \text{ g}} = 0.125 \text{ mol} \]

c. molar mass Al₂O₃ = 101.96 g

\[ 4.00 \text{ g} \times \frac{1 \text{ mol}}{101.96 \text{ g}} = 0.0392 \text{ mol} \]

d. molar mass of Fe₂O₃ = 159.7 g

\[ 24.1 \text{ g} \times \frac{1 \text{ mol}}{159.7 \text{ g}} = 0.151 \text{ mol} \]

e. millimolar mass of Li₂CO₃ = 73.89 mg

\[ 125 \text{ mg} \times \frac{1 \text{ mmol}}{73.89 \text{ mg}} = 1.69 \text{ mmol} = 1.69 \times 10^{-3} \text{ mol} \]

f. molar mass of Fe = 55.85 g; 2.25 kg = 2250 g

\[ 2250 \text{ g} \times \frac{1 \text{ mol}}{55.85 \text{ g}} = 40.3 \text{ mol} \]

34.

a. molar mass of NaH₂PO₄ = 120.0 g

\[ 4.26 \times 10^{-3} \text{ g} \times \frac{1 \text{ mol}}{120.0 \text{ g}} = 3.55 \times 10^{-5} \text{ mol} \]
36.

a. molar mass of CO = 28.01 g  
\[ 0.00471 \text{ mol} \times \frac{28.01 \text{ g}}{1 \text{ mol}} = 0.132 \text{ g} \]

b. molar mass of AuCl₃ = 303.4 g  
\[ 1.75 \times 10^{-6} \text{ mol AuCl}_3 \times \frac{303.4 \text{ g}}{1 \text{ mol}} = 5.31 \times 10^{-4} \text{ g} \]

c. molar mass of FeCl₃ = 162.2 g  
\[ 228 \text{ mol FeCl}_3 \times \frac{162.2 \text{ g}}{1 \text{ mol}} = 3.70 \times 10^4 \text{ g} \]

d. molar mass of K₃PO₄ = 212.3 g; 2.98 millimol = 0.00298 mol 
\[ 0.00298 \text{ mol K}_3\text{PO}_4 \times \frac{212.3 \text{ g}}{1 \text{ mol}} = 0.633 \text{ g} \]

e. molar mass of LiCl = 42.39 g  
\[ 2.71 \times 10^{-3} \text{ mol LiCl} \times \frac{42.39 \text{ g}}{1 \text{ mol}} = 0.115 \text{ g} \]

f. molar mass of NH₃ = 17.03 g  
\[ 6.55 \text{ mol NH}_3 \times \frac{17.03 \text{ g}}{1 \text{ mol}} = 112 \text{ g} \]

38.

a. molar mass NaOCl = 74.44 g  
\[ 0.00421 \text{ mol} \times \frac{74.44 \text{ g}}{1 \text{ mol}} = 0.313 \text{ g} \]
b. molar mass BaH$_2$ = 139.3 g  
\[0.998 \text{ mol} \times \frac{139.3 \text{ g}}{1 \text{ mol}} = 139 \text{ g}\]

c. molar mass AlF$_3$ = 83.98 g  
\[1.99 \times 10^{-2} \text{ mol} \times \frac{83.98 \text{ g}}{1 \text{ mol}} = 1.67 \text{ g}\]

d. molar mass MgCl$_2$ = 95.21 g  
\[0.119 \text{ mol} \times \frac{95.21 \text{ g}}{1 \text{ mol}} = 11.3 \text{ g}\]

e. molar mass Pb = 207.2 g  
\[225 \text{ mol} \times \frac{207.2 \text{ g}}{1 \text{ mol}} = 4.66 \times 10^4 \text{ g}\]

f. molar mass CO$_2$ = 44.01 g  
\[0.101 \text{ mol} \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 4.45 \text{ g}\]

40.

a. 6.37 mol CO \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.84 \times 10^{24} \text{ molecules CO}

b. molar mass of CO = 28.01 g  
\[6.37 \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{28.01 \text{ g}} = 1.37 \times 10^{23} \text{ molecules CO}\]

c. molar mass of H$_2$O = 18.02 g  
\[2.62 \times 10^{-6} \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{18.02 \text{ g}} = 8.76 \times 10^{16} \text{ molecules H}_2\text{O}\]

d. 2.62 \times 10^{-6} \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.58 \times 10^{18} \text{ molec. H}_2\text{O}

e. molar mass of C$_6$H$_6$ = 78.11 g  
\[5.23 \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{78.11 \text{ g}} = 4.03 \times 10^{22} \text{ molecules C}_6\text{H}_6\]

42.

a. molar mass of Na$_2$SO$_4$ = 142.1 g  
\[2.01 \text{ g} \text{ Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.1 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol Na}_2\text{SO}_4} = 0.0141 \text{ mol S}\]
b. molar mass of Na$_2$SO$_3$ = 126.1 g

\[
2.01 \text{ g Na}_2\text{SO}_3 \times \frac{1 \text{ mol Na}_2\text{SO}_3}{126.1 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol Na}_2\text{SO}_3} = 0.0159 \text{ mol S}
\]

c. molar mass of Na$_2$S = 78.05 g

\[
2.01 \text{ g Na}_2\text{S} \times \frac{1 \text{ mol Na}_2\text{S}}{78.05 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol Na}_2\text{S}} = 0.0258 \text{ mol S}
\]

d. molar mass of Na$_2$S$_2$O$_3$ = 158.1 g

\[
2.01 \text{ g Na}_2\text{S}_2\text{O}_3 \times \frac{1 \text{ mol Na}_2\text{S}_2\text{O}_3}{158.1 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol Na}_2\text{S}} = 0.0127 \text{ mol S}
\]

44. less than

46.

a. mass of Cu present = 2(63.55 g) = 127.1 g

mass of O present = 16.00 g

molar mass of Cu$_2$O = 143.1 g

\[
\% \text{ Cu} = \frac{127.1 \text{ g Cu}}{143.1 \text{ g}} \times 100 = 88.82\% \text{ Cu}
\]

\[
\% \text{ O} = \frac{16.00 \text{ g O}}{143.1 \text{ g}} \times 100 = 11.18\% \text{ O}
\]

b. mass of Cu present = 63.55 g

mass of O present = 16.00 g

molar mass of CuO = 79.55 g

\[
\% \text{ Cu} = \frac{63.55 \text{ g Cu}}{79.55 \text{ g}} \times 100 = 79.89\% \text{ Cu}
\]

\[
\% \text{ O} = \frac{16.00 \text{ g O}}{79.55 \text{ g}} \times 100 = 20.11\% \text{ O}
\]

c. mass of Fe present = 55.85 g

mass of O present = 16.00 g

molar mass of FeO = 71.85 g

\[
\% \text{ Fe} = \frac{55.85 \text{ g Fe}}{71.85 \text{ g}} \times 100 = 77.73\% \text{ Fe}
\]

\[
\% \text{ O} = \frac{16.00 \text{ g O}}{71.85 \text{ g}} \times 100 = 22.27\% \text{ O}
\]
d. mass of Fe present \(= 2(55.85 \text{ g}) = 111.7 \text{ g}\)

mass of O present \(= 3(16.00 \text{ g}) = 48.00 \text{ g}\)

molar mass of Fe\(_2\)O\(_3\) = 159.7 g

\[
\% \text{ Fe} = \frac{111.7 \text{ g Fe}}{159.7 \text{ g}} \times 100 = 69.94\% \text{ Fe}
\]

\[
\% \text{ O} = \frac{48.00 \text{ g O}}{159.7 \text{ g}} \times 100 = 30.06\% \text{ O}
\]

e. mass of N present = 14.01 g

mass of O present = 16.00 g

molar mass of NO = 30.01 g

\[
\% \text{ N} = \frac{14.01 \text{ g N}}{30.01 \text{ g}} \times 100 = 46.68\% \text{ N}
\]

\[
\% \text{ O} = \frac{16.00 \text{ g O}}{30.01 \text{ g}} \times 100 = 53.32\% \text{ O}
\]

f. mass of N present = 14.01 g

mass of O present = 2(16.00 g) = 32.00 g

molar mass of NO\(_2\) = 46.01 g

\[
\% \text{ N} = \frac{14.01 \text{ g N}}{46.01 \text{ g}} \times 100 = 30.45\% \text{ N}
\]

\[
\% \text{ O} = \frac{32.00 \text{ g O}}{46.01 \text{ g}} \times 100 = 69.55\% \text{ O}
\]

48.

a. molar mass of CuBr\(_2\) = 223.4

\[
\% \text{ Cu} = \frac{63.55 \text{ g Cu}}{223.4 \text{ g}} \times 100 = 28.45\% \text{ Cu}
\]

b. molar mass of CuBr = 143.5 g

\[
\% \text{ Cu} = \frac{63.55 \text{ g Cu}}{143.5 \text{ g}} \times 100 = 44.29\% \text{ Cu}
\]

c. molar mass of FeCl\(_2\) = 126.75 g

\[
\% \text{ Fe} = \frac{55.85 \text{ g Fe}}{126.75 \text{ g}} \times 100 = 44.06\% \text{ Fe}
\]

d. molar mass of FeCl\(_3\) = 162.2 g

\[
\% \text{ Fe} = \frac{55.85 \text{ g Fe}}{162.2 \text{ g}} \times 100 = 34.43\% \text{ Fe}
\]
8. Chapter 8: Chemical Composition

c. molar mass of CoI₂ = 312.7 g
   \[ \% \text{ Co} = \frac{58.93 \text{ g Co}}{312.7 \text{ g}} \times 100 = 18.85\% \text{ Co} \]

d. molar mass of CoI₃ = 439.6 g
   \[ \% \text{ Co} = \frac{58.93 \text{ g Co}}{439.6 \text{ g}} \times 100 = 13.41\% \text{ Co} \]

e. molar mass of SnO = 134.7 g
   \[ \% \text{ Sn} = \frac{118.7 \text{ g Sn}}{134.7 \text{ g}} \times 100 = 88.12\% \text{ Sn} \]

f. molar mass of SnO₂ = 150.7 g
   \[ \% \text{ Sn} = \frac{118.7 \text{ g Sn}}{150.7 \text{ g}} \times 100 = 78.77\% \text{ Sn} \]

50.

a. molar mass of FeCl₃ = 162.2 g
   \[ \% \text{ Fe} = \frac{55.85 \text{ g Fe}}{162.2 \text{ g}} \times 100 = 34.43\% \text{ Fe} \]

b. molar mass of OF₂ = 54.00 g
   \[ \% \text{ O} = \frac{16.00 \text{ g O}}{54.00} \times 100 = 29.63\% \text{ O} \]

c. molar mass of C₆H₆ = 78.11 g
   \[ \% \text{ C} = \frac{72.06 \text{ g C}}{78.11 \text{ g}} \times 100 = 92.25\% \text{ C} \]

d. molar mass of NH₄ClO₄ = 117.5 g
   \[ \% \text{ N} = \frac{14.01 \text{ g N}}{117.5 \text{ g}} \times 100 = 11.92\% \text{ N} \]

e. molar mass of Ag₂O = 231.8 g
   \[ \% \text{ Ag} = \frac{215.8 \text{ g Ag}}{231.8 \text{ g}} \times 100 = 93.10\% \text{ Ag} \]

f. molar mass of CoCl₂ = 129.83 g
   \[ \% \text{ Co} = \frac{58.93 \text{ g Co}}{129.83 \text{ g}} \times 100 = 45.39\% \text{ Co} \]

g. molar mass of N₂O₄ = 92.02 g
   \[ \% \text{ N} = \frac{28.02 \text{ g N}}{92.02 \text{ g}} \times 100 = 30.45\% \text{ N} \]
h. molar mass of \( \text{MnCl}_2 \) = 125.84 g

\[
\% \text{ Mn} = \frac{54.94 \text{ g Mn}}{125.8 \text{ g}} \times 100 = 43.66\% \text{ Mn}
\]

52.

a. molar mass of \( \text{NH}_4\text{Cl} \) = 53.49 g; molar mass of \( \text{NH}_4^+ \) ion = 18.04 g

\[
\% \text{ NH}_4^+ = \frac{18.04 \text{ g NH}_4^+}{53.49 \text{ g NH}_4\text{Cl}} \times 100 = 33.73\% \text{ NH}_4^+
\]

b. molar mass of \( \text{CuSO}_4 \) = 159.62 g; molar mass of \( \text{Cu}^{2+} \) = 63.55 g

\[
\% \text{ Cu}^{2+} = \frac{63.55 \text{ g Cu}^{2+}}{159.62 \text{ g CuSO}_4} \times 100 = 39.81\% \text{ Cu}^{2+}
\]

c. molar mass of \( \text{AuCl}_3 \) = 303.4 g; molar mass of \( \text{Au}^{3+} \) ion = 197.0 g

\[
\% \text{ Au}^{3+} = \frac{197.0 \text{ g Au}^{3+}}{303.4 \text{ g AuCl}_3} \times 100 = 64.93\% \text{ Au}^{3+}
\]

d. molar mass of \( \text{AgNO}_3 \) = 169.9 g; molar mass of \( \text{Ag}^+ \) ion = 107.9 g

\[
\% \text{ Ag}^+ = \frac{107.9 \text{ g Ag}^+}{169.9 \text{ g AgNO}_3} \times 100 = 63.51\% \text{ Ag}^+
\]

54. The empirical formula indicates the smallest whole number ratio of the number and type of atoms present in a molecule. For example, \( \text{NO}_2 \) and \( \text{N}_2\text{O}_4 \) both have two oxygen atoms for every nitrogen atom and therefore have the same empirical formula

56.

a. yes (Each of these has the empirical formula CH.)

b. no (The number of hydrogen atoms is wrong.)

c. yes (Both have the empirical formula NO.)

d. no (The number of hydrogen and oxygen atoms is wrong.)

58. Assume we have 100.0 g of the compound so that the percentages become masses.

\[
98.55 \text{ g Ba} \times \frac{1 \text{ mol}}{137.3 \text{ g}} = 0.7178 \text{ mol Ba}
\]

\[
1.447 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 1.4355 \text{ mol H}
\]

Dividing both of these numbers of moles by the smaller number of moles (0.7178 mol Ba) gives

\[
\frac{0.7178 \text{ mol Ba}}{0.7178} = 1.000 \text{ mol Ba}
\]

\[
\frac{1.4355 \text{ mol H}}{0.7178 \text{ mol}} = 2.000 \text{ mol H}
\]

The empirical formula is \( \text{BaH}_2 \).
60. Assume we have 100.0 g of the compound, so that the percentages become masses.

\[
28.03 \text{ g Ca} \times \frac{1 \text{ mol}}{40.08 \text{ g}} = 0.6994 \text{ mol Ca}
\]

\[
22.38 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 1.3988 \text{ mol O}
\]

\[
49.59 \text{ g Cl} \times \frac{1 \text{ mol}}{35.45 \text{ g}} = 1.3989 \text{ mol Cl}
\]

Dividing each number of moles by the smallest number of moles (0.6994 mol Ca) gives

\[
\frac{0.6994 \text{ mol Ca}}{0.6994 \text{ mol}} = 1.000 \text{ mol Ca}
\]

\[
\frac{1.3988 \text{ mol O}}{0.6994 \text{ mol}} = 2.000 \text{ mol O}
\]

\[
\frac{1.3939 \text{ mol Cl}}{0.6994 \text{ mol}} = 2.000 \text{ mol Cl}
\]

The empirical formula is \(\text{CaO}_2\text{Cl}_2\), which is more commonly written as \(\text{Ca} (\text{OCl})_2\).

62. Consider 100.0 g of the compound.

\[
29.16 \text{ g N} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 2.081 \text{ mol N}
\]

\[
8.392 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 8.325 \text{ mol H}
\]

\[
12.50 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = 1.041 \text{ mol C}
\]

\[
49.95 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 3.122 \text{ mol O}
\]

Dividing each number of moles by the smallest (1.041 mol C) gives

\[
\frac{2.081 \text{ mol N}}{1.041 \text{ mol}} = 1.999 \text{ mol N}
\]

\[
\frac{8.325 \text{ mol H}}{1.041 \text{ mol}} = 7.997 \text{ mol H}
\]

\[
\frac{1.041 \text{ mol C}}{1.041 \text{ mol}} = 1.000 \text{ mol C}
\]

\[
\frac{3.121 \text{ mol O}}{1.041 \text{ mol}} = 2.998 \text{ mol O}
\]

The empirical formula is \(\text{N}_2\text{H}_6\text{CO}_3\) [i.e., \((\text{NH}_4)_2\text{CO}_3\)].
64. Consider 100.0 g of the compound.

\[
55.06 \text{ g Co} \times \frac{1 \text{ mol}}{58.93 \text{ g}} = 0.9343 \text{ mol Co}
\]

If the sulfide of cobalt is 55.06% Co, then it is 44.94% S by mass.

\[
44.94 \text{ g S} \times \frac{1 \text{ mol}}{32.07 \text{ g}} = 1.401 \text{ mol S}
\]

Dividing each number of moles by the smaller (0.9343 mol Co) gives

\[
\frac{0.09343 \text{ mol Co}}{0.9343} = 1.000 \text{ mol Co}
\]

\[
\frac{1.401 \text{ mol S}}{0.9343 \text{ mol}} = 1.500 \text{ mol S}
\]

Multiplying by two to convert to whole numbers of moles gives the empirical formula for the compound as Co$_2$S$_3$.

66. 2.50 g Al × \(\frac{1 \text{ mol}}{26.98 \text{ g}}\) = 0.09266 mol Al

\[
5.28 \text{ g F} \times \frac{1 \text{ mol}}{19.00 \text{ g}} = 0.2779 \text{ mol F}
\]

Dividing each number of moles by the smaller number of moles gives

\[
\frac{0.09266 \text{ mol Al}}{0.09266 \text{ mol}} = 1.000 \text{ mol Al}
\]

\[
\frac{0.2779 \text{ mol F}}{0.09266 \text{ mol}} = 2.999 \text{ mol F}
\]

The empirical formula is just AlF$_3$. Note the similarity between this problem and question 65; they differ in the way the data is given. In question 65, you were given the mass of the product and first had to calculate how much fluorine had reacted.

68. Consider 100.0 g of the compound.

\[
32.13 \text{ g Al} \times \frac{1 \text{ mol}}{26.98 \text{ g}} = 1.191 \text{ mol Al}
\]

\[
67.87 \text{ g F} \times \frac{1 \text{ mol}}{19.00 \text{ g}} = 3.572 \text{ mol F}
\]

Dividing each number of moles by the smaller number (1.191 mol Al) gives

\[
\frac{1.191 \text{ mol Al}}{1.191 \text{ mol}} = 1.000 \text{ mol Al}
\]

\[
\frac{3.572 \text{ mol F}}{1.191 \text{ mol}} = 2.999 \text{ mol F}
\]
The empirical formula is AlF₃. Compare this question to questions 65 and 66: the three questions illustrate the different forms in which data for calculating empirical formulas may occur.

70. Consider 100.0 g of the compound.

\[ 59.78 \text{ g Li} \times \frac{1 \text{ mol}}{6.941 \text{ g}} = 8.613 \text{ mol Li} \]

\[ 40.22 \text{ g N} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 2.871 \text{ mol N} \]

Dividing each number of moles by the smaller number of moles (2.871 mol N) gives

\[ \frac{8.613 \text{ mol Li}}{2.871 \text{ mol}} = 3.000 \text{ mol Li} \]

\[ \frac{2.871 \text{ mol N}}{2.871 \text{ mol}} = 1.000 \text{ mol N} \]

The empirical formula is Li₃N.

72. Consider 100.0 g of the compound.

\[ 15.77 \text{ g Al} \times \frac{1 \text{ mol}}{26.98 \text{ g}} = 0.5845 \text{ mol Al} \]

\[ 28.11 \text{ g S} \times \frac{1 \text{ mol}}{32.07 \text{ g}} = 0.8765 \text{ mol S} \]

\[ 56.12 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 3.508 \text{ mol O} \]

Dividing each number of moles by the smallest number of moles (0.5845 mol Al) gives

\[ \frac{0.5845 \text{ mol Al}}{0.5845 \text{ mol}} = 1.000 \text{ mol Al} \]

\[ \frac{0.8765 \text{ mol S}}{0.5845 \text{ mol}} = 1.500 \text{ mol S} \]

\[ \frac{3.508 \text{ mol O}}{0.5845 \text{ mol}} = 6.002 \text{ mol O} \]

Multiplying these relative numbers of moles by two to give whole numbers gives the empirical formula as Al₂S₃O₁₂ [i.e., Al₂(SO₄)₃].

74. Compound 1: Assume 100.0 g of the compound.

\[ 22.55 \text{ g P} \times \frac{1 \text{ mol}}{30.97 \text{ g}} = 0.7281 \text{ mol P} \]

\[ 77.45 \text{ g Cl} \times \frac{1 \text{ mol}}{35.45 \text{ g}} = 2.185 \text{ mol Cl} \]

Dividing each number of moles by the smaller (0.7281 mol P) indicates that the formula of Compound 1 is PCl₃.
Compound 2: Assume 100.0 g of the compound.

\[ 14.87 \text{ g P} \times \frac{1 \text{ mol}}{30.97 \text{ g}} = 0.4801 \text{ mol P} \]

\[ 85.13 \text{ g Cl} \times \frac{1 \text{ mol}}{35.45 \text{ g}} = 2.401 \text{ mol Cl} \]

Dividing each number of moles by the smaller (0.4801 mol P) indicates that the formula of Compound 2 is PCl₅.

76. If only the empirical formula is known, the molar mass of the substance must be determined before the molecular formula can be calculated.

78. empirical formula mass of CH = 13 g

\[ n = \frac{\text{molar mass}}{\text{empirical formula mass}} = \frac{78 \text{ g}}{13 \text{ g}} = 6 \]

The molecular formula is (CH)₆ or C₆H₆.

80. empirical formula mass of C₂H₅O = 46 g

\[ n = \frac{\text{molar mass}}{\text{empirical formula mass}} = \frac{90 \text{ g}}{46 \text{ g}} = \approx 2 \]

molecular formula is \( (C₂H₅O)₂ = C₄H₁₀O₂ \)

82. For NO₂, molar mass = 14.01 + 2(16.00) = 46.01 g.

\[ \% \text{ N} = \frac{14.01 \text{ g N}}{46.01 \text{ g}} \times 100 = 30.45\% \text{ N} \]

\[ \% \text{ O} = \frac{2(16.00 \text{ g O})}{46.01 \text{ g}} \times 100 = 69.55\% \text{ O} \]

For N₂O₄, molar mass = 2(14.01 g) + 4(16.00 g) = 92.02 g.

\[ \% \text{ N} = \frac{2(14.01 \text{ g N})}{92.02 \text{ g}} \times 100 = 30.45\% \text{ N} \]

\[ \% \text{ O} = \frac{4(16.00 \text{ g O})}{92.02 \text{ g}} \times 100 = 69.55\% \text{ O} \]

84.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass (g)</th>
<th>Molar Mass</th>
<th>Atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al</td>
<td>5.00</td>
<td>0.185</td>
<td>(1.12 \times 10^{23})</td>
</tr>
<tr>
<td>Fe</td>
<td>0.140</td>
<td>0.00250</td>
<td>(1.51 \times 10^{21})</td>
</tr>
<tr>
<td>Cu</td>
<td>(2.7 \times 10^{-2})</td>
<td>4.3</td>
<td>(2.6 \times 10^{24})</td>
</tr>
<tr>
<td>Mg</td>
<td>0.00250</td>
<td>1.03 × 10^{-4}</td>
<td>(6.19 \times 10^{19})</td>
</tr>
<tr>
<td>Na</td>
<td>0.062</td>
<td>2.7 × 10^{-3}</td>
<td>(1.6 \times 10^{23})</td>
</tr>
<tr>
<td>U</td>
<td>(3.95 \times 10^{-18})</td>
<td>1.66 × 10^{-20}</td>
<td>(1.00 \times 10^{4})</td>
</tr>
</tbody>
</table>
86. mass of 2 mol X = 2(41.2 g) = 82.4 g  
mass of 1 mol Y = 57.7 g = 57.7 g  
mass of 3 mol Z = 3(63.9 g) = 191.7 g  
molar mass of X₂YZ₃ = 331.8 g  

\[
\% X = \frac{82.4 \text{ g}}{331.8 \text{ g}} \times 100 = 24.8\% \text{ X}
\]

\[
\% Y = \frac{57.7 \text{ g}}{331.8 \text{ g}} \times 100 = 17.4\% \text{ Y}
\]

\[
\% Z = \frac{191.7 \text{ g}}{331.8 \text{ g}} \times 100 = 57.8\% \text{ Z}
\]

If the molecular formula were actually X₄Y₂Z₆, the percentage composition would be the same, and the relative mass of each element present would not change. The molecular formula is always a whole-number multiple of the empirical formula.

88. For the first compound (restricted amount of oxygen)

\[
2.118 \text{ g Cu} \times \frac{1 \text{ mol}}{63.55 \text{ g}} = 0.03333 \text{ mol Cu}
\]

\[
0.2666 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 0.01666 \text{ mol O}
\]

Since the number of moles of Cu (0.03333 mol) is twice the number of moles of O (0.01666 mol), the empirical formula is Cu₂O.

For the second compound (stream of pure oxygen)

\[
2.118 \text{ g Cu} \times \frac{1 \text{ mol}}{63.55 \text{ g}} = 0.03333 \text{ mol Cu}
\]

\[
0.5332 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 0.03333 \text{ mol O}
\]

Since the numbers of moles are the same, the empirical formula is CuO.

90.

a. molar mass H₂O = 18.02 g

\[
4.21 \text{ g} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.41 \times 10^{23} \text{ molecules}
\]

The sample contains \(1.41 \times 10^{23}\) oxygen atoms and \(2(1.41 \times 10^{23}) = 2.82 \times 10^{23}\) hydrogen atoms.
b. molar mass $\text{CO}_2 = 44.01$ g

\[
6.81 \text{ g} \times \frac{1 \text{ mol}}{44.01 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 9.32 \times 10^{22} \text{ molecules}
\]

The sample contains $9.32 \times 10^{22}$ carbon atoms and $2(9.32 \times 10^{22}) = 1.86 \times 10^{23}$ oxygen atoms.

c. molar mass $\text{C}_6\text{H}_6 = 78.11$ g

\[
0.000221 \text{ g} \times \frac{1 \text{ mol}}{78.11 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.70 \times 10^{18} \text{ molec.}
\]

The sample contains $6(1.70 \times 10^{18}) = 1.02 \times 10^{19}$ atoms of each element.

d. $2.26 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.36 \times 10^{24} \text{ molecules}$

atoms C = $12(1.36 \times 10^{24}) = 1.63 \times 10^{25}$ atoms

atoms H = $22(1.36 \times 10^{24}) = 2.99 \times 10^{25}$ atoms

atoms O = $11(1.36 \times 10^{24}) = 1.50 \times 10^{25}$ atoms

92.

a. molar mass of $\text{C}_3\text{O}_2 = 3(12.01 \text{ g}) + 2(16.00 \text{ g}) = 68.03$ g

\[
\% \text{ C} = \frac{36.03 \text{ g C}}{68.03 \text{ g}} \times 100 = 52.96\% \text{ C}
\]

\[
7.819 \text{ g} \text{C}_3\text{O}_2 \times \frac{52.96 \text{ g C}}{100.0 \text{ g} \text{C}_3\text{O}_2} = 4.141 \text{ g C}
\]

\[
4.141 \text{ g C} \times \frac{6.022 \times 10^{23} \text{ molecules}}{12.01 \text{ g C}} = 2.076 \times 10^{23} \text{ C atoms}
\]

b. molar mass of CO = $12.01$ g + $16.00$ g = $28.01$ g

\[
\% \text{ C} = \frac{12.01 \text{ g C}}{28.01 \text{ g}} \times 100 = 42.88\% \text{ C}
\]

\[
1.53 \times 10^{21} \text{ molecules CO} \times \frac{1 \text{ C atom}}{1 \text{ molecule CO}} = 1.53 \times 10^{21} \text{ C atoms}
\]

\[
1.53 \times 10^{21} \text{ C atoms} \times \frac{12.01 \text{ g C}}{6.022 \times 10^{23} \text{ C atoms}} = 0.0305 \text{ g C}
\]
c. molar mass of C₆H₆O = 6(12.01 g) + 6(1.008 g) + 16.00 g = 94.11 g

\[
\% C = \frac{72.06 \text{ g C}}{94.11 \text{ g}} \times 100 = 76.57\% \text{ C}
\]

\[
0.200 \text{ mol C}_6\text{H}_6\text{O} \times \frac{6 \text{ mol C}}{1 \text{ mol C}_6\text{H}_6\text{O}} = 1.20 \text{ mol C}
\]

\[
1.20 \text{ mol C} \times \frac{12.01 \text{ g}}{1 \text{ mol C}} = 14.4 \text{ g C}
\]

\[
14.4 \text{ g C} \times \frac{6.022 \times 10^{23} \text{ C atoms}}{12.01 \text{ g C}} = 7.22 \times 10^{23} \text{ C atoms}
\]

94. \ 2.24 \text{ g Co} \times \frac{55.85 \text{ g Fe}}{58.93 \text{ g Co}} = 2.12 \text{ g Fe}

96. \ 5.00 \text{ g Te} \times \frac{200.6 \text{ g Hg}}{127.6 \text{ g Te}} = 7.86 \text{ g Hg}

98. \ 153.8 \text{ g CCl}_4 = 6.022 \times 10^{23} \text{ molecules CCl}_4

\[
1 \text{ molecule} \times \frac{153.8 \text{ g CCl}_4}{6.022 \times 10^{23} \text{ molecules}} = 2.554 \times 10^{-22} \text{ g}
\]

100.

a. molar mass of C₂H₂O₂N = 2(12.01 g) + 5(1.008 g) + 2(16.00 g) + 14.01 g = 75.07 g

\[
5.000 \text{ g} \times \frac{14.01 \text{ g N}}{75.07 \text{ g}} = 0.9331 \text{ g N}
\]

b. molar mass of Mg₃N₂ = 3(24.31 g) + 2(14.01 g) = 100.95 g

\[
5.000 \text{ g} \times \frac{28.02 \text{ g N}}{100.95 \text{ g}} = 1.388 \text{ g N}
\]

c. molar mass of Ca(NO₃)₂ = 40.08 g + 2(14.01 g) + 6(16.00 g) = 164.10 g

\[
5.000 \text{ g} \times \frac{28.02 \text{ g N}}{164.10 \text{ g}} = 0.8537 \text{ g N}
\]

d. molar mass of N₂O₄ = 2(14.01 g) + 4(16.00 g) = 92.02 g

\[
5.000 \text{ g} \times \frac{28.02 \text{ g N}}{92.02 \text{ g}} = 1.522 \text{ g N}
\]
102. Consider 100.0 g of the compound.

\[
16.39 \text{ g Mg} \times \frac{1 \text{ mol}}{24.31 \text{ g}} = 0.6742 \text{ mol Mg}
\]

\[
18.89 \text{ g N} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 1.348 \text{ mol N}
\]

\[
64.72 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 4.045 \text{ mol O}
\]

Dividing each number of moles by the smallest number of moles

\[
\frac{0.6742 \text{ mol Mg}}{0.6742 \text{ mol}} = 1.000 \text{ mol Mg}
\]

\[
\frac{1.348 \text{ mol N}}{0.6742 \text{ mol}} = 1.999 \text{ mol N}
\]

\[
\frac{4.045 \text{ mol O}}{0.6742 \text{ mol}} = 5.999 \text{ mol O}
\]

The empirical formula is MgN\textsubscript{2}O\textsubscript{6} [i.e., Mg(NO\textsubscript{3})\textsubscript{2}].

104. We use the average mass because this average is a weighted average and takes into account both the masses and the relative abundances of the various isotopes.

106. \[
1.98 \times 10^{13} \text{ amu} \times \frac{1 \text{ Na atom}}{22.99 \text{ amu}} = 8.61 \times 10^{11} \text{ Na atoms}
\]

\[
3.01 \times 10^{23} \text{ Na atoms} \times \frac{22.99 \text{ amu}}{1 \text{ Na atom}} = 6.92 \times 10^{24} \text{ amu}
\]

108.

a. \[
5.0 \text{ mol K} \times \frac{39.10 \text{ g}}{1 \text{ mol}} = 195 \text{ g} = 2.0 \times 10^{2} \text{ g K}
\]

b. \[
0.000305 \text{ mol Hg} \times \frac{200.6 \text{ g}}{1 \text{ mol}} = 0.0612 \text{ g Hg}
\]

c. \[
2.31 \times 10^{-5} \text{ mol Mn} \times \frac{54.94 \text{ g}}{1 \text{ mol}} = 1.27 \times 10^{-3} \text{ g Mn}
\]

d. \[
10.5 \text{ mol P} \times \frac{30.97 \text{ g}}{1 \text{ mol}} = 325 \text{ g P}
\]

e. \[
4.9 \times 10^{4} \text{ mol Fe} \times \frac{55.85 \text{ g}}{1 \text{ mol}} = 2.7 \times 10^{6} \text{ g Fe}
\]

f. \[
125 \text{ mol Li} \times \frac{6.941 \text{ g}}{1 \text{ mol}} = 868 \text{ g Li}
\]

g. \[
0.01205 \text{ mol F} \times \frac{19.00 \text{ g}}{1 \text{ mol}} = 0.2290 \text{ g F}
\]
110.  

   a.  mass of 1 mol Fe = (55.85 g) = 55.85 g  
       mass of 1 mol S = (32.07 g) = 32.07 g  
       mass of 4 mol O = 4(16.00 g) = 64.00 g  
       molar mass of FeSO₄ = 151.92 g

   b.  mass of 1 mol Hg = (200.6 g) = 200.6 g  
       mass of 2 mol I = 2(126.9 g) = 253.8 g  
       molar mass of HgI₂ = 454.4 g

   c.  mass of 1 mol Sn = (118.7 g) = 118.7 g  
       mass of 2 mol O = 2(16.00 g) = 32.00 g  
       molar mass of SnO₂ = 150.7 g

   d.  mass of 1 mol Co = (58.93 g) = 58.93 g  
       mass of 2 mol Cl = 2(35.45 g) = 70.90 g  
       molar mass of CoCl₂ = 129.83 g

   e.  mass of 1 mol Cu = (63.55 g) = 63.55 g  
       mass of 2 mol N = 2(14.01 g) = 28.02 g  
       mass of 6 mol O = 6(16.00 g) = 96.00 g  
       molar mass of Cu(NO₃)₂ = 187.57 g

112.  

   a.  molar mass of (NH₄)₂S = 68.15 g  
       21.2 g × \( \frac{1 \text{ mol}}{68.15 \text{ g}} \) = 0.311 mol (NH₄)₂S

   b.  molar mass of Ca(NO₃)₂ = 164.1 g  
       44.3 g × \( \frac{1 \text{ mol}}{164.1 \text{ g}} \) = 0.270 mol Ca(NO₃)₂

   c.  molar mass of Cl₂O = 86.9 g  
       4.35 g × \( \frac{1 \text{ mol}}{86.9 \text{ g}} \) = 0.0501 mol Cl₂O

   d.  1.0 lb = 454 g; molar mass of FeCl₃ = 162.2  
       454 g × \( \frac{1 \text{ mol}}{162.2 \text{ g}} \) = 2.8 mol FeCl₃

   e.  1.0 kg = 1.0 × 10³ g; molar mass of FeCl₃ = 162.2 g  
       1.0 × 10³ g × \( \frac{1 \text{ mol}}{162.2 \text{ g}} \) = 6.2 mol FeCl₃
114.
   a. molar mass of CuSO₄ = 159.62 g
      \[ 2.6 \times 10^{-2} \text{ mol} \times \frac{159.62 \text{ g}}{1 \text{ mol}} = 4.2 \text{ g CuSO₄} \]

   b. molar mass of C₂F₄ = 100.0 g
      \[ 3.05 \times 10^{3} \text{ mol} \times \frac{100.0 \text{ g}}{1 \text{ mol}} = 3.05 \times 10^{5} \text{ g C₂F₄} \]

   c. 7.83 mmol = 0.00783 mol; molar mass of C₃H₈ = 68.11 g
      \[ 0.00783 \text{ mol} \times \frac{68.11 \text{ g}}{1 \text{ mol}} = 0.533 \text{ g C₃H₈} \]

   d. molar mass of BiCl₃ = 315.3 g
      \[ 6.30 \text{ mol} \times \frac{315.3 \text{ g}}{1 \text{ mol}} = 1.99 \times 10^{3} \text{ g BiCl₃} \]

   e. molar mass of C₁₂H₂₂O₁₁ = 342.3 g
      \[ 12.2 \text{ mol} \times \frac{342.3 \text{ g}}{1 \text{ mol}} = 4.18 \times 10^{3} \text{ g C₁₂H₂₂O₁₁} \]

116.
   a. molar mass of C₆H₁₂O₆ = 180.2 g
      \[ 3.45 \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{180.2 \text{ g}} = 1.15 \times 10^{22} \text{ molecules C₆H₁₂O₆} \]

   b. \[ 3.45 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 2.08 \times 10^{24} \text{ molecules C₆H₁₂O₆} \]

   c. molar mass of ICl₅ = 304.2 g
      \[ 25.0 \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{304.2 \text{ g}} = 4.95 \times 10^{22} \text{ molecules ICl₅} \]

   d. molar mass of B₂H₆ = 27.67 g
      \[ 1.00 \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{27.67 \text{ g}} = 2.18 \times 10^{22} \text{ molecules B₂H₆} \]

   e. 1.05 mmol = 0.00105 mol
      \[ 0.00105 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ formula units}}{1 \text{ mol}} = 6.32 \times 10^{20} \text{ formula units} \]
118. 

a. mass of Ca present = 3(40.08 g) = 120.24 g 
   mass of P present = 2(30.97 g) = 61.94 g 
   mass of O present = 8(16.00 g) = 128.00 g 
   molar mass of $Ca_3(PO_4)_2 = 310.18$ g 
   
   \[
   \% Ca = \frac{120.24 \text{ g Ca}}{310.18 \text{ g}} \times 100 = 38.76\% \text{ Ca}
   \]

   \[
   \% P = \frac{61.94 \text{ g P}}{310.18 \text{ g}} \times 100 = 19.97\% \text{ P}
   \]

   \[
   \% O = \frac{128.00 \text{ g O}}{310.18 \text{ g}} \times 100 = 41.27\% \text{ O}
   \]

b. mass of Cd present = 112.4 g = 112.4 g 
   mass of S present = 32.07 g = 32.07 g 
   mass of O present = 4(16.00 g) = 64.00 g 
   molar mass of $CdSO_4 = 208.5$ g 
   
   \[
   \% Cd = \frac{112.4 \text{ g Cd}}{208.5 \text{ g}} \times 100 = 53.91\% \text{ Cd}
   \]

   \[
   \% S = \frac{32.07 \text{ g S}}{208.5 \text{ g}} \times 100 = 15.38\% \text{ S}
   \]

   \[
   \% O = \frac{64.00 \text{ g O}}{208.5 \text{ g}} \times 100 = 30.70\% \text{ O}
   \]

c. mass of Fe present = 2(55.85 g) = 111.7 g 
   mass of S present = 3(32.07 g) = 96.21 g 
   mass of O present = 12(16.00 g) = 192.0 g 
   molar mass of $Fe_2(SO_4)_3 = 399.9$ g 
   
   \[
   \% Fe = \frac{111.7 \text{ g Fe}}{399.9 \text{ g}} \times 100 = 27.93\% \text{ Fe}
   \]

   \[
   \% S = \frac{96.21 \text{ g S}}{399.9 \text{ g}} \times 100 = 24.06\% \text{ S}
   \]

   \[
   \% O = \frac{192.0 \text{ g O}}{399.9 \text{ g}} \times 100 = 48.01\% \text{ O}
   \]
d. mass of Mn present = 54.94 g = 54.94 g
mass of Cl present = 2(35.45 g) = 70.90 g
molar mass of MnCl₂ = 125.84 g
\[ \% \text{ Mn} = \frac{54.94 \text{ g Mn}}{125.84 \text{ g}} \times 100 = 43.66\% \text{ Mn} \]
\[ \% \text{ Cl} = \frac{70.90 \text{ g Cl}}{125.84 \text{ g}} \times 100 = 56.34\% \text{ Cl} \]

e. mass of N present = 2(14.01 g) = 28.02 g
mass of H present = 8(1.008 g) = 8.064 g
mass of C present = 12.01 g = 12.01 g
mass of O present = 3(16.00 g) = 48.00 g
molar mass of (NH₄)₂CO₃ = 96.09 g
\[ \% \text{ N} = \frac{28.02 \text{ g N}}{96.09 \text{ g}} \times 100 = 29.16\% \text{ N} \]
\[ \% \text{ H} = \frac{8.064 \text{ g H}}{96.09 \text{ g}} \times 100 = 8.392\% \text{ H} \]
\[ \% \text{ C} = \frac{12.01 \text{ g C}}{96.09 \text{ g}} \times 100 = 12.50\% \text{ C} \]
\[ \% \text{ O} = \frac{48.00 \text{ g O}}{96.09 \text{ g}} \times 100 = 49.95\% \text{ O} \]

f. mass of Na present = 22.99 g = 22.99 g
mass of H present = 1.008 g = 1.008 g
mass of C present = 12.01 g = 12.01 g
mass of O present = 3(16.00 g) = 48.00 g
molar mass of NaHCO₃ = 84.01 g
\[ \% \text{ Na} = \frac{22.99 \text{ g Na}}{84.01 \text{ g}} \times 100 = 27.37\% \text{ Na} \]
\[ \% \text{ H} = \frac{1.008 \text{ g H}}{84.01 \text{ g}} \times 100 = 1.200\% \text{ H} \]
\[ \% \text{ C} = \frac{12.01 \text{ g C}}{84.01 \text{ g}} \times 100 = 14.30\% \text{ C} \]
\[ \% \text{ O} = \frac{48.00 \text{ g O}}{84.01 \text{ g}} \times 100 = 57.14\% \text{ O} \]
g. mass of C present = 12.01 g = 12.01 g
   mass of O present = 2(16.00 g) = 32.00 g
   molar mass of CO$_2$ = 44.01 g
   $\%$ C = $\frac{12.01 \text{ g C}}{44.01 \text{ g}} \times 100 = 27.29\%$ C
   $\%$ O = $\frac{32.00 \text{ g O}}{44.01 \text{ g}} \times 100 = 72.71\%$ O

h. mass of Ag present = 107.9 g = 107.9 g
   mass of N present = 14.01 g = 14.01 g
   mass of O present = 3(16.00 g) = 48.00 g
   molar mass of AgNO$_3$ = 169.9 g
   $\%$ Ag = $\frac{107.9 \text{ g Ag}}{169.9 \text{ g}} \times 100 = 63.51\%$ Ag
   $\%$ N = $\frac{14.01 \text{ g N}}{169.9 \text{ g}} \times 100 = 8.246\%$ N
   $\%$ O = $\frac{48.00 \text{ g O}}{169.9 \text{ g}} \times 100 = 28.25\%$ O

120.

a. $\%$ Fe = $\frac{55.85 \text{ g Fe}}{151.92 \text{ g}} \times 100 = 36.76\%$ Fe

b. $\%$ Ag = $\frac{215.8 \text{ g Ag}}{231.8 \text{ g}} \times 100 = 93.10\%$ Ag

c. $\%$ Sr = $\frac{87.62 \text{ g Sr}}{158.5 \text{ g}} \times 100 = 55.28\%$ Sr

d. $\%$ C = $\frac{48.04 \text{ g C}}{86.09 \text{ g}} \times 100 = 55.80\%$ C

e. $\%$ C = $\frac{12.01 \text{ g C}}{32.04 \text{ g}} \times 100 = 37.48\%$ C

f. $\%$ Al = $\frac{53.96 \text{ g Al}}{101.96 \text{ g}} \times 100 = 52.92\%$ Al

g. $\%$ K = $\frac{39.10 \text{ g K}}{106.55 \text{ g}} \times 100 = 36.70\%$ K

h. $\%$ K = $\frac{39.10 \text{ g K}}{74.55 \text{ g}} \times 100 = 52.45\%$ K
122. \[ \frac{0.2990 \text{ g C}}{12.01 \text{ g}} = 0.02490 \text{ mol C} \]

\[ \frac{0.05849 \text{ g H}}{1.008 \text{ g}} = 0.05803 \text{ mol H} \]

\[ \frac{0.2318 \text{ g N}}{14.01 \text{ g}} = 0.01655 \text{ mol N} \]

\[ \frac{0.1328 \text{ g O}}{16.00 \text{ g}} = 0.008300 \text{ mol O} \]

Dividing each number of moles by the smallest number of moles (0.008300 mol) gives:

\[ \frac{0.02490 \text{ mol C}}{0.008300} = 3.000 \text{ mol C} \]

\[ \frac{0.05803 \text{ mol H}}{0.008300} = 6.992 \text{ mol H} \]

\[ \frac{0.01655 \text{ mol N}}{0.008300} = 1.994 \text{ mol N} \]

\[ \frac{0.008300 \text{ mol O}}{0.008300} = 1.000 \text{ mol O} \]

The empirical formula is C\textsubscript{3}H\textsubscript{7}N\textsubscript{2}O.

124. Mass of oxygen in compound = 4.33 g – 4.01 g = 0.32 g O

\[ \frac{4.01 \text{ g Hg}}{200.6 \text{ g}} = 0.0200 \text{ mol Hg} \]

\[ \frac{0.32 \text{ g O}}{16.00 \text{ g}} = 0.020 \text{ mol O} \]

Since the numbers of moles are equal, the empirical formula is HgO.

126. Assume we have 100.0 g of the compound.

\[ \frac{65.95 \text{ g Ba}}{137.3 \text{ g}} = 0.4803 \text{ mol Ba} \]

\[ \frac{34.05 \text{ g Cl}}{35.45 \text{ g}} = 0.9605 \text{ mol Cl} \]

Dividing each number of moles by the smaller number gives:

\[ \frac{0.4803 \text{ mol Ba}}{0.4803 \text{ mol}} = 1.000 \text{ mol Ba} \]

\[ \frac{0.9605 \text{ mol Cl}}{0.4803 \text{ mol}} = 2.000 \text{ mol Cl} \]

The empirical formula is then BaCl\textsubscript{2}.
CHAPTER 9

Chemical Quantities

CHAPTER ANSWERS

2. The coefficients of the balanced chemical equation for a reaction indicate the relative numbers of moles of each reactant that combine during the process as well as the number of moles of each product formed.

4. Balanced chemical equations tell us in what molar ratios substances combine to form products; not, in what mass proportions they combine.

6.
   a. \((NH_4)_2CO_3(s) \rightarrow 2NH_3(g) + CO_2(g) + H_2O(g)\)

      One formula unit of solid ammonium carbonate decomposes to produce two molecules of ammonia gas, one molecule of carbon dioxide gas, and one molecule of water vapor. One mole of solid ammonium carbonate decomposes into two moles of gaseous ammonia, one mole of carbon dioxide gas, and one mole of water vapor.

   b. \(6Mg(s) + P_4(s) \rightarrow 2Mg_3P_2(s)\)

      Six atoms of magnesium metal react with one molecule of solid phosphorus \((P_4)\) to make two formula units of solid magnesium phosphide. Six moles of magnesium metal react with one mole of phosphorus solid \((P_4)\) to produce two moles of solid magnesium phosphide.

   c. \(4Si(s) + S_8(s) \rightarrow 2Si_2S_4(l)\)

      Four atoms of solid silicon react with one molecule of solid sulfur \((S_8)\) to form two molecules of liquid disilicon tetrasulfide. Four moles of solid silicon react with one mole of solid sulfur \((S_8)\) to form two moles of liquid disilicon tetrasulfide.

   d. \(C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)\)

      One molecule of liquid ethanol burns with three molecules of oxygen gas to produce two molecules of carbon dioxide gas and three molecules of water vapor. One mole of liquid ethanol burns with three moles of oxygen gas to produce two moles of gaseous carbon dioxide and three moles of water vapor.

8. Balanced chemical equations tell us in what molar ratios substances combine to form products; not in what mass proportions they combine. How could 2 g of reactant produce a total of 3 g of products?
10. \( \text{S(s)} + 2\text{H}_2\text{SO}_4(l) \rightarrow 3\text{SO}_2(g) + 2\text{H}_2\text{O}(l) \)

For \( \text{SO}_2 \), \( \frac{3 \text{ mol SO}_2}{1 \text{ mol S}} \)

For \( \text{H}_2\text{O} \), \( \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol S}} \)

For \( \text{H}_2\text{SO}_4 \), \( \frac{2 \text{ mol H}_2\text{SO}_4}{1 \text{ mol S}} \)

12.

a. \( 2\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \)

\[ 5.0 \text{ mol C}_2\text{H}_6 \times \frac{7 \text{ mol O}_2}{2 \text{ mol C}_2\text{H}_6} = 17.5 \text{ mol O}_2 \] (18 mol O_2)

b. \( \text{P}_4(s) + 5\text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(g) \)

\[ 5.0 \text{ mol P}_4 \times \frac{5 \text{ mol O}_2}{1 \text{ mol P}_4} = 25 \text{ mol O}_2 \]

c. \( \text{CaO(s)} + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) \)

\[ 5.0 \text{ mol CaO} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CaO}} = 5.0 \text{ mol CO}_2 \]

d. \( 4\text{Fe(s)} + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \)

\[ 5.0 \text{ mol Fe} \times \frac{3 \text{ mol O}_2}{4 \text{ mol Fe}} = 3.75 \text{ mol O}_2 \] (3.8 mol O_2)

14.

a. \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl(s)} \)

molar mass of \( \text{NH}_4\text{Cl} \), 53.49 g

\[ 0.50 \text{ mol NH}_3 \times \frac{1 \text{ mol NH}_4\text{Cl}}{1 \text{ mol NH}_3} = 0.50 \text{ mol NH}_4\text{Cl} \]

\[ 0.50 \text{ mol NH}_4\text{Cl} \times \frac{53.49 \text{ g NH}_4\text{Cl}}{1 \text{ mol NH}_4\text{Cl}} = 27 \text{ g NH}_4\text{Cl} \]
b. \( \text{CH}_4(g) + 4\text{S}(g) \rightarrow \text{CS}_2(l) + 2\text{H}_2\text{S}(g) \)

molar masses: \( \text{CS}_2, 76.15 \text{ g}; \text{H}_2\text{S}, 34.09 \text{ g} \)

\[
0.50 \text{ mol } \text{S} \times \frac{1 \text{ mol } \text{CS}_2}{4 \text{ mol } \text{S}} = 0.125 \text{ mol } \text{CS}_2 = (0.13 \text{ mol } \text{CS}_2)
\]

\[
0.125 \text{ mol } \text{CS}_2 \times \frac{76.15 \text{ g } \text{CS}_2}{1 \text{ mol } \text{CS}_2} = 9.5 \text{ g } \text{CS}_2
\]

\[
0.50 \text{ mol } \text{S} \times \frac{2 \text{ mol } \text{H}_2\text{S}}{4 \text{ mol } \text{S}} = 0.25 \text{ mol } \text{H}_2\text{S}
\]

\[
0.25 \text{ mol } \text{H}_2\text{S} \times \frac{34.09 \text{ g } \text{H}_2\text{S}}{1 \text{ mol } \text{H}_2\text{S}} = 8.5 \text{ g } \text{H}_2\text{S}
\]

c. \( \text{PCl}_3(l) + 3\text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{PO}_3(aq) + 3\text{HCl}(aq) \)

molar masses: \( \text{H}_3\text{PO}_3, 81.99 \text{ g}; \text{HCl}, 36.46 \text{ g} \)

\[
0.50 \text{ mol } \text{PCl}_3 \times \frac{1 \text{ mol } \text{H}_3\text{PO}_3}{1 \text{ mol } \text{PCl}_3} = 0.50 \text{ mol } \text{H}_3\text{PO}_3
\]

\[
0.50 \text{ mol } \text{H}_3\text{PO}_3 \times \frac{81.99 \text{ g } \text{H}_3\text{PO}_3}{1 \text{ mol } \text{H}_3\text{PO}_3} = 41 \text{ g } \text{H}_3\text{PO}_3
\]

\[
0.50 \text{ mol } \text{PCl}_3 \times \frac{3 \text{ mol } \text{HCl}}{1 \text{ mol } \text{PCl}_3} = 1.5 \text{ mol } \text{HCl}
\]

\[
1.5 \text{ mol } \text{HCl} \times \frac{36.46 \text{ g } \text{HCl}}{1 \text{ mol } \text{HCl}} = 54.7 = 55 \text{ g } \text{HCl}
\]

d. \( \text{NaOH}(s) + \text{CO}_2(g) \rightarrow \text{NaHCO}_3(s) \)

molar mass of \( \text{NaHCO}_3 = 84.01 \text{ g} \)

\[
0.50 \text{ mol } \text{NaOH} \times \frac{1 \text{ mol } \text{NaHCO}_3}{1 \text{ mol } \text{NaOH}} = 0.50 \text{ mol } \text{NaHCO}_3
\]

\[
0.50 \text{ mol } \text{NaHCO}_3 \times \frac{84.01 \text{ g } \text{NaHCO}_3}{1 \text{ mol } \text{NaHCO}_3} = 42 \text{ g } \text{NaHCO}_3
\]

16. Before doing the calculations, the equations must be balanced.

a. \( 4\text{KO}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow 3\text{O}_2(g) + 4\text{KOH}(s) \)

\[
0.625 \text{ mol } \text{KOH} \times \frac{3 \text{ mol } \text{O}_2}{4 \text{ mol } \text{KOH}} = 0.469 \text{ mol } \text{O}_2
\]

b. \( \text{SeO}_2(g) + 2\text{H}_2\text{Se}(g) \rightarrow 3\text{Se}(s) + 2\text{H}_2\text{O}(g) \)

\[
0.625 \text{ mol } \text{H}_2\text{O} \times \frac{3 \text{ mol } \text{Se}}{2 \text{ mol } \text{H}_2\text{O}} = 0.938 \text{ mol } \text{Se}
\]
c. \(2\text{CH}_3\text{CH}_2\text{OH}(l) + \text{O}_2(g) \rightarrow 2\text{CH}_3\text{CHO}(aq) + 2\text{H}_2\text{O}(l)\)

\[
0.625 \text{ mol H}_2\text{O} \times \frac{2 \text{ mol CH}_3\text{CHO}}{2 \text{ mol H}_2\text{O}} = 0.625 \text{ mol CH}_3\text{CHO}
\]

d. \(\text{Fe}_2\text{O}_3(s) + 2\text{Al}(s) \rightarrow 2\text{Fe}(l) + \text{Al}_2\text{O}_3(s)\)

\[
0.625 \text{ mol Al}_2\text{O}_3 \times \frac{2 \text{ mol Fe}}{1 \text{ mol Al}_2\text{O}_3} = 1.25 \text{ mol Fe}
\]

18. Stoichiometry is the process of using a chemical equation to calculate the relative masses of reactants and products involved in a reaction.

