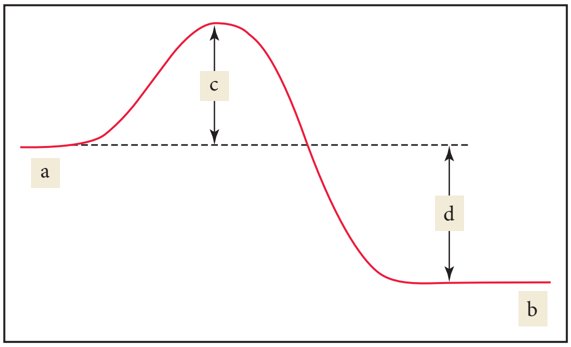
57.



59. 17 s^{-1}

61. 61.90 kJ/mol

63. $E_a = 251$ kJ/mol, $A = 7.93 \times 10^{11}\ s^{-1}$

65. $E_a = 23.0$ kJ/mol, $A = 8.05 \times 10^{10}\ s^{-1}$

67. **a.** 122 kJ/mol
  b. 0.101 s^{-1}

69. 47.85 kJ/mol

71. **a**

73. The mechanism is valid.

75. **a.** $Cl_2(g) + CHCl_3(g) \longrightarrow HCl(g) + CCl_4(g)$
  b. $Cl(g)$, $CCl_3(g)$
  c. Rate $= k[Cl_2]^{1/2}[CHCl_3]$

77. Heterogeneous catalysts require a large surface area because catalysis can only happen at the surface. A greater surface area means greater opportunity for the substrate to react, which results in a faster reaction.

79. 10^{12}

81. **a.** first order, $k = 0.0462\ hr^{-1}$
  b. 15 hr
  c. 5.0×10^1 hr

83. 0.0531 M/s

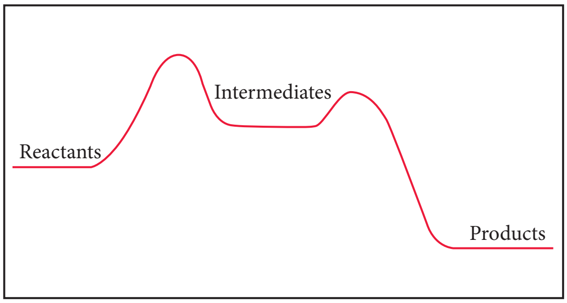
85. rate $= 4.5 \times 10^{-4}\ [CH_3CHO]^2$, $k = 4.5 \times 10^{-4}$, 0.37 atm

87. 219 torr

89. 1×10^{-7} s

91. 1.6×10^2 seconds

93. **a.** 2
  b.



  c. first step **d.** exothermic

95. **a.** 5.41 s
  b. 2.2 s for 25%, 5.4 s for 50%
  c. 0.28 at 10 s, 0.077 at 20 s

97. **a.** $E_a = 89.5$ kJ/mol, $A = 4.22 \times 10^{11}\ s^{-1}$
  b. $2.5 \times 10^{-5}\ M^{-1}\ s^{-1}$
  c. 6.0×10^{-4} M/s

99. **a.** No
  b. No bond is broken, and the two radicals attract each other.
  c. Formation of diatomic gases from atomic gases.

101. 1.35×10^4 years

103. **a.** Both are valid. For both, all steps sum to overall reaction and the predicted rate law is consistent with experimental data.
  b. Buildup of $I(g)$

105. rate $= k_2[(k_1/k_{-1})[Br_2]]^{1/2}[H_2]$
The rate law is 3/2 order overall.

107. **a.** 0%
  b. 25%
  c. 33%

109. 174 kJ

111. **a.** second order
  b. $CH_3NC + CH_3NC \underset{k_2}{\overset{k_1}{\rightleftharpoons}} CH_3NC^* + CH_3NC$ (fast)

$CH_3NC^* \xrightarrow{k_3} CH_3CN$ (slow)

Rate $= k_3[CH_3NC^*]$

$k_1[CH_3NC]^2 = k_2[CH_3NC^*][CH_3NC]$

$$[CH_3NC^*] = \frac{k_1}{k_2}[CH_3NC]$$

$$\text{Rate} = k_3 \times \frac{k_1}{k_2}[CH_3NC]$$

Rate $= k[CH_3NC]$

113. Rate $= k[A]^2$

$$Rate = -\frac{d[A]}{dt}$$

$$\frac{d[A]}{dt} = -k[A]^2$$

$$-\frac{d[A]}{[A]^2} = k\,dt$$

$$\int_{[A]_0}^{[A]} -\frac{1}{[A]^2}d[A] = \int_0^t k\,dt$$

$$\left[\frac{1}{[A]}\right]_{[A]_0}^{[A]} = k\,[t]_0^t$$

$$\frac{1}{[A]} - \frac{1}{[A]_0} = kt$$

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$$

115. Rate $= k\,[CO][Cl_2]^{\frac{3}{2}}$

117. $[Cl_2] = 0.0084$ mol/L, $[NO] = 0.017$ mol/L

119. B is first order and A is second order.
B will be linear if you plot ln[B] versus time; A will be linear if you plot 1/[A] versus time.

Chapter 14

21. **a.** $K = \dfrac{[SbCl_3][Cl_2]}{[SbCl_5]}$ **b.** $K = \dfrac{[NO]^2[Br_2]}{[BrNO]^2}$
  c. $K = \dfrac{[CS_2][H_2]^4}{[CH_4][H_2S]^2}$ **d.** $K = \dfrac{[CO_2]^2}{[CO]^2[O_2]}$

23. The concentration of the reactants will be greater. No, this is not dependent on initial concentrations; it is dependent on the value of K_c.

25. **a.** figure v
b. The change in the decrease of reactants and increase of products would be faster.
c. No, catalysts affect kinetics, not equilibrium.

27. **a.** 4.42×10^{-5}, reactants favored
b. 1.50×10^{2}, products favored
c. 1.96×10^{-9}, reactants favored

29. 1.3×10^{-29}

31. **a.** 2.56×10^{-23}
b. 1.3×10^{22}
c. 81.9

33. **a.** $K_c = \dfrac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$
b. $K_c = [O_2]^3$
c. $K_c = \dfrac{[H_3O^+][F^-]}{[HF]}$
d. $K_c = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$

35. 136

37.

T (K)	$[N_2]$	$[H_2]$	$[NH_3]$	K_c
500	0.115	0.105	0.439	1.45×10^{-3}
575	0.110	0.249	0.128	9.6
775	0.120	0.140	4.39×10^{-3}	0.0584

39. 234 torr

41. 18

43. 3.3×10^{2}

45. 764

47. More solid will form.

49. Additional solid will not dissolve.

51. **a.** [A] = 0.20 M, [B] = 0.80 M
b. [A] = 0.33 M, [B] = 0.67 M
c. [A] = 0.38 M, [B] = 1.2 M

53. $[N_2O_4] = 0.0115$ M, $[NO_2] = 0.0770$ M

55. $0.1\underline{9}9$ M

57. 1.9×10^{-3} M

59. 7.84 torr

61. **a.** [A] = 0.38 M, [B] = 0.62 M, [C] = 0.62 M
b. [A] = 0.90 M, [B] = 0.095 M, [C] = 0.095 M
c. [A] = 1.0 M, [B] = 3.2×10^{-3} M, [C] = 3.2×10^{-3} M

63. **a.** shift left **b.** shift right
c. shift right

65. **a.** shift right **b.** no effect
c. no effect **d.** shift left

67. **a.** shift right **b.** shift left
c. no effect

69. Increase temperature $\longrightarrow$ shift right, decrease temperature $\longrightarrow$ shift left. Increasing the temperature will increase the equilibrium constant.

71. **b, d**

73. **a.** 1.7×10^{2}
b. $\dfrac{[Hb{-}CO]}{[Hb{-}O_2]} = 0.85$ or $17/20$
CO is highly toxic, as it blocks O_2 uptake by hemoglobin. CO at a level of 0.1% will replace nearly half of the O_2 in blood.

75. **a.** 1.68 atm
b. 1.41 atm

77. 0.406 g

79. **b, c, d**

81. 0.0144 atm

83. 3.1×10^{2} g, 20% yield

85. 0.12 atm

87. 0.72 atm

89. 0.017 g

91. 0.226

93. **a.** 29.3
b. 86.3 torr

95. $P_{NO} = P_{Cl_2} = 429$ torr

97. 1.27×10^{-2}

99. $K_P = 5.1 \times 10^{-2}$

101. Yes, because the volume affects Q.

103. a = 1, b = 2

Chapter 15

33. **a.** acid, $HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$
b. acid, $NH_4^+(aq) \rightleftharpoons H^+(aq) + NH_3(aq)$
c. base, $KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$
d. acid, $HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$

35. **a.** $\underset{\text{acid}}{H_2CO_3(aq)} + \underset{\text{base}}{H_2O(l)} \rightleftharpoons \underset{\text{conj. acid}}{H_3O^+(aq)} + \underset{\text{conj. base}}{HCO_3^-(aq)}$
b. $\underset{\text{acid}}{NH_3(aq)} + \underset{\text{base}}{H_2O(l)} \rightleftharpoons \underset{\text{conj. acid}}{NH_4^+(aq)} + \underset{\text{conj. base}}{OH^-(aq)}$
c. $\underset{\text{acid}}{HNO_3(aq)} + \underset{\text{base}}{H_2O(l)} \rightleftharpoons \underset{\text{conj. acid}}{H_3O^+(aq)} + \underset{\text{conj. base}}{NO_3^-(aq)}$
d. $\underset{\text{acid}}{C_5H_5N(aq)} + \underset{\text{base}}{H_2O(l)} \rightleftharpoons \underset{\text{conj. acid}}{C_5H_5NH^+(aq)} + \underset{\text{conj. base}}{OH^-(aq)}$

37. **a.** Cl^- **b.** HSO_3^- **c.** CHO_2^- **d.** F^-

39. $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$
$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq)$

41. **a.** strong **b.** strong
c. strong **d.** weak, $K_a = \dfrac{[H_3O^+][HSO_3^-]}{[H_2SO_3]}$

43. **a, b, c**

45. **a.** F^- **b.** NO_2^- **c.** ClO^-

47. **a.** 8.3×10^{-7}, basic
b. 1.2×10^{-10}, acidic
c. 2.9×10^{-13}, acidic

49. **a.** pH = 7.77, pOH = 6.23
b. pH = 7.00, pOH = 7.00
c. pH = 5.66, pOH = 8.34

51.

$[H_3O^+]$	$[OH^-]$	pH	Acidic or Basic
7.1×10^{-4}	1.4×10^{-11}	3.15	Acidic
3.7×10^{-9}	2.7×10^{-6}	8.43	Basic
7.9×10^{-12}	1.3×10^{-3}	11.1	Basic
6.3×10^{-4}	1.6×10^{-11}	3.20	Acidic

53. $[H_3O^+] = 1.5 \times 10^{-7}$ M, pH = 6.81

55. pH = 1.36, 1.35, 1.34. A difference of 1 in the second significant digit in a concentration value produces a difference of 0.01 in pH. Therefore, the second significant digit in value of the concentration corresponds to the hundredths place in a pH value.

57. **a.** $[H_3O^+] = 0.25$ M, $[OH^-] = 4.0 \times 10^{-14}$ M, pH = 0.60
b. $[H_3O^+] = 0.015$ M, $[OH^-] = 6.7 \times 10^{-13}$ M, pH = 1.82
c. $[H_3O^+] = 0.072$ M, $[OH^-] = 1.4 \times 10^{-13}$ M, pH = 1.14
d. $[H_3O^+] = 0.105$ M, $[OH^-] = 9.5 \times 10^{-14}$ M, pH = 0.979

59. **a.** 1.8 g **b.** 0.57 g **c.** 0.045 g

61. 2.21

63. $[H_3O^+] = 2.5 \times 10^{-3}$ M, pH = 2.59

65. **a.** 1.82 (approximation valid)
b. 2.18 (approximation breaks down)
c. 2.72 (approximation breaks down)

67. 2.75

69. 6.8×10^{-6}

71. 0.0063%

73. **a.** 0.42% **b.** 0.60% **c.** 1.3% **d.** 1.9%

75. 3.61×10^{-5}

77. **a.** pH = 2.03, percent ionization = 3.7%
b. pH = 2.24, percent ionization = 5.7%
c. pH = 2.40, percent ionization = 8.0%

79. **a.** 0.939 **b.** 1.07 **c.** 2.19 **d.** 3.02

81. **a.** $[OH^-] = 0.15$ M, $[H_3O^+] = 6.7 \times 10^{-14}$ M, pH = 13.17, pOH = 0.83
b. $[OH^+] = 0.003$ M, $[H_3O^+] = 3.3 \times 10^{-12}$ M, pH = 11.48, pOH = 2.52
c. $[OH^-] = 9.6 \times 10^{-4}$ M, $[H_3O^+] = 1.0 \times 10^{-11}$ M, pH = 10.98, pOH = 3.02
d. $[OH^-] = 8.7 \times 10^{-5}$ M, $[H_3O^+] = 1.1 \times 10^{-10}$ M, pH = 9.93, pOH = 4.07

83. 13.842

85. 0.104 L

87. **a.** $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$,

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

b. $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$, $K_b = \dfrac{[H_2CO_3][OH^-]}{[HCO_3^-]}$

c. $CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$, $K_b = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$

89. $[OH^-] = 1.6 \times 10^{-3}$ M, pOH = 2.79, pH = 11.21

91. 7.48

93. 6.7×10^{-7}

95. **a.** neutral
b. basic,
$ClO^-(aq) + H_2O(l) \rightleftharpoons HClO(aq) + OH^-(aq)$
c. basic,
$CN^-(aq) + H_2O(l) \rightleftharpoons HCN(aq) + OH^-(aq)$
d. neutral

97. $[OH^-] = 2.0 \times 10^{-6}$ M, pH = 8.30

99. **a.** acidic,
$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$
b. neutral
c. acidic, $Co(H_2O)_6^{3+}(aq) + H_2O(l) \rightleftharpoons Co(H_2O)_5(OH)^{2+}(aq) + H_3O^+(aq)$
d. acidic, $CH_2NH_3^+(aq) + H_2O(l) \rightleftharpoons CH_2NH_2(aq) + H_3O^+(aq)$

101. **a.** acidic **b.** basic **c.** neutral
d. acidic **e.** acidic

103. NaOH, $NaHCO_3$, NaCl, NH_4ClO_2, NH_4Cl

105. **a.** 5.13 **b.** 8.87 **c.** 7.0

107. $[K^+] = 0.15$ M, $[F^-] = 0.15$ M, $[HF] = 2.1 \times 10^{-6}$ M, $[OH^-] = 2.1 \times 10^{-6}$ M; $[H_3O^+] = 4.8 \times 10^{-9}$ M

109. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$,

$$K_{a_i} = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]}$$

$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$,

$$K_{a_2} = \frac{[H_3O^+][HPO_4^{2-}]}{[H_2PO_4^-]}$$

$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$,

$$K_{a_3} = \frac{[H_3O^+][PO_4^{3-}]}{[HPO_4^{2-}]}$$

111. **a.** $[H_3O^+] = 0.048$ M, pH = 1.32
b. $[H_3O^+] = 0.12$ M, pH = 0.92

113. $[H_2SO_3] = 0.418$ M
$[HSO_3^-] = 0.082$ M
$[SO_3^{2-}] = 6.4 \times 10^{-8}$ M
$[H_3O^+] = 0.082$ M

115. **a.** $[H_3O^+] = 0.50$ M, pH = 0.30
b. $[H_3O^+] = 0.11$ M, pH = 0.96 (*x is small* approximation breaks down)
c. $[H_3O^+] = 0.059$ M, pH = 1.23

117. **a.** HCl, weaker bond
b. HF, bond polarity
c. H_2Se, weaker bond

119. **a.** H_2SO_4, more oxygen atoms bonded to S
b. $HClO_2$, more oxygen atoms bonded to Cl
c. HClO, Cl has higher electronegativity
d. CCl_3COOH, Cl has higher electronegativity

121. S^{2-}, its conjugate acid (H_2S), is a weaker acid than H_2S

123. **a.** Lewis acid **b.** Lewis acid
c. Lewis base **d.** Lewis base

125. **a.** acid: Fe^{3+}, base: H_2O
b. acid: Zn^{2+}, base: NH_3
c. acid: BF_3, base: $(CH_3)_3N$

127. **a.** weak **b.** strong
c. weak **d.** strong

129. If blood became acidic, the H^+ concentration would increase. According to Le Châtelier's principle, equilibrium would be shifted to the left and the concentration of oxygenated Hb would decrease.

131. All acid will be neutralized.

133. $[H_3O^+]$(Great Lakes) = 3×10^{-5} M, $[H_3O^+]$(West Coast) = 4×10^{-6} M. The rain over the Great Lakes is about eight times more concentrated.

135. 2.7

137. **a.** 2.000 **b.** 1.52 **c.** 12.95
d. 11.12 **e.** 5.03

139. **a.** 1.260 **b.** 8.22 **c.** 0.824
d. 8.57 **e.** 1.171

141. **a.** $CN^-(aq) + H^+(aq) \rightleftharpoons HCN(aq)$
b. $NH_4^+(aq) + OH^-(aq) \rightleftharpoons NH_3(aq) + H_2O(l)$
c. $CN^-(aq) + NH_4^+(aq) \rightleftharpoons HCN(aq) + NH_3(aq)$
d. $HSO_4^-(aq) + C_2H_3O_2^-(aq) \rightleftharpoons SO_4^{2-}(aq) + HC_2H_3O_2(aq)$
e. no reaction between the major species

143. 0.794

145. $K_a = 8.3 \times 10^{-4}$

147. 6.79

149. 2.14

151. $[A^-] = 4.5 \times 10^{-5}$ M
$[H^+] = 2.2 \times 10^{-4}$ M
$[HA_2^-] = 1.8 \times 10^{-4}$ M

153. 9.28

155. 50.1 g $NaHCO_3$

157. **b**

159. $CH_3COOH < CH_2ClCOOH < CHCl_2COOH < CCl_3COOH$

Chapter 16

27. **d**

29. **a.** 3.62
b. 9.11

31. pure water: 2.1%, in $NaC_7H_5O_2$: 0.065%. The percent ionization in the sodium benzoate solution is much smaller because the presence of the benzoate ion shifts the equilibrium to the left.

33. **a.** 2.14 **b.** 8.32 **c.** 3.46

35. $HCl + NaC_2H_3O_2 \longrightarrow HC_2H_3O_2 + NaCl$
$NaOH + HC_2H_3O_2 \longrightarrow NaC_2H_3O_2 + H_2O$

37. **a.** 3.62 **b.** 9.11

39. **a.** 7.60 **b.** 11.18 **c.** 4.61

41. **a.** 3.86 **b.** 8.95

43. 3.5

45. 3.7 g

47. **a.** 4.74 **b.** 4.68 **c.** 4.81

49. **a.** initial 7.00 after 1.70
b. initial 4.71 after 4.56
c. initial 10.78 after 10.66

51. 1.2 g; 2.7 g

53. **a.** yes **b.** no
c. yes **d.** no
e. no

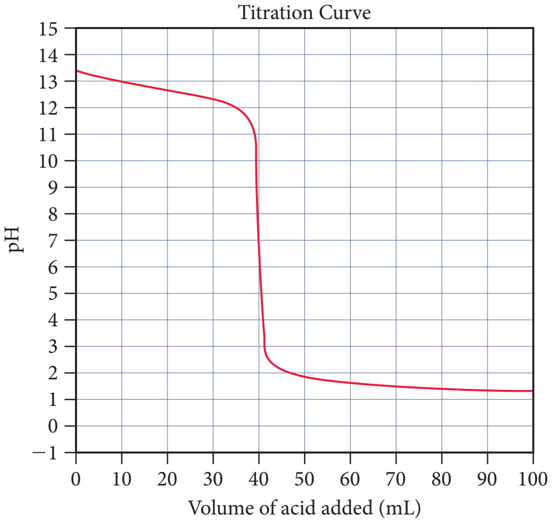
55. **a.** 7.4 **b.** 0.3 g **c.** 0.14 g

57. KClO/HClO = 0.79

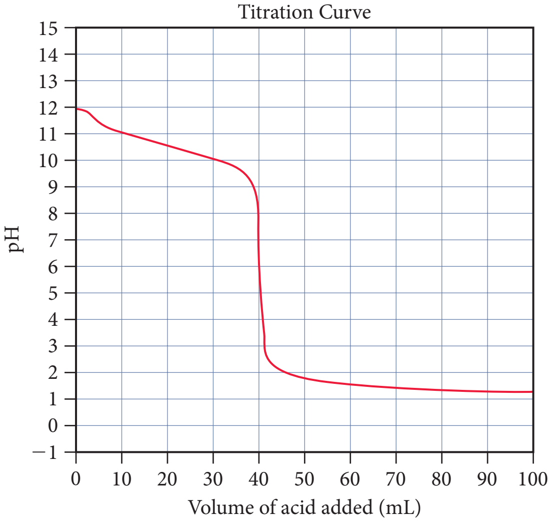
59. **a.** does not exceed capacity
b. does not exceed capacity
c. does not exceed capacity
d. does not exceed capacity

61. **i.** (a) pH = 8, (b) pH = 7
ii. (a) weak acid, (b) strong acid

63. **a.** 40.0 mL HI for both
b. KOH: neutral, CH_3NH_2: acidic
c. CH_3NH_2
d. Titration of KOH with HI:



Titration of CH_3NH_2 with HI:



65. **a.** pH = 9, added base = 30 mL
b. 0 mL **c.** 15 mL
d. 30 mL **e** 30 mL

67. **a.** 0.757 **b.** 30.6 mL **c.** 1.038
d. 7 **e** 12.15

69. **a.** 13.06 **b.** 28.8 mL **c.** 12.90
d. 7 **e.** 2.07

71. **a.** 2.86 **b.** 16.8 mL **c.** 4.37
d. 4.74 **e.** 8.75 **f.** 12.17

73. **a.** 11.94 **b.** 29.2 mL **c.** 11.33
d. 10.64 **e.** 5.87 **f.** 1.90

75. **i.** (a) **ii.** (b)

77. $pK_a = 3$, 82 g/mol

79. First equivalence: 22.7 mL
Second equivalence: 45.4 mL

81. The indicator will appear red. The pH range is 4 to 6.

83. **a.** phenol red, *m*-nitrophenol
b. alizarin, bromothymol blue, phenol red
c. alizarin yellow R

85. **a.** $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$, $K_{sp} = [Ba^{2+}][SO_4^{2-}]$
b. $PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\ Br^-(aq)$, $K_{sp} = [Pb^{2+}][Br^-]^2$
c. $Ag_2CrO_4(s) \rightleftharpoons 2\ Ag^+(aq) + CrO_4^{2-}(aq)$, $K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

87. **a.** 7.31×10^{-7} M **b.** 3.72×10^{-5} M
c. 3.32×10^{-4} M

89. **a.** 1.07×10^{-21} **b.** 7.14×10^{-7}
c. 7.44×10^{-11}

91. AX_2

93. 2.07×10^{-5} g/100 mL

95. **a.** 0.0183 M **b.** 0.00755 M
c. 0.00109 M

97. **a.** 5×10^{14} M **b.** 5×10^{8} M
c. 5×10^{4} M

99. **a.** more soluble, CO_3^{2-} is basic
b. more soluble, S^{2-} is basic
c. not, neutral
d. not, neutral

101. precipitate will form, CaF_2

103. precipitate will form, $Mg(OH)_2$

105. **a.** 0.018 M **b.** 1.4×10^{-7} M
c. 1.1×10^{-5} M

107. **a.** $BaSO_4$, 1.1×10^{-8} M
b. 3.0×10^{-8} M

109. 8.7×10^{-10} M

111. 5.6×10^{16}

113. 4.03

115. 3.57

117. HCl, 4.7 g

119. **a.** $NaOH(aq) + KHC_8H_4O_4(aq) \longrightarrow Na^+(aq) + K^+(aq) + C_8H_4O_4^{2-}(aq) + H_2O(l)$
b. 0.1046 M

121. 4.73

123. 176 g/mol; 1.0×10^{-4}

125. 14.2 L

127. 1.6×10^{-7} M

129. 8.0×10^{-8} M

131. 6.29

133. 0.172 M

135. The ratio by mass of dimethyl ammonium chloride to dimethyl amine needed is 3.6.

137. 0.18 M benzoic acid, 0.41 M sodium benzoate

139. 51.6 g

141. 1.8×10^{-11} (based on this data)

143. **a.** 5.5×10^{-25} M **b.** 5.5×10^{-4} M

145. 1.38 L

147. 12.97

149. **a.** $pH < pK_a$ **b.** $pH > pK_a$
c. $pH = pK_a$ **d.** $pH > pK_a$

151. **b**

153. **a.** no difference **b.** less soluble
c. more soluble

Chapter 17

27. **a, c**

29. System B has the greatest entropy. There is only one energetically equivalent arrangement for System A. However, the particles of System B may exchange positions for a second energetically equivalent arrangement.

31. **a.** $\Delta S > 0$ **b.** $\Delta S < 0$
c. $\Delta S < 0$ **d.** $\Delta S < 0$

33. **a.** $\Delta S_{sys} > 0$, $\Delta S_{surr} > 0$, spontaneous at all temperatures
b. $\Delta S_{sys} < 0$, $\Delta S_{surr} < 0$, nonspontaneous at all temperatures
c. $\Delta S_{sys} < 0$, $\Delta S_{surr} < 0$, nonspontaneous at all temperatures
d. $\Delta S_{sys} > 0$, $\Delta S_{surr} > 0$, spontaneous at all temperatures

35. **a.** 1.29×10^3 J/K **b.** 5.00×10^3 J/K
c. -3.83×10^2 J/K **d.** -1.48×10^3 J/K

37. **a.** −649 J/K, nonspontaneous
b. 649 J/K, spontaneous
c. 123 J/K, spontaneous
d. −76 J/K, nonspontaneous

39. **a.** 1.93×10^5 J, nonspontaneous
b. -1.93×10^5 J, spontaneous
c. -3.7×10^4 J, spontaneous
d. 4.7×10^4 J, nonspontaneous

41. -2.247×10^6 J, spontaneous

43.

ΔH	ΔS	ΔG	Low Temperature	High Temperature
−	+	−	Spontaneous	Spontaneous
−	−	Temperature dependent	Spontaneous	Nonspontaneous
+	+	Temperature dependent	Nonspontaneous	Spontaneous
+	−	+	Nonspontaneous	Nonspontaneous

45. It increases.

47. **a.** $CO_2(g)$, greater molar mass and complexity
b. $CH_3OH(g)$, gas phase
c. $CO_2(g)$, greater molar mass and complexity
d. $SiH_4(g)$, greater molar mass
e. $CH_3CH_2CH_3(g)$, greater molar mass and complexity
f. $NaBr(aq)$, aqueous

49. **a.** He, Ne, SO_2, NH_3, CH_3CH_2OH. From He to Ne there is an increase in molar mass; beyond that, the molecules increase in complexity.
b. $H_2O(s)$, $H_2O(l)$, $H_2O(g)$; increase in entropy in going from solid to liquid to gas phase.
c. CH_4, CF_4, CCl_4; increasing entropy with increasing molar mass.

51. **a.** −120.8 J/K, decrease in moles of gas
b. 133.9 J/K, increase in moles of gas
c. −42.0 J/K, small change because moles of gas stay constant
d. −390.8 J/K, decrease in moles of gas

53. −89.3 J/K, decrease in moles of gas

55. $\Delta H^\circ_{\text{rxn}} = -1277$ kJ, $\Delta S^\circ_{\text{rxn}} = 313.6$ J/K, $\Delta G^\circ_{\text{rxn}} = -1.370 \times 10^3$ kJ; yes

57. **a.** $\Delta H^\circ_{\text{rxn}} = 57.2$ kJ, $\Delta S^\circ_{\text{rxn}} = 175.8$ J/K, $\Delta G^\circ_{\text{rxn}} = 4.8 \times 10^3$ J/mol; nonspontaneous, becomes spontaneous at high temperatures

b. $\Delta H^\circ_{\text{rxn}} = 176.2$ kJ, $\Delta S^\circ_{\text{rxn}} = 285.1$ J/K, $\Delta G^\circ_{\text{rxn}} = 91.2$ kJ; nonspontaneous, becomes spontaneous at high temperatures

c. $\Delta H^\circ_{\text{rxn}} = 98.8$ kJ, $\Delta S^\circ_{\text{rxn}} = 141.5$ J/K, $\Delta G^\circ_{\text{rxn}} = 56.6$ kJ; nonspontaneous, becomes spontaneous at high temperatures

d. $\Delta H^\circ_{\text{rxn}} = -91.8$ kJ, $\Delta S^\circ_{\text{rxn}} = -198.1$ J/K, $\Delta G^\circ_{\text{rxn}} = -32.8$ kJ; spontaneous

59. **a.** 2.8 kJ
b. 91.2 kJ
c. 56.4 kJ
d. −32.8 kJ
Values are comparable. The method using ΔH° and ΔS° can be used to determine how ΔG° changes with temperature.