20.

a. molar mass \(\text{Li}_2\text{CO}_3 = 73.89 \text{ g}\); \(2.36 \text{ mg} = 0.00236 \text{ g}\)

\[
0.00236 \text{ g Li}_2\text{CO}_3 \times \frac{1 \text{ mol Li}_2\text{CO}_3}{73.89 \text{ g Li}_2\text{CO}_3} = 3.19 \times 10^{-5} \text{ mol Li}_2\text{CO}_3
\]

b. molar mass \(\text{U} = 238.0 \text{ g}\)

\[
1.92 \times 10^{-3} \text{ g U} \times \frac{1 \text{ mol U}}{238.0 \text{ g U}} = 8.07 \times 10^{-6} \text{ mol U}
\]

c. molar mass \(\text{PbCl}_2 = 278.1 \text{ g}\); \(3.21 \text{ kg} = 3.21 \times 10^3 \text{ g}\)

\[
3.21 \times 10^3 \text{ g PbCl}_2 \times \frac{1 \text{ mol PbCl}_2}{278.1 \text{ g}} = 11.5 \text{ mol PbCl}_2
\]

d. molar mass \(\text{C}_6\text{H}_12\text{O}_6 = 180.2 \text{ g}\)

\[
4.62 \text{ g C}_6\text{H}_12\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_12\text{O}_6}{180.2 \text{ g C}_6\text{H}_12\text{O}_6} = 0.0256 \text{ mol C}_6\text{H}_12\text{O}_6
\]

e. molar mass \(\text{KOH} = 56.11 \text{ g}\)

\[
7.75 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} = 0.138 \text{ mol KOH}
\]

22.

a. molar mass of \(\text{CuI} = 190.5 \text{ g}\)

\[
0.624 \text{ mol CuI} \times \frac{190.5 \text{ g CuI}}{1 \text{ mol CuI}} = 119 \text{ g CuI}
\]

b. molar mass of \(\text{Br}_2 = 159.8 \text{ g}\)

\[
4.24 \text{ mol Br}_2 \times \frac{159.8 \text{ g Br}_2}{1 \text{ mol Br}_2} = 678 \text{ g Br}_2
\]

c. molar mass of \(\text{XeF}_4 = 207.4 \text{ g}\)

\[
0.000211 \text{ mol XeF}_4 \times \frac{207.4 \text{ g XeF}_4}{1 \text{ mol XeF}_4} = 0.0438 \text{ g XeF}_4
\]
d. molar mass of \( \text{C}_2\text{H}_4 \) = 28.05 g

\[
9.11 \text{ mol C}_2\text{H}_4 \times \frac{28.05 \text{ g C}_2\text{H}_4}{1 \text{ mol C}_2\text{H}_4} = 256 \text{ g C}_2\text{H}_4
\]

e. molar mass of \( \text{NH}_3 \) = 17.03 g; \( 1.21 \text{ millimol} = 0.00121 \text{ mol} \)

\[
0.00121 \text{ mol NH}_3 \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 0.0206 \text{ g NH}_3
\]

f. molar mass of \( \text{NaOH} \) = 40.00 g

\[
4.25 \text{ mol NaOH} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 170 \text{ g NaOH}
\]

g. molar mass of \( \text{KI} \) = 166.0 g

\[
1.27 \times 10^{-6} \text{ mol KI} \times \frac{166.0 \text{ g KI}}{1 \text{ mol KI}} = 2.11 \times 10^{-4} \text{ g KI}
\]

24. Before any calculations are done, the equations must be balanced.

a. \( \text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g) \)

molar mass \( \text{C}_2\text{H}_5\text{OH} = 46.07 \text{ g} \)

\[
5.00 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.07 \text{ g C}_2\text{H}_5\text{OH}} \times \frac{3 \text{ mol O}_2}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 0.326 \text{ mol O}_2
\]

b. \( \text{P}_4(s) + 5\text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(g) \)

molar mass \( \text{P}_4 = 123.88 \text{ g} \)

\[
5.00 \text{ g P}_4 \times \frac{1 \text{ mol P}_4}{123.88 \text{ g}} \times \frac{5 \text{ mol O}_2}{1 \text{ mol P}_4} = 0.202 \text{ mol O}_2
\]

c. \( \text{MgO(s) + CO}_2(g) \rightarrow \text{MgCO}_3(s) \)

molar mass \( \text{MgO} = 40.31 \text{ g} \)

\[
5.00 \text{ g MgO} \times \frac{1 \text{ mol MgO}}{40.31 \text{ g MgO}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol MgO}} = 0.124 \text{ mol CO}_2
\]

d. \( 2\text{Fe(s) + O}_2(g) \rightarrow 2\text{FeO(s)} \)

molar mass \( \text{Fe} = 55.85 \text{ g} \)

\[
5.00 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol O}_2}{2 \text{ mol Fe}} = 0.0448 \text{ mol O}_2
\]
26.  

(a) \[2\text{BCl}_3(s) + 3\text{H}_2(g) \rightarrow 2\text{B}(s) + 6\text{HCl}(g)\]

molar masses: \(\text{BCl}_3, 117.16 \text{ g}\); \(\text{B}, 10.81 \text{ g}\); \(\text{HCl}, 36.46 \text{ g}\)

\[
15.0 \text{ g BCl}_3 \times \frac{1 \text{ mol BCl}_3}{117.16 \text{ g BCl}_3} = 0.128 \text{ mol BCl}_3
\]

\[
0.128 \text{ mol BCl}_3 \times \frac{2 \text{ mol B}}{2 \text{ mol BCl}_3} \times \frac{10.81 \text{ g B}}{1 \text{ mol B}} = 1.38 \text{ g B}
\]

\[
0.128 \text{ mol BCl}_3 \times \frac{6 \text{ mol HCl}}{2 \text{ mol BCl}_3} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 14.0 \text{ g HCl}
\]

(b) \[2\text{Cu}_2\text{S}(s) + 3\text{O}_2(g) \rightarrow 2\text{Cu}_2\text{O}(s) + 2\text{SO}_2(g)\]

molar masses: \(\text{Cu}_2\text{S}, 159.17 \text{ g}\); \(\text{Cu}_2\text{O}, 143.1 \text{ g}\); \(\text{SO}_2, 64.07 \text{ g}\)

\[
15.0 \text{ g Cu}_2\text{S} \times \frac{1 \text{ mol Cu}_2\text{S}}{159.17 \text{ g Cu}_2\text{S}} = 0.09424 \text{ mol Cu}_2\text{S}
\]

\[
0.09424 \text{ mol Cu}_2\text{S} \times \frac{2 \text{ mol Cu}_2\text{O}}{2 \text{ mol Cu}_2\text{S}} \times \frac{143.1 \text{ g Cu}_2\text{O}}{1 \text{ mol Cu}_2\text{O}} = 13.5 \text{ g Cu}_2\text{O}
\]

\[
0.09424 \text{ mol Cu}_2\text{S} \times \frac{2 \text{ mol SO}_2}{2 \text{ mol Cu}_2\text{S}} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = 6.04 \text{ g SO}_2
\]

(c) \[2\text{Cu}_2\text{O}(s) + \text{Cu}_2\text{S}(s) \rightarrow 6\text{Cu}(s) + \text{SO}_2(g)\]

molar masses: \(\text{Cu}_2\text{S}, 159.17 \text{ g}\); \(\text{Cu}, 63.55 \text{ g}\); \(\text{SO}_2, 64.07 \text{ g}\)

\[
15.0 \text{ g Cu}_2\text{S} \times \frac{1 \text{ mol Cu}_2\text{S}}{159.17 \text{ g Cu}_2\text{S}} = 0.09424 \text{ mol Cu}_2\text{S}
\]

\[
0.09424 \text{ mol Cu}_2\text{S} \times \frac{6 \text{ mol Cu}}{1 \text{ mol Cu}_2\text{S}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 35.9 \text{ g Cu}
\]

\[
0.09424 \text{ mol Cu}_2\text{S} \times \frac{1 \text{ mol SO}_2}{1 \text{ mol Cu}_2\text{S}} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = 6.04 \text{ g SO}_2
\]

(d) \[\text{CaCO}_3(s) + \text{SiO}_2(s) \rightarrow \text{CaSiO}_3(s) + \text{CO}_2(g)\]

molar masses: \(\text{SiO}_2, 60.09 \text{ g}\); \(\text{CaSiO}_3, 116.17 \text{ g}\); \(\text{CO}_2, 44.01 \text{ g}\)

\[
15.0 \text{ g SiO}_2 \times \frac{1 \text{ mol SiO}_2}{60.09 \text{ g SiO}_2} = 0.2496 \text{ mol SiO}_2
\]

\[
0.2496 \text{ mol SiO}_2 \times \frac{1 \text{ mol CaSiO}_3}{1 \text{ mol SiO}_2} \times \frac{116.17 \text{ g CaSiO}_3}{1 \text{ mol CaSiO}_3} = 29.0 \text{ g CaSiO}_3
\]

\[
0.2496 \text{ mol SiO}_2 \times \frac{1 \text{ mol CO}_2}{1 \text{ mol SiO}_2} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 11.0 \text{ g CO}_2
\]
28. \[ 2S(s) + 3O_2(g) \rightarrow 2SO_3(g) \quad \text{SO}_3(g) + H_2O(l) \rightarrow H_2SO_4(l) \]

It would make things simpler if the first equation were expressed in terms of one mole of SO$_3$ since the second equation is expressed in terms of 1 mole of SO$_3$. To do this, divide the first equation by two:

\[ S(s) + \frac{3}{2}O_2(g) \rightarrow SO_3(g) \quad \text{SO}_3(g) + H_2O(l) \rightarrow H_2SO_4(l) \]

By doing this, we now have the simpler relationship that one mole of S will produce one mole of H$_2$SO$_4$.

molar masses: S, 32.07 g; H$_2$SO$_4$, 98.09 g

\[ 1.25 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.03898 \text{ mol S} \]

\[ 0.03988 \text{ mol S} \times \frac{98.09 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 3.82 \text{ g H}_2\text{SO}_4 \]

30. \[ 2\text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + H_2O(g) + \text{CO}_2(g) \]

molar masses: \text{NaHCO}_3, 84.01 g; \text{Na}_2\text{CO}_3, 106.0 g

\[ 1.52 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} = 0.01809 \text{ mol NaHCO}_3 \]

\[ 0.01809 \text{ mol NaHCO}_3 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{2 \text{ mol NaHCO}_3} = 0.009047 \text{ mol Na}_2\text{CO}_3 \]

\[ 0.009047 \text{ mol Na}_2\text{CO}_3 \times \frac{106.0 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} = 0.959 \text{ g Na}_2\text{CO}_3 \]

32. \[ C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2\text{CO}_2(g) \]

molar masses: \text{C}_6\text{H}_{12}\text{O}_6, 180.2 g; \text{C}_2\text{H}_5\text{OH}, 46.07 g

\[ 5.25 \text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.2 \text{ g C}_6\text{H}_{12}\text{O}_6} = 0.02913 \text{ mol C}_6\text{H}_{12}\text{O}_6 \]

\[ 0.02913 \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{2 \text{ mol C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = 0.5826 \text{ mol C}_2\text{H}_5\text{OH} \]

\[ 0.5286 \text{ mol C}_2\text{H}_5\text{OH} \times \frac{46.07 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 2.68 \text{ g ethyl alcohol} \]
34. $\text{NH}_4\text{Cl(s)} + \text{NaOH(s)} \rightarrow \text{NH}_3(g) + \text{NaCl(s)} + \text{H}_2\text{O(g)}$

molar masses: $\text{NH}_4\text{Cl}$, 53.49 g; $\text{NH}_3$, 17.03 g

$1.39 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} = 0.02599 \text{ mol NH}_4\text{Cl}$

$0.02599 \text{ mol NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_3}{1 \text{ mol NH}_4\text{Cl}} = 0.02599 \text{ mol NH}_3$

$0.02599 \text{ mol NH}_3 \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 0.443 \text{ g NH}_3$

36. $4\text{HgS(s)} + 4\text{CaO(s)} \rightarrow 4\text{Hg(l)} + 3\text{CaS(s)} + \text{CaSO}_4(s)$

molar masses: HgS, 232.7 g Hg, 200.6 g; 10.0 kg = 1.00 $\times$ 10$^4$ g

$1.00 \times 10^4 \text{ g HgS} \times \frac{1 \text{ mol HgS}}{232.7 \text{ g HgS}} = 42.97 \text{ mol HgS}$

$42.97 \text{ mol HgS} \times \frac{4 \text{ mol Hg}}{4 \text{ mol HgS}} = 42.97 \text{ mol Hg}$

$42.97 \text{ mol Hg} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} = 8.62 \times 10^3 \text{ g Hg} = 8.62 \text{ kg Hg}$

38. $\text{C}_{12}\text{H}_{22}\text{O}_{11}(s) \rightarrow 12\text{C(s)} + 11\text{H}_2\text{O(g)}$

molar masses: $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, 342.3 g; C, 12.01

$1.19 \text{ g } \text{C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{1 \text{ mol } \text{C}_{12}\text{H}_{22}\text{O}_{11}}{342.3 \text{ g } \text{C}_{12}\text{H}_{22}\text{O}_{11}} = 3.476 \times 10^{-3} \text{ mol } \text{C}_{12}\text{H}_{22}\text{O}_{11}$

$3.476 \times 10^{-3} \text{ mol } \text{C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{12 \text{ mol C}}{1 \text{ mol } \text{C}_{12}\text{H}_{22}\text{O}_{11}} = 0.04172 \text{ mol C}$

$0.04172 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.501 \text{ g C}$

40. $2\text{Mg(s)} + \text{O}_2(g) \rightarrow 2\text{MgO(s)}$

molar masses: Mg, 24.31 g; MgO, 40.31 g

$1.25 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 5.14 \times 10^{-2} \text{ mol Mg}$

$5.14 \times 10^{-2} \text{ mol Mg} \times \frac{2 \text{ mol MgO}}{2 \text{ mol Mg}} = 5.14 \times 10^{-2} \text{ mol MgO}$

$5.14 \times 10^{-2} \text{ mol MgO} \times \frac{40.31 \text{ g MgO}}{1 \text{ mol MgO}} = 2.07 \text{ g MgO}$
42. To determine the limiting reactant, first calculate the number of moles of each reactant present. Then determine how these numbers of moles correspond to the stoichiometric ratio indicated by the balanced chemical equation for the reaction.

44. A reactant is present *in excess* if there is more of that reactant present than is needed to combine with the limiting reactant for the process. By definition, the limiting reactant cannot be present in excess. An excess of any reactant does not affect the theoretical yield for a process; the theoretical yield is determined by the limiting reactant.

46.

a. \( S(s) + 2\text{H}_2\text{SO}_4(aq) \rightarrow 3\text{SO}_2(g) + 2\text{H}_2\text{O}(l) \)

Molar masses: S, 32.07 g; H\(_2\)SO\(_4\), 98.09 g; SO\(_2\), 64.07 g; H\(_2\)O, 18.02 g

\[
5.00 \text{ g S} \times \frac{1 \text{ mol}}{32.07 \text{ g}} = 0.1559 \text{ mol S}
\]

\[
5.00 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ mol}}{98.09 \text{ g}} = 0.05097 \text{ mol H}_2\text{SO}_4
\]

According to the balanced chemical equation, we would need twice as much sulfuric acid as sulfur for complete reaction of both reactants. We clearly have much less sulfuric acid present than sulfur; sulfuric acid is the limiting reactant. The calculation of the masses of products produced is based on the number of moles of the sulfuric acid.

\[
0.05097 \text{ mol H}_2\text{SO}_4 \times \frac{3 \text{ mol SO}_2}{2 \text{ mol H}_2\text{SO}_4} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = 4.90 \text{ g SO}_2
\]

\[
0.05097 \text{ mol H}_2\text{SO}_4 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2\text{SO}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.918 \text{ g H}_2\text{O}
\]

b. \( \text{MnO}_2(s) + 2\text{H}_2\text{SO}_4(aq) \rightarrow \text{Mn(SO}_4)_2 + 2\text{H}_2\text{O}(l) \)

Molar masses: \( \text{MnO}_2, 86.94 \text{ g}; \text{H}_2\text{SO}_4 98.09 \text{ g}; \text{Mn(SO}_4)_2, 247.1 \text{ g}; \text{H}_2\text{O}, 18.02 \text{ g} \)

\[
5.00 \text{ g MnO}_2 \times \frac{1 \text{ mol}}{86.94 \text{ g}} = 0.05751 \text{ mol MnO}_2
\]

\[
5.00 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ mol}}{98.09 \text{ g}} = 0.05097 \text{ mol H}_2\text{SO}_4
\]

According to the balanced chemical equation, we would need twice as much sulfuric acid as manganese(IV) oxide for complete reaction of both reactants. We do not have this much sulfuric acid, so sulfuric acid must be the limiting reactant. The amount of each product produced will be based on the sulfuric acid reacting completely.

\[
0.05097 \text{ mol H}_2\text{SO}_4 \times \frac{1 \text{ mol Mn(SO}_4)_2}{2 \text{ mol H}_2\text{SO}_4} \times \frac{247.1 \text{ g Mn(SO}_4)_2}{1 \text{ mol Mn(SO}_4)_2} = 6.30 \text{ g Mn(SO}_4)_2
\]

\[
0.05097 \text{ mol H}_2\text{SO}_4 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2\text{SO}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.918 \text{ g H}_2\text{O}
\]
c. \(2\text{H}_2\text{S}(g) + 3\text{O}_2(g) \rightarrow 2\text{SO}_2(g) + 2\text{H}_2\text{O}(l)\)

Molar masses: \(\text{H}_2\text{S}, 34.09 \text{ g} \); \(\text{O}_2, 32.00 \text{ g} \); \(\text{SO}_2, 64.07 \text{ g} \); \(\text{H}_2\text{O}, 18.02 \text{ g} \)

\[5.00 \text{ g } \text{H}_2\text{S} \times \frac{1 \text{ mol}}{34.09 \text{ g}} = 0.1467 \text{ mol H}_2\text{S}\]

\[5.00 \text{ g } \text{O}_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.1563 \text{ mol O}_2\]

According to the balanced equation, we would need 1.5 times as much \(\text{O}_2\) as \(\text{H}_2\text{S}\) for complete reaction of both reactants. We don’t have that much \(\text{O}_2\), so \(\text{O}_2\) must be the limiting reactant that will control the masses of each product produced.

\[0.1563 \text{ mol O}_2 \times \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = 6.67 \text{ g SO}_2\]

\[0.1563 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 1.88 \text{ g H}_2\text{O}\]

d. \(3\text{AgNO}_3(aq) + \text{Al}(s) \rightarrow 3\text{Ag}(s) + \text{Al(NO}_3)_3(aq)\)

Molar masses: \(\text{AgNO}_3, 169.9 \text{ g} \); \(\text{Al}, 26.98 \text{ g} \); \(\text{Ag}, 107.9 \text{ g} \); \(\text{Al(NO}_3)_3, 213.0 \text{ g} \)

\[5.00 \text{ g } \text{AgNO}_3 \times \frac{1 \text{ mol}}{169.9 \text{ g}} = 0.02943 \text{ mol AgNO}_3\]

\[5.00 \text{ g } \text{Al} \times \frac{1 \text{ mol}}{26.98 \text{ g}} = 0.1853 \text{ mol Al}\]

According to the balanced chemical equation, we would need three moles of \(\text{AgNO}_3\) for every mole of \(\text{Al}\) for complete reaction of both reactants. We in fact have fewer moles of \(\text{AgNO}_3\) than aluminum, so \(\text{AgNO}_3\) must be the limiting reactant. The amount of product produced is calculated from the number of moles of the limiting reactant present:

\[0.02943 \text{ mol AgNO}_3 \times \frac{3 \text{ mol Ag}}{3 \text{ mol AgNO}_3} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 3.18 \text{ g Ag}\]

\[0.02943 \text{ mol AgNO}_3 \times \frac{1 \text{ mol Al(NO}_3)_3}{3 \text{ mol AgNO}_3} \times \frac{213.0 \text{ g Al(NO}_3)_3}{1 \text{ mol Al(NO}_3)_3} = 2.09 \text{ g Al(NO}_3)_3\]

48.

a. \(\text{CS}_2(l) + 3\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{SO}_2(g)\)

Molar masses: \(\text{CS}_2, 76.15 \text{ g} \); \(\text{O}_2, 32.00 \text{ g} \); \(\text{CO}_2, 44.01 \text{ g} \)

\[1.00 \text{ g } \text{CS}_2 \times \frac{1 \text{ mol}}{76.15 \text{ g}} = 0.01313 \text{ mol CS}_2\]

\[1.00 \text{ g } \text{O}_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.03125 \text{ mol O}_2\]
From the balanced chemical equation, we would need three times as much oxygen as carbon disulfide for complete reaction of both reactants. We do not have this much oxygen, and so oxygen must be the limiting reactant.

\[0.03125 \text{ mol O}_2 \times \frac{1 \text{ mol CO}_2}{3 \text{ mol O}_2} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 0.458 \text{ g CO}_2\]

b. \(2\text{NH}_3(\text{g}) + \text{CO}_2(\text{g}) \rightarrow \text{CN}_2\text{H}_4\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l})\)

Molar masses: \(\text{NH}_3, 17.03 \text{ g} ; \text{CO}_2, 44.01 \text{ g} ; \text{H}_2\text{O}, 18.02 \text{ g}\)

\[1.00 \text{ g NH}_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 0.05872 \text{ mol NH}_3\]

\[1.00 \text{ g CO}_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 0.02272 \text{ mol CO}_2\]

The balanced chemical equation tells us that we would need twice as many moles of ammonia as carbon dioxide for complete reaction of both reactants. We have more than this amount of ammonia present, so the reaction will be limited by the amount of carbon dioxide present.

\[0.02272 \text{ mol CO}_2 \times \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol CO}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.409 \text{ g H}_2\text{O}\]

c. \(\text{H}_2(\text{g}) + \text{MnO}_2(\text{s}) \rightarrow \text{MnO}(\text{s}) + \text{H}_2\text{O}(\text{l})\)

Molar masses: \(\text{H}_2, 2.016 \text{ g} ; \text{MnO}_2, 86.94 \text{ g} ; \text{H}_2\text{O}, 18.02 \text{ g}\)

\[1.00 \text{ g H}_2 \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 0.496 \text{ mol H}_2\]

\[1.00 \text{ g MnO}_2 \times \frac{1 \text{ mol}}{86.94 \text{ g}} = 0.0115 \text{ mol MnO}_2\]

Because the coefficients of both reactants in the balanced chemical equation are the same, we would need equal amounts of both reactants for complete reaction. Therefore manganese(IV) oxide must be the limiting reactant and controls the amount of product obtained.

\[0.0115 \text{ mol MnO}_2 \times \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol MnO}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.207 \text{ g H}_2\text{O}\]

d. \(\text{I}_2(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{ICl}(\text{g})\)

Molar masses: \(\text{I}_2, 253.8 \text{ g} ; \text{Cl}_2, 70.90 \text{ g} ; \text{ICl}, 162.35 \text{ g}\)

\[1.00 \text{ g I}_2 \times \frac{1 \text{ mol}}{253.8 \text{ g}} = 0.00394 \text{ mol I}_2\]

\[1.00 \text{ g Cl}_2 \times \frac{1 \text{ mol}}{70.90 \text{ g}} = 0.0141 \text{ mol Cl}_2\]
From the balanced chemical equation, we would need equal amounts of I₂ and Cl₂ for complete reaction of both reactants. As we have much less iodine than chlorine, iodine must be the limiting reactant.

\[
0.00394 \text{ mol } I_2 \times \frac{2 \text{ mol } ICl}{1 \text{ mol } I_2} \times \frac{162.35 \text{ g } ICl}{1 \text{ mol } ICl} = 1.28 \text{ g } ICl
\]

50. a. \( \text{CO(g) + 2H}_2(\text{g}) \rightarrow \text{CH}_3\text{OH(l)} \)
   CO is the limiting reactant; 11.4 mg CH₃OH

b. \( 2\text{Al}(s) + 3I_2(s) \rightarrow 2\text{AlI}_3(s) \)
   I₂ is the limiting reactant; 10.7 mg AlI₃

c. \( \text{Ca(OH)}_2(aq) + 2\text{HBr}(aq) \rightarrow \text{CaBr}_2(aq) + 2\text{H}_2\text{O(l)} \)
   HBr is the limiting reactant; 12.4 mg CaBr₂; 2.23 mg H₂O

d. \( 2\text{Cr}(s) + 2\text{H}_3\text{PO}_4(aq) \rightarrow 2\text{CrPO}_4(s) + 3\text{H}_2(g) \)
   \( \text{H}_3\text{PO}_4 \) is the limiting reactant; 15.0 mg CrPO₄; 0.309 mg H₂

52. \( 2\text{NH}_3(g) + \text{CO}_2(g) \rightarrow \text{CN}_2\text{H}_4\text{O}(s) + \text{H}_2\text{O(l)} \)
   molar masses: NH₃, 17.03 g; CO₂, 44.01 g; CN₂H₄O, 60.06 g

\[
100. \text{ g NH}_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 5.872 \text{ mol NH}_3
\]

\[
100. \text{ g CO}_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 2.272 \text{ mol CO}_2
\]

CO₂ is the limiting reactant that determines the yield of product.

\[
2.272 \text{ mol CO}_2 \times \frac{1 \text{ mol CN}_2\text{H}_4\text{O}}{1 \text{ mol CO}_2} \times \frac{60.06 \text{ g CN}_2\text{H}_4\text{O}}{1 \text{ mol CN}_2\text{H}_4\text{O}} = 136 \text{ g CN}_2\text{H}_4\text{O}
\]

54. \( 4\text{Fe(s) + 3O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \)
   Molar masses: Fe, 55.85 g; Fe₂O₃, 159.7 g

\[
1.25 \text{ g Fe} \times \frac{1 \text{ mol}}{55.85 \text{ g}} = 0.0224 \text{ mol Fe present}
\]

Calculate how many mol of O₂ are required to react with this amount of Fe

\[
0.0224 \text{ mol Fe} \times \frac{3 \text{ mol O}_2}{4 \text{ mol Fe}} = 0.0168 \text{ mol O}_2
\]

Because we have more O₂ than this, Fe must be the limiting reactant.

\[
0.0224 \text{ mol Fe} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol Fe}} \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 1.79 \text{ g Fe}_2\text{O}_3
\]
56. \(2\text{CuSO}_4(\text{aq}) + 5\text{KI(}\text{aq}) \rightarrow 2\text{CuI(s)} + \text{KI}_3(\text{aq}) + 2\text{K}_2\text{SO}_4(\text{aq})\)

molar masses: \(\text{CuSO}_4, 159.6 \text{ g}\); \(\text{KI}, 166.0 \text{ g}\); \(\text{CuI}, 190.5 \text{ g}\); \(\text{KI}_3, 419.8 \text{ g}\); \(\text{K}_2\text{SO}_4, 174.3 \text{ g}\)

\[
0.525 \text{ g CuSO}_4 \times \frac{1 \text{ mol}}{159.6 \text{ g}} = 3.29 \times 10^{-3} \text{ mol CuSO}_4
\]

\[
2.00 \text{ g KI} \times \frac{1 \text{ mol}}{166.0 \text{ g}} = 0.0120 \text{ mol KI}
\]

To determine the limiting reactant, let’s calculate what amount of KI would be needed to react with the given amount of CuSO\(_4\) present.

\[
3.29 \times 10^{-3} \text{ mol CuSO}_4 \times \frac{5 \text{ mol KI}}{2 \text{ mol CuSO}_4} = 8.23 \times 10^{-3} \text{ mol KI}
\]

As we have more KI present than the amount required to react with the CuSO\(_4\) present, CuSO\(_4\) must be the limiting reactant that will control the amount of products produced.

\[
3.29 \times 10^{-3} \text{ mol CuSO}_4 \times \frac{2 \text{ mol CuI}}{2 \text{ mol CuSO}_4} \times \frac{190.5 \text{ g CuI}}{1 \text{ mol CuI}} = 0.627 \text{ g CuI}
\]

\[
3.29 \times 10^{-3} \text{ mol CuSO}_4 \times \frac{1 \text{ mol KI}_3}{2 \text{ mol CuSO}_4} \times \frac{419.8 \text{ g KI}_3}{1 \text{ mol KI}_3} = 0.691 \text{ g KI}_3
\]

\[
3.29 \times 10^{-3} \text{ mol CuSO}_4 \times \frac{2 \text{ mol K}_2\text{SO}_4}{2 \text{ mol CuSO}_4} \times \frac{174.3 \text{ g K}_2\text{SO}_4}{1 \text{ mol K}_2\text{SO}_4} = 0.573 \text{ g K}_2\text{SO}_4
\]

58. \(\text{SiO}_2(\text{s}) + 3\text{C(}\text{s}) \rightarrow 2\text{CO(}\text{g}) + \text{SiC(}\text{s})\)

molar masses: \(\text{SiO}_2, 60.09 \text{ g}\); \(\text{SiC, 40.10 g}\); \(1.0 \text{ kg} = 1.0 \times 10^3 \text{ g}\)

\[
1.0 \times 10^3 \text{ g SiO}_2 \times \frac{1 \text{ mol}}{60.09 \text{ g}} = 16.64 \text{ mol SiO}_2
\]

From the balanced chemical equation, if 16.64 mol of SiO\(_2\) were to react completely (an excess of carbon is present), then 16.64 mol of SiC should be produced (the coefficients of SiO\(_2\) and SiC are the same).

\[
16.64 \text{ mol SiC} \times \frac{40.01 \text{ g}}{1 \text{ mol}} = 6.7 \times 10^2 \text{ g SiC} = 0.67 \text{ kg SiC}
\]

60. If the reaction is performed in a solvent, the product may have substantial solubility in the solvent and the reaction may come to equilibrium before the full yield of product is achieved (See Chapter 17.). Loss of product may occur through operator error.
62. \( 2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g) \)

molar masses: \( \text{HgO}, 216.6 \text{ g} \); \( \text{Hg}, 200.6 \text{ g} \)

\[
1.25 \text{ g HgO} \times \frac{1 \text{ mol}}{216.6 \text{ g}} = 0.005771 \text{ mol HgO}
\]

\[
0.005771 \text{ mol HgO} \times \frac{2 \text{ mol Hg}}{2 \text{ mol HgO}} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} = 1.16 \text{ g (theoretical yield)}
\]

\[
\% \text{ yield} = \frac{1.09 \text{ g actual}}{1.16 \text{ g theoretical}} \times 100 = 94.0\% \text{ of theory}
\]

64. \( 2\text{LiOH}(s) + \text{CO}_2(g) \rightarrow \text{Li}_2\text{CO}_3(s) + \text{H}_2\text{O}(g) \)

molar masses: \( \text{LiOH}, 23.95 \text{ g} \); \( \text{CO}_2, 44.01 \text{ g} \)

\[
155 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 142 \text{ g CO}_2
\]

As the cartridge has only absorbed 102 g CO2 out of a total capacity of 142 g CO2, the cartridge has absorbed

\[
\frac{102 \text{ g}}{142 \text{ g}} \times 100 = 71.8\% \text{ of its capacity.}
\]

66. \( \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) \)

molar masses: \( \text{SO}_4^{2-}, 96.07 \text{ g} \); \( \text{BaCl}_2, 208.2 \text{ g} \); \( \text{BaSO}_4, 233.4 \text{ g} \)

\[
1.12 \text{ g SO}_4^{2-} \times \frac{1 \text{ mol SO}_4^{2-}}{96.07 \text{ g}} = 0.01166 \text{ mol SO}_4^{2-}
\]

\[
5.02 \text{ g BaCl}_2 \times \frac{1 \text{ mol BaCl}_2}{208.2 \text{ g}} = 0.02411 \text{ mol BaCl}_2 = 0.02411 \text{ mol Ba}^{2+}
\]

\( \text{SO}_4^{2-} \) is the limiting reactant.

\[
0.01166 \text{ mol SO}_4^{2-} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol SO}_4^{2-}} \times \frac{233.4 \text{ g BaSO}_4}{1 \text{ mol BaSO}_4} = 2.72 \text{ g BaSO}_4
\]

\[
\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{2.02 \text{ g}}{2.72 \text{ g}} \times 100 = 74.3\%
\]
68. \( \text{NaCl}(aq) + \text{NH}_3(aq) + H_2O(l) + \text{CO}_2(s) \rightarrow \text{NH}_4\text{Cl}(aq) + \text{NaHCO}_3(s) \)

Molar masses: \( \text{NH}_3, 17.03 \text{ g} \); \( \text{CO}_2, 44.01 \text{ g} \); \( \text{NaHCO}_3, 84.01 \text{ g} \)

\[
10.0 \text{ g} \ \text{NH}_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 0.5872 \text{ mol} \ \text{NH}_3
\]

\[
15.0 \text{ g} \ \text{CO}_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 0.3408 \text{ mol} \ \text{CO}_2
\]

CO\(_2\) is the limiting reactant.

\[
0.3408 \text{ mol} \ \text{CO}_2 \times \frac{1 \text{ mol} \ \text{NaHCO}_3}{1 \text{ mol} \ \text{CO}_2} = 0.3408 \text{ mol} \ \text{NaHCO}_3
\]

\[
0.3408 \text{ mol} \ \text{NaHCO}_3 \times \frac{84.01 \text{ g}}{1 \text{ mol}} = 28.6 \text{ g} \ \text{NaHCO}_3
\]

70. \( \text{C}_6\text{H}_12\text{O}_6(s) + 6\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \)

Molar masses: glucose, 180.2 g; \( \text{CO}_2, 44.01 \text{ g} \)

\[
1.00 \text{ g} \ \text{glucose} \times \frac{1 \text{ mol}}{54.09 \times 10^{-3} \text{ mol}} = 5.549 \times 10^{-3} \text{ mol} \ \text{glucose}
\]

\[
5.549 \times 10^{-3} \text{ mol} \ \text{glucose} \times \frac{6 \text{ mol} \ \text{CO}_2}{1 \text{ mol} \ \text{glucose}} = 3.33 \times 10^{-2} \text{ mol} \ \text{CO}_2
\]

\[
3.33 \times 10^{-2} \text{ mol} \ \text{CO}_2 \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 1.47 \text{ g} \ \text{CO}_2
\]

72. \( \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) \)

Millimolar ionic masses: \( \text{Ba}^{2+}, 137.3 \text{ mg} \); \( \text{SO}_4^{2-}, 96.07 \text{ mg} \); \( \text{BaCl}_2, 208.2 \text{ mg} \)

\[
150 \text{ mg} \ \text{SO}_4^{2-} \times \frac{1 \text{ mmol}}{96.07 \text{ mg}} = 1.56 \text{ millimol} \ \text{SO}_4^{2-}
\]

As barium ion and sulfate ion react on a 1:1 stoichiometric basis, then 1.56 millimol of barium ion is needed, which corresponds to 1.56 millimol of \( \text{BaCl}_2 \).

\[
1.56 \text{ millimol} \ \text{BaCl}_2 \times \frac{208.2 \text{ mg}}{1 \text{ mmol}} = 325 \text{ milligrams} \ \text{BaCl}_2 \ \text{needed}
\]

74.

a. \( \text{UO}_2(s) + 4\text{HF}(aq) \rightarrow \text{UF}_4(aq) + 2\text{H}_2\text{O}(l) \)

One molecule (formula unit) of uranium(IV) oxide will combine with four molecules of hydrofluoric acid, producing one uranium(IV) fluoride molecule and two water molecules. One mole of uranium(IV) oxide will combine with four moles of hydrofluoric acid to produce one mole of uranium(IV) fluoride and two moles of water.
b. \(2\text{NaC}_2\text{H}_3\text{O}_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{HC}_2\text{H}_3\text{O}_2(aq)\)

Two molecules (formula units) of sodium acetate react exactly with one molecule of sulfuric acid, producing one molecule (formula unit) of sodium sulfate and two molecules of acetic acid. Two moles of sodium acetate will combine with one mole of sulfuric acid, producing one mole of sodium sulfate and two moles of acetic acid.

c. \(\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)\)

One magnesium atom will react with two hydrochloric acid molecules (formula units) to produce one molecule (formula unit) of magnesium chloride and one molecule of hydrogen gas. One mole of magnesium will combine with two moles of hydrochloric acid, producing one mole of magnesium chloride and one mole of gaseous hydrogen.

d. \(\text{B}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{B(OH)}_3(aq)\)

One molecule of diboron trioxide will react exactly with three molecules of water, producing two molecules of boron trihydroxide (boric acid). One mole of diboron trioxide will combine with three moles of water to produce two moles of boron trihydroxide (boric acid).

76. 

For \(\text{O}_2\): \(\frac{5 \text{ mol O}_2}{1 \text{ mol C}_3\text{H}_8}\)

For \(\text{CO}_2\): \(\frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8}\)

For \(\text{H}_2\text{O}\): \(\frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_3\text{H}_8}\)

78.

a. \(\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)\)

molar mass of \(\text{NH}_3 = 17.01 \text{ g}\)

\[1.00 \text{ g NH}_3 \times \frac{1 \text{ mol}}{17.01 \text{ g}} = 0.0588 \text{ mol NH}_3\]

\[0.0588 \text{ mol NH}_3 \times \frac{1 \text{ mol NH}_4\text{Cl}}{1 \text{ mol NH}_3} = 0.0588 \text{ mol NH}_4\text{Cl}\]

b. \(\text{CaO}(s) + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s)\)

molar mass \(\text{CaO} = 56.08 \text{ g}\)

\[1.00 \text{ g CaO} \times \frac{1 \text{ mol}}{56.08 \text{ g}} = 0.0178 \text{ mol CaO}\]

\[0.0178 \text{ mol CaO} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CaO}} = 0.0178 \text{ mol CaCO}_3\]

c. \(4\text{Na}(s) + \text{O}_2(g) \rightarrow 2\text{Na}_2\text{O}(s)\)

molar mass \(\text{Na} = 22.99 \text{ g}\)

\[1.00 \text{ g Na} \times \frac{1 \text{ mol}}{22.99 \text{ g}} = 0.0435 \text{ mol Na}\]

\[0.0435 \text{ mol Na} \times \frac{2 \text{ mol Na}_2\text{O}}{4 \text{ mol Na}} = 0.0217 \text{ mol Na}_2\text{O}\]
100  Chapter 9: Chemical Quantities

d. \[2P(s) + 3Cl_2(g) \rightarrow 2PCl_3(l)\]

molar mass \(P = 30.97 \text{ g}\)

\[1.00 \text{ g } P \times \frac{1 \text{ mol}}{30.97 \text{ g }} = 0.0323 \text{ mol } P\]

\[0.0323 \text{ mol } P \times \frac{2 \text{ mol } PCl_3}{2 \text{ mol } P} = 0.0323 \text{ mol } PCl_3\]

80. a. molar mass \(HNO_3 = 63.0 \text{ g}\)

\[5.0 \text{ mol } HNO_3 \times \frac{63.0 \text{ g}}{1 \text{ mol }} = 3.2 \times 10^2 \text{ g } HNO_3\]

b. molar mass \(Hg = 200.6 \text{ g}\)

\[0.000305 \text{ mol } Hg \times \frac{200.6 \text{ g}}{1 \text{ mol }} = 0.0612 \text{ g } Hg\]

c. molar mass \(K_2CrO_4 = 194.2 \text{ g}\)

\[2.31 \times 10^{-5} \text{ mol } K_2CrO_4 \times \frac{194.2 \text{ g}}{1 \text{ mol }} = 4.49 \times 10^{-3} \text{ g } K_2CrO_4\]

d. molar mass \(AlCl_3 = 133.3 \text{ g}\)

\[10.5 \text{ mol } AlCl_3 \times \frac{133.3 \text{ g}}{1 \text{ mol }} = 1.40 \times 10^3 \text{ g } AlCl_3\]

e. molar mass \(SF_6 = 146.1 \text{ g}\)

\[4.9 \times 10^4 \text{ mol } SF_6 \times \frac{146.1 \text{ g}}{1 \text{ mol }} = 7.2 \times 10^6 \text{ g } SF_6\]

f. molar mass \(NH_3 = 17.01 \text{ g}\)

\[125 \text{ mol } NH_3 \times \frac{17.01 \text{ g}}{1 \text{ mol }} = 2.13 \times 10^3 \text{ g } NH_3\]

g. molar mass \(Na_2O_2 = 77.98 \text{ g}\)

\[0.01205 \text{ mol } Na_2O_2 \times \frac{77.98 \text{ g}}{1 \text{ mol }} = 0.9397 \text{ g } Na_2O_2\]
82. \[ 2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \]

molar masses: \( \text{SO}_2, 64.07 \text{ g} \); \( \text{SO}_3, 80.07 \text{ g} \); 150 kg = \( 1.5 \times 10^5 \text{ g} \)

\[ 1.5 \times 10^5 \text{ g SO}_2 \times \frac{1 \text{ mol}}{64.07 \text{ g}} = 2.34 \times 10^3 \text{ mol SO}_2 \]

\[ 2.34 \times 10^3 \text{ mol SO}_2 \times \frac{2 \text{ mol SO}_3}{2 \text{ mol SO}_2} = 2.34 \times 10^3 \text{ mol SO}_3 \]

\[ 2.34 \times 10^3 \text{ mol SO}_3 \times \frac{80.07 \text{ g}}{1 \text{ mol}} = 1.9 \times 10^5 \text{ g SO}_3 = 1.9 \times 10^2 \text{ kg SO}_3 \]

84. \[ 2\text{Na}_2\text{O}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow 4\text{NaOH}(aq) + \text{O}_2(g) \]

molar masses: \( \text{Na}_2\text{O}_2, 77.98 \text{ g} \); \( \text{O}_2, 32.00 \text{ g} \)

\[ 3.25 \text{ g Na}_2\text{O}_2 \times \frac{1 \text{ mol}}{77.98 \text{ g}} = 0.0417 \text{ mol Na}_2\text{O}_2 \]

\[ 0.0417 \text{ mol Na}_2\text{O}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol Na}_2\text{O}_2} = 0.0209 \text{ mol O}_2 \]

\[ 0.0209 \text{ mol O}_2 \times \frac{32.00 \text{ g}}{1 \text{ mol}} = 0.669 \text{ g O}_2 \]

86. \[ \text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \]

molar masses: \( \text{Zn}, 65.38 \text{ g} \); \( \text{H}_2, 2.016 \text{ g} \)

\[ 2.50 \text{ g Zn} \times \frac{1 \text{ mol}}{65.38 \text{ g}} = 0.03824 \text{ mol Zn} \]

\[ 0.03824 \text{ mol Zn} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} = 0.03824 \text{ mol H}_2 \]

\[ 0.03824 \text{ mol H}_2 \times \frac{2.016 \text{ g}}{1 \text{ mol}} = 0.0771 \text{ g H}_2 \]

88.

a. \[ 2\text{Na}(s) + \text{Br}_2(l) \rightarrow 2\text{NaBr}(s) \]

molar masses: \( \text{Na}, 22.99 \text{ g} \); \( \text{Br}_2, 159.8 \text{ g} \); \( \text{NaBr}, 102.9 \text{ g} \)

\[ 5.0 \text{ g Na} \times \frac{1 \text{ mol}}{22.99 \text{ g}} = 0.2175 \text{ mol Na} \]

\[ 5.0 \text{ g Br}_2 \times \frac{1 \text{ mol}}{159.8 \text{ g}} = 0.03129 \text{ mol Br}_2 \]
Intuitively, we would suspect that Br₂ is the limiting reactant because there is much less Br₂ than Na on a mole basis. To prove that Br₂ is the limiting reactant, the following calculation is needed:

\[
0.03129 \text{ mol Br}_2 \times \frac{2 \text{ mol Na}}{1 \text{ mol Br}_2} = 0.06258 \text{ mol Na.}
\]

Clearly there is more Na than this present, so Br₂ limits the reaction extent and the amount of NaBr formed.

\[
0.03129 \text{ mol Br}_2 \times \frac{2 \text{ mol NaBr}}{1 \text{ mol Br}_2} = 0.06258 \text{ mol NaBr}
\]

\[
0.06258 \text{ mol NaBr} \times \frac{102.9 \text{ g}}{1 \text{ mol}} = 6.4 \text{ g NaBr}
\]

b. Zn(s) + CuSO₄(aq) → ZnSO₄(aq) + Cu(s)

molar masses: Zn, 65.38 g; Cu, 63.55 g; ZnSO₄, 161.5 g; CuSO₄, 159.6 g

\[
5.0 \text{ g Zn} \times \frac{1 \text{ mol}}{65.38 \text{ g}} = 0.07648 \text{ mol Zn}
\]

\[
5.0 \text{ g CuSO}_4 \times \frac{1 \text{ mol}}{159.6 \text{ g}} = 0.03132 \text{ mol CuSO}_4
\]

As the coefficients of Zn and CuSO₄ are the same in the balanced chemical equation, an equal number of moles of Zn and CuSO₄ would be needed for complete reaction. There is less CuSO₄ present, so CuSO₄ must be the limiting reactant.

\[
0.03132 \text{ mol CuSO}_4 \times \frac{1 \text{ mol ZnSO}_4}{1 \text{ mol CuSO}_4} = 0.03132 \text{ mol ZnSO}_4
\]

\[
0.03132 \text{ mol ZnSO}_4 \times \frac{161.5 \text{ g}}{1 \text{ mol}} = 5.1 \text{ g ZnSO}_4
\]

\[
0.03132 \text{ mol CuSO}_4 \times \frac{1 \text{ mol Cu}}{1 \text{ mol CuSO}_4} = 0.03132 \text{ mol Cu}
\]

\[
0.03132 \text{ mol Cu} \times \frac{63.55 \text{ g}}{1 \text{ mol}} = 2.0 \text{ g Cu}
\]

c. NH₄Cl(aq) + NaOH(aq) → NH₃(g) + H₂O(l) + NaCl(aq)

molar masses: NH₄Cl, 53.49 g; NaOH, 40.00 g; NH₃, 17.03 g; H₂O, 18.02 g; NaCl, 58.44 g

\[
5.0 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol}}{53.49 \text{ g}} = 0.09348 \text{ mol NH}_4\text{Cl}
\]

\[
5.0 \text{ g NaOH} \times \frac{1 \text{ mol}}{40.00 \text{ g}} = 0.1250 \text{ mol NaOH}
\]

As the coefficients of NH₄Cl and NaOH are both one in the balanced chemical equation for the reaction, an equal number of moles of NH₄Cl and NaOH would be needed for complete reaction. There is less NH₄Cl present, so NH₄Cl must be the limiting reactant.
As the coefficients of the products in the balanced chemical equation are also all one, if 0.09348 mol of NH₄Cl (the limiting reactant) reacts completely, then 0.09348 mol of each product will be formed.

\[
0.09348 \text{ mol NH}_3 \times \frac{17.03 \text{ g}}{1 \text{ mol}} = 1.6 \text{ g NH}_3
\]

\[
0.09348 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 1.7 \text{ g H}_2\text{O}
\]

\[
0.09348 \text{ mol NaCl} \times \frac{58.44 \text{ g}}{1 \text{ mol}} = 5.5 \text{ g NaCl}
\]

d. \(\text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g)\)
molar masses: Fe₂O₃, 159.7 g; CO, 28.01 g; Fe, 55.85 g; CO₂, 44.01 g

\[
5.0 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol}}{159.7 \text{ g}} = 0.03131 \text{ mol Fe}_2\text{O}_3
\]

\[
5.0 \text{ g CO} \times \frac{1 \text{ mol}}{28.01 \text{ g}} = 0.1785 \text{ mol CO}
\]

Because there is considerably less Fe₂O₃ than CO on a mole basis, let’s see if Fe₂O₃ is the limiting reactant.

\[
0.03131 \text{ mol Fe}_2\text{O}_3 \times \frac{3 \text{ mol CO}}{1 \text{ mol Fe}_2\text{O}_3} = 0.09393 \text{ mol CO}
\]

There is 0.1785 mol of CO present, but we have determined that only 0.09393 mol CO would be needed to react with all the Fe₂O₃ present, so Fe₂O₃ must be the limiting reactant. CO is present in excess.

\[
0.03131 \text{ mol Fe}_2\text{O}_3 \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 3.5 \text{ g Fe}
\]

\[
0.03131 \text{ mol Fe}_2\text{O}_3 \times \frac{3 \text{ mol CO}_2}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 4.1 \text{ g CO}_2
\]

90. \(\text{N}_2\text{H}_4(l) + \text{O}_2(g) \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(g)\)
molar masses: N₂H₄, 32.05 g; O₂, 32.00 g; N₂, 28.02 g; H₂O, 18.02 g

\[
20.0 \text{ g N}_2\text{H}_4 \times \frac{1 \text{ mol}}{32.05 \text{ g}} = 0.624 \text{ mol N}_2\text{H}_4
\]

\[
20.0 \text{ g O}_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.625 \text{ mol O}_2
\]
The two reactants are present in nearly the required ratio for complete reaction (due to the 1:1 stoichiometry of the reaction and the very similar molar masses of the substances). We will consider N₂H₄ as the limiting reactant in the following calculations.

\[
0.624 \text{ mol N}_2\text{H}_4 \times \frac{1 \text{ mol N}_2}{1 \text{ mol N}_2\text{H}_4} \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 17.5 \text{ g } N_2
\]

\[
0.624 \text{ mol N}_2\text{H}_4 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol N}_2\text{H}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 22.5 \text{ g } H_2O
\]

92. \[\frac{12.5 \text{ g theory}}{100 \text{ g theory}} \times \frac{40 \text{ g actual}}{100 \text{ g theory}} = 5.0 \text{ g}\]
Cumulative Review Chapters 8 & 9

ANSWERS

2. On a microscopic basis, one mole of a substance represents Avogadro’s number (6.022 × 10^23) of individual units (atoms or molecules) of the substance. On a macroscopic, more practical basis, one mole of a substance represents the amount of substance present when the molar mass of the substance in grams is taken (for example 12.01 g of carbon will be one mole of carbon). Chemists have chosen these definitions so that there will be a simple relationship between measurable amounts of substances (grams) and the actual number of atoms or molecules present and so that the number of particles present in samples of different substances can be compared easily. For example, it is known that carbon and oxygen react by the reaction

\[ \text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}). \]

Chemists understand this equation to mean that one carbon atom reacts with one oxygen molecule to produce one molecule of carbon dioxide and also that one mole (12.01 g) of carbon will react with one mole (32.00 g) of oxygen to produce one mole (44.01 g) of carbon dioxide.