61. **a.** −72.5 kJ, spontaneous
b. −11.4 kJ, spontaneous
c. 9.1 kJ, nonspontaneous

63. −29.4 kJ

65. **a.** 19.3 kJ
b. (i) 2.9 kJ (ii) −2.9 kJ
c. The partial pressure of iodine is very low.

67. 11.9 kJ

69. **a.** 1.48×10^{90} **b.** 2.09×10^{-26}

71. **a.** −24.8 kJ **b.** 0 **c.** 9.4 kJ

73. **a.** 1.90×10^{47} **b.** 1.51×10^{-13}

75. $\Delta H^\circ = 50.6$ kJ
$\Delta S^\circ = 226\ \text{J} \cdot \text{K}$

77. 4.8

79. **a.** + **b.** − **c.** −

81. **a.** $\Delta G^\circ = 175.2$ kJ, $K = 1.95 \times 10^{-31}$, nonspontaneous
b. 133 kJ, yes

83. Cl_2: $\Delta H^\circ_{\text{rxn}} = -182.1$ kJ, $\Delta S^\circ_{\text{rxn}} = -134.4$ J/K,
$\Delta G^\circ_{\text{rxn}} = -142.0$ kJ $K = 7.94 \times 10^{24}$
Br_2: $\Delta H^\circ_{\text{rxn}} = -121.6$ kJ, $\Delta S^\circ_{\text{rxn}} = -134.2$ J/K,
$\Delta G^\circ_{\text{rxn}} = -81.6$ kJ $K = 2.02 \times 10^{14}$
I_2: $\Delta H^\circ_{\text{rxn}} = -48.3$ kJ, $\Delta S^\circ_{\text{rxn}} = -132.2$ J/K,
$\Delta G^\circ_{\text{rxn}} = -8.9$ kJ $K = 37$
Cl_2 is the most spontaneous, I_2 is the least. Spontaneity is determined by the standard enthalpy of formation of the dihalogenated ethane. Higher temperatures make the reactions less spontaneous.

85. **a.** 107.8 kJ **b.** 5.0×10^{-7} atm
c. spontaneous at higher temperatures, $T = 923.4$ K

87. **a.** 2.22×10^5 **b.** 94.4 mol

89. **a.** $\Delta G^\circ = -689.6$ kJ, ΔG° becomes less negative
b. $\Delta G^\circ = -665.2$ kJ, ΔG° becomes less negative
c. $\Delta G^\circ = -632.4$ kJ, ΔG° becomes less negative
d. $\Delta G^\circ = -549.3$ kJ, ΔG° becomes less negative

91. With one exception, the formation of any oxide of nitrogen at 298 K requires more moles of gas as reactants than are formed as products. For example, 1 mol of N_2O requires 0.5 mol of O_2 and 1 mol of N_2, 1 mol of N_2O_3 requires 1 mol of N_2 and 1.5 mol of O_2, and so on. The exception is NO, where 1 mol of NO requires 0.5 mol of O_2 and 0.5 mol of N_2:

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

This reaction has a positive ΔS because what is essentially mixing of the N and O has taken place in the product.

93. 15.0 kJ

95. **a.** Positive, the process is spontaneous. It is slow unless a spark is applied.
b. Positive, although the change in the system is not spontaneous; the overall change, which includes such processes as combustion or water flow to generate electricity, is spontaneous.
c. Positive, the acorn-oak tree system is becoming more ordered, so the processes associated with growth are not spontaneous. But they are driven by spontaneous processes such as the generation of heat by the sun and the reactions that produce energy in the cell.

97. At 18.3 mmHg $\Delta G = 0$, at 760 mmHg
$\Delta G^\circ = 55.4$ kJ

99. **a.** 3.24×10^{-3}

b.

$$\begin{array}{r}
NH_3 + ATP + H_2O \longrightarrow NH_3{-}P_i + ADP \\
NH_3{-}P_i + C_5H_8O_4N^- \longrightarrow C_5H_9O_3N_2 + P_i + H_2O \\
\hline
NH_3 + C_5H_8O_4N^- + ATP \longrightarrow C_5H_9O_3N_2 + ADP + P_i \\
\Delta G^\circ = -16.3\text{ kJ}, K = 7.20 \times 10^2
\end{array}$$

101. **a.** −95.3 kJ/mol. Since the number of moles of reactants and products are the same, the decrease in volume affects the entropy of both equally, so there is no change in ΔG.
b. 102.8 kJ/mol. The entropy of the reactants (1.5 mol) is decreased more than the entropy of the product (1 mol). Since the product is relatively more favored at lower volume, ΔG is less positive.
c. 204.2 kJ/mol. The entropy of the product (1 mol) is decreased more than the entropy of the reactant (0.5 mol). Since the product is relatively less favored, ΔG is more positive.

103. $\Delta H^\circ = -93$ kJ, $\Delta S^\circ = -2.0 \times 10^2$ J/K

105. ΔS_{vap} diethyl ether = 86.1 J/mol K,
ΔS_{vap} acetone = 88.4 J/mol K,
ΔS_{vap} benzene = 87.3 J/mol K,
ΔS_{vap} chloroform = 88.0 J/mol K
Because water and ethanol hydrogen bond they are more ordered in the liquid and we expect ΔS_{vap} to be more positive. Ethanol 38600/351.0 = 110 J/mol K, H_2O = 40700/373.2 = 109 J/mol K

107. c
109. b
111. c
113. ΔG°_{rxn} is negative and ΔG_{rxn} is positive.

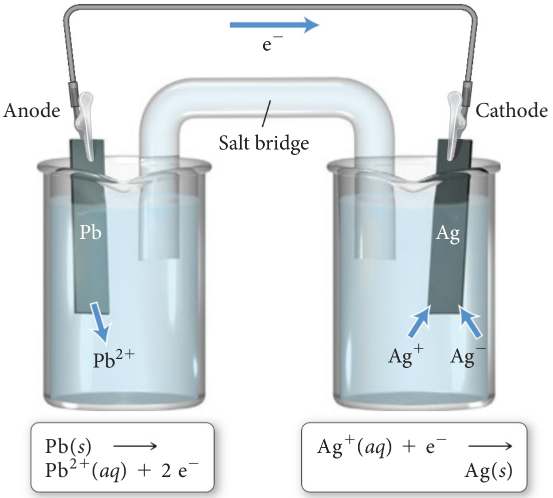
Chapter 18

37. **a.** $3\,K(s) + Cr^{3+}(aq) \longrightarrow Cr(s) + 3\,K^{+}(aq)$
b. $2\,Al(s) + 3\,Fe^{2+}(aq) \longrightarrow 2\,Al^{3+}(aq) + 3\,Fe(s)$
c. $2\,BrO_3^{-}(aq) + 3\,N_2H_4(g) \longrightarrow 2\,Br^{-}(aq) + 3\,N_2(g) + 6\,H_2O(l)$

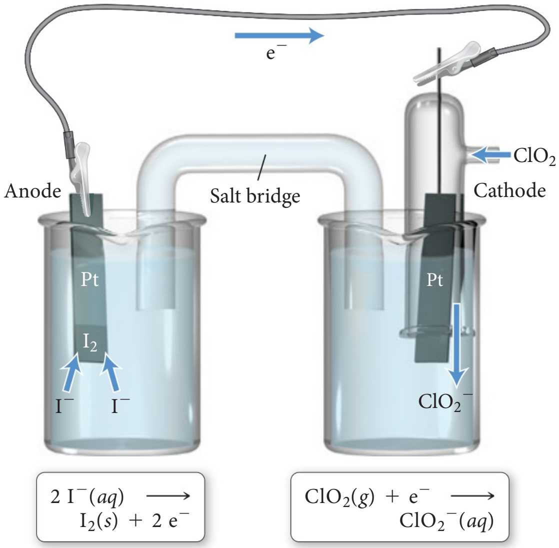
39. **a.** $PbO_2(s) + 2\,I^{-}(aq) + 4\,H^{+}(aq) \longrightarrow Pb^{2+}(aq) + I_2(s) + 2\,H_2O(l)$
b. $5\,SO_3^{2-}(aq) + 2\,MnO_4^{-}(aq) + 6\,H^{+}(aq) \longrightarrow 5\,SO_4^{2-}(aq) + 2\,Mn^{2+}(aq) + 3\,H_2O(l)$
c. $S_2O_3^{2-}(aq) + 4\,Cl_2(g) + 5\,H_2O(l) \longrightarrow 2\,SO_4^{2-}(aq) + 8\,Cl^{-}(aq) + 10\,H^{+}(aq)$

41. **a.** $H_2O_2(aq) + 2\,ClO_2(aq) + 2\,OH^{-}(aq) \longrightarrow O_2(g) + 2\,ClO_2^{-}(aq) + 2\,H_2O(l)$
b. $Al(s) + MnO_4^{-}(aq) + 2\,H_2O(l) \longrightarrow Al(OH)_4^{-}(aq) + MnO_2(s)$
c. $Cl_2(g) + 2\,OH^{-}(aq) \longrightarrow Cl^{-}(aq) + ClO^{-}(aq) + H_2O(l)$

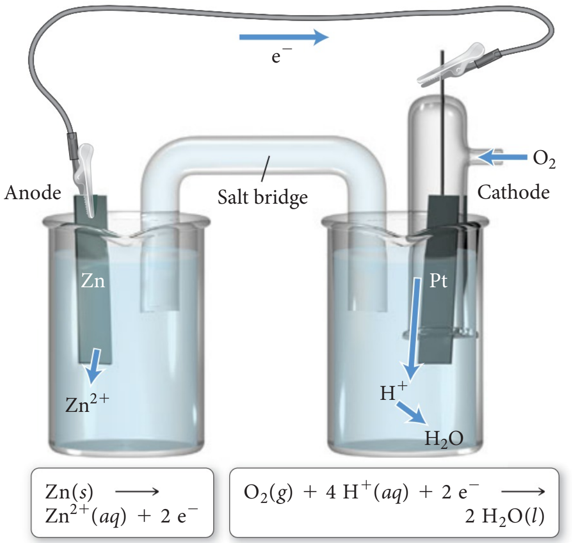
43. **a.**


e⁻
Anode
Salt bridge
Cathode
Pb
Ag
Pb²⁺
Ag⁺
Ag⁺
Pb(s) ⟶ Pb²⁺(aq) + 2 e⁻
Ag⁺(aq) + e⁻ ⟶ Ag(s)

b.

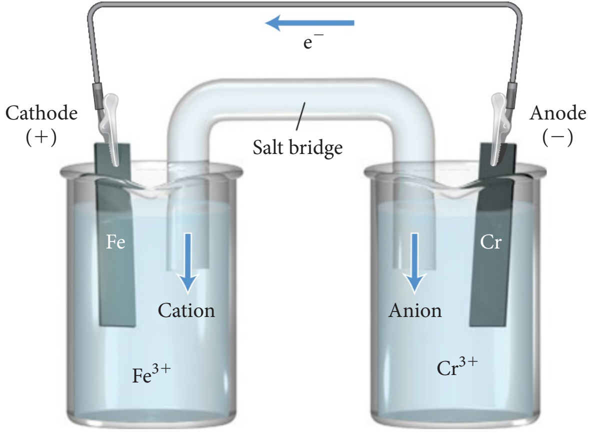

e⁻
Anode
Salt bridge
ClO₂
Cathode
Pt
Pt
I₂
I⁻
I⁻
ClO₂⁻
2 I⁻(aq) ⟶ I₂(s) + 2 e⁻
ClO₂(g) + e⁻ ⟶ ClO₂⁻(aq)

c.


e⁻
Anode
Salt bridge
O₂
Cathode
Zn
Pt
Zn²⁺
H⁺
H₂O
Zn(s) ⟶ Zn²⁺(aq) + 2 e⁻
O₂(g) + 4 H⁺(aq) + 2 e⁻ ⟶ 2 H₂O(l)

45. **a.** 0.93 V **b.** 0.41 V **c.** 1.99 V

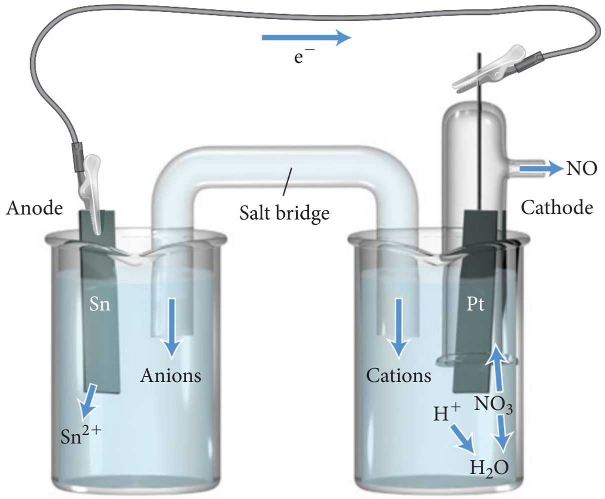
47. **a, c, d**


e⁻
Cathode (+)
Salt bridge
Anode (−)
Fe
Cr
Cation
Anion
Fe³⁺
Cr³⁺

b. $Cr(s) + Fe^{3+}(aq) \longrightarrow Cr^{3+}(aq) + Fe(s)$, $E^\circ_{cell} = 0.69\ V$

49. **a.** $Pb(s)\,|\,Pb^{2+}(aq)\,\|\,Ag^{+}(aq)\,|\,Ag(s)$
b. $Pt(s),\ I_2(s)\,|\,I^{-}(aq)\,\|\,ClO_2^{-}(aq)\,|\,ClO_2(g)\,|\,Pt(s)$
c. $Zn(s)\,|\,Zn^{2+}(aq)\,\|\,H_2O(l)\,|\,H^{+}(aq)\,|\,O_2(g)\,|\,Pt(s)$

51.


e⁻
Anode
Salt bridge
NO
Cathode
Sn
Pt
Anions
Cations
Sn²⁺
H⁺
NO₃⁻
H₂O

$3\,Sn(s) + 2\,NO_3^{-}(aq) + 8\,H^{+}(aq) \longrightarrow 3\,Sn^{2+}(aq) + 2\,NO(g) + 4\,H_2O(l)$, $E^\circ_{cell} = 1.10\ V$

53. **b, c** occur spontaneously in the forward direction
55. aluminum

57. **a.** yes, $2\ Al(s) + 6\ H^+(aq) \longrightarrow 2\ Al^{3+}(aq) + 3\ H_2(g)$
b. no
c. yes, $Pb(s) + 2\ H^+(aq) \longrightarrow Pb^{2+}(aq) + H_2(g)$

59. **a.** yes, $3\ Cu(s) + 2\ NO_3^-(aq) + 8\ H^+(aq) \longrightarrow 3\ Cu^{2+}(aq) + 2\ NO(g) + 4\ H_2O(l)$
b. no

61. **a.** -1.70 V, nonspontaneous **b.** 1.97 V, spontaneous
c. -1.51 V, nonspontaneous

63. a

65. **a.** -432 kJ **b.** 52 kJ **c.** -1.7×10^2 kJ

67. **a.** 5.31×10^{75} **b.** 7.7×10^{-10} **c.** 6.3×10^{29}

69. 5.6×10^5

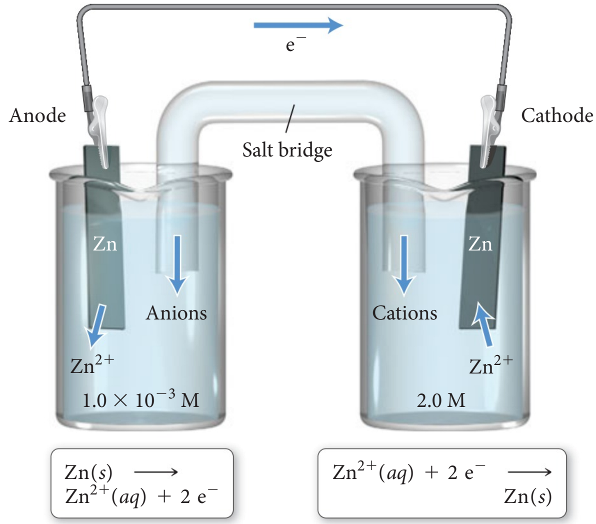
71. $\Delta G^\circ = -7.\underline{9}7$ kJ, $E^\circ_{cell} = 0.041$ V

73. **a.** 1.04 V **b.** 0.97 V **c.** 1.11 V

75. 1.87 V

77. **a.** 0.56 V **b.** 0.52 V
c. $[Ni^{2+}] = 0.003$ M, $[Zn^{2+}] = 1.60$ M

79.


e⁻
Anode
Cathode
Salt bridge
Zn
Zn
Anions
Cations
Zn²⁺
Zn²⁺
1.0 × 10⁻³ M
2.0 M
Zn(s) ⟶ Zn²⁺(aq) + 2 e⁻
Zn²⁺(aq) + 2 e⁻ ⟶ Zn(s)

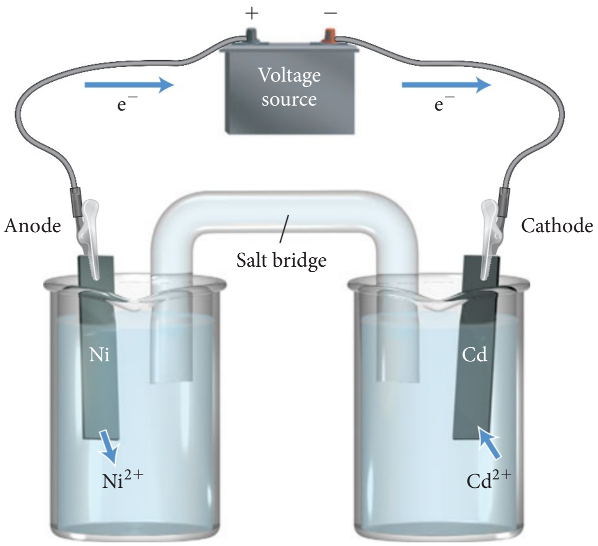
81. $\dfrac{[Sn^{2+}](ox)}{[Sn^{2+}](red)} = 4.2 \times 10^{-4}$

83. 0.3762

85. 1.038 V

87. **a, c** would prevent the corrosion of iron

89.


+
−
Voltage source
e⁻
e⁻
Anode
Cathode
Salt bridge
Ni
Cd
Ni²⁺
Cd²⁺

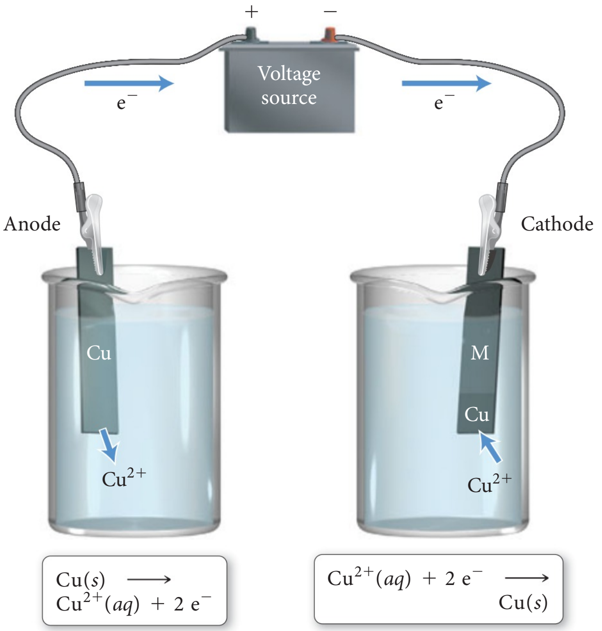
minimum voltage = 0.17 V

91. oxidation: $2\ Br^-(l) \longrightarrow Br_2(g) + 2\ e^-$
reduction: $K^+(l) + e^- \longrightarrow K(l)$

93. oxidation: $2\ Br^-(l) \longrightarrow Br_2(g) + 2\ e^-$
reduction: $K^+(l) + e^- \longrightarrow K(l)$

95. **a.** anode: $2\ Br^- \longrightarrow Br_2(l) + 2\ e^-$
cathode: $2\ H_2O(l) + 2\ e^- \longrightarrow H_2(g) + 2\ OH^-(aq)$
b. anode: $2\ I^-(aq) \longrightarrow I_2(s) + 2\ e^-$
cathode: $Pb^{2+}(aq) + 2\ e^- \longrightarrow Pb(s)$
c. anode: $2\ H_2O(l) \longrightarrow O_2(g) + 4\ H^+(aq) + 4\ e^-$
cathode: $2\ H_2O(l) + 2\ e^- \longrightarrow H_2(g) + 2\ OH^-(aq)$

97.


+
−
Voltage source
e⁻
e⁻
Anode
Cathode
Cu
M
Cu
Cu²⁺
Cu²⁺
Cu(s) ⟶ Cu²⁺(aq) + 2 e⁻
Cu²⁺(aq) + 2 e⁻ ⟶ Cu(s)

99. 1.8×10^2 s

101. 1.2×10^3 A

103. $2\ MnO_4^-(aq) + 5\ Zn(s) + 16\ H^+(aq) \longrightarrow 2\ Mn^{2+}(aq) + 5\ Zn^{2+}(aq) + 8\ H_2O(l)$
34.9 mL

105. The drawing should show that several Al atoms dissolve into solution as Al^{3+} ions and that several Cu^{2+} ions are deposited on the Al surface as solid Cu.

107. **a.** 68.3 mL
b. cannot be dissolved
c. cannot be dissolved

109. 0.25

111. There are no paired reactions that produce more than about 5 or 6 V.

113. **a.** 2.83 V
b. 2.71 V
c. 16 hr

115. 176 hr

117. 0.71 V

119. **a.** $\Delta G^\circ = 461$ kJ, $K = 1.4 \times 10^{-81}$
b. $\Delta G^\circ = 2.7 \times 10^2$ kJ, $K = 2.0 \times 10^{-48}$

121. MCl_4

123. 51.3%

125. pH = 0.85

127. 0.83 M

129. 4.1×10^5 L

131. 435 s

133. 8.39% U

135. The overall cell reaction for both cells is $2\,\text{Cu}^{+}(aq) \longrightarrow \text{Cu}^{2+}(aq) + \text{Cu}(s)$. The difference in E° is because n = 1 for the first cell and n = 2 for the second cell. For both cells, $\Delta G^\circ = -35.1$ kJ.

137. a

139. $\Delta G^\circ_{\text{rxn}}$ is positive and E°_{cell} is negative.

Chapter 19

31. **a.** ${}^{234}_{92}\text{U} \longrightarrow {}^{4}_{2}\text{He} + {}^{230}_{90}\text{Th}$
b. ${}^{230}_{90}\text{Th} \longrightarrow {}^{4}_{2}\text{He} + {}^{226}_{88}\text{Ra}$
c. ${}^{214}_{82}\text{Pb} \longrightarrow {}^{0}_{-1}\text{e} + {}^{214}_{83}\text{Bi}$
d. ${}^{13}_{7}\text{N} \longrightarrow {}^{0}_{+1}\text{e} + {}^{13}_{6}\text{C}$
e. ${}^{51}_{24}\text{Cr} + {}^{0}_{-1}\text{e} \longrightarrow {}^{51}_{23}\text{V}$

33. ${}^{232}_{90}\text{Th} \longrightarrow {}^{4}_{2}\text{He} + {}^{228}_{88}\text{Ra}$
${}^{228}_{88}\text{Ra} \longrightarrow {}^{0}_{-1}\text{e} + {}^{228}_{89}\text{Ac}$
${}^{228}_{89}\text{Ac} \longrightarrow {}^{0}_{-1}\text{e} + {}^{228}_{90}\text{Th}$
${}^{228}_{90}\text{Th} \longrightarrow {}^{4}_{2}\text{He} + {}^{224}_{88}\text{Ra}$

35. **a.** ${}^{221}_{87}\text{Fr}$
b. ${}^{0}_{-1}\text{e}$
c. ${}^{0}_{+1}\text{e}$
d. ${}^{0}_{-1}\text{e}$

37. **a.** stable, N/Z ratio is close to 1, acceptable for low Z atoms
b. not stable, N/Z ratio is much too high for low Z atom
c. not stable, N/Z ratio is less than 1, much too low
d. stable, N/Z ratio is acceptable for this Z

39. Sc, V, and Mn, each have odd numbers of protons. Atoms with an odd number of protons typically have less stable isotopes than those with an even number of protons.

41. **a.** beta decay
b. positron emission
c. positron emission
d. positron emission

43. **a.** Cs-125
b. Fe-62

45. 2.34×10^{9} years

47. 0.57 g

49. 10.8 hrs

51. 2.66×10^{3} yr

53. 2.4×10^{4} yr

55. 2.7×10^{9} yr

57. ${}^{235}_{92}\text{U} + {}^{1}_{0}\text{n} \longrightarrow {}^{144}_{54}\text{Xe} + {}^{90}_{38}\text{Sr} + 2\,{}^{1}_{0}\text{n}$

59. ${}^{2}_{1}\text{H} + {}^{2}_{1}\text{H} \longrightarrow {}^{3}_{2}\text{He} + {}^{1}_{0}\text{n}$

61. ${}^{238}_{92}\text{U} + {}^{1}_{0}\text{n} \longrightarrow {}^{239}_{92}\text{U}$
${}^{239}_{92}\text{U} \longrightarrow {}^{239}_{93}\text{Np} + {}^{0}_{-1}\text{e}$
${}^{239}_{93}\text{Np} \longrightarrow {}^{239}_{94}\text{Pu} + {}^{0}_{-1}\text{e}$

63. ${}^{249}_{98}\text{Cf} + {}^{12}_{6}\text{C} \longrightarrow {}^{257}_{104}\text{Rf} + 4\,{}^{1}_{0}\text{n}$

65. 9.0×10^{13} J

67. **a.** mass defect = 0.13701 amu
binding energy = 7.976 MeV/nucleon
b. mass defect = 0.54369 amu
binding energy = 8.732 MeV/nucleon
c. mass defect = 1.16754 amu
binding energy = 8.431 MeV/nucleon

69. 7.228×10^{10} J/g U-235

71. 7.84×10^{10} J/g H-2

73. radiation: 25 J, fall: 370 J

75. 68 mi

77. **a.** ${}^{1}_{1}\text{p} + {}^{9}_{4}\text{Be} \longrightarrow {}^{6}_{3}\text{Li} + {}^{4}_{2}\text{He}$
1.03×10^{11} J/mol
b. ${}^{209}_{83}\text{Bi} + {}^{64}_{28}\text{Ni} \longrightarrow {}^{272}_{111}\text{Rg} + {}^{1}_{0}\text{n}$
1.141×10^{13} J/mol
c. ${}^{179}_{74}\text{W} + {}^{0}_{-1}\text{e} \longrightarrow {}^{179}_{73}\text{Ta}$
7.59×10^{10} J/mol

79. **a.** ${}^{114}_{44}\text{Ru} \longrightarrow {}^{0}_{-1}\text{e} + {}^{114}_{45}\text{Rh}$
b. ${}^{216}_{88}\text{Ra} \longrightarrow {}^{0}_{+1}\text{e} + {}^{216}_{87}\text{Fr}$
c. ${}^{58}_{30}\text{Zn} \longrightarrow {}^{0}_{+1}\text{e} + {}^{58}_{29}\text{Cu}$
d. ${}^{31}_{10}\text{Ne} \longrightarrow {}^{0}_{-1}\text{e} + {}^{31}_{11}\text{Na}$

81. 2.9×10^{21} beta emissions, 3700 Ci

83. 1.6×10^{-5} L

85. 4.94×10^{7} kJ/mol

87. 7.72 MeV

89. ${}^{14}\text{N}$

91. 0.15%

93. 1.24×10^{21} atoms

95. 2.42×10^{-12} m

97. −0.7 MeV, there is no coulombic barrier for collision with a neutron.

99. **a.** 1.164×10^{10} kJ
b. 0.1299 g

101. U-235 forms Pb-207 in seven α-decays and four β-decays and Th-232 forms Pb-208 in six α-decays and four β-decays.

103. 3.0×10^{2} K

105. ${}^{21}_{9}\text{F} \longrightarrow {}^{21}_{10}\text{Ne} + {}^{0}_{-1}\text{e}$

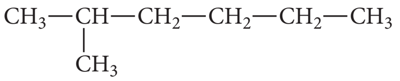
107. Nuclide A is more dangerous because the half-life is shorter (18.5 days) and so it decays faster.

109. Iodine is used by the thyroid gland to make hormones. Normally we ingest iodine in foods, especially iodized salt. The thyroid gland cannot tell the difference between stable and radioactive iodine and will absorb both. KI tablets work by blocking radioactive iodine from entering the thyroid. When a person takes KI, the stable iodine in the tablet gets absorbed by the thyroid. Because KI contains so much stable iodine, the thyroid gland becomes "full" and cannot absorb any more iodine—either stable or radioactive—for the next 24 hours.

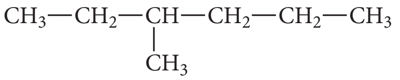
Chapter 20

35. **a.** alkane
b. alkene
c. alkyne
d. alkene

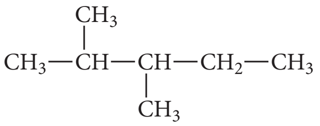
37. $CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$



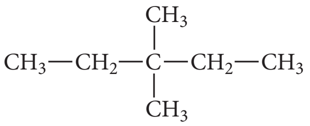
$CH_3-\underset{CH_3}{CH}-CH_2-CH_2-CH_2-CH_3$



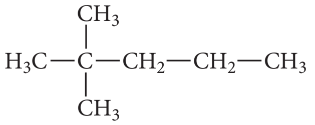
$CH_3-CH_2-\underset{CH_3}{CH}-CH_2-CH_2-CH_3$



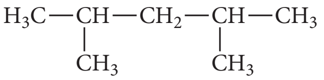
$CH_3-\overset{CH_3}{CH}-\underset{CH_3}{CH}-CH_2-CH_3$



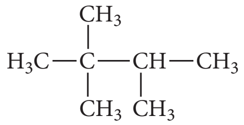
$CH_3-CH_2-\overset{CH_3}{\underset{CH_3}{C}}-CH_2-CH_3$



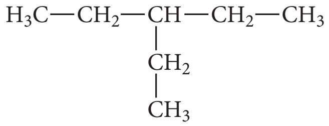
$H_3C-\overset{CH_3}{\underset{CH_3}{C}}-CH_2-CH_2-CH_3$



$H_3C-\underset{CH_3}{CH}-CH_2-\underset{CH_3}{CH}-CH_3$



$H_3C-\overset{CH_3}{\underset{CH_3}{C}}-\underset{CH_3}{CH}-CH_3$



$H_3C-CH_2-\underset{\underset{CH_3}{CH_2}}{CH}-CH_2-CH_3$

39. **a.** no **b.** yes **c.** yes **d.** no

41. **a.** enantiomers **b.** same

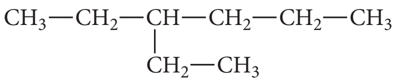
c. enantiomers

43. **a.** pentane **b.** 2-methylbutane

c. 4-isopropyl-2-methylheptane

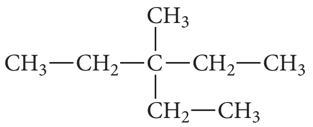
d. 4-ethyl-2-methylhexane

45. **a.**



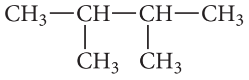
$CH_3-CH_2-\underset{CH_2-CH_3}{CH}-CH_2-CH_2-CH_3$

b.



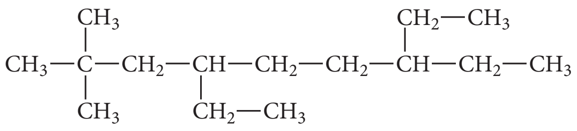
$CH_3-CH_2-\overset{CH_3}{\underset{CH_2-CH_3}{C}}-CH_2-CH_3$

c.



$CH_3-\underset{CH_3}{CH}-\underset{CH_3}{CH}-CH_3$

d.



$CH_3-\overset{CH_3}{\underset{CH_3}{C}}-CH_2-\underset{CH_2-CH_3}{CH}-CH_2-CH_2-\overset{CH_2-CH_3}{CH}-CH_2-CH_3$

47. **a.** $CH_3CH_2CH_3 + 5\ O_2 \longrightarrow 3\ CO_2 + 4\ H_2O$

b. $CH_3CH_2CH{=}CH_2 + 6\ O_2 \longrightarrow 4\ CO_2 + 4\ H_2O$

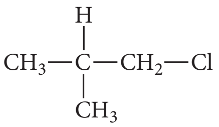
c. $2\ CH{\equiv}CH + 5\ O_2 \longrightarrow 4\ CO_2 + 2\ H_2O$

49. **a.** CH_3CH_2Br

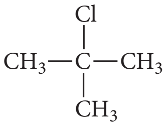
b. $CH_3CH_2CH_2Cl$, $CH_3CHClCH_3$

c. $CHCl_2Br$

d.



$CH_3-\overset{H}{\underset{CH_3}{C}}-CH_2-Cl$



$CH_3-\overset{Cl}{\underset{CH_3}{C}}-CH_3$

51. $CH_2{=}CH-CH_2-CH_2-CH_2-CH_3$

$CH_3-CH{=}CH-CH_2-CH_2-CH_3$

$CH_3-CH_2-CH{=}CH-CH_2-CH_3$

53. **a.** 1-butene **b.** 3,4-dimethyl-2-pentene

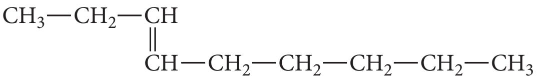
c. 3-isopropyl-1-hexene **d.** 2,4-dimethyl-3-hexene

55. **a.** 2-butyne **b.** 4,4-dimethyl-2-hexyne

c. 3-isopropyl-1-hexyne **d.** 3,6-dimethyl-4-nonyne

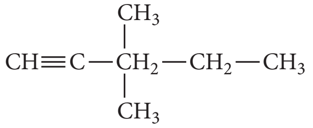
57. **a.** $CH_3-CH_2-CH-C{\equiv}C-CH_2-CH_2-CH_3$

b.



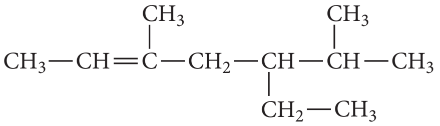
$CH_3-CH_2-CH{=}CH-CH_2-CH_2-CH_2-CH_2-CH_3$

c.



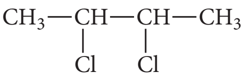
$CH{\equiv}C-\overset{CH_3}{\underset{CH_3}{CH_2}}-CH_2-CH_3$

d.



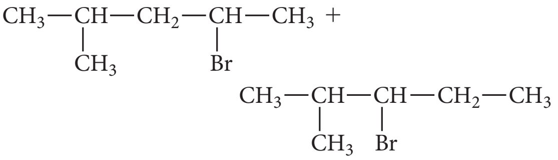
$CH_3-CH{=}\overset{CH_3}{C}-CH_2-\underset{CH_2-CH_3}{CH}-\overset{CH_3}{CH}-CH_3$

59. **a.**



$CH_3-\underset{Cl}{CH}-\underset{Cl}{CH}-CH_3$

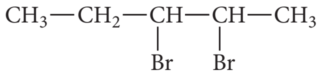
b.



$CH_3-\underset{CH_3}{CH}-CH_2-\underset{Br}{CH}-CH_3 +$

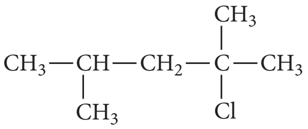
$CH_3-\underset{CH_3}{CH}-\underset{Br}{CH}-CH_2-CH_3$

c.



$CH_3-CH_2-\underset{Br}{CH}-\underset{Br}{CH}-CH_3$

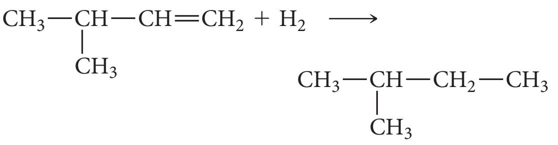
d.



$CH_3-\underset{CH_3}{CH}-CH_2-\overset{CH_3}{\underset{Cl}{C}}-CH_3$

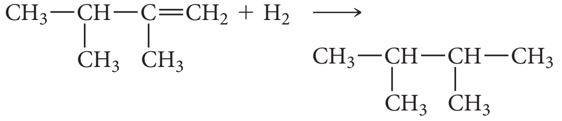
61. **a.** $CH_2{=}CH-CH_3 + H_2 \longrightarrow CH_3-CH_2-CH_3$

b.



$CH_3-\underset{CH_3}{CH}-CH{=}CH_2 + H_2 \longrightarrow CH_3-\underset{CH_3}{CH}-CH_2-CH_3$

c.



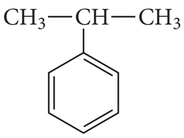
$CH_3-\underset{CH_3}{CH}-\underset{CH_3}{C}{=}CH_2 + H_2 \longrightarrow CH_3-\underset{CH_3}{CH}-\underset{CH_3}{CH}-CH_3$

63. **a.** methylbenzene or toluene
b. bromobenzene
c. chlorobenzene

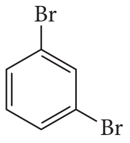
65. **a.** 3,5-dimethyl-7-phenylnonane
b. 2-phenyl-3-octene
c. 4,5-dimethyl-6-phenyl-2-octyne

67. **a.** 1,4-dibromobenzene or *p*-dibromobenzene
b. 1,3-diethylbenzene or *m*-diethylbenzene
c. 1-chloro-2-fluorobenzene or *o*-chlorofluorobenzene

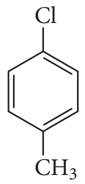
69. **a.**



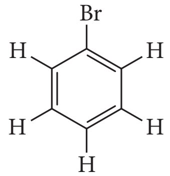
b.



c.

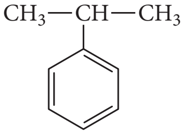


71. **a.**



+ HBr

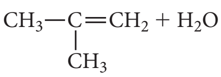
b.



+ HCl

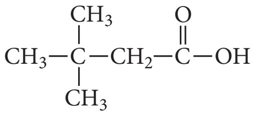
73. **a.** 1-propanol **b.** 4-methyl-2-hexanol
c. 2,6-dimethyl-4-heptanol **d.** 3-methyl-3-pentanol

75. **a.** $CH_3CH_2CH_2Br + H_2O$ **b.**



$CH_3-\underset{\underset{CH_3}{|}}{C}=CH_2 + H_2O$

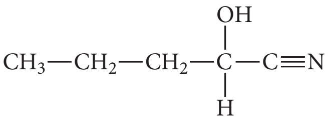
c. $CH_3CH_2ONa + \frac{1}{2}H_2$ **d.**



$CH_3-\overset{CH_3}{\overset{|}{\underset{\underset{CH_3}{|}}{C}}}-CH_2-\overset{O}{\overset{\|}{C}}-OH$

77. **a.** butanone **b.** pentanal
c. 3,5,5-trimethylhexanal **d.** 4-methyl-2-hexanone

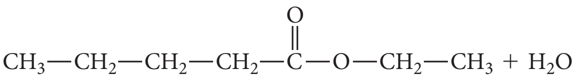
79.



$CH_3-CH_2-CH_2-\overset{OH}{\overset{|}{\underset{\underset{H}{|}}{C}}}-C\equiv N$

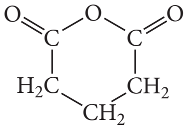
81. **a.** methylbutanoate **b.** propanoic acid
c. 5-methylhexanoic acid **d.** ethylpentanoate

83. **a.**



$CH_3-CH_2-CH_2-CH_2-\overset{O}{\overset{\|}{C}}-O-CH_2-CH_3 + H_2O$

b.



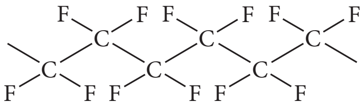
+ H_2O

85. **a.** ethyl propyl ether **b.** ethyl pentyl ether
c. dipropyl ether **d.** butyl ethyl ether

87. **a.** diethylamine **b.** methylpropylamine
c. butylmethylpropylamine

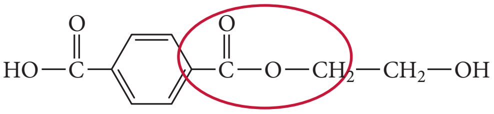
89. **a.** acid–base, $(CH_3)_2NH_2^+(aq) + Cl^-(aq)$
b. condensation, $CH_3CH_2CONHCH_2CH_3(aq) + H_2O$
c. acid–base, $CH_3NH_3^+(aq) + HSO_4^-(aq)$

91. …



…

93.



95. **a.** ester, methyl 3-methylbutanoate
b. ether, ethyl 2-methylbutyl ether
c. aromatic, 1-ethyl-3-methylbenzene or *m*-ethylmethylbenzene
d. alkyne, 5-ethyl-4-methyl-2-heptyne
e. aldehyde, butanal
f. alcohol, 2-methyl-1-propanol

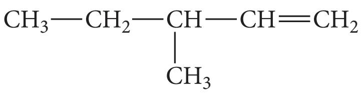
97. **a.** 5-isobutyl-3-methylnonane
b. 5-methyl-3-hexanone
c. 3-methyl-2-butanol
d. 4-ethyl-3,5-dimethyl-1-hexyne

99. **a.** isomers **b.** isomers **c.** same

101. 558 g

103. **a.** combustion **b.** alkane substitution
c. alcohol elimination **d.** aromatic substitution

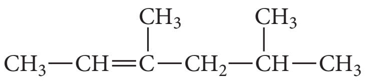
105. **a.**



$CH_3-CH_2-\underset{\underset{CH_3}{|}}{CH}-CH=CH_2$

Can exist as a stereoisomer

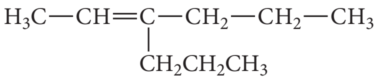
b.



$CH_3-CH=\overset{CH_3}{\overset{|}{C}}-CH_2-\overset{CH_3}{\overset{|}{CH}}-CH_3$

Can exist as a stereoisomer

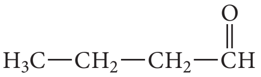
c.



$H_3C-CH=\underset{\underset{CH_2CH_2CH_3}{|}}{C}-CH_2-CH_2-CH_3$

Can exist as a stereoisomer

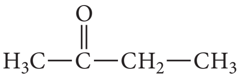
107. 1.



$H_3C-CH_2-CH_2-\overset{O}{\overset{\|}{C}}H$

Aldehyde

2.



$H_3C-\overset{O}{\overset{\|}{C}}-CH_2-CH_3$

Ketone

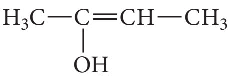
3. $H_3C-CH=CH-O-CH_3$
Alkene, ether

4. $H_2C=CH-O-CH_2-CH_3$
Alkene, ether

5. $H_2C=CH-CH_2-O-CH_3$
Alkene, ether

6. $H_3C-CH=CH-CH_2-OH$
Alkene, alcohol

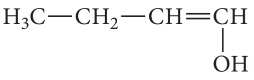
7.



$H_3C-\underset{\underset{OH}{|}}{C}=CH-CH_3$

Alkene, alcohol

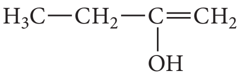
8.



$H_3C-CH_2-CH=\underset{\underset{OH}{|}}{CH}$

Alkene, alcohol

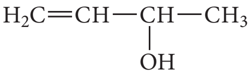
9.



$H_3C-CH_2-\underset{\underset{OH}{|}}{C}=CH_2$

Alkene, alcohol

10.



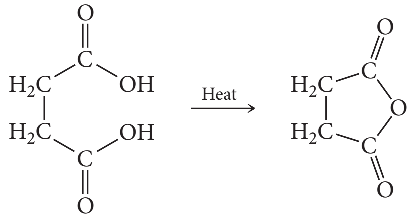
$H_2C=CH-\underset{\underset{OH}{|}}{CH}-CH_3$

Alkene, alcohol

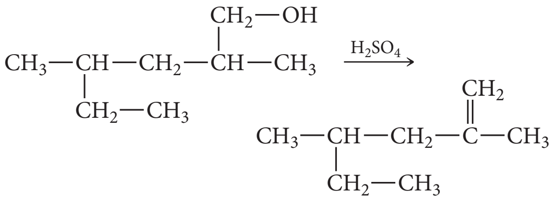
11. $H_2C=CH-CH_2-CH_2-OH$
Alkene, alcohol

109. In the acid form of the carboxylic acid, electron withdrawal by the C═O enhances acidity. The conjugate base, the carboxylate anion, is stabilized by resonance so the two O atoms are equivalent and bear the negative charge equally.

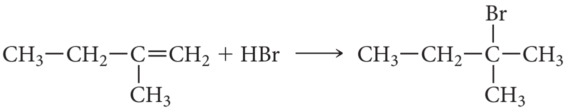
111. a.



b.



c.

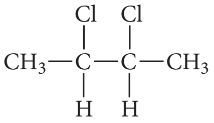


$CH_3-CH_2-\underset{CH_3}{C}=CH_2 + HBr \longrightarrow CH_3-CH_2-\overset{Br}{\underset{CH_3}{C}}-CH_3$

113. a. 3 : 1

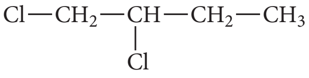
b. 2° hydrogen atoms are more reactive. The reactivity of 2° hydrogens to 1° hydrogens is 11 : 3.

115.



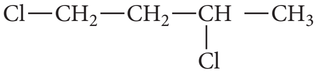
$CH_3-\overset{Cl}{\underset{H}{C}}-\overset{Cl}{\underset{H}{C}}-CH_3$

Chiral



$Cl-CH_2-\underset{Cl}{CH}-CH_2-CH_3$

Chiral



$Cl-CH_2-CH_2-\underset{Cl}{CH}-CH_3$

Chiral

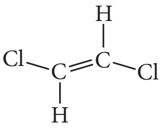
117. The first propagation step for F is very rapid and exothermic because of the strength of the H—F bond that forms. For I the first propagation step is endothermic and slow because the H—I bond that forms is relatively weak.

119.

$Cl-C\equiv C-Cl$

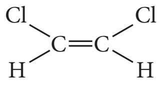
No dipole moment

trans



$ClCH_2-CH_2Cl$

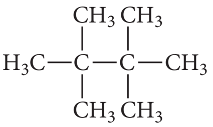
cis



Cl_2CHCH_3

Dipole moment

121.



$H_3C-\overset{CH_3}{\underset{CH_3}{C}}-\overset{CH_3}{\underset{CH_3}{C}}-CH_3$

2,2,3,3-tetramethylbutane

Chapter 21

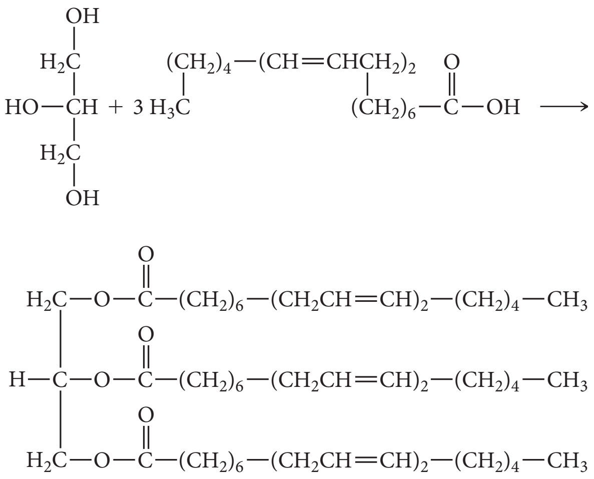
31. c. saturated fatty acid; **d.** steroid

33. a. saturated fatty acid **b.** not a fatty acid

c. not a fatty acid

d. monounsaturated fatty acid

35.



Triglyceride is expected to be an oil.

37. a. monosaccharide;

c. disaccharide

39. a. aldose, hexose

b. aldose, pentose

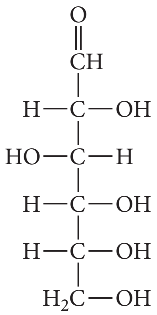
c. ketose, tetrose

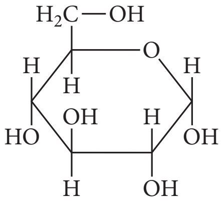
d. aldose, tetrose

41. a. 5 **b.** 3

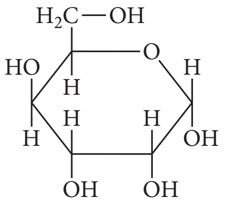
c. 1 **d.** 3

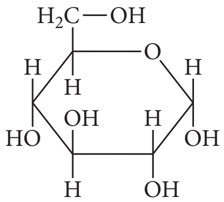
43.



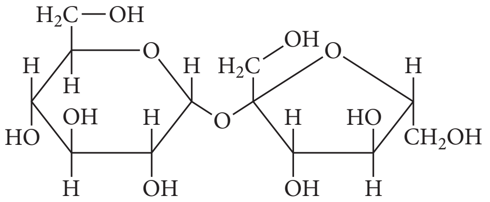


45.

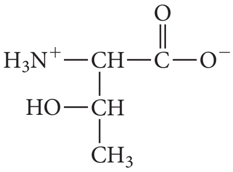


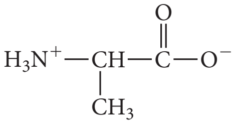


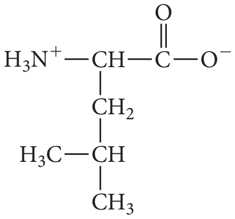
47.

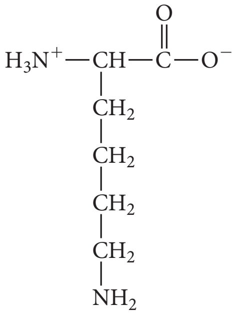


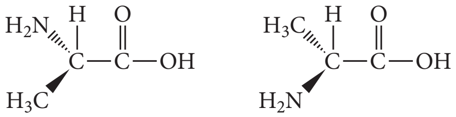
Glucose Fructose

49. a.

$H_3N^+{-}CH(CH(OH)CH_3){-}C(=O){-}O^-$

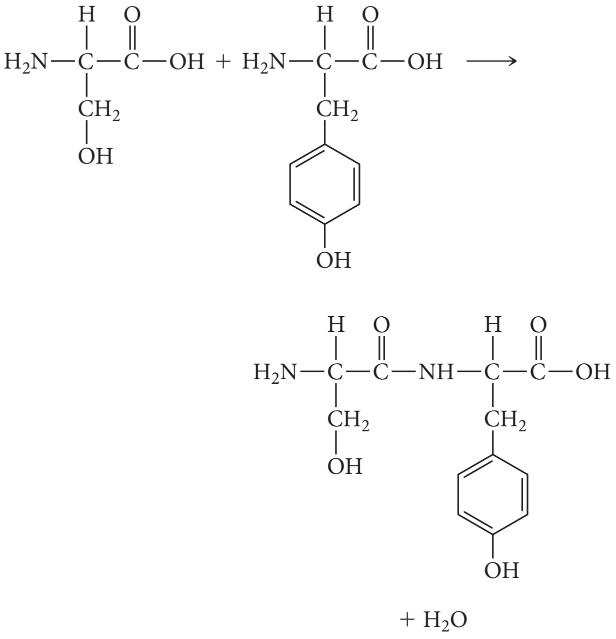
b.

$H_3N^+{-}CH(CH_3){-}C(=O){-}O^-$

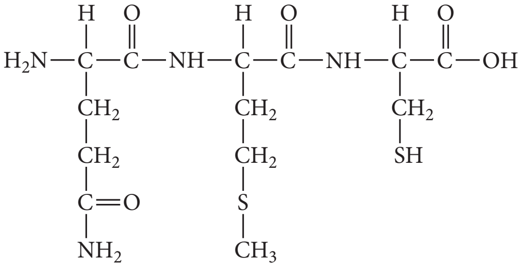
c.

$H_3N^+{-}CH(CH_2CH(CH_3)CH_3){-}C(=O){-}O^-$

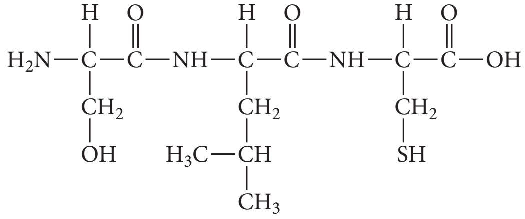
d.

$H_3N^+{-}CH(CH_2CH_2CH_2CH_2NH_2){-}C(=O){-}O^-$

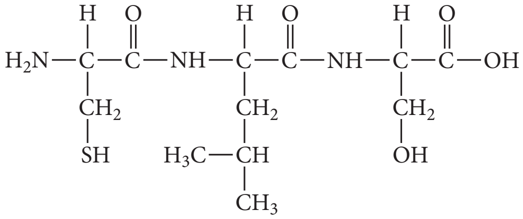
51.


53. 6, Ser-Gly-Cys, Ser-Cys-Gly, Gly-Ser-Cys, Gly-Cys-Ser, Cys-Ser-Gly, Cys-Gly-Ser

55.


57. a.

$H_2N{-}CH(CH_2CH_2C(=O)NH_2){-}C(=O){-}NH{-}CH(CH_2CH_2SCH_3){-}C(=O){-}NH{-}CH(CH_2SH){-}C(=O){-}OH$

b.

$H_2N{-}CH(CH_2OH){-}C(=O){-}NH{-}CH(CH_2CH(CH_3)CH_3){-}C(=O){-}NH{-}CH(CH_2SH){-}C(=O){-}OH$

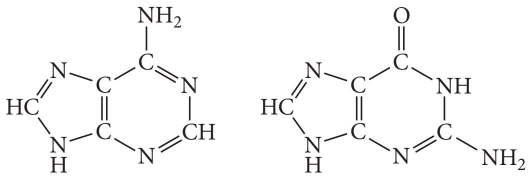
c.

$H_2N{-}CH(CH_2SH){-}C(=O){-}NH{-}CH(CH_2CH(CH_3)CH_3){-}C(=O){-}NH{-}CH(CH_2OH){-}C(=O){-}OH$

59. tertiary

61. primary

63. a. A;

c. T

65.


67. A C A T G C G

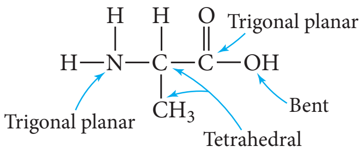
69. 154 codons, 462 nucleotides

71. a. protein

b. carbohydrate

c. lipid

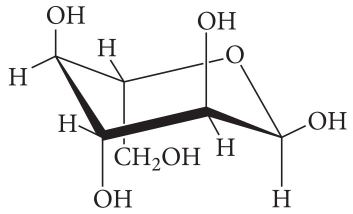
73. A codon is composed of three nucleotides. A codon codes for a specific amino acid while a gene codes for an entire protein.

75.


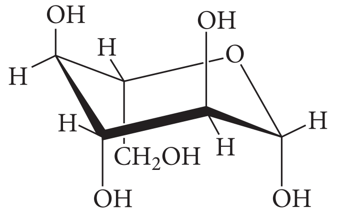
77. valine, leucine, isoleucine, phenylalanine

79. Gly-Arg-Ala-Leu-Phe-Gly-Asn-Lys-Trp-Glu-Cys

81.



a.



b.

83. As the temperature increases the favorable entropy for uncoiling a chain becomes dominant. On cooling the favorable enthalpy of forming hydrogen bonds between paired bases is dominant.

85. When the fake thymine nucleotide is added to the replicating DNA, the chain cannot continue to form, because the $-N{=}N^+{=}NH$ group on the sugar prevents future phosphate linkages.

87. $V_{max} = 47.6$, $K_t = 1.68$

89. $H_3N^+CH_2COO^- + H^+ \rightleftharpoons$ $(H_3N^+CH_2COOH\ [HA]/[A^-] = 2,$
$H_3N^+CH_2COO^- \rightleftharpoons$
$H_2NCH_2COO^- + H^+\ [HA]/[A^-] = 0.4$, pH $= 6.0$

91. A three-base codon codes for a single amino acid. If there are only three bases, there could be 27 different three-base codon arrangements. Therefore, you could theoretically code for the 20 different amino acids needed.