4. The molar mass of a compound is the mass in grams of one mole of the compound (6.022 × 10^23 molecules of the compound), and is calculated by summing the average atomic masses of all the atoms present in a molecule of the compound. For example, a molecule of the compound H₃PO₄ contains three hydrogen atoms, one phosphorus atom, and four oxygen atoms. The molar mass is obtained by adding up the average atomic masses of these atoms; molar mass H₃PO₄ = 3(1.008 g) + 1(30.97 g) + 4(16.00 g) = 97.99 g

6. The empirical formula of a compound represents the lowest ratio of the relative number of atoms of each type present in a molecule of the compound whereas the molecular formula represents the actual number of atoms of each type present in a real molecule of the compound. For example, both acetylene (molecular formula C₂H₂) and benzene (molecular formula C₆H₆) have the same relative number of carbon and hydrogen atoms (one hydrogen for each carbon atom) and so have the same empirical formula (CH). Once the empirical formula of a compound has been determined, it is also necessary to determine the molar mass of the compound before the actual molecular formula can be calculated. As real molecules cannot contain fractional parts of atoms, the molecular formula is always a whole number multiple of the empirical formula. For the examples above, the molecular formula of acetylene is twice the empirical formula, and the molecular formula of benzene is six times the empirical formula (both factors are integers).

8. In question 7, we chose to calculate the percentage composition of phosphoric acid, H₃PO₄; 3.086% H, 31.60% P, 65.31% O. We could convert this percentage composition data into “experimental” data by first choosing a mass of sample to be “analyzed” and then calculating what mass of each element is present in this size sample using the percentage of each element. For example, suppose we choose our sample to have a mass of 2.417 g. Then the masses of H, P, and O present in this sample would be given by the following:

\[ \text{g H} = (2.417 \text{ g sample}) \times \frac{3.086 \text{ g H}}{100.0 \text{ g sample}} = 0.07459 \text{ g H} \]
\[ \text{g P} = (2.417 \text{ g sample}) \times \frac{31.60 \text{ g P}}{100.0 \text{ g sample}} = 0.7638 \text{ g P} \]

\[ \text{g O} = (2.417 \text{ g sample}) \times \frac{65.31 \text{ g O}}{100.0 \text{ g sample}} = 1.579 \text{ g O} \]

Note that \((0.07459 \text{ g} + 0.7638 \text{ g} + 1.579 \text{ g}) = 2.41739 = 2.417 \text{ g}.\)

So our new problem could be worded as follows, “A 2.417 g sample of a compound has been analyzed and was found to contain 0.07459 g H, 0.7638 g of P, and 1.579 g of oxygen. Calculate the empirical formula of the compound.”

\[ \text{mol H} = (0.07459 \text{ g H}) \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.07400 \text{ mol H} \]

\[ \text{mol P} = (0.7638 \text{ g P}) \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.02466 \text{ mol P} \]

\[ \text{mol O} = (1.579 \text{ g O}) \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.09869 \text{ mol O} \]

Dividing each of these numbers of moles by the smallest number of moles (0.02466 mol P) gives the following:

\[ \frac{0.07400 \text{ mol H}}{0.02466} = 3.001 \text{ mol H} \]

\[ \frac{0.02466 \text{ mol P}}{0.02466} = 1.000 \text{ mol P} \]

\[ \frac{0.09869 \text{ mol O}}{0.02466} = 4.002 \text{ mol O} \]

The empirical formula is (not surprisingly) just \(\text{H}_3\text{PO}_4\)!

10. The mole ratios for a reaction are based on the coefficients of the balanced chemical equation for the reaction. These coefficients show in what proportions molecules (or moles of molecules) combine. For a given amount of \(\text{C}_2\text{H}_5\text{OH}\), the following mole ratios could be constructed and would enable you to calculate the number of moles of each product, or of the second reactant, that would be involved.

\[ \text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g) \]

To calculate mol \(\text{CO}_2\) produced:

\[ \frac{2 \text{ mol CO}_2}{1 \text{ mol C}_2\text{H}_5\text{OH}} \]

To calculate mol \(\text{H}_2\text{O}\) produced:

\[ \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \]

To calculate mol \(\text{O}_2\) required:

\[ \frac{3 \text{ mol O}_2}{1 \text{ mol C}_2\text{H}_5\text{OH}} \]
We could then calculate the numbers of moles of the other substances if 0.65 mol of C₂H₅OH were to be combusted as follows:

\[
\text{mol CO}_2 \text{ produced} = (0.65 \text{ mol C}_2\text{H}_5\text{OH}) \times \frac{2 \text{ mol CO}_2}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 1.3 \text{ mol CO}_2
\]

\[
\text{mol H}_2\text{O produced} = (0.65 \text{ mol C}_2\text{H}_5\text{OH}) \times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 1.95 = 2.0 \text{ mol H}_2\text{O}
\]

\[
\text{mol O}_2 \text{ required} = (0.65 \text{ mol C}_2\text{H}_5\text{OH}) \times \frac{3 \text{ mol O}_2}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 1.95 = 2.0 \text{ mol O}_2
\]

12. Although we can calculate specifically the exact amounts of each reactant needed for a chemical reaction, oftentimes reaction mixtures are prepared using more or less arbitrary amounts of the reagents. However, regardless of how much of each reagent may be used for a reaction, the substances still react stoichiometrically according to the mole ratios derived from the balanced chemical equation for the reaction. When arbitrary amounts of reactants are used, there will be one reactant that stoichiometrically is present in the least amount. This substance is called the **limiting reactant** for the experiment. It is the limiting reactant that controls how much product is formed regardless of how much of the other reactants are present. The limiting reactant limits the amount of product that can form in the experiment because once the limiting reactant has reacted completely, the reaction must stop. We say that the other reactants in the experiment are present in excess, which means that a portion of those reactants will still be present unchanged after the reaction has ended and the limiting reactant has been used up completely.

14. The **theoretical yield** for an experiment is the mass of product calculated based on the limiting reactant for the experiment being completely consumed. The **actual yield** for an experiment is the mass of product actually collected by the experimenter. Obviously, any experiment is restricted by the skills of the experimenter and by the inherent limitations of the experimental method being used. For these reasons, the actual yield is often less than the theoretical yield (most scientific writers report the actual or percentage yield for their experiments as an indication of the usefulness of their experiments). Although one would expect that the actual yield should never be more than the theoretical yield, in real experiments sometimes this happens. However, an actual yield greater than a theoretical yield is usually taken to mean that something is wrong in either the experiment (for example, impurities may be present or the reaction may not occur as envisioned) or in the calculations.

16.

a. element

b. \( \% \text{ C} = \frac{12.01 \text{ g}}{76.51 \text{ g}} \times 100 = 15.77\% \text{ C} \)

c. \( \% \text{ Al} = \frac{26.98 \text{ g}}{213.0 \text{ g}} \times 100 = 12.67\% \text{ Al} \)

d. \( \% \text{ C} = \frac{3(12.01 \text{ g})}{44.09 \text{ g}} \times 100 = 81.72\% \text{ C} \)

e. element
f. \( \% \text{Ca} = \frac{40.08 \text{ g}}{100.09 \text{ g}} \times 100 = 40.04\% \text{Ca} \)

g. \( \% \text{C} = \frac{12.01 \text{ g}}{28.01 \text{ g}} \times 100 = 42.88\% \text{C} \)

h. \( \% \text{C} = \frac{12.01 \text{ g}}{44.01 \text{ g}} \times 100 = 27.29\% \text{C} \)

i. \( \% \text{Ca} = \frac{40.08 \text{ g} \text{Ca}}{110.98 \text{ g}} \times 100 = 36.11\% \text{Ca} \)

j. element

18.

a. molar masses: \( \text{SiC}, 40.10 \text{ g}; \text{SiCl}_4, 169.9 \text{ g} \)

\[ 12.5 \text{ g SiC} \times \frac{1 \text{ mol}}{40.10 \text{ g}} = 0.3117 \text{ mol SiC} \]

for \( \text{SiCl}_4 \): \( 0.3117 \text{ mol SiC} \times \frac{1 \text{ mol SiCl}_4}{1 \text{ mol SiC}} \times \frac{169.9 \text{ g SiCl}_4}{1 \text{ mol SiCl}_4} = 53.0 \text{ g SiCl}_4 \)

for \( \text{C} \): \( 0.3117 \text{ mol SiC} \times \frac{1 \text{ mol C}}{1 \text{ mol SiC}} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 3.75 \text{ g C} \)

b. molar masses: \( \text{Li}_2\text{O}, 29.88 \text{ g}; \text{LiOH}, 23.95 \text{ g} \)

\[ 12.5 \text{ g Li}_2\text{O} \times \frac{1 \text{ mol}}{29.88 \text{ g}} = 0.4183 \text{ mol Li}_2\text{O} \]

\[ 0.4183 \text{ mol Li}_2\text{O} \times \frac{2 \text{ mol LiOH}}{1 \text{ mol Li}_2\text{O}} \times \frac{23.95 \text{ g LiOH}}{1 \text{ mol LiOH}} = 20.0 \text{ g LiOH} \]

c. molar masses: \( \text{Na}_2\text{O}_2, 77.98 \text{ g}; \text{NaOH}, 40.00 \text{ g}; \text{O}_2, 32.00 \text{ g} \)

\[ 12.5 \times \frac{1 \text{ mol}}{77.98 \text{ g}} = 0.1603 \text{ mol Na}_2\text{O}_2 \]

for \( \text{NaOH} \): \( 0.1603 \text{ mol Na}_2\text{O}_2 \times \frac{4 \text{ mol NaOH}}{2 \text{ mol Na}_2\text{O}_2} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 12.8 \text{ g NaOH} \)

for \( \text{O}_2 \): \( 0.1603 \text{ mol Na}_2\text{O}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol Na}_2\text{O}_2} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 2.56 \text{ g O}_2 \)
d. molar masses: SnO$_2$, 150.7 g; Sn, 118.7 g; H$_2$O, 18.02 g

\[
12.5 \text{ g SnO}_2 \times \frac{1 \text{ mol}}{150.7 \text{ g}} = 0.08295 \text{ mol SnO}_2
\]

for Sn: \(0.08295 \text{ mol SnO}_2 \times \frac{1 \text{ mol Sn}}{1 \text{ mol SnO}_2} \times \frac{118.7 \text{ g Sn}}{1 \text{ mol Sn}} = 9.84 \text{ g Sn}\)

for H$_2$O: \(0.08295 \text{ mol SnO}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol SnO}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 2.99 \text{ g H}_2\text{O}\)

20. molar masses: C, 12.01 g; CO, 28.01 g; CO$_2$, 44.01 g

\[
5.00 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.416 \text{ mol C}
\]

for CO: \(0.416 \text{ mol C} \times \frac{2 \text{ mol CO}}{2 \text{ mol C}} \times \frac{28.01 \text{ g CO}}{1 \text{ mol CO}} = 11.7 \text{ g CO}\)

for CO$_2$: \(0.416 \text{ mol C} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 18.3 \text{ g CO}_2\)
CHAPTER 10

Energy

CHAPTER ANSWERS

2. potential

4. The total energy of the universe is constant. Energy cannot be created or destroyed, but can only be converted from one form to another.

6. Ball A initially possesses potential energy by virtue of its position at the top of the hill. As Ball A rolls down the hill, its potential energy is converted to kinetic energy and frictional (heat) energy. When Ball A reaches the bottom of the hill and hits Ball B, it transfers its kinetic energy to Ball B. Ball A then has only the potential energy corresponding to its new position.

8. The hot tea is at a higher temperature, which means the particles in the hot tea have higher average kinetic energies. When the tea spills on the skin, energy flows from the hot tea to the skin, until the tea and skin are at the same temperature. This sudden inflow of energy causes the burn.

10. Temperature is the concept by which we express the thermal energy contained in a sample. We cannot measure the motions of the particles/kinetic energy in a sample of matter directly. We know, however, that if two objects are at different temperatures, the one with the higher temperature has molecules that have higher average kinetic energies than those in the object at the lower temperature.

12. When the chemical system evolves energy, the energy evolved from the reacting chemicals is transferred to the surroundings.

14. exactly equal to

16. internal

18. losing

20. gaining

22.

a. \[ \frac{1 \text{ J}}{4.184 \text{ cal}} \]

b. \[ \frac{4.184 \text{ cal}}{1 \text{ J}} \]

c. \[ \frac{1 \text{ kcal}}{1000 \text{ cal}} \]

d. \[ \frac{1000 \text{ J}}{1 \text{ kJ}} \]

24. 6540 J = 6.54 kJ for ten times more water
26. 
   a. \(7518 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 7.518 \text{ kcal}\)
   b. \(7.518 \times 10^3 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 7.518 \text{ kcal}\)
   c. \(1 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 0.001 \text{ kcal}\)
   d. \(655,200 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 655.2 \text{ kcal}\)

28. 
   a. \(243,000 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 243 \text{ kJ}\) to 3 significant figures
   b. \(4.184 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.004184 \text{ kJ}\)
   c. \(0.251 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.000251 \text{ kJ}\)
   d. \(450.3 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.4503 \text{ kJ}\)

30. 
   a. \(89.74 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 21.45 \text{ kcal}\)
   b. \(1.756 \times 10^4 \text{ J} \times \frac{1 \text{ kcal}}{1000 \text{ J}} = 17.56 \text{ kJ}\)
   c. \(1.756 \times 10^4 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 4.20 \text{ kcal}\)
   d. \(1.00 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 239 \text{ cal}\)

32. \[Q = s \times m \times \Delta T\]

   Temperature increase = \(41.5 - 27.2 = 14.3 \, ^\circ\text{C}\)
   
   \[Q = (0.89 \, \text{J/g} \, ^\circ\text{C}) \times (29.2 \, \text{g}) \times (14.3 \, ^\circ\text{C}) = 3.7 \times 10^2 \, \text{J} = 89 \, \text{cal}\]

34. \[Q = s \times m \times \Delta T\]

   Specific heat capacity of mercury is \(0.14 \, \text{J/g} \, ^\circ\text{C}\)

   \[100. \, \text{J} = (0.14 \, \text{J/g} \, ^\circ\text{C}) \times 25 \, \text{g} \times \Delta T\]

   \[\Delta T = 28.6 \, ^\circ\text{C} = 29 \, ^\circ\text{C}\]
36. \[ \frac{0.24 \text{ J}}{\text{g °C}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 0.057 \text{ cal/g °C} \]

38. \[ Q = s \times m \times \Delta T \]
\[ 133 \text{ J} = s \times (5.00 \text{ g}) \times (55.1 - 25.2 \text{ °C}) \]
\[ s = 0.89 \text{ J/g °C} \]
The specific heat is 0.89 J/g °C, so the element is most likely aluminum.

40. A calorimeter is an insulated device in which reactions are performed and temperature changes measured, enabling the calculation of heat flows. (See Figure 10.6.)

42.

a. molar mass of S = 32.07 g
\[ 1.00 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.0312 \text{ mol S} \]
From the balanced chemical equation, combustion of 0.0312 mol S would produce 0.0312 mol SO₂
\[ 0.0312 \text{ mol} \times \frac{-296 \text{ kJ}}{\text{mol}} = -9.23 \text{ kJ} \]

b. From the balanced chemical equation, combustion of 0.0312 mol S would produce 0.0312 mol SO₂
\[ 0.501 \text{ mol} \times \frac{-296 \text{ kJ}}{\text{mol}} = -148 \text{ kJ} \]

c. The enthalpy change would be the same in magnitude, but opposite in sign = +296 kJ/mol.

44. It is important when dealing with enthalpy changes to make sure you understand the language in which the number representing \( \Delta H \) is given. Sometimes the \( \Delta H \) value is given “for the equation as written,” including all the stoichiometric coefficients of the balanced equation. At other times, the \( \Delta H \) value may be given on a per-mole basis for one specific reactant or product. In this question, the \( \Delta H \) value is given for the reaction as written, in which one mole of methane reacts with two moles of dioxygen to produce one mole of carbon dioxide and two moles of water. \( \Delta H \) would be –890 kJ for each mole of dioxygen reacts or when one mole of water is produced.

a. –445 kJ/mol of water (because 2 moles of water are formed in the reaction as given)

b. –445 kJ/mol of dioxygen (because 2 moles of dioxygen react in the equation as given)

46. The desired equation, \( 2C(s) + O_2(g) \rightarrow 2CO(g) \), can be generated by taking twice the first equation and adding it to the reverse of the second equation:
\[ 2 \times [C(s) + O_2(g) \rightarrow CO_2(g)] \]
\[ \Delta H = 2 \times -393 \text{ kJ} = -786 \text{ kJ} \text{ (equation doubled)} \]
\[ 2CO_2(g) \rightarrow 2CO(g) + O_2(g) \]
\[ \Delta H = -(-566 \text{ kJ}) = +566 \text{ kJ} \text{ (equation reversed)} \]
\[ 2C(s) + O_2(g) \rightarrow 2CO(g) \]
\[ \Delta H = (-786) + (+566) = -220 \text{ kJ} \]
48. The desired equation can be generated as follows:

\[
\text{NO}(g) + \text{O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g) \quad \Delta H = -199 \text{ kJ}
\]

\[
\frac{1}{2} \times [3\text{O}_2(g) \rightarrow 2\text{O}_3(g)] \quad \Delta H = \frac{1}{2} \times -(-427 \text{ kJ}) = +213.5 \text{ kJ}
\]

\[
\frac{1}{2} \times [2\text{O}(g) \rightarrow \text{O}_2(g)] \quad \Delta H = \frac{1}{2} \times -(+495 \text{ kJ}) = -247.5 \text{ kJ}
\]

\[
\text{NO}(g) + \text{O}(g) \rightarrow \text{NO}_2(g) \quad \Delta H = (-199) + (+213.5) + (-247.5) = -233 \text{ kJ}
\]

50. Once everything in the universe is at the same temperature, no further thermodynamic work can be done. Even though the total energy of the universe will be the same, the energy will have been dispersed evenly, making it effectively useless.

52. Concentrated sources of energy, such as petroleum, are being used to disperse the energy they contain, thus making it unavailable for further use.

54. Petroleum consists mainly of hydrocarbons, which are molecules containing chains of carbon atoms with hydrogen atoms attached to the chains. The fractions are based on the number of carbon atoms in the chains. For example, gasoline is a mixture of hydrocarbons with five to ten carbon atoms in the chains whereas asphalt is a mixture of hydrocarbons with 25 or more carbon atoms in the chains. Different fractions have different physical properties and uses, but all can be combusted to produce energy. (See Table 10.3.)

56. Tetraethyl lead was used as an additive for gasoline to promote smoother running of engines. It is no longer widely because of concerns about the lead being released to the environment as the leaded gasoline is burned.

58. The greenhouse effect is a warming effect due to the presence of gases in the atmosphere that absorb infrared radiation that has reached the earth from the sun and do not allow it to pass back into space. A limited greenhouse effect is desirable because it moderates the temperature changes in the atmosphere that would otherwise be more drastic between daytime when the sun is shining and nighttime. Having too high a concentration of greenhouse gases, however, will elevate the temperature of the earth too much, affecting climate, crops, the polar ice caps, temperature of the oceans, and so on. Carbon dioxide produced by combustion reactions is our greatest concern as a greenhouse gas.

60. If a proposed reaction involves either or both of those phenomena, the reaction will tend to be favorable.

62. Formation of a solid precipitate represents a concentration of matter.

64. The molecules in liquid water are moving around freely and are therefore more “disordered” than when the molecules are held rigidly in a solid lattice in ice. The entropy increases during melting.

66.

\[
a. \quad 462.4 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 110.5 \text{ kcal}
\]

\[
b. \quad 18.28 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 4.369 \text{ kcal}
\]

\[
c. \quad 1.014 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 0.2424 \text{ kcal}
\]

\[
d. \quad 190.5 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 45.53 \text{ kcal}
\]
68. Temperature increase = 75.0 – 22.3 = 52.7 °C

\[ \frac{145 \text{ g} \times 4.184 \text{ J}}{\text{g} \times \text{°C}} \times 52.7^\circ\text{C} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 7641.5 \text{ cal} = 7.65 \text{ kcal} \]

70. The specific heat capacity of iron is 0.45 J/g °C. Then 50. joules is the heat that is applied to the sample of iron and must equal the product of the mass of iron, the specific heat capacity of the iron, and the temperature change undergone by the iron (which is what we want).

\[ Q = s \times m \times \Delta T \]

\[ 50. \text{ J} = (0.45 \text{ J/g } \circ\text{C})(10. \text{ g})(\Delta T) \]

\[ \Delta T = 11^\circ\text{C} \]

72. 2.5 kg water = 2500 g

Temperature change = 55.0 – 18.5 = 36.5°C

\[ Q = s \times m \times \Delta T \]

\[ Q = 4.184 \text{ J/g } \circ\text{C} \times 2500 \text{ g} \times 36.5 ^\circ\text{C} = 3.8 \times 10^5 \text{ J} \]

74. Let \( T_f \) represent the final temperature reached by the system.

For the hot water, heat lost = 50.0 g \times 4.184 J/g °C \times (100. - T^\circ\text{C})

For the cold water, heat gained = 50.0 g \times 4.184 J/g °C \times (T_f - 25 °C)

The heat lost by the hot water must equal the heat gained by the cold water; therefore

\[ 50.0 \text{ g} \times 4.184 \text{ J/g } \circ\text{C} \times (100. - T_f ^\circ\text{C}) = 50.0 \text{ g} \times 4.184 \text{ J/g } \circ\text{C} \times (T_f - 25 ^\circ\text{C}) \]

Solving this equation for \( T_f \) gives \( T_f = 62.5 ^\circ\text{C} = 63 ^\circ\text{C} \)

76. 9.0 J (It requires twice as much heat to warm a sample of twice the mass over the same temperature interval.)

78. Because, for any substance, \( Q = s \times m \times \Delta T \), we can solve this equation for the temperature change, \( \Delta T \). The results are tabulated:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat Capacity</th>
<th>Temperature Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>water (l)</td>
<td>4.184 J/g °C</td>
<td>23.9 °C</td>
</tr>
<tr>
<td>water (s)</td>
<td>2.03 J/g °C</td>
<td>49.3 °C</td>
</tr>
<tr>
<td>water (g)</td>
<td>2.0 J/g °C</td>
<td>50. °C</td>
</tr>
<tr>
<td>aluminum</td>
<td>0.89 J/g °C</td>
<td>( 1.1 \times 10^2 ^\circ\text{C} )</td>
</tr>
<tr>
<td>iron</td>
<td>0.45 J/g °C</td>
<td>( 2.2 \times 10^2 ^\circ\text{C} )</td>
</tr>
<tr>
<td>mercury</td>
<td>0.14 J/g °C</td>
<td>( 7.1 \times 10^2 ^\circ\text{C} )</td>
</tr>
<tr>
<td>carbon</td>
<td>0.71 J/g °C</td>
<td>( 1.4 \times 10^2 ^\circ\text{C} )</td>
</tr>
<tr>
<td>silver</td>
<td>0.24 J/g °C</td>
<td>( 4.2 \times 10^2 ^\circ\text{C} )</td>
</tr>
<tr>
<td>gold</td>
<td>0.13 J/g °C</td>
<td>( 7.7 \times 10^2 ^\circ\text{C} )</td>
</tr>
</tbody>
</table>
80. 
   a. The combustion of gasoline releases heat, so this is an exothermic process. 
   b. \( \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l) \); Heat is released when water vapor condenses, so this is an exothermic process. 
   c. To convert a solid to a gas, heat must be absorbed, so this is an endothermic process. 
   d. Heat must be added (absorbed) in order to break a bond, so this is an endothermic process. 

82. Reversing the first equation and dividing by 6 we get 
   \[
   \frac{1}{6} \text{D} \rightarrow \frac{1}{6} \text{A} + \text{B} \quad \Delta H = +403 \text{ kJ/6}
   \]
   or \( \frac{1}{7} \text{D} \rightarrow \frac{1}{7} \text{A} + \text{B} \quad \Delta H = +67.2 \text{ kJ} \)

   Dividing the second equation by 2 we get 
   \[
   \frac{1}{2} \text{E} + \text{F} \rightarrow \frac{1}{2} \text{A} \quad \Delta H = -105.2 \text{ kJ/2} = -52.6 \text{ kJ}
   \]

   Dividing the third equation by 2 we get 
   \[
   \frac{1}{2} \text{C} \rightarrow \frac{1}{2} \text{E} + \frac{1}{2} \text{D} \quad \Delta H = +64.8 \text{ kJ/2} = +32.4 \text{ kJ}
   \]

   Adding these equations together we get 
   \[
   \frac{1}{2} \text{C} + \text{F} \rightarrow \text{A} + \text{B} + \text{D} \quad \Delta H = 47.0 \text{ kJ}
   \]

84. If 400 kcal is burned per hour walking 4.0 mph, then the total amount of heat burned while walking at 4.0 mph will be given by 
   \[
   Q = 400 \text{ kcal/hr} \times t
   \]
   where \( t \) is the amount of time spent walking (in hours). One gram of fat is consumed for every 7.7 kcal of heat. Therefore, one pound of fat requires 
   \[
   7.7 \text{ kcal/g} \times 454 \text{ g/lb} = 3500 \text{ kcal/lb}
   \]

   Since we want to lose one pound, \( Q = 3500 \text{ kcal} \) and 
   \[
   t = \frac{Q}{400 \text{ kcal/hr}} = \frac{3500 \text{ kcal}}{400 \text{ kcal/hr}} = 8.75 \text{ hr} = 9 \text{ hr}
   \]
CHAPTER 11

Modern Atomic Theory

CHAPTER ANSWERS

2. Rutherford’s experiments determined that the atom had a nucleus containing positively charged particles called protons and neutral particles called neutrons. He established that the nucleus was very small compared to the overall size of the atom. He was not able to determine where the electrons were in the atom or what they were doing.

4. The different forms of electromagnetic radiation are similar in that they all exhibit the same type of wave-like behavior and are propagated through space at the same speed (the speed of light). The types of electromagnetic radiation differ in their frequency (and wavelength) and in the resulting amount of energy carried per photon.

6. The speed of electromagnetic radiation represents how fast a given wave moves through space. The frequency of electromagnetic radiation represents how many complete cycles of the wave pass a given point per second. These two concepts are not the same.

8. more

10. exactly equal to

12. A photon having an energy corresponding to the energy difference between the two states is emitted by an atom in an excited state when it returns to its ground state.

14. higher; lower

16. When excited hydrogen atoms emit their excess energy, the photons of radiation emitted are always of exactly the same wavelength and energy. We consider this to mean that the hydrogen atom possesses only certain allowed energy states and that the photons emitted correspond to the atom changing from one of these allowed energy states to another of the allowed energy states. The energy of the photon emitted corresponds to the energy difference in the allowed states. If the hydrogen atom did not possess discrete energy levels, then we would expect the photons emitted to have random wavelengths and energies.

18. The energy of an emitted photon is identical to the energy change within the atom that gave rise to the emitted photon.

20. When the hydrogen atoms are excited by the applied voltage, they absorb only photons of energy corresponding to the different energy states within the atom. When the hydrogen atoms relax, they emit photons corresponding to the differences between these same energy states, which correspond to light of only certain characteristic wavelengths.

22. orbit

24. Bohr’s theory explained the experimentally observed line spectrum of hydrogen exactly. Bohr’s theory was ultimately discarded because when attempts were made to extend the theory to atoms other than hydrogen, the calculated properties did not correspond closely to experimental measurements.
26. An orbit represents a definite, exact circular pathway around the nucleus in which an electron can be found. An orbital represents a region of space in which there is a high probability of finding the electron.

28. The firefly analogy is intended to demonstrate the concept of a probability map for electron density. In the wave mechanical model of the atom, we cannot say specifically where the electron is in the atom; we can only say where there is a high probability of finding the electron. The analogy is to imagine a time-exposure photograph of a firefly in a closed room. Most of the time, the firefly will be found near the center of the room.

30. Pictures we draw to represent orbitals should only be interpreted as probability maps. They are not meant to represent that the electron moves only on the surface of, or within, the region drawn in the picture. As the mathematical probability of finding the electron never actually becomes zero on moving outward from the nucleus, scientists have decided that pictures of orbitals should represent a 90\% probability that the electron will be found inside the region depicted in the drawing (for 100\% probability, the orbital would have to encompass all space).

32. The \( p \) orbitals, in general, have two lobes and are sometimes described as having a “dumbbell” shape. The \( 2p \) and \( 3p \) orbitals are similar in shape, and in fact there are three equivalent \( 2p \) or \( 3p \) orbitals in the \( 2p \) or \( 3p \) subshell. The orbitals differ in size, mean distance from the nucleus, and energy.

34. \( n = 1 \)

36.

<table>
<thead>
<tr>
<th>Value of ( n )</th>
<th>Possible Subshells</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s</td>
</tr>
<tr>
<td>2</td>
<td>2s, 2p</td>
</tr>
<tr>
<td>3</td>
<td>3s, 3p, 3d</td>
</tr>
<tr>
<td>4</td>
<td>4s, 4p, 4d, 4f</td>
</tr>
</tbody>
</table>

38. Electrons have an intrinsic spin (they spin on their own axes). Geometrically, there are only two directions possible for spin (clockwise or counter-clockwise). This means only two electrons having opposite senses or directions of spin can occupy an orbital. This idea is called the Pauli Exclusion Principle.

40. Increases; as you move out from the nucleus, there is more space and room for more sublevels.

42. Opposite

44. Answer depends on student choices.

46. When a hydrogen atom is in its ground state, the electron is found in the \( 1s \) orbital. The \( 1s \) orbital has the lowest energy of all the possible hydrogen orbitals.

48. The elements in a given vertical column of the periodic table have the same valence electron configuration. Having the same valence electron configuration causes the elements in a given group to have similar chemical properties.

50.

a. C
b. P
52.
   a. Br
   b. Ca
   c. Cl
   d. O

54.
   a. 1s(↑↓)  2s(↑↓)  2p(↑↓)(↑↓)(↑↓)  3s(↑↓)  3p(↑)(   )
   b. 1s(↑↓)  2s(↑↓)  2p(↑↓)(↑↓)(↑↓)  3s(↑↓)  3p(↑)(   )(↑)
   c. 1s(↑↓)  2s(↑↓)  2p(↑↓)(↑↓)(↑↓)  3s(↑↓)  3p(↑↓)(↑↓)(↑↓)  4s(↑↓)
   d. 1s(↑↓)  2s(↑↓)  2p(↑↓)(↑↓)(↑↓)  3s(↑↓)  3p(↑↓)(↑↓)(↑↓)

56. For the representative elements (those filling s and p subshells), the group number gives the number of valence electrons.
   a. five
   b. seven
   c. one
   d. six

58. The properties of Rb and Sr suggest that they are members of Groups 1 and 2, respectively, and so must be filling the 5s orbital. The 5s orbital is lower in energy (and fills before) the 4d orbitals.

60.
   a. Mg
   b. Rb
   c. Sc
   d. P

62.
   a. [Ne] 3s^2 3p^3
   b. [Ne] 3s^2 3p^5
   c. [Ne] 3s^2
   d. [Ar] 4s^2 3d^{10}

64.
   a. one
   b. two
   c. zero
66. The position of the element (both in terms of vertical column and horizontal row) indicates which set of orbitals is being filled last.
   a. 5f
   b. 5f
   c. 4f
   d. 6p

68.
   a. Group 7
   b. Group 8
   c. Group 4
   d. Group 5

70. The metallic elements lose electrons and form positive ions (cations); the nonmetallic elements gain electrons and form negative ions (anions). Remember that the electron itself is negatively charged.

72. All exist as diatomic molecules (F₂, Cl₂, Br₂, I₂); all are nonmetals; all have relatively high electronegativities; all form 1- ions in reacting with metallic elements.

74. Elements at the left of a period (horizontal row) lose electrons more readily. At the left of a period (given principal energy level), the nuclear charge is the smallest, and the electrons are least tightly held.

76. The elements of a given period (horizontal row) have valence electrons in the same principal energy level. Nuclear charge, however, increases across a period going from left to right. Atoms at the left side have smaller nuclear charges and hold onto their valence electrons less tightly.

78. The nuclear charge increases from left to right within a period, pulling progressively more tightly on the valence electrons.

80. Ionization energies decrease in going from top to bottom within a vertical group; ionization energies increase in going from left to right within a horizontal period.
   a. Li
   b. Ca
   c. Cl
   d. S

82. Atomic size increases in going from top to bottom within a vertical group; atomic size decreases in going from left to right within a horizontal period.
   a. Na
   b. S
   c. N
   d. F

84. speed of light
86. photons
88. quantized
90. orbital
92. transition metal
94. spins
96. 
   a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ 
      $[\text{Ar}] 4s^1$
      $1s(\uparrow \downarrow) 2s(\uparrow \downarrow) 2p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow)$
      $3s(\uparrow \downarrow) 3p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) 4s(\uparrow)$
   
   b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
      $[\text{Ar}] 4s^2 3d^2$
      $1s(\uparrow \downarrow) 2s(\uparrow \downarrow) 2p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow)$
      $3s(\uparrow \downarrow) 3p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) 4s(\uparrow \downarrow)$
      $3d(\uparrow)(\uparrow)(\uparrow)(\uparrow)$
   
   c. $1s^2 2s^2 2p^6 3s^2 3p^2$
      $[\text{Ne}] 3s^2 3p^2$
      $1s(\uparrow \downarrow) 2s(\uparrow \downarrow) 2p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow)$
      $3s(\uparrow \downarrow) 3p(\uparrow)(\uparrow)(\uparrow)$
   
   d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
      $[\text{Ar}] 4s^2 3d^6$
      $1s(\uparrow \downarrow) 2s(\uparrow \downarrow) 2p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow)$
      $3s(\uparrow \downarrow) 3p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) 4s(\uparrow \downarrow)$
      $3d(\uparrow)(\uparrow)(\uparrow)(\uparrow)(\uparrow)$
   
   e. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
      $[\text{Ar}] 4s^2 3d^{10}$
      $1s(\uparrow \downarrow) 2s(\uparrow \downarrow) 2p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow)$
      $3s(\uparrow \downarrow) 3p(\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) 4s(\uparrow \downarrow)$
      $3d(\uparrow)(\uparrow)(\uparrow)(\uparrow)(\uparrow)$

98. 
   a. $ns^2$
   b. $ns^2 np^5$
   c. $ns^2 np^3$
   d. $ns^1$
   e. $ns^2 np^4$

100. 
   a. $\lambda = \frac{h}{mv}$
   
   $\lambda = \frac{6.63 \times 10^{-34} \text{ J s}}{(9.1 \times 10^{-31} \text{ kg})(0.90 \times (3.00 \times 10^8 \text{ m/s})]}$
   
   $\lambda = 2.7 \times 10^{-12} \text{ m (0.0027 nm)}$
   
   b. $4.4 \times 10^{-34} \text{ m}$
   
   c. $2 \times 10^{-35} \text{ m}$

   The wavelengths for the ball and the person are \textit{infinitesimally small} whereas the wavelength for the electron is nearly the same order of magnitude as the diameter of a typical atom.
102. Light is emitted from the hydrogen atom only at certain fixed wavelengths. If the energy levels of hydrogen were continuous, a hydrogen atom would emit energy at all possible wavelengths.

104. The third principal energy level of hydrogen is divided into three sublevels (3s, 3p, and 3d); there is a single 3s orbital; there is a set of three 3p orbitals; there is a set of five 3d orbitals. See Figures 11.24–11.28 for the shapes of these orbitals.

106.
   a. incorrect; the \( n = 1 \) shell has only the 1s subshell
   b. correct
   c. incorrect; the \( n = 3 \) shell has only 3s, 3p, and 3d subshells
   d. correct
   e. correct
   f. correct

108.
   a. \( 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5 \)
   b. \( 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 \)
   c. \( 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 \)
   d. \( 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 6s^2 \)

110.
   a. five (2s, 2p)
   b. seven (3s, 3p)
   c. one (3s)
   d. three (3s, 3p)

112.
   a. \([Kr]\) 5s2 4d2
   b. \([Kr]\) 5s2 4d10 5p5
   c. \([Ar]\) 4s2 3d10 4p2
   d. \([Xe]\) 6s1

114.
   a. Se
   b. Se
   c. Rb
   d. V

116. metals, low; nonmetals, high
118. Atomic size increases in going from top to bottom within a vertical group; atomic size decreases in going from left to right within a horizontal period.

a. Ca
b. P
c. K
CHAPTER 12
Chemical Bonding

CHAPTER ANSWERS
2. bond energy
4. covalent
6. In H₂ and HF, the bonding is covalent in nature with an electron pair being shared between the atoms. In H₂, the two atoms are identical (the sharing is equal). In HF, the two atoms are different (the sharing is unequal) and as a result the bond is polar. Both of these are in marked contrast to the situation in NaF: NaF is an ionic compound—an electron has been completely transferred from sodium to fluorine, producing separate ions.

8. A bond is polar if the centers of positive and negative charge do not coincide at the same point. The bond has a negative end and a positive end. Polar bonds will exist in any molecule with nonidentical bonded atoms (although the molecule as a whole may not be polar if the bond dipoles cancel each other). Two simple examples are HF and HCl; in both cases, the negative center of charge is closer to the halogen atom.

10. The level of polarity in a polar covalent bond is determined by the difference in electronegativity of the atoms in the bond.

12. 
   a. I is most electronegative, Rb is least electronegative
   b. Mg is most electronegative, Ca and Sr have similar electronegativities
   c. Br is most electronegative, K is least electronegative

14. Generally, covalent bonds between atoms of different elements are polar.
   a. ionic
   b. polar covalent
   c. covalent

16. For a bond to be polar covalent, the atoms involved in the bond must have different electronegativities (must be of different elements).
   a. nonpolar covalent (atoms of the same element)
   b. nonpolar covalent (atoms of the same element)
   c. polar covalent (atoms of different elements)
   d. polar covalent (atoms of different elements)

18. The atom with the larger electronegativity will be more negative relative to the other atom.
   a. F
   b. neither (similar electronegativity)
20. The greater the electronegativity difference between two atoms, the more ionic will be the bond between those two atoms.
   a. Ca–Cl
   b. Ba–Cl
   c. Fe–I
   d. Be–F

22. The presence of strong bond dipoles and a large overall dipole moment in water make it a polar substance overall. Among the properties of water dependent on its dipole moment are its freezing point, melting point, vapor pressure, and its ability to dissolve many substances.

24. In a diatomic molecule containing two different elements, the more electronegative atom will be the negative end of the molecule, and the less electronegative atom will be the positive end.
   a. H
   b. Cl
   c. I

26. In the figures, the arrow points toward the more electronegative atom.
   a. δ+ P → F δ–
   b. δ+ P → O δ–
   c. δ+ P → C δ–
   d. P and H have similar electronegativities.

28. In the figures, the arrow points toward the more electronegative atom.
   a. δ+ H → C δ–
   b. δ+ N → O δ–
   c. δ+ S → N δ–
   d. δ+ C → N δ–

30. preceding

32. Atoms in covalent molecules gain a configuration like that of a noble gas by sharing one or more pairs of electrons between atoms. Such shared pairs of electrons “belong” to each of the atoms of the bond at the same time. In ionic bonding, one atom completely gives over one or more electrons to another atom, and the resulting ions behave independently of one another.

34.
   a. Br−, Kr (Br has one electron less than Kr.)
   b. Cs+, Xe (Cs has one electron more than Xe.)
   c. P3+, Ar (P has three fewer electrons than Ar.)
   d. S2−, Ar (S has two fewer electrons than Ar.)
36. Atoms or ions with the same number of electrons are said to be isoelectronic.
   a. F\textsuperscript{−}, O\textsuperscript{2−}, N\textsuperscript{3−}
   b. Cl\textsuperscript{−}, S\textsuperscript{2−}, P\textsuperscript{3−}
   c. F\textsuperscript{−}, O\textsuperscript{2−}, N\textsuperscript{3−}
   d. Br\textsuperscript{−}, Se\textsuperscript{2−}, As\textsuperscript{3−}

38. 
   a. Na\textsubscript{2}S: Na has one electron more than a noble gas; S has two fewer electrons than a noble gas.
   b. BaSe: Ba has two electrons more than a noble gas; Se has two fewer electrons than a noble gas.
   c. MgBr\textsubscript{2}: Mg has two electrons more than a noble gas; Br has one electron less than a noble gas.
   d. Li\textsubscript{3}N: Li has one electron more than a noble gas; N has three fewer electrons than a noble gas.
   e. KH: K has one electron more than a noble gas; H has one electron less than a noble gas.

40. 
   a. Sr\textsuperscript{2+}, [Kr]; O\textsuperscript{2−}, [Ne]
   b. Ca\textsuperscript{2+}, [Ar]; H\textsuperscript{−}, [He]
   c. K\textsuperscript{+}, [Ar]; P\textsuperscript{3−}, [Ar]
   d. Ba\textsuperscript{2+}, [Xe]; Se\textsuperscript{2−}, [Kr]

42. An ionic solid such as NaCl consists of an array of alternating positively and negatively charged ions; that is, each positive ion has as its nearest neighbors a group of negative ions, and each negative ion has a group of positive ions surrounding it. In most ionic solids, the ions are packed as tightly as possible.

44. In forming an anion, an atom gains additional electrons in its outermost (valence) shell. Additional electrons in the valence shell increase the repulsive forces between electrons, so the outermost shell becomes larger to accommodate this.

46. Relative ionic sizes are given in Figure 12.9. Within a given horizontal row of the periodic chart, negative ions tend to be larger than positive ions because the negative ions contain a larger number of electrons in the valence shell. Within a vertical group of the periodic table, ionic size increases from top to bottom. In general, positive ions are smaller than the atoms from which they come, whereas negative ions are larger than the atoms from which they come.
   a. F\textsuperscript{−}
   b. Cl\textsuperscript{−}
   c. Ca
   d. I\textsuperscript{−}
48. Relative ionic sizes are given in Figure 12.9. Within a given horizontal row of the periodic chart, negative ions tend to be larger than positive ions because the negative ions contain a larger number of electrons in the valence shell. Within a vertical group of the periodic table, ionic size increases from top to bottom. In general, positive ions are smaller than the atoms from which they come, whereas negative ions are larger than the atoms from which they come.

   a. $I^-$
   b. $F^-$
   c. $F^-$

50. When atoms form covalent bonds, they try to attain a valence electronic configuration similar to that of the following noble gas element. When the elements in the first few horizontal rows of the periodic table form covalent bonds, they will attempt to gain configurations similar to the noble gases helium (two valence electrons, duet rule), and neon and argon (eight valence electrons, octet rule).

52. These elements attain a total of eight valence electrons, making the valence electron configurations similar to those of the noble gases Ne and Ar.

54. When two atoms in a molecule are connected by a triple bond, the atoms share three pairs of electrons (six electrons) in completing their outermost shells. A simple molecule containing a triple bond is acetylene, $C_2H_2$ (H:C:::C:H).

56.

   a. $\text{Rb}^-$
   b. $\text{\ddot{C}l}$
   c. $\text{\ddot{K}r}$
   d. $\text{Ba}$
   e. $\text{\ddot{P}h}$
   f. $\text{\ddot{A}t}$

58.

   a. phosphorus provides 5; each hydrogen provides 1; total valence electrons = 8
   b. each N provides 5; oxygen provides 6; total valence electrons = 16
   c. each carbon provides 4; each hydrogen provides 1; total valence electrons = 26
60.  

a. Each hydrogen provides 1 valence electron; total valence electrons = 2  
\[ \text{H} \text{H} \]

b. Hydrogen provides 1 valence electron; chlorine provides 7 valence electrons; total valence electrons = 8  
\[ \text{H} \text{Cl} \]

c. Carbon provides 4 valence electrons; each fluorine provides 7 valence electrons; total valence electrons = 32  
\[ \text{:F} \text{C} \text{F} \]

d. Each carbon provides 4 valence electrons; each fluorine provides 7 valence electrons; total valence electrons = 50  
\[ \text{:F} \text{C} \text{F} \]

62.  

a. Each H provides 1 valence electron. S provides 6 valence electrons. Total valence electrons = 8  
\[ \text{H} \text{S} \text{H} \text{H} \text{H} \text{S} \text{H} \] or sometimes drawn as follows to indicate the shape \[ \text{H} \text{S} \text{H} \text{H} \text{H} \text{S} \text{H} \]

b. Si provides 4 valence electrons. Each F provides 7 valence electrons. Total valence electrons = 32  
\[ \text{:F} \text{Si} \text{F} \text{F} \]

c. Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 12  
\[ \text{H} \text{C} \text{C} \text{H} \text{H} \text{C} \text{C} \text{H} \text{H} \text{H} \] or \[ \text{H} \text{C} \text{C} \text{H} \text{H} \text{C} \text{C} \text{H} \text{H} \text{H} \]
d. Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 20

\[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{H}
\end{array}
\]

64.

a. N provides 5 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 17 (an odd number). Note that there are several Lewis structures possible because of the odd number of electrons (the unpaired electron can be located on different atoms in the Lewis structure).

\[
\begin{align*}
\text{O} & \equiv \text{N} \equiv \text{O} \\
\text{O} & \equiv \text{N} \equiv \text{O} \\
\text{O} & \equiv \text{N} \equiv \text{O} \\
\text{O} & \equiv \text{N} \equiv \text{O}
\end{align*}
\]

b. Each H provides 1 valence electron. S provides 6 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 32

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{S} \\
\text{H}
\end{array}
\]

c. Each N provides 5 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 34

\[
\begin{align*}
\text{O} & \equiv \text{N} \equiv \text{O} \\
\text{O} & \equiv \text{N} \equiv \text{O} \\
\text{O} & \equiv \text{N} \equiv \text{O} \\
\text{O} & \equiv \text{N} \equiv \text{O}
\end{align*}
\]

66.

a. Cl provides 7 valence electrons. Each O provides 6 valence electrons. The 1– charge means 1 additional electron. Total valence electrons = 26

\[
\begin{array}{c}
\text{:O} \\
\text{O} \\
\text{Cl} \\
\text{:O}
\end{array}
\]

b. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 14

\[
\begin{array}{c}
\text{O} \\
\text{O}
\end{array}
\]
c. Each C provides 4 valence electrons. Each H provides 1 valence electron. Each O provides 6 valence electrons. The 1– charge means 1 additional valence electron. Total valence electrons = 24

\[
\begin{align*}
\text{H} & : \text{O} : \\
\text{H} & = \text{C} = \text{C} = \text{O} :
\end{align*}
\]

\[
\begin{align*}
\text{H} & : \text{O} : \\
\text{H} & = \text{C} = \text{C} = \text{O} :
\end{align*}
\]

68.


\[
\begin{align*}
\text{H} & : \text{O} : \\
\text{H} & = \text{O} = \text{P} = \text{O} :
\end{align*}
\]

b. Each H provides 1 valence electron. P provides 5 valence electrons. Each O provides 6 valence electrons. The 1– charge means one additional valence electron. Total valence electrons = 32

\[
\begin{align*}
\text{H} & : \text{O} : \\
\text{H} & = \text{O} = \text{P} = \text{O} :
\end{align*}
\]

c. P provides 5 valence electrons. Each O provides 6 valence electrons. The 3– charge means three additional valence electrons. Total valence electrons = 32

\[
\begin{align*}
\text{H} & : \text{O} : \\
\text{H} & = \text{O} = \text{P} = \text{O} :
\end{align*}
\]

70. The geometric structure of NH₃ is that of a trigonal pyramid. The nitrogen atom of NH₃ is surrounded by four electron pairs (three are bonding; one is a lone pair). The H–N–H bond angle is somewhat less than 109.5° (due to the presence of the lone pair).

72. The geometric structure of SiF₄ is that of a tetrahedron. The silicon atom of SiF₄ is surrounded by four bonding electron pairs. The F–Si–F bond angle is the characteristic angle of the tetrahedron, 109.5°.

74. The general molecular structure of a molecule is determined by (1) how many electron pairs surround the central atom in the molecule and (2) which of those electron pairs are used for bonding to the other atoms of the molecule. Nonbonding electron pairs on the central atom do, however, cause minor changes in the bond angles compared to the ideal regular geometric structure.
76. You will remember from high school geometry that two points in space are all that is needed to define a straight line. A diatomic molecule represents two points (the nuclei of the atoms) in space.

78. In NF₃, the nitrogen atom has *four* pairs of valence electrons whereas in BF₃ there are only *three* pairs of valence electrons around the boron atom. The nonbonding electron pair on nitrogen in NF₃ pushes the three F atoms out of the plane of the N atom.

80.
   a. four electron pairs arranged in a tetrahedral arrangement with some distortion due to the nonbonding pair
   b. four electron pairs in a tetrahedral arrangement
   c. four electron pairs in a tetrahedral arrangement

82.
   a. tetrahedral (four electron pairs on C and four atoms attached)
   b. nonlinear, V-shaped (four electron pairs on S, but only two atoms attached)
   c. tetrahedral (four electron pairs on Ge and four atoms attached)

84.
   a. basically tetrahedral around the P atom (The hydrogen atoms are attached to two of the oxygen atoms and do not affect greatly the geometrical arrangement of the oxygen atoms around the phosphorus.)
   b. tetrahedral (4 electron pairs on Cl and 4 atoms attached)
   c. trigonal pyramidal (4 electron pairs on S and 3 atoms attached)

86.
   a. approximately 109.5° (the molecule is V-shaped or nonlinear)
   b. approximately 109.5° (the molecule is trigonal pyramidal)
   c. 109.5°
   d. approximately 120° (the double bond makes the molecule flat)

88. The ethylene molecule contains a double bond between the carbon atoms. This makes the molecule planar (flat) with H–C–H and H–C–C bond angles of approximately 120°. The 1,2-dibromoethane molecule would *not* be planar, however. Each carbon would have four bonding pairs of electrons around it, and so the orientation around each carbon atom would be basically tetrahedral with bond angles of approximately 109.5° (assuming all bonds are similar).

90. double

92. The bond with the larger electronegativity difference will be the more polar bond. See Figure 12.3 for electronegativities.
   a. S–F
   b. P–O
   c. C–H
94. The bond energy of a chemical bond is the quantity of energy required to break the bond and separate the atoms.