Chapter 22

17. **a.** +4 **b.** +4 **c.** +4

19. $Ca_3Al_2(SiO_4)_3$

21. 4

23. tetrahedrons stand alone, orthosilicates

25. amphibole or double-chain structure;
Ca^{2+}, Mg^{2+}, Fe^{2+}, Al^{3+}

27. 950 g

29. NCl_3 has a lone pair that BCl_3 lacks, giving it a trigonal pyramidal shape, as opposed to BCl_3's trigonal planar shape.

31. **a.** 6 vertices, 8 faces **b.** 12 vertices, 20 faces

33. *closo*-Boranes have the formula $B_nH_n^{2-}$ and form fully closed polyhedra, *nido*-boranes have the formula B_nH_{n+4} and consist of a cage missing a corner, and *arachno*-boranes have the formula B_nH_{n+6} and consist of a cage missing two or three corners.

35. Graphite consists of covalently bonded sheets that are held to each other by weak interactions, allowing them to slip past each other. Diamond is not a good lubricant, because it is an extremely strong network covalent solid, where all of the carbon atoms are covalently bonded.

37. Activated charcoal consists of fine particles, rather than a lump of charcoal, and subsequently has a much higher surface area.

39. Ionic carbides are composed of carbon, generally in the form of the carbide ion, C_2^{2-}, and low-electronegativity metals, such as the alkali and alkaline earth metals. Covalent carbides are composed of carbon and low-electronegativity nonmetals or metalloids, such as silicon.

41. **a.** solid $\longrightarrow$ gas
b. gas $\longrightarrow$ liquid $\longrightarrow$ solid
c. solid $\longrightarrow$ gas

43. **a.** $CO(g) + CuO(s) \longrightarrow CO_2(g) + Cu(s)$
b. $SiO_2(s) + 3\ C(s) \longrightarrow SiC(s) + 2\ CO(g)$
c. $S(s) + CO(g) \longrightarrow COS(g)$

45. **a.** +2 **b.** +4 **c.** +4/3

47. Fixing nitrogen refers to converting N_2 to a nitrogen-containing compound.

49. White phosphorus consists of P_4 molecules in a tetrahedral shape with the atoms at the corners of the tetrahedron. This allotrope is unstable because of the strain from the bond angles. Red phosphorus is much more stable because one bond of the tetrahedron is broken, allowing the phosphorus atoms to make chains with bond angles that are less strained.

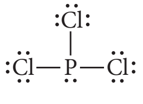
51. saltpeter: 13.86% N by mass
Chile saltpeter: 16.48% N by mass

53. HN_3 has a positive ΔG°_f, meaning that it spontaneously decomposes into H_2 and N_2 at room temperature. There are no temperatures at which HN_3 will be stable. ΔH_f is positive and ΔS_f is negative, so ΔG_f will always be negative.

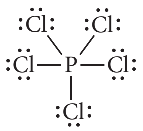
55. **a.** $NH_4NO_3(aq) + \text{heat} \longrightarrow N_2O(g) + 2\ H_2O(l)$
b. $3\ NO_2(g) + H_2O(l) \longrightarrow 2\ HNO_3(l) + NO(g)$
c. $2\ PCl_3(l) + O_2(g) \longrightarrow 2\ POCl_3(l)$

57. NO_3^-, NO_2^-, N_3^-, $N_2H_5^+$, NH_4^+

59.



Trigonal pyramidal



Trigonal bipyramidal

61. $CO(NH_2)_2 + 2\ H_2O \longrightarrow (NH_4)_2CO_3$
14 g

63. P_4O_6 forms if there is only a limited amount of oxygen available, while P_4O_{10} will form with greater amounts of oxygen.

65. The major source of oxygen is the fractionation of air by which air is cooled and liquefied and oxygen is separated from the other components.

67. **a.** superoxide **b.** oxide **c.** peroxide

69. Initially, liquid sulfur becomes less viscous when heated because the S_8 rings have greater thermal energy, which overcomes intermolecular forces. Above 150 °C the rings break and the broken rings entangle one another, causing greater viscosity.

71. **a.** 4×10^{-22} g **b.** 4.0×10^{-19} g

73. $2\ FeS_2(s) \xrightarrow{heat} 2\ FeS(s) + S_2(g)$
510 L

75. **a.** +2, linear **b.** +6, octahedral
c. +6, square pyramidal

77. $Cl_2(g) + 2\ Br^-(aq) \longrightarrow 2\ Cl^-(aq) + Br_2(l)$
Oxidizing agent: Cl_2
Reducing agent: Br^-

79. No, there is not enough HF to dissolve all of the SiO_2. HF is the limiting reagent. 1.6 g SiO_2.

81. 8 kg from lignite, 40 kg from bituminous

83. Chlorine is much more electronegative than iodine, allowing it to withdraw an electron and ionize in solution much more easily.

85. **a.** $rate_{HCl}/rate_{Cl_2} = 1.395$
b. $rate_{HCl}/rate_{HF} = 0.7407$
c. $rate_{HCl}/rate_{HI} = 1.873$

87. $4\ Na_2O_2 + 3\ Fe \longrightarrow 4\ Na_2O + Fe_3O_4$

89. The bond length of the O_2 species increases as electrons are added because they are added to the π^* antibonding orbital. O_2^{2-} is diamagnetic.

91. 2.0 mol of C—C bonds, 715 kJ/mol, 6.9×10^2 kJ/mol, This value calculated from the bond energy is too low because it doesn't include van der Waals attractions between C atoms not directly bonded to each other.

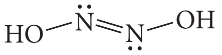
93. −50 kJ/mol

95. **a.** −13.6 kJ/mol **b.** −11.0 kJ/mol
c. −24.8 kJ/mol
Fe_2O_3 is the most exothermic because it has the highest oxidation state and is therefore able to oxidize the most CO per mol Fe.

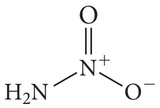
97. **a.** $\ddot{O}=C=C=C=\ddot{O}$ **b.** *sp*
c. −92 kJ/mol

99. **a.** 7.6×10^{-22} **b.** 1.2×10^{-8}
c. $[N_2H_4] = 0.009$ M, $[N_2H_5^+] = 0.0025$ M, $[N_2H_6^{2+}] = 7.0 \times 10^{-13}$ M

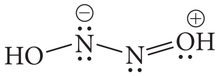
101. The acid is

a. 

and the base is

b. 

The acid is weaker than nitrous acid because of electron donation by resonance in contributing structures such as

c. 

The base is weaker than ammonia because of electron withdrawal by the electronegative nitro group.

103. The triple bond in nitrogen is much stronger than the double bond in oxygen, so it is much harder to break. This makes it less likely that the bond in nitrogen will be broken.

105. Sodium dinitrogen phosphate (NaH_2PO_4) can act as a weak base or a weak acid. A buffer can be made by mixing it with either Na_2HPO_4 or with Na_3PO_4, depending on the desired pH of the buffer solution.

107. F is extremely small and so there is a huge driving force to fill the octet by adding an electron, giving a −1 oxidation state. Other halogens have access to *d* orbitals, which allows for more hybridization and oxidation state options.

109. SO_3 cannot be a reducing agent, because the oxidation state of S is +6, the highest possible oxidation state for S. Reducing agents need to be able to be oxidized. SO_2 can be a reducing agent or an oxidizing agent, because the oxidation state of S is +4.

Chapter 23

15. Metals are typically opaque, are good conductors of heat and electricity, and are ductile and malleable, meaning they can be drawn into wires and flattened into sheets.

17. aluminum, iron, calcium, magnesium, sodium, potassium

19. Fe: hematite (Fe_2O_3), magnetite (Fe_3O_4)
Hg: cinnabar (HgS)
V: vanadite $[Pb_5(VO_4)Cl]$, carnotite $[K_2(UO_2)_2(VO_4)_2 \cdot 3\ H_2O]$
Nb: columbite $[Fe(NbO_3)_2]$

21. $MgCO_3(s) + heat \longrightarrow MgO(s) + CO_2(g)$
$Mg(OH)_2(s) + heat \longrightarrow MgO(s) + H_2O(g)$

23. The flux is a material that will react with the gangue to form a substance with a low melting point. MgO is the flux.

25. Hydrometallurgy is used to separate metals from ores by selectively dissolving the metal in a solution, filtering out impurities, and then reducing the metal to its elemental form.

27. The Bayer process is a hydrometallurgical process by which Al_2O_3 is selectively dissolved, leaving other oxides as solids. The soluble form of aluminum is $Al(OH)_4^-$.

29. Sponge powdered iron contains many small holes in the iron particles due to the escaping of the oxygen when the iron is reduced. Water atomized powdered iron has much smoother and denser particles as the powder is formed from molten iron.

31. **a.** 50% Cr, 50% V by moles; 50.5% Cr, 49.5% V by mass
b. 25% Fe, 75% V by moles; 26.8% Fe, 73.2% V by mass
c. 25% Cr, 25% Fe, 50% V by moles; 24.8% Cr, 26.6% Fe, 48.6% V by mass

33. Cr and Fe are very close to each other in mass, so their respective atomic radii are probably close enough to form an alloy. Also, they both form body-centered cubic structures.

35. A: solid, 20% Cr, 80% Fe
B: liquid, 50% Cr, 50% Fe

37. A: solid (20% Co and 80% Cu overall. Two phases; one is the Cu structure with 4% Co, and the other is the Co structure with 7% Cu. There will be more of the Cu structure).
B: solid (Co structure), 90% Co, 10% Cu

39. C would fill interstitial holes; Mn and Si would substitute for Fe.

41. **a.** Mo_2N **b.** CrH_2

43. **a.** zinc **b.** copper **c.** manganese

45. −19.4 kJ/mol

47. When Cr is added to steel it reacts with oxygen in steel to prevent it from rusting. A Cr steel alloy would be used in any situation where the steel might be easily oxidized, such as when it comes in contact with water.

49. rutile: 33.3% Ti by moles, 59.9% Ti by mass
ilmenite: 20.0% Ti by moles, 31.6% Ti by mass

51. Titanium must be arc-melted in an inert atmosphere because the high temperature and flow of electrons would cause the metal to oxidize in a normal atmosphere.

53. TiO_2 is the most important industrial product of titanium and it is often used as a pigment in white paint.

55. The Bayer process is a hydrometallurgical process used to separate Al_2O_3 from other oxides. The Al_2O_3 is selectively dissolved by hot, concentrated NaOH. The other oxides are removed as solids and the Al_2O_3 precipitates out of solution when the solution is neutralized.

57. cobalt and tungsten

59. 3.3 kg Fe, 2.0 kg Ti

61. Four atoms surround a tetrahedral hole and six atoms surround an octahedral hole. The octahedral hole is larger because it is surrounded by a greater number of atoms.

63. Mn has one more electron orbital available for bonding than does chromium.

65. Ferromagnetic atoms, like paramagnetic ones, have unpaired electrons. However, in ferromagnetic atoms, these electrons align with their spin oriented in the same direction, resulting in a permanent magnetic field.

67. The nuclear charge of the last three is relatively high because of the lanthanide series in which the 4*f* subshell falls between them and the other six metals of the group.

69. **a.** 16.0 cm **b.** 4.95 cm **c.** 14%

71. 92%

73. 5.4×10^7

75. First, roast to form the oxide.

$$4\,CoAsS(s) + 9\,O_2(g) \longrightarrow 4\,CoO(s) + 4\,SO_2(g) + As_4O_6(s)$$

Then reduce the oxide with coke.

$$CoO(s) + C(s) \longrightarrow Co(s) + CO(g)$$

The oxides of arsenic are relatively volatile and can be separated, but they are poisonous.

77. Au and Ag are found in elemental form because of their low reactivity. Na and Ca are group 1 and group 2 metals, respectively, and are highly reactive as they readily lose their valence electrons to obtain octets.

Chapter 24

17. **a.** $[Ar]\,4s^23d^8$, $[Ar]\,3d^8$
b. $[Ar]\,4s^23d^5$, $[Ar]\,3d^3$
c. $[Kr]\,5s^24d^1$, $[Kr]\,5s^14d^1$
d. $[Xe]\,6s^24f^{14}5d^3$, $[Xe]\,4f^{14}5d^3$

19. **a.** +5 **b.** +7 **c.** +4

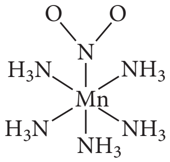
21. **a.** +3, 6 **b.** +2, 6 **c.** +2, 4 **d.** +1, 2

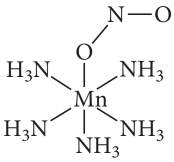
23. **a.** hexaaquachromium(III)
b. tetracyanocuprate(II)
c. pentaaminebromoiron(III) sulfate
d. aminetetraaquahydroxycobalt(III) chloride

25. **a.** $[Cr(NH_3)_6]^{3+}$
b. $K_3[Fe(CN)_6]$
c. $[Cu(en)(SCN)_2]$
d. $[Pt(H_2O)_4][PtCl_6]$

27. **a.** $[Co(NH_3)_3(CN)_3]$, triaminetricyanocobalt(III)
b. $[Cr(en)_3]^{3+}$, tris(ethylenediamine)chromium(III)

29.



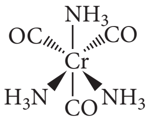


31. $[Fe(H_2O)_5Cl]Cl \cdot H_2O$, pentaaquachloroiron(II) chloride monohydrate
$[Fe(H_2O)_4Cl_2] \cdot 2\,H_2O$, tetraaquadichloroiron(II) dihydrate

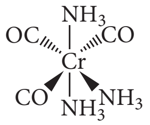
33. **b, c, e**

35. **a.** 3 **b.** No geometric isomers.

37. **a.**

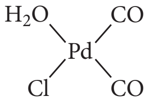


Fac

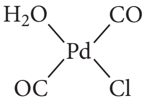


Mer

b.



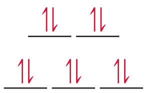
Cis



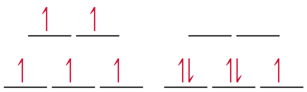
Trans

39. *cis* isomer is optically active

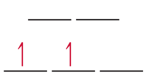
41. **a.**



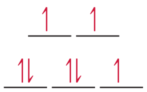
b.



c.



d.



43. 163 kJ/mol

45. $[Co(CN)_6]^{3-} \longrightarrow$ 290 nm, colorless
$[Co(NH_3)_6]^{3+} \longrightarrow$ 440 nm, yellow
$[CoF_6]^{3-} \longrightarrow$ 770 nm, green

47. weak

49. **a.** 4
b. 3
c. 1

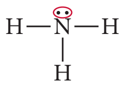
51. 3

53. porphyrin

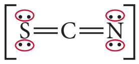
55. Water is a weak field ligand that forms a high-spin complex with hemoglobin. Because deoxyhemoglobin is weak field it absorbs large wavelength light and appears blue. Oxyhemoglobin is a low-spin complex and absorbs small wavelength light, so O_2 must be a strong field ligand.

57. **a.** [Ar] $4s^1 3d^5$, [Ar] $3d^5$, [Ar] $3d^4$, [Ar] $3d^3$
b. [Ar] $4s^1 3d^{10}$, [Ar] $3d^{10}$, [Ar] $3d^9$

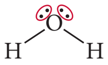
59. **a.**



b.

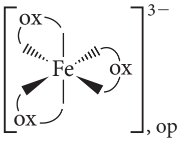


c.



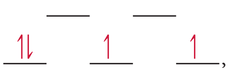
61. $[MA_2B_2C_2]$ all cis; A trans and B and C cis; B trans and A and C cis; C trans and A and B cis; all trans.
$[MA_2B_3C]$ will have fac–mer isomers.
$[MAB_2C_3]$ will have fac–mer isomers.
$[MAB_3C_2]$ will have fac–mer isomers.
$[MA_3B_2C]$ will have fac–mer isomers.
$[MA_2BC_3]$ will have fac–mer isomers.
$[MA_3BC_2]$ will have fac–mer isomers.
$[MABC_2]$ will have AB cis–trans isomers.
$[MAB_4C]$ will have AC cis–trans isomers.
$[MA_4BC]$ will have BC cis–trans isomers.
$[MABC_4]$ will have AB cis–trans isomers.

63.



, optical isomers

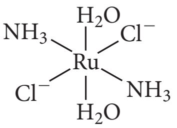
65.



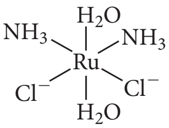
, paramagnetic

67.

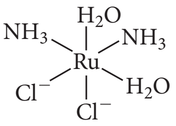
1.



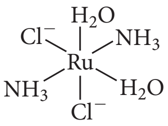
2.



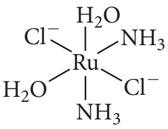
3.



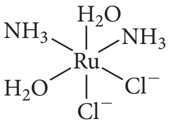
4.



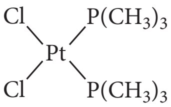
5.



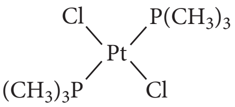
Only structure 3. is chiral. This is its mirror image.



69.



cis-dichlorobis (trimethyl phosphine) platinum(II)



trans-dichlorobis (trimethyl phosphine) platinum(II)

71. ____ d_{z^2}
____ ____ $d_{x^2-y^2}$ and d_{xy}
____ ____ d_{xz} and d_{yz}

73. **a.** 2×10^{-8} M
b. 6.6×10^{-3} M
c. NiS will dissolve more easily in the ammonia solution because the formation of the complex ion is favorable, removing Ni^{2+} ions from the solution and allowing more NiS to dissolve.

75. Prepare a solution that contains both $[MCl_6]^{3-}$ and $[MBr_6]^{3-}$ and see if any complex ions that contain both Cl and Br form. If they do it would demonstrate that these complexes are labile.

77. pH = 10.1

79. Au

Appendix IV:

Answers to In-Chapter Practice Problems

Chapter 1

1.1. **a.** The composition of the copper is not changing; thus, being hammered flat is a physical change that signifies a physical property.

b. The dissolution and color change of the nickel indicate that it is undergoing a chemical change and exhibiting a chemical property.

c. Sublimation is a physical change indicative of a physical property.

d. When a match ignites, a chemical change begins as the match reacts with oxygen to form carbon dioxide and water. Flammability is a chemical property.

1.2. **a.** 29.8 °C

b. 302.9 K

1.3. 21.4 g/cm^3. This matches the density of platinum.

1.3. *For More Practice* 4.50 g/cm^3 The metal is titanium.

1.4. The thermometer shown has markings every 1 °F; thus, the first digit of uncertainty is 0.1. The answer is 103.1 °F.

1.5. **a.** Each figure in this number is significant by rule 1: three significant figures.

b. This is a defined quantity that has an unlimited number of significant figures.

c. Both 1's are significant (rule 1) and the interior zero is significant as well (rule 2): three significant figures.

d. Only the two 9's are significant, the leading zeroes are not (rule 3): two significant figures.

e. There are five significant figures because the 1, 4, and 5 are nonzero (rule 1) and the trailing zeroes are after a decimal point so they are significant as well (rule 4).

f. The number of significant figures is ambiguous because the trailing zeroes occur before an implied decimal point (rule 4). Assume two significant figures.

1.6. **a.** 0.381

b. 121.0

c. 1.174

d. 8

1.7. 3.15 yd

1.8. 2.446 gal

1.9. 1.61×10^6 cm^3

1.9. *For More Practice* 3.23×10^3 kg

1.10. 1.03 kg

1.10. *For More Practice* 2.9×10^{-2} cm^3

1.11. 0.855 cm

1.12. 2.70 g/cm^3

Chapter 2

2.1. For the first sample:

$$\frac{\text{mass of oxygen}}{\text{mass of carbon}} = \frac{17.2\text{ g O}}{12.9\text{ g C}} = 1.33 \text{ or } 1.33 : 1$$

For the second sample:

$$\frac{\text{mass of oxygen}}{\text{mass of carbon}} = \frac{10.5\text{ g O}}{7.88\text{ g C}} = 1.33 \text{ or } 1.33 : 1$$

The ratios of oxygen to carbon are the same in the two samples of carbon monoxide, so these results are consistent with the law of definite proportions.

2.2.

$$\frac{\text{mass of hydrogen to 1 g of oxygen in hydrogen peroxide}}{\text{mass of hydrogen to 1 g of oxygen in water}}$$

$$= \frac{0.250}{0.125} = 2.00$$

The ratio of the mass of hydrogen from one compound to the mass of hydrogen in the other is equal to 2. This is a simple whole number and therefore consistent with the law of multiple proportions.

2.3. **a.** $Z = 6$, $A = 13$, $^{13}_{6}\text{C}$

b. 19 protons, 20 neutrons

2.4. **a.** N^{3-} **b.** Rb^+

2.5. 24.31 amu

2.5. *For More Practice* 70.92 amu

2.6. 4.65×10^{-2} mol Ag

2.7. 0.563 mol Cu

2.7. *For More Practice* 22.6 g Ti

2.8. 1.3×10^{22} C atoms

2.8. *For More Practice* 6.87 g W

2.9. $l = 1.72$ cm

2.9. *For More Practice* 2.90×10^{24} Cu atoms

Chapter 3

3.1. **a.** C_5H_{12}

b. HgCl

c. CH_2O

3.2. **a.** molecular element

b. molecular compound

c. atomic element

d. ionic compound

e. ionic compound

3.3. K_2S

3.4. AlN

3.5. silver nitride

3.5. *For More Practice* Rb_2S

3.6. iron(II) sulfide
3.6. *For More Practice* RuO_2
3.7. tin(II) chlorate
3.7. *For More Practice* $Co_3(PO_4)_2$
3.8. dinitrogen pentoxide
3.8. *For More Practice* PBr_3
3.9. hydrofluoric acid
3.10. nitrous acid
3.10. *For More Practice* $HClO_4$
3.11. sulfurous acid
3.12. 164.10 amu
3.13. 5.839×10^{20} $C_{13}H_{18}O_2$ molecules
3.13. *For More Practice* 1.06 g H_2O
3.14. 53.29%
3.14. *For More Practice* 74.19% Na
3.15. 83.9 g Fe_2O_3
3.15. *For More Practice* 8.6 g Na
3.16. 4.0 g O
3.16. *For More Practice* 3.60 g C
3.17. CH_2O
3.18. $C_{13}H_{18}O_2$
3.19. C_6H_6
3.19. *For More Practice* $C_2H_8N_2$
3.20. C_2H_5
3.21. C_2H_4O
3.22. $SiO_2(s) + 3\,C(s) \longrightarrow SiC(s) + 2\,CO(g)$
3.23. $2\,C_2H_6(g) + 7\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$
3.24. $Pb(NO_3)_2(aq) + 2\,KCl(aq) \longrightarrow PbCl_2(s) + 2\,KNO_3(aq)$

Chapter 4

4.1. 4.08 g HCl
4.2. 22 kg HNO_3
4.3. H_2 is the limiting reagent, since it produces the least amount of NH_3. Therefore, 29.4 kg NH_3 is the theoretical yield.
4.4. CO is the limiting reagent, since it only produces 114 g Fe. Therefore, 114 g Fe is the theoretical yield: percentage yield = 63.4% yield
4.5. 0.214 M $NaNO_3$
4.5. *For More Practice* 44.6 g KBr
4.6. 402 g $C_{12}H_{22}O_{11}$
4.6. *For More Practice* 221 mL of KCl solution
4.7. 667 mL
4.7. *For More Practice* 0.105 L
4.8. 51.4 mL HNO_3 solution
4.8. *For More Practice* 0.170 g CO_2
4.9. **a.** insoluble
b. insoluble
c. soluble
d. soluble
4.10. $NH_4Cl(aq) + Fe(NO_3)_3(aq) \longrightarrow$ NO REACTION
4.11. $2\,NaOH(aq) + CuBr_2(aq) \longrightarrow Cu(OH)_2(s) + 2\,NaBr(aq)$
4.12. $2\,H^+(aq) + 2\,I^-(aq) + Ba^{2+}(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l) + Ba^{2+}(aq) + 2\,I^-(aq)$
$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
4.13. $H_2SO_4(aq) + 2\,LiOH(aq) \longrightarrow 2\,H_2O(l) + Li_2SO_4(aq)$
$H^+(aq) + OH^-(aq) \longrightarrow H_2O(aq)$
4.14. 9.03×10^{-2} M H_2SO_4
4.14. *For More Practice* 24.5 mL NaOH solution
4.15. $2\,HBr(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\,KBr(aq)$
4.15. *For More Practice* $2\,H^+(aq) + S^{2-}(aq) \longrightarrow H_2S(g)$
4.16. **a.** $Cr = 0$
b. $Cr^{3+} = +3$
c. $Cl^- = -1, C = +4$
d. $Br = -1$, $Sr = +2$
e. $O = -2$, $S = +6$
f. $O = -2$, $N = +5$
4.17. Sn is oxidized and N is reduced.
4.17. *For More Practice* **b.** Reaction b is the only redeox reaction. Al is oxidized and O is reduced.
4.18. **a.** This is a redox reaction in which Li is the reducing agent (it is oxidized) and Cl_2 is the oxidizing reagent (it is reduced).
b. This is a redox reaction in which Al is the reducing agent and Sn^{2+} is the oxidizing agent.
c. This is not a redox reaction because no oxidation states change.
d. This is a redox reaction in which C is the reducing agent and O_2 is the oxidizing agent.
4.19. $2\,C_2H_5SH(l) + 9\,O_2(g) \longrightarrow 4\,CO_2(g) + 2\,SO_2(g) + 6\,H_2O(g)$

Chapter 5

5.1. 15.0 psi
5.1. *For More Practice* 80.6 kPa
5.2. 2.1 atm at a depth of approximately 11 m.
5.3. 123 mL
5.4. 11.3 L
5.5. 1.63 atm, 23.9 psi
5.6. 16.1 L
5.6. *For More Practice* 976 mmHg
5.7. $d = 4.91$ g/L
5.7. *For More Practice* 44.0 g/mol
5.8. 70.7 g/mol
5.9. 0.0610 mol H_2
5.10. 4.2 atm
5.11. 12.0 mg H_2
5.12. 82.3 g Ag_2O
5.12. *For More Practice* 7.10 g Ag_2O
5.13. 6.53 L O_2
5.14. $u_{\text{rms}} = 238$ m/s
5.15. $\dfrac{\text{rate}_{H_2}}{\text{rate}_{Kr}} = 6.44$

Chapter 6

6.1. $\Delta E = 71\ \text{J}$

6.2. $C_s = 0.38 \dfrac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}$

The specific heat capacity of gold is 0.128 J/g · °C; therefore, the rock cannot be pure gold.

6.2. *For More Practice* $T_f = 42.1\ {}^\circ\text{C}$

6.3. 37.8 grams Cu

6.4. $-122\ \text{J}$

6.4. *For More Practice* $\Delta E = -998\ \text{J}$

6.5. $\Delta E_{\text{reaction}} = -3.91 \times 10^3\ \text{kJ/mol}\ C_6H_{14}$

6.5. *For More Practice* $C_{\text{cal}} = 4.55 \dfrac{\text{kJ}}{{}^\circ\text{C}}$

6.6. **a.** endothermic, positive ΔH
b. endothermic, positive ΔH
c. exothermic, negative ΔH

6.7. $-2.06 \times 10^3\ \text{kJ}$

6.7. *For More Practice* 33 g C_4H_{10}
99 g CO_2

6.8. $\Delta H_{\text{rxn}} = -68\ \text{kJ}$

6.9. $N_2O(g) + NO_2(g) \longrightarrow 3\ NO(g),\ \Delta H_{\text{rxn}} = +157.6\ \text{kJ}$

6.9. *For More Practice*
$3\ H_2(g) + O_3(g) \longrightarrow 3\ H_2O(g),\ \Delta H_{\text{rxn}} = -868.1\ \text{kJ}$

6.10. **a.** $Na(s) + \frac{1}{2}Cl_2(g) \longrightarrow NaCl(s),\ \Delta H^\circ_f = -411.2\ \text{kJ/mol}$
b. $Pb(s) + N_2(g) + 3\ O_2(g) \longrightarrow Pb(NO_3)_2(s),\ \Delta H^\circ_f = -451.9\ \text{kJ/mol}$

6.11. $\Delta H^\circ_{\text{rxn}} = -851.5\ \text{kJ}$

6.12. $\Delta H^\circ_{\text{rxn}} = -1648.4\ \text{kJ}$
111 kJ emitted (−111 kJ)

6.13. $1.2 \times 10^2\ \text{kg}\ CO_2$

Chapter 7

7.1. $5.83 \times 10^{14}\ \text{s}^{-1}$

7.2. 2.64×10^{20} photons

7.2. *For More Practice* 435 nm

7.3. **a.** blue < green < red
b. red < green < blue
c. red < green < blue

7.4. $6.1 \times 10^6\ \text{m/s}$

7.5. For the 5*d* orbitals:
$n = 5$
$l = 2$
$m_l = -2, -1, 0, 1, 2$
The 5 integer values for m_l signify that there are five 5*d* orbitals.

7.6. **a.** l cannot equal 3 if $n = 3$. $l = 2$
b. m_l cannot equal -2 if $l = -1$. Possible values for $m_l = -1, 0$, or 1
c. l cannot be 1 if $n = 1$. $l = 0$

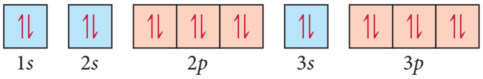
7.7. 397 nm

7.7. *For More Practice* $n = 1$

Chapter 8

8.1. **a.** Cl $1s^22s^22p^63s^23p^5$ or [Ne] $3s^23p^5$
b. Si $1s^22s^22p^63s^23p^2$ or [Ne] $3s^23p^2$
c. Sr $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$ or [Kr] $5s^2$
d. O $1s^22s^22p^4$ or [He] $2s^22p^4$

8.2. There are no unpaired electrons.



8.3. $1s^22s^22p^63s^23p^3$ or [Ne] $3s^23p^3$. The 5 electrons in the $3s^23p^3$ orbitals are the valence electrons, while the 10 electrons in the $1s^22s^22p^6$ orbitals belong to the core.

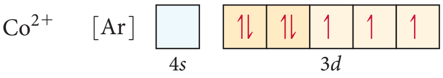
8.4. [Xe] $6s^24f^{14}5d^{10}6p^3$

8.4. *For More Practice* [Kr] $5s^24d^{10}5p^5$

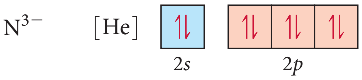
8.5. **a.** Sn
b. cannot predict
c. W
d. Se

8.5. *For More Practice* Rb > Ca > Si > S > F

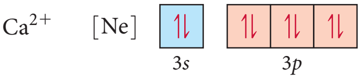
8.6. **a.** [Ar] $4s^03d^7$. Co^{2+} is paramagnetic.



b. [He] $2s^22p^6$. N^{3-} is diamagnetic.



c. [Ne] $3s^23p^6$. Ca^{2+} is diamagnetic.



8.7. **a.** K **b.** F^- **c.** Cl^-

8.7. *For More Practice* $Cl^- > Ar > Ca^{2+}$

8.8. **a.** I **b.** Ca
c. cannot predict **d.** F

8.8. *For More Practice* F > S > Si > Ca > Rb

8.9. **a.** Sn
b. cannot predict based on simple trends (Po is larger)
c. Bi
d. B

8.9. *For More Practice* Cl < Si < Na < Rb

8.10. **a.** $2\ Al(s) + 3\ Cl_2(g) \longrightarrow 2\ AlCl_3(s)$
b. $2\ Li(s) + 2\ H_2O(l) \longrightarrow 2\ Li^+(aq) + 2\ OH^-(aq) + H_2(g)$
c. $H_2(g) + Br_2(l) \longrightarrow 2\ HBr(g)$

Chapter 9

9.1. Mg_3N_2

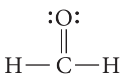
9.2. KI < LiBr < CaO

9.2. *For More Practice* $MgCl_2$

9.3. **a.** pure covalent
b. ionic
c. polar covalent

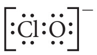
9.4. $:C \equiv O:$

9.5.



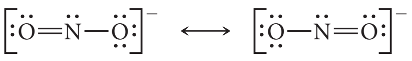
$$H-\overset{\overset{\Large :\ddot{O}:}{\|}}{C}-H$$

9.6.



$$[:\ddot{Cl}:\ddot{O}:]^-$$

9.7.



$$[:\ddot{O}=\ddot{N}-\ddot{O}:]^- \longleftrightarrow [:\ddot{O}-\ddot{N}=\ddot{O}:]^-$$

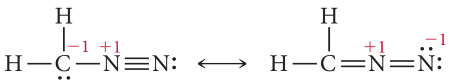
9.8.

Structure	**A**			**B**			**C**		
	$:\ddot{N}=N=\ddot{O}:$			$:N\equiv N-\ddot{O}:$			$:\ddot{N}-N\equiv O:$		
number of valence e^-	5	5	6	5	5	6	5	5	6
number of nonbonding e^-	−4	−0	−4	−2	−0	−6	−6	−0	−2
1/2(number of bonding e^-)	−2	−4	−2	−3	−4	−1	−1	−4	−3
Formal charge	**−1**	**+1**	**0**	**0**	**+1**	**−1**	**−2**	**+1**	**+1**

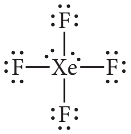
Structure B contributes the most to the correct overall structure of N_2O.

9.8. *For More Practice* The nitrogen is +1, the singly bonded oxygen atoms are −1, and the double-bonded oxygen atom has no formal charge.

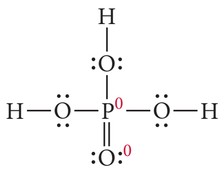
9.9.



9.10.



9.10. *For More Practice*



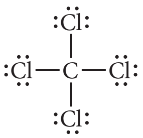
9.11. $CH_3OH(g) + \frac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$

$\Delta H_{rxn} = -641$ kJ

9.11. *For More Practice* $\Delta H_{rxn} = -8.0 \times 10^1$ kJ

Chapter 10

10.1. tetrahedral



10.2. bent

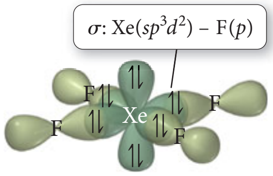
10.3. linear

10.4.

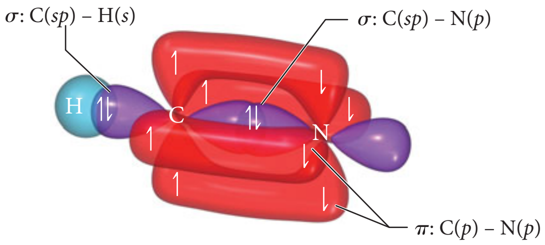
Atom	**Number of Electron Groups**	**Number of Lone Pairs**	**Molecular Geometry**
Carbon (left)	4	0	Tetrahedral
Carbon (right)	3	0	Trigonal planar
Oxygen	4	2	Bent

10.5. The molecule is nonpolar.

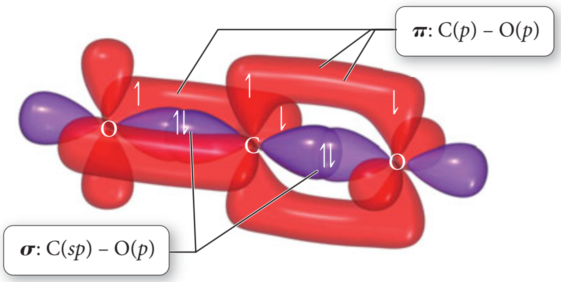
10.6. The xenon atom has six electron groups and therefore has an octahedral electron geometry. An octahedral electron geometry corresponds to sp^3d^2 hybridization (refer to Table 10.3).



10.7. Since there are only two electron groups around the central atom (C), the electron geometry is linear. According to Table 10.3, the corresponding hybridization on the carbon atom is sp.



10.8. Since there are only two electron groups about the central atom (C), the electron geometry is linear. The hybridization on C is sp (refer to Table 10.3).

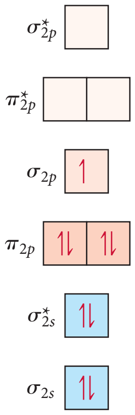


10.8. *For More Practice* There are five electron groups about the central atom (I); therefore, the electron geometry is trigonal bipyramidal and the corresponding hybridization of I is sp^3d (refer to Table 10.3).

10.9. H_2^+ bond order $= +\frac{1}{2}$

Since the bond order is positive, the H_2^+ ion should be stable; however, the bond order of H_2^+ is lower than the bond order of H_2 (bond order = 1). Therefore, the bond in H_2^+ is weaker than in H_2.

10.10. The bond order of N_2^+ is 2.5, which is lower than that of the N_2 molecule (bond order =3); therefore, the bond is weaker. The MO diagram shows that the N_2^+ ion has one unpaired electron and is therefore paramagnetic.



10.10. *For More Practice* The bond order of Ne_2 is 0, which indicates that dineon does not exist.

10.11. The bond order of NO is +2.5. The MO diagram shows that the ion has one unpaired electron and is therefore paramagnetic.

Chapter 11

11.1. **b, c**

11.2. HF has a higher boiling point than HCl because, unlike HCl, HF is able to form hydrogen bonds. The hydrogen bond is the strongest of the intermolecular forces and requires more energy to break.

11.3. 5.83×10^3 kJ

11.3. *For More Practice* 49 °C

11.4. 33.8 KJ/mol

11.5. 7.04×10^3 torr

11.6. 29.4°

11.7. $7.18 \frac{\text{g}}{\text{cm}^3}$

Chapter 12

12.1. **a.** not soluble
 b. soluble
 c. not soluble
 d. not soluble

12.2. 2.7×10^{-4} M

12.3. 42.5 g $C_{12}H_{22}O_{11}$

12.3. *For More Practice* 3.3×10^4 L

12.4. **a.** $M = 0.415$ M
 b. $m = 0.443$ m
 c. % by mass = 13.2%
 d. $\chi C_{12}H_{22}O_{11} = 0.00793$
 e. mole percent = 0.793%

12.5. 0.600 M

12.5. *For More Practice* 0.651 *m*

12.6. 22.5 torr

12.6. *For More Practice* 0.144

12.7. **a.** $P_{benzene} = 26.6$ torr
 $P_{toluene} = 20.4$ torr
 b. 47.0 torr
 c. 52.5% benzene; 47.5% toluene
 The vapor will be richer in the more volatile component, which in this case is benzene.

12.8. $T_f = -4.8$ °C

12.9. 101.84 °C

12.10. 11.8 atm

12.11. −0.60 °C

12.12. 0.014 mol NaCl

Chapter 13

13.1. $\frac{\Delta[H_2O_2]}{\Delta t} = -4.40 \times 10^{-3}$ M/s

$\frac{\Delta[I_3^-]}{\Delta t} = 4.40 \times 10^{-3}$ M/s

13.2. **a.** Rate $= k[CHCl_3][Cl_2]^{1/2}$. (Fractional-order reactions are not common but are occasionally observed.)
 b. $3.5\ M^{-1/2} \cdot s^{-1}$

13.3. 5.78×10^{-2} M

13.4. 0.0277 M

13.5. 1.64×10^{-3} M

13.6. 79.2 s

13.7. $2.07 \times 10^{-5} \frac{\text{L}}{\text{mol} \cdot \text{s}}$

13.8. $6.13 \times 10^{-4} \frac{\text{L}}{\text{mol} \cdot \text{s}}$

13.9. $2\,A + B \longrightarrow A_2B$
 Rate $= k[A]^2$

Chapter 14

14.1. $K = \frac{[CO_2]^3[H_2O]^4}{[C_3H_8][O_2]^5}$

14.2. 2.1×10^{-13}

14.2. *For More Practice* 1.4×10^2

14.3. 6.2×10^2

14.4. $K_c = \frac{[Cl_2]^2}{[HCl]^4[O_2]}$

14.5. 9.4

14.6. 1.1×10^{-6}

14.7. $Q_c = 0.0196$
 Reaction proceeds to the left.

14.8. 0.033 M

14.9. $[N_2] = 4.45 \times 10^{-3}$ M
 $[O_2] = 4.45 \times 10^{-3}$ M
 $[NO] = 1.1 \times 10^{-3}$ M

14.10. $[N_2O_4] = 0.005$ M
 $[NO_2] = 0.041$ M

14.11. $P_{I_2} = 0.0027$ atm
$P_{Cl_2} = 0.0027$ atm
$P_{ICl_2} = 0.246$ atm

14.12. 1.67×10^{-7} M

14.13. 6.78×10^{-6} M

14.14. Adding Br_2 increases the concentration of Br_2, causing a shift to the left (away from the Br_2). Adding BrNO increases the concentration of BrNO, causing a shift to the right.

14.15. Decreasing the volume causes the reaction to shift right. Increasing the volume causes the reaction to shift left.

14.16. If we increase the temperature, the reaction shifts to the left. If we decrease the temperature, the reaction shifts to the right.

Chapter 15

15.1. **a.** H_2O donates a proton to C_5H_5N, making it the acid. The conjugate base is therefore OH^-. Since C_5H_5N accepts the proton, it is the base and becomes the conjugate acid $C_5H_5NH^+$.
b. Since HNO_3 donates a proton to H_2O, it is the acid, making NO_3^- the conjugate base. Since H_2O is the proton acceptor, it is the base and becomes the conjugate acid, H_3O^+.

15.2. **a.** Since $[H_3O^+] < [OH^-]$ the solution is basic.
b. $[H_3O^+] = 1.0 \times 10^{-7}$ M
Neutral solution.
c. $[H_3O^+] = 1.2 \times 10^{-5}$ M
Since $[H_3O^+] > [OH^-]$ the solution is acidic.

15.3. **a.** 8.02 (basic)
b. 11.85 (basic)

15.4. 4.3×10^{-9} M

15.5. 9.4×10^{-3} M

15.6. 3.28

15.7. 2.72

15.8. 1.8×10^{-6}

15.9. 0.85%

15.10. 4.0×10^{-7} M

15.11. $[OH^-] = 0.020$ M
pH = 12.30

15.12. $[OH^-] = 1.2 \times 10^{-2}$ M
pH = 12.08

15.13. **a.** weak base **b.** pH-neutral

15.14. 9.07

15.15. **a.** pH-neutral **b.** weak acid
c. weak acid

15.16. **a.** basic **b.** acidic
c. pH-neutral **d.** acidic

15.17. 3.83

15.18. $[SO_4^{2-}] = 0.00386$ M
pH = 1.945

15.19. 5.6×10^{-11} M

Chapter 16

16.1. 4.44

16.1. *For More Practice* 3.44

16.2. 9.14

16.3. 4.87

16.3. *For More Practice* 4.65

16.4. 9.68

16.4. *For More Practice* 9.56

16.5. hypochlorous acid (HClO); 2.4 g NaClO

16.6. 1.74

16.7. 18.08

16.8. 2.30×10^{-6} M

16.9. 5.3×10^{-13}

16.10. 1.21×10^{-5} M

16.11. $FeCO_3$ will be more soluble in an acidic solution than $PbBr_2$ because the CO_3^{2-} ion is a basic anion, whereas Br^- is the conjugate base of a strong acid (HBr) and is therefore pH-neutral.

16.12. $Q > K_{sp}$; therefore, a precipitate forms.

16.13. 2.9×10^{-6} M

16.14. **a.** AgCl precipitates first; $[NaCl] = 7.1 \times 10^{-9}$ M
b. $[Ag^+]$ is 1.5×10^{-8} M when $PbCl_2$ begins to precipitate, and $[Pb^{2+}]$ is 0.085 M.

16.15. 9.6×10^{-6} M

Chapter 17

17.1. **a.** positive
b. negative
c. positive

17.2. **a.** −548 J/K
b. ΔS_{sys} is negative.
c. ΔS_{univ} is negative, and the reaction is not spontaneous.

17.2. *For More Practice* 375 K

17.3. $\Delta G = -101.6 \times 10^3$ J
Therefore, the reaction is spontaneous. Since both ΔH and ΔS are negative, as the temperature increases ΔG will become more positive.

17.4. −153.2 J/K

17.5. $\Delta G^\circ_{rxn} = -36.3$ kJ
Since ΔG°_{rxn} is negative, the reaction is spontaneous at this temperature.

17.6. $\Delta G^\circ_{rxn} = -42.1$ kJ
Since the value of ΔG°_{rxn} at the lowered temperature is more negative (or less positive) (which is −36.3 kJ), the reaction is more spontaneous.

17.7. $\Delta G^\circ_{rxn} = -689.6$ kJ
Since ΔG°_{rxn} is negative, the reaction is spontaneous at this temperature.

17.7. *For More Practice* $\Delta G^\circ_{rxn} = -689.7$ kJ (at 25°)
The value calculated for ΔG°_{rxn} from the tabulated values (−689.6 kJ) is the same, to within 1 in the

least significant digit, as the value calculated using the equation for $\Delta G^\circ_{\text{rxn}}$.

$\Delta G^\circ_{\text{rxn}} = -649.7 \text{ kJ (at 500.0 K)}$

You could not calculate $\Delta G^\circ_{\text{rxn}}$ at 500.0 K using tabulated $\Delta G^\circ_{\text{f}}$ values because the tabulated values of free energy are calculated at a standard temperature of 298 K, much lower than 500 K.

17.8. $+107.1 \text{ kJ}$

17.9. $\Delta G_{\text{rxn}} = -129 \text{ kJ}$

The reaction is more spontaneous under these conditions than under standard conditions because ΔG_{rxn} is more negative than $\Delta G^\circ_{\text{rxn}}$.

17.10. -10.9 kJ

Chapter 18

18.1. $2\,\text{Cr}(s) + 4\,\text{H}^+(aq) \longrightarrow 2\,\text{Cr}^{2+}(aq) + 2\,\text{H}_2(g)$

18.2. $\text{Cu}(s) + 4\,\text{H}^+(aq) + 2\,\text{NO}_3^-(aq) \longrightarrow \text{Cu}^{2+}(aq) + 2\,\text{NO}_2(g) + 2\,\text{H}_2\text{O}(l)$

18.3. $3\,\text{ClO}^-(aq) + 2\,\text{Cr(OH)}_4^-(aq) + 2\,\text{OH}^-(aq) \longrightarrow 3\,\text{Cl}^-(aq) + 2\,\text{CrO}_4^{2-}(aq) + 5\,\text{H}_2\text{O}(l)$

18.4. $+0.60 \text{ V}$

18.5. **a.** The reaction *will* be spontaneous as written.
b. The reaction *will not* be spontaneous as written.

18.6. $\Delta G^\circ = -3.63 \times 10^5 \text{ J}$

Since ΔG° is negative, the reaction is spontaneous.

18.7. 4.5×10^3

18.8. 1.08 V

18.9. *Anode:* $2\,\text{H}_2\text{O}(l) \longrightarrow \text{O}_2(g) + 4\,\text{H}^+(aq) + 4\,\text{e}^-$
Cathode: $2\,\text{H}_2\text{O}(l) + 2\,\text{e}^- \longrightarrow \text{H}_2(g) + 2\,\text{OH}^-(aq)$

18.10. 6.0×10^1 min

Chapter 19

19.1. ${}^{216}_{84}\text{Po} \longrightarrow {}^{212}_{82}\text{Pb} + {}^{4}_{2}\text{He}$

19.2. **a.** ${}^{235}_{92}\text{U} \longrightarrow {}^{231}_{90}\text{Th} + {}^{4}_{2}\text{He}$
${}^{231}_{90}\text{Th} \longrightarrow {}^{231}_{91}\text{Pa} + {}^{0}_{-1}\text{e}$
${}^{231}_{91}\text{Pa} \longrightarrow {}^{227}_{89}\text{Ac} + {}^{4}_{2}\text{He}$
b. ${}^{22}_{11}\text{Na} \longrightarrow {}^{22}_{10}\text{Ne} + {}^{0}_{+1}\text{e}$
c. ${}^{76}_{36}\text{Kr} + {}^{0}_{-1}\text{e} \longrightarrow {}^{76}_{35}\text{Br}$

19.2. *For More Practice* Positron emission
$({}^{40}_{19}\text{K} \longrightarrow {}^{40}_{18}\text{Ar} + {}^{0}_{+1}\text{e})$ or electron capture
$({}^{40}_{19}\text{K} + {}^{0}_{+1}\text{e} \longrightarrow {}^{40}_{18}\text{Ar})$

19.3. **a.** positron emission
b. beta decay
c. positron emission

19.4. 10.7 yr

19.5. $t = 964$ yr

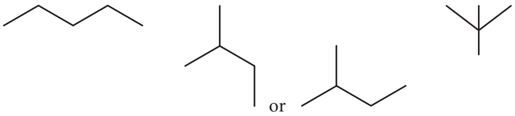
No, the C-14 content suggests that the scroll is from about A.D. 1000, not 500 B.C.

19.6. 1.0×10^9 yr

19.7. Mass defect $= 1.934$ amu
Nuclear binding energy $= 7.569$ MeV/nucleon

Chapter 20

20.1.



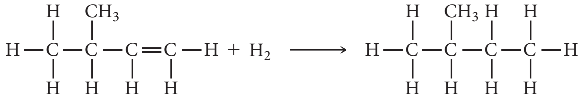
20.2. 3-methylhexane

20.3. 3,5-dimethylheptane

20.4. 2,3,5-trimethylhexane

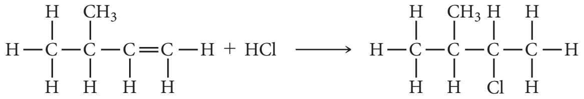
20.5. **a.** 4,4-dimethyl-2-pentyne
b. 3-ethyl-4,6-dimethyl-1-heptene

20.6. **a.** 2-methylbutane



$$\begin{array}{ccccccccccc} & & H & & CH_3 & & & & & & \\ & & | & & | & & & & & & \\ H & - & C & - & C & - & C & = & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & \end{array} + H_2 \longrightarrow \begin{array}{ccccccccccc} & & H & & CH_3 & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & \end{array}$$

b. 2-chloro-3-methylbutane

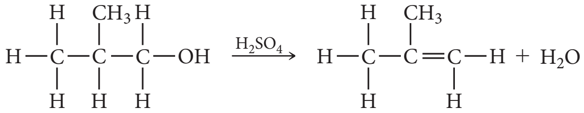


$$\begin{array}{ccccccccccc} & & H & & CH_3 & & & & & & \\ & & | & & | & & & & & & \\ H & - & C & - & C & - & C & = & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & \end{array} + HCl \longrightarrow \begin{array}{ccccccccccc} & & H & & CH_3 & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & Cl & & H & & \end{array}$$

20.7. **a.** Alcohol reacting with an active metal.

$$CH_3CH_2OH + Na \longrightarrow CH_3CH_2ONa + \frac{1}{2}H_2$$

b. dehydration reaction

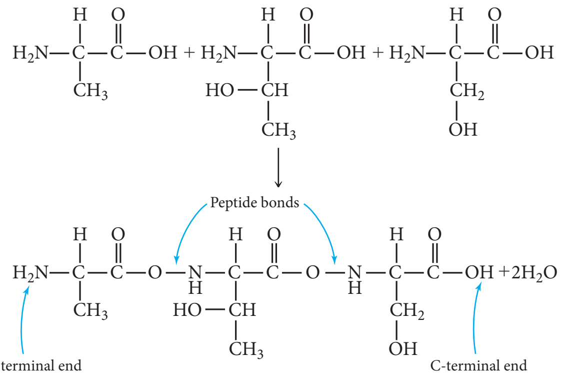


$$\begin{array}{ccccccccc} & & H & & CH_3 & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & OH \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array} \xrightarrow{H_2SO_4} \begin{array}{ccccccccc} & & H & & CH_3 & & & & \\ & & | & & | & & & & \\ H & - & C & - & C & = & C & - & H \\ & & | & & & & | & & \\ & & H & & & & H & & \end{array} + H_2O$$

Chapter 21

21.1. Fructose exhibits optical isomerism. It contains three chiral carbons.

21.2.



Chapter 22

22.1. $KAlSi_3O_8$

22.2. $x = 2$

22.3. Orthosilicate (or neosilicate): each of the two Be ions has a charge of 2+ for a total of 4+, and the SiO_4 unit has a charge of 4−.

22.4. Inosilicate (or pyroxene): Ca and Mg each have a charge of 2+ for a total of 4+, and the Si_2O_6 unit has a charge of 4− (two SiO_3^{2-} units).

22.5. $2\,H_2S(g) + 3\,O_2(g) \longrightarrow 2\,H_2O(g) + SO_2(g)$

S changes from the −2 to +4 oxidation state.

22.6. The oxidation state for Cl is +7 in ClO_4^- and −1 in Cl^-.

22.7. $I_2(s) + 5\,F_2(g) \longrightarrow 2\,IF_5(g)$

22.8. The electron geometry is tetrahedral and the shape is bent for ICl_2^+.

22.9. The electron geometry is octahedral for BrF_5, and the molecular geometry is square pyramidal.

22.10. The oxidation number changes from −1 to 0 for the oxidation of the Cl in HCl to Cl_2 and from +5 to +4 for the reduction of the Cl in $NaClO_3$ to ClO_2. The oxidizing agent is $NaClO_3$, and the reducing agent is HCl.

Chapter 23

23.1. At 50 mol % Ni and 1000 °C this is a solid phase with half of the atoms each Ni and Cu.

23.2. At 50 mol % Ni and 1400 °C this is a liquid phase with half of the atoms each Ni and Cu.

23.3. At 900 °C and 60 mol % Cr this is a two-phase region with more Ni-rich face-centered cubic crystals than Cr-rich body-centered cubic crystals. The Ni-rich phase is about 42 mol % Cr and 58 mol % Ni. The Cr-rich phase is about 94 mol % Cr and 6 mol % Ni.

23.4. At 900 °C and 98 mol % Cr this is a single-phase region with 100 mol % of the Cr-rich body-centered cubic crystals that contains 2% Ni.

Chapter 24

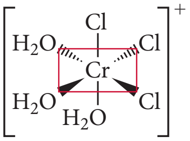
24.1. [Xe] $6s^2 4f^{14} 5d^6$

24.2. [Kr] $5s^0 4d^3$ or [Kr] $4d^3$

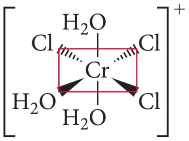
24.3. pentaamminecarbonylmanganese(II) sulfate

24.4. sodium tetrachloroplatinate(II)

24.5. The complex ion $[Cr(H_2O)_3Cl_3]^+$ fits the general formula MA_3B_3, which results in fac and mer isomers.

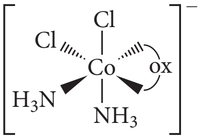


Fac

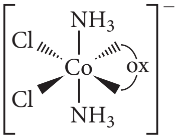


Mer

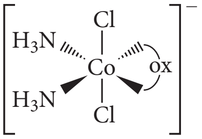
24.6. The oxalate ligand is a small bidentate ligand, so it will have to occupy two adjacent (cis) positions of the octahedron. There are three ways to arrange the two NH_3 and two Cl^- ligands in the four remaining positions. One has both NH_3 and both Cl^- in cis positions (cis isomer). Another has the NH_3 ligands in a trans arrangement with both Cl^- in cis positions (*trans*-ammine isomer). The third has both NH_3 ligands cis and the Cl^- ligands trans (*trans*-chloro isomer).



Cis isomer



Trans (in NH_3)



Trans (in Cl^-)

24.7. Both the fac and mer isomers are superimposable (by rotating 180°) on their mirror images, so neither one is optically active.