96. In each case, the element higher up within a group on the periodic table has the higher electronegativity.
   a. Be
   b. N
   c. F

98. For a bond to be polar covalent, the atoms involved in the bond must have different electronegativities (must be of different elements).
   a. polar covalent (different elements)
   b. nonpolar covalent (two atoms of the same element)
   c. polar covalent (different elements)
   d. nonpolar covalent (atoms of the same element)

100. In a diatomic molecule containing two different elements, the more electronegative atom will be the negative end of the molecule, and the less electronegative atom will be the positive end.
   a. oxygen
   b. bromine
   c. iodine

102.
   a. Al \(1s^2 2s^2 2p^6 3s^2 3p^1\)
      \(\text{Al}^{3+} 1s^2 2s^2 2p^6\)
      Ne has the same configuration as Al\(^{3+}\).
   b. Br \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5\)
      \(\text{Br}^- 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6\)
      Kr has the same configuration as Br\(^-\).
   c. Ca \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2\)
      \(\text{Ca}^{2+} 1s^2 2s^2 2p^6 3s^2 3p^6\)
      Ar has the same configuration as Ca\(^{2+}\).
   d. Li \(1s^2 2s^1\)
      \(\text{Li}^+ 1s^2\)
      He has the same configuration as Li\(^+\).
   e. F \(1s^2 2s^2 2p^5\)
      \(\text{F}^- 1s^2 2s^2 2p^6\)
      Ne has the same configuration as F\(^-\).
104.
  a. \( \text{Na}_2\text{Se} \): Na has one electron more than a noble gas; Se has two fewer electrons than a noble gas.
  b. \( \text{RbF} \): Rb has one electron more than a noble gas; F has one electron less than a noble gas.
  c. \( \text{K}_2\text{Te} \): K has one electron more than a noble gas; Te has two fewer electrons than a noble gas.
  d. \( \text{BaSe} \): Ba has two electrons more than a noble gas; Se has two fewer electrons than a noble gas.
  e. \( \text{KAt} \): K has one electron more than a noble gas; At has one electron less than a noble gas.
  f. \( \text{FrCl} \): Fr has one electron more than a noble gas; Cl has one electron less than a noble gas.

106. Relative ionic sizes are indicated in Figure 12.9.
  a. \( \text{Na}^+ \)
  b. \( \text{Al}^{3+} \)
  c. \( \text{F}^- \)
  d. \( \text{Na}^+ \)

108.
  a. H provides 1; N provides 5; each O provides 6; total valence electrons = 24
  b. each H provides 1; S provides 6; each O provides 6; total valence electrons = 32
  c. each H provides 1; P provides 5; each O provides 6; total valence electrons = 32
  d. H provides 1; Cl provides 7; each O provides 6; total valence electrons = 32

110.
  a. \( \text{N}_2\text{H}_4 \): Each N provides 5 valence electrons. Each H provides 1 valence electron. Total valence electrons = 14
    \[
    \begin{array}{c}
    \text{H} \\
    \text{N} \\
    \text{N} \\
    \text{H} \\
    \text{H} \\
    \end{array}
    \]
  b. \( \text{C}_2\text{H}_6 \): Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 14
    \[
    \begin{array}{c}
    \text{H} \\
    \text{H} \\
    \text{H} \\
    \text{C} \\
    \text{C} \\
    \text{H} \\
    \end{array}
    \]
  c. \( \text{NCl}_3 \): N provides 5 valence electrons. Each Cl provides 7 valence electrons. Total valence electrons = 26
    \[
    \begin{array}{c}
    \text{Cl} \\
    \text{Cl} \\
    \text{N} \\
    \text{Cl} \\
    \end{array}
    \]
d. **SiCl\textsubscript{4}**: Si provides 4 valence electrons. Each Cl provides 7 valence electrons. Total valence electrons = 32

\[ \text{Si} - \underbrace{\text{Cl} - \text{Cl} - \text{Cl} - \text{Cl}}_{32} \]

112.

a. **NO\textsubscript{3}\textsuperscript{–}**: N provides 5 valence electrons. Each O provides 6 valence electrons. The 1– charge means one additional valence electron. Total valence electrons = 24

\[ \overline{\text{O}} - \text{N} - \overline{\text{O}} - \overline{\text{O}} - \overline{\text{O}} \]

b. **CO\textsubscript{3}\textsuperscript{2–}**: C provides 4 valence electrons. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 24

\[ \overline{\text{O}} - \text{C} - \overline{\text{O}} - \overline{\text{O}} - \overline{\text{O}} \]

c. **NH\textsubscript{4}\textsuperscript{+}**: N provides 5 valence electrons. Each H provides 1 valence electron. The 1+ charge means one less valence electron. Total valence electrons = 8

\[ \overline{\text{H}} - \text{N} - \text{H} - \text{H} \]

114.

a. four electron pairs arranged tetrahedrally about C
b. four electron pairs arranged tetrahedrally about Ge
c. three electron pairs arranged trigonally (planar) around B

116.

a. **ClO\textsubscript{3}–**, trigonal pyramid (lone pair on Cl)
b. **ClO\textsubscript{2}–**, nonlinear (V-shaped, two lone pairs on Cl)
c. **ClO\textsubscript{4}–**, tetrahedral (all pairs on Cl are bonding)

118.

a. nonlinear (V-shaped)
b. trigonal planar
c. basically trigonal planar around the C (the H is attached to one of the O atoms and distorts the shape around the carbon only slightly)
d. linear
120. Ionic compounds tend to be hard, crystalline substances with relatively high melting and boiling points. Covalently bonded substances tend to be gases, liquids, or relatively soft solids with much lower melting and boiling points.
Cumulative Review Chapters 10–12

ANSWERS

2. Temperature is a measure of the random motions of the components of a substance. In other words, temperature is a measure of the average kinetic energy of the particles in a sample. The molecules in warm water must be moving faster than the molecules in cold water (the molecules have the same mass, so if the temperature is higher, the average velocity of the particles must be higher in the warm water). Heat is the energy that flows because of a difference in temperature.

4. Thermodynamics is the study of energy and energy changes. The first law of thermodynamics is the law of conservation of energy: The energy of the universe is constant. Energy cannot be created or destroyed, only transferred from one place to another or from one form to another. The internal energy of a system, $E$, represents the total of the kinetic and potential energies of all the particles in a system. A flow of heat may be produced when there is a change in internal energy in the system, but it is not correct to say that the system “contains” the heat. Part of the internal energy is converted to heat energy during the process (under other conditions, the change in internal energy might be expressed as work rather than as heat flow).

6. The enthalpy change represents the heat energy that flows (at constant pressure) on a molar basis when a reaction occurs. The enthalpy change is indeed a state function (that we make great use of in Hess’s Law calculations). Enthalpy changes are typically measured in insulated reaction vessels called calorimeters (a simple calorimeter is shown in Figure 10.6 in the text).

8. Consider petroleum. A gallon of gasoline contains concentrated, stored energy. We can use that energy to make our car move, but when we do, the energy stored in the gasoline is dispersed to the environment. Although the energy is still there (it is conserved), it is no longer in a concentrated, useful form. So although the energy content of the universe remains constant, the energy that is now stored in concentrated forms in oil, coal, wood, and other sources is gradually being dispersed to the universe where it can do no work.

10. A “driving force” is an effect that tends to make a process occur. Two important driving forces are dispersion of energy during a process or dispersion of matter during a process (“energy spread” and “matter spread”). For example, a log burns in a fireplace because the energy contained in the log is dispersed to the universe when it burns. If we put a teaspoon of sugar into a glass of water, the dissolving of the sugar is a favorable process because the matter of the sugar is dispersed when it dissolves. Entropy is a measure of the randomness or disorder in a system. The entropy of the universe is constantly increasing because of “matter spread” and “energy spread.” A spontaneous process is one that occurs without outside intervention; the spontaneity of a reaction depends on the energy spread and matter spread if the reaction takes place. A reaction that disperses energy and also disperses matter will always be spontaneous. Reactions that require an input of energy may still be spontaneous if the matter spread is large enough.

12. molar mass $C_2H_5OH = 46.07 \text{ g}$:

$$1.25 \text{ g } C_2H_5OH \times \frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g}} \times \frac{-1409 \text{ kJ}}{1 \text{ mol } C_2H_5OH} = -38.2 \text{ kJ}$$
14. An atom is said to be in its ground state when it is in its lowest possible energy state. When an atom possesses more energy than its ground state energy, the atom is said to be in an excited state. An atom is promoted from its ground state to an excited state by absorbing energy; when the atom returns from an excited state to its ground state it emits the excess energy as electromagnetic radiation. Atoms do not gain or emit radiation randomly, but rather do so only in discrete bundles of radiation called photons. The photons of radiation emitted by atoms are characterized by the wavelength (color) of the radiation. Longer wavelength photons carry less energy than shorter wavelength photons. The energy of a photon emitted by an atom corresponds exactly to the difference in energy between two allowed energy states in an atom. Thus, we can use an observable phenomenon (emission of light by excited atoms) to gain insight into the energy changes taking place within the atom.

16. Bohr pictured the electron moving in only certain circular orbits around the nucleus. Each particular orbit (corresponding to a particular distance from the nucleus) had associated with it a particular energy (resulting from the attraction between the nucleus and the electron). When an atom absorbs energy, the electron moves from its ground state in the orbit closest to the nucleus \((n = 1)\) to an orbit farther away from the nucleus \((n = 2, 3, 4, \ldots)\). When an excited atom returns to its ground state, corresponding to the electron moving from an outer orbit to the orbit nearest the nucleus, the atom emits the excess energy as radiation. As the Bohr orbits are of fixed distances from the nucleus and from each other, when an electron moves from one fixed orbit to another, the energy change is of a definite amount. This corresponds to a photon of a particular characteristic wavelength and energy being emitted. The original Bohr theory worked very well for hydrogen: Bohr even predicted emission wavelengths for hydrogen that had not yet been seen, but were subsequently found at the exact wavelengths Bohr had calculated. However, when the simple Bohr model for the atom was applied to the emission spectra of other elements, the theory could not predict or explain the observed emission spectra.

18. The lowest energy hydrogen atomic orbital is called the \(1s\) orbital. The \(1s\) orbital is spherical in shape (that is, the electron density around the nucleus is uniform in all directions from the nucleus). The \(1s\) orbital represents a probability map of electron density around the nucleus for the first principal energy level. The orbital does not have a sharp edge (it appears fuzzy) because the probability of finding the electron does not drop off suddenly with distance from the nucleus. The orbital does not represent just a spherical surface on which the electron moves (this would be similar to Bohr’s original theory). When we draw a picture to represent the \(1s\) orbital we are indicating that the probability of finding the electron within this region of space is greater than 90%. We know that the likelihood of finding the electron within this orbital is very high, but we still don’t know exactly where in this region the electron is at a given instant in time.

20. The third principal energy level of hydrogen is divided into three sublevels, the \(3s\), \(3p\), and \(3d\) sublevels. The \(3s\) subshell consists of the single \(3s\) orbital. Like the other \(s\) orbitals, the \(3s\) orbital is spherical in shape. The \(3p\) subshell consists of a set of three equal-energy \(3p\) orbitals. Each of these \(3p\) orbitals has the same shape (“dumbbell”), but each of the \(3p\) orbitals is oriented in a different direction in space. The \(3d\) subshell consists of a set of five \(3d\) orbitals that have the shapes indicated in Figure 11.28, and which are oriented in various directions around the nucleus. Students sometimes say that the \(3d\) orbitals have the shape of a “four-leaf clover”. The fourth principal energy level of hydrogen is divided into four sublevels, the \(4s\), \(4p\), \(4d\), and \(4f\) orbitals. The \(4s\) subshell consists of the single \(4s\) orbital. The \(4p\) subshell consists of a set of three \(4p\) orbitals. The \(4d\) subshell consists of a set of five \(4d\) orbitals. The shapes of the \(4s\), \(4p\), and \(4d\) orbitals are the same as the shapes of the orbitals of the third principal energy level (the orbitals of the fourth principal energy level are larger and further from the nucleus than the orbitals of the third level, however). The fourth principal energy level, because it is further from the nucleus,
also contains a 4f subshell consisting of seven 4f orbitals (the shapes of the 4f orbitals are beyond the scope of this text).

22. Atoms have a series of principal energy levels symbolized by the letter $n$. The $n = 1$ level is the closest to the nucleus, and the energies of the levels increase as the value of $n$ increases going out from the nucleus. Each principal energy level is divided into a set of sublevels of different characteristic shapes (designated by the letters $s$, $p$, $d$, and $f$). Each sublevel is further subdivided into a set of orbitals. Each $s$ subshell consists of a single $s$ orbital, each $p$ subshell consists of a set of three $p$ orbitals, each $d$ subshell consists of a set of five $d$ orbitals, etc. A given orbital can be empty or it can contain one or two electrons, but never more than two electrons (if an orbital contains two electrons, then the electrons must have opposite intrinsic spins). The shape we picture for an orbital represents only a probability map for finding electrons; the shape does not represent a trajectory or pathway for electron movements.

24. The valence electrons are the electrons in an atom’s outermost shell. The valence electrons are those most likely to be involved in chemical reactions because they are at the outside edge of the atom.

26. From the column and row location of an element, you should be able to determine what the valence shell of an element has for its electronic configuration. For example, the element in the third horizontal row in the second vertical column, has $3s^2$ as its valence configuration. We know that the valence electrons are in the $n = 3$ shell because the element is in the third horizontal row. We know that the valence electrons are $s$ electrons because the first two electrons in a horizontal row are always in an $s$ subshell. We know that there are two electrons because the element is the second element in the horizontal row. As an additional example, the element in the seventh vertical column of the second horizontal row in the periodic table has valence configuration $2s^22p^5$.

28. The ionization energy of an atom represents the energy required to remove an electron from the atom. As one goes from top to bottom in a vertical group in the periodic table, the ionization energies decrease (it becomes easier to remove an electron). As one goes down the table within a group, the valence electrons are farther and farther from the nucleus and are less tightly held. The ionization energies increase when going from left to right within a horizontal row in the periodic table. The left-hand side of the periodic table is where the metallic elements that lose electrons relatively easily are found. The right-hand side of the periodic table is where the nonmetallic elements are found and, rather than losing electrons, these elements tend to gain electrons. Within a given horizontal row in the periodic table, the valence electrons are all in the same principal energy shell. However, as you go from left to right in the horizontal row, the nuclear charge that holds onto the electrons is increasing one unit with each successive element, making it that much more difficult to remove an electron. The relative sizes of atoms also vary systematically with the location of an element in the periodic table. Within a given vertical group, the atoms get progressively larger when going from the top of the group to the bottom. The valence electrons of the atoms are in progressively higher principal energy shells (and are progressively farther from the nucleus) as we go down in a group. In going from left to right within a horizontal row in the periodic table, the atoms get progressively smaller. Although all the elements in a given horizontal row in the periodic table have their valence electrons in the same principal energy shell, the nuclear charge is progressively increasing from left to right, making the given valence shell progressively smaller as the electrons are drawn more closely to the nucleus.

30. Ionic bonding results when elements of very different electronegativities react with each other. Typically a metallic element reacts with a nonmetallic element, the metallic element losing electrons and forming positive ions and the nonmetallic element gaining electrons and forming negative ions. Sodium chloride, NaCl, is an example of a typical ionic compound. The aggregate
form of such a compound consists of a crystal lattice of alternating positively and negatively charged ions. A given positive ion is attracted by several surrounding negatively charged ions, and a given negative ion is attracted by several surrounding positively charged ions. Similar electrostatic attractions go on in three dimensions throughout the crystal of ionic solid, leading to a very stable system (with very very melting and boiling points, for example). We know that ionic-bonded solids do not conduct electricity in the solid state (because the ions are held tightly in place by all the attractive forces), but such substances are strong electrolytes when melted or when dissolved in water (either process sets the ions free to move around).

32. Electronegativity represents the relative ability of an atom in a molecule to attract shared electrons towards itself. In order for a bond to be polar, one of the atoms in the bond must attract the shared electron pair towards itself and away from the other atom of the bond. This can only happen if one atom of the bond is more electronegative than the other (that is, that there is a considerable difference in electronegativity for the two atoms of the bond). If two atoms in a bond have the same electronegativity, then the two atoms pull the electron pair equally, and the bond is nonpolar covalent. If two atoms sharing a pair of electrons have vastly different electronegativities, the electron pair will be pulled so strongly by the more electronegative atom that a negative ion may be formed (as well as a positive ion for the second atom), and ionic bonding will result. If the difference in electronegativity between two atoms sharing an electron pair is somewhere in between these two extremes (equal sharing of the electron pair and formation of ions), then a polar covalent bond will result.

34. It has been observed over many experiments that when an active metal like sodium or magnesium reacts with a nonmetal, the sodium atoms always form Na\(^+\) ions, and the magnesium atoms always form Mg\(^{2+}\) ions. It has been further observed that aluminum always forms only the Al\(^{3+}\) ion. When nitrogen, oxygen, or fluorine form simple ions, the ions that are formed are always N\(^3-\), O\(^2-\), and F\(^-\), respectively. Clearly the facts that these elements always form the same ions and that those ions all contain eight electrons in the outermost shell led scientists to speculate that there must be something fundamentally stable about a species that has eight electrons in its outermost shell (like the noble gas neon). The repeated observation that so many elements, when reacting, tend to attain an electronic configuration that is isoelectronic with a noble gas led chemists to speculate that all elements try to attain such a configuration for their outermost shells. In general, when atoms of a metal react with atoms of a nonmetal, the metal atoms lose electrons until they have the configuration of the preceding noble gas, and the nonmetal atoms gain electrons until they have the configuration of the following noble gas. Covalently and polar covalently bonded molecules also strive to attain pseudo-noble gas electronic configurations. For a covalently bonded molecule like F\(_2\), in which neither fluorine atom has a greater tendency than the other to gain or lose electrons completely, each F atom provides one electron of the pair of electrons that constitutes the covalent bond. Each F atom also feels the influence of the other F atom’s electron in the shared pair, and each F atom effectively fills its outermost shell. Similarly, in polar covalently bonded molecules like HF or HCl, the shared pair of electrons between the atoms effectively completes the outer electron shell of each atom simultaneously to give each atom a noble gas-like electronic configuration.

36. Bonding between atoms to form a molecule involves only the valence electrons of the atoms (not the inner core electrons). So when we draw the Lewis structure of a molecule, we show only these valence electrons (both bonding valence electrons and nonbonding valence electrons, however). The most important requisite for the formation of a stable compound (which we try to demonstrate when we write Lewis structures) is that each atom of a molecule attains a noble gas electron configuration. When we write Lewis structures, we arrange the bonding and nonbonding valence electrons to try to complete the octet (or duet) for as many atoms as is possible.
38. Obviously, you could choose practically any molecule for your discussion. Let’s illustrate the method for ammonia, \( \text{NH}_3 \). First count up the total number of valence electrons available in the molecule (without regard to what atom they officially “belong”). Remember that for the representative elements the number of valence electrons is indicated by which group the element is found in on the periodic table. For \( \text{NH}_3 \), because nitrogen is in Group 5, one nitrogen atom would contribute five valence electrons. Because hydrogen atoms have only one electron each, the three hydrogen atoms provide an additional three valence electrons for a total of eight valence electrons overall. Next write down the symbols for the atoms in the molecule, and use one pair of electrons (represented by a line) to form a bond between each pair of bound atoms.

\[
\text{H} \quad \text{N} \quad \text{H}
\]

These three bonds use six of the eight valence electrons. Since each hydrogen already has its duet in what we have drawn so far while the nitrogen atom has only six electrons around it so far, the final two valence electrons must represent a lone pair on the nitrogen.

\[
\text{H} \quad \text{N} \quad \text{H}
\]

40. There are several types of exceptions to the octet rule described in the text. The octet rule is really a “rule of thumb” that we apply to molecules unless we have some evidence that a molecule does not follow the rule. There are some common molecules that, from experimental measurements, we know do not follow the octet rule. For example, in \( \text{BF}_3 \) the boron atom has only six valence electrons, whereas in \( \text{BeF}_2 \) the beryllium atom has only four valence electrons. Other molecules that are exceptions to the octet rule include any molecule with an odd number of valence electrons (such as \( \text{NO} \) or \( \text{NO}_2 \)). You can’t get an octet (an even number) of electrons around each atom in a molecule with an odd number of valence electrons. Even the oxygen gas we breathe is an exception to the octet rule. Although we can write a Lewis structure for \( \text{O}_2 \) satisfying the octet rule for each oxygen, we know from experiment that \( \text{O}_2 \) contains unpaired electrons (which would not be consistent with a structure in which all the electrons were paired up.)

42. | Number of Valence Pairs | Bond Angle | Example(s) |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>180°</td>
<td>( \text{BeF}_2, \text{BeH}_2 )</td>
</tr>
<tr>
<td>3</td>
<td>120°</td>
<td>( \text{BCl}_3 )</td>
</tr>
<tr>
<td>4</td>
<td>109.5°</td>
<td>( \text{CH}_4, \text{CCl}_4, \text{GeF}_4 )</td>
</tr>
<tr>
<td>Element</td>
<td>Configuration</td>
<td>Ion configuration</td>
</tr>
<tr>
<td>---------</td>
<td>---------------</td>
<td>------------------</td>
</tr>
<tr>
<td>F</td>
<td>1s²2s²2p⁵</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>Sr</td>
<td>1s²2s²2p⁶3s²3p⁴4s³3d¹⁰4p⁶5s²</td>
<td>1s²2s²2p⁶3s²3p⁴4s³3d¹⁰4p⁶</td>
</tr>
<tr>
<td>Al</td>
<td>1s²2s²2p³3s³3p⁴</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>N</td>
<td>1s²2s²2p⁵</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>P</td>
<td>1s²2s²2p⁶3s²3p⁴</td>
<td>1s²2s²2p⁶3s²3p⁶</td>
</tr>
<tr>
<td>I</td>
<td>[Kr]5s²4d¹⁰5p⁶</td>
<td>[Kr]5s²4d¹⁰5p⁶ = [Xe]</td>
</tr>
<tr>
<td>Rb</td>
<td>1s²2s²2p⁶3s³3p⁴4s³3d¹⁰4p⁶5s¹</td>
<td>1s²2s²2p⁶3s³3p⁴4s³3d¹⁰4p⁶</td>
</tr>
<tr>
<td>Na</td>
<td>1s²2s²2p⁵3s²</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>Cl</td>
<td>1s²2s²2p⁶3s²3p⁴</td>
<td>1s²2s²2p⁶3s²3p⁶</td>
</tr>
<tr>
<td>Ca</td>
<td>1s²2s²2p⁶3s²3p⁴4s²</td>
<td>1s²2s²2p⁶3s²3p⁶</td>
</tr>
</tbody>
</table>
CHAPTER ANSWERS

2. Solids and liquids have essentially fixed volumes and cannot be compressed easily. Gases have volumes that depend on their conditions and can be compressed or expanded by changes in those conditions. Although the particles of matter in solids are essentially fixed in position (the solid is rigid), the particles in liquids and gases are free to move.

4. Figure 13.2 in the text shows a simple mercury barometer; a tube filled with mercury is inverted over a reservoir (containing mercury) that is open to the atmosphere. When the tube is inverted, the mercury falls to the level at which the pressure of the atmosphere is sufficient to support the column of mercury. One standard atmosphere of pressure is taken to be the pressure capable of supporting a column of mercury to a height of 760.0 mm above the reservoir level.

6. Pressure units include mm Hg, torr, pascals, and psi. The unit “mm Hg” is derived from the barometer since, in a traditional mercury barometer, we measure the height of the mercury column (in millimeters) above the reservoir of mercury.

8.

\[
a. \quad 745 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.980 \text{ atm} \\
b. \quad 771 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.01 \text{ atm} \\
c. \quad 69.6 \text{ cm Hg} \times \frac{1 \text{ atm}}{76 \text{ cm Hg}} = 0.912 \text{ atm} \\
d. \quad 2.59 \times 10^{-3} \text{ mm} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 3.41 \times 10^{-6} \text{ atm}
\]

10.

\[
a. \quad 792 \text{ torr} = 792 \text{ mm Hg} \\
b. \quad 71.4 \text{ cm Hg} \times \frac{10 \text{ mm}}{1 \text{ cm}} = 714 \text{ mm Hg} \\
c. \quad 0.981 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 746 \text{ mm Hg} \\
d. \quad 1.09 \text{ kPa} \times \frac{760 \text{ mm Hg}}{101.325 \text{ kPa}} = 8.18 \text{ mm Hg}
\]
12.

a. \(2.07 \times 10^6 \text{ Pa} \times \frac{1 \text{ kPa}}{1000 \text{ Pa}} = 2.07 \times 10^3 \text{ kPa}\)

b. \(795 \text{ mm Hg} \times \frac{101.325 \text{ kPa}}{760 \text{ mm}} = 106 \text{ kPa}\)

c. \(10.9 \text{ atm} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 1.10 \times 10^3 \text{ kPa}\)

d. \(659 \text{ torr} = 659 \text{ mm Hg}\)
\(659 \text{ torr} \times \frac{101.325 \text{ kPa}}{760 \text{ mm}} = 87.9 \text{ kPa}\)

14. Additional mercury increases the pressure on the gas sample causing the volume of the gas upon which the pressure is exerted to decrease (Boyle’s Law)

16. \(PV = k; \ P_1V_1 = P_2V_2\)

18.

a. \(P_1 = 1.09 \text{ atm} \quad P_2 = ?\)
\(V_1 = 525 \text{ mL} \quad V_2 = 122 \text{ mL}\)
\(P_2 = \frac{PV_1}{V_2} = \frac{(1.09 \text{ atm})(525 \text{ mL})}{(122 \text{ mL})} = 4.69 \text{ atm}\)

b. \(P_1 = 760. \text{ mm Hg} \quad P_2 = ?\)
\(V_1 = 25.2 \text{ L} \quad V_2 = 1.01 \text{ L}\)
\(P_2 = \frac{PV_1}{V_2} = \frac{(760. \text{ mm})(25.2 \text{ L})}{(1.01 \text{ L})} = 1.90 \times 10^4 \text{ mm Hg}\)

c. \(P_1 = 20.2 \text{ atm} \quad P_2 = ?\)
\(V_1 = 2.79 \text{ mL} \quad V_2 = 209 \text{ mL}\)
\(P_2 = \frac{PV_1}{V_2} = \frac{(20.2 \text{ atm})(2.79 \text{ mL})}{(209 \text{ mL})} = 0.270 \text{ atm}\)

20.

a. \(P_1 = 1.07 \text{ atm} \quad P_2 = 2.14 \text{ atm}\)
\(V_1 = 291 \text{ mL} \quad V_2 = ? \text{ mL}\)
\(V_2 = \frac{PV_1}{P_2} = \frac{(1.07 \text{ atm})(291 \text{ mL})}{(2.14 \text{ atm})} = 146 \text{ mL}\)

b. \(P_1 = 755 \text{ mm Hg} \quad P_2 = 3.51 \text{ atm} = 2668 \text{ mm Hg}\)
\(V_1 = 1.25 \text{ L} \quad V_2 = ? \text{ L}\)
\(V_2 = \frac{PV_1}{P_2} = \frac{(755 \text{ mm Hg})(1.25 \text{ L})}{(2668 \text{ mm Hg})} = 0.354 \text{ L}\)
c. \[ P_1 = 101.4 \text{ kPa} = 760.6 \text{ mm Hg} \quad \text{and} \quad P_2 = ? \text{ mm Hg} \]
\[ V_1 = 2.71 \text{ L} \quad \text{and} \quad V_2 = 3.00 \text{ L} \]
\[
\frac{P_2}{V_2} = \frac{(760.6 \text{ mm Hg})(2.71 \text{ L})}{(3.00 \text{ L})} = 687 \text{ mm Hg}
\]

22. \[ P_1 = P_1 \quad \text{and} \quad P_2 = 2 \times P_1 \]
\[ V_1 = 1.04 \text{ L} \quad \text{and} \quad V_2 = ? \text{ L} \]
\[
V_2 = \frac{P_1V_1}{P_2} = \frac{(P_1)(1.04 \text{ L})}{(2 \times P_1)} = \frac{1.04 \text{ L}}{2} = 0.520 \text{ L}
\]

24. \[ P_2 = \frac{P_1V_1}{V_2} = \frac{(1.00 \text{ atm})(27.2 \text{ mL})}{(1.00 \text{ mL})} = 27.2 \text{ atm}
\]

26. Charles’s Law indicates that an ideal gas decreases by \( \frac{1}{273} \) of its volume for every degree Celsius its temperature is lowered. This means an ideal gas would approach a volume of zero at \(-273^\circ\text{C}\).

28. \[ V = kT; \quad \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
30. \[ V_1 = 25.2 \text{ mL} \quad \text{and} \quad V_2 = ? \text{ mL} \]
\[ T_1 = 29^\circ\text{C} = 302 \text{ K} \quad \text{and} \quad T_2 = 151^\circ\text{C} = 424 \text{ K} \]
\[
V_2 = \frac{V_1T_2}{T_1} = \frac{(25.2 \text{ mL})(424 \text{ K})}{(302 \text{ K})} = 35.4 \text{ mL}
\]

32.

a. \[ V_1 = 73.5 \text{ mL} \quad \text{and} \quad V_2 = ? \text{ L} \]
\[ T_1 = 0^\circ\text{C} = 273 \text{ K} \quad \text{and} \quad T_2 = 25^\circ\text{C} = 298 \text{ K} \]
\[
V_2 = \frac{V_1T_2}{T_1} = \frac{(73.5 \text{ mL})(298 \text{ K})}{(273 \text{ K})} = 80.2 \text{ mL}
\]

b. \[ V_1 = 15.2 \text{ L} \quad \text{and} \quad V_2 = 10.0 \text{ L} \]
\[ T_1 = 25^\circ\text{C} = 298 \text{ K} \quad \text{and} \quad T_2 = 9^\circ\text{C} \]
\[
T_2 = \frac{V_1T_1}{V_2} = \frac{(10.0 \text{ L})(298 \text{ K})}{(15.2 \text{ L})} = 196 \text{ K} = -77^\circ\text{C}
\]

b. \[ V_1 = 1.75 \text{ mL} \quad \text{and} \quad V_2 = ? \text{ mL} \]
\[ T_1 = 2.3 \text{ K} \quad \text{and} \quad T_2 = 0^\circ\text{C} = 273 \text{ K} \]
\[
V_2 = \frac{V_1T_2}{T_1} = \frac{(1.75 \text{ mL})(273 \text{ K})}{(2.3 \text{ K})} = 208 \text{ mL} (2.1 \times 10^2 \text{ mL})
\]
34.   
   a. \( V_1 = 2.01 \times 10^2 \text{ L} \) \hspace{1cm} \( V_2 = 5.00 \text{ L} \)  
      \( T_1 = 1150^\circ \text{C} = 1423 \text{ K} \) \hspace{1cm} \( T_2 = ?^\circ \text{C} \)  
      \( T_2 = \frac{V_2 T_1}{V_1} = \frac{(5.00 \text{ L})(1423 \text{ K})}{(201 \text{ L})} = 35.4 \text{ K} = -238^\circ \text{C} \)  
   b. \( V_1 = 44.2 \text{ mL} \) \hspace{1cm} \( V_2 = ? \text{ mL} \)  
      \( T_1 = 298 \text{ K} \) \hspace{1cm} \( T_2 = 0 \)  
      \( V_2 = \frac{V_1 T_2}{T_1} = \frac{(44.2 \text{ mL})(0 \text{ K})}{(298 \text{ K})} = 0 \text{ mL} \) (0 K is absolute zero)  
   c. \( V_1 = 44.2 \text{ mL} \) \hspace{1cm} \( V_2 = ? \text{ mL} \)  
      \( T_1 = 298 \text{ K} \) \hspace{1cm} \( T_2 = 0^\circ \text{C} = 273 \text{ K} \)  
      \( V_2 = \frac{V_1 T_2}{T_1} = \frac{(44.2 \text{ mL})(273 \text{ K})}{(298 \text{ K})} = 40.5 \text{ mL} \)  

36. If the temperature is decreased by a factor of two, the volume also will decrease by a factor of two. The new volume of the sample will be half its original volume.

38. \( V_2 = \frac{V_1 T_2}{T_1} \)  

<table>
<thead>
<tr>
<th>Temp, °C</th>
<th>90</th>
<th>80</th>
<th>70</th>
<th>60</th>
<th>50</th>
<th>40</th>
<th>30</th>
<th>20</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume, mL</td>
<td>124</td>
<td>121</td>
<td>117</td>
<td>113</td>
<td>110</td>
<td>107</td>
<td>103</td>
<td>100</td>
</tr>
</tbody>
</table>

40. \( V = an; \) \( V_1/n_1 = V_2/n_2 \)

42. molar mass of He = 4.003 g  
\[ 1.00 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} = 0.250 \text{ mol He} \]  
\( V_2 = 22.4 \text{ L} \times \frac{0.250 \text{ mol}}{1.00 \text{ mol}} = 5.60 \text{ L} \)

44. molar mass of Ar = 39.95 g  
\[ 2.71 \text{ g Ar} \times \frac{1 \text{ mol}}{39.95 \text{ g}} = 0.0678 \text{ mol Ar} \]  
\[ 4.21 \text{ L} \times \frac{1.29 \text{ mol}}{0.0678 \text{ mol}} = 80.1 \text{ L} \]

46. Real gases most closely approach ideal gas behavior under conditions of relatively high temperatures (0°C or higher) and relatively low pressures (1 atm or lower).
48. For an ideal gas, $PV = nRT$ is true under any conditions. Consider a particular sample of gas (so that $n$ remains constant) at a particular fixed pressure (so that $P$ remains constant also). Suppose that at temperature $T_1$ the volume of the gas sample is $V_1$. Then for this set of conditions, the ideal gas equation would be given by

$$PV_1 = nRT_1.$$ 

If we then change the temperature of the gas sample to a new temperature $T_2$, the volume of the gas sample changes to a new volume $V_2$. For this new set of conditions, the ideal gas equation would be given by

$$PV_2 = nRT_2.$$ 

If we make a ratio of these two expressions for the ideal gas equation for this gas sample and cancel out terms that are constant for this situation ($P$, $n$, and $R$), we get

$$\frac{PV_1}{PV_2} = \frac{nRT_1}{nRT_2}$$

This can be rearranged to the familiar form of Charles’s law

$$\frac{V_1}{V_2} = \frac{T_1}{T_2}$$

50.

a. $P = 782 \text{ mm Hg} = 1.03 \text{ atm}; \ T = 27^\circ \text{C} = 300 \text{ K}$

$$V = \frac{nRT}{P} = \frac{(0.210 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})}{(1.03 \text{ atm})} = 5.02 \text{ L}$$

b. $V = 644 \text{ mL} = 0.644 \text{ L}$

$$P = \frac{nRT}{V} = \frac{(0.0921 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(303 \text{ K})}{(0.644 \text{ L})} = 3.56 \text{ atm}$$

$$= 2.70 \times 10^3 \text{ mm Hg}$$

c. $P = 745 \text{ mm} = 0.980 \text{ atm}$

$$T = \frac{PV}{nR} = \frac{(0.980 \text{ atm})(11.2 \text{ L})}{(0.401 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})} = 334 \text{ K}$$

52. $P = 764 \text{ mm Hg} = 1.005 \text{ atm}; \text{ molar mass Ar = 39.95 g}; \ 1.05 \text{ g Ar} = 0.02628 \text{ mol}$

$$T = \frac{PV}{nR} = \frac{(1.005 \text{ atm})(5.21 \text{ L})}{(0.02628 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})} = 2428 \text{ K} = 2.43 \times 10^3 \text{ K}$$

54. molar mass He = 4.003 g; 5.21 kg He = 1302 mol; $27^\circ \text{C} = 300 \text{ K}$

$$P = \frac{nRT}{V} = \frac{(1302 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})}{(245 \text{ L})} = 131 \text{ atm}$$
56. molar mass Ne = 20.18 g; 25°C = 298 K; 50°C = 323 K

\[ 1.25 \text{ g Ne} \times \frac{1 \text{ mol}}{20.18 \text{ g}} = 0.06194 \text{ mol} \]

\[ \frac{P}{V} = \frac{nRT}{V} = \frac{(0.06194 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(10.1 \text{ L})} = 0.150 \text{ atm} \]

\[ \frac{P}{V} = \frac{nRT}{V} = \frac{(0.06194 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(323 \text{ K})}{(10.1 \text{ L})} = 0.163 \text{ atm} \]

58. molar mass O₂ = 32.00 g; 784 mm Hg = 1.032 atm

\[ 4.25 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.1328 \text{ mol} \]

\[ T = \frac{PV}{nR} = \frac{(1.032 \text{ atm})(2.51 \text{ L})}{(0.1328 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})} = 238 \text{ K} = -35°C \]

60. Molar mass of O₂ = 32.00 g; 56.2 kg = 5.62 × 10⁴ g; T = 21°C = 294 K

\[ 5.62 \times 10^4 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 1.76 \times 10^3 \text{ mol O}_2 \]

\[ \frac{P}{V} = \frac{nRT}{V} = \frac{(1.76 \times 10^3 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(294 \text{ K})}{(125 \text{ L})} = 340 \text{ atm} \]

62. molar mass Ar = 39.95 g; 29°C = 302 K; 42°C = 315 K

\[ 1.29 \text{ g Ar} \times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}} = 0.03329 \text{ mol Ar} \]

\[ \frac{P}{V} = \frac{nRT}{V} = \frac{(0.03329 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(302 \text{ K})}{(2.41 \text{ L})} = 0.332 \text{ atm} \]

\[ \frac{P}{V} = \frac{nRT}{V} = \frac{(0.03329 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(315 \text{ K})}{(2.41 \text{ L})} = 0.346 \text{ atm} \]

64. Molar mass of H₂ = 2.016 g; 255 mL = 0.255 L; 100 °C = 373 K

\[ 2.51 \text{ g H}_2 \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 1.245 \text{ mol H}_2 \]

\[ \frac{P}{V} = \frac{nRT}{V} = \frac{(1.245 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(373 \text{ K})}{(0.255 \text{ L})} = 150. \text{ atm} \]

66. As a gas is bubbled through water the bubbles of gas become saturated with water vapor, thus forming a gaseous mixture. The total pressure in a sample of gas that has been collected by bubbling through water is made up of two components; the pressure of the gas of interest and the pressure of water vapor. The partial pressure of the gas of interest is then the total pressure of the sample minus the vapor pressure of water.
68. molar masses: Ne, 20.18 g; Ar, 39.95 g; 27°C = 300 K

\[1.28 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} = 0.06343 \text{ mol Ne}\]

\[2.49 \text{ g Ar} \times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}} = 0.06233 \text{ mol Ar}\]

\[P_{\text{neon}} = \frac{n_{\text{neon}}RT}{V} = \frac{(0.06343 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})}{9.87 \text{ L}} = 0.1582 \text{ atm}\]

\[P_{\text{argon}} = \frac{n_{\text{argon}}RT}{V} = \frac{(0.06233 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})}{9.87 \text{ L}} = 0.1555 \text{ atm}\]

\[P_{\text{total}} = 0.1582 \text{ atm} + 0.1555 \text{ atm} = 0.314 \text{ atm}\]

70. 759 mm Hg = 0.9986 atm; 25°C = 298 K

\[n = \frac{PV}{RT} = \frac{(0.9986 \text{ atm})(2.41 \text{ L})}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 0.0984 \text{ mol}\]

The number of moles of an ideal gas required to fill a given-size container to a particular pressure at a particular temperature does not depend on the specific identity of the gas. Since the neon gas is being measured under the same conditions of pressure, temperature, and volume, then 0.0984 mol of neon will be required.

72. molar masses: He, 4.003 g; Ar, 39.95 g; 273°C = 546 K

\[1.15 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} = 0.2873 \text{ mol He}\]

\[2.91 \text{ g Ar} \times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}} = 0.07284 \text{ mol Ar}\]

\[P_{\text{helium}} = \frac{n_{\text{helium}}RT}{V} = \frac{(0.2873 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(546 \text{ K})}{5.25 \text{ L}} = 2.452 \text{ atm}\]

\[P_{\text{argon}} = \frac{n_{\text{argon}}RT}{V} = \frac{(0.07284 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(546 \text{ K})}{5.25 \text{ L}} = 0.6216 \text{ atm}\]

\[P_{\text{total}} = 0.1582 \text{ atm} + 0.1555 \text{ atm} = 3.07 \text{ atm}\]

74. 1.032 atm = 784.3 mm Hg; molar mass of Zn = 65.38 g

\[P_{\text{hydrogen}} = 784.3 \text{ mm Hg} - 32 \text{ mm Hg} = 752.3 \text{ mm Hg} = 0.990 \text{ atm}\]

\[V = 240 \text{ mL} = 0.240 \text{ L}; \ T = 30^\circ \text{C} + 273 = 303 \text{ K}\]

\[n_{\text{hydrogen}} = \frac{PV}{RT} = \frac{(0.990 \text{ atm})(0.240 \text{ L})}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(303 \text{ K})} = 0.00956 \text{ mol hydrogen}\]

\[0.00956 \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} = 0.00956 \text{ mol of Zn must have reacted}\]
0.00956 mol Zn \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 0.625 \text{ g Zn must have reacted}

76. A theory is successful if it explains known experimental observations. Theories that have been successful in the past may not be successful in the future (for example, as technology evolves, more sophisticated experiments may be possible in the future).

78. pressure

80. no

82. If the temperature of a sample of gas is increased, the average kinetic energy of the particles of gas increases. This means that the speeds of the particles increase. If the particles have a higher speed, they will hit the walls of the container more frequently and with greater force, thereby increasing the pressure.

84. Standard Temperature and Pressure, STP = 0°C, 1 atm pressure. These conditions were chosen because they are easy to attain and reproduce experimentally. The barometric pressure within a laboratory is likely to be near 1 atm most days, and 0°C can be attained with a simple ice bath.

86. Molar mass of Al = 26.98 g

\[1.55 \text{ g Al} \times \frac{1 \text{ mol}}{26.98 \text{ g}} = 0.0574 \text{ mol Al}\]

From the balanced chemical equation, the amount of oxygen needed is

\[0.0574 \text{ mol Al} \times \frac{3 \text{ mol O}_2}{2 \text{ mol Al}} = 0.0861 \text{ mol O}_2\]

\[0.0861 \text{ mol O}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 1.93 \text{ L at STP.}\]

88. Molar mass of Mg = 24.31 g; STP: 1.00 atm, 273 K

\[1.02 \text{ g Mg} \times \frac{1 \text{ mol}}{24.31 \text{ g}} = 0.0420 \text{ mol Mg}\]

As the coefficients for Mg and Cl\(_2\) in the balanced equation are the same, for 0.0420 mol of Mg reacting we will need 0.0420 mol of Cl\(_2\).

\[V = 0.0420 \text{ mol Cl}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 0.941 \text{ L Cl}_2 \text{ at STP.}\]

90. molar mass CaC\(_2\) = 64.10 g; 25 °C = 298 K

\[2.49 \text{ g CaC}_2 \times \frac{1 \text{ mol}}{64.10 \text{ g}} = 0.03885 \text{ mol CaC}_2\]

From the balanced chemical equation for the reaction, 0.03885 mol of CaC\(_2\) reacting completely would generate 0.03885 mol of acetylene, C\(_2\)H\(_2\).

\[V = \frac{nRT}{P} = \frac{(0.03885 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(1.01 \text{ atm})} = 0.941 \text{ L}\]

\[V = \frac{nRT}{P} = \frac{(0.03885 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(273 \text{ K})}{(1.00 \text{ atm})} = 0.870 \text{ L at STP}\]
92. Molar mass of Mg\textsubscript{3}N\textsubscript{2} = 100.95 g; \(T = 24 \, ^\circ\text{C} = 297 \, \text{K}; \, P = 752 \, \text{mm Hg} = 0.989 \, \text{atm}\)

\[
10.3 \, \text{g Mg}_3\text{N}_2 \times \frac{1 \, \text{mol}}{100.95 \, \text{g}} = 0.102 \, \text{mol Mg}_3\text{N}_2
\]

From the balanced chemical equation, the amount of NH\textsubscript{3} produced will be

\[
0.102 \, \text{mol Mg}_3\text{N}_2 \times \frac{2 \, \text{mol NH}_3}{1 \, \text{mol Mg}_3\text{N}_2} = 0.204 \, \text{mol NH}_3
\]

\[
V = \frac{nRT}{P} = \frac{(0.204 \, \text{mol})(0.08206 \, \text{L atm mol}^{-1}\, \text{K}^{-1})(297 \, \text{K})}{(0.989 \, \text{atm})} = 5.03 \, \text{L}
\]

This assumes that the ammonia was collected dry.

94. Molar masses: O\textsubscript{2}, 32.00 g; N\textsubscript{2}, 28.02 g; \(T = 35 \, ^\circ\text{C} = 308 \, \text{K}; \, P = 755 \, \text{mm Hg} = 0.993 \, \text{atm}\)

\[
26.2 \, \text{g O}_2 \times \frac{1 \, \text{mol O}_2}{32.00 \, \text{g O}_2} = 0.819 \, \text{mol O}_2
\]

\[
35.1 \, \text{g N}_2 \times \frac{1 \, \text{mol N}_2}{28.02 \, \text{g N}_2} = 1.25 \, \text{mol N}_2
\]

Total moles = 0.819 mol + 1.25 mol = 2.07 mol

\[
V = \frac{nRT}{P} = \frac{(2.07 \, \text{mol})(0.08206 \, \text{L atm mol}^{-1}\, \text{K}^{-1})(308 \, \text{K})}{(0.993 \, \text{atm})} = 52.7 \, \text{L}
\]

96. \(P_1 = 1.47 \, \text{atm} \quad P_2 = 1.00 \, \text{atm} \) (Standard Pressure)

\(V_1 = 145 \, \text{mL} \quad V_2 = ? \, \text{mL}\)

\(T_1 = 44^\circ \text{C} = 317 \, \text{K} \quad T_2 = 0^\circ \text{C} = 273 \, \text{K} \) (Standard Temperature)

\[
V_2 = \frac{T_2PV_1}{T_1P_2} = \frac{(273 \, \text{K})(1.47 \, \text{atm})(145 \, \text{mL})}{(317 \, \text{K})(1.00 \, \text{atm})} = 184 \, \text{mL}
\]

98. Molar masses: He, 4.003 g; Ne, 20.18 g

\[
6.25 \, \text{g He} \times \frac{1 \, \text{mol He}}{4.003 \, \text{g He}} = 1.561 \, \text{mol He}
\]

\[
4.97 \, \text{g Ne} \times \frac{1 \, \text{mol Ne}}{20.18 \, \text{g Ne}} = 0.2463 \, \text{mol Ne}
\]

\(n_{\text{total}} = 1.561 \, \text{mol} + 0.2463 \, \text{mol} = 1.807 \, \text{mol}\)

As 1 mol of an ideal gas occupies 22.4 L at STP, the volume is given by

\[
1.807 \, \text{mol} \times \frac{22.4 \, \text{L}}{1 \, \text{mol}} = 40.48 \, \text{L} = 40.5 \, \text{L}.
\]

The partial pressure of a given gas in a mixture will be proportional to what fraction of the total number of moles of gas the given gas represents

\[
P_{\text{He}} = \frac{1.561 \, \text{mol He}}{1.807 \, \text{mol total}} \times 1.00 \, \text{atm} = 0.8639 \, \text{atm} = 0.864 \, \text{atm}
\]
\[ P_{\text{Ne}} = \frac{0.2463 \text{ mol Ne}}{1.807 \text{ mol total}} \times 1.00 \text{ atm} = 0.1363 \text{ atm} = 0.136 \text{ atm} \]

100. \[ 2\text{C}_2\text{H}_2(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) + 4\text{CO}_2(\text{g}) \]

Molar mass \( \text{C}_2\text{H}_2 = 26.04 \text{ g} \)

\[ 1.00 \text{ g} \text{C}_2\text{H}_2 \times \frac{1 \text{ mol}}{26.04 \text{ g}} = 0.0384 \text{ mol C}_2\text{H}_2 \]

From the balanced chemical equation, \( 2 \times 0.0384 = 0.0768 \text{ mol of CO}_2 \) will be produced.

\[ 0.0768 \text{ mol CO}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 1.72 \text{ L at STP} \]

102. \( 125 \text{ mL} = 0.125 \text{ L} \)

\[ 0.125 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.00558 \text{ mol H}_2 \]

From the balanced chemical equation, one mole of zinc is required for each mole of hydrogen produced. Therefore, \( 0.00558 \text{ mol of Zn will be required.} \)

\[ 0.00558 \text{ mol Zn} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol}} = 0.365 \text{ g Zn} \]

104. twice

106.

a. \( PV = k; P_1V_1 = P_2V_2 \)

b. \( V = kT; V_1/T_1 = V_2/T_2 \)

c. \( V = an; V_1/n_1 = V_2/n_2 \)

d. \( PV = nRT \)

e. \( P_1V_1/T_1 = P_2V_2/T_2 \)

108. First determine what volume the helium in the tank would have if it were at a pressure of 755 mm Hg (corresponding to the pressure the gas will have in the balloons).

\[ 8.40 \text{ atm} = 6384 \text{ mm Hg} \]

\[ V_2 = (25.2 \text{ L}) \times \frac{6384 \text{ mm Hg}}{755 \text{ mm Hg}} = 213 \text{ L} \]

Allowing for the fact that 25.2 L of He will have to remain in the tank, this leaves \( 213 - 25.2 = 187.8 \text{ L of He for filling the balloons.} \)

\[ 187.8 \text{ L He} \times \frac{1 \text{ balloon}}{1.50 \text{ L He}} = 125 \text{ balloons} \]
110. According to the balanced chemical equation, when 1 mol of \((\text{NH}_4)_2\text{CO}_3\) reacts, a total of 4 moles of gaseous substances is produced.

\[
\text{molar mass } (\text{NH}_4)_2\text{CO}_3 = 96.09 \text{ g; } 453 \, ^\circ\text{C} = 726 \text{ K}
\]

\[
52.0 \text{ g} \times \frac{1 \text{ mol}}{96.09 \text{ g}} = 0.541 \text{ mol}
\]

As 0.541 mol of \((\text{NH}_4)_2\text{CO}_3\) reacts, \(4(0.541) = 2.16 \text{ mol of gaseous products result.}\)

\[
V = \frac{nRT}{P} = \frac{(2.16 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(726 \text{ K})}{(1.04 \text{ atm})} = 124 \text{ L}
\]

112. \(\text{CaCO}_3(s) + 2\text{H}^+(aq) \rightarrow \text{Ca}^{2+}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)\)

\[
\text{molar mass } \text{CaCO}_3 = 100.1 \text{ g; } 60^\circ\text{C} + 273 = 333 \text{ K}
\]

\[
10.0 \text{ g CaCO}_3 \times \frac{1 \text{ mol}}{100.1 \text{ g}} = 0.0999 \text{ mol CaCO}_3 = 0.0999 \text{ mol CO}_2 \text{ also}
\]

\[
P_{\text{carbon dioxide}} = P_{\text{total}} - P_{\text{water vapor}}
\]

\[
P_{\text{carbon dioxide}} = 774 \text{ mm Hg} - 149.4 \text{ mm Hg} = 624.6 \text{ mm Hg} = 0.822 \text{ atm}
\]

\[
V_{\text{wet}} = \frac{nRT}{P} = \frac{(0.0999 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(333 \text{ K})}{(0.822 \text{ atm})} = 3.32 \text{ L wet CO}_2
\]

\[
V_{\text{dry}} = 3.32 \text{ L} \times \frac{624.6 \text{ mm Hg}}{774 \text{ mm Hg}} = 2.68 \text{ L}
\]

114. \(2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)\)

\[
\text{molar mass } \text{KClO}_3 = 122.6 \text{ g; } 25^\circ\text{C} + 273 = 298 \text{ K; } 630. \text{ torr} = 0.829 \text{ atm}
\]

\[
50.0 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.6 \text{ g KClO}_3} = 0.408 \text{ mol KClO}_3
\]

\[
0.408 \text{ mol KClO}_3 \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = 0.612 \text{ mol O}_2
\]

\[
V = \frac{nRT}{P} = \frac{(0.612 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(0.829 \text{ atm})} = 18.1 \text{ L O}_2
\]

116.

a. \(752 \text{ mm Hg} \times \frac{101,325 \text{ Pa}}{760 \text{ mm Hg}} = 1.00 \times 10^5 \text{ Pa}\)

b. \(458 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 4.52 \text{ atm}\)

c. \(1.43 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 1.09 \times 10^3 \text{ mm Hg}\)

d. \(842 \text{ torr} = 842 \text{ mm Hg}\)
118.

a. \[645 \text{ mm Hg} \times \frac{101,325 \text{ Pa}}{760 \text{ mm Hg}} = 8.60 \times 10^4 \text{ Pa}\]

b. \[221 \text{ kPa} = 221 \times 10^3 \text{ Pa} = 2.21 \times 10^5 \text{ Pa}\]

c. \[0.876 \text{ atm} \times \frac{101,325 \text{ Pa}}{1 \text{ atm}} = 8.88 \times 10^4 \text{ Pa}\]

d. \[32 \text{ torr} \times \frac{101,325 \text{ Pa}}{760 \text{ torr}} = 4.3 \times 10^3 \text{ Pa}\]

120.

a. \[1.00 \text{ mm Hg} = 1.00 \text{ torr}\]
\[V = 255 \text{ mL} \times \frac{1.00 \text{ torr}}{2.00 \text{ torr}} = 128 \text{ mL}\]

b. \[1.0 \text{ atm} = 101.325 \text{ kPa}\]
\[V = 1.3 \text{ L} \times \frac{1.0 \text{ kPa}}{101.325 \text{ kPa}} = 1.3 \times 10^{-2} \text{ L}\]

c. \[1.0 \text{ mm Hg} = 0.133 \text{ kPa}\]
\[V = 1.3 \text{ L} \times \frac{1.0 \text{ kPa}}{0.133 \text{ kPa}} = 9.8 \text{ L}\]

122. \[1.52 \text{ L} = 1.52 \times 10^3 \text{ mL}\]
\[755 \text{ mm Hg} \times \frac{1.52 \times 10^3 \text{ mL}}{450 \text{ mL}} = 2.55 \times 10^3 \text{ mm Hg}\]

124.

a. \[74 \degree C + 273 = 347 \text{ K}; -74 \degree C + 273 = 199 \text{ K}\]
\[100 \text{ mL} \times \frac{199 \text{ K}}{347 \text{ K}} = 57.3 \text{ mL}\]

b. \[100 \degree C + 273 = 373 \text{ K}\]
\[373 \text{ K} \times \frac{600 \text{ mL}}{500 \text{ mL}} = 448 \text{ K} (175 \degree C)\]

c. zero (the volume of any gas sample becomes zero at 0 K)

126. \[12 \degree C + 273 = 285 \text{ K}; 192 \degree C + 273 = 465 \text{ K}\]
\[75.2 \text{ mL} \times \frac{465 \text{ K}}{285 \text{ K}} = 123 \text{ mL}\]
128. For a given gas, the number of moles present in a sample is directly proportional to the mass of the sample. The problem, therefore, can be solved even though the gas is not identified (so that its molar mass is not known).

\[ 23.2 \text{ g} \times \frac{10.4 \text{ L}}{93.2 \text{ L}} = 2.59 \text{ g} \]

130.

a. \[ V = 21.2 \text{ mL} = 0.0212 \text{ L} \]

\[ T = \frac{PV}{nR} = \frac{(1.034 \text{ atm})(0.0212 \text{ L})}{(0.00432 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})} = 61.8 \text{ K} \]

b. \[ V = 1.73 \text{ mL} = 0.00173 \text{ L} \]

\[ P = \frac{nRT}{V} = \frac{(0.000115 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(182 \text{ K})}{(0.00173 \text{ L})} = 0.993 \text{ atm} \]

c. \[ P = 1.23 \text{ mm Hg} = 0.00162 \text{ atm}; \quad T = 152 ^\circ \text{C} + 273 = 425 \text{ K} \]

\[ V = \frac{nRT}{P} = \frac{(0.773 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(425 \text{ K})}{(0.00162 \text{ atm})} = 1.66 \times 10^4 \text{ L} \]

132. \[ 27 ^\circ \text{C} + 273 = 300 \text{ K} \]

The number of moles of gas it takes to fill the 100 L tanks to 120 atm at 27 \(^\circ\)C is independent of the identity of the gas.

\[ n = \frac{PV}{RT} = \frac{(120 \text{ atm})(100. \text{ L})}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})} = 487 \text{ mol} \]

487 mol of \textit{any} gas will fill the tanks to the required specifications.

molar masses: \text{CH}_4, 16.0 \text{ g}; \text{N}_2, 28.0 \text{ g}; \text{CO}_2, 44.0 \text{ g} \\
for \text{CH}_4: (487 \text{ mol})(16.0 \text{ g/mol}) = 7792 \text{ g} = 7.79 \text{ kg CH}_4 \\\nfor \text{N}_2: (487 \text{ mol})(28.0 \text{ g/mol}) = 13,636 \text{ g} = 13.6 \text{ kg N}_2 \\\nfor \text{CO}_2: (487 \text{ mol})(44.0 \text{ g/mol}) = 21,428 \text{ g} = 21.4 \text{ kg CO}_2 \]

134. molar mass of \text{O}_2 = 32.00 \text{ g}; 55 \text{ mg} = 0.055 \text{ g} \\
\[ n = 0.055 \text{ g} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.0017 \text{ mol} \]

\[ V = 100. \text{ mL} = 0.100 \text{ L}; \quad T = 26 ^\circ \text{C} + 273 = 299 \text{ K} \]

\[ P = \frac{nRT}{V} = \frac{(0.0017 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(299 \text{ K})}{(0.100 \text{ L})} = 0.42 \text{ atm} \]
136. \( P_1 = 1.13 \text{ atm} \quad P_2 = 1.89 \text{ atm} \n\)
\( V_1 = 100 \text{ mL} = 0.100 \text{ L} \quad V_2 = 500 \text{ mL} = 0.500 \text{ L} \n\)
\( T_1 = 300 \text{ K} \quad T_2 = ? \n\)
\[ T_2 = \frac{T_1 P_1 V_1}{P_1 V_1} = \frac{(300 \text{ K})(1.89 \text{ atm})(0.500 \text{ L})}{(1.13 \text{ atm})(0.100 \text{ L})} = 2.51 \times 10^3 \text{ K} \]

Note that the calculation could have been carried through with the two volumes expressed in milliliters because the universal gas constant does not appear explicitly in this form of the ideal gas equation.

138. molar masses: \( \text{N}_2, 28.02 \text{ g} \); \( \text{He}, 4.003 \text{ g} \); STP: 1.00 atm, 273 K
\[ 12.1 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 0.432 \text{ mol N}_2 \]
\[ 4.05 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} = 1.01 \text{ mol He} \]
Total moles of gas = 0.432 mol + 1.01 mol = 1.44 mol
\[ V = \frac{nRT}{P} = \frac{(1.44 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(273 \text{ K})}{(1.00 \text{ atm})} = 32.3 \text{ L} \]

140. \( \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \)

molar mass of \( \text{NH}_3 = 17.03 \text{ g} \); 11 °C + 273 = 284 K
\[ 5.00 \text{ g } \text{NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} = 0.294 \text{ mol NH}_3 \text{ to be produced} \]
\[ 0.294 \text{ mol NH}_3 \times \frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3} = 0.147 \text{ mol N}_2 \text{ required} \]
\[ 0.294 \text{ mol NH}_3 \times \frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} = 0.441 \text{ mol H}_2 \text{ required} \]
\[ V_{\text{nitrogen}} = \frac{nRT}{P} = \frac{(0.147 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(284 \text{ K})}{(0.998 \text{ atm})} = 3.43 \text{ L N}_2 \]
\[ V_{\text{hydrogen}} = \frac{nRT}{P} = \frac{(0.441 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(284 \text{ K})}{(0.998 \text{ atm})} = 10.3 \text{ L H}_2 \]

142. \( 2\text{Cu}_2\text{S(s)} + 3\text{O}_2(g) \rightarrow 2\text{Cu}_2\text{O(s)} + 2\text{SO}_2(g) \)

molar mass \( \text{Cu}_2\text{S} = 159.2 \text{ g} \); 27.5 °C + 273 = 301 K
\[ 25 \text{ g Cu}_2\text{S} \times \frac{1 \text{ mol Cu}_2\text{S}}{159.2 \text{ g Cu}_2\text{S}} = 0.1570 \text{ mol Cu}_2\text{S} \]
\[ 0.1570 \text{ mol Cu}_2\text{S} \times \frac{3 \text{ mol O}_2}{2 \text{ mol Cu}_2\text{S}} = 0.2355 \text{ mol O}_2 \]
\[ V_{\text{oxygen}} = \frac{nRT}{P} = \frac{(0.2355 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(301 \text{ K})}{(0.998 \text{ atm})} = 5.8 \text{ L O}_2 \]

0.1570 mol Cu\textsubscript{2}S \times \frac{2 \text{ mol SO}_2}{2 \text{ mol Cu}_2\text{S}} = 0.1570 \text{ mol SO}_2

\[ V_{\text{sulfur dioxide}} = \frac{nRT}{P} = \frac{(0.1570 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(301 \text{ K})}{(0.998 \text{ atm})} = 3.9 \text{ L SO}_2 \]

144. One mole of any ideal gas occupies 22.4 L at STP.

\[ 35 \text{ mol N}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 7.8 \times 10^2 \text{ L} \]

146. molar masses: He, 4.003 g; Ar, 39.95 g; Ne, 20.18 g

\[ 5.0 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} = 1.249 \text{ mol He} \]

\[ 1.0 \text{ g Ar} \times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}} = 0.02503 \text{ mol Ar} \]

\[ 3.5 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} = 0.1734 \text{ mol Ne} \]

Total moles of gas = 1.249 + 0.02503 + 0.1734 = 1.447 mol

22.4 L is the volume occupied by one mole of any ideal gas at STP. This would apply even if the gas sample is a mixture of individual gases.

\[ 1.447 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 32 \text{ L total volume for the mixture} \]

The partial pressure of each individual gas in the mixture will be related to what fraction on a mole basis each gas represents in the mixture.

\[ P_{\text{He}} = 1.00 \text{ atm} \times \frac{1.249 \text{ mol He}}{1.447 \text{ mol total}} = 0.86 \text{ atm} \]

\[ P_{\text{Ar}} = 1.00 \text{ atm} \times \frac{0.02503 \text{ mol Ar}}{1.447 \text{ mol total}} = 0.017 \text{ atm} \]

\[ P_{\text{Ne}} = 1.00 \text{ atm} \times \frac{0.1734 \text{ mol Ne}}{1.447 \text{ mol total}} = 0.12 \text{ atm} \]
148. The solution is only 50% H₂O₂. Therefore 125 g solution = 62.5 g H₂O₂

molar mass of H₂O₂ = 34.02 g; \( T = 27^\circ C = 300 \text{ K}; \ P = 764 \text{ mm Hg} = 1.01 \text{ atm} \)

\[
62.5 \text{ g H}_2\text{O}_2 \times \frac{1 \text{ mol}}{34.02 \text{ g}} = 1.84 \text{ mol H}_2\text{O}_2
\]

\[
1.84 \text{ mol H}_2\text{O}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}_2} = 0.920 \text{ mol O}_2
\]

\[
V = \frac{nRT}{P} = \frac{(0.920 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})}{(1.01 \text{ atm})} = 22.4 \text{ L}
\]
CHAPTER 14

Liquids and Solids

CHAPTER ANSWERS

2. Water exerts its cooling effect in nature in many ways. Water, as perspiration, helps cool the human body (the evaporation of water from skin is an endothermic process; the heat required for evaporation comes from the body). Large bodies of natural water (e.g., the oceans) have a cooling effect on nearby land masses (the interior of the United States, away from the oceans, tends to be hotter than coastal regions). In industry, water is used as a coolant in many situations. Some nuclear power plants, for example, use water to cool the reactor core. Many office buildings are air-conditioned in summer by circulating cold water systems.

4. Since it requires so much more energy to vaporize water than to melt ice, this suggests that the gaseous state is significantly different from the liquid state, but that the liquid and solid states are relatively similar.

6. See Figure 14.2.

8. When a solid is heated, the molecules begin to vibrate/move more quickly. When enough energy has been added to overcome the intermolecular forces that hold the molecules in a crystal lattice, the solid melts. As the liquid is heated, the molecules begin to move more quickly and more randomly. When enough energy has been added, molecules having sufficient kinetic energy will begin to escape from the surface of the liquid. Once the pressure of vapor coming from the liquid is equal to the pressure above the liquid, the liquid boils. Only intermolecular forces need to be overcome in this process: no chemical bonds are broken.

10. intramolecular; intermolecular

12. fusion; vaporization
14.  
   a. More energy is required to separate the molecules of a liquid into the freely-moving and widely separated molecules of a vapor/gas.

   b. \( 1.00 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{2.60 \text{ kJ}}{1 \text{ mol Na}} = 0.113 \text{ kJ} \)

   c. \( 1.00 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{97.0 \text{ kJ}}{1 \text{ mol Na}} = 4.22 \text{ kJ} \)

   d. \(-4.22 \text{ kJ}\) (The sign is changed for the opposite process.)

16. molar mass \( \text{H}_2\text{O} = 18.02 \text{ g} \)

   melt: \( 25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.02 \text{ kJ}}{1 \text{ mol H}_2\text{O}} = 8.35 \text{ kJ} \)

   vaporize: \( 37.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{40.6 \text{ kJ}}{1 \text{ mol H}_2\text{O}} = 84.4 \text{ kJ} \)

   heating liquid: \( 55.2 \text{ g H}_2\text{O} \times \frac{4.184 \text{ J/g °C}}{100 \text{ °C}} = 23,073 \text{ J} = 23.1 \text{ kJ} \)

18. The molar heat of fusion of aluminum is the heat required to melt 1 mol.

   \( 3.95 \text{ kJ} \div 26.98 \text{ g Al} = 107 \text{ kJ/mol} \)

   \( 10.0 \text{ g Al} \times 3.95 \text{ kJ/g} = 39.5 \text{ kJ} \text{ required to melt 10.0 g Al} \)

   \( 10.0 \text{ mol Al} \times 107 \text{ kJ/mol} = 1.07 \times 10^3 \text{ kJ} \text{ required to melt 10.0 mol Al} \)

20. Dipole-dipole forces are relatively stronger at short distances; they are short-range forces. Molecules must first closely approach one another before dipole-dipole forces can cause attraction among molecules.

22. Water molecules are able to form strong hydrogen bonds with each other. These bonds are an especially strong form of dipole-dipole forces and are possible only when hydrogen atoms are bonded to the most electronegative elements (N, O, and F). The particularly strong intermolecular forces in \( \text{H}_2\text{O} \) require much higher temperatures (higher energies) to be overcome in order to permit the liquid to boil. We take the fact that water has a much higher boiling point than the other hydrogen compounds of the Group 6 elements as proof that a special force is at play in water (hydrogen bonding).

24. London dispersion forces are instantaneous dipole forces that arise when the electron cloud of an atom is momentarily distorted by a nearby dipole, temporarily separating the centers of positive and negative charge in the atom.

26.  
   a. London dispersion forces (non polar molecules)

   b. hydrogen bonding (H bonded to N)

   c. London dispersion forces (nonpolar molecules)

   d. dipole-dipole forces (polar molecules)
28. An increase in the heat of fusion is observed for an increase in the size of the halogen atom involved (the electron cloud of a larger atom is more easily polarized by an approaching dipole, thus giving larger London dispersion forces).

30. For a homogeneous mixture to be able to form at all, the forces between molecules of the two substances being mixed must be at least comparable in magnitude to the intermolecular forces within each separate substance. Apparently in the case of a water-ethanol mixture, the forces that exist when water and ethanol are mixed are stronger than water-water or ethanol-ethanol forces in the separate substances. This allows ethanol and water molecules to approach each other more closely in the mixture than either substance’s molecules could approach a like molecule in separate substances. There is strong hydrogen bonding in both ethanol and water.

32. Vapor pressure is the pressure of vapor present at equilibrium above a liquid in a sealed container at a particular temperature. When a liquid is placed in a closed container, molecules of the liquid evaporate freely into the empty space above the liquid. As the number of molecules present in the vapor state increases with time, vapor molecules begin to rejoin the liquid state (condense). Eventually a dynamic equilibrium between evaporation and condensation is reached in which the net number of molecules present in the vapor phase becomes constant with time.

34. A liquid is injected at the bottom of the column of mercury and rises to the surface of the mercury where the liquid evaporates into the vacuum above the mercury column. As the liquid evaporates, the pressure of vapor increases in the space above the mercury, and presses down on the mercury. The level of mercury, therefore, drops, and the amount by which the mercury level drops (in mm Hg) is equivalent to the vapor pressure of the liquid.

In the picture, the left tube represents a barometer—a tube of mercury inverted into a dish of mercury with a vacuum above the mercury column; the height of the mercury column represents the pressure of the atmosphere. In the remaining three tubes, liquids of different volatilities are admitted to the bottom of the tube of mercury. They rise through the mercury and evaporate into the vacuum above the column of mercury. As the pressure of vapor builds up above the mercury column, the height of the mercury in the tube drops. Note that diethyl ether, (C₂H₅)₂O, shows the highest vapor pressure because it is the most volatile of the three liquids.

36.

a. HF: Although both substances are capable of hydrogen bonding, water has two O–H bonds that can be involved in hydrogen bonding versus only one F–H bond in HF.
b. CH₃OCH₃: As there is no H attached to the O atom, no hydrogen bonding can exist. Moreover, as there is no hydrogen bonding possible, the molecule should be relatively more volatile than CH₃CH₂OH even though it contains the same number of atoms of each element.

c. CH₃SH: Hydrogen bonding is not as important for a S–H bond (because S has a lower electronegativity than O). As there is little hydrogen bonding, the molecule is relatively more volatile than CH₃OH.

38. Both substances have the same molar mass. However, ethyl alcohol contains a hydrogen atom directly bonded to an oxygen atom. Therefore, hydrogen bonding can exist in ethyl alcohol whereas only weak dipole-dipole forces can exist in dimethyl ether. Dimethyl ether is more volatile; ethyl alcohol has a higher boiling point.

40. **Ionic** solids have as their fundamental particles positive and negative **ions**; a simple example is sodium chloride, in which Na⁺ and Cl⁻ ions are held together by strong electrostatic forces. **Molecular** solids have molecules as their fundamental particles, with the molecules being held together in the crystal by dipole-dipole forces, hydrogen bonding forces, or London dispersion forces (depending on the identity of the substance). Simple examples of molecular solids include ice (H₂O) and ordinary table sugar (sucrose).

**Atomic** solids have simple atoms as their fundamental particles, with the atoms being held together in the crystal either by covalent bonding (as in graphite or diamond) or by metallic bonding (as in copper or other metals).

42. The interparticle forces in ionic solids (the ionic bond) are much stronger than the interparticle forces in molecular solids (dipole-dipole forces, London forces, etc.). The difference in intermolecular forces is most clearly shown in the great differences in melting points and boiling points between ionic and molecular solids. For example, table salt and ordinary sugar are both crystalline solids that appear very similar. Yet sugar can be melted easily in a saucepan during the making of candy whereas even the full heat of a stove will not melt salt.

44. Ionic solids consist of a crystal lattice of basically alternating positively and negatively charged ions. A given ion is surrounded by several ions of the opposite charge, all of which electrostatically attract it strongly. This pattern repeats itself throughout the crystal. The existence of these strong electrostatic forces throughout the crystal means a great deal of energy must be applied to overcome the forces and melt the solid.

46. Ordinary ice contains nonlinear, highly polar water molecules. In addition, there is extensive, strong hydrogen bonding possible between water molecules in ordinary ice. Dry ice, on the other hand, consists of linear, nonpolar carbon dioxide molecules, and only very weak intermolecular forces are possible.

48. Although ions exist in the solid, liquid, or dissolved states, in the solid state the ions are rigidly held in place in the crystal lattice and cannot **move** so as to conduct an electrical current.

50. The presence of atoms of a second element in a given metal’s crystal lattice changes the properties of the metal. Frequently the alloy is stronger than either of the original components because the irregularities introduced into the crystal lattice by the presence of a second element’s atoms prevent the crystal from being deformed as easily. The properties of iron may be modified by alloying it with many different substances, particularly with carbon, nickel, and cobalt. Steels with relatively high carbon content are exceptionally strong whereas steels with low carbon content are softer, more malleable, and more ductile.

52. j
54. f
56. d
58. a
60. l

62. Dimethyl ether has the larger vapor pressure. No hydrogen bonding is possible because the O atom does not have a hydrogen atom attached. Hydrogen bonding can occur only when a hydrogen atom is directly attached to a strongly electronegative atom (such as N, O, or F). Hydrogen bonding is possible in ethanol (ethanol contains an \(\cdot\)OH group).

64.
   a. \(\text{H}_2\): London dispersion forces are the only intermolecular forces present in these nonpolar molecules. Typically London forces become larger with increasing atomic size (As the atoms become bigger, the edge of the electron cloud lies farther from the nucleus and becomes more easily distorted.).
   b. \(\text{Xe}\): Only the relatively weak London forces could exist in a crystal of Xe atoms whereas in NaCl strong ionic forces exist. In diamond, strong covalent bonding exists between carbon atoms.
   c. \(\text{Cl}_2\): Only London forces exist among such nonpolar molecules. London forces become larger with increasing atomic size.

66. Steel is a general term applied to alloys consisting primarily of iron but with small amounts of other substances added. Whereas pure iron itself is relatively soft, malleable, and ductile, steels are typically much stronger and harder and much less subject to damage.

68. Water is the solvent in which cellular processes take place in living creatures. Water in the oceans moderates the Earth’s temperature. Water is used in industry as a cooling agent. Water serves as a means of transportation on the Earth’s oceans. The liquid range is 0°C to 100°C at 1 atm pressure.

70. At higher altitudes, the boiling points of liquids such as water are lower because there is a lower atmospheric pressure above the liquid. The temperature at which food cooks is determined by the temperature to which the water in the food can be heated before it escapes as steam. Thus food cooks at a lower temperature at high elevations where the boiling point of water is lowered.

72. Heat of fusion (melt); heat of vaporization (boil).

   The heat of vaporization is always larger because virtually all of the intermolecular forces must be overcome to form a gas. In a liquid, considerable intermolecular forces remain. Thus going from a solid to liquid requires less energy than going from the liquid to the gas.

74. Dipole-dipole interactions are typically about 1% as strong as a covalent bond. Dipole-dipole interactions represent electrostatic attractions between portions of molecules that carry only a partial positive or negative charge. Such forces require the molecules that are interacting to come near enough to each other.

76. London dispersion forces are relatively weak forces that arise among noble gas atoms and in nonpolar molecules. London forces are due to instantaneous dipoles that develop when one atom (or molecule) momentarily distorts the electron cloud of another atom (or molecule). London forces are typically weaker than either permanent dipole-dipole forces or covalent bonds.

78. For every mole of liquid water that evaporates, several kilojoules of heat must be absorbed to provide kinetic energy to overcome attractive forces among the molecules. This heat is absorbed by the water from its surroundings.
80. In NH₃, strong hydrogen bonding can exist. In CH₄, because the molecule is nonpolar, only the relatively weak London dispersion forces exist.

82. In a crystal of ice, strong hydrogen bonding forces are present whereas in the crystal of a nonpolar substance like oxygen only the much weaker London forces exist.

84. Ice floats on liquid water; water expands when it is frozen.

86. Although they are at the same temperature, steam at 100°C contains a larger amount of energy (equal to the heat of vaporization of water) than hot water.

88. Hydrogen bonding is a special case of dipole-dipole interactions that occur among molecules containing hydrogen atoms bonded to highly electronegative atoms such as fluorine, oxygen, or nitrogen. The bonds are very polar, and the small size of the hydrogen atom (compared to other atoms) allows the dipoles to approach each other very closely. Examples: H₂O, NH₃, HF, etc.

90. Evaporation and condensation are opposite processes. Evaporation is an endothermic process; condensation is an exothermic process. Evaporation requires an input of energy to provide the increased kinetic energy possessed by the molecules when they are in the gaseous state. Evaporation occurs when the molecules in a liquid are moving fast enough and randomly enough that molecules are able to escape from the surface of the liquid and enter the vapor phase.
CHAPTER 15

Solutions

CHAPTER ANSWERS

2. A heterogeneous mixture does not have a uniform composition; the composition varies in different places within the mixture. Examples of non–homogeneous mixtures include salad dressing (mixture of oil, vinegar, water, herbs, and spices) and granite (combination of minerals).

4. solid

6. “Like dissolves like.” The hydrocarbons in oil have intermolecular forces that are very different from those in water, and so the oil spreads out rather than dissolving in the water.

8. Molecules with similar intermolecular forces tend to mix together. Polar molecules mix with polar molecules. Nonpolar molecules mix with nonpolar molecules. Examples depend on student choices.

10. unsaturated

12. large

14. 100.0

16.

a. \( \frac{0.00224 \text{ g KI}}{0.00224 \text{ g KI} + 10.0 \text{ g water}} \times 100 = 0.0224\% \)

b. \( \frac{2.24 \text{ g KI}}{2.24 \text{ g KI} + 10.0 \text{ g water}} \times 100 = 18.3\% \)

c. \( \frac{2.24 \text{ g KI}}{2.24 \text{ g KI} + 1000 \text{ g water}} \times 100 = 0.223\% \)

d. \( \frac{224 \text{ g KI}}{224 \text{ g KI} + 1000 \text{ g water}} \times 100 = 18.3\% \)

18.

a. \( 125 \text{ g solution} \times \frac{2.49 \text{ g NaCl}}{100 \text{ g solution}} = 3.11 \text{ g NaCl} \)

   \( 125 \text{ g solution} - 3.11 \text{ g NaCl} = 121.89 \text{ g water} (122 \text{ g water}) \)

b. \( 35.2 \text{ mg solution} \times \frac{4.95 \text{ mg NaCl}}{100 \text{ mg solution}} = 1.74 \text{ mg NaCl} \)

   \( 35.2 \text{ mg solution} - 1.74 \text{ mg NaCl} = 33.46 \text{ mg water} (33.5 \text{ mg water}) \)
c. \[ 1000. \text{ g solution} \times \frac{6.21 \text{ g NaCl}}{100. \text{ g solution}} = 62.1 \text{ g NaCl} \]

1000. g solution – 62.1 g NaCl = 937.9 g water (938 g water)

d. \[ 29.2 \text{ g solution} \times \frac{0.100 \text{ g NaCl}}{100. \text{ g solution}} = 0.0292 \text{ g NaCl} \]

29.2 g solution – 0.0292 g NaCl = 29.17 g water (29.2 g water)

20. \% Cu = \[ \frac{71.2 \text{ g Cu}}{(71.2 \text{ g} + 28.7 \text{ g} + 1.03 \text{ g})} \times 100 = 70.5\% \text{ Cu} \]

\% Zn = \[ \frac{28.7 \text{ g Zn}}{(71.2 \text{ g} + 28.7 \text{ g} + 1.03 \text{ g})} \times 100 = 28.4\% \text{ Zn} \]

\% Sn = \[ \frac{1.03 \text{ g Sn}}{(71.2 \text{ g} + 28.7 \text{ g} + 1.03 \text{ g})} \times 100 = 1.02\% \text{ Sn} \]

22. \[ \frac{67.1 \text{ g CaCl}_2}{(67.1 \text{ g CaCl}_2 + 275 \text{ g water})} \times 100 = 19.6\% \text{ CaCl}_2 \]

24. To say that the solution is 6.25\% KBr by mass, means that 100. g of the solution will contain 6.25 g KBr.

125 g solution \times \frac{6.25 \text{ g KBr}}{100. \text{ g solution}} = 7.81 \text{ g KBr}

26. molar mass O\text{2} = 32.00 \text{ g}

1.00 g O\text{2} \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.03125 \text{ mol O}

From the balanced chemical equation, it will take 2(0.03125) = 0.0625 mol H\text{2O}_2 to produce this quantity of oxygen.

molar mass H\text{2O}_2 = 34.02 \text{ g}

0.0625 mol H\text{2O}_2 \times \frac{34.02 \text{ g H\text{2O}_2}}{1 \text{ mol H\text{2O}_2}} = 2.13 \text{ g H\text{2O}_2}

2.13 g H\text{2O}_2 \times \frac{100. \text{ g solution}}{3 \text{ g H\text{2O}_2}} \approx 71 \text{ g}

28. \[ 1000 \text{ g} \times \frac{0.95 \text{ g stabilizer}}{100. \text{ g}} = 9.5 \text{ g} \]

30. 0.110 mol; 0.220 mol

32. The molarity represents the number of moles of solute per liter of solution; choice b is the only scenario that fulfills this definition.
34. Molarity = \frac{\text{moles of solute}}{\text{liters of solution}}

a. 250 mL = 0.25 L

\[ M = \frac{0.50 \text{ mol KBr}}{0.25 \text{ L}} = 2.0 \text{ M} \]

b. 500 mL = 0.50 L

\[ M = \frac{0.50 \text{ mol KBr}}{0.50 \text{ L}} = 1.0 \text{ M} \]

c. 750 mL = 0.75 L

\[ M = \frac{0.50 \text{ mol KBr}}{0.75 \text{ L}} = 0.67 \text{ M} \]

d. \[ M = \frac{0.50 \text{ mol KBr}}{1.0 \text{ L}} = 0.50 \text{ M} \]

36. Molarity = \frac{\text{moles of solute}}{\text{liters of solution}}

a. molar mass of KNO₃ = 101.11 g

\[ 1.25 \text{ g KNO}_3 \times \frac{1 \text{ mol}}{101.11 \text{ g}} = 0.01236 \text{ mol KNO}_3 \]

\[ M = \frac{0.01236 \text{ mol}}{0.115 \text{ L}} = 0.108 \text{ M} \]

b. 12.5 g KNO₃ \times \frac{1 \text{ mol}}{101.11 \text{ g}} = 0.1236 \text{ mol KNO}_3

\[ M = \frac{0.1236 \text{ mol}}{0.115 \text{ L}} = 1.075 \text{ M} = 1.08 \text{ M} \]

c. 1.25 mg KNO₃ \times \frac{1 \text{ mol}}{101.11 \text{ g}} = 0.01236 \text{ mmol KNO}_3

\[ M = \frac{0.01236 \text{ mmol}}{1.15 \text{ mL}} = 0.0108 \text{ M} \]

d. 1250 g KNO₃ \times \frac{1 \text{ mol}}{101.11 \text{ g}} = 12.36 \text{ mol KNO}_3

\[ M = \frac{12.36 \text{ mol}}{115 \text{ L}} = 0.108 \text{ M} \]
38. molar mass of NaCl = 58.44 g; 100 mL = 0.100 L

\[
2.71 \text{ g NaCl} \times \frac{1 \text{ mol}}{58.44 \text{ g}} = 0.04637 \text{ mol NaCl}
\]

\[
M = \frac{0.04637 \text{ mol}}{0.100 \text{ L}} = 0.464 \text{ M}
\]

40. molar mass of I₂ = 253.8 g; 225 mL = 0.225 L

\[
5.15 \text{ g I₂} \times \frac{1 \text{ mol}}{253.8 \text{ g}} = 0.0203 \text{ mol I₂}
\]

\[
M = \frac{0.0203 \text{ mol I₂}}{0.225 \text{ L solution}} = 0.0902 \text{ M}
\]

42. molar mass of NaOH = 40.00 g

\[
495 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.0 \text{ g}} = 12.4 \text{ mol NaOH}
\]

\[
M = \frac{12.4 \text{ mol NaOH}}{20.0 \text{ L solution}} = 0.619 \text{ M}
\]

44.

a. molar mass of HNO₃ = 63.02 g; 127 mL = 0.127 L

\[
0.127 \text{ L solution} \times \frac{0.105 \text{ mol HNO₃}}{1.00 \text{ L solution}} = 0.0133 \text{ mol HNO₃}
\]

\[
0.0133 \text{ mol HNO₃} \times \frac{63.02 \text{ g HNO₃}}{1 \text{ mol}} = 0.838 \text{ g HNO₃}
\]

b. molar mass of NH₃ = 17.03 g; 155 mL = 0.155 L

\[
0.155 \text{ L solution} \times \frac{15.1 \text{ mol NH₃}}{1.00 \text{ L solution}} = 2.34 \text{ mol NH₃}
\]

\[
2.34 \text{ mol NH₃} \times \frac{17.03 \text{ g NH₃}}{1 \text{ mol NH₃}} = 39.9 \text{ g NH₃}
\]

c. molar mass KSCN = 97.19 g

\[
2.51 \text{ L solution} \times \frac{2.01 \times 10^{⁻³} \text{ mol KSCN}}{1.00 \text{ L solution}} = 5.05 \times 10^{⁻³} \text{ mol KSCN}
\]

\[
5.05 \times 10^{⁻³} \text{ mol KSCN} \times \frac{97.19 \text{ g KSCN}}{1 \text{ mol}} = 0.490 \text{ g KSCN}
\]
d. molar mass of HCl = 36.46 g  
\[
12.2 \text{ mL} = 0.0122 \text{ L} 
\]
\[
0.0122 \text{ L solution} \times \frac{2.45 \text{ mol HCl}}{1.00 \text{ L solution}} = 0.0299 \text{ mol HCl} 
\]
\[
0.0299 \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol}} = 1.09 \text{ g HCl} 
\]

46.

a. molar mass of KBr = 119.0 g; 173 mL = 0.173 L
\[
0.173 \text{ L solution} \times \frac{1.24 \text{ mol KBr}}{1.00 \text{ L solution}} = 0.215 \text{ mol KBr} 
\]
\[
0.215 \text{ mol KBr} \times \frac{119.0 \text{ g KBr}}{1 \text{ mol}} = 25.6 \text{ g KBr} 
\]

b. molar mass of HCl = 36.46 g
\[
2.04 \text{ L solution} \times \frac{12.1 \text{ mol HCl}}{1.00 \text{ L solution}} = 24.7 \text{ mol HCl} 
\]
\[
24.7 \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol}} = 901 \text{ g HCl} 
\]

c. molar mass of NH₃ = 17.03 g  
25 mL = 0.025 L
\[
0.025 \text{ L solution} \times \frac{3.0 \text{ mol NH}_3}{1.00 \text{ L solution}} = 0.075 \text{ mol NH}_3 
\]
\[
0.075 \text{ mol NH}_3 \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol}} = 1.3 \text{ g NH}_3 
\]

d. molar mass CaCl₂ = 111.0 g; 125 mL = 0.125 L
\[
0.125 \text{ L solution} \times \frac{0.552 \text{ mol CaCl}_2}{1.00 \text{ L solution}} = 0.0690 \text{ mol CaCl}_2 
\]
\[
0.0690 \text{ mol CaCl}_2 \times \frac{111.0 \text{ g CaCl}_2}{1 \text{ mol}} = 7.66 \text{ g CaCl}_2 
\]

48. Molar mass of NaCl = 58.44 g; 1 lb = 453.59 g
\[
453.59 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g}} = 7.76 \text{ mol NaCl} 
\]
\[
7.76 \text{ mol NaCl} \times \frac{1.00 \text{ L solution}}{1.0 \text{ mol NaCl}} = 7.76 \text{ L} = 7.8 \text{ L of solution} 
\]
50.

a. 10.2 mL = 0.0102 L

\[ 0.0102 \text{ L} \times \frac{0.451 \text{ mol AlCl}_3}{1.00 \text{ L}} \times \frac{1 \text{ mol Al}^{3+}}{1 \text{ mol AlCl}_3} = 4.60 \times 10^{-3} \text{ mol Al}^{3+} \]

\[ 0.0102 \text{ L} \times \frac{0.451 \text{ mol AlCl}_3}{1.00 \text{ L}} \times \frac{3 \text{ mol Cl}^-}{1 \text{ mol AlCl}_3} = 1.38 \times 10^{-2} \text{ mol Cl}^- \]

b. 5.51 L = 0.0551 L

\[ 0.00551 \text{ L} \times \frac{0.103 \text{ mol Na}_3\text{PO}_4}{1.00 \text{ L}} \times \frac{3 \text{ mol Na}^+}{1 \text{ mol Na}_3\text{PO}_4} = 1.70 \text{ mol Na}^+ \]

\[ 0.00551 \text{ L} \times \frac{0.103 \text{ mol Na}_3\text{PO}_4}{1.00 \text{ L}} \times \frac{1 \text{ mol PO}_4^{3-}}{1 \text{ mol Na}_3\text{PO}_4} = 0.568 \text{ mol PO}_4^{3-} \]

c. 1.75 mL = 0.00175 L

\[ 0.00175 \text{ L} \times \frac{1.25 \text{ mol CuCl}_2}{1.00 \text{ L}} \times \frac{1 \text{ mol Cu}^{2+}}{1 \text{ mol CuCl}_2} = 2.19 \times 10^{-3} \text{ mol Cu}^{2+} \]

\[ 0.00175 \text{ L} \times \frac{1.25 \text{ mol CuCl}_2}{1.00 \text{ L}} \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol CuCl}_2} = 4.38 \times 10^{-3} \text{ mol Cl}^- \]

d. 25.2 mL = 0.0252 L

\[ 0.0252 \text{ L} \times \frac{0.00157 \text{ mol Ca(OH)}_2}{1.00 \text{ L}} \times \frac{1 \text{ mol Ca}^{2+}}{1 \text{ mol Ca(OH)}_2} = 3.96 \times 10^{-5} \text{ mol Ca}^{2+} \]

\[ 0.0252 \text{ L} \times \frac{0.00157 \text{ mol Ca(OH)}_2}{1.00 \text{ L}} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca(OH)}_2} = 7.91 \times 10^{-5} \text{ mol OH}^- \]

52. Molar mass of Na$_2$CO$_3$ = 106.0 g; 250 mL = 0.250 L

\[ 0.250 \text{ L} \times \frac{0.0500 \text{ mol Na}_2\text{CO}_3}{1.00 \text{ L}} \times \frac{106 \text{ g}}{1 \text{ mol}} = 1.33 \text{ g Na}_2\text{CO}_3 \]

54. 1/2

56. $M_1 \times V_1 = M_2 \times V_2$

a. $M_1 = 0.251 \text{ M}$

\[ V_1 = 10.0 \text{ mL} \]

\[ V_2 = (10.0 + 25.0) = 35.0 \text{ mL} \]

\[ M_2 = \frac{(0.251 \text{ M})(10.0 \text{ mL})}{(35.0 \text{ mL})} = 0.0717 \text{ M} \]

b. $M_1 = 3.00 \text{ M}$

\[ V_1 = 125 \text{ mL} \]

\[ V_2 = (125 + 97.5) = 222.5 \text{ mL} \]

\[ M_2 = \frac{(3.00 \text{ M})(125 \text{ mL})}{(222.5 \text{ mL})} = 1.69 \text{ M} \]
c. \[ M_1 = 0.851 \, M \quad M_2 = ? \]
\[ V_1 = 25.0 \, \text{mL} \quad V_2 = 500. \, \text{mL} \]
\[ M_2 = \frac{(0.851 \, M)(25 \, \text{mL})}{(500. \, \text{mL})} = 0.0426 \, M \]

\[ V_1 = 25.0 \, \text{mL} \quad V_2 = 50.0 \, \text{mL} \]
\[ M_2 = \frac{(1.25 \, M)(25 \, \text{mL})}{(50.0 \, \text{mL})} = 0.625 \, M \]

58. \[ M_1 = 19.4 \, M \quad M_2 = 3.00 \, M \]
\[ V_1 = ? \, \text{mL} \quad V_2 = 3.50 \, \text{L} \]
\[ M_1 = \frac{(3.00 \, M)(3.50 \, \text{mL})}{(19.4 \, M)} = 0.541 \, \text{L} (541 \, \text{mL}) \]

60. \[ M_1 = 0.211 \, M \quad M_2 = ? \]
\[ V_1 = 75 \, \text{mL} \quad V_2 = 125 \, \text{mL} \]
\[ M_2 = \frac{(0.211 \, M)(75 \, \text{mL})}{(125 \, \text{mL})} = 0.127 \, M = 0.13 \, M \]

62. \[ \frac{(100. \, \text{mL})(1.25 \, M)}{(12.1 \, M)} = 10.3 \, \text{mL} \]

64. \[ \text{Na}_2\text{CO}_3(aq) + \text{CaCl}_2(aq) \rightarrow \text{CaCO}_3(s) + 2\text{NaCl}(s) \]

\[
\text{mmole Ca}^{2+} \text{ ion: } 37.2 \, \text{mL} \times \frac{0.105 \, \text{mmol Ca}^{2+}}{1.00 \, \text{mL}} = 3.91 \, \text{mmol Ca}^{2+} \\
\]

From the balanced chemical equation, 3.91 mmol CO$_3^{2-}$ will be needed to precipitate this quantity of Ca$^{2+}$ ion.

\[
3.91 \, \text{mmol CO}_3^{2-} \times \frac{1.00 \, \text{mL}}{0.125 \, \text{mmol}} = 31.2 \, \text{mL} \\
\]

66. \[ \text{molar mass Na}_2\text{C}_2\text{O}_4 = 134.0 \, \text{g} \]
\[ 37.5 \, \text{mL} = 0.0375 \, \text{L} \]

\[
\text{moles Ca}^{2+} \text{ ion} = 0.0375 \, \text{L} \times \frac{0.104 \, \text{mol Ca}^{2+}}{1.00 \, \text{L}} = 0.00390 \, \text{mol Ca}^{2+} \text{ ion} \\
\]

\[ \text{Ca}^{2+}(aq) + \text{C}_2\text{O}_4^{2-}(aq) \rightarrow \text{CaC}_2\text{O}_4(s) \]

As the precipitation reaction is of 1:1 stoichiometry, then 0.00390 mol of C$_2$O$_4^{2-}$ ion is needed. Moreover, each formula unit of Na$_2$C$_2$O$_4$ contains one C$_2$O$_4^{2-}$ ion, so 0.00390 mol of Na$_2$C$_2$O$_4$ is required.

\[ 0.00390 \, \text{mol Na}_2\text{C}_2\text{O}_4 \times \frac{134.0 \, \text{g}}{1 \, \text{mol}} = 0.523 \, \text{g Na}_2\text{C}_2\text{O}_4 \text{ required} \]
68. 10.0 mL = 0.0100 L

\[
0.0100 \, \text{L} \times \frac{0.250 \, \text{mol AlCl}_3}{1.00 \, \text{L}} = 2.50 \times 10^{-3} \, \text{mol AlCl}_3
\]

\[
\text{AlCl}_3(aq) + 3\text{NaOH(s)} \rightarrow \text{Al(OH)}_3(s) + 3\text{NaCl(aq)}
\]

\[
2.50 \times 10^{-3} \, \text{mol AlCl}_3 \times \frac{3 \, \text{mol NaOH}}{1 \, \text{mol AlCl}_3} = 7.50 \times 10^{-3} \, \text{mol NaOH}
\]

Molar mass NaOH = 40.00 g

\[
7.50 \times 10^{-3} \, \text{mol NaOH} \times \frac{40.00 \, \text{g NaOH}}{1 \, \text{mol}} = 0.300 \, \text{g NaOH}
\]

70. \text{NaOH}(aq) + \text{HC}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{Na}^+\text{C}_2\text{H}_3\text{O}_2(aq)

\[
24.8 \, \text{mL} \times \frac{0.833 \, \text{mmol acetic acid}}{1.00 \, \text{mL}} = 20.66 \, \text{mmol acetic acid}
\]

From the balanced chemical equation, \text{NaOH} and acetic acid react on a 1:1 basis, so 20.66 mmol of \text{NaOH} will be needed for the reaction.

\[
20.66 \, \text{mmol NaOH} \times \frac{1.00 \, \text{mL}}{0.125 \, \text{mmol NaOH}} = 165 \, \text{mL of NaOH solution required}
\]

72. 7.2 mL = 0.0072 L

\[
0.0072 \, \text{L} \times \frac{2.5 \times 10^{-3} \, \text{mol NaOH}}{1.00 \, \text{L}} = 1.8 \times 10^{-5} \, \text{mol NaOH}
\]

\[
\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)
\]

\[
1.8 \times 10^{-5} \, \text{mol OH}^- \times \frac{1 \, \text{mol H}^+}{1 \, \text{mol OH}^-} = 1.8 \times 10^{-5} \, \text{mol H}^+
\]

100 mL = 0.100 L

\[
M = \frac{1.8 \times 10^{-5} \, \text{mol H}^+}{0.100 \, \text{L}} = 1.8 \times 10^{-4} \, \text{M H}^+(aq)
\]

74. Experimentally, neutralization reactions are usually performed with volumetric glassware that is calibrated in milliliters rather than liters. For convenience in calculations for such reactions, the arithmetic is often performed in terms of milliliters and millimoles rather than in liters and moles: 1 mmol = 0.001 mol. Note that the number of moles of solute per liter of solution, the molarity, is numerically equivalent to the number of millimoles of solute per milliliter of solution.

a. \text{HNO}_3(aq) + \text{NaOH(aq)} \rightarrow \text{NaNO}_3(aq) + \text{H}_2\text{O}(l)

\[
12.7 \, \text{mL} \times \frac{0.501 \, \text{mmol}}{1.00 \, \text{mL}} = 6.36 \, \text{mmol NaOH present in the sample}
\]

\[
6.36 \, \text{mmol NaOH} \times \frac{1 \, \text{mmol HNO}_3}{1 \, \text{mmol NaOH}} = 6.36 \, \text{mmol HNO}_3 \text{ required to react}
\]
6.36 mmol HNO₃ × \frac{1.00 \text{ mL}}{0.101 \text{ mmol HNO}_3} = 63.0 \text{ mL HNO}_3 \text{ required}

b. \[2\text{HNO}_3(\text{aq}) + \text{Ba(OH)}_2(\text{aq}) \rightarrow \text{Ba(NO}_3)_2 + 2\text{H}_2\text{O(l)}\]

\[24.9 \text{ mL} \times \frac{0.00491 \text{ mmol}}{1.00 \text{ mL}} = 0.122 \text{ mmol Ba(OH)}_2 \text{ present in the sample}\]

\[0.122 \text{ mmol Ba(OH)}_2 \times \frac{2 \text{ mmol HNO}_3}{1 \text{ mmol Ba(OH)}_2} = 0.244 \text{ mmol HNO}_3 \text{ required}\]

\[0.244 \text{ mmol HNO}_3 \times \frac{1.00 \text{ mL}}{0.101 \text{ mmol HNO}_3} = 2.42 \text{ mL HNO}_3 \text{ is required}\]

c. \[\text{HNO}_3(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq})\]

\[49.1 \text{ mL} \times \frac{0.103 \text{ mmol}}{1.00 \text{ mL}} = 5.06 \text{ mmol NH}_3 \text{ present in the sample}\]

\[5.06 \text{ mmol NH}_3 \times \frac{1 \text{ mmol HNO}_3}{1 \text{ mmol NH}_3} = 5.06 \text{ mmol HNO}_3 \text{ required}\]

\[5.06 \text{ mmol HNO}_3 \times \frac{1.00 \text{ mL}}{0.101 \text{ mmol HNO}_3} = 50.1 \text{ mL HNO}_3 \text{ required}\]

d. \[\text{KOH(\text{aq}) + HNO}_3(\text{aq}) \rightarrow \text{KNO}_3(\text{aq}) + \text{H}_2\text{O(l)}\]

\[1.21 \text{ L} \times \frac{0.102 \text{ mol}}{1.00 \text{ L}} = 0.123 \text{ mol KOH present in the sample}\]

\[0.123 \text{ mol KOH} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol KOH}} = 0.123 \text{ mol HNO}_3 \text{ required}\]

\[0.123 \text{ mol HNO}_3 \times \frac{1.00 \text{ L}}{0.101 \text{ mol HNO}_3} = 1.22 \text{ L HNO}_3 \text{ required}\]

76. one normal

78. 1.53 equivalents OH⁻ ion are needed to react with 1.53 equivalents of H⁺ ion. By definition, one equivalent of OH⁻ ion exactly neutralizes one equivalent of H⁺ ion.

80. \[N = \frac{\text{number of equivalents of solute}}{\text{number of liters of solution}}\]

a. equivalent weight NaOH = molar mass NaOH = 40.00 g

\[0.113 \text{ g NaOH} \times \frac{1 \text{ equiv NaOH}}{40.00 \text{ g}} = 2.83 \times 10^{-3} \text{ equiv NaOH}\]

\[10.2 \text{ mL} = 0.0102 \text{ L}\]

\[N = \frac{2.83 \times 10^{-3} \text{ equiv}}{0.0102 \text{ L}} = 0.277 N\]
b. equivalent weight \( \text{Ca(OH)}_2 \) = \( \frac{\text{molar mass}}{2} \) = \( \frac{74.10 \text{ g}}{2} \) = \( 37.05 \text{ g} \)

\[
12.5 \text{ mg} \times \frac{1 \text{ g}}{10^3 \text{ mg}} \times \frac{1 \text{ equiv}}{37.05 \text{ g}} = 3.37 \times 10^{-4} \text{ equiv Ca(OH)}_2
\]

100. mL = 0.100 L

\[
N = \frac{3.37 \times 10^{-3} \text{ equiv}}{0.100 \text{ L}} = 3.37 \times 10^{-2} \text{ } N
\]

c. equivalent weight \( \text{H}_2\text{SO}_4 \) = \( \frac{\text{molar mass}}{2} \) = \( \frac{98.09 \text{ g}}{2} \) = \( 49.05 \text{ g} \)

\[
12.4 \text{ g} \times \frac{1 \text{ equiv}}{49.05 \text{ g}} = 0.253 \text{ equiv H}_2\text{SO}_4
\]

155 mL = 0.155 L

\[
N = \frac{0.253 \text{ equiv}}{0.155 \text{ L}} = 1.63 \text{ N}
\]

82.

a. 0.134 \( M \) \( \text{NaOH} \) \times \frac{1 \text{ equiv NaOH}}{1 \text{ mol NaOH}} = 0.134 \( N \) \( \text{NaOH} \)

b. 0.00521 \( M \) \( \text{Ca(OH)}_2 \) \times \frac{2 \text{ equiv Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 0.0104 \( N \) \( \text{Ca(OH)}_2 \)

c. 4.42 \( M \) \( \text{H}_3\text{PO}_4 \) \times \frac{3 \text{ equiv H}_3\text{PO}_4}{1 \text{ mol H}_3\text{PO}_4} = 13.3 \( N \) \( \text{H}_3\text{PO}_4 \)

84. Molar mass of \( \text{Ca(OH)}_2 \) = 74.10 g

\[
5.21 \text{ mg} \text{ Ca(OH)}_2 \times \frac{1 \text{ g}}{10^3 \text{ mg}} \times \frac{1 \text{ mol}}{74.10 \text{ g}} = 7.03 \times 10^{-5} \text{ mol Ca(OH)}_2
\]

1000. mL = 1.000 L (volumetric flask volume; 4 significant figures).

\[
M = \frac{7.03 \times 10^{-5} \text{ mol}}{1.000 \text{ L}} = 7.03 \times 10^{-5} \text{ M Ca(OH)}_2
\]

\[
N = 7.03 \times 10^{-5} \text{ M Ca(OH)}_2 \times \frac{2 \text{ equiv Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 1.41 \times 10^{-4} \text{ N Ca(OH)}_2
\]

86. \( \text{H}_2\text{SO}_4(aq) + 2\text{NaOH(aq)} \to \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)} \)

\[
N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}}
\]

\[
(0.104 \text{ } N)(V_{\text{acid}}) = (0.152 \text{ } N)(15.2 \text{ mL})
\]

\[
V_{\text{acid}} = 22.2 \text{ mL}
\]
The 0.104 M sulfuric acid solution is twice as concentrated as the 0.104 N sulfuric acid solution (1 mole = 2 equivalents), so half as much will be required to neutralize the same quantity of NaOH = 11.1 mL

88. \[2\text{NaOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2 \text{H}_2\text{O}(l)\]

\[27.34 \text{ mL NaOH} \times \frac{0.1021 \text{ mmol}}{1.00 \text{ mL}} = 2.791 \text{ mmol NaOH}\]

\[2.791 \text{ mmol NaOH} \times \frac{1 \text{ mmol H}_2\text{SO}_4}{2 \text{ mmol NaOH}} = 1.396 \text{ mmol H}_2\text{SO}_4\]

\[M = \frac{1.396 \text{ mmol H}_2\text{SO}_4}{25.00 \text{ mL}} = 0.05583 \text{ M H}_2\text{SO}_4 = 0.1117 \text{ N H}_2\text{SO}_4\]

90. Molarity is defined as the number of moles of solute contained in 1 liter of total solution volume (solute plus solvent after mixing). In the first case where 50. g of NaCl is dissolved in 1.0 L of water, the total volume after mixing is not known, and the molarity cannot be calculated. In the second example, the final volume after mixing is known, and the molarity can be calculated simply.