24.8. 288 kJ/mol

24.9. five unpaired electrons

24.10. one unpaired electron

Glossary

accuracy A term that refers to how close a measured value is to the actual value. (1.7)

acid A molecular compound that is able to donate an H^+ ion (proton) when dissolved in water, thereby increasing the concentration of H^+. (3.6)

acid ionization constant (K_a) The equilibrium constant for the ionization reaction of a weak acid; used to compare the relative strengths of weak acids. (15.4)

acid–base reaction (neutralization reaction) A reaction in which an acid reacts with a base and the two neutralize each other, producing water. (4.8)

acid–base titration A laboratory procedure in which a basic (or acidic) solution of unknown concentration is reacted with an acidic (or basic) solution of known concentration, in order to determine the concentration of the unknown. (16.4)

acidic solution A solution containing an acid that creates additional H_3O^+ ions, causing $[H_3O^+]$ to increase. (15.4)

activated carbon Very fine carbon particles with high surface area. (22.5)

activated complex (transition state) A high-energy intermediate state between reactant and product. (13.5)

activation energy An energy barrier in a chemical reaction that must be overcome for the reactants to be converted into products. (13.5)

active site The specific area of an enzyme at which catalysis occurs. (13.7)

actual yield The amount of product actually produced by a chemical reaction. (4.3)

addition polymer A polymer in which the monomers simply link together without the elimination of any atoms. (20.14)

addition reaction A type of organic reaction in which two substituents are added across a double bond. (20.10)

alcohol A member of the family of organic compounds that contain a hydroxyl functional group ($-OH$). (3.11, 20.9)

aldehyde A member of the family of organic compounds that contain a carbonyl functional group ($C{=}O$) bonded to two R groups, one of which is a hydrogen atom. (20.10)

aldose A sugar that is an aldehyde. (21.3)

aliphatic hydrocarbons Organic compounds in which carbon atoms are joined in straight or branched chains. (20.3)

alkali metals Highly reactive metals in group 1A of the periodic table. (2.7)

alkaline battery A dry-cell battery that employs slightly different half-reactions in a basic medium. (18.7)

alkaline earth metals Fairly reactive metals in group 2A of the periodic table. (2.7)

alkaloid Organic bases found in plants; they are often poisonous. (15.2)

alkane A hydrocarbon containing only single bonds. (3.11)

alkene A hydrocarbon containing one or more carbon–carbon double bonds. (3.11)

alkyne A hydrocarbon containing one or more carbon–carbon triple bonds. (3.11)

alloy A metallic material that contains more than one element. (23.4)

alpha (α) decay The form of radioactive decay that occurs when an unstable nucleus emits a particle composed of two protons and two neutrons. (19.3)

alpha (α) particle A low-energy particle released during alpha decay; equivalent to a He-4 nucleus. (19.3)

α-helix A pattern in the secondary structure of a protein that occurs when the amino acid chain is wrapped tightly in a coil with the side chains extending outward. (21.5)

aluminosilicates Members of a family of compounds in which aluminum atoms substitute for silicon atoms in some of the silicon lattice sites of the silica structure. (22.3)

amino acids Organic compounds that contain a carbon atom, called the α-carbon, bonded to four different groups: an amine group, an R group, a carboxylic acid group, and a hydrogen atom. (21.4)

ammonia NH_3, the strong smelling compound in which nitrogen displays its lowest oxidation state (-3). (22.6)

amorphous solid A solid in which atoms or molecules do not have any long-range order. (1.3, 11.2)

ampere (A) The SI unit for electrical current; $1\ A = 1\ C/s$. (18.3)

amphoteric Able to act as either an acid or a base. (15.3)

amplitude The vertical height of a crest (or depth of a trough) of a wave; a measure of wave intensity. (7.2)

angular momentum quantum number (l) An integer that determines the shape of an orbital. (7.5)

anion A negatively charged ion. (2.6)

anode The electrode in an electrochemical cell where oxidation occurs; electrons flow away from the anode. (18.3)

antibonding orbital A molecular orbital that is higher in energy than any of the atomic orbitals from which it was formed. (10.8)

aqueous solution A solution in which water acts as the solvent. (4.4, 12.2)

***arachno*-boranes** Boranes with the formula B_nH_{n+6}, consisting of a cage of boron atoms that is missing two or three corners. (22.4)

arc-melting A method in which the solid metal is melted with an arc from a high-voltage electric source in a controlled atmosphere to prevent oxidation. (23.5)

Arrhenius definitions (of acids and bases) The definitions of an acid as a substance that produces H^+ ions in aqueous solution and a base as a substance that produces OH^- ions in aqueous solution. (4.8, 15.3)

Arrhenius equation An equation that relates the rate constant of a reaction to the temperature, the activation energy, and the frequency factor; $k = Ae^{\frac{-E_a}{RT}}$. (13.5)

Arrhenius plot A plot of the natural log of the rate constant ($\ln k$) versus the inverse of the temperature in kelvins ($1/T$) that yields a straight line with a slope of $-E_a/R$ and a y-intercept of $\ln A$. (13.5)

atmosphere (atm) A unit of pressure based on the average pressure of air at sea level; $1\ atm = 101{,}325\ Pa$. (5.2)

atom A submicroscopic particle that constitutes the fundamental building block of ordinary matter; the smallest identifiable unit of an element. (1.1)

atomic element Those elements that exist in nature with single atoms as their basic units. (3.4)

atomic mass (atomic weight) The average mass in amu of the atoms of a particular element based on the relative abundance of the various isotopes; it is numerically equivalent to the mass in grams of one mole of the element. (2.8)

atomic mass unit (amu) A unit used to express the masses of atoms and subatomic particles, defined as $1/12$ the mass of a carbon atom containing 6 protons and 6 neutrons. (2.6)

atomic number (*Z*) The number of protons in an atom; the atomic number defines the element. (2.6)

atomic solids Solids whose composite units are atoms; they include nonbonding atomic solids, metallic atomic solids, and network covalent solids. (11.12)

atomic theory The theory that each element is composed of tiny indestructible particles called atoms, that all atoms of a given element have the same mass and other properties, and that atoms combine in simple, whole-number ratios to form compounds. (1.2, 2.3)

aufbau principle The principle that indicates the pattern of orbital filling in an atom. (8.3)

autoionization The process by which water acts as an acid and a base with itself. (15.4)

Avogadro's law The law that states that the volume of a gas is directly proportional to its amount in moles ($V \propto n$). (5.3)

Avogadro's number The number of ^{12}C atoms in exactly 12 g of ^{12}C; equal to 6.0221421×10^{23}. (2.9)

balanced see *chemical equation* (3.10)

ball-and-stick model A representation of the arrangement of atoms in a molecule that shows how the atoms are bonded to each other and the overall shape of the molecule. (3.3)

band gap An energy gap that exists between the valence band and conduction band of semiconductors and insulators. (11.13)

band theory A model for bonding in atomic solids that comes from molecular orbital theory in which atomic orbitals combine and become delocalized over the entire crystal. (11.13)

barometer An instrument used to measure atmospheric pressure. (5.2)

base ionization constant (K_b) The equilibrium constant for the ionization reaction of a weak base; used to compare the relative strengths of weak bases. (15.7)

basic solution A solution containing a base that creates additional OH^- ions, causing the $[OH^-]$ to increase. (15.4)

beta (β) decay The form of radioactive decay that occurs when an unstable nucleus emits an electron. (19.3)

beta (β) particle A medium-energy particle released during beta decay; equivalent to an electron. (19.3)

β-pleated sheet A pattern in the secondary structure of a protein that occurs when the amino acid chain is extended and forms a zigzag pattern. (21.5)

bidentate Describes ligands that donate two electron pairs to the central metal. (24.3)

bimolecular An elementary step in a reaction that involves two particles, either the same species or different, that collide and go on to form products. (13.6)

binary acid An acid composed of hydrogen and a nonmetal. (3.6)

binary compound A compound that contains only two different elements. (3.5)

biochemistry The study of the chemistry occurring in living organisms. (21.1)

biological effectiveness factor (RBE) A correction factor multiplied by the dose of radiation exposure in rad to obtain the dose in rem. (19.11)

black phosphorus An allotrope of phosphorus with a structure similar to that of graphite; the most thermodynamically stable form. (22.6)

body-centered cubic A unit cell that consists of a cube with one atom at each corner and one atom at the center of the cube. (11.11)

boiling point The temperature at which the vapor pressure of a liquid equals the external pressure. (11.5)

boiling point elevation The effect of a solute that causes a solution to have a higher boiling point than the pure solvent. (12.7)

bomb calorimeter A piece of equipment designed to measure ΔE_{rxn} for combustion reactions at constant volume. (6.4)

bond energy The energy required to break 1 mol of the bond in the gas phase. (9.10)

bond length The average length of a bond between two particular atoms in a variety of compounds. (9.10)

bond order For a molecule, the number of electrons in bonding orbitals minus the number of electrons in nonbonding orbitals divided by two; a positive bond order implies that the molecule is stable. (10.8)

bonding orbital A molecular orbital that is lower in energy than any of the atomic orbitals from which it was formed. (10.8)

bonding pair A pair of electrons shared between two atoms. (9.5)

boranes Compounds composed of boron and hydrogen. (22.4)

Born–Haber cycle A hypothetical series of steps based on Hess's law that represents the formation of an ionic compound from its constituent elements. (9.4)

Boyle's law The law that states that volume of a gas is inversely proportional to its pressure $\left(V \propto \dfrac{1}{P}\right)$. (5.3)

brass A widely used alloy that contains copper and zinc. (23.5)

Brønsted-Lowry definitions (of acids and bases) The definitions of an acid as a proton (H^+ ion) donor and a base as a proton acceptor. (15.3)

bronze An alloy of copper and tin that has been used for thousands of years. (23.5)

buffer A solution containing significant amounts of both a weak acid and its conjugate base (or a weak base and its conjugate acid) that resists pH change by neutralizing added acid or added base. (16.2)

buffer capacity The amount of acid or base that can be added to a buffer without destroying its effectiveness. (16.3)

calcination The heating of an ore in order to decompose it and drive off a volatile product. (23.3)

calorie (cal) A unit of energy defined as the amount of energy required to raise one gram of water 1 °C; equal to 4.184 J. (6.1)

Calorie (Cal) Shorthand notation for the kilocalorie (kcal), or 1000 calories; also called the nutritional calorie, the unit of energy used on nutritional labels. (6.1)

calorimetry The experimental procedure used to measure the heat evolved in a chemical reaction. (6.4)

capillary action The ability of a liquid to flow against gravity up a narrow tube due to adhesive and cohesive forces. (11.4)

carbohydrate A polyhydroxyl aldehyde or ketone. (21.3)

carbon black A fine powdered form of carbon. (22.5)

carbonyl group A functional group consisting of a carbon atom double-bonded to an oxygen atom ($C{=}O$). (20.10)

carboxylic acid An organic acid containing the functional group $—COOH$. (15.2, 20.11)

catalyst A substance that is not consumed in a chemical reaction but increases the rate of the reaction by providing an alternate mechanism in which the rate-determining step has a smaller activation energy. (13.7)

cathode The electrode in an electrochemical cell where reduction occurs; electrons flow toward the cathode. (18.3)

cathode rays A stream of electrons produced when a high electrical voltage is applied between two electrodes within a partially evacuated tube. (2.4)

cation A positively charged ion. (2.6)

cell potential (cell emf) (E_{cell}) The potential difference between the cathode and the anode in an electrochemical cell. (18.3)

cellulose A polysaccharide that consists of glucose units bonded together by β-glycosidic linkages; the main structural component of plants, and the most abundant organic substance on Earth. (21.3)

Celsius (°C) scale The temperature scale most often used by scientists (and by most countries other than the United States), on which pure water freezes at 0 °C and boils at 100 °C (at atmospheric pressure). (1.6)

chain reaction A series of reactions in which previous reactions cause future ones; in a fission bomb, neutrons produced by the fission of one uranium nucleus induce fission in other uranium nuclei. (19.7)

charcoal A fuel similar to coal made by heating wood in the absence of air. (22.5)

Charles's law The law that states that the volume of a gas is directly proportional to its temperature ($V \propto T$). (5.3)

chelate A complex ion that contains either a bi- or polydentate ligand. (24.3)

chelating agent The coordinating ligand of a chelate. (24.3)

chemical bond The sharing or transfer of electrons to attain stable electron configurations for the bonding atoms. (9.3)

chemical change A change that alters the molecular composition of a substance; see also *chemical reaction.* (1.4)

chemical energy The energy associated with the relative positions of electrons and nuclei in atoms and molecules. (6.1)

chemical equation A symbolic representation of a chemical reaction; a balanced equation contains equal numbers of the atoms of each element on both sides of the equation. (3.10)

chemical formula A symbolic representation of a compound that indicates the elements present in the compound and the relative number of atoms of each. (3.3)

chemical property A property that a substance displays only by changing its composition via a chemical change. (1.4)

chemical reaction A process by which one or more substances are converted to one or more different substances; see also *chemical change.* (3.10)

chemical symbol A one- or two-letter abbreviation for an element that is listed directly below its atomic number on the periodic table. (2.6)

chemistry The science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules. (1.1)

chiral molecule A molecule that is not superimposable on its mirror image and thus exhibits optical isomerism. (20.3)

chromosome The DNA-containing structures that occur in the nuclei of living cells. (21.6)

cis–trans isomerism Another term for geometric isomerism; cis-isomers have the same functional group on the same side of a bond and trans-isomers have the same functional group on opposite sides of a bond. (20.5)

Claus process An industrial process for obtaining sulfur through the oxidation of hydrogen sulfide. (22.8)

Clausius–Clapeyron equation An equation that displays the exponential relationship between vapor pressure and temperature; $\ln(P_{vap}) = \frac{-\Delta H_{vap}}{R}\left(\frac{1}{T}\right) + \ln \beta$. (11.5)

***closo*-boranes** Boranes that have the formula $B_{12}H_{12}{}^{2-}$ and form the full icosohedral shape. (22.4)

coal A solid, black fuel with high carbon content, the product of the decomposition of ancient plant material. (22.5)

codon A sequence of three bases in a nucleic acid that codes for one amino acid. (21.6)

coffee-cup calorimeter A piece of equipment designed to measure ΔH_{rxn} for reactions at constant pressure. (6.5)

coke A solid formed by heating coal in the absence of air that consists primarily of carbon and ash. (22.5)

colligative property A property that depends on the amount of a solute but not on the type. (12.7)

collision model A model of chemical reactions in which a reaction occurs after a sufficiently energetic collision between two reactant molecules. (13.5)

colloidal dispersion (colloid) A mixture in which a dispersed substance is finely dived but not truly dissolved in a dispersing medium. (12.8)

combustion analysis A method of obtaining empirical formulas for unknown compounds, especially those containing carbon and hydrogen, by burning a sample of the compound in pure oxygen and analyzing the products of the combustion reaction. (3.9)

combustion reaction A type of chemical reaction in which a substance combines with oxygen to form one or more oxygen-containing compounds; the reaction often causes the evolution of heat and light in the form of a flame. (3.10)

common ion effect The tendency for a common ion to decrease the solubility of an ionic compound or to decrease the ionization of a weak acid or weak base. (16.2)

common name A traditional name of a compound that gives little or no information about its chemical structure; for example, the common name of $NaHCO_3$ is "baking soda." (3.5)

complementary Capable of precise pairing; in particular, the bases of nucleic acids. (21.6)

complementary properties Those properties that exclude one another, that is, the more you know about one, the less you know about the other. For example, the wave nature and particle nature of the electron are complementary. (7.4)

complete ionic equation An equation that lists individually all ofthe ions present as either reactants or products in a chemical reaction. (4.7)

complex carbohydrate Another term for a polysaccharide based on the fact that it is made up of many simple sugars. (21.3)

complex ion An ion that contains a central metal ion bound to one or more ligands. (16.8, 24.3)

compound A substance composed of two or more elements in fixed, definite proportions. (1.3)

concentrated solution A solution that contains a large amount of solute relative to the amount of solvent. (4.4, 12.5)

condensation The phase transition from gas to liquid. (11.5)

condensation polymer A polymer formed by elimination of an atom or small group of atoms (usually water) between pairs of monomers during polymerization. (20.14)

condensation reaction A reaction in which two or more organic compounds are joined, often with the loss of water or some other small molecule. (20.11)

conjugate acid–base pair Two substances related to each other by the transfer of a proton. (15.3)

constructive interference The interaction of waves from two sources that align with overlapping crests, resulting in a wave of greater amplitude. (7.2)

contact process An industrial method for the production of sulfuric acid. (22.8)

conversion factor A factor used to convert between two different units; a conversion factor can be constructed from any two quantities known to be equivalent. (1.8)

coordinate covalent bond The bond formed when a ligand donates electrons to an empty orbital of a metal in a complex ion. (24.3)

coordination compound A neutral compound made when a complex ion combines with one or more counterions. (24.3)

coordination isomers Isomers of complex ions that occur when a coordinated ligand exchanges places with the uncoordinated counter-ion. (24.4)

coordination number (secondary valence) The number of molecules or ions directly bound to the metal atom in a complex ion. (24.3)

coordination number The number of atoms with which each atom in a crystal lattice is in direct contact. (11.11)

core electrons Those electrons in a complete principal energy level and those in complete *d* and *f* sublevels. (8.4)

corrosion The gradual, nearly always undesired oxidation of metals that occurs when they are exposed to oxidizing agents in the environment. (18.9)

covalent bond A chemical bond in which two atoms share electrons that interact with the nuclei of both atoms, lowering the potential energy of each through electrostatic interactions. (3.2, 9.2)

covalent carbides Binary compounds composed of carbon combined with low-electronegativity nonmetals or metalloids. (22.5)

covalent radius (bonding atomic radius) Defined in nonmetals as one-half the distance between two atoms bonded together, and in metals as one-half the distance between two adjacent atoms in a crystal of the metal. (8.6)

critical mass The necessary amount of a radioactive isotope required to produce a self-sustaining fission reaction. (19.7)

critical point The temperature and pressure above which a supercritical fluid exists. (11.8)

critical pressure The pressure required to bring about a transition to a liquid at the critical temperature. (11.5)

critical temperature The temperature above which a liquid cannot exist, regardless of pressure. (11.5)

crystalline lattice The regular arrangement of atoms in a crystalline solid. (11.11)

crystalline solid (crystal) A solid in which atoms, molecules, or ions are arranged in patterns with long-range, repeating order. (1.3, 11.2)

cubic closest packing A closest-packed arrangement in which the third layer of atoms is offset from the first; the same structure as the face-centered cubic. (11.11)

cyclotron A particle accelerator in which a charged particle is accelerated in an evacuated ring-shaped tube by an alternating voltage applied to each semi-circular half of the ring. (19.10)

Dalton's law of partial pressures The law stating that the sum of the partial pressures of the components in a gas mixture must equal the total pressure. (5.6)

de Broglie relation The observation that the wavelength of a particle is inversely proportional to its momentum $\lambda = \frac{h}{mv}$. (7.4)

decanting A method of separating immiscible liquids by pouring the top layer into another container. (1.3)

degenerate A term describing two or more electron orbitals with the same value of *n* that have the same energy. (8.3)

density (*d*) The ratio of an object's mass to its volume. (1.6)

deposition The phase transition from gas to solid. (11.6)

derived unit A unit that is a combination of other base units. For example, the SI unit for speed is meters per second (m/s), a derived unit. (1.6)

destructive interference The interaction of waves from two sources aligned so that the crest of one overlaps the trough of the other, resulting in cancellation. (7.2)

deterministic A characteristic of the classical laws of motion, which imply that present circumstances determine future events. (7.4)

dextrorotatory Capable of rotating the plane of polarization of light clockwise. (20.3)

diamagnetic The state of an atom or ion that contains only paired electrons and is, therefore, slightly repelled by an external magnetic field. (8.7, 10.8)

diamond An elemental form of carbon with a crystal structure that consists of carbon atoms connected to four other carbon atoms at the corners of a tetrahedron, creating a strong network covalent solid. (22.5)

diffraction The phenomena by which a wave emerging from an aperture spreads out to form a new wave front. (7.2)

diffusion The process by which a gas spreads through a space occupied by another gas. (5.9)

dilute solution A solution that contains a very small amount of solute relative to the amount of solvent. (4.4, 12.5)

dimensional analysis The use of units as a guide to solving problems. (1.8)

dimer The product that forms from the reaction of two monomers. (20.14)

diode A device that allows the flow of electrical current in only one direction. (11.13)

dipeptide Two amino acids linked together. (21.4)

dipole moment A measure of the separation of positive and negative charge in a molecule. (9.6)

dipole–dipole force An intermolecular force exhibited by polar molecules that results from the uneven charge distribution. (11.3)

diprotic acid An acid that contains two ionizable protons. (4.8, 15.4)

disaccharide A carbohydrate composed of two monosaccharides. (21.3)

dispersion force (London force) An intermolecular force exhibited by all atoms and molecules that results from fluctuations in the electron distribution. (11.3)

distillation The process by which mixtures of miscible liquids are separated by heating the mixture to boil off the more volatile liquid. The vaporized component is then recondensed and collected in a separate flask. (1.3)

disubstituted benzene A benzene in which two hydrogen atoms have been replaced by other atoms. (20.7)

double bond The bond that forms when two electrons are shared between two atoms. (9.5)

dry-cell battery A battery that does not contain a large amount of liquid water, often using the oxidation of zinc and the reduction of MnO_2 to provide the electrical current. (18.7)

duet A Lewis structure with two dots, signifying a filled outer electron shell for the elements H and He. (9.3)

dynamic equilibrium The point at which the rate of the reverse reaction or process equals the rate of the forward reaction or process. (11.5, 12.4, 14.2)

effective nuclear charge (Z_{eff}) The actual nuclear charge experienced by an electron, defined as the charge of the nucleus plus the charge of the shielding electrons. (8.3)

effusion The process by which a gas escapes from a container into a vacuum through a small hole. (5.9)

electrical charge A fundamental property of certain particles that causes them to experience a force in the presence of electric fields. (2.4)

electrical current The flow of electric charge. (18.3)

electrochemical cell A device in which a chemical reaction either produces or is carried out by an electrical current. (18.3)

electrolysis The process by which electrical current is used to drive an otherwise nonspontaneous redox reaction. (18.8)

electrolyte A substance that dissolves in water to form solutions that conduct electricity. (4.5)

electrolytic cell An electrochemical cell that uses electrical current to drive a nonspontaneous chemical reaction. (18.3)

electromagnetic radiation A form of energy embodied in oscillating electric and magnetic fields. (7.2)

electromagnetic spectrum The range of the wavelengths of all possible electromagnetic radiation. (7.2)

electrometallurgy The use of electrolysis to produce metals from their compounds. (23.3)

electromotive force (emf) The force that results in the motion of electrons due to a difference in potential. (18.3)

electron A negatively charged, low-mass particle found outside the nucleus of all atoms that occupies most of the atom's volume but contributes almost none of its mass. (2.4)

electron affinity (EA) The energy change associated with the gaining of an electron by an atom in its gaseous state. (8.7)

electron capture The form of radioactive decay that occurs when a nucleus assimilates an electron from an inner orbital. (19.3)

electron configuration A notation that shows the particular orbitals that are occupied by electrons in an atom. (8.3)

electron geometry The geometrical arrangement of electron groups in a molecule. (10.3)

electron groups A general term for lone pairs, single bonds, multiple bonds, or lone electrons in a molecule. (10.2)

electron spin A fundamental property of electrons; spin can have a value of $\pm\frac{1}{2}$. (7.5)

electronegativity The ability of an atom to attract electrons to itself in a covalent bond. (9.6)

element A substance that cannot be chemically broken down into simpler substances. (1.3)

elementary step An individual step in a reaction mechanism. (13.6)

emission spectrum The range of wavelengths emitted by a particular element; used to identify the element. (7.3)

empirical formula A chemical formula that shows the simplest whole number ratio of atoms in the compound. (3.3)

empirical formula molar mass The sum of the masses of all the atoms in an empirical formula. (3.9)

enantiomers (optical isomers) Two molecules that are nonsuperimposable mirror images of one another. (20.3, 24.4)

endothermic reaction A chemical reaction that absorbs heat from its surroundings; for an endothermic reaction, $\Delta H > 0$. (6.5)

endpoint The point of pH change where an indicator changes color. (16.4)

energy The capacity to do work. (1.5, 6.1)

English system The system of units used in the United States and various other countries in which the inch is the unit of length, the pound is the unit of force, and the ounce is the unit of mass. (1.6)

enthalpy (*H*) The sum of the internal energy of a system and the product of its pressure and volume; the energy associated with the breaking and forming of bonds in a chemical reaction. (6.5)

entropy A thermodynamic function that is proportional to the number of energetically equivalent ways to arrange the components of a system to achieve a particular state; a measure of the energy randomization or energy dispersal in a system. (12.2, 17.3)

enzyme A biochemical catalyst made of protein that increases the rates of biochemical reactions. (13.7, 21.4)

equilibrium constant (*K*) The ratio, at equilibrium, of the concentrations of the products of a reaction raised to their stoichiometric coefficients to the concentrations of the reactants raised to their stoichiometric coefficients. (14.3)

equivalence point The point in a titration at which the added solute completely reacts with the solute present in the solution; for acid–base titrations, the point at which the amount of acid is stoichiometrically equal to the amount of base in solution. (4.8, 16.4)

ester A family of organic compounds with the general structure $R{-}COO{-}R$. (20.11)

ester linkage The bonds that form between a carboxylic acid and an alcohol to form an ester, such as those in triglycerides. (21.2)

ether A member of the family of organic compounds of the form $R{-}O{-}R'$. (20.12)

exact numbers Numbers that have no uncertainty and thus do not limit the number of significant figures in any calculation. (1.7)

exothermic reaction A chemical reaction that releases heat to its surroundings; for an exothermic reaction, $\Delta H < 0$. (6.5)

experiment A highly controlled procedure designed to generate observations that may support a hypothesis or prove it wrong. (1.2)

exponential factor A number between zero and one that represents the fraction of molecules that have enough energy to make it over the activation barrier on a given approach. (13.5)

extensive property A property that depends on the amount of a given substance, such as mass. (1.6)

extractive metallurgy The process by which an elemental metal must be extracted from the compounds in which it is found. (23.3)

face-centered cubic A crystal structure whose unit cell consists of a cube with one atom at each corner and one atom in the center of every face. (11.11)

Fahrenheit (°F) scale The temperature scale that is most familiar in the United States, on which pure water freezes at 32 °F and boils at 212 °F. (1.6)

family A group of organic compounds with the same functional group. (3.11)

family (group) Columns within the main group elements in the periodic table that contain elements that exhibit similar chemical properties. (2.7)

Faraday's constant (F) The charge in coulombs of 1 mol of electrons: $F = \frac{96{,}485\ \text{C}}{\text{mol e}^-}$. (18.5)

fatty acid A carboxylic acid with a long hydrocarbon tail. (21.2)

ferromagnetic The state of an atom or ion that is very strongly attracted by an external magnetic field. (23.5)

fertilizer A material containing large amounts of nitrogen or phosphorus that is used to increase plant growth. (22.6)

fibrous protein A protein with a relatively linear structure; fibrous proteins tend to be insoluble in aqueous solutions. (21.5)

film-badge dosimeter A device for monitoring exposure to radiation consisting of photographic film held in a small case that is pinned to clothing. (19.5)

filtration A procedure used to separate a mixture composed of an insoluble solid and a liquid by pouring it through filter paper or some other porous membrane or layer. (1.3)

first law of thermodynamics The law stating that the total energy of the universe is constant. (6.2)

flux In pyrometallurgy, material that will react with the gangue to form a substance with a low melting point. (23.3)

formal charge The charge that an atom in a Lewis structure would have if all the bonding electrons were shared equally between the bonded atoms. (9.8)

formation constant (K_f) The equilibrium constant associated with reactions for the formation of complex ions. (16.8)

formula mass The average mass of a molecule of a compound in amu. (3.7)

formula unit The smallest, electrically neutral collection of ions in an ionic compound. (3.4)

Frasch process An industrial process for the recovery of sulfur that uses superheated water to liquefy sulfur deposits in Earth's crust and bring the molten sulfur to the surface. (22.8)

free energy of formation (ΔG_f°) The change in free energy when 1 mol of a compound forms from its constituent elements in their standard states. (17.7)

free radical A molecule or ion with an odd number of electrons in its Lewis structure. (9.9)

freezing The phase transition from liquid to solid. (11.6)

freezing point depression The effect of a solute that causes a solution to have a lower melting point than the pure solvent. (12.7)

frequency (ν) For waves, the number of cycles (or complete wavelengths) that pass through a stationary point in one second. (7.2)

frequency factor The number of times that reactants approach the activation energy per unit time. (13.5)

fuel cell A voltaic cell that uses the oxidation of hydrogen and the reduction of oxygen, forming water, to provide electrical current. (18.7)

fullerenes Carbon clusters, such as C_{60}, bonded in roughly spherical shapes containing from 36 to over 100 carbon atoms. (22.5)

functional group A characteristic atom or group of atoms that imparts certain chemical properties to an organic compound. (3.11)

gamma (γ) rays The form of electromagnetic radiation with the shortest wavelength and highest energy. (7.2, 19.3)

gamma (γ) ray emission The form of radioactive decay that occurs when an unstable nucleus emits extremely high frequency electromagnetic radiation. (19.3)

gangue The undesirable minerals that are separated from specific ores. (23.3)

gas A state of matter in which atoms or molecules have a great deal of space between them and are free to move relative to one another; lacking a definite shape or volume, a gas conforms to those of its container. (1.3)

gas-evolution reaction A reaction in which two aqueous solutions are mixed and a gas forms, resulting in bubbling. (4.8)

Geiger-Müller counter A device used to detect radioactivity that uses argon atoms that become ionized in the presence of energetic particles to produce an electrical signal. (19.5)

gene A sequence of codons within a DNA molecule that codes for a single protein. (21.6)

geometric isomerism A form of stereoisomerism involving the orientation of functional groups in a molecule that contains bonds incapable of rotating. (20.5)

geometric isomers For complex ions, isomers that result when the ligands bonded to the metal have a different spatial arrangement. (24.4)