92. \[75 \text{ g solution} \times \frac{25 \text{ g NaCl}}{100. \text{ g solution}} = 18.75 \text{ g NaCl}\]

\[\text{new } \% = \frac{18.75 \text{ g NaCl}}{575 \text{ g solution}} \times 100 = 3.26 = 3.3\%\]

94. molar mass NaHCO$_3$ = 84.01 g; 25.2 mL = 0.0252 L

\[\text{NaHCO}_3(s) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)\]

\[\text{mol HCl} = \text{mol NaHCO}_3 \text{ required} = 0.0252 \text{ L} \times \frac{6.01 \text{ mol HCl}}{1.00 \text{ L}} = 0.151 \text{ mol}\]

\[0.151 \text{ mol} \times \frac{84.01 \text{ g}}{1 \text{ mol}} = 12.7 \text{ g NaHCO}_3 \text{ required}\]

96. molar mass H$_2$O = 18.0 g

\[1.0 \text{ L water} = 1.0 \times 10^3 \text{ mL water} \equiv 1.0 \times 10^3 \text{ g water}\]

\[1.0 \times 10^3 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = 56 \text{ mol H}_2\text{O}\]

98. 500 mL HCl solution = 0.500 L HCl solution

\[0.500 \text{ L solution} \times \frac{0.100 \text{ mol HCl}}{1.00 \text{ L solution}} = 0.0500 \text{ mol HCl}\]

\[0.0500 \text{ mol HCl} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 1.12 \text{ L HCl gas at STP}\]
100. \(10.0 \ \text{g HCl} \times \frac{100 \ \text{g solution}}{33.1 \ \text{g HCl}} = 30.21 \ \text{g solution}\)

\(30.21 \ \text{g solution} \times \frac{1.00 \ \text{mL solution}}{1.147 \ \text{g solution}} = 26.3 \ \text{mL solution}\)

102. molar mass \(\text{CaCl}_2 = 111.0 \ \text{g}\)

\(14.2 \ \text{g CaCl}_2 \times \frac{1 \ \text{mol CaCl}_2}{111.0 \ \text{g CaCl}_2} = 0.128 \ \text{mol CaCl}_2\)

\(50.0 \ \text{mL} = 0.0500 \ \text{L}\)

\(M = \frac{0.128 \ \text{mol CaCl}_2}{0.0500 \ \text{L}} = 2.56 \ M\)

104.

a. \(\frac{5.0 \ \text{g KNO}_3}{(5.0 \ \text{g KNO}_3 + 75 \ \text{g H}_2\text{O})} \times 100 = 6.3\% \ \text{KNO}_3\)

b. \(2.5 \ \text{mg} = 0.0025 \ \text{g}\)

\(\frac{0.0025 \ \text{g KNO}_3}{(0.0025 \ \text{g KNO}_3 + 1.0 \ \text{g H}_2\text{O})} \times 100 = 0.25\% \ \text{KNO}_3\)

c. \(\frac{11 \ \text{g KNO}_3}{(11 \ \text{g KNO}_3 + 89 \ \text{g H}_2\text{O})} \times 100 = 11\% \ \text{KNO}_3\)

d. \(\frac{11 \ \text{g KNO}_3}{(11 \ \text{g KNO}_3 + 49 \ \text{g H}_2\text{O})} \times 100 = 18\% \ \text{KNO}_3\)

106. \(\% \ \text{C} = \frac{5.0 \ \text{g C}}{(5.0 \ \text{g C} + 1.5 \ \text{g Ni} + 100. \ \text{g Fe})} \times 100 = 4.7\% \ \text{C}\)

\(\% \ \text{Ni} = \frac{1.5 \ \text{g Ni}}{(5.0 \ \text{g C} + 1.5 \ \text{g Ni} + 100. \ \text{g Fe})} \times 100 = 1.4\% \ \text{Ni}\)

\(\% \ \text{Fe} = \frac{100. \ \text{g Fe}}{(5.0 \ \text{g C} + 1.5 \ \text{g Ni} + 100. \ \text{g Fe})} \times 100 = 93.9\% \ \text{Fe}\)

108. To say that the solution is 5.5\% by mass \(\text{Na}_2\text{CO}_3\) means that 5.5 g of \(\text{Na}_2\text{CO}_3\) are contained in every 100 g of the solution.

\(500. \ \text{g solution} \times \frac{5.5 \ \text{g Na}_2\text{CO}_3}{100. \ \text{g solution}} = 28 \ \text{g Na}_2\text{CO}_3\)

110. For \(\text{NaCl}\): \(125 \ \text{g solution} \times \frac{7.5 \ \text{g NaCl}}{100. \ \text{g solution}} = 9.4 \ \text{g NaCl}\)

For \(\text{KBr}\): \(125 \ \text{g solution} \times \frac{2.5 \ \text{g KBr}}{100. \ \text{g solution}} = 3.1 \ \text{g KBr}\)
112.  
   a. 25 mL = 0.025 L  
       \[ M = \frac{0.10 \text{ mol CaCl}_2}{0.025 \text{ L solution}} = 4.0 \text{ M} \]  
   b. \[ M = \frac{2.5 \text{ mol KBr}}{2.5 \text{ L solution}} = 1.0 \text{ M} \]  
   c. 755 mL = 0.755 L  
       \[ M = \frac{0.55 \text{ mol NaNO}_3}{0.755 \text{ L solution}} = 0.73 \text{ M} \]  
   d. \[ M = \frac{4.5 \text{ mol Na}_2\text{SO}_4}{1.25 \text{ L solution}} = 3.6 \text{ M} \]  

114. molar mass C$_{12}$H$_{22}$O$_{11}$ = 342.3 g; 450. mL = 0.450 L  
    \[ 125 \text{ g C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{1 \text{ mol}}{342.3 \text{ g}} = 0.3652 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11} \]  
    \[ M = \frac{0.3652 \text{ mol}}{0.450 \text{ L solution}} = 0.812 \text{ M} \]  

116. molar mass NaCl = 58.44 g  
    \[ 1.5 \text{ g NaCl} \times \frac{1 \text{ mol}}{58.44 \text{ g}} = 0.0257 \text{ mol NaCl} \]  
    \[ M = \frac{0.0257 \text{ mol NaCl}}{1.00 \text{ L solution}} = 0.026 \text{ M} \]  

118.  
   a. 4.25 L solution \( \times \) \( \frac{0.105 \text{ mol KCl}}{1.00 \text{ L solution}} = 0.446 \text{ mol KCl} \)  
       molar mass KCl = 74.6 g  
       \[ 0.446 \text{ mol KCl} \times \frac{74.6 \text{ g KCl}}{1 \text{ mol KCl}} = 33.3 \text{ g KCl} \]  
   b. 15.1 mL = 0.0151 L  
       \[ 0.0151 \text{ L solution} \times \frac{0.225 \text{ mol NaNO}_3}{1.00 \text{ L solution}} = 3.40 \times 10^{-3} \text{ mol NaNO}_3 \]  
       molar mass NaNO$_3$ = 85.00 g  
       \[ 3.40 \times 10^{-3} \text{ mol} \times \frac{85.00 \text{ g NaNO}_3}{1 \text{ mol NaNO}_3} = 0.289 \text{ g NaNO}_3 \]
c. 25 mL = 0.025 L

\[ \frac{3.0 \text{ mol HCl}}{1.00 \text{ L solution}} \times 0.025 \text{ L solution} = 0.075 \text{ mol HCl} \]

Molar mass HCl = 36.46 g

\[ 0.075 \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 2.7 \text{ g HCl} \]

d. 100. mL = 0.100 L

\[ \frac{0.505 \text{ mol H}_2\text{SO}_4}{1.00 \text{ L solution}} \times 0.100 \text{ L solution} = 0.0505 \text{ mol H}_2\text{SO}_4 \]

Molar mass H\textsubscript{2}SO\textsubscript{4} = 98.09 g

\[ 0.0505 \text{ mol H}_2\text{SO}_4 \times \frac{98.09 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 4.95 \text{ g H}_2\text{SO}_4 \]

120.

a. 1.25 L \times \frac{0.250 \text{ mol Na}_3\text{PO}_4}{1.00 \text{ L}} = 0.3125 \text{ mol Na}_3\text{PO}_4

0.3125 \text{ mol Na}_3\text{PO}_4 \times \frac{3 \text{ mol Na}^+}{1 \text{ mol Na}_3\text{PO}_4} = 0.938 \text{ mol Na}^+

0.3125 \text{ mol Na}_3\text{PO}_4 \times \frac{1 \text{ mol PO}_4^{3-}}{1 \text{ mol Na}_3\text{PO}_4} = 0.313 \text{ mol PO}_4^{3-}

b. 3.5 mL = 0.0035 L

0.0035 L \times \frac{6.0 \text{ mol H}_2\text{SO}_4}{1.00 \text{ L}} = 0.021 \text{ mol H}_2\text{SO}_4

0.021 \text{ mol H}_2\text{SO}_4 \times \frac{2 \text{ mol H}^+}{1 \text{ mol H}_2\text{SO}_4} = 0.042 \text{ mol H}^+

0.021 \text{ mol H}_2\text{SO}_4 \times \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol H}_2\text{SO}_4} = 0.021 \text{ mol SO}_4^{2-}

c. 25 mL = 0.025 L

0.025 L \times \frac{0.15 \text{ mol AlCl}_3}{1.00 \text{ L}} = 0.00375 \text{ mol AlCl}_3

0.00375 \text{ mol AlCl}_3 \times \frac{1 \text{ mol Al}^{3+}}{1 \text{ mol AlCl}_3} = 0.0038 \text{ mol Al}^{3+}

0.00375 \text{ mol AlCl}_3 \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol AlCl}_3} = 0.011 \text{ mol Cl}^-$

Copyright © Houghton Mifflin Company. All rights reserved.
d. \[ 1.50 \text{ L} \times \frac{1.25 \text{ mol BaCl}_2}{1.00 \text{ L}} = 1.875 \text{ mol BaCl}_2 \]

\[ 1.875 \text{ mol BaCl}_2 \times \frac{1 \text{ mol Ba}^{2+}}{1 \text{ mol BaCl}_2} = 1.88 \text{ mol Ba}^{2+} \]

\[ 1.875 \text{ mol BaCl}_2 \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol BaCl}_2} = 3.75 \text{ mol Cl}^- \]

122. \[ M_1 \times V_1 = M_2 \times V_2 \]

a. \[ M_1 = 0.200 \text{ M} \quad M_2 = ? \]
\[ V_1 = 125 \text{ mL} \quad V_2 = 125 + 150 = 275 \text{ mL} \]
\[ M_2 = \frac{(0.200 \text{ M})(125 \text{ mL})}{275 \text{ mL}} = 0.0909 \text{ M} \]

b. \[ M_1 = 0.250 \text{ M} \quad M_2 = ? \]
\[ V_1 = 155 \text{ mL} \quad V_2 = 155 + 150 = 305 \text{ mL} \]
\[ M_2 = \frac{(0.250 \text{ M})(155 \text{ mL})}{305 \text{ mL}} = 0.127 \text{ M} \]

c. \[ M_1 = 0.250 \text{ M} \quad M_2 = ? \]
\[ V_1 = 0.500 \text{ L} = 500 \text{ mL} \quad V_2 = 500 + 150 = 650 \text{ mL} \]
\[ M_2 = \frac{(0.250 \text{ M})(500 \text{ mL})}{650 \text{ mL}} = 0.192 \text{ M} \]

d. \[ M_1 = 18.0 \text{ M} \quad M_2 = ? \]
\[ V_1 = 15 \text{ mL} \quad V_2 = 15 + 150 = 165 \text{ mL} \]
\[ M_2 = \frac{(18.0 \text{ M})(15 \text{ mL})}{165 \text{ mL}} = 1.6 \text{ M} \]

124. \[ M_1 \times V_1 = M_2 \times V_2 \]

\[ M_1 = 5.4 \text{ M} \quad M_2 = ? \]
\[ V_1 = 50. \text{ mL} \quad V_2 = 300. \text{ mL} \]
\[ M_2 = \frac{(5.4 \text{ M})(50 \text{ mL})}{300 \text{ mL}} = 0.90 \text{ M} \]
126.  25.0 mL = 0.0250 L

\[ 0.0250 \text{ L NiCl}_2 \text{ solution} \times \frac{0.20 \text{ mol NiCl}_2}{1.00 \text{ L NiCl}_2 \text{ solution}} = 0.00500 \text{ mol NiCl}_2 \]

\[ 0.00500 \text{ mol NiCl}_2 \times \frac{1 \text{ mol Na}_2\text{S}}{1 \text{ mol NiCl}_2} = 0.00500 \text{ mol Na}_2\text{S} \]

\[ 0.00500 \text{ mol Na}_2\text{S} \times \frac{1.00 \text{ L Na}_2\text{S solution}}{0.10 \text{ mol Na}_2\text{S}} = 0.050 \text{ L} = 50. \text{ mL Na}_2\text{S solution} \]

128.  \( \text{HNO}_3(aq) + \text{NaOH}(aq) \rightarrow \text{NaNO}_3(aq) + \text{H}_2\text{O}(l) \)

\[ 35.0 \text{ mL} = 0.0350 \text{ L} \]

\[ 0.0350 \text{ L} \times \frac{0.150 \text{ mol NaOH}}{1.00 \text{ L}} = 5.25 \times 10^{-3} \text{ mol NaOH} \]

\[ 5.25 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol NaOH}} = 5.25 \times 10^{-3} \text{ mol HNO}_3 \]

\[ 5.25 \times 10^{-3} \text{ mol HNO}_3 \times \frac{1.00 \text{ L}}{0.150 \text{ mol HNO}_3} = 0.0350 \text{ L} = 35.0 \text{ mL HNO}_3 \]

130.

a.  equivalent weight \( \text{HCl} = \text{molar mass HCl} = 36.46 \text{ g} \); 500. mL = 0.500 L

\[ 15.0 \text{ g HCl} \times \frac{1 \text{ equiv HCl}}{36.46 \text{ g HCl}} = 0.411 \text{ equiv HCl} \]

\[ N = \frac{0.411 \text{ equiv}}{0.500 \text{ L}} = 0.822 \text{ N} \]

b.  equivalent weight \( \text{H}_2\text{SO}_4 = \frac{\text{molar mass}}{2} \) = \( \frac{98.09 \text{ g}}{2} \) = 49.05 g; 250. mL = 0.250 L

\[ 49.0 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ equiv H}_2\text{SO}_4}{49.05 \text{ g H}_2\text{SO}_4} = 0.999 \text{ equiv H}_2\text{SO}_4 \]

\[ N = \frac{0.999 \text{ equiv}}{0.250 \text{ L}} = 4.00 \text{ N} \]

c.  equivalent weight \( \text{H}_3\text{PO}_4 = \frac{\text{molar mass}}{3} \) = \( \frac{98.0 \text{ g}}{3} \) = 32.67 g; 100. mL = 0.100 L

\[ 10.0 \text{ g H}_3\text{PO}_4 \times \frac{1 \text{ equiv H}_3\text{PO}_4}{32.67 \text{ g H}_3\text{PO}_4} = 0.3061 \text{ equiv H}_3\text{PO}_4 \]

\[ N = \frac{0.3061 \text{ equiv}}{0.100 \text{ L}} = 3.06 \text{ N} \]
132. molar mass NaH₂PO₄ = 120.0 g; 500. mL = 0.500 L

\[
5.0 \text{ g NaH}_2\text{PO}_4 \times \frac{1 \text{ mol NaH}_2\text{PO}_4}{120.0 \text{ g NaH}_2\text{PO}_4} = 0.04167 \text{ mol NaH}_2\text{PO}_4
\]

\[
M = \frac{0.04167 \text{ mol}}{0.500 \text{ L}} = 0.08333 \text{M NaH}_2\text{PO}_4 = 0.083 \text{M NaH}_2\text{PO}_4
\]

\[
0.08333 \text{M NaH}_2\text{PO}_4 \times \frac{2 \text{ equiv NaH}_2\text{PO}_4}{1 \text{ mol NaH}_2\text{PO}_4} = 0.1667 \text{N NaH}_2\text{PO}_4 = 0.17 \text{N NaH}_2\text{PO}_4
\]

134. \( N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}} \)

\[
N_{\text{acid}} \times (10.0 \text{ mL}) = (3.5 \times 10^{-2} \text{ N})(27.5 \text{ mL})
\]

\[
N_{\text{acid}} = 9.6 \times 10^{-2} \text{ N HNO}_3
\]
Cumulative Review Chapters 13–15

ANSWERS

2. The pressure of the atmosphere represents the mass of the gases in the atmosphere pressing down on the surface of the earth. The device most commonly used to measure the pressure of the atmosphere is the mercury barometer shown in Figure 13.2 in the text.

A simple experiment to demonstrate the pressure of the atmosphere is shown in Figure 13.1 in the text. Some water is added to a metal can, and the can is heated until the water boils (boiling represents when the pressure of the vapor coming from the water is equal to the atmospheric pressure). The can is then stoppered. As the steam in the can cools, it condenses to liquid water, which lowers the pressure of gas inside the can. The pressure of the atmosphere outside the can is then much larger than the pressure inside the can, and the can collapses.

4. In simple terms, Boyle’s law states that the volume of a gas sample will decrease if you squeeze harder on it. Imagine squeezing hard on a tennis ball with your hand; the ball collapses as the gas inside is forced into a smaller volume by your hand. Of course, to be perfectly correct, the temperature and amount of gas (moles) must remain the same as you adjust the pressure for Boyle’s law to hold true. There are two mathematical statements of Boyle’s law you should remember. The first is

\[ P \times V = \text{constant}, \]

which basically is the definition of Boyle’s law [in order for the product \((P \times V)\) to remain constant, if one of these terms increases the other must decrease]. The second formula is the one more commonly used in solving problems,

\[ P_1 \times V_1 = P_2 \times V_2. \]

With this second formulation, we can determine pressure-volume information about a given sample under two sets of conditions. These two mathematical formulas are just two different ways of saying the same thing: If the pressure on a sample of gas increases, the volume of the sample of gas will decrease. A graph of Boyle’s law data is given as Figure 13.5. This type of graph \((xy = k)\) is known to mathematicians as a hyperbola.

6. Charles’s law simply says that if you heat a sample of gas, the volume of the sample will increase. That is, when the temperature of a gas increases, the volume of the gas also increases (assuming the pressure and amount of gas remains the same). Charles’s law is a direct proportionality when the temperature is expressed in kelvins (if you increase \(T\), this increases \(V\)), whereas Boyle’s law is an inverse proportionality (if you increase \(P\), this decreases \(V\)). There are two mathematical statements of Charles’s law with which you should be familiar. The first statement is

\[ V = kT. \]
This is simply a definition (the volume of a gas sample is directly related to its Kelvin temperature; if you increase the temperature, the volume increases). The working formulation of Charles’s law we use in problem solving is given as:

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

With this formulation, we can determine volume-temperature information for a given gas sample under two sets of conditions. Charles’s law holds true only if the amount of gas remains the same (obviously the volume of a gas sample would increase if there were more gas present) and also if the pressure remains the same (a change in pressure also changes the volume of a gas sample).

8. Avogadro’s law tells us that, with all other things being equal, two moles of gas are twice as big as one mole of gas! That is, the volume of a sample of gas is directly proportional to the number of moles or molecules of gas present (at constant temperature and pressure). If we wanted to compare the volumes of two samples of the same gas as an indication of the amount of gas present in the samples, we would have to make certain that the two samples of gas were at the same pressure and temperature. The volume of a sample of gas would vary with either temperature or pressure, or both. Avogadro’s law holds true for comparing gas samples that are under the same conditions. Avogadro’s law is a direct proportionality; the greater the number of gas molecules you have in a sample, the larger the sample’s volume will be.

10. The “partial” pressure of an individual gas in a mixture of gases represents the pressure the gas would have in the same container at the same temperature if it were the only gas present. The total pressure in a mixture of gases is the sum of the individual partial pressures of the gases present in the mixture. Because the partial pressures of the gases in a mixture are additive (i.e., the total pressure is the sum of the partial pressures), this suggests that the total pressure in a container is a function of only the number of molecules present in the container and not of the identity of the molecules or any other property of the molecules (such as their inherent atomic size).

12. The main postulates of the kinetic-molecular theory for gases are as follows: (a) gases consist of tiny particles (atoms or molecules), and the size of these particles is negligible compared to the bulk volume of a gas sample; (b) the particles in a gas are in constant random motion, colliding with each other and with the walls of the container; (c) the particles in a gas sample do not exert any attractive or repulsive forces on one another; (d) the average kinetic energy of the particles in a sample of gas is directly related to the absolute temperature of the gas sample. The pressure exerted by a gas is a result of the molecules colliding with (and pushing on) the walls of the container. The pressure increases with temperature because at a higher temperature the molecules are moving faster and hit the walls of the container with greater force. A gas fills whatever volume is available to it because the molecules in a gas are in constant random motion. If the motion of the molecules is random, they eventually will move out into whatever volume is available until the distribution of molecules is uniform. At constant pressure, the volume of a gas sample increases as the temperature is increased because with each collision having greater force, the container must expand so that the molecules (and therefore the collisions) are farther apart if the pressure is to remain constant.

14. Solids and liquids are much more condensed states of matter than are gases. The molecules are much closer together in solids and liquids and interact with each other to a much greater extent. Solids and liquids have much greater densities than do gases and are much less compressible because there is so little room between the molecules in the solid and liquid states (solids and liquids effectively have native volumes of their own, and their volumes are not affected nearly as much by the temperature or pressure). Although solids are more rigid than liquids, the solid and liquid states have much more in common with each other than either of these states has with the
gaseous state. We know this is true because it typically takes only a few kilojoules of energy to melt 1 mol of a solid (not much change has to take place in the molecules), whereas it may take ten times more energy to vaporize a liquid (as there is a great change between the liquid and gaseous states).

16. The normal boiling point of water, that is, water’s boiling point at a pressure of exactly 760 mm Hg, is 100 °C (You will recall that the boiling point of water was used to set one of the reference temperatures of the Celsius temperature scale.). Water remains at 100 °C while boiling until all the water has boiled away because the additional heat energy being added to the sample is used to overcome attractive forces among the water molecules as they go from the condensed, liquid state to the gaseous state. The normal (760 mm Hg) freezing point of water is exactly 0 °C (Again, this property of water was used as one of the reference points for the Celsius temperature scale.). A cooling curve for water is given in Figure 14.2. Notice how the curve shows that the amount of heat needed to boil the sample is much larger than the amount needed to melt the sample.

18. Dipole-dipole forces are a type of intermolecular force that can exist between molecules with permanent dipole moments. Molecules with permanent dipole moments try to orient themselves so that the positive end of one polar molecule can attract the negative end of another polar molecule. Dipole-dipole forces are not nearly as strong as ionic or covalent bonding forces (only about 1% as strong as covalent bonding forces) because electrostatic attraction is related to the magnitude of the charges of the attracting species. As polar molecules have only a “partial” charge at each end of the dipole, the magnitude of the attractive force is not as large. The strength of such forces also drops rapidly as molecules become farther apart and is important only in the solid and liquid states (such forces are negligible in the gaseous state because the molecules are too far apart). Hydrogen bonding is an especially strong sort of dipole-dipole attractive force that can exist when hydrogen atoms are directly bonded to the most strongly electronegative atoms (N, O, and F). Because the hydrogen atom is so small, dipoles involving N–H, O–H, and F–H bonds can approach each other much more closely than can dipoles involving other atoms. As the magnitude of dipole-dipole forces is dependent on distance, unusually strong attractive forces can exist in such molecules. We take the fact that the boiling point of water is higher than that of the other covalent hydrogen compounds of the Group 6 elements as evidence for the special strength of hydrogen bonding (It takes more energy to vaporize water because of the extra-strong forces holding together the molecules in the liquid state.).

20. Vaporization of a liquid requires an input of energy because the intermolecular forces that hold the molecules together in the liquid state must be overcome. The high heat of vaporization of water is essential to life on Earth because much of the excess energy striking the Earth from the sun is dissipated in vaporizing water. Condensation is the opposite process to vaporization; that is, condensation refers to the process by which molecules in the vapor state form a liquid. In a closed container containing a liquid and some empty space above the liquid, an equilibrium is set up between vaporization and condensation. The liquid in such a sealed container never completely evaporates. When the liquid is first placed in the container, the liquid phase begins to evaporate into the empty space. As the number of molecules in the vapor phase begins to get large, however, some of these molecules begin to re-enter the liquid phase. Eventually, every time a molecule of liquid somewhere in the container enters the vapor phase, somewhere else in the container a molecule of vapor re-enters the liquid. There is no further net change in the amount of liquid phase (although molecules are continually moving between the liquid and vapor phases). The pressure of the vapor in such an equilibrium situation is characteristic for the liquid at each particular temperature (For example, the vapor pressures of water are tabulated at different temperatures in Table 13.2.). A simple experiment to determine vapor pressure is shown in Figure 14.10. Samples of a liquid are injected into a sealed tube containing mercury. Because mercury is so dense, the liquids float to the top of the mercury where they evaporate. As the vapor pressures
of the liquids develop to the saturation point, the level of mercury in the tube changes as an index of the magnitude of the vapor pressures. Typically, liquids with strong intermolecular forces have small vapor pressures (they have more difficulty in evaporating) than do liquids with very weak intermolecular forces: For example, the components of gasoline (weak forces) have much higher vapor pressures and evaporate more easily than does water (strong forces).

22. The simple model we use to explain many properties of metallic elements is called the electron sea model. In this model we picture a regular lattice array of metal cations in sort of a “sea” of mobile valence electrons. The electrons can move easily to conduct heat or electricity through the metal, and the lattice of cations can be deformed fairly easily, allowing the metal to be hammered into a sheet or stretched to make a wire. An alloy contains a mixture of elements that overall has metallic properties. Substitutional alloys consist of a host metal in which some of the atoms in the metal’s crystalline structure are replaced by atoms of other metallic elements of comparable size to the atoms of the host metal. For example, sterling silver consists of an alloy in which approximately 7% of the silver atoms have been replaced by copper atoms. Brass and pewter are also substitutional alloys. An interstitial alloy is formed when other smaller atoms enter the interstices (holes) between atoms in the host metal’s crystal structure. Steel is an interstitial alloy in which carbon atoms typically enter the interstices of a crystal of iron atoms. The presence of the interstitial carbon atoms markedly changes the properties of the iron, making it much harder, more malleable, and more ductile. Depending on the amount of carbon introduced into the iron crystals, the properties of the resulting steel can be controlled carefully.

24. A saturated solution is one that contains as much solute as can dissolve at a particular temperature. To say that a solution is saturated does not necessarily mean that the solute is present at a high concentration. For example, magnesium hydroxide dissolves only to a very small extent before the solution is saturated, whereas it takes a great deal of sugar to form a saturated solution (and the saturated solution is extremely concentrated). A saturated solution is one that is in equilibrium with undissolved solute. As molecules of solute dissolve from the solid in one place in the solution, dissolved molecules rejoin the solid phase in another place in the solution. As with the development of vapor pressure above a liquid (see Question 20 above), formation of a solution reaches a state of dynamic equilibrium. Once the rates of dissolving and “undissolving” become equal there will be no further net change in the concentration of the solution, and the solution will be saturated.

26. Adding more solvent to a solution so as to dilute the solution does not change the number of moles of solute present, but only changes the volume in which the solute is dispersed. If we are using the molarity of the solution to describe its concentration, the number of liters is changed when we add solvent, and the number of moles per liter (the molarity) changes, but the actual number of moles of solute does not change. For example, 125 mL of 0.551 M NaCl contains 68.9 millimol of NaCl. The solution will still contain 68.9 millimol of NaCl after the 250 mL of water is added to it, only now the 68.9 millimol of NaCl will be dispersed in a total volume of 375 mL. This gives the new molarity as 68.9 mmol/375 mL = 0.184 M. The volume and the concentration have changed, but the number of moles of solute in the solution has not changed.

28. \( P_1 \times V_1 = P_2 \times V_2 \)
   a. The pressure is being doubled, so the volume will become half its original value: 128 mL.
   b. The volume is being doubled, so the pressure will become half of its original value. Note that the actual pressures are not even mentioned in the problem. The fact that the volume increases by a factor of two means that the pressure must decrease by a factor of two regardless of the actual pressures.
30.  

a. \[ P_1 = 1.51 \text{ atm} \quad P_2 = 1.00 \text{ atm} \]
   \[ 2V_1 = 245 \text{ mL} \quad V_2 = ? \text{ mL} \]
   \[ 2T_1 = 29 \, ^\circ \text{C} = 302 \text{ K} \quad T_2 = 273 \text{ K} \]
   \[ V_2 = \frac{P_1V_1T_2}{P_1T_1} = \frac{(1.51 \text{ atm})(245 \text{ mL})(273 \text{ K})}{(1.00 \text{ atm})(302 \text{ K})} = 334 \text{ mL} \]

b. molar mass of N\(_2\) = 28.02 g
   
   \[ 21.6 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 0.771 \text{ mol N}_2 \]
   
   As one mole of an ideal gas occupies a volume of 22.4 L at STP
   
   \[ 0.771 \text{ mol N}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 17.3 \text{ L} \]

c. molar masses: He, 4.003 g; Ne, 20.18 g
   
   \[ 1.62 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} = 0.4046 \text{ mol He} \]
   
   \[ 2.41 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} = 0.1194 \text{ mol Ne} \]
   
   At STP, the total pressure is 1.00 atm.
   
   \[ P_{\text{He}} = \frac{0.4046 \text{ mol He}}{(0.4046 \text{ mol He} + 0.1194 \text{ mol Ne})} \times 1.00 \text{ atm} = 0.772 \text{ atm} \]
   
   \[ P_{\text{Ne}} = \frac{0.1194 \text{ mol Ne}}{(0.4046 \text{ mol He} + 0.1194 \text{ mol Ne})} \times 1.00 \text{ atm} = 0.228 \text{ atm} \]

d. molar mass of He, 4.003 g 11.2 mL = 0.0112 L
   
   One mole of an ideal gas has a volume of 22.4 L at STP.
   
   \[ 0.0112 \text{ L} \times \frac{1 \text{ mol He}}{22.4 \text{ L}} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 0.00200 \text{ g He} \]

32. molar masses: CaCO\(_3\), 100.09 g; CO\(_2\), 44.01 g
   
   \[ 1.25 \text{ g CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g}} = 0.01249 \text{ mol CaCO}_3 \]
   
   \[ 0.01249 \text{ mol CaCO}_3 \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} = 0.01249 \text{ mol CO}_2 \]
   
   \[ 0.01249 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 0.550 \text{ g CO}_2 \]
34. 
   a. mass of solution = 1.01 + 2.11 + 1.55 + 151 = 155.67 g
      \[ \frac{1.01 \text{ g NaCl}}{155.67 \text{ g}} \times 100 = 0.649\% \text{ NaCl} \]
   b. mass of solution (from part a) = 155.67 g
      molar mass NaCl = 58.44 g
      \[ \text{mol NaCl} = \frac{\text{mol Na}^+}{\text{mol NaCl}} = \frac{1.01 \text{ g NaCl}}{58.44 \text{ g NaCl}} = 0.01728 \text{ mol} \]
      volume = \[ \frac{155.67 \text{ g}}{1.02 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.1526 \text{ L} \]
      \[ M = \frac{0.01728 \text{ mol Na}^+}{0.1526 \text{ L}} = 0.113 M \text{ Na}^+ \text{ ion} \]
   c. molar masses: NaCl, 58.44; KCl, 74.55 g; CaCl₂, 111.0 g
      The total chloride ion in the solution comes from three different sources. We need to determine the number of moles of chloride ion contributed to the solution from each of the sources.
      from NaCl:
      \[ \text{mol Cl}^- = \text{mol NaCl} = \frac{1.01 \text{ g}}{58.44 \text{ g NaCl}} = 0.01728 \text{ mol} \]
      from KCl:
      \[ \text{mol Cl}^- = \text{mol KCl} = \frac{2.11 \text{ g}}{74.55 \text{ g KCl}} = 0.02830 \text{ mol} \]
      from CaCl₂:
      \[ \text{mol Cl}^- = 2 \times \text{mol CaCl₂} = 2 \times 1.55 \text{ g} \times \frac{1 \text{ mol CaCl₂}}{111.0 \text{ g CaCl₂}} = 0.02792 \text{ mol} \]
      total \( \text{mol Cl}^- = 0.01728 + 0.02830 + 0.02792 = 0.0735 \text{ mol} \)
      volume of solution (from part b) = 0.1526 L
      \[ M = \frac{0.0735 \text{ mol Cl}^-}{0.1526 \text{ L}} = 0.482 M \]

36. \( M_1 \times V_1 = M_2 \times V_2 \)
   a. \[ M_2 = \frac{(5.05 \text{ mL})(0.201 M)}{(5.05 + 10.0 \text{ mL})} = 0.0674 M \]
b. \[ M_2 = \frac{(25.2 \text{ mL})(0.561 \ M)}{(100. \text{ mL})} = 0.141 \ M \]

c. \[ M_2 = \frac{(8.33 \text{ mL})(12.1 \ M)}{(125 \text{ mL})} = 0.806 \ M \]

38.

a. \[ \frac{(41.5 \text{ mL})(0.118 \ M)(1)}{(0.242 \ M)(2)} = 10.1 \text{ mL H}_2\text{SO}_4 \]

b. \[ \frac{(27.1 \text{ mL})(0.121 \ M)(3)}{(0.242 \ M)(2)} = 20.3 \text{ mL H}_2\text{SO}_4 \]
CHAPTER 16

Acids and Bases

CHAPTER ANSWERS

2. acid; base

4. Conjugate acid–base pairs differ from each other by one proton (one hydrogen ion, \( \text{H}^+ \)). For example, \( \text{CH}_3\text{COOH} \) (acetic acid), differs from its conjugate base, \( \text{CH}_3\text{COO}^- \) (acetate ion), by a single \( \text{H}^+ \) ion.

\[
\text{CH}_3\text{COOH}(aq) \rightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}^+(aq)
\]

6. When an acid is dissolved in water, the hydronium ion (\( \text{H}_3\text{O}^+ \)) is formed. The hydronium ion is the conjugate *acid* of water (\( \text{H}_2\text{O} \)).

8.

a. \( \text{HSO}_4^- \) and \( \text{SO}_4^{2-} \) represent a conjugate acid–base pair. (\( \text{HSO}_4^- \) is the acid, \( \text{SO}_4^{2-} \) is the base; they differ by one proton.)

b. \( \text{HBr} \) and \( \text{Br}^- \) represent a conjugate acid–base pair. (\( \text{HBr} \) is the acid, \( \text{Br}^- \) is the base; they differ by one proton.)

c. \( \text{H}_2\text{PO}_4^- \) and \( \text{PO}_4^{3-} \) are not a conjugate acid–base pair; they differ by two protons. (\( \text{H}_2\text{PO}_4^- \) is the conjugate acid of \( \text{HPO}_4^{2-} \) and also the conjugate base of \( \text{H}_3\text{PO}_4 \); \( \text{HPO}_4^{2-} \) is the conjugate acid of \( \text{PO}_4^{3-} \)).

d. \( \text{HNO}_3 \) and \( \text{NO}_2^- \) are not a conjugate acid–base pair; they differ by an oxygen atom as well as a proton. (\( \text{NO}_3^- \) is the conjugate base of \( \text{HNO}_3 \); \( \text{NO}_2^- \) is the conjugate base of \( \text{HNO}_2 \).)

10.

a. \( \text{NH}_3(aq)(\text{base}) + \text{H}_2\text{O}(l)(\text{acid}) \iff \text{NH}_4^+(aq)(\text{acid}) + \text{OH}^-(aq)(\text{base}) \)

b. \( \text{NH}_4^+(aq)(\text{acid}) + \text{H}_2\text{O}(l)(\text{base}) \iff \text{NH}_3(aq)(\text{base}) + \text{H}_3\text{O}^+(aq)(\text{acid}) \)

c. \( \text{NH}_2^- (aq)(\text{base}) + \text{H}_2\text{O}(l)(\text{acid}) \rightarrow \text{NH}_3(aq)(\text{acid}) + \text{OH}^-(aq)(\text{base}) \)

12. The conjugate *acid* of the species indicated would have *one additional proton*:

a. \( \text{HBrO}_3 \)

b. \( \text{HF} \)

c. \( \text{HSO}_3^- \)

d. \( \text{H}_2\text{SO}_3 \)

14. The conjugate *bases* of the species indicated would have *one less proton*:

a. \( \text{BrO}^- \)

b. \( \text{HSO}_3^- \)

c. \( \text{SO}_3^{2-} \)
16. 
   a. $\text{CN}^- + \text{H}_2\text{O} \rightleftharpoons \text{HCN} + \text{OH}^-$  
   b. $\text{CO}_3^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^- + \text{OH}^-$  
   c. $\text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{PO}_4^{2-} + \text{H}_3\text{O}^+$  
   d. $\text{NH}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{OH}^-$  

18. To say that an acid is \textit{weak} in aqueous solution means that the acid does not easily transfer protons to water (and does not fully ionize). If an acid does not lose protons easily, then the acid’s anion must be a strong attractor of protons (good at holding on to protons).

20. A strong acid is one that loses its protons easily and fully ionizes in water. This means that the acid’s conjugate base must be poor at attracting and holding on to protons and is therefore a relatively weak base. A weak acid is one that resists loss of its protons and does not ionize well in water. This means that the acid’s conjugate base attracts and holds onto protons tightly and is a relatively strong base.

22. $\text{H}_2\text{SO}_4$ (sulfuric): $\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{HSO}_4^- + \text{H}_3\text{O}^+$  
   $\text{HCl}$ (hydrochloric): $\text{HCl} + \text{H}_2\text{O} \rightarrow \text{Cl}^- + \text{H}_3\text{O}^+$  
   $\text{HNO}_3$ (nitric): $\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{NO}_3^- + \text{H}_3\text{O}^+$  
   $\text{HClO}_4$ ( perchloric): $\text{HClO}_4 + \text{H}_2\text{O} \rightarrow \text{ClO}_4^- + \text{H}_3\text{O}^+$

24. An oxyacid is an acid containing a particular element that is bonded to one or more oxygen atoms. $\text{HNO}_3$, $\text{H}_2\text{SO}_4$, $\text{HClO}_4$ are oxyacids. $\text{HCl}$, $\text{HF}$, $\text{HBr}$ are not oxyacids.

26. Bases that are \textit{weak} have relatively strong conjugate acids:
   a. $\text{SO}_4^{2-}$ is a moderately weak base; $\text{HSO}_4^-$ is a moderately strong acid.
   b. $\text{Br}^-$ is a very weak base; $\text{HBr}$ is a strong acid.
   c. $\text{CN}^-$ is a fairly strong base; $\text{HCN}$ is a weak acid.
   d. $\text{CH}_3\text{COO}^-$ is a fairly strong base; $\text{CH}_3\text{COOH}$ is a weak acid.

28. For example, $\text{HCO}_3^-$ can behave as an acid if it reacts with something that more strongly gains protons than does $\text{HCO}_3^-$ itself. $\text{HCO}_3^-$ would behave as an acid when reacting with hydroxide ion (a much stronger base).
   
   $$\text{HCO}_3^-(aq) + \text{OH}^-(aq) \rightarrow \text{CO}_3^{2-}(aq) + \text{H}_2\text{O}(l).$$

   On the other hand, $\text{HCO}_3^-$ would behave as a base when reacted with something that more readily loses protons than does $\text{HCO}_3^-$ itself. For example, $\text{HCO}_3^-$ would behave as a base when reacting with hydrochloric acid (a much stronger acid).
   
   $$\text{HCO}_3^-(aq) + \text{HCl}(aq) \rightarrow \text{H}_2\text{CO}_3(aq) + \text{Cl}^-(aq)$$

   For $\text{H}_3\text{PO}_4^-$, similar equations can be written:
   
   $$\text{H}_3\text{PO}_4^-(aq) + \text{OH}^-(aq) \rightarrow \text{HPO}_4^{2-}(aq) + \text{H}_2\text{O}(l)$$  
   $$\text{H}_3\text{PO}_4^-(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{H}_3\text{PO}_4(aq) + \text{H}_2\text{O}(l)$$
30. The hydrogen ion concentration and the hydroxide ion concentration of water are not independent, they are related by the equilibrium

\[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq) \]

for which \( K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \) at 25 °C.

If the concentration of one of these ions is increased by addition of a reagent producing \( \text{H}^+ \) or \( \text{OH}^- \), then the concentration of the complementary ion will have to decrease so that the value of \( K_w \) will hold true. So if an acid is added to a solution, the concentration of hydroxide ion in the solution will decrease to a lower value. Similarly, if a base is added to a solution, then the concentration of hydrogen ion will have to decrease to a lower value.

32. \( K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \) at 25 °C

a. \( [\text{H}^+] = \frac{1.0 \times 10^{-14}}{4.22 \times 10^{-2} \, M} = 2.4 \times 10^{-11} \, M \); basic
b. \( [\text{H}^+] = \frac{1.0 \times 10^{-14}}{7.98 \times 10^{-9} \, M} = 1.3 \times 10^{-6} \, M \); acidic
c. \( [\text{H}^+] = \frac{1.0 \times 10^{-14}}{0.0104 \, M} = 9.6 \times 10^{-13} \, M \); basic
d. \( [\text{H}^+] = \frac{1.0 \times 10^{-14}}{6.55 \times 10^{-7} \, M} = 1.5 \times 10^{-8} \, M \); basic

34. \( K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \) at 25 °C

a. \( [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.34 \times 10^{-2} \, M} = 7.5 \times 10^{-13} \, M \); solution is acidic
b. \( [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{6.99 \times 10^{-7} \, M} = 1.4 \times 10^{-8} \, M \); solution is acidic
c. \( [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{4.01 \times 10^{-7} \, M} = 2.5 \times 10^{-6} \, M \); solution is basic
d. \( [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{4.02 \times 10^{-13} \, M} = 2.5 \times 10^{-2} \, M \); solution is basic

36.

a. \( [\text{OH}^-] = 0.105 \, M \)
b. \( [\text{OH}^-] = 5.22 \times 10^{-5} \, M \)
c. \( [\text{OH}^-] = 8.41 \times 10^{-2} \, M \)

38. Answer depends on student choice.

40. The pH of a solution is defined as the negative of the logarithm of the hydrogen ion concentration, \( \text{pH} = -\log[\text{H}^+] \). Mathematically, the negative sign in the definition causes the pH to decrease as the hydrogen ion concentration increases.

42. \( \text{pH} = -\log[\text{H}^+] \)

a. \( \text{pH} = -\log[0.0010 \, M] = 3.000 \); solution is acidic
b. \( \text{pH} = -\log[2.19 \times 10^{-4} \text{ M}] = 3.660; \) solution is acidic

c. \( \text{pH} = -\log[9.18 \times 10^{-11} \text{ M}] = 10.037; \) solution is basic

d. \( \text{pH} = -\log[4.71 \times 10^{-7} \text{ M}] = 6.327; \) solution is acidic

44.

a. \( \text{pH} = 1.983; \) solution is acidic

b. \( \text{pH} = 12.324; \) solution is basic

c. \( \text{pH} = 11.368; \) solution is basic

d. \( \text{pH} = 3.989; \) solution is acidic

46. \( \text{pOH} = 14 - \text{pH} \)

a. \( \text{pOH} = 14 - 7.45 = 6.55; \) solution is basic

b. \( \text{pOH} = 14 - 1.89 = 12.11; \) solution is acidic

c. \( \text{pOH} = 14 - 13.15 = 0.85; \) solution is basic

d. \( \text{pOH} = 14 - 5.55 = 8.45; \) solution is acidic

48.

a. \( \text{pH} = -\log[1.91 \times 10^{-2} \text{ M}] = 1.719; \) solution is acidic

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.91 \times 10^{-2}} \text{ M} = 5.2 \times 10^{-13} \text{ M} \]

b. \( \text{pH} = -\log[4.83 \times 10^{-7} \text{ M}] = 6.316; \) solution is acidic

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{4.83 \times 10^{-7}} \text{ M} = 2.1 \times 10^{-8} \text{ M} \]

c. \( \text{pH} = -\log[8.92 \times 10^{-11} \text{ M}] = 10.050; \) solution is basic

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{8.92 \times 10^{-11}} \text{ M} = 1.1 \times 10^{-4} \text{ M} \]

d. \( \text{pH} = -\log[6.14 \times 10^{-5} \text{ M}] = 4.212; \) solution is acidic

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{6.14 \times 10^{-5}} \text{ M} = 1.6 \times 10^{-10} \text{ M} \]

50. \( \text{pH} = 14 - \text{pOH} \quad [\text{H}^+] = \{\text{inv}\} \{\log\}[-\text{pH}] \)

a. \( \text{pH} = 14 - 4.96 = 9.04 \quad [\text{H}^+] = \{\text{inv}\} \{\log\}[-9.04] = 9.1 \times 10^{-10} \text{ M} \)

b. \( [\text{H}^+] = \{\text{inv}\} \{\log\}[-5.17] = 6.8 \times 10^{-6} \text{ M} \)

c. \( [\text{H}^+] = \{\text{inv}\} \{\log\}[-9.30] = 5.0 \times 10^{-10} \text{ M} \)

d. \( \text{pH} = 14 - 3.41 = 10.59 \quad [\text{H}^+] = \{\text{inv}\} \{\log\}[-10.59] = 2.6 \times 10^{-11} \text{ M} \)

52. \( \text{pH} + \text{pOH} = 14 \quad [\text{OH}^-] = \{\text{inv}\} \{\log\}[-\text{pOH}] \)

a. \( \text{pOH} = 14 - 1.01 = 12.99 \quad [\text{OH}^-] = \{\text{inv}\} \{\log\}[-12.99] = 1.0 \times 10^{-13} \text{ M} \)

b. \( [\text{OH}^-] = \{\text{inv}\} \{\log\}[-1.01] = 9.8 \times 10^{-2} \text{ M} \)

c. \( [\text{OH}^-] = \{\text{inv}\} \{\log\}[-11.22] = 6.0 \times 10^{-12} \text{ M} \)
d. \( \text{pOH} = 14 - 12.49 = 1.51 \) \( [\text{OH}^-] = \{\text{inv}\} \{\log\} [-1.51] = 3.1 \times 10^{-2} \) M

54.

a. \( \text{pH} = 14 - \text{pOH} = 14 - 11.31 = 2.69 \)

b. \( \text{pOH} = -\log[7.22 \times 10^{-5} \text{ M}] = 4.141 \) \( \text{pH} = 14 - \text{pOH} = 14 - 4.141 = 9.86 \)

c. \( \text{pH} = -\log[9.93 \times 10^{-4} \text{ M}] = 3.003 \)

d. \( \text{pOH} = -\log[1.49 \times 10^{-8} \text{ M}] = 7.827 \) \( \text{pH} = 14 - \text{pOH} = 14 - 7.827 = 6.17 \)

56. The solution contains water molecules, \( \text{H}_3\text{O}^+ \) ions (protons), and \( \text{NO}_3^- \) ions. Because \( \text{HNO}_3 \) is a strong acid that is completely ionized in water, there are no \( \text{HNO}_3 \) molecules present.

58.

a. \( \text{HNO}_3 \) is a strong acid and completely ionized so \( [\text{H}^+] = 1.21 \times 10^{-3} \text{ M} \) and \( \text{pH} = 2.917 \).

b. \( \text{HClO}_4 \) is a strong acid and completely ionized so \( [\text{H}^+] = 0.000199 \text{ M} \) and \( \text{pH} = 3.701 \).

c. \( \text{HCl} \) is a strong acid and completely ionized so \( [\text{H}^+] = 5.01 \times 10^{-5} \text{ M} \) and \( \text{pH} = 4.300 \).

d. \( \text{HBr} \) is a strong acid and completely ionized so \( [\text{H}^+] = 0.00104 \text{ M} \) and \( \text{pH} = 2.983 \).

60. A buffered solution consists of a mixture of a weak acid and its conjugate base. One example of a buffered solution is a mixture of acetic acid (\( \text{CH}_3\text{COOH} \)) and sodium acetate (\( \text{NaCH}_3\text{COO} \)).

62. The weak acid component of a buffered solution is capable of reacting with added strong base. For example, using the buffered solution given as an example in question 60, acetic acid would consume added sodium hydroxide as follows:

\[
\text{CH}_3\text{COOH}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCH}_3\text{COO}(aq) + \text{H}_2\text{O}(l).
\]

Acetic acid neutralizes the added \( \text{NaOH} \) and prevents it from having much effect on the overall \( \text{pH} \) of the solution.

64. \( \text{HCl}: \quad \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2^- \rightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \)

\( \text{NaOH}: \quad \text{OH}^- + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \)

66.

a. \( \text{NaOH} \) is completely ionized, so \( [\text{OH}^-] = 0.10 \text{ M} \).

\( \text{pOH} = -\log[0.10] = 1.00 \)

\( \text{pH} = 14 - 1.00 = 13.00 \)

b. \( \text{KOH} \) is completely ionized, so \( [\text{OH}^-] = 2.0 \times 10^{-4} \text{ M} \).

\( \text{pOH} = -\log[2.0 \times 10^{-4}] = 3.70 \)

\( \text{pH} = 14 - 3.70 = 10.30 \)

c. \( \text{CsOH} \) is completely ionized, so \( [\text{OH}^-] = 6.2 \times 10^{-3} \text{ M} \).

\( \text{pOH} = -\log[6.2 \times 10^{-3}] = 2.21 \)

\( \text{pH} = 14 - 2.21 = 11.79 \)

d. \( \text{NaOH} \) is completely ionized, so \( [\text{OH}^-] = 0.0001 \text{ M} \).

\( \text{pOH} = -\log[0.0001] = 4.0 \)

\( \text{pH} = 14 - 4.0 = 10.0 \)
68. b, c, and d

70. a, c, and e represent strong acids; b and d are typical weak acids.

72. Ordinarily, in calculating the pH of strong acid solutions, the major contribution to the concentration of hydrogen ion present is from the dissolved strong acid. We ordinarily neglect the small amount of hydrogen ion present in such solutions due to the ionization of water. With $1.0 \times 10^{-7}$ M HCl solution, however, the amount of hydrogen ion present due to the ionization of water is comparable to that present due to the addition of acid (HCl) and must be considered in the calculation of pH.

74. accepts

76. base

78. carboxyl (–COOH)  

$$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_3\text{O}^+$$

80. $1.0 \times 10^{-14}$

82. higher

84. pH

86. weak acid

88.

a. H$_2$O and OH$^-$ represent a conjugate acid–base pair (H$_2$O is the acid, having one more proton than the base, OH$^-$).

b. H$_2$SO$_4$ and SO$_4^{2-}$ are not a conjugate acid–base pair (they differ by two protons). The conjugate base of H$_2$SO$_4$ is HSO$_4^-$; the conjugate acid of SO$_4^{2-}$ is also HSO$_4^-$.

c. H$_3$PO$_4$ and H$_2$PO$_4^-$ represent a conjugate acid–base pair (H$_3$PO$_4$ is the acid, having one more proton than the base H$_2$PO$_4^-$).

d. HC$_2$H$_3$O$_2$ and C$_2$H$_3$O$_2^-$ represent a conjugate acid–base pair (HC$_2$H$_3$O$_2$ is the acid, having one more proton than the base C$_2$H$_3$O$_2^-$).

90. The conjugate acid of the species indicated would have one additional proton:

a. NH$_4^+$

b. NH$_3$

c. H$_3$O$^+$

d. H$_2$O

92. When an acid ionizes in water, a proton is released to the water as an H$_3$O$^+$ ion.

a. CH$_3$CH$_2$COOH + H$_2$O $\rightleftharpoons$ CH$_3$CH$_2$COO$^-$ + H$_3$O$^+$

b. NH$_4^+$ + H$_2$O $\rightleftharpoons$ NH$_3$ + H$_3$O$^+$

c. H$_2$SO$_4$ + H$_2$O $\rightarrow$ HSO$_4^-$ + H$_3$O$^+$

d. H$_3$PO$_4$ + H$_2$O $\rightleftharpoons$ H$_2$PO$_4^-$ + H$_3$O$^+$

94. K$_w$ = [H$^+$][OH$^-$] = $1.0 \times 10^{-14}$ at 25°C

a. $[H^+] = \frac{1.0 \times 10^{-14}}{4.22 \times 10^{-3}} M = 2.4 \times 10^{-12} M$; solution is basic
b. \[
[H^+] = \frac{1.0 \times 10^{-14}}{1.01 \times 10^{-13}} = 9.9 \times 10^{-2} \text{ M}; \text{ solution is acidic}
\]
c. \[
[H^+] = \frac{1.0 \times 10^{-14}}{3.05 \times 10^{-5}} = 3.3 \times 10^{-9} \text{ M}; \text{ solution is basic}
\]
d. \[
[H^+] = \frac{1.0 \times 10^{-14}}{6.02 \times 10^{-6}} = 1.7 \times 10^{-9} \text{ M}; \text{ solution is basic}
\]

96.

a. \([\text{OH}^-] = 0.0000032 \text{ M} \text{ is more basic}
\]
b. \([\text{OH}^-] = 1.54 \times 10^{-8} \text{ M} \text{ is more basic}
\]
c. \([\text{OH}^-] = 4.02 \times 10^{-7} \text{ M} \text{ is more basic}
\]

98. \(pOH = –\log[\text{OH}^-]\)  \(pH = 14 – pOH\)

a. \(pOH = –\log[1.4 \times 10^{-6} \text{ M}] = 5.85; \text{ pH} = 14 – 5.85 = 8.15; \text{ solution is basic}\)
b. \(pOH = –\log[9.35 \times 10^{-9} \text{ M}] = 8.029 = 8.03; \text{ pH} = 14 – 8.029 = 5.97; \text{ solution is acidic}\)
c. \(pOH = –\log[2.21 \times 10^{-1} \text{ M}] = 0.656 = 0.66; \text{ pH} = 14 – 0.656 = 13.34; \text{ solution is basic}\)
d. \(pOH = –\log[7.98 \times 10^{-12} \text{ M}] = 11.10; \text{ pH} = 14 – 11.098 = 2.90; \text{ solution is acidic}\)

100.

a. \([\text{OH}^-] = \frac{1.0 \times 10^{-14}}{5.72 \times 10^{-4} \text{ M}} = 1.75 \times 10^{-11} \text{ M} = 1.8 \times 10^{-11} \text{ M}
\]
\(pOH = –\log[1.75 \times 10^{-11} \text{ M}] = 10.76\)
\(pH = 14 – 10.76 = 3.24\)

b. \([H^+] = \frac{1.0 \times 10^{-14}}{8.91 \times 10^{-5} \text{ M}} = 1.12 \times 10^{-10} \text{ M} = 1.1 \times 10^{-10} \text{ M}
\]
\(pH = –\log[1.12 \times 10^{-10} \text{ M}] = 9.95\)
\(pOH = 14 – 9.95 = 4.05\)

c. \([\text{OH}^-] = \frac{1.0 \times 10^{-14}}{2.87 \times 10^{-12} \text{ M}} = 3.48 \times 10^{-3} \text{ M} = 3.5 \times 10^{-3} \text{ M}
\]
\(pOH = –\log[3.48 \times 10^{-3} \text{ M}] = 2.46\)
\(pH = 14 – 2.46 = 11.54\)

d. \([H^+] = \frac{1.0 \times 10^{-14}}{7.22 \times 10^{-8} \text{ M}} = 1.39 \times 10^{-7} \text{ M} = 1.4 \times 10^{-7} \text{ M}
\]
\(pH = –\log[1.39 \times 10^{-7} \text{ M}] = 6.86\)
\(pOH = 14 – 6.86 = 7.14\)

102. \(\text{pH} = 14 – \text{pOH}\)  \([H^+] = \{\text{inv}\} \{\log\}[-\text{pH}]\)

a. \([H^+] = \{\text{inv}\} \{\log\}[-5.41] = 3.9 \times 10^{-6} \text{ M}\)

b. \(\text{pH} = 14 – 12.04 = 1.96\)  \([H^+] = \{\text{inv}\} \{\log\}[-1.96] = 1.1 \times 10^{-2} \text{ M}\)
c. \([\text{H}^+] = \text{inv}\{\log\{-11.91\}\} = 1.2 \times 10^{-12}\ M\)

d. \[
\text{pH} = 14 - 3.89 = 10.11
\]
\[
[\text{H}^+] = \text{inv}\{\log\{-10.11\}\} = 7.8 \times 10^{-11}\ M
\]

104.

a. HClO₄ is a strong acid and completely ionized so \([\text{H}^+] = 1.4 \times 10^{-3}\ M\) and pH = 2.85.

b. HCl is a strong acid and completely ionized so \([\text{H}^+] = 3.0 \times 10^{-5}\ M\) and pH = 4.52.

c. HNO₃ is a strong acid and completely ionized so \([\text{H}^+] = 5.0 \times 10^{-2}\ M\) and pH = 1.30.

d. HCl is a strong acid and completely ionized so \([\text{H}^+] = 0.0010\ M\) and pH = 3.00.
CHAPTER 17

Equilibrium

CHAPTER ANSWERS


4. The symbol $E_a$ represents the activation energy of the reaction. The activation energy is the minimum energy two colliding molecules must possess in order for the collision to result in reaction. If molecules do not possess energies equal to or greater than $E_a$, a collision between these molecules will not result in a reaction.

6. Enzymes are biochemical catalysts that accelerate the complicated biochemical reactions in cells that would ordinarily be too slow to sustain life at normal body temperatures.

8. A state of equilibrium is attained when two opposing processes are exactly balanced so there is no further observable net change in the system.

10. Chemical equilibrium occurs when two opposing chemical reactions reach the same speed in a closed system. When a state of chemical equilibrium has been reached, the concentrations of reactants and products present in the system remain constant with time, and the reaction appears to “stop.” A chemical reaction that reaches a state of equilibrium is indicated by using a double arrow ($\rightleftharpoons$). The points of the double arrow point in opposite directions to indicate that two opposite processes are going on.

12. The two curves come together when a state of chemical equilibrium has been reached, after which point the forward and reverse reactions are occurring at the same rate so there is no further net change in concentration.

14. The equilibrium constant is a ratio of concentration of products to concentration of reactants with all concentrations measured at equilibrium. Depending on the amount of reactant present at the beginning of an experiment, there may be different absolute amounts of reactants and products present at equilibrium, but the ratio will always be the same for a given reaction at a given temperature. For example, the ratios (4/2) and (6/3) are different absolutely in terms of the numbers involved, but each of these ratios has the value of 2.

16.

- $a. \quad K = \frac{[\text{CO}(g)][\text{H}_2(g)]}{[\text{CH}_4(g)][\text{H}_2\text{O}(g)]}$

- $b. \quad K = \frac{[\text{O}_3(g)]^3}{[\text{O}_2(g)]}$

- $c. \quad K = \frac{[\text{CO}_2(g)]^6[\text{H}_2\text{O}(g)]^6}{[\text{C}_2\text{H}_6(g)]^6[O_(g)]}$
18.

a. \[ K = \frac{[\text{CH}_2\text{OH}]}{[\text{CO}][\text{H}_2]^2} \]

b. \[ K = \frac{[\text{NO}]^2[\text{O}_2]}{[\text{NO}_2]^2} \]

c. \[ K = \frac{[\text{PBr}_3]^4}{[\text{P}_2][\text{Br}_2]^4} \]

20. \[ 4\text{HCl}(g) + \text{O}_2(g) \rightleftharpoons 2\text{H}_2\text{O}(g) + 2\text{Cl}_2(g) \]

\[ K = \frac{[\text{H}_2\text{O}]^2[\text{Cl}_2]^2}{[\text{HCl}]^4[\text{O}_2]} = \frac{[8.3 \times 10^{-11}]^2[1.3 \times 10^{-10}]^2}{[7.1 \times 10^{-5}]^4[4.9 \times 10^{-5}]} = 9.4 \times 10^{-20} \]

22. \[ \text{I}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{ICl}(g) \]

\[ K = \frac{[\text{ICl}]^2}{[\text{I}_2][\text{Cl}_2]} = 1.2 \times 10^{10} \]

24. Equilibrium constants represent ratios of the concentrations of products and reactants present at the point of equilibrium. The concentration of a pure solid or of a pure liquid is constant and is determined by the density of the solid or liquid. For example, suppose you had a liter of water. Within that liter of water are 55.5 mol of water (the number of moles of water that is contained in one liter of water does not vary).

26.

a. \[ K = \frac{[\text{SF}_6(g)]}{[\text{F}_2(g)]^6} \]

b. \[ K = \frac{[\text{HCl}(g)]^2}{[\text{H}_2\text{S}(g)][\text{Cl}_2(g)]} \]

c. \[ K = \frac{[\text{Cl}_2\text{O}(g)]}{[\text{SO}_2(g)][\text{Cl}_2(g)]^2} \]

28.

a. \[ K = \frac{[\text{S}_2\text{Cl}_2(g)]}{[\text{CS}_2(g)][\text{Cl}_2(g)]^3} \]

b. \[ K = \frac{1}{[\text{Xe}(g)][\text{F}_2(g)]^5} \]

c. \[ K = \frac{1}{[\text{O}_2(g)]^3} \]
30. When an additional amount of one of the reactants is added to an equilibrium system, the system shifts to the right and adjusts so as to consume some of the added reactant. This results in a net increase in the amount of product compared to the equilibrium system before the additional reactant was added, and so the amount of CO$_2$(g) in the system will be higher than if the additional CO(g) had not been added. The numerical value of the equilibrium constant does not change when a reactant is added; the concentrations of all reactants and products adjust until the correct value of $K$ is once again achieved.

32. If heat is applied to an endothermic reaction (i.e., the temperature is raised), the equilibrium is shifted to the right. More product will be present at equilibrium than if the temperature had not been increased. The value of $K$ increases.

34. 
   a. shifts right (system reacts to get rid of excess chlorine)
   b. no change (S is solid)
   c. shifts right (system reacts to replace HCl)
   d. no change

36. 
   a. no change (B is solid)
   b. shifts right (system reacts to replace removed C)
   c. shifts left (system reacts by shifting in direction of fewer mol of gas)
   d. shifts right (the reaction is endothermic as written)

38. An increase in temperature favors the forward reaction for the endothermic reactions.
   a. no
   b. yes
   c. no
   d. yes

40. For an endothermic reaction, an increase in temperature will shift the position of equilibrium to the right (toward products).

42. CO(g) + 2H$_2$(g) $\rightleftharpoons$ CH$_3$OH(l)
   
   add additional CO(g) or H$_2$(g): The system will react in the forward direction to remove the excess.
   decrease the volume of the system: The system will react in the direction of fewer moles of gas.

44. A small equilibrium constant implies that not much product forms before equilibrium is reached. The reaction would not be a good source of the products unless Le Châtelier’s principle can be used to force the reaction to the right.

46. 
\[ K = \frac{[SO_2][NO_2]}{[SO_3][NO]} = \frac{[0.997 \text{ M}][0.781 \text{ M}]}{[0.0205 \text{ M}][0.0301 \text{ M}]} = 1.26 \times 10^3 \]
48. \[ K = 45.1 = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}]} = \frac{[0.998]^2}{[\text{H}_2][0.209]} \]
\[ [\text{H}_2] = 1.06 \times 10^{-1} \text{ M} \]

50. \[ K = 2.4 \times 10^{-3} = \frac{[\text{H}_2][\text{O}_2]}{[\text{H}_2\text{O}]^2} = \frac{[1.9 \times 10^{-2}]^2[\text{O}_2]}{[1.1 \times 10^{-1}]^2} \]
\[ [\text{O}_2] = 8.0 \times 10^{-2} \text{ M} \]

52. \[ K = 8.1 \times 10^{-3} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{[0.0021]M^2}{[\text{N}_2\text{O}_4]} \]
\[ [\text{N}_2\text{O}_4] = 5.4 \times 10^{-4} \text{ M} \]

54. solubility product, \( K_{sp} \)

56. Stirring or grinding the solute increases the speed with which the solute dissolves, but the ultimate amount of solute that dissolves is fixed by the equilibrium constant for the dissolving process, \( K_{sp} \), which changes only with temperature. Therefore only the temperature will affect the solubility.

58.

a. \( \text{Ni(OH)}_2(s) \rightleftharpoons \text{Ni}^{2+}(aq) + 2\text{OH}^-(aq) \)
\[ K_{sp} = [\text{Ni}^{2+}(aq)][\text{OH}^-(aq)]^2 \]

b. \( \text{Cr}_2\text{S}_3(s) \rightleftharpoons 2\text{Cr}^{3+}(aq) + 3\text{S}^2-(aq) \)
\[ K_{sp} = [\text{Cr}^{3+}(aq)]^2[\text{S}^2-(aq)]^3 \]

c. \( \text{Hg(OH)}_2(s) \rightleftharpoons \text{Hg}^{2+}(aq) + 2\text{OH}^-(aq) \)
\[ K_{sp} = [\text{Hg}^{2+}(aq)][\text{OH}^-(aq)]^2 \]

d. \( \text{Ag}_2\text{CO}_3(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) \)
\[ K_{sp} = [\text{Ag}^+(aq)]^2[\text{CO}_3^{2-}(aq)] \]

60. \( \text{Ag}_2\text{CrO}_4(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{CrO}_4^{2-}(aq) \)

Molar mass \( \text{Ag}_2\text{CrO}_4 = 315.8 \text{ g} \)

Let \( x \) represent the solubility of \( \text{Ag}_2\text{CrO}_4 \) in mol/L. Then \( [\text{CrO}_4^{2-}] = x \), and \( [\text{Ag}^+] = 2x \) from the stoichiometry of the equation.

\[ K_{sp} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}] = 1.1 \times 10^{-12} = (2x)^2(x) = 4x^3 \]
then the molar solubility of \( \text{Ag}_2\text{CrO}_4 = x = 6.5 \times 10^{-5} \text{ M} \)

gram solubility = \( 6.5 \times 10^{-5} \text{ mol/L} \times \frac{315.8 \text{ g}}{1 \text{ mol}} = 0.021 \text{ g/L} \)

62. \( \text{Ni(OH)}_2(s) \rightleftharpoons \text{Ni}^{2+}(aq) + 2\text{OH}^-(aq) \)

molar mass \( \text{Ni(OH)}_2 = 92.71 \text{ g} \)

let \( x \) represent the molar solubility of \( \text{Ni(OH)}_2 \); then \( [\text{Ni}^{2+}] = x \) and \( [\text{OH}^-] = 2x \).

\[ K_{sp} = [\text{Ni}^{2+}][\text{OH}^-]^2 = 2.0 \times 10^{-15} = [x][2x]^2 = 4x^3 \]
then the molar solubility of \( \text{Ni(OH)}_2 = x = 7.9 \times 10^{-6} \text{ M} \)

gram solubility = \( 7.98 \times 10^{-6} \text{ mol/L} \times \frac{92.71 \text{ g}}{1 \text{ mol}} = 7.4 \times 10^{-4} \text{ g/L} \)
64. \( \text{Cu}_2\text{S(s)} \rightleftharpoons 2\text{Cu}^+(aq) + \text{S}^2-(aq) \)

If \( \text{Cu}_2\text{S} \) dissolves to the extent of \( 1.8 \times 10^{-16} \text{ M} \), then \([\text{Cu}^+]\) will be \( 2(1.8 \times 10^{-16} \text{ M}) \) and \([\text{S}^2-]\) will be \( 1.8 \times 10^{-16} \text{ M} \).

\[ K_{sp} = [\text{Cu}^+]^2[\text{S}^2-] = [2(1.8 \times 10^{-16} \text{ M})]^2[1.8 \times 10^{-16} \text{ M}] = 2.3 \times 10^{-47} \]

66. \( \text{Cr(OH)}_3(s) \rightleftharpoons \text{Cr}^{3+}(aq) + 3\text{OH}-(aq) \)

If \( \text{Cr(OH)}_3 \) dissolves to the extent of \( 8.21 \times 10^{-5} \text{ M} \), then \([\text{Cr}^{3+}]\) will be \( 8.21 \times 10^{-5} \text{ M} \) and \([\text{OH}^-]\) will be \( 3(8.21 \times 10^{-5} \text{ M}) \) in a saturated solution.

\[ K_{sp} = [\text{Cr}^{3+}][\text{OH}^-]^3 = [8.21 \times 10^{-5} \text{ M}][8.21 \times 10^{-5} \text{ M}]^3 = 1.23 \times 10^{-15} \]

68. \( \text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \)

\[ K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2 \]

If \( \text{PbCl}_2 \) dissolves to the extent of \( 3.6 \times 10^{-2} \text{ M} \), then \([\text{Pb}^{2+}] = 3.6 \times 10^{-2} \text{ M} \) and \([\text{Cl}^-] = 2 \times (3.6 \times 10^{-2}) = 7.2 \times 10^{-2} \text{ M} \).

\[ K_{sp} = (3.6 \times 10^{-2} \text{ M})(7.2 \times 10^{-2} \text{ M})^2 = 1.9 \times 10^{-4} \]

Molar mass \( \text{PbCl}_2 = 278.1 \text{ g} \)

\[ \frac{3.6 \times 10^{-2} \text{ mol}}{1 \text{ L}} \times \frac{278.1 \text{ g}}{1 \text{ mol}} = 10. \text{ g/L} \]

70. \( \text{Fe(OH)}_3(s) \rightleftharpoons \text{Fe}^{3+}(aq) + 3\text{OH}^-(aq) \)

\[ K_{sp} = [\text{Fe}^{3+}][\text{OH}^-]^3 = 4 \times 10^{-38} \]

Let \( x \) represent the number of moles of \( \text{Fe(OH)}_3 \) that dissolve per liter; then \([\text{Fe}^{3+}] = x \).

The amount of hydroxide ion that would be produced by the dissolving of \( \text{Fe(OH)}_3 \) would then be \( 3x \), but pure water itself contains hydroxide ion at the concentration of \( 1.0 \times 10^{-7} \text{ M} \) (See Chapter 16.). The total concentration of hydroxide ion is then \([\text{OH}^-] = (3x + 1.0 \times 10^{-7}) \). As \( x \) must be a very small number [because \( \text{Fe(OH)}_3 \) is not very soluble], we can save ourselves a lot of arithmetic if we use the approximation that

\[ (3x + 1.0 \times 10^{-7}) = 1.0 \times 10^{-7} \]

\[ K_{sp} = [x][1.0 \times 10^{-7}]^3 = 4 \times 10^{-38} \]

\[ x = 4 \times 10^{-17} \text{ M} \]

Molar mass \( \text{Fe(OH)}_3 = 106.9 \text{ g} \)

\[ \frac{4 \times 10^{-17} \text{ mol}}{1 \text{ L}} \times \frac{106.9 \text{ g}}{1 \text{ mol}} = 4 \times 10^{-15} \text{ g/L} \]

72. An increase in temperature increases the fraction of molecules that possess sufficient energy for a collision to result in a reaction.
78. When we say that a chemical equilibrium is *dynamic* we are recognizing the fact that, even though
the reaction has appeared macroscopically to have stopped, on a microscopic basis the forward and
reverse reactions are still taking place at the same speed.

80. heterogeneous

82. position

84. Heat is considered a *product* of an exothermic process. Adding a product to a system in
equilibrium causes the reverse reaction to occur (producing additional reactants).

86. An equilibrium reaction may come to many *positions* of equilibrium, but at each possible position
of equilibrium the numerical value of the equilibrium constant is fulfilled. If different amounts of
reactant are taken in different experiments, the *absolute amounts* of reactant and product present
at the point of equilibrium reached will differ from one experiment to another, but the *ratio* that
defines the equilibrium constant will be the same.

88. \[ \text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g) \]

\[ K = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = 4.5 \times 10^{-3} \]

The concentration of PCl_5 is to be twice the concentration of PCl_3: \([\text{PCl}_5] = 2 \times [\text{PCl}_3]\)

\[ K = \frac{[\text{PCl}_3][\text{Cl}_2]}{2 \times [\text{PCl}_5]} = 4.5 \times 10^{-3} \]

\[ K = \frac{[\text{Cl}_2]}{2} = 4.5 \times 10^{-3} \quad \text{and} \quad [\text{Cl}_2] = 9.0 \times 10^{-3} \text{ M} \]

90. As all of the metal carbonates indicated have the metal ion in the +2 oxidation state, we can
illustrate the calculations for a general metal carbonate, MCO_3:

\[ \text{MCO}_3(s) \rightleftharpoons \text{M}^{2+}(aq) + \text{CO}_3^{2-}(aq) \quad K_{sp} = [\text{M}^{2+}(aq)][\text{CO}_3^{2-}(aq)] \]

If we then let \( x \) represent the number of moles of MCO_3 that dissolve per liter, then \([\text{M}^{2+}(aq)] = x \)
and \([\text{CO}_3^{2-}(aq)] = x \) also because the reaction is of 1:1 stoichiometry. Therefore,

\[ K_{sp} = [\text{M}^{2+}(aq)][\text{CO}_3^{2-}(aq)] = x^2 \] for each salt. Solving for \( x \) gives the following results.

\([\text{BaCO}_3] = x = 7.1 \times 10^{-5} \text{ M}\)

\([\text{CdCO}_3] = x = 2.3 \times 10^{-6} \text{ M}\)

\([\text{CaCO}_3] = x = 5.3 \times 10^{-5} \text{ M}\)

\([\text{CoCO}_3] = x = 3.9 \times 10^{-7} \text{ M}\)

92. Although a small solubility product generally implies a small solubility, comparisons of solubility
based directly on \( K_{sp} \) values are valid only if the salts produce the same numbers of positive and
negative ions per formula when they dissolve. For example, one can compare the solubilities of
AgCl(s) and NiS(s) directly using \( K_{sp} \), because each salt produces one positive and one negative
ion per formula when dissolved. One could not directly compare AgCl(s) with a salt such as
Ca_3(PO_4)_2, however.

94. At higher temperatures, the average kinetic energy of the reactant molecules is larger. At higher
temperatures, the probability that a collision between molecules will be energetic enough for
reaction to take place is larger. On a molecular basis, a higher temperature means a given
molecule will be moving faster.
96. 
   a.  \[ K = \frac{[HBr]^2}{[H_2][Br_2]} \]
   b.  \[ K = \frac{[H_S]^2}{[H_2][S_2]} \]
   c.  \[ K = \frac{[HCN]^2}{[H_2][C_2N_2]} \]

98.  \[ K = \frac{[Br]^2}{[Br_2]} = \frac{[0.034 M]^2}{[0.97 M]} = 1.2 \times 10^{-3} \]

100. 
   a.  \[ K = \frac{1}{[O_2]^3} \]
   b.  \[ K = \frac{1}{[NH_3][HCl]} \]
   c.  \[ K = \frac{1}{[O_2]} \]

102. An exothermic reaction is one that liberates heat energy. Increasing the temperature (adding heat) for such a reaction is fighting against the reaction’s own tendency to liberate heat. The net effect of raising the temperature will be a shift to the left to decrease the amount of product. If it is desired to increase the amount of product in an exothermic reaction, heat must be removed from the system. Changing the temperature does change the numerical value of the equilibrium constant for a reaction.

104. The reaction is exothermic as written. An increase in temperature (addition of heat) will shift the reaction to the left (toward reactants).

106.  \[ K = \frac{[NH_3]^2}{[N_2][H_2]} = 1.3 \times 10^{-2} = \frac{[NH_3]^2}{[0.1 M][0.1 M]^3} \]

   \[ [NH_3]^2 = 1.3 \times 10^{-6} \]
   \[ [NH_3] = 1.1 \times 10^{-3} M \]

108. 
   a.  \[ \text{Cu(OH)}_2(s) \rightleftharpoons \text{Cu}^{2+}(aq) + 2\text{OH}^- (aq) \]
      \[ K_{sp} = [\text{Cu}^{2+}][\text{OH}^-]^2 \]
   b.  \[ \text{Cr(OH)}_3(s) \rightleftharpoons \text{Cr}^{3+}(aq) + 3\text{OH}^- (aq) \]
      \[ K_{sp} = [\text{Cr}^{3+}][\text{OH}^-]^3 \]
   c.  \[ \text{Ba(OH)}_2(s) \rightleftharpoons \text{Ba}^{2+}(aq) + 2\text{OH}^- (aq) \]
      \[ K_{sp} = [\text{Ba}^{2+}][\text{OH}^-]^2 \]
d. \( \text{Sn(OH)}_2(s) \rightleftharpoons \text{Sn}^{2+}(aq) + 2\text{OH}^-(aq) \)

\[ K_{sp} = [\text{Sn}^{2+}][\text{OH}^-]^2 \]

110. molar mass \( \text{AgCl} = 143.4 \text{ g} \)

\[ 9.0 \times 10^{-4} \text{ g AgCl/L} \times \frac{1 \text{ mol AgCl}}{143.4 \text{ g AgCl}} = 6.28 \times 10^{-6} \text{ mol AgCl/L} \]

\( \text{AgCl(s)} \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq) \)

\[ K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (6.28 \times 10^{-6} \text{M})(6.28 \times 10^{-6} \text{M}) = 3.9 \times 10^{-11} \]

112. molar mass \( \text{Ni(OH)}_2 = 92.71 \text{ g} \)

\[ 0.14 \text{ g Ni(OH)}_2 \times \frac{1 \text{ mol}}{92.71 \text{ g Ni(OH)}_2} = 1.510 \times 10^{-3} \text{ M} \]

\( \text{Ni(OH)}_2(s) \rightleftharpoons \text{Ni}^{2+}(aq) + 2\text{OH}^- (aq) \)

\[ K_{sp} = [\text{Ni}^{2+}][\text{OH}^-]^2 \]

If \( 1.510 \times 10^{-3} \text{ M} \) of \( \text{Ni(OH)}_2 \) dissolves, then \([\text{Ni}^{2+}] = 1.510 \times 10^{-3} \text{ M} \) and \([\text{OH}^-] = 2 \times (1.510 \times 10^{-3} \text{ M}) = 3.020 \times 10^{-3} \text{ M} \).

\[ K_{sp} = (1.510 \times 10^{-3} \text{ M})(3.020 \times 10^{-3} \text{ M})^2 = 1.4 \times 10^{-8} \]

114. The activation energy is the minimum energy two colliding molecules must possess in order for the collision to result in reaction. If molecules do not possess energies equal to or greater than \( E_a \), a collision between these molecules will not result in a reaction.

116. Once a system has reached equilibrium the net concentration of product no longer increases because molecules of product already present react to form the original reactants. This is not to say that the same product molecules are necessarily always present.

118.

a. \[ K = [\text{H}_2\text{O(g)}][\text{CO}_2(g)] \]

b. \[ K = [\text{CO}_2] \]

c. \[ K = \frac{1}{[\text{O}_2]} \]
ANSWERS

2. A conjugate acid–base pair consists of two species related to each other by donation or acceptance of a single proton, $H^+$. An acid has one more $H^+$ than its conjugate base; a base has one less $H^+$ than its conjugate acid.

**Brønsted-Lowry acids:**

- $\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{Cl}^-(aq) + \text{H}_3\text{O}^+(aq)$
- $\text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{HSO}_4^-(aq) + \text{H}_3\text{O}^+(aq)$
- $\text{H}_3\text{PO}_4(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{PO}_4^-(aq) + \text{H}_3\text{O}^+(aq)$
- $\text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq)$

**Brønsted-Lowry bases:**

- $\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$
- $\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq) + \text{OH}^-(aq)$
- $\text{NH}_2^-(aq) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_3(aq) + \text{OH}^-(aq)$
- $\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{PO}_4(aq) + \text{OH}^-(aq)$

4. The strength of an acid is a direct result of the position of the acid’s ionization equilibrium. Strong acids are those whose ionization equilibrium positions lie far to the right whereas weak acids are those whose equilibrium positions lie only slightly to the right. For example, HCl, HNO$_3$, and HClO$_4$ are all strong acids, which means they are completely ionized in aqueous solution (the position of equilibrium is very far to the right):

- $\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{Cl}^-(aq) + \text{H}_3\text{O}^+(aq)$
- $\text{HNO}_3(aq) + \text{H}_2\text{O}(l) \rightarrow \text{NO}_3^-(aq) + \text{H}_3\text{O}^+(aq)$
- $\text{HClO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{ClO}_4^-(aq) + \text{H}_3\text{O}^+(aq)$

As these are very strong acids, we know their anions (Cl$^-$, NO$_3^-$, ClO$_4^-$) must be very weak bases and that solutions of the sodium salts of these anions would not be appreciably basic. As these acids have a strong tendency to lose protons, there is very little tendency for the anions (bases) to gain protons.

6. The pH of a solution is defined as the negative of the base ten logarithm of the hydrogen ion concentration in the solution; that is

$$\text{pH} = -\log_{10}[\text{H}^+]$$
In pure water, the amount of \( \text{H}^+ (aq) \) ion present is equal to the amount of \( \text{OH}^- (aq) \) ion, so we say that pure water is \textit{neutral}. As \([\text{H}^+] = 1.0 \times 10^{-7} \text{M}\) in pure water, this means the pH of pure water is \(-\log[1.0 \times 10^{-7} \text{M}] = 7.00\). Solutions in which the hydrogen ion concentration is greater than \(1.0 \times 10^{-7} \text{M}\) (pH < 7.00) are \textit{acidic}; solutions in which the hydrogen ion concentration is less than \(1.0 \times 10^{-7} \text{M}\) (pH > 7.00) are \textit{basic}. The pH scale is logarithmic. When the pH changes by one unit, this corresponds to a change in the hydrogen ion concentration by a factor of \textit{ten}.

In some instances, it may be more convenient to speak directly about the hydroxide ion concentration present in a solution, and so an analogous logarithmic expression is defined for the hydroxide ion concentration:

\[
p\text{OH} = -\log_{10}[\text{OH}^-].
\]

The concentrations of hydrogen ion and hydroxide ion in water (and in aqueous solutions) are \textit{not} independent of one another, but rather are related by the dissociation equilibrium constant for water,

\[
K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at 25°C}.
\]

From this constant it is obvious that \(p\text{H} + p\text{OH} = 14.00\) for water (or an aqueous solution) at 25°C.

8. Chemists envision that a reaction can take place between molecules only if the molecules physically \textit{collide} with each other. Furthermore, when molecules collide, the molecules must collide with enough force for the reaction to be successful (there must be enough energy to break bonds in the reactants), and the colliding molecules must be positioned with the correct relative orientation for the products (or intermediates) to form. Reactions tend to be faster if higher concentrations are used for the reaction because, if there are more molecules present per unit volume, there will be more collisions between molecules in a given time period. Reactions are faster at higher temperatures because at higher temperatures the reactant molecules have a higher average kinetic energy, and the number of molecules that will collide with sufficient force to break bonds increases.

10. Chemists define equilibrium as the exact balancing of two exactly opposing processes. When a chemical reaction is begun by combining pure reactants, the only process possible initially is

\[
\text{reactants} \rightarrow \text{products}
\]

However, for many reactions, as the concentration of product molecules increases, it becomes more likely that product molecules will collide and react with each other

\[
\text{products} \rightarrow \text{reactants}
\]

giving back molecules of the original reactants. At some point in the process the rates of the forward and reverse reactions become equal, and the system attains chemical equilibrium. To an outside observer, the system appears to have stopped reacting. On a microscopic basis, though, both the forward and reverse processes are still going on. Every time additional molecules of the product form, however, somewhere else in the system molecules of product react to give back molecules of reactant.

Once the point is reached that product molecules are reacting at the same speed at which they are forming, there is no further net change in concentration. A graph showing how the rates of the forward and reverse reactions change with time is given in the text as Figure 17.8. At the start of the reaction, the rate of the forward reaction is at its maximum whereas the rate of the reverse reaction is zero. As the reaction proceeds, the rate of the forward reaction gradually decreases as the concentration of reactants decreases, whereas the rate of the reverse reaction increases as the
concentration of products increases. Once the two rates have become equal, the reaction has reached a state of equilibrium.

12. The equilibrium constant for a reaction is a ratio of the concentration of products present at the point of equilibrium to the concentration of reactants still present. A ratio means that we have one number divided by another number (for example, the density of a substance is the ratio of a substance’s mass to its volume). As the equilibrium constant is a ratio, there are an infinite number of sets of data that can give the same ratio: for example, the ratios 8/4, 6/3, 100/50 all have the same value, 2. The actual concentrations of products and reactants will differ from one experiment to another involving a particular chemical reaction, but the ratio of the amount of product to reactant at equilibrium should be the same for each experiment.

Consider this simple example: Suppose we have a reaction for which $K = 4$, and we begin this reaction with 100 reactant molecules. At the point of equilibrium, there should be 80 molecules of product and 20 molecules of reactant remaining ($80/20 = 4$). Suppose we perform another experiment involving the same reaction, only this time we begin the experiment with 500 molecules of product present and 100 molecules of reactant remaining ($400/100 = 4$). As we began the two experiments with different numbers of reactant molecules, it’s not troubling that there are different absolute numbers of product and reactant molecules present at equilibrium; however, the ratio, $K$, is the same for both experiments. We say that these two experiments represent two different positions of equilibrium. An equilibrium position corresponds to a particular set of experimental equilibrium concentrations that fulfill the value of the equilibrium constant. Any experiment that is performed with a different amount of starting material will come to its own unique equilibrium position, but the equilibrium constant ratio, $K$, will be the same for a given reaction regardless of the starting amounts.

14. Your paraphrase of Le Châtelier’s principle should go something like this, “When you make any change to a system in equilibrium, this throws the system temporarily out of equilibrium, and the system responds by reacting in whichever direction will be able to reach a new position of equilibrium.” There are various changes that can be made to a system in equilibrium. Following are examples:

a. the concentration of one of the reactants is increased.

Consider the reaction: $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$

Suppose the reactants have already reacted, and a position of equilibrium has been reached that fulfills the value of $K$ for the reaction. At this point there will be present particular amounts of each reactant and of product. Suppose then one additional mole of $\text{O}_2$ is added to the system from outside. At the instant the additional $\text{O}_2$ is added, the system will not be in equilibrium; there will be too much $\text{O}_2$ present in the system to be compatible with the amounts of $\text{SO}_2$ and $\text{SO}_3$ present. The system will respond by reacting to get rid of some of the excess $\text{O}_2$ until the value of the ratio $K$ is again fulfilled. If the system reacts to get rid of the excess $\text{O}_2$, additional product $\text{SO}_3$ will form. The net result is more $\text{SO}_3$ produced than if the change had not been made.

b. The concentration of one of the products is decreased by selectively removing it from the system.

Consider the reaction: $\text{CH}_3\text{COOH} + \text{CH}_3\text{OH} \rightleftharpoons \text{H}_2\text{O} + \text{CH}_3\text{COOCH}_3$

This reaction is typical of many reactions involving organic chemical substances in which two organic molecules react to form a larger molecule with a molecule of water split out during the combination. This type of reaction on its own tends to come to equilibrium with
only part of the starting materials being converted to the desired organic product (which effectively would leave the experimenter with a mixture of materials). A technique that is used by organic chemists to increase the effective yield of the desired organic product is to separate the two products (if the products are separated, they cannot react to give back the reactants). One method used is to add a drying agent to the mixture. Such a drying agent chemically or physically absorbs the water from the system, removing it from equilibrium. If the water is removed, the reverse reaction cannot take place, and the reaction proceeds to a greater extent in the forward direction than if the drying agent had not been added. In other situations, an experimenter may separate the products of the reaction by distillation (if the boiling points make this possible). Again, if the products have been separated, then the reverse reaction will not be possible, and the forward reaction will occur to a greater extent.

c. The reaction system is compressed to a smaller volume.

Consider the example: \[3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})\]

For equilibria involving gases, when the volume of the reaction system is compressed suddenly, the pressure in the system increases. However, if the reacting system can relieve some of this increased pressure by reacting, it will do so. This will happen by the reaction occurring in whichever direction will give the smaller number of moles of gas (If the number of moles of gas is decreased in a particular volume, the pressure will decrease.).

For the reaction above, there are two moles of the gas on the right side of the equation, but there is a total of four moles on the left side. If this system at equilibrium were to be suddenly compressed to a smaller volume, the reaction would proceed farther to the right (in favor of more ammonia being produced).

d. The temperature is increased for an endothermic reaction.

Consider the reaction: \[2\text{NaHCO}_3 + \text{heat} \rightleftharpoons \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2\]

Although a change in temperature actually does change the value of the equilibrium constant, we can simplify reactions involving temperature changes by treating heat energy as if it were a chemical substance. For this endothermic reaction, heat is one of the reactants. As we saw in the example in part (a) of this question, increasing the concentration of one of the reactants for a system at equilibrium causes the reaction to proceed farther to the right, forming additional product. Similarly for the endothermic reaction given above, increasing the temperature causes the reaction to proceed farther in the direction of products than if no change had been made. It is as if there were too much “heat” to be compatible with the amount of substances present. The substances react to get rid of some of the energy.

e. The temperature is decreased for an exothermic process.

Consider the reaction: \[\text{PCl}_3 + \text{Cl}_2 \rightleftharpoons \text{PCl}_5 + \text{heat}\]

As discussed in part (d) above, although changing the temperature at which a reaction is performed does change the numerical value of \(K\), we can simplify our discussion of this reaction by treating heat energy as if it were a chemical substance. Heat is a product of this reaction. If we are going to lower the temperature of this reaction system, the only way to accomplish this is to remove energy from the system. Lowering the temperature of the system is really working with this system in its attempt to release heat energy. So lowering the temperature should favor the production of product more than if no change were made.
16. Specific answer depends on student choice of examples. In general, for a weak acid, HA, and a weak base, B:

\[
\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^- \quad \text{B} + \text{H}_2\text{O} \rightleftharpoons \text{HB}^+ + \text{OH}^- 
\]

18. 

a. \(\text{NH}_3(\text{aq})(\text{base}) + \text{H}_2\text{O}(\text{l})(\text{acid}) \rightleftharpoons \text{NH}_4^+(\text{aq})(\text{acid}) + \text{OH}^-(\text{aq})(\text{base})\)

b. \(\text{H}_2\text{SO}_4(\text{aq})(\text{acid}) + \text{H}_2\text{O}(\text{l})(\text{base}) \rightleftharpoons \text{HSO}_4^-(\text{aq})(\text{base}) + \text{H}_3\text{O}^+(\text{aq})(\text{acid})\)

c. \(\text{O}_2^-(\text{s})(\text{base}) + \text{H}_2\text{O}(\text{l})(\text{acid}) \rightleftharpoons \text{OH}^-(\text{aq})(\text{acid}) + \text{OH}^-(\text{aq})(\text{base})\)

d. \(\text{NH}_2^- (\text{aq})(\text{base}) + \text{H}_2\text{O}(\text{l})(\text{acid}) \rightleftharpoons \text{NH}_3(\text{aq})(\text{acid}) + \text{OH}^- (\text{aq})(\text{base})\)

e. \(\text{H}_2\text{PO}_4^-(\text{aq})(\text{acid}) + \text{OH}^- (\text{aq})(\text{base}) \rightleftharpoons \text{HPO}_4^{2-}(\text{aq})(\text{base}) + \text{H}_2\text{O}(\text{l})(\text{acid})\)

20. 

a. \(\text{HNO}_3\) is a strong acid, so

\[
[\text{H}^+] = 0.00141 \, \text{M} \\
\text{pH} = –\log(0.00141) = 2.851 \\
\text{pOH} = 14 – 2.851 = 11.15
\]

b. \(\text{NaOH}\) is a strong base, so

\[
[\text{OH}^-] = 2.13 \times 10^{-3} \, \text{M} \\
\text{pOH} = –\log(2.13 \times 10^{-3}) = 2.672 \\
\text{pH} = 14 – 2.672 = 11.33
\]

c. \(\text{HCl}\) is a strong acid, so

\[
[\text{H}^+] = 0.00515 \, \text{M} \\
\text{pH} = –\log(0.00515) = 2.288 \\
\text{pOH} = 14 – 2.288 = 11.71
\]

d. \(\text{Ca(OH)}_2\) is a strong but not very soluble base. Each formula unit of \(\text{Ca(OH)}_2\) produces two formula units of \(\text{OH}^-\) ion.

\[
[\text{OH}^-] = 2 \times 5.65 \times 10^{-5} \, \text{M} = 1.13 \times 10^{-4} \, \text{M} \\
\text{pOH} = –\log(1.13 \times 10^{-4}) = 3.947 \\
\text{pH} = 14 – 3.947 = 10.05
\]

22. \(2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})\)

\[
K = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} = \frac{(0.42)^2}{[1.4 \times 10^{-3}]^2[4.5 \times 10^{-4}]} = 2.0 \times 10^8
\]
24. MgCO₃ ⇌ Mg²⁺(aq) + CO₃²⁻(aq)

\[ K_{sp} = [Mg^{2+}][CO_3^{2-}] \]

let \( x \) represent the number of moles of MgCO₃ that dissolve per liter. Then [Mg²⁺] = \( x \) and [CO₃²⁻] = \( x \) also

\[ K_{sp} = [x][x] = x^2 = 6.82 \times 10^{-6} \]

\( x = [MgCO_3] = 2.61 \times 10^{-3} \text{ M} \)

molar mass of MgCO₃ = 84.32 g

\[ \frac{2.61 \times 10^{-3} \text{ mol}}{1 \text{ L}} \times \frac{84.32 \text{ g}}{1 \text{ mol}} = 0.220 \text{ g/L} \]
CHAPTER 18

Oxidation–Reduction
Reactions/Electrochemistry

CHAPTER ANSWERS

2. Oxidation can be defined as the loss of electrons by an atom, molecule, or ion. Oxidation can also be defined as an increase in oxidation state for an element, but because elements can increase their oxidation states only by losing electrons, the two definitions are equivalent. The following equation shows the oxidation of copper metal to copper(II) ion.

\[ \text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^- \]

Reduction can be defined as the gaining of electrons by an atom, molecule, or ion. Reduction may also be defined as a decrease in oxidation state for an element, but naturally such a decrease takes place by the gaining of electrons (so the two definitions are equivalent). The following equation shows the reduction of sulfur atoms to sulfide ion.

\[ \text{S} + 2\text{e}^- \rightarrow \text{S}^{2-} \]

4. Each of these reactions involves a free element on one side of the equation. On the other side of the equation the element is combined in an ionic compound.
   a. potassium is being oxidized, oxygen is being reduced
   b. iodine is being oxidized, chlorine is being reduced
   c. cobalt is being oxidized, chlorine is being reduced
   d. carbon is being oxidized, oxygen is being reduced.

6. Each of these reactions involves a free element on one side of the equation. On the other side of the equation, the element is combined in a compound.
   a. sulfur is being oxidized, oxygen is being reduced
   b. phosphorus is being oxidized, oxygen is being reduced
   c. hydrogen is being oxidized, carbon is being reduced
   d. boron is being oxidized, hydrogen is being reduced.

8. Oxidation numbers represent a “relative charge” one atom has compared to another in a compound. In an element, all the atoms are equivalent.

10. Fluorine is always assigned a negative oxidation state \((-1)\) because all other elements are less electronegative. The other halogens are usually assigned an oxidation state of \(-1\) in compounds. In interhalogen compounds such as ClF, fluorine is assigned oxidation state \(-1\) (F is more electronegative than Cl). Chlorine, therefore, must be assigned a +1 oxidation state in this instance.

12. 3–
14. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element in the following answers is given in parentheses after the element and its oxidation state.
   a. Cr +3 (Rule 6); Cl –1 (Rule 5)
   b. Cu +1 (Rule 6); O –2 (Rule 3)
   c. Cu +2 (Rule 6); O –2 (Rule 3)
   d. 0 (Rule 1)

16. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element in the following answers is given in parentheses after the element and its oxidation state.
   a. Al +3 (Rule 2); P +5 (Rule 6); O –2 (Rule 3)
   b. Mn +4 (Rule 6); O –2 (Rule 3)
   c. Ba +2 (Rule 2); C +4 (Rule 6); O –2 (Rule 3)
   d. Cl +1 (Rule 6); F –1 (Rule 5)

18. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element in the following answers is given in parentheses after the element and its oxidation state.
   a. H +1 (Rule 4); P +5 (Rule 6); O –2 (Rule 3)
   b. H +1 (Rule 4); Br +1 (Rule 6); O –2 (Rule 3)
   c. H +1 (Rule 4); N +5 (Rule 6); O –2 (Rule 3)
   d. H +1 (Rule 4); Cl +7 (Rule 6); O –2 (Rule 3)

20. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element in the following answers is given in parentheses after the element and its oxidation state.
   a. K +1 (Rule 2); Cl +5 (Rule 6); O –2 (Rule 3)
   b. 0 (Rule 1)
   c. C +2 (Rule 6); O –2 (Rule 3)
   d. Na +1 (Rule 2); I +5 (Rule 6); O –2 (Rule 3)

22. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element in the following answers is given in parentheses after the element and its oxidation state.
   a. Fe +3 (Rule 6); O –2 (Rule 3)
   b. Al +3 (Rule 2); C +4 (Rule 6); O –2 (Rule 3)
   c. Ba +2 (Rule 2); Cr +6 (Rule 6); O –2 (Rule 3)
   d. Ca +2 (Rule 2); H +1 (Rule 4); C +4 (Rule 6); O –2 (Rule 3)

24. Electrons are negative; when an atom gains electrons, it gains one negative charge for each electron gained. For example, in the reduction reaction Cl + e → Cl, the oxidation state of chlorine decreases from 0 to –1 as the electron is gained.
26. Answer depends on student example chosen.

28. An oxidizing agent oxidizes another species by gaining the electrons lost by the other species; therefore, an oxidizing agent itself decreases in oxidation state. A reducing agent increases its oxidation state when acting on another atom or molecule.

30. 
   a. \[ \text{Mn}(s) + 2\text{HCl}(aq) \rightarrow \text{MnCl}_2(aq) + \text{H}_2(g) \]
      manganese is being oxidized, hydrogen is being reduced
   b. \[ 2\text{Fe}_2\text{S}_3(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) + 6\text{S}(s) \]
      sulfur is being oxidized, oxygen is being reduced
   c. \[ 2\text{Al}(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(s) + 3\text{H}_2(g) \]
      aluminum is being oxidized, hydrogen is being reduced
   d. \[ 2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \]
      nitrogen is being oxidized, oxygen is being reduced

32. 
   a. \[ 4\text{KClO}_3(s) + \text{C}_6\text{H}_12\text{O}_6(s) \rightarrow 4\text{KCl}(s) + 6\text{H}_2\text{O}(l) + 6\text{CO}_2(g) \]
      carbon is being oxidized, chlorine is being reduced
   b. \[ 2\text{C}_8\text{H}_{18}(l) + 25\text{O}_2(g) \rightarrow 16\text{CO}_2(g) + 18\text{H}_2\text{O}(l) \]
      carbon is being oxidized, oxygen is being reduced
   c. \[ \text{PCl}_3(g) + \text{Cl}_2(g) \rightarrow \text{PCl}_5(g) \]
      phosphorus is being oxidized, chlorine is being reduced
   d. \[ \text{Ca}(s) + \text{H}_2(g) \rightarrow \text{CaH}_2(g) \]
      calcium is being oxidized, hydrogen is being reduced

34. Iron is reduced \([+3 \text{ in Fe}_2\text{O}_3(s), 0 \text{ in Fe}(l)]\); carbon is oxidized \([+2 \text{ in CO}(g), +4 \text{ in CO}_2(g)]\). \text{Fe}_2\text{O}_3(s)\) is the oxidizing agent; \text{CO}(g)\) is the reducing agent.

36. 
   a. chlorine is being reduced, iodine is being oxidized; chlorine is the oxidizing agent, iodide ion is the reducing agent
   b. iron is being reduced, iodine is being oxidized; iron(III) is the oxidizing agent, iodide ion is the reducing agent
   c. copper is being reduced, iodine is being oxidized; copper(II) is the oxidizing agent, iodide ion is the reducing agent

38. Oxidation–reduction reactions are often more complicated than “regular” reactions. Frequently the coefficients necessary to balance the number of electrons transferred come out to be large numbers. We also have to make certain that we account for the electrons being transferred.