Gibbs free energy (G) A thermodynamic state function related to enthalpy and entropy by the equation $G = H - TS$; chemical systems tend toward lower Gibbs free energy, also called the *chemical potential*. (17.5)

globular protein A protein that folds into a roughly spherical shape so that its polar side chains are oriented outward and its nonpolar side chains toward the interior; globular proteins tend to be soluble in water. (21.5)

glycogen A highly branched form of starch. (21.3)

glycolipid A triglyceride composed of a fatty acid, a hydrocarbon chain, and a sugar molecule as the polar section. (21.2)

glycosidic linkage A bond between carbohydrates that results from a dehydration reaction. (21.3)

graphite An elemental form of carbon consisting of flat sheets of carbon atoms, bonded together as interconnected hexagonal rings held together by intermolecular forces, that can easily slide past each other. (22.5)

Haber-Bosch process The industrial process for producing ammonia from nitrogen gas and hydrogen gas. (22.6)

half-cell One half of an electrochemical cell where either oxidation or reduction occurs. (18.3)

half-life ($t_{1/2}$) The time required for the concentration of a reactant or the amount of a radioactive isotope to fall to one-half of its initial value. (13.4)

halogens Highly reactive nonmetals in group 7A of the periodic table. (2.7)

heat (q) The flow of energy caused by a temperature difference. (6.1)

heat capacity (C) The quantity of heat required to change a system's temperature by 1 °C. (6.3)

heat of fusion (ΔH_{fus}) The amount of heat required to melt 1 mole of a solid. (11.6)

heat of hydration ($\Delta H_{\text{hydration}}$) The enthalpy change that occurs when 1 mole of gaseous solute ions are dissolved in water. (12.3)

heat of reaction (ΔH_{rxn}) The enthalpy change for a chemical reaction. (6.5)

heat of vaporization (ΔH_{vap}) The amount of heat required to vaporize one mole of a liquid to a gas. (11.5)

Heisenberg's uncertainty principle The principle stating that due to the wave-particle duality, it is fundamentally impossible to precisely determine both the position and velocity of a particle at a given moment in time. (7.4)

Henderson-Hasselbalch equation An equation used to easily calculate the pH of a buffer solution from the initial concentrations of the buffer components, assuming that the "*x is small*" approximation is valid: $\text{pH} = \text{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]}$. (16.2)

Henry's law An equation that expresses the relationship between solubility of a gas and pressure: $S_{gas} = k_H P_{gas}$. (12.4)

Hess's law The law stating that if a chemical equation can be expressed as the sum of a series of steps, then ΔH_{rxn} for the overall equation is the sum of the heats of reactions for each step. (6.7)

heterogeneous catalysis Catalysis in which the catalyst and the reactants exist in different phases. (13.7)

heterogeneous mixture A mixture in which the composition varies from one region to another. (1.3)

hexagonal closest packing A closest-packed arrangement in which the atoms of the third layer align exactly over those in the first layer. (11.11)

hexose A six-carbon sugar. (21.3)

high-spin complex A complex ion with weak field ligands that have the same number of unpaired electrons as the free metal ion. (24.5)

homogeneous catalysis Catalysis in which the catalyst exists in the same phase as the reactants. (13.7)

homogeneous mixture A mixture with the same composition throughout. (1.3)

Hund's rule The principle stating that when electrons fill degenerate orbitals, they first fill them singly with parallel spins. (8.3)

hybrid orbitals Orbitals formed from the combination of standard atomic orbitals that correspond more closely to the actual distribution of electrons in a chemically bonded atom. (10.7)

hybridization A mathematical procedure in which standard atomic orbitals are combined to form new, hybrid orbitals. (10.7)

hydrate An ionic compound that contains a specific number of water molecules associated with each formula unit. (3.5)

hydrazine N_2H_4, a nitrogen and hydrogen compound in which nitrogen has a negative oxidation state (-2). (22.6)

hydrocarbon An organic compound that contains only carbon and hydrogen. (3.11)

hydrogen azide A nitrogen and hydrogen compound with a higher hydrogen-to-nitrogen ratio than ammonia or hydrazine. (22.6)

hydrogen bond A strong dipole–dipole attractive force between a hydrogen bonded to O, N, or F and one of these electronegative atoms on a neighboring molecule. (11.3)

hydrogenation The catalyzed addition of hydrogen to alkene double bonds to make single bonds. (13.7)

hydrolysis The splitting of a chemical bond with water, resulting in the addition of H and OH to the products. (21.3)

hydrometallurgy The use of an aqueous solution to extract metals from their ores. (23.3)

hydronium ion H_3O^+, the ion formed from the association of a water molecule with an H^+ ion donated by an acid. (4.8, 15.3)

hypothesis A tentative interpretation or explanation of an observation. A good hypothesis is *falsifiable*. (1.2)

hypoxia A physiological condition caused by low levels of oxygen, marked by dizziness, headache, shortness of breath, and eventually unconsciousness or even death in severe cases. (5.6)

ideal gas constant The proportionality constant of the ideal gas law, R, equal to 8.314 J/mol · K or 0.08206 L · atm/mol · K. (5.4)

ideal gas law The law that combines the relationships of Boyle's, Charles's, and Avogadro's laws into one comprehensive equation of state with the proportionality constant R in the form $PV = nRT$. (5.4)

ideal solution A solution that follows Raoult's law at all concentrations for both solute and solvent. (12.6)

indeterminacy The principle that present circumstances do not necessarily determine future evens in the quantum-mechanical realm. (7.4)

indicator A dye whose color depends on the pH of the solution it is dissolved in; often used to detect the endpoint of a titration. (4.8, 16.4)

infrared (IR) radiation Electromagnetic radiation emitted from warm objects, with wavelengths slightly larger than those of visible light. (7.2)

insoluble Incapable of dissolving in water or being extremely difficult of solution. (4.5)

integrated rate law A relationship between the concentrations of the reactants in a chemical reaction and time. (13.4)

intensive property A property such as density that is independent of the amount of a given substance. (1.6)

interference The superposition of two or more waves overlapping in space, resulting in either an increase in amplitude (constructive interference) or a decrease in amplitude (destructive interference). (7.2)

interhalogen compounds A class of covalent compounds that contain two different halogens. (22.9)

internal energy (E) The sum of the kinetic and potential energies of all of the particles that compose a system. (6.2)

International System of Units (SI) The standard unit system used by scientists, based on the metric system. (1.6)

interstitial alloy An alloy in which small, usually nonmetallic atoms fit between the metallic atoms of a crystal. (23.4)

ion An atom or molecule with a net charge caused by the loss or gain of electrons. (2.6)

ion product constant for water (K_w) The equilibrium constant for the autoionization of water. (15.5)

ion–dipole force An intermolecular force between an ion and the oppositely charged end of a polar molecule. (11.3)

ionic bond A chemical bond formed between two oppositely charged ions, generally a metallic cation and a nonmetallic anion, that are attracted to one another by electrostatic forces. (3.2, 9.2)

ionic carbides Binary compounds composed of carbon combined with low-electronegativity metals. (22.5)

ionic compound A compound composed of cations and anions bound together by electrostatic attraction. (3.4)

ionic solids Solids whose composite units are ions; they generally have high melting points. (11.12)

ionization energy (IE) The energy required to remove an electron from an atom or ion in its gaseous state. (8.7)

ionizing power The ability of radiation to ionize other molecules and atoms. (19.3)

irreversible reaction A reaction that does not achieve the theoretical limit of available free energy. (17.7)

isotopes Atoms of the same element with the same number of protons but different numbers of neutrons and consequently different masses. (2.6)

joule (J) The SI unit for energy: equal to $1\ \text{kg} \cdot \text{m}^2/\text{s}^2$. (6.1)

kelvin (K) The SI standard unit of temperature. (1.6)

Kelvin scale The temperature scale that assigns 0 K (−273 °C or −459 °F) to the coldest temperature possible, absolute zero, which is the temperature at which molecular motion virtually stops: $1\ \text{K} = 1\ °\text{C}$. (1.6)

ketone A member of the family of organic compounds that contain a carbonyl functional group ($C = O$) bonded to two R groups, neither of which is a hydrogen atom. (20.10)

ketose A sugar that is a ketone. (21.3)

kilogram (kg) The SI standard unit of mass defined as the mass of a block of metal kept at the International Bureau of Weights and Measures at Sèvres, France. (1.6)

kilowatt-hour (kWh) An energy unit used primarily to express large amounts of energy produced by the flow of electricity; equal to 3.60×10^6 J. (6.1)

kinetic energy The energy associated with motion of an object. (1.5, 6.1)

kinetic molecular theory A model of an ideal gas as a collection of point particles in constant motion undergoing completely elastic collisions. (5.8)

lanthanide contraction The trend toward leveling off in size of the atoms in the third and fourth transition rows due to the ineffective shielding of the *f* sublevel electrons. (24.2)

lattice energy The energy associated with forming a crystalline lattice from gaseous ions. (9.4)

law see *scientific law*

law of conservation of energy A law stating that energy can neither be created nor destroyed, only converted from one form to another. (1.5, 6.1)

law of conservation of mass A law stating that matter is neither created nor destroyed in a chemical reaction. (1.2)

law of definite proportions A law stating that all samples of a given compound have the same proportions of their constituent elements. (2.3)

law of mass action The relationship between the balanced chemical equation and the expression of the equilibrium constant. (14.3)

law of multiple proportions A law stating that when two elements (A and B) form two different compounds, the masses of element B that combine with one gram of element A can be expressed as a ratio of small whole numbers. (2.3)

Le Châtelier's principle The principle stating that when a chemical system at equilibrium is disturbed, the system shifts in a direction that minimizes the disturbance. (14.9)

leaching The process by which a metal is separated out of a mixture by selectively dissolving it into solution. (23.3)

lead–acid storage battery A battery that uses the oxidation of lead and the reduction of lead(IV) oxide in sulfuric acid to provide electrical current. (18.7)

lever rule The rule that states that in a two-phase region, whichever phase is closest to the composition of the alloy is the more abundant phase. (23.4)

levorotatory Capable of rotating the polarization of light counterclockwise. (20.3)

Lewis acid An atom, ion, or molecule that is an electron pair acceptor. (15.11)

Lewis base An atom, ion, or molecule that is an electron pair donor. (15.11)

Lewis electron-dot structures (Lewis structures) A drawing that represents chemical bonds between atoms as shared or transferred electrons; the valence electrons of atoms are represented as dots. (9.1)

Lewis theory A simple model of chemical bonding using diagrams that represent bonds between atoms as lines or pairs of dots. In this theory, atoms bond together to obtain stable octets (eight valence electrons). (9.1)

ligand A neutral molecule or an ion that acts as a Lewis base with the central metal ion in a complex ion. (16.8, 24.3)

limiting reactant The reactant that has the smallest stoichiometric amount in a reactant mixture and consequently limits the amount of product in a chemical reaction. (4.3)

linear accelerator A particle accelerator in which a charged particle is accelerated in an evacuated tube by a potential difference between the ends of the tube or by alternating charges in sections of the tube. (19.10)

linear geometry The molecular geometry of three atoms with a 180° bond angle due to the repulsion of two electron groups. (10.2)

linkage isomers Isomers of complex ions that occur when some ligands coordinate to the metal in different ways. (24.4)

lipid A member of the class of biochemical compounds that are insoluble in water but soluble in nonpolar solvents, including fatty acids, triglycerides, and steroids. (21.2)

lipid bilayer A double-layered structure made of phospholipids or glycolipids, in which the polar heads of the molecules interact with the environment and the nonpolar tails interact with each other; a component of many cellular membranes. (21.2)

liquid A state of matter in which atoms or molecules pack about as closely as they do in solid matter but are free to move relative to each other, giving a fixed volume but not a fixed shape. (1.3)

liter (L) A unit of volume equal to $1000\ \text{cm}^3$ or 1.057 qt. (1.6)

lithium ion battery A battery that produces electrical current in the form of motion of lithium ions from the anode to the cathode. (18.7)

lone pair A pair of electrons associated with only one atom. (9.5)

low-spin complex A complex ion with strong field ligands that have fewer unpaired electrons than the free metal ion. (24.5)

magic numbers Certain numbers of nucleons (N or $Z = 2$, 8, 20, 28, 50, 82, and $N = 126$) that confer unique stability. (19.4)

magnetic quantum number (m_l) An integer that specifies the orientation of an orbital. (7.5)

main-group elements Those elements found in the *s* or *p* blocks of the periodic table, whose properties tend to be predictable based on their position in the table. (2.7, 22.2)

manometer An instrument used to determine the pressure of a gaseous sample, consisting of a liquid-filled U-shaped tube with one end exposed to the ambient pressure and the other end connected to the sample. (5.2)

mass A measure of the quantity of matter making up an object. (1.6)

mass defect The difference in mass between the nucleus of an atom and the sum of the separated particles that make up that nucleus. (19.8)

mass number (A) The sum of the number of protons and neutrons in an atom. (2.6)

mass percent composition (mass percent) An element's percentage of the total mass of a compound containing the element. (3.8)

mass spectrometry An experimental method of determining the precise mass and relative abundance of isotopes in a given sample using an instrument called a *mass spectrometer*. (2.8)

matter Anything that occupies space and has mass. (1.3)

mean free path The average distance that a molecule in a gas travels between collisions. (5.9)

melting (fusion) The phase transition from solid to liquid. (11.6)

melting point The temperature at which the molecules of a solid have enough thermal energy to overcome intermolecular forces and become a liquid. (11.6)

metals A large class of elements that are generally good conductors of heat and electricity, malleable, ductile, lustrous, and tend to lose electrons during chemical changes. (2.7)

metallic atomic solids Atomic solids held together by metallic bonds; they have variable melting points. (11.12)

metallic bonding The type of bonding that occurs in metal crystals, in which metal atoms donate their electrons to an electron sea, delocalized over the entire crystal lattice. (9.2)

metallic carbides Binary compounds composed of carbon combined with metals that have a metallic lattice with holes small enough to fit carbon atoms. (22.5)

metalloids A category of elements found on the boundary between the metals and nonmetals of the periodic table, with properties intermediate between those of both groups; also called *semimetals*. (2.7)

metallurgy The part of chemistry that includes all the processes associated with mining, separating, and refining metals and the subsequent production of pure metals and mixtures of metals called alloys. (23.1)

meter (m) The SI standard unit of length; equivalent to 39.37 inches. (1.6)

metric system The system of measurements used in most countries in which the meter is the unit of length, the kilogram is the unit of mass, and the second is the unit of time. (1.6)

microwaves Electromagnetic radiation with wavelengths slightly longer than those of infrared radiation; used for radar and in microwave ovens. (7.2)

milliliter (mL) A unit of volume equal to 10^{-3} L or 1 cm^3. (1.6)

millimeter of mercury (mmHg) A common unit of pressure referring to the air pressure required to push a column of mercury to a height of 1 mm in a barometer; 760 mmHg = 1 atm. (5.2)

mineral A homogenous, naturally occurring, crystalline inorganic solid. (22.2)

miscibility The ability to mix without separating into two phases. (11.3)

miscible The ability of two or more substances to be soluble in each other in all proportions. (12.2)

mixture A substance composed of two or more different types of atoms or molecules that can be combined in variable proportions. (1.3)

molality (*m*) A means of expressing solution concentration as the number of moles of solute per kilogram of solvent. (12.5)

molar heat capacity The amount of heat required to raise the temperature of one mole of a substance by 1 °C. (6.3)

molar mass The mass in grams of one mole of atoms of an element; numerically equivalent to the atomic mass of the element in amu. (2.9)

molar solubility The solubility of a compound in units of moles per liter. (16.5)

molar volume The volume occupied by one mole of a gas; the molar volume of an ideal gas at STP is 22.4 L. (5.5)

molarity (M) A means of expressing solution concentration as the number of moles of solute per liter of solution. (4.4, 12.5)

mole (mol) A unit defined as the amount of material containing 6.0221421×10^{23} (Avogadro's number) particles. (2.9)

mole fraction (χ_A) The number of moles of a component in a mixture divided by the total number of moles in the mixture. (5.6)

mole fraction (χ_{solute}) A means of expressing solution concentration as the number of moles of solute per moles of solution. (12.5)

mole percent A means of expressing solution concentration as the mole fraction multiplied by 100%. (12.5)

molecular compound Compounds composed of two or more covalently bonded nonmetals. (3.4)

molecular element Those elements that exist in nature with diatomic or polyatomic molecules as their basic unit. (3.4)

molecular equation An equation showing the complete neutral formula for each compound in a reaction. (4.7)

molecular formula A chemical formula that shows the actual number of atoms of each element in a molecule of a compound. (3.3)

molecular geometry The geometrical arrangement of atoms in a molecule. (10.3)

molecular orbital theory An advanced model of chemical bonding in which electrons reside in molecular orbitals delocalized over the entire molecule. In the simplest version, the molecular orbitals are simply linear combinations of atomic orbitals. (10.8)

molecular solids Solids whose composite units are molecules; they generally have low melting points. (11.12)

molecularity The number of reactant particles involved in an elementary step. (13.6)

molecule Two or more atoms joined chemically in a specific geometrical arrangement. (1.1)

monodentate Describes ligands that donate only one electron pair to the central metal. (24.3)

monoprotic acid An acid that contains only one ionizable proton. (15.4)

monosaccharide The simplest carbohydrates, with three to eight carbon atoms and only one aldehyde or ketone group. (21.3)

nanotubes Long, tubular structures consisting of interconnected C_6 rings. (22.5)

natural abundance The relative percentage of a particular isotope in a naturally occurring sample with respect to other isotopes of the same element. (2.6)

Nernst equation The equation relating the cell potential of an electrochemical cell to the standard cell potential and the reaction quotient; $E_{\text{cell}} = E^\circ_{\text{cell}} - \dfrac{0.0592\text{ V}}{n}\log Q$. (18.6)

net ionic equation An equation that shows only the species that actually change during the reaction. (4.7)

network covalent atomic solids Atomic solids held together by covalent bonds; they have high melting points. (11.12)

neutral The state of a solution where the concentrations of H_3O^+ and OH^- are equal. (15.4)

neutron An electrically neutral subatomic particle found in the nucleus of an atom, with a mass almost equal to that of a proton. (2.5)

nickel–cadmium (NiCad) battery A battery that consists of an anode composed of solid cadmium and a cathode composed of NiO(OH)(s) in a KOH solution. (18.7)

nickel–metal hydride (NiMH) battery A battery that uses the same cathode reaction as the NiCad battery but a different anode reaction, the oxidation of hydrogens in a metal alloy. (18.7)

***nido*-boranes** Boranes that have the formula B_nH_{n+4} and consist of a cage of boron atoms missing one corner. (22.4)

nitrogen narcosis A physiological condition caused by an increased partial pressure of nitrogen, resulting in symptoms similar to those of intoxication. (5.6)

noble gases The group 8A elements, which are largely unreactive (inert) due to their stable filled *p* orbitals. (2.7)

node A point where the wave function (ψ), and therefore the probability density (ψ^2) and radial distribution function, all go through zero (7.6)

nonbonding atomic solids Atomic solids held together by dispersion forces; they have low melting points. (11.12)

nonbonding orbital An orbital whose electrons remain localized on an atom. (10.8)

nonelectrolyte A compound that does not dissociate into ions when dissolved in water. (4.5)

nonmetal A class of elements that tend to be poor conductors of heat and electricity and usually gain electrons during chemical reactions. (2.7)

nonvolatile Not easily vaporized. (11.5)

normal boiling point The temperature at which the vapor pressure of a liquid equals 1 atm. (11.5)

n-type semiconductor A semiconductor that employs negatively charged electrons in the conduction band as the charge carriers. (11.13)

nuclear binding energy The amount of energy that would be required to break apart the nucleus into its component nucleons. (19.8)

nuclear equation An equation that represents nuclear processes such as radioactivity. (19.3)

nuclear fission The splitting of the nucleus of an atom, resulting in a tremendous release of energy. (19.7)

nuclear fusion The combination of two light nuclei to form a heavier one. (19.9)

nuclear theory The theory that most of the atom's mass and all of its positive charge is contained in a small, dense nucleus. (2.5)

nucleons The particles that compose the nucleus and that are protons and neutrons. (19.4)

nucleotides The individual units composing nucleic acids; each consists of a phosphate group, a sugar, and a nitrogenous base. (21.6)

nucleus The very small, dense core of the atom that contains most of the atom's mass and all of its positive charge; it is composed of protons and neutrons. (2.5)

nuclide A particular isotope of an atom. (19.3)

octahedral arrangement The molecular geometry of seven atoms with 90° bond angles. (10.2)

octahedral hole A space that exists in the middle of six atoms on two adjacent close-packed sheets of atoms in a crystal lattice. (23.4)

octet A Lewis structure with eight dots, signifying a filled outer electron shell for *s* and *p* block elements. (9.3)

octet rule The tendency for most bonded atoms to possess or share eight electrons in their outer shell to obtain stable electron configurations and lower their potential energy. (9.3)

optical isomers Two molecules that are nonsuperimposable mirror images of one another. (20.3, 24.4)

orbital A probability distribution map, based on the quantum mechanical model of the atom, used to describe the likely position of an electron in an atom; also an allowed energy state for an electron. (7.5)

orbital diagram A diagram that gives information similar to an electron configuration but symbolizes an electron as an arrow in a box representing an orbital, with the arrow's direction denoting the electron's spin. (8.3)

ore A rock that contains a high concentration of a specific mineral. (23.2)

organic chemistry The study of carbon-based compounds. (20.1)

organic molecule A molecule containing carbon combined with several other elements including hydrogen, nitrogen, oxygen, or sulfur. (20.1)

orthosilicates Silicates in which tetrahedral SO_4^{4-} ions stand alone. (22.3)

osmosis The flow of solvent from a solution of lower solute concentration to one of higher solute concentration. (12.7)

osmotic pressure The pressure required to stop osmotic flow. (12.7)

Ostwald process An industrial process used for commercial preparation of nitric acid. (22.6)

overall order The sum of the orders of all reactants in a chemical reaction. (13.3)

oxidation The loss of one or more electrons; also the gaining of oxygen or the loss of hydrogen. (4.9)

oxidation state (oxidation number) A positive or negative whole number that represents the "charge" an atom in a compound would have if all shared electrons were assigned to the atom with a greater attraction for those electrons. (4.9)

oxidation–reduction (redox) reaction Reactions in which electrons are transferred from one reactant to another and the oxidation states of certain atoms are changed. (4.9)

oxidizing agent A substance that causes the oxidation of another substance; an oxidizing agent gains electrons and is reduced. (4.9)

oxyacid An acid composed of hydrogen and an oxyanion. (3.6)

oxyanion A polyatomic anion containing a nonmetal covalently bonded to one or more oxygen atoms. (3.5)

oxygen toxicity A physiological condition caused by an increased level of oxygen in the blood, resulting in muscle twitching, tunnel vision, and convulsions. (5.6)

ozone O_3, an allotrope of oxygen that is a toxic blue diamagnetic gas with a strong odor. (22.7)

packing efficiency The percentage of volume of a unit cell occupied by the atoms, assumed to be spherical. (11.11)

paramagnetic The state of an atom or ion that contains unpaired electrons and is, therefore, attracted by an external magnetic field. (8.7, 10.8)

partial pressure (P_n) The pressure due to any individual component in a gas mixture. (5.6)

parts by mass A unit for expressing solution concentration as the mass of the solute divided by the mass of the solution multiplied by a multiplication factor. (12.5)

parts by volume A unit for expressing solution concentration as the volume of the solute divided by the volume of the solution multiplied by a multiplication factor. (12.5)

parts per billion (ppb) A unit for expressing solution concentration in parts by mass where the multiplication factor is 10^9. (12.5)

parts per million (ppm) A unit for expressing solution concentration in parts by mass where the multiplication factor is 10^6. (12.5)

pascal (Pa) The SI unit of pressure, defined as $1\ \text{N/m}^2$. (5.2)

Pauli exclusion principle The principle that no two electrons in an atom can have the same four quantum numbers. (8.3)

penetrating power The ability of radiation to penetrate matter. (19.3)

penetration The phenomenon of some higher-level atomic orbitals having significant amounts of probability within the space occupied by orbitals of lower energy level. For example, the 2*s* orbital penetrates into the 1*s* orbital. (8.3)

peptide bond The bond that forms between the amine end of one amino acid and the carboxylic end of another. (21.4)

percent by mass A unit for expressing solution concentration in parts by mass with a multiplication factor of 100%. (12.5)

percent ionic character The ratio of a bond's actual dipole moment to the dipole moment it would have if the electron were transferred completely from one atom to the other, multiplied by 100%. (9.6)

percent ionization The concentration of ionized acid in a solution divided by the initial concentration of acid multiplied by 100%. (15.6)

percent yield The percentage of the theoretical yield of a chemical reaction that is actually produced; the ratio of the actual yield to the theoretical yield multiplied by 100%. (4.3)

periodic law A law based on the observation that when the elements are arranged in order of increasing mass, certain sets of properties recur periodically. (2.7)

periodic property A property of an element that is predictable based on an element's position in the periodic table. (8.1)

permanent dipole A permanent separation of charge; a molecule with a permanent dipole always has a slightly negative charge at one end and a slightly positive charge at the other. (11.3)

pH The negative log of the concentration of H_3O^+ in a solution; the pH scale is a compact way to specify the acidity of a solution. (15.4)

phase With regard to waves and orbitals, the phase is the sign of the amplitude of the wave, which can be positive or negative. (7.6)

phase diagram A map of the phase of a substance as a function of pressure and temperature. (11.8)

phenyl group A benzene ring treated as a substituent. (20.7)

phosphine PH_3, a colorless, poisonous gas that smells like decaying fish and has an oxidation state of -3 for phosphorus. (22.6)

phospholipid Compound similar in structure to a triglyceride but with one fatty acid replaced by a phosphate group. (21.2)

phosphorescence The long-lived emission of light that sometimes follows the absorption of light by certain atoms and molecules. (19.2)

photoelectric effect The observation that many metals emit electrons when light falls upon them. (7.2)

photon (quantum) The smallest possible packet of electromagnetic radiation with an energy equal to $h\nu$. (7.2)

physical change A change that alters only the state or appearance of a substance but not its chemical composition. (1.4)

physical property A property that a substance displays without changing its chemical composition. (1.4)

pi (π) bond The bond that forms between two *p* orbitals that overlap side to side. (10.7)

p–n junctions Tiny areas in electronic circuits that have p-type semiconductors on one side and n-type on the other. (11.13)

polar covalent bond A covalent bond between two atoms with significantly different electronegativities, resulting in an uneven distribution of electron density. (9.6)

polyatomic ion An ion composed of two or more atoms. (3.5)

polydentate Describes ligands that donate more than one electron pair to the central metal. (24.3)

polypeptide A chain of amino acids joined together by peptide bonds. (21.4)

polyprotic acid An acid that contains more than one ionizable proton and releases them sequentially. (4.8, 15.9)

polysaccharide A long, chainlike molecule composed of many monosaccharide units bonded together. (21.3)

positron The particle released in positron emission; equal in mass to an electron but opposite in charge. (19.3)

positron emission The form of radioactive decay that occurs when an unstable nucleus emits a positron. (19.3)

positron emission tomography (PET) A specialized imaging technique that employs positron-emitting nuclides, such as fluorine-18, as a radiotracer. (19.12)

potential difference A measure of the difference in potential energy (usually in joules) per unit of charge (coulombs). (18.3)

potential energy The energy associated with the position or composition of an object. (1.5, 6.1)

powder metallurgy A process by which metallic components are made from powdered metal. (23.3)

precipitate A solid, insoluble ionic compound that forms in, and separates from, a solution. (4.6)

precipitation reaction A reaction in which a solid, insoluble product forms upon mixing two solutions. (4.6)

precision A term that refers to how close a series of measurements are to one another or how reproducible they are. (1.7)

prefix multipliers Multipliers that change the value of the unit by powers of ten. (1.6)

pressure A measure of force exerted per unit area; in chemistry, most commonly the force exerted by gas molecules as they strike the surfaces around them. (5.1)

pressure–volume work The work that occurs when a volume change takes place against an external pressure. (6.3)

primary structure The sequence of amino acids in a protein chain. (21.5)

primary valence The oxidation state on the central metal atom in a complex ion. (24.3)

principal level (shell) The group of orbitals with the same value of *n*. (7.5)

principal quantum number (*n*) An integer that specifies the overall size and energy of an orbital. The higher the quantum number *n*, the greater the average distance between the electron and the nucleus and the higher its energy. (7.5)

probability density The probability (per unit volume) of finding the electron at a point in space as expressed by a three-dimensional plot of the wave function squared (ψ^2). (7.6)

products The substances produced in a chemical reaction; they appear on the right-hand side of a chemical equation. (3.10)

proton A positively charged subatomic particle found in the nucleus of an atom. (2.5)

p-type semiconductor A semiconductor that employs positively charged "holes" in the valence band as the charge carriers. (11.13)

pure substance A substance composed of only one type of atom or molecule. (1.3)

pyrometallurgy A technique of extractive metallurgy in which heat is used to extract a metal from its mineral. (23.3)

pyrosilicates Silicates in which two SO_4^{4-} tetrahedral ions share a corner. (22.3)

pyroxenes Silicates in which SO_4^{4-} tetrahedral ions bond together to form chains. (22.3)

qualitative analysis A systematic way to determine the ions present in an unknown solution. (16.7)

quantitative analysis A systematic way to determine the amounts of substances in a solution or mixture. (16.7)

quantum number One of four interrelated numbers that determine the shape and energy of orbitals, as specified by a solution of the Schrödinger equation. (7.5)

quantum-mechanical model A model that explains the behavior of absolutely small particles such as electrons and photons. (7.1)

quartz A silicate crystal that has a formula unit of SiO_2. (22.3)

quaternary structure The way that subunits fit together in a multimeric protein. (21.5)

racemic mixture An equimolar mixture of two optical isomers that does not rotate the plane of polarization of light at all. (20.3)

radio waves The form of electromagnetic radiation with the longest wavelengths and smallest energy. (7.2)

radioactive The state of those unstable atoms that emit subatomic particles or high-energy electromagnetic radiation. (19.1)

radioactivity The emission of subatomic particles or high-energy electromagnetic radiation by the unstable nuclei of certain atoms. (2.5, 19.1)

radiocarbon dating A form of radiometric dating based on the C-14 isotope. (19.6)

radiometric dating A technique used to estimate the age of rocks, fossils, or artifacts that depends on the presence of radioactive isotopes and their predictable decay with time. (19.6)

radiotracer A radioactive nuclide that has been attached to a compound or introduced into a mixture in order to track the movement of the compound or mixture within the body. (19.12)

random coils Sections of a protein's secondary structure that have less-regular patterns than α-helixes or β-pleated sheets. (21.5)

random error Error that has equal probability of being too high or too low. (1.7)