40. Under ordinary conditions it is impossible to have “free” electrons that are not part of some atom, ion, or molecule. For this reason, the total number of electrons lost by the species being oxidized must equal the total number of electrons gained by the species being reduced.
42.  
   a.  \( \text{Cl}^- (aq) \rightarrow \text{Cl}_2(g) + 2e^- \)  
      balance chlorine: \( 2\text{Cl}^- (aq) \rightarrow \text{Cl}_2(g) \)  
      balance charge: \( 2\text{Cl}^- (aq) \rightarrow \text{Cl}_2(g) + 2e^- \)  
      balanced half-reaction: \( 2\text{Cl}^- (aq) \rightarrow \text{Cl}_2(g) + 2e^- \)  
   b.  \( \text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq) + e^- \)  
   c.  \( \text{Fe}(s) \rightarrow \text{Fe}^{3+}(aq) + 3e^- \)  
   d.  \( \text{Cu}^{2+}(aq) + e^- \rightarrow \text{Cu}^+(aq) \)  

44.  
   a.  \( \text{NO}_3^-(aq) \rightarrow \text{NH}_4^+(aq) \)  
      balance oxygen: \( \text{NO}_3^-(aq) \rightarrow \text{NH}_4^+(aq) + 3\text{H}_2\text{O(l)} \)  
      balance hydrogen: \( 10\text{H}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{NH}_4^+(aq) + 3\text{H}_2\text{O(l)} \)  
      balance charge: \( 8e^- + 10 \text{H}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{NH}_4^+(aq) + 3\text{H}_2\text{O(l)} \)  
      balanced half-reaction: \( 8e^- + 10 \text{H}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{NH}_4^+(aq) + 3\text{H}_2\text{O(l)} \)  
   b.  \( \text{C}_2\text{N}_2(g) \rightarrow \text{HCN}(aq) \)  
      balance nitrogen: \( \text{C}_2\text{N}_2(g) \rightarrow 2\text{HCN}(aq) \)  
      balance hydrogen: \( 2\text{H}^+(aq) + \text{C}_2\text{N}_2(g) \rightarrow 2\text{HCN}(aq) \)  
      balance charge: \( 2e^- + 2\text{H}^+(aq) + \text{C}_2\text{N}_2(g) \rightarrow 2\text{HCN}(aq) \)  
      balanced half-reaction: \( 2e^- + 2\text{H}^+(aq) + \text{C}_2\text{N}_2(g) \rightarrow 2\text{HCN}(aq) \)  
   c.  \( \text{ClO}_3^-(aq) \rightarrow \text{Cl}^-(aq) \)  
      balance oxygen: \( \text{ClO}_3^-(aq) \rightarrow \text{Cl}^-(aq) + 3\text{H}_2\text{O(l)} \)  
      balance hydrogen: \( 6\text{H}^+(aq) + \text{ClO}_3^-(aq) \rightarrow \text{Cl}^-(aq) + 3\text{H}_2\text{O(l)} \)  
      balance charge: \( 6e^- + 6\text{H}^+(aq) + \text{ClO}_3^-(aq) \rightarrow \text{Cl}^-(aq) + 3\text{H}_2\text{O(l)} \)  
      balanced half-reaction: \( 6e^- + 6\text{H}^+(aq) + \text{ClO}_3^-(aq) \rightarrow \text{Cl}^-(aq) + 3\text{H}_2\text{O(l)} \)  
   d.  \( \text{MnO}_2(s) \rightarrow \text{Mn}^{2+}(aq) \)  
      balance oxygen: \( \text{MnO}_2(s) \rightarrow \text{Mn}^{2+}(aq) + 2\text{H}_2\text{O(l)} \)  
      balance hydrogen: \( 4\text{H}^+(aq) + \text{MnO}_2(s) \rightarrow \text{Mn}^{2+}(aq) + 2\text{H}_2\text{O(l)} \)  
      balance charge: \( 2e^- + 4\text{H}^+(aq) + \text{MnO}_2(s) \rightarrow \text{Mn}^{2+}(aq) + 2\text{H}_2\text{O(l)} \)  
      balanced half-reaction: \( 2e^- + 4\text{H}^+(aq) + \text{MnO}_2(s) \rightarrow \text{Mn}^{2+}(aq) + 2\text{H}_2\text{O(l)} \)
46. For simplicity, the physical states of the substances have been omitted until the final balanced equation is given.

a. \( \text{Al}(s) + \text{H}^+(aq) \rightarrow \text{Al}^{3+}(aq) + \text{H}_2(g) \)

\[ \text{Al} \rightarrow \text{Al}^{3+} \]
Balance charge: \( \text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^- \)

\[ \text{H}^+ \rightarrow \text{H}_2 \]
Balance hydrogen: \( 2\text{H}^+ \rightarrow \text{H}_2 \)
Balance charge: \( 2\text{e}^- + 2\text{H}^+ \rightarrow \text{H}_2 \)

Combine half-reactions:

\[ 3 \times (2\text{e}^- + 2\text{H}^+ \rightarrow \text{H}_2) \]
\[ 2 \times (\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-) \]

\[ 2\text{Al}(s) + 6\text{H}^+(aq) \rightarrow 2\text{Al}^{3+}(aq) + 3\text{H}_2(g) \]

b. \( \text{S}^{2-}(aq) + \text{NO}_3^-(g) \rightarrow \text{S}(s) + \text{NO}(g) \)

\[ \text{S}^{2-} \rightarrow \text{S} \]
Balance charge: \( \text{S}^{2-} \rightarrow \text{S} + 2\text{e}^- \)

\[ \text{NO}_3^- \rightarrow \text{NO} \]
Balance oxygen: \( \text{NO}_3^- \rightarrow \text{NO} + 2\text{H}_2\text{O} \)
Balance hydrogen: \( 4\text{H}^+ + \text{NO}_3^- \rightarrow \text{NO} + 2\text{H}_2\text{O} \)
Balance charge: \( 3\text{e}^- + 4\text{H}^+ + \text{NO}_3^- \rightarrow \text{NO} + 2\text{H}_2\text{O} \)

Combine half-reactions:

\[ 3 \times (\text{S}^{2-} \rightarrow \text{S} + 2\text{e}^-) \]
\[ 2 \times (3\text{e}^- + 4\text{H}^+ + \text{NO}_3^- \rightarrow \text{NO} + 2\text{H}_2\text{O}) \]

\[ 8\text{H}^+ + 3\text{S}^{2-}(aq) + 2\text{NO}_3^-(g) \rightarrow 3\text{S}(s) + 2\text{NO}(g) + 4\text{H}_2\text{O} \]

c. \( \text{I}_2(aq) + \text{Cl}_2(aq) \rightarrow \text{IO}_3^-(aq) + \text{HCl}(g) \)

\[ \text{I}_2 \rightarrow \text{IO}_3^- \]
Balance iodine: \( \text{I}_2 \rightarrow 2\text{IO}_3^- \)
Balance oxygen: \( 6\text{H}_2\text{O} + \text{I}_2 \rightarrow \text{IO}_3^- \)
Balance hydrogen: \( 6\text{H}_2\text{O} + \text{I}_2 \rightarrow \text{IO}_3^- + 12\text{H}^+ \)
Balance charge: \( 6\text{H}_2\text{O} + \text{I}_2 \rightarrow \text{IO}_3^- + 12\text{H}^+ + 10\text{e}^- \)

\[ \text{Cl}_2 \rightarrow \text{HCl} \]
Balance chlorine: \( \text{Cl}_2 \rightarrow 2\text{HCl} \)
Balance hydrogen: \( 2\text{H}^+ + \text{Cl}_2 \rightarrow 2\text{HCl} \)
Balance charge: \( 2\text{e}^- + 2\text{H}^+ + \text{Cl}_2 \rightarrow 2\text{HCl} \)
Combine half-reactions:
\[5 \times (2e^- + 2H^+ + Cl_2 \rightarrow 2HCl)\]
\[6H_2O + I_2 \rightarrow IO_3^- + 12H^+ + 10e^-\]
\[6H_2O(l) + 2I_2(aq) + 5Cl_2(aq) \rightarrow 2IO_3^-(aq) + 10HCl(g) + 2H^+(aq)\]
d. \[\text{AsO}_4^{2-}(aq) + S^{2-}(aq) \rightarrow \text{AsO}_3^-(s) + S(s)\]
\[\text{AsO}_4^{2-} \rightarrow \text{AsO}_3^-\]
Balance oxygen: \[\text{AsO}_4^{2-} \rightarrow \text{AsO}_3^- + H_2O\]
Balance hydrogen: \[2H^+ + \text{AsO}_4^{2-} \rightarrow \text{AsO}_3^- + H_2O\]
Balance charge: \[2e^- + 2H^+ + \text{AsO}_4^{2-} \rightarrow \text{AsO}_3^- + H_2O\]
\[S^{2-} \rightarrow S\]
Balance charge: \[S^{2-} \rightarrow S + 2e^-\]
\[2H^+(aq) \text{AsO}_4^{2-}(aq) + S^{2-}(aq) \rightarrow \text{AsO}_3^-(s) + S(s) + H_2O(l)\]
48. \[\text{Cu}(s) + 2\text{HNO}_3(aq) + 2H^+(aq) \rightarrow \text{Cu}^{2+}(aq) + 2\text{NO}_2(g) + 2H_2O(l)\]
\[\text{Mg}(s) + 2\text{HNO}_3(aq) \rightarrow \text{Mg(NO}_3)_2(aq) + \text{H}_2(g)\]
50. A salt bridge typically consists of a \(U\)-shaped tube filled with an inert electrolyte (one involving ions that are not part of the oxidation–reduction reaction). A salt bridge is used to complete the electrical circuit in a cell. Any method that allows transfer of charge without allowing bulk mixing of the solutions may be used (another common method is to set up one half-cell in a porous cup that is then placed in the beaker containing the second half-cell).
52. Reduction takes place at the cathode, and oxidation takes place at the anode.
54.
\[\text{Pb}^{2+}(aq) \text{ion is reduced; } \text{Zn}(s) \text{ is oxidized.}\]
The reaction at the anode is \[\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^-\].
The reaction at the cathode is \[\text{Pb}^{2+}(aq) + 2e^- \rightarrow \text{Pb}(s)\]
56. \[\text{Cd} + 2\text{OH}^- \rightarrow \text{Cd(OH)}_2 + 2e^- \text{ (oxidation)}\]
\[\text{NiO}_2 + 2\text{H}_2\text{O} + 2e^- \rightarrow \text{Ni(OH)}_2 + 2\text{OH}^- \text{ (reduction)}\]
58. Aluminum is a very reactive metal when freshly isolated in the pure state. However, on standing for even a relatively short period of time, aluminum metal forms a thin coating of Al₂O₃ on its surface from reaction with atmospheric oxygen. This coating of Al₂O₃ is much less reactive than the metal and serves to protect the surface of the metal from further attack.

60. The magnesium is used for cathodic protection of the steel pipeline. Magnesium is more reactive than iron and will be oxidized in preference to the iron of the pipeline.

62. The main recharging reaction for the lead storage battery is

\[ 2\text{PbSO}_4(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Pb}(s) + \text{PbO}_2(s) + 2\text{H}_2\text{SO}_4(aq). \]

A major side reaction is the electrolysis of water

\[ 2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g). \]

This results in production of an explosive mixture of hydrogen and oxygen, which accounts for many accidents in recharging of such batteries.

64. Electrolysis is applied in electroplating by making the item to be plated the cathode in a cell containing a solution of ions of the desired plating metal.

66. loss; oxidation state

68. electronegative

70. An oxidizing agent is an atom, molecule, or ion that causes the oxidation of another species. During this process, the oxidizing agent is reduced.

72. lose

74. separate from

76. oxidation

78. An electrolysis reaction results when an electrical current from an outside source is used to cause an otherwise nonspontaneous reaction to occur. An example is the electrolysis of water: \( 2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g). \) This reaction takes place only if an electrical current of sufficient voltage is passed through the water.

80. hydrogen; oxygen

82. oxidation

84.

a. \( 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \)
   iron is oxidized; oxygen is reduced

b. \( 2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s) \)
   aluminum is oxidized; chlorine is reduced

c. \( 6\text{Mg}(s) + \text{P}_4(s) \rightarrow 2\text{Mg}_3\text{P}_2(s) \)
   magnesium is oxidized; phosphorus is reduced

86.

a. aluminum is oxidized; hydrogen is reduced

b. hydrogen is reduced; iodine is oxidized

c. copper is oxidized; hydrogen is reduced
88.

a. \( \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)

b. \( \text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{OH}(l) \)

c. \( \text{SnO}_2(s) + 2\text{C}(s) \rightarrow \text{Sn}(s) + 2\text{CO}(g) \)

d. \( \text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g) \)

90. Each of these reactions involves a metallic element in the form of the free element on one side of the equation. On the other side of the equation, the metallic element is combined in an ionic compound. If a metallic element goes from the free metal to the ionic form, the metal is oxidized (loses electrons).

a. sodium is oxidized; oxygen is reduced

b. iron is oxidized; hydrogen is reduced

c. oxygen \((\text{O}^2-)\) is oxidized; aluminum \((\text{Al}^{3+})\) is reduced (This reaction is the reverse of the type discussed above.)

d. magnesium is oxidized; nitrogen is reduced

92. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element in the following answers is given in parentheses after the element and its oxidation state.

a. \( \text{H} +1 \) (Rule 4); \( \text{N} -3 \) (Rule 6)

b. \( \text{C} +2 \) (Rule 6); \( \text{O} -2 \) (Rule 3)

c. \( \text{C} +4 \) (Rule 6); \( \text{O} -2 \) (Rule 3)

d. \( \text{N} +3 \) (Rule 6); \( \text{F} -1 \) (Rule 5)

94. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element is that given in parentheses after the element and its oxidation state.

a. \( \text{Mn} +4 \) (Rule 6); \( \text{O} -2 \) (Rule 3)

b. \( \text{Ba} +2 \) (Rule 2); \( \text{Cr} +6 \) (Rule 6); \( \text{O} -2 \) (Rule 3)

c. \( \text{H} +1 \) (Rule 4); \( \text{S} +4 \) (Rule 6); \( \text{O} -2 \) (Rule 3)

d. \( \text{Ca} +2 \) (Rule 2); \( \text{P} +5 \) (Rule 6); \( \text{O} -2 \) (Rule 3)

96. The rules for assigning oxidation states are given in Section 18.2 of the text. The rule that applies for each element is that given in parentheses after the element and its oxidation state.

a. \( \text{Bi} +3 \) (Rule 7); \( \text{O} -2 \) (Rule 3)

b. \( \text{P} +5 \) (Rule 7); \( \text{O} -2 \) (Rule 3)

c. \( \text{N} +3 \) (Rule 7); \( \text{O} -2 \) (Rule 3)

d. \( \text{Hg} +1 \) (Rule 7)

98.

a. \( 2\text{B}_2\text{O}_3(s) + 6\text{Cl}_2(g) \rightarrow 4\text{BCl}_3(l) + 3\text{O}_2(g) \)

  oxygen is oxidized \((-2\) to 0\); chlorine is reduced \((0\) to \(-1\))
b. \( \text{GeH}_4(g) + \text{O}_2(g) \rightarrow \text{Ge}(s) + 2\text{H}_2\text{O}(g) \)
germanium is oxidized (–4 to 0); oxygen is reduced (0 to –2)
c. \( \text{C}_2\text{H}_4(g) + \text{Cl}_2(g) \rightarrow \text{C}_2\text{H}_4\text{Cl}_2(l) \)
carbon is oxidized (–2 to –1); chlorine is reduced (0 to –1)
d. \( \text{O}_2(g) + 2\text{F}_2(g) \rightarrow 2\text{OF}_2(g) \)
oxygen is oxidized (0 to +2); fluorine is reduced (0 to –1)

100.
a. \( \text{SiO}_2(s) \rightarrow \text{Si}(s) \)
Balance oxygen: \( \text{SiO}_2(s) \rightarrow \text{Si}(s) + 2\text{H}_2\text{O}(l) \)
Balance hydrogen: \( \text{SiO}_2(s) + 4\text{H}^+(aq) \rightarrow \text{Si}(s) + 2\text{H}_2\text{O}(l) \)
Balance charge: \( \text{SiO}_2(s) + 4\text{H}^+(aq) + 4\text{e}^- \rightarrow \text{Si}(s) + 2\text{H}_2\text{O}(l) \)
Balanced half-reaction: \( \text{SiO}_2(s) + 4\text{H}^+(aq) + 4\text{e}^- \rightarrow \text{Si}(s) + 2\text{H}_2\text{O}(l) \)
b. \( \text{S}(s) \rightarrow \text{H}_2\text{S}(g) \)
Balance hydrogen: \( \text{S}(s) + 2\text{H}^+(aq) \rightarrow \text{H}_2\text{S}(g) \)
Balance charge: \( \text{S}(s) + 2\text{H}^+(aq) + 2\text{e}^- \rightarrow \text{H}_2\text{S}(g) \)
Balanced half-reaction: \( \text{S}(s) + 2\text{H}^+(aq) + 2\text{e}^- \rightarrow \text{H}_2\text{S}(g) \)
c. \( \text{NO}_3^-(aq) \rightarrow \text{HNO}_2(aq) \)
Balance oxygen: \( \text{NO}_3^-(aq) \rightarrow \text{HNO}_2(aq) + \text{H}_2\text{O}(l) \)
Balance hydrogen: \( \text{NO}_3^-(aq) + 3\text{H}^+(aq) \rightarrow \text{HNO}_2(aq) + \text{H}_2\text{O}(l) \)
Balance charge: \( \text{NO}_3^-(aq) + 3\text{H}^+(aq) + 2\text{e}^- \rightarrow \text{HNO}_2(aq) + \text{H}_2\text{O}(l) \)
Balanced half-reaction: \( \text{NO}_3^-(aq) + 3\text{H}^+(aq) + 2\text{e}^- \rightarrow \text{HNO}_2(aq) + \text{H}_2\text{O}(l) \)
d. \( \text{NO}_3^-(aq) \rightarrow \text{NO}(g) \)
Balance oxygen: \( \text{NO}_3^-(aq) \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) \)
Balance hydrogen: \( \text{NO}_3^-(aq) + 4\text{H}^+(aq) \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) \)
Balance charge: \( \text{NO}_3^-(aq) + 4\text{H}^+(aq) + 3\text{e}^- \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) \)
Balanced half-reaction: \( \text{NO}_3^-(aq) + 4\text{H}^+(aq) + 3\text{e}^- \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) \)

102. For simplicity, the physical states of the substances have been omitted until the final balanced equation is given.

For the reduction of the permanganate ion, \( \text{MnO}_4^- \), in acid solution, the half-reaction is always the same:
\[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \]
Balance oxygen: \( \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \)
Balance hydrogen: \( 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \)
Balance charge: \( 8\text{H}^+ + \text{MnO}_4^- + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \)
a. $C_2O_4^{2-} \rightarrow CO_2$
Balance carbon: $C_2O_4^{2-} \rightarrow 2CO_2$
Balance charge: $C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$
Combine half-reactions:

$$5 \times (C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-)$$
$$2 \times (8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O)$$

$$16H^+(aq) + 2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(g)$$

b. $Fe^{2+} \rightarrow Fe^{3+}$
Balance charge: $Fe^{2+} \rightarrow Fe^{3+} + e^-$
Combine half-reactions:

$$5 \times (Fe^{2+} \rightarrow Fe^{3+} + e^-)$$
$$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$8H^+(aq) + MnO_4^-(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$$

c. $Cl^- \rightarrow Cl_2$
Balance chlorine: $2Cl^- \rightarrow Cl_2$
Balance charge: $2Cl^- \rightarrow Cl_2 + 2e^-$
Combine half-reactions:

$$5 \times (2Cl^- \rightarrow Cl_2 + 2e^-)$$
$$2 \times (8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O)$$

$$16H^+(aq) + 2MnO_4^-(aq) + 10Cl^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Cl_2(g)$$

104. A galvanic cell is a battery. A spontaneous oxidation–reduction reaction is separated physically into the two half-reactions, and the electrons being transferred between the two half-cells are made available as an electrical current.
CHAPTER 19

Radioactivity and Nuclear Energy

CHAPTER ANSWERS

2. The radius of a typical atomic nucleus is on the order of $10^{-13}$ cm, which is about one hundred thousand times smaller than the radius of an atom overall.

4. The mass number represents the total number of protons and neutrons in a nucleus.

6. The atomic number ($Z$) is written in such formulas as a left subscript, whereas the mass number ($A$) is written as a left superscript. That is, the general symbol for a nuclide is $^{A}Z X$. As an example, consider the isotope of oxygen with eight protons and eight neutrons. Its symbol would be $^{16}O$.

8. a neutron

10. Answer depends on students’ choice of equation. In general, the atomic number of the parent nucleus increases by one unit while the mass number remains the same.

$$^{14}_{6}C \rightarrow ^{0}_{-1}e + ^{14}_{7}N$$

12. Gamma rays are high-energy photons of electromagnetic radiation. Gamma rays are not considered to be particles. When a nucleus produces only gamma radiation, the atomic number and mass number of the nucleus do not change. Gamma rays represent the energy changes associated with transitions and rearrangement of the particles within the nucleus.

14. Electron capture occurs when one of the inner orbital electrons is pulled into and becomes part of the nucleus.

16. $^{7}_{3}Li$ and $^{6}_{3}Li$; Since there is a much larger abundance of $^{7}_{3}Li$, the average mass number will be closer to seven.

18. The approximate atomic molar mass could be calculated as follows:

$$0.79(24) + 0.10(25) + 0.11(26) = 24.3.$$ 

This is only an approximation because the mass numbers, rather than the actual isotopic masses, were used. The fact that the approximate mass calculated is slightly above 24 shows that the isotope of mass number 24 predominates.

20.

a. electron

b. positron

c. neutron

d. proton
22. 
   a. \(^0\)\(_{-1}\)e
   b. \(^{14}\)_7\(\text{N}\)
   c. \(^0\)\(_{-1}\)e

24. 
   a. \(^{218}\)_86\(\text{Rn}\)
   b. \(^0\)\(_{+1}\)e (positron)
   c. \(^{137}\)_56\(\text{Ba}\)

26. 
   a. \(^{137}\)_55\(\text{Cs} \rightarrow ^0\)\(_{-1}\)e + \(^{137}\)_56\(\text{Ba}\)
   b. \(^3\)_1\(\text{H} \rightarrow ^0\)\(_{-1}\)e + \(^3\)_2\(\text{He}\)
   c. \(^{216}\)_84\(\text{Po} \rightarrow ^0\)\(_{-1}\)e + \(^{216}\)_85\(\text{At}\)

28. 

<table>
<thead>
<tr>
<th>Parent Nuclide</th>
<th>Particle Emitted</th>
</tr>
</thead>
<tbody>
<tr>
<td>Th-232</td>
<td>(\alpha)</td>
</tr>
<tr>
<td>Ra-228</td>
<td>(\beta)</td>
</tr>
<tr>
<td>Ac-228</td>
<td>(\beta)</td>
</tr>
<tr>
<td>Th-228</td>
<td>(\beta)</td>
</tr>
<tr>
<td>Ra-224</td>
<td>(\alpha)</td>
</tr>
<tr>
<td>Rn-220</td>
<td>(\alpha)</td>
</tr>
<tr>
<td>Po-216</td>
<td>(\alpha)</td>
</tr>
<tr>
<td>Pb-212</td>
<td>(\alpha)</td>
</tr>
<tr>
<td>Bi-212</td>
<td>(\beta)</td>
</tr>
<tr>
<td>Po-212</td>
<td>(\beta)</td>
</tr>
<tr>
<td>Tl-208</td>
<td>(\alpha)</td>
</tr>
<tr>
<td>Pb-208</td>
<td>(\beta)</td>
</tr>
</tbody>
</table>

30. In a nuclear bombardment process, a target nucleus is bombarded with high-energy particles (typically subatomic particles or small atoms) from a particle accelerator. This may result in the transmutation of the target nucleus into some other element. For example, nitrogen-14 may be transmuted into oxygen-17 by bombardment with alpha particles. There is often considerable repulsion between the target nucleus and the particles being used for bombardment (especially if the bombarding particle is positively charged like the target nucleus). Using accelerators to increases the kinetic energy of the bombarding particles can overcome this repulsion.

32. \(^{27}\)_{13}\(\text{Al} + ^4\)_2\(\text{He} \rightarrow ^{30}\)_15\(\text{P} + ^1\)_0\(\text{n}\)
34. The half-life of a nucleus is the time required for one-half of the original sample of nuclei to decay. A given isotope of an element always has the same half-life although different isotopes of the same element may have greatly different half-lives. Nuclei of different elements typically have different half-lives.

36. $^{226}_{88}$Ra is the most stable (longest half-life); $^{224}_{88}$Ra is the “hottest” (shortest half-life)

38. With a half-life of 2.6 hours, strontium-87 is the hottest; with a half-life of 45.1 days, iron-59 is the most stable to decay.

40.

<table>
<thead>
<tr>
<th>Time, min</th>
<th>0</th>
<th>2.5</th>
<th>5.0</th>
<th>7.5</th>
<th>10.0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass, µg</td>
<td>100</td>
<td>50</td>
<td>25</td>
<td>12.5</td>
<td>6.25</td>
</tr>
</tbody>
</table>

After four half-lives a little over 6 µg will remain.

42. For an administered dose of 100 µg, 0.39 µg remains after 2 days. The fraction remaining is $0.39/100 = 0.0039$; on a percentage basis, less than 0.4% of the original radioisotope remains.

44. Carbon-14 is produced in the upper atmosphere by the bombardment of ordinary nitrogen with neutrons from space:

$$^{14}_7\text{N} + ^1_0\text{n} \rightarrow ^{14}_6\text{C} + ^1_1\text{H}$$

46. We assume that the concentration of C-14 in the atmosphere is effectively constant. A living organism is constantly replenishing C-14 either through the processes of metabolism (sugars ingested in foods contain C-14), or photosynthesis (carbon dioxide contains C-14). When a plant dies, it can no longer replenish, and as the C-14 undergoes radioactive decay, its amount decreases with time.

48.

a. thyroid gland
b. heart muscle
c. bones, heart, liver, lungs
d. circulatory system

50. fission, fusion, fusion, fission

52. $^1_0\text{n} + ^{235}_{92}\text{U} \rightarrow ^{142}_{56}\text{Ba} + ^{91}_{36}\text{Kr} + 3^1_0\text{n}$ is one possibility.

54. A critical mass of a fissionable material is the amount needed to provide a high-enough internal neutron flux to sustain the chain reaction (enough neutrons are produced to cause the continuous fission of further material). A sample with less than a critical mass is still radioactive but cannot sustain a chain reaction.

56. An actual nuclear explosion of the type produced by a nuclear weapon cannot occur in a nuclear reactor because the concentration of the fissionable materials is not sufficient to form a supercritical mass. However, since many reactors are cooled by water, which can decompose into hydrogen and oxygen gases, a chemical explosion is possible that could scatter the radioactive material used in the reactor.
58. Breeder reactors use the more common uranium-238 isotope, which is bombarded with neutrons. Through a three-step process, the uranium-238 is converted to plutonium-239, which is fissionable and can then be used to fuel another reactor. The major controversy over breeder reactors concerns the extreme chemical toxicity of plutonium and the difficulty in handling it.

60. In one type of fusion reactor, two $^1_1\text{H}$ atoms are fused to produce $^4_2\text{He}$. Because the hydrogen nuclei are positively charged, extremely high energies (temperatures of 40 million K) are needed to overcome the repulsion between the nuclei as they are shot into each other.

62. In the theory of stellar nucleosynthesis, it is considered that the nucleus began as a cloud of neutrons that exploded (the Big Bang). After this initial explosion, neutrons were thought to have decomposed into protons and electrons

$$^1_0\text{n} \rightarrow ^1_1\text{H} + ^0_1\text{e}$$

The products of this decomposition were then thought to have combined to form large clouds of hydrogen atoms. As the hydrogen clouds became larger, gravitational forces caused these clouds to contract and heat up. Eventually the clouds of hydrogen were so dense and so hot that fusion of hydrogen nuclei into helium nuclei took place, with a great release of energy. When the tendency for the hydrogen clouds to expand from the heat of fusion was counter-balanced by the gravitational forces of the cloud, a small star had formed. In addition to the fusion of hydrogen nuclei into helium mentioned already, as the star’s hydrogen supply is exhausted, the helium present in the star also begins to undergo fusion into nuclei of other elements.

64. Somatic damage is directly to the organism itself, causing nearly immediate sickness or death to the organism. Genetic damage is to the genetic machinery of the organism, which will be manifested in future generations of offspring.

66. Gamma rays penetrate long distances, but seldom cause ionization of biological molecules. Alpha particles, because they are much heavier although less penetrating, are very effective at ionizing biological molecules and leaving a dense trail of damage in the organism. Isotopes that release alpha particles can be ingested or breathed into the body where the damage from the alpha particles will be more acute.

68. Nuclear waste may remain radioactive for thousands of years, and much of it is chemically poisonous as well as radioactive. Most reactor waste is still in “temporary storage.” Various suggestions have been made for a more permanent solution, such as casting the spent fuel into glass bricks to contain it, and then storing the bricks in corrosion-proof metal containers deep underground. No agreement on a permanent solution to the disposal of nuclear waste has yet been reached.

70. radioactive
72. mass
74. neutron; proton
76. radioactive decay
78. mass number
80. transuranium
82. half-life
84. radiotracers
86. chain
88. breeder

90. \( 4.5 \times 10^9 \) dollars ($4.5 billion)

92. \( 3.5 \times 10^{-11} \) J/atom; \( 8.9 \times 10^{10} \) J/g

94. \(^{90}\text{Zr} , ^{91}\text{Zr} , ^{92}\text{Zr} , ^{94}\text{Zr} , \) and \(^{96}\text{Zr} \)

96. \(^{27}\text{Al} \): 13 protons, 14 neutrons

\(^{28}\text{Al} \): 13 protons, 15 neutrons

\(^{29}\text{Al} \): 13 protons, 16 neutrons

98.

a. \(^0\text{e}^{-}\)

b. \(^{74}\text{Se} \)

c. \(^{240}\text{U} \)

100. \(^4\text{Be} + \(^2\text{He} \rightarrow ^{12}\text{C} + \, ^0\text{n} \)

102. \(^{238}\text{U} + \, ^0\text{n} \rightarrow ^{239}\text{U} \)

\(^{239}\text{U} \rightarrow ^{239}\text{Np} + ^0\text{e}^{-}\)

\(^{239}\text{Np} \rightarrow ^{239}\text{Pu} + ^0\text{e}^{-}\)
CHAPTER 20

Organic Chemistry

CHAPTER ANSWERS

2. A given carbon atom can be attached to a maximum of four other atoms. Carbon atoms have four valence electrons. By making four bonds, carbon atoms exactly complete their valence octet.

4. A triple bond represents the sharing of six electrons (three pairs of electrons). The simplest example of an organic molecule containing a triple bond is acetylene, H–C≡C–H.

6. $\overset{\equiv}{C} \overset{\equiv}{C}$

8. An unsaturated hydrocarbon is one that contains an area of multiple bonding (double bond or triple bond) between some of the carbon atoms. Such a hydrocarbon is called “unsaturated” because each carbon atom of the multiple bond could bond to additional atoms (rather than to each other). Specific examples will depend on students’ choice.

10. 

a. 
$$\text{H}_3\text{C} \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_3$$

b. 
$$\text{H}_3\text{C} \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_3$$

c. 
$$\text{H}_3\text{C} \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_3$$

d. 
$$\text{H}_3\text{C} \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_2 \overset{\equiv}{-} \text{CH}_3$$

12. The general formula for the alkanes is $\text{C}_n\text{H}_{2n+2}$.

a. $2(3) + 2 = 8$

b. $2(11) + 2 = 24$

c. $2(20) + 2 = 42$

d. $2(5) + 2 = 12$

14. 

a. hexane  
$$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$$

$$\begin{array}{ccccccc}
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{H} \\
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\end{array}$$
Chapter 20: Organic Chemistry

b. dodecane \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3 \)

\[
\begin{array}{c}
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} \\
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\end{array}
\]

c. butane \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3 \)

\[
\begin{array}{c}
\text{H} & \text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{C} & \text{C} & \text{H} \\
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\end{array}
\]

d. octane \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3 \)

\[
\begin{array}{c}
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} \\
\text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\end{array}
\]

16. A branched alkane contains one or more shorter carbon atom chains attached to the side of the main (longest) carbon atom chain. The simplest branched alkane is 2-methylpropane.

\[
\begin{array}{c}
\text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
\text{H} & \text{CH}_3 & \text{H} \\
\end{array}
\]

18. Structures depend on students’ choices.

20. The root name is derived from the number of carbon atoms in the longest continuous chain of carbon atoms.

22. The numbers indicate to which carbon of the longest continuous chain of carbon atoms the substituents are attached. The longest continuous chain is numbered from the end closest to the first substituent so as to give the lowest possible locator numbers.

24. Multiple substituents are listed in alphabetical order, disregarding any prefix.

26. Look for the longest continuous chain of carbon atoms.

   a. 2,3,4-trimethylpentane
   b. 2,3-dimethylpentane
   c. 3,4-dimethylhexane
   d. 4,5-dimethylheptane

28.

   a. 

   \[
   \begin{array}{c}
   \text{H} & \text{H} & \text{CH}_3 & \text{H} & \text{H} & \text{H} & \text{H} \\
   \text{H} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{H} \\
   \text{H} & \text{CH}_3 & \text{CH}_3 & \text{H} & \text{H} & \text{H} & \text{H} \\
   \end{array}
   \]
30. C atoms    Use
C_5–C_12    gasoline
C_10–C_18    kerosene, jet fuel
C_15–C_25    diesel fuel, heating oil, lubrication
C_25–    asphalt

32. Tetraethyl lead was added to gasolines to prevent “knocking” of high efficiency automobile engines. The use of tetraethyl lead is being phased out because of the danger to the environment posed by the lead in this substance.

34. The combustion of alkanes has been used as a source of heat, light, and mechanical energy.

\[
C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) + \text{heat}
\]

36. Equations depend on student choice. Here are some examples.

Substitution: \( CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g) \)

Dehydrogenation: \( C_2H_6(g) \xrightarrow{\text{catalyst}} C_2H_4(g) + H_2(g) \)

38.

a. \( 2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g) \)

b. \( CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(l) + HCl(g) \)

c. \( CHCl_3(l) + Cl_2(g) \rightarrow CCl_4(l) + HCl(g) \)

40. An alkyne is a hydrocarbon containing a carbon-carbon triple bond: general formula is \( C_nH_{2n-2} \).

42. The location of a double or triple bond in the longest chain of an alkene or alkyne is indicated by giving the number of the lowest number carbon atom involved in the double or triple bond.
44. Hydrogenation converts unsaturated compounds to saturated (or less unsaturated) compounds. In the case of a liquid vegetable oil, this is likely to convert the oil to a solid.

\[ \text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g}) \]

46. 
   a. 2-chloro-1-butene
   b. 3-chloro-1-butene
   c. 3-chloro-2-butene
   d. 1-chloro-2-butene

48. Shown are carbon skeletons:

\[ \text{C} \equiv \text{C} \equiv \text{C} \equiv \text{C} \equiv \text{C} \]

50. For benzene, a set of equivalent Lewis structures can be drawn that differ only in the location of the three double bonds in the ring. Experimentally, however, benzene does not demonstrate the chemical properties expected for molecules having any double bonds. We say that the “extra” electrons that would go into making the second bond of the three double bonds are delocalized around the entire benzene ring; this delocalization of the electrons explains benzene’s unique properties.

52. When named as a substituent, the benzene ring is called the phenyl group. Two examples are:

3-phenyl-1-butene

2-phenylhexane

54. ortho—refers to adjacent substituents (1,2–); meta—refers to two substituents with one unsubstituted carbon atom between them (1,3–); para—refers to two substituents with two unsubstituted carbon atoms between them (1,4–).

56. 
   a. anthracene
   b. 1,3,5-trimethylbenzene
   c. 1,4-dinitrobenzene
   d. 4-bromotoluene (4-bromo-1-methylbenzene)

58. 
   a. organic acids
   b. aldehydes
c. ketones
d. ethers

60. Primary alcohols have one hydrocarbon fragment (alkyl group) bonded to the carbon atom where the –OH group is attached. Secondary alcohols have two such alkyl groups attached, and tertiary alcohols contain three such alkyl groups. Examples are:

ethanol (primary)
\[
\text{CH}_3\text{–CH}_2\text{–OH}
\]
2-propanol (secondary)
\[
\begin{array}{c}
\text{CH}_3 \\
\text{OH} \\
\end{array}
\text{CH}_2\text{–CH}_3
\]
2-methyl-2-propanol (tertiary)
\[
\begin{array}{c}
\text{CH}_3 \\
\text{OH} \\
\end{array}
\text{CH}_3\text{–C–CH}_3
\]

62. Specific examples will depend on students’ choices. Below are examples for alcohols with five carbon atoms.

primary (1-pentanol)
\[
\text{CH}_3\text{–CH}_2\text{–CH}_2\text{–CH}_2\text{–CH}_2\text{–OH}
\]
secondary (2-pentanol)
\[
\begin{array}{c}
\text{CH}_3 \\
\text{OH} \\
\end{array}
\text{CH}_2\text{–CH}_2\text{–CH–CH}_3
\]
tertiary (2-methyl-2-butanol)
\[
\begin{array}{c}
\text{CH}_3 \\
\text{OH} \\
\end{array}
\text{CH}_3\text{–C–CH}_2\text{–CH}_3
\]

64. The reaction is
\[
\text{C}_6\text{H}_{12}\text{O}_6 \xrightarrow{\text{yeast}} 2\text{CH}_3\text{–CH}_2\text{–OH} + 2\text{CO}_2
\]
The yeast necessary for the fermentation process are killed if the concentration of ethanol is more than 13%. More concentrated ethanol solutions are most commonly made by distillation.

66. methanol (CH₃OH) - starting material for synthesis of acetic acid and many plastics
ethylene glycol (CH₂OH–CH₂OH) - automobile antifreeze
isopropyl alcohol (2-propanol, CH₃–CH(OH)–CH₃) - rubbing alcohol
68. Aldehydes and ketones both contain the carbonyl group C=O.

\[ \text{C} = \text{O} \]

Aldehydes and ketones differ in the location of the carbonyl function. Aldehydes contain the carbonyl group at the end of a hydrocarbon chain (The carbon atom of the carbonyl group is bonded at most to only one other carbon atom.). The carbonyl group of ketones represents one of the interior carbon atoms of a chain (The carbon atom of the carbonyl group is bonded to two other carbon atoms.).

70.

a. CH₃–CH₂–CH₂–COOH
b. CH₃–CH₂–C(=O)–CH₃

72. In addition to their systematic names (based on the hydrocarbon root with the ending –one), ketones also can be named by naming the groups attached to either side of the carbonyl carbon as alkyl groups followed by the word “ketone”. Examples are:

CH₃–C(=O)–CH₂CH₃  methyl ethyl ketone (2-butanone, butanone)
CH₃CH₂–C(=O)–CH₂CH₃  diethyl ketone (3-pentanone)

74.

a.

\[ \begin{array}{c} \text{O} \\ \text{H} \end{array} \]

b.

\[ \begin{array}{c} \text{O} \\ \text{CH₃} \end{array} \]

76. Carboxylic acids are typically weak acids.

\[ \text{CH₃–CH₂–COOH}_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{CH₃–CH₂–COO}^−_{(aq)} \]

78.

a. CH₃–CH₂–CH₂–CHO
80. Acetylsalicylic acid is synthesized from salicylic acid (behaving as an alcohol through its \(-\text{OH}\) group) and acetic acid.

\[
\text{CH}_3\text{C}_6\text{H}_4\text{OH} + \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{C}_6\text{H}_4\text{O} \text{C} - \text{CH}_3 + \text{H}_2\text{O}
\]

82. (a) 
\[
\text{CH}_3\text{C}_6\text{H}_4\text{OH} + \text{CH}_3\text{C}_6\text{H}_4\text{OH}
\]
(b) 
\[
\text{CH}_3\text{C}_6\text{H}_4\text{O} \text{C} - \text{CH}_3
\]
(c) 
\[
\text{NO}_2\text{C}_6\text{H}_4\text{COOH}
\]
(d) 
\[
\text{CH}_3\text{C}_6\text{H}_4\text{O} \text{C} - \text{CH}_3
\]

84. In addition polymerization, the monomer units simply add together to form the polymer with no other products. Polyethylene and polytetrafluoroethylene (Teflon) are common examples.

86. A polyester is formed from the reaction of a dialcohol (two \(-\text{OH}\) groups) with a diacid (two \(-\text{COOH}\) groups). One \(-\text{OH}\) group of the alcohol forms an **ester linkage** with one of the \(-\text{COOH}\) groups of the acid. As the resulting dimer still possesses an \(-\text{OH}\) and a \(-\text{COOH}\) group, the dimer can undergo further esterification reactions. Dacron is a common polyester.
90. unsaturated
92. straight-chain or normal
94. -ane
96. number
98. anti-knocking
100. substitution
102. hydrogenation
104. functional
106. carbon monoxide
108. carbonyl
110. carboxyl
112. addition
114. Structures depend on student choices.
116.
   a. 2-chlorobutane
   b. 1,2-dibromoethane
   c. triiodomethane (common name: iodoform)
   d. 2,3,4-trichloropentane
   e. 2,2-dichloro-4-isopropylheptane
118.
   a. 
      \[
      \text{CH}_3\text{CHCH}_2\text{CH}_2\text{CH}_2\text{CH}_3
      \]
      \[
      \text{CH}_3 \quad \text{CH}_3
      \]
   b. 
      \[
      \text{HO-CH}_2\text{CHCH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3
      \]
      \[
      \text{CH}_3 \quad \text{Cl}
      \]
   c. 
      \[
      \text{CH}_2\text{CCH}_2\text{CH}_2\text{CH}_2\text{CH}_3
      \]
      \[
      \text{Cl}
      \]
   d. 
      \[
      \text{Cl-CH}_2\text{CCHCH}_2\text{CH}_2\text{CH}_3
      \]
120. primary
CH₃─CH₂─CH₂─CH₂─CH₂─CH₂─OH

secondary
CH₃─CH₂─CH₂─CH₂─CH─CH₃

OH

tertiary
CH₃
CH₃─CH₂─C─CH₂─CH₃

OH

122.
HO─CH₂─CH─CH─CH─CH─CHO

| OH | OH | OH | OH |

124.

a. CH₃─C─CH₂─CH₂─CH₂─CH₂─CH₃

O

b. CH₃─CH₂─CH─CH₂─CHO

CH₃

c. CH₃─CH₂─CH₂─CH─CH₂─OH

CH₃

d. OH
CH₂─CH─CH₂

HO | OH
Chapter 20: Organic Chemistry

126.

\[
\text{HO-C=O} \quad \text{CH}_3-\text{CH-N-H} + \text{HO-C=O} \quad \rightarrow \quad \text{CH}_3-\text{CH-N-C=O} + \text{H}_2\text{O}
\]

128.

a. \(\text{CH}_3-\text{CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3\)

b. \(\text{CH}_3\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3\)

c. \(\text{CH}_3\text{-CH}_2\text{-CH}_2\text{-CH}_3\)

d. \(\text{CH}_3\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3\)

130.

a. 2,3-dimethylbutane

b. 3,3-diethylpentane

c. 2,3,3-trimethylhexane

d. 2,3,4,5,6-pentamethylheptane

132.

a. \(\text{CH}_3\text{Cl}(g)\)

b. \(\text{H}_2(g)\)

c. \(\text{HCl}(g)\)

134. \(\text{CH}=\text{C-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3\) 1-octyne

\(\text{CH}_3\text{-C}=\text{C-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3\) 2-octyne

\(\text{CH}_3\text{-CH}_2\text{-C}=\text{C-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3\) 3-octyne

\(\text{CH}_3\text{-CH}_2\text{-CH}_2\text{-C}=\text{C-CH}_2\text{-CH}_2\text{-CH}_3\) 4-octyne

136.

a. carboxylic acid

b. ketone
c. ester
d. alcohol (phenol)

138.
a. 3-methylpentanal
\[ \text{CH}_3-\text{CH}_2-\text{CH}-\text{CH}_2-\text{CHO} \]
\[ \text{CH}_3 \]
b. 3-methyl-2-pentanone
\[ \text{CH}_3-\text{CH}_2-\text{CH}-\text{C}-\text{CH}_3 \]
\[ \text{CH}_3 \]
c. methyl phenyl ketone
\[ \text{CH}--\text{CH}_3 \]
\[ \text{O} \]
d. 2-hydroxybutanal
\[ \text{CH}_3-\text{CH}_2-\text{CH}-\text{CHO} \]
\[ \text{OH} \]
e. propanal
\[ \text{CH}_3-\text{CH}_2-\text{CHO} \]

140.
a. 
\[ \text{CH}_3-\text{CH}-\text{CH}_2-\text{COOH} \]
\[ \text{CH}_3 \]
b. 
\[ \text{CH}--\text{OH} \]
c. 
\[ \text{CH}_3-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{COOH} \]
d. 
\[ \text{CH}_3-\text{COOH} \]
CHAPTER 21

Biochemistry

CHAPTER ANSWERS

2. The essential elements are those required for the structure and processes of the human body. The essential elements and their functions are listed in Table 21.1. The top five essential elements (percent by mass) are oxygen (present in water and many organic compounds in the body), carbon (present in all organic compounds), hydrogen (present in water and in many organic and inorganic compounds), nitrogen (present in both inorganic and organic compounds), and calcium (constituent of bones/teeth and essential to some enzymes).

4. Fibrous, globular

6. If the side chain in the amino acid is polar, this will tend to make the amino acid hydrophilic. If the side chain is mostly nonpolar, this will tend to make the amino acid hydrophobic. The arrangement of hydrophilic and hydrophobic side chains will therefore affect the overall shape of the protein in an aqueous medium.

8. The amino acid will be hydrophilic if the R group is polar and hydrophobic if the R group is nonpolar. Serine is a good example of an amino acid in which the R group is polar. Leucine is a good example of an amino acid with a nonpolar R group.

10. There are six tripeptides possible.

\[
\begin{align*}
\text{cys-ala-phe} & \quad \text{ala-cys-phe} & \quad \text{phe-ala-cys} \\
\text{cys-phe-ala} & \quad \text{ala-phe-cys} & \quad \text{phe-cys-ala}
\end{align*}
\]

12. The “peptide linkage” is an amide linkage:

\[
\begin{array}{c}
\text{H} \\
\text{C} \quad \text{N} \\
\text{O}
\end{array}
\]

14. Long, thin, resilient proteins, such as hair, typically contain elongated, elastic \( \alpha \)-helical protein molecules. Other proteins, such as silk, which in bulk form sheets or plates, typically contain protein molecules having the \( \beta \)-pleated sheet secondary structure. Proteins that do not have a structural function in the body, such as hemoglobin, typically have a globular structure.

16. Silk consists of a sheet structure in which the individual chains of amino acids are lined up lengthwise next to each other to form the sheet.

18. The strands of protein in hair are cross-linked by means of disulfide linkages between cysteine amino acid units in adjacent strands. In a permanent wave, the hair is first treated to break these linkages. Then, after the hair is rolled onto curlers, the linkages are allowed to reform. As the hair is on the curlers, the disulfide linkages form in a different place along the protein strands, which imparts a curl to the hair when the curler is removed.

20. Oxygen is transported by the protein hemoglobin.
22. Antibodies are special proteins that are synthesized in the body in response to foreign substances such as bacteria or viruses.

24. Amino acids contain both a weak acid and a weak base group, and thus they can neutralize both bases and acids, respectively.

26. The molecule acted upon by an enzyme is referred to as the enzyme’s substrate. When we say that an enzyme is specific for a particular substrate, we mean that the enzyme will catalyze the reactions of that molecule and that molecule only.

28. The lock-and-key model for enzymes indicates that the structures of an enzyme and its substrate must be complementary so that the substrate can approach and attach itself along the length of the enzyme at the enzyme’s active sites. A given enzyme is intended to act upon a particular substrate. The substrate attaches itself to the enzyme, is acted upon, and then moves away from the enzyme. If a different molecule has a similar structure to the substrate, this other molecule may also be capable of attaching itself to the enzyme. But since this molecule is not the enzyme’s proper substrate, the enzyme may not be able to act upon the molecule, and the molecule may remain attached to the enzyme, preventing proper substrate molecules from approaching the enzyme (irreversible inhibition). If the enzyme cannot act upon its proper substrate, then the enzyme is said to be inhibited. Irreversible inhibition might be a desirable feature in an antibiotic, which would bind to the enzymes of a bacteria and prevent the bacteria from reproducing, thereby preventing or curing an infection.

30. Sugars contain an aldehyde or ketone functional group (carbonyl group) as well as several –OH groups (hydroxyl group).

32.

a. glucose

```
  CHO
 /    |
H--C--OH
 /  |
HO--C--H
 |  |
H--C--OH
 |  |
H--C--OH
   |
CH₂OH
```

b. ribose

```
  CHO
 /    |
H--C--OH
 /  |
H--C--OH
 /  |
H--C--OH
   |
CH₂OH
```
34. Although starch and cellulose are both polymers of glucose, the glucose rings in the two substances are linked together differently, which results in different properties. In particular, humans do not have an enzyme capable of catalyzing the breakdown of the cellulose chain into glucose units whereas we do have an enzyme capable of breaking down starch.

36. ribose (aldopentose); arabinose (aldopentose); ribulose (ketopentose); glucose (aldohexose); mannose (aldohexose); galactose (aldohexose); fructose (ketohexose).

38. A nucleotide consists of a five-carbon sugar (ribose in RNA, deoxyribose in DNA) to which a phosphate group and one of five characteristic organic bases is bonded.

40. Uracil (RNA only); cytosine (DNA, RNA); thymine (DNA only); adenine (DNA, RNA); guanine (DNA, RNA)

42. An overall DNA molecule consists of two chains of nucleotides, with the organic bases on the nucleotides arranged in complementary pairs (cytosine with guanine and adenine with thymine). The structures and properties of the organic bases are such that these pairs fit together well and allow the two chains of nucleotides to form the double helix structure. When DNA replicates, it is assumed the double helix unwinds, and then new molecules of the organic bases come in and pair up with their respective partner on the separated nucleotide chains, thereby replicating the original structure. (See Figure 21.20.)

44. Messenger RNA molecules are synthesized to be complementary to a portion (gene) of the DNA molecule in the cell, and serve as the template or pattern upon which a protein will be constructed (a particular group of nitrogen bases on m-RNA is able to accommodate and specify a particular amino acid in a particular location in the protein). Transfer RNA molecules are much smaller than m-RNA, and their structure accommodates only a single specific amino acid molecule. Transfer RNA molecules “find” their specific amino acid in the cellular fluids and bring it to m-RNA where it is added to the protein molecule being synthesized.
46. A triglyceride typically consists of a glycerol backbone to which three separate fatty acid molecules are attached by ester linkages.

\[ R - C - O - CH_2 - O - \quad R' \]

\[ O - CH - O - C - \quad R'' - C - O - CH_2 \]

48. “Soaps” are the salts of long-chain organic acids (“fatty acids”), most commonly either the sodium or potassium salt. Soaps are prepared by treating a fat or oil (a triglyceride) with a strong base such as NaOH or KOH. This breaks the ester linkages in the triglyceride, releasing three fatty acid anions and glycerol.

\[ R - C - O - CH_2 - O - CH_2OH \quad RCOONa \]

\[ O - CH - O - C - R' \quad CHOH \quad RCOONa \]

\[ R'' - C - O - CH_2 \quad CH_2OH \quad R''COONa \]

50. Soaps have both a nonpolar nature (due to the long chain of the fatty acid) and an ionic nature (due to the charge on the carboxyl group). In water, soap anions form aggregates called micelles in which the water-repelling hydrocarbon chains are oriented towards the interior of the aggregate, and the ionic, water-attracting carboxyl groups are oriented towards the outside. Most dirt has a greasy nature. A soap micelle is able to interact with a grease molecule, pulling the grease molecule into the hydrocarbon interior of the micelle. When the clothing is rinsed, the micelle containing the grease is washed away. (See Figures 21.22 and 21.23.)

52. Cholesterol is the naturally occurring steroid from which the body synthesizes other needed steroids. As cholesterol is insoluble in water, it is thought that having too large a concentration of this substance in the bloodstream may lead to its deposition and build-up on the walls of blood vessels, causing their eventual blockage.

54. The bile acids are synthesized from cholesterol in the liver and are stored in the gall bladder. Bile acids such as cholic acid act as emulsifying agents for lipids and aid in their digestion.

56. i

58. m

60. u

62. f

64. g

66. r

68. p

70. o

72. b

74. d

76. a
78. nucleotides
80. ester
82. thymine, guanine
84. transfer, messenger
86. lipids
88. unsaturated, saturated
90. ionic, nonpolar
92. fatty (long chain)
94. progesterone
96. The primary structure of a protein refers to the specific identity and ordering of amino acids in a protein’s polypeptide chain. The primary structure is sometimes referred to as the protein’s amino acid sequence.
98. tendons, bone (with mineral constituents), skin, cartilage, hair, fingernails.
100. Proteins contain both acidic (–COOH) and basic (–NH₂) groups in their side chains that can neutralize both acids and bases.
102. pentoses (five carbons); hexoses (six carbons); trioses (three carbons)
104. In a strand of DNA, the phosphate group and the sugar molecule of adjacent nucleotides become bound to each other. The chain portion of the DNA molecule, therefore, consists of alternating phosphate groups and sugar molecules. The nitrogen bases are found protruding from the side of this phosphate-sugar chain, bonded to the sugar molecules.
106. Phospholipids are esters of glycerol. Two fatty acids are bonded to the –OH groups of the glycerol backbone with the third –OH group of glycerol bonded to a phosphate group. Having the two fatty acids but also the polar phosphate group makes the phospholipid lecithin a good emulsifying agent.