Raoult's law An equation used to determine the vapor pressure of a solution; $P_{\text{soln}} = X_{\text{solv}}P^{\circ}_{\text{solv}}$. (12.6)

rate constant (*k*) A constant of proportionality in the rate law. (13.3)

rate law A relationship between the rate of a reaction and the concentration of the reactants. (13.3)

rate-determining step The step in a reaction mechanism that occurs much more slowly than any of the other steps. (13.6)

reactants The starting substances of a chemical reaction; they appear on the left-hand side of a chemical equation. (3.10)

reaction intermediates Species that are formed in one step of a reaction mechanism and consumed in another. (13.6)

reaction mechanism A series of individual chemical steps by which an overall chemical reaction occurs. (13.6)

reaction order (*n*) A value in the rate law that determines how the rate depends on the concentration of the reactants. (13.3)

reaction quotient (Q_c) The ratio, at any point in the reaction, of the concentrations of the products of a reaction raised to their stoichiometric coefficients to the concentrations of the reactants raised to their stoichiometric coefficients. (14.7)

recrystallization A technique used to purify solids in which the solid is put into hot solvent until the solution is saturated; when the solution cools, the purified solute comes out of solution. (12.4)

red phosphorus An allotrope of phosphorus similar in structure to white phosphorus but with one of the bonds between two phosphorus atoms in the tetrahedron broken; red phosphorus is more stable than white. (22.6)

reducing agent A substance that causes the reduction of another substance; a reducing agent loses electrons and is oxidized. (4.9)

reduction The gaining of one or more electrons; also the gaining of hydrogen or the loss of oxygen. (4.9)

refine To purify, particularly a metal. (23.3)

refining A process in which the crude material is purified. (23.3)

rem A unit of the dose of radiation exposure that stands for roentgen equivalent man, where a roentgen is defined as the amount of radiation that produces 2.58×10^{-4} C of charge per kg of air. (19.11)

resonance hybrid The actual structure of a molecule that is intermediate between two or more resonance structures. (9.8)

resonance structures Two or more valid Lewis structures that are shown with double-headed arrows between them to indicate that the actual structure of the molecule is intermediate between them. (9.8)

reversible As applied to a reaction, the ability to proceed in either the forward or the reverse direction. (14.2)

reversible reaction A reaction that achieves the theoretical limit with respect to free energy and will change direction upon an infinitesimally small change in a variable (such as temperature or pressure) related to the reaction. (17.7)

roasting Heating that causes a chemical reaction between a furnace atmosphere and a mineral in order to process ores. (23.3)

salt An ionic compound formed in a neutralization reaction by the replacement of an H^+ ion from the acid with a cation from the base. (4.8)

salt bridge An inverted, U-shaped tube containing a strong electrolyte such as KNO_3 that connects the two half-cells, allowing a flow of ions that neutralizes the charge build-up. (18.3)

saturated fat A triglyceride with no double bonds in the hydrocarbon chain; saturated fats tend to be solid at room temperature. (21.2)

saturated hydrocarbon A hydrocarbon containing no double bonds in the carbon chain. (20.4)

saturated solution A solution in which the dissolved solute is in dynamic equilibrium with any undissolved solute; any added solute will not dissolve. (12.4)

scientific law A brief statement or equation that summarizes past observations and predicts future ones. (1.2)

scientific method An approach to acquiring knowledge about the natural world that begins with observations and leads to the formation of testable hypotheses. (1.2)

scintillation counter A device for the detection of radioactivity using a material that emits ultraviolet or visible light in response to excitation by energetic particles. (19.5)

second (s) The SI standard unit of time, defined as the duration of 9,192,631,770 periods of the radiation emitted from a certain transition in a cesium-133 atom. (1.6)

second law of thermodynamics A law stating that for any spontaneous process, the entropy of the universe increases ($\Delta S_{\text{univ}} > 0$). (17.3)

secondary structure The regular periodic or repeating patterns in the arrangement of protein chains. (21.5)

secondary valence The number of molecules or ions directly bound to the metal atom in a complex ion; also called the *coordination number*. (24.3)

seesaw The molecular geometry of a molecule with trigonal bipyramidal electron geometry and one lone pair in an axial position. (10.3)

selective precipitation A process involving the addition of a reagent to a solution that forms a precipitate with one of the dissolved ions but not the others. (16.6)

semiconductor A material with intermediate electrical conductivity that can be changed and controlled. (2.7)

semipermeable membrane A membrane that selectively allows some substances to pass through but not others. (12.7)

shielding The effect on an electron of repulsion by electrons in lower-energy orbitals that screen it from the full effects of nuclear charge. (8.3)

sigma (σ) bond The resulting bond that forms between a combination of any two *s*, *p*, or hybridized orbitals that overlap end to end. (10.7)

significant figures (significant digits) In any reported measurement, the non-place-holding digits that indicate the precision of the measured quantity. (1.7)

silica A silicate crystal that has a formula unit of SiO_2, also called *quartz*. (22.3)

silicates Covalent atomic solids that contain silicon, oxygen, and various metal atoms. (22.3)

simple cubic A unit cell that consists of a cube with one atom at each corner. (11.11)

slag In pyrometallurgy, the waste liquid solution that is formed between the flux and gangue; usually a silicate material. (23.3)

smelting A form of roasting in which the product is liquefied, which aids in the separation. (23.3)

solid A state of matter in which atoms or molecules are packed close to one another in fixed locations with definite volume. (1.3)

solubility The amount of a substance that will dissolve in a given amount of solvent. (12.2)

solubility product constant (K_{sp}) The equilibrium expression for a chemical equation representing the dissolution of a slightly to moderately soluble ionic compound. (16.5)

soluble Able to dissolve to a significant extent, usually in water. (4.5)

solute The minority component of a solution. (4.4, 12.1)

solution A homogenous mixture of two substances. (4.4, 12.1)

solvent The majority component of a solution. (4.4, 12.1)

space-filling molecular model A representation of a molecule that shows how the atoms fill the space between them. (3.3)

specific heat capacity (C_s) The amount of heat required to raise the temperature of 1 g of a substance by 1 °C. (6.3)

spectator ion Ions in a complete ionic equation that do not participate in the reaction and therefore remain in solution. (4.7)

spin quantum number (m_s) The fourth quantum number, which denotes the electron's spin as either $\frac{1}{2}$ (up arrow) or $-\frac{1}{2}$ (down arrow). (8.3)

spontaneous process A process that occurs without ongoing outside intervention. (17.2)

square planar The molecular geometry of a molecule with octahedral electron geometry and two lone pairs. (10.3)

square pyramidal The molecular geometry of a molecule with octahedral electron geometry and one lone pair. (10.3)

standard cell potential (standard emf) (E°_{cell}) The cell potential for a system in standard states (solute concentration of 1 M and gaseous reactant partial pressure of 1 atm). (18.3)

standard change in free energy ($\Delta G^\circ_{\text{rxn}}$) The change in free energy for a process when all reactants and products are in their standard states. (17.7)

standard enthalpy change (ΔH°) The change in enthalpy for a process when all reactants and products are in their standard states. (6.8)

standard enthalpy of formation (ΔH°_f) The change in enthalpy when 1 mol of a compound forms from its constituent elements in their standard states. (6.8)

standard entropy change (ΔS_{rxn}) The change in entropy for a process when all reactants and products are in their standard states. (17.6)

standard entropy change for a reaction ($\Delta S^\circ_{\text{rxn}}$) The change in entropy for a process in which all reactants and products are in their standard states. (17.6)

Standard Hydrogen Electrode (SHE) The half-cell consisting of an inert platinum electrode immersed in 1 M HCl with hydrogen gas at 1 atm bubbling through the solution; used as the standard of a cell potential of zero. (18.4)

standard molar entropy (S°) A measure of the energy dispersed into one mole of a substance at a particular temperature. (17.6)

standard state For a gas the standard state is the pure gas at a pressure of exactly 1 atm; for a liquid or solid the standard state is the pure substance in its most stable form at a pressure of 1 atm and the temperature of interest (often taken to be 25 °C); for a substance in solution the standard state is a concentration of exactly 1 M. (6.8)

standard temperature and pressure (STP) The conditions of $T = 0\,°\text{C}$ (273 K) and $P = 1$ atm; used primarily in reference to a gas. (5.5)

starch A polysaccharide that consists of glucose units bonded together by α-glycosidic linkages; the main energy storage medium for plants. (21.3)

state A classification of the form of matter as a solid, liquid, or gas. (1.3)

state function A function whose value depends only on the state of the system, not on how the system got to that state. (6.2)

stereoisomers Molecules in which the atoms are bonded in the same order but have a different spatial arrangement. (20.3, 24.4)

steroid A lipid composed of four fused hydrocarbon rings. (21.2)

stock solution A highly concentrated form of a solution used in laboratories to make less concentrated solutions via dilution. (4.4)

stoichiometry The numerical relationships between amounts of reactants and products in a balanced chemical equation. (4.2)

strong acid An acid that completely ionizes in solution. (4.5, 15.4)

strong base A base that completely dissociates in solution. (15.7)

strong electrolyte A substance that completely dissociates into ions when dissolved in water. (4.5)

strong force Of the four fundamental forces of physics, the one that is the strongest but acts over the shortest distance; the strong force is responsible for holding the protons and neutrons together in the nucleus of an atom. (19.4)

strong-field complex A complex ion in which the crystal field splitting is large. (24.5)

structural formula A molecular formula that shows how the atoms in a molecule are connected or bonded to each other. (3.3, 20.3)

structural isomers Molecules with the same molecular formula but different structures. (20.3, 24.4)

sublevel (subshell) Those orbitals in the same principle level with the same value of *n* and *l*. (7.5)

sublimation The phase transition from solid to gas. (11.6)

substitutional alloy An alloy in which one metal atom substitutes for another in the crystal structure. (23.4)

substrate The reactant molecule of a biochemical reaction that binds to an enzyme at the active site. (13.7)

supersaturated solution An unstable solution in which more than the equilibrium amount of solute is dissolved. (12.4)

surface tension The energy required to increase the surface area of a liquid by a unit amount; responsible for the tendency of liquids to minimize their surface area, giving rise to a membrane-like surface. (11.4)

surroundings In thermodynamics, everything in the universe that exists outside the system under investigation. (6.1)

system In thermodynamics, the portion of the universe that is singled out for investigation. (6.1)

systematic error Error that tends toward being consistently either too high or too low. (1.7)

systematic name An official name for a compound, based on well-established rules, that can be determined by examining its chemical structure. (3.5)

temperature A measure of the average kinetic energy of the atoms or molecules that compose a sample of matter. (1.6)

termolecular An elementary step of a reaction in which three particles collide and go on to form products. (13.6)

tertiary structure The large-scale bends and folds produced by interactions between the R groups of amino acids that are separated by large distances in the linear sequence of a protein chain. (21.5)

tetrahedral geometry The molecular geometry of five atoms with 109.5° bond angles. (10.2)

tetrahedral hole A space that exists directly above the center point of three closest-packed metal atoms in one plane and a fourth metal located directly above the center point in the adjacent plane in a crystal lattice. (23.4)

theoretical yield The greatest possible amount of product that can be made in a chemical reaction based on the amount of limiting reactant. (4.3)

theory A proposed explanation for observations and laws, based on well-established and tested hypotheses, that presents a model of the way nature works and predicts behavior beyond the observations and laws on which it was based. (1.2)

thermal energy A type of kinetic energy associated with the temperature of an object, arising from the motion of individual atoms or molecules in the object; see also *heat*. (1.5, 6.1)

thermal equilibrium The point at which there is no additional net transfer of heat between a system and its surroundings. (6.3)

thermochemistry The study of the relationship between chemistry and energy. (6.1)

thermodynamics The general study of energy and its interconversions. (6.2)

third law of thermodynamics The law stating that the entropy of a perfect crystal at absolute zero (0 K) is zero. (17.6)

titration A laboratory procedure in which a substance in a solution of known concentration is reacted with another substance in a solution of unknown concentration in order to determine the unknown concentration; see also *acid–base titration*. (4.8)

transition elements (transition metals) Those elements found in the *d* block of the periodic table whose properties tend to be less predictable based simply on their position in the table. (2.7)

transmutation The transformation of one element into another as a result of nuclear reactions. (19.10)

triglyceride Triesters composed of glycerol with three fatty acids attached. (21.2)

trigonal bipyramidal The molecular geometry of six atoms with 120° bond angles between the three equatorial electron groups and 90° bond angles between the two axial electron groups and the trigonal plane. (10.2)

trigonal planar geometry The molecular geometry of four atoms with 120° bond angles in a plane. (10.2)

trigonal pyramidal The molecular geometry of a molecule with tetrahedral electron geometry and one lone pair. (10.3)

triple bond The bond that forms when three electron pairs are shared between two atoms. (9.5)

triple point The unique set of conditions at which all three phases of a substance are equally stable and in equilibrium. (11.8)

triprotic acid An acid that contains three ionizable protons. (15.4)

T-shaped The molecular geometry of a molecule with trigonal bipyramidal electron geometry and two lone pairs in axial positions. (10.3)

two-phase region The region between the two phases in a metal alloy phase diagram, where the amount of each phase depends upon the composition of the alloy. (23.4)

Tyndall effect The scattering of light by a colloidal dispersion. (12.8)

ultraviolet (UV) radiation Electromagnetic radiation with slightly smaller wavelengths than visible light. (7.2)

unimolecular Describes a reaction that involves only one particle that goes on to form products. (13.6)

unit cell The smallest divisible unit of a crystal that, when repeated in three dimensions, reproduces the entire crystal lattice. (11.11)

units Standard quantities used to specify measurements. (1.6)

unsaturated fat A triglyceride with one or more double bonds in the hydrocarbon chain; unsaturated fats tend to be liquid at room temperature. (21.2)

unsaturated hydrocarbon A hydrocarbon that includes one or more double or triple bonds. (20.5)

unsaturated solution A solution containing less than the equilibrium amount of solute; any added solute will dissolve until equilibrium is reached. (12.4)

valence bond theory An advanced model of chemical bonding in which electrons reside in quantum-mechanical orbitals localized on individual atoms that are a hybridized blend of standard atomic orbitals; chemical bonds result from an overlap of these orbitals. (10.6)

valence electrons Those electrons that are important in chemical bonding. For main-group elements, the valence electrons are those in the outermost principal energy level. (8.4)

valence shell electron pair repulsion (VSEPR) theory A theory that allows prediction of the shapes of molecules based on the idea that electrons—either as lone pairs or as bonding pairs—repel one another. (10.2)

van der Waals equation The extrapolation of the ideal gas law that considers the effects of intermolecular forces and particle volume in a nonideal gas: $P + a\left(\frac{n}{V}\right)^2 \times (V - nb) = nRT$. (5.9)

van der Waals radius (nonbonding atomic radius) Defined as one-half the distance between the centers of adjacent, nonbonding atoms in a crystal. (8.6)

van't Hoff factor (*i*) The ratio of moles of particles in a solution to moles of formula units dissolved. (12.7)

vapor pressure The partial pressure of a vapor in dynamic equilibrium with its liquid. (5.6, 11.5)

vaporization The phase transition from liquid to gas. (11.5)

viscosity A measure of the resistance of a liquid to flow. (11.4)

visible light Those frequencies of electromagnetic radiation that can be detected by the human eye. (7.2)

volatile Tending to vaporize easily. (1.3, 11.5)

voltaic (galvanic) cell An electrochemical cell that produces electrical current from a spontaneous chemical reaction. (18.3)

volume (*V*) A measure of space. Any unit of length, when cubed (raised to the third power), becomes a unit of volume. (1.6)

washing soda The hydrated crystal of sodium carbonate, $Na_2CO_3 \cdot 10\ H_2O$. (22.5)

wave function (*ψ*) A mathematical function that describes the wavelike nature of the electron. (7.5)

wavelength (λ) The distance between adjacent crests of a wave. (7.2)

weak acid An acid that does not completely ionize in water. (4.5, 15.4)

weak base A base that only partially ionizes in water. (15.7)

weak electrolyte A substance that does not completely ionize in water and only weakly conducts electricity in solution. (4.5)

weak-field complex A complex ion in which the crystal field splitting is small. (24.5)

white phosphorus An unstable allotrope of phosphorus consisting of P_4 molecules in a tetrahedral shape, with the phosphorus atoms at the corners of the tetrahedron. (22.6)

work (*w*) The result of a force acting through a distance. (1.5, 6.1)

X-rays Electromagnetic radiation with wavelengths slightly longer than those of gamma rays; used to image bones and internal organs. (7.2)

X-ray diffraction A powerful laboratory technique that allows for the determination of the arrangement of atoms in a crystal and the measuring of the distance between them. (11.10)

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🟨 Metals 🟪 Metalloids 🟩 Nonmetals

Main groups (1A–2A); Transition metals (3B–2B); Main groups (3A–8A)

	1A[a] 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	 9	 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.94	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** [98]	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
6	55 **Cs** 132.91	56 **Ba** 137.33	57 **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** [208.98]	85 **At** [209.99]	86 **Rn** [222.02]
7	87 **Fr** [223.02]	88 **Ra** [226.03]	89 **Ac** [227.03]	104 **Rf** [261.11]	105 **Db** [262.11]	106 **Sg** [266.12]	107 **Bh** [264.12]	108 **Hs** [269.13]	109 **Mt** [268.14]	110 **Ds** [271]	111 **Rg** [272]	112 **Cn** [285]	113	114 **Fl** [289]	115	116 **Lv** [292]	117*	118

Lanthanide series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** [145]	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
Actinide series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** [237.05]	94 **Pu** [244.06]	95 **Am** [243.06]	96 **Cm** [247.07]	97 **Bk** [247.07]	98 **Cf** [251.08]	99 **Es** [252.08]	100 **Fm** [257.10]	101 **Md** [258.10]	102 **No** [259.10]	103 **Lr** [262.11]

[a] The labels on top (1A, 2A, etc.) are common American usage. The labels below these (1, 2, etc.) are those recommended by the International Union of Pure and Applied Chemistry.

Atomic masses in brackets are the masses of the longest-lived or most important isotope of radioactive elements.

*Element 117 is currently under review by IUPAC.

List of Elements with Their Symbols and Atomic Masses

Element	Symbol	Atomic Number	Atomic Mass
Actinium	Ac	89	227.03[a]
Aluminum	Al	13	26.98
Americium	Am	95	243.06[a]
Antimony	Sb	51	121.76
Argon	Ar	18	39.95
Arsenic	As	33	74.92
Astatine	At	85	209.99[a]
Barium	Ba	56	137.33
Berkelium	Bk	97	247.07[a]
Beryllium	Be	4	9.012
Bismuth	Bi	83	208.98
Bohrium	Bh	107	264.12[a]
Boron	B	5	10.81
Bromine	Br	35	79.90
Cadmium	Cd	48	112.41
Calcium	Ca	20	40.08
Californium	Cf	98	251.08[a]
Carbon	C	6	12.01
Cerium	Ce	58	140.12
Cesium	Cs	55	132.91
Chlorine	Cl	17	35.45
Chromium	Cr	24	52.00
Cotalt	Co	27	58.93
Copernicium	Cn	112	285[a]
Copper	Cu	29	63.55
Curium	Cm	96	247.07[a]
Darmstadtium	Ds	110	271[a]
Dubnium	Db	105	262.11[a]
Dysprosium	Dy	66	162.50
Einsteinium	Es	99	252.08[a]
Erbium	Er	68	167.26
Europium	Eu	63	151.96
Fermium	Fm	100	257.10[a]
Flerovium	Fl	114	289[a]
Fluorine	F	9	19.00
Francium	Fr	87	223.02[a]
Gadolinium	Gd	64	157.25
Gallium	Ga	31	69.72
Germanium	Ge	32	72.63
Gold	Au	79	196.97
Hafnium	Hf	72	178.49
Hassium	Hs	108	269.13[a]
Helium	He	2	4.003
Holmium	Ho	67	164.93
Hydrogen	H	1	1.008
Indium	In	49	114.82
Iodine	I	53	126.90
Iridium	Ir	77	192.22
Iron	Fe	26	55.85
Krypton	Kr	36	83.80
Lanthanum	La	57	138.91
Lawrencium	Lr	103	262.11[a]
Lead	Pb	82	207.2
Lithium	Li	3	6.94
Livermorium	Lv	116	292[a]
Lutetium	Lu	71	174.97
Magnesium	Mg	12	24.31
Manganese	Mn	25	54.94
Meitnerium	Mt	109	268.14[a]
Mendelevium	Md	101	258.10[a]
Mercury	Hg	80	200.59
Molybdenum	Mo	42	95.96
Neodymium	Nd	60	144.24
Neon	Ne	10	20.18
Neptunium	Np	93	237.05[a]
Nickel	Ni	28	58.69
Niobium	Nb	41	92.91
Nitrogen	N	7	14.01
Nobelium	No	102	259.10[a]
Osmium	Os	76	190.23
Oxygen	O	8	16.00
Palladium	Pd	46	106.42
Phosphorus	P	15	30.97
Platinum	Pt	78	195.08
Plutonium	Pu	94	244.06[a]
Polonium	Po	84	208.98[a]
Potassium	K	19	39.10
Praseodymium	Pr	59	140.91
Promethium	Pm	61	145[a]
Protactinium	Pa	91	231.04
Radium	Ra	88	226.03[a]
Radon	Rn	86	222.02[a]
Rhenium	Re	75	186.21
Rhodium	Rh	45	102.91
Roentgenium	Rg	111	272[a]
Rubidium	Rb	37	85.47
Ruthenium	Ru	44	101.07
Rutherfordium	Rf	104	261.11[a]
Samarium	Sm	62	150.36
Scandium	Sc	21	44.96
Seaborgium	Sg	106	266.12[a]
Selenium	Se	34	78.96
Silicon	Si	14	28.09
Silver	Ag	47	107.87
Sodium	Na	11	22.99
Strontium	Sr	38	87.62
Sulfur	S	16	32.06
Tantalum	Ta	73	180.95
Technetium	Tc	43	98[a]
Tellurium	Te	52	127.60
Terbium	Tb	65	158.93
Thallium	Tl	81	204.38
Thorium	Th	90	232.04
Thulium	Tm	69	168.93
Tin	Sn	50	118.71
Titanium	Ti	22	47.87
Tungsten	W	74	183.84
Uranium	U	92	238.03
Vanadium	V	23	50.94
Xenon	Xe	54	131.293
Ytterbium	Yb	70	173.05
Yttrium	Y	39	88.91
Zinc	Zn	30	65.38
Zirconium	Zr	40	91.22
*[b]		113	284[a]
*[b]		115	288[a]

[a] Mass of longest-lived or most important isotope.

[b] The names of these elements have not yet been decided.

Conversion Factors and Relationships

Length
SI unit: meter (m)
1 m = 1.0936 yd
1 cm = 0.39370 in
1 in = 2.54 cm (exactly)
1 km = 0.62137 mi
1 mi = 5280 ft
= 1.6093 km
1 Å = 10^{-10} m

Temperature
SI unit: kelvin (K)
0 K = −273.15 °C
= −459.67 °F
K = °C + 273.15
$$°\text{C} = \frac{(°\text{F} - 32)}{1.8}$$
°F = 1.8 (°C) + 32

Energy (derived)
SI unit: joule (J)
$1\ \text{J} = 1\ \text{kg}\cdot\text{m}^2/\text{s}^2$
= 0.23901 cal
= 1 C · V
$= 9.4781 \times 10^{-4}\ \text{Btu}$
1 cal = 4.184 J
$1\ \text{eV} = 1.6022 \times 10^{-19}\ \text{J}$

Pressure (derived)
SI unit: pascal (Pa)
$1\ \text{Pa} = 1\ \text{N/m}^2$
$= 1\ \text{kg/(m}\cdot\text{s}^2)$
1 atm = 101,325 Pa
= 760 torr
$= 14.70\ \text{lb/in}^2$
$1\ \text{bar} = 10^5\ \text{Pa}$
1 torr = 1 mmHg

Volume (derived)
SI unit: cubic meter (m^3)
$1\ \text{L} = 10^{-3}\ \text{m}^3$
$= 1\ \text{dm}^3$
$= 10^3\ \text{cm}^3$
= 1.0567 qt
1 gal = 4 qt
= 3.7854 L
$1\ \text{cm}^3 = 1\ \text{mL}$
$1\ \text{in}^3 = 16.39\ \text{cm}^3$
1 qt = 32 fluid oz

Mass
SI unit: kilogram (kg)
1 kg = 2.2046 lb
1 lb = 453.59 g
= 16 oz
$1\ \text{amu} = 1.66053873 \times 10^{-27}\ \text{kg}$
1 ton = 2000 lb
= 907.185 kg
1 metric ton = 1000 kg
= 2204.6 lb

Geometric Relationships
$\pi = 3.14159\ldots$
Circumference of a circle $= 2\pi r$
Area of a circle $= \pi r^2$
Surface area of a sphere $= 4\pi r^2$
Volume of a sphere $= \frac{4}{3}\pi r^3$
Volume of a cylinder $= \pi r^2 h$

Fundamental Constants

Atomic mass unit	1 amu 1 g	$= 1.66053873 \times 10^{-27}$ kg $= 6.02214199 \times 10^{23}$ amu
Avogadro's number	N_A	$= 6.02214179 \times 10^{23}$/mol
Bohr radius	a_0	$= 5.29177211 \times 10^{-11}$ m
Boltzmann's constant	k	$= 1.38065052 \times 10^{-23}$ J/K
Electron charge	e	$= 1.60217653 \times 10^{-19}$ C
Faraday's constant	F	$= 9.64853383 \times 10^{4}$ C/mol
Gas constant	R	= 0.08205821 (L · atm/(mol · K) = 8.31447215 J/(mol · K)
Mass of an electron	m_e	$= 5.48579909 \times 10^{-4}$ amu $= 9.10938262 \times 10^{-31}$ kg
Mass of a neutron	m_n	= 1.00866492 amu $= 1.67492728 \times 10^{-27}$ kg
Mass of a proton	m_p	= 1.00727647 amu $= 1.67262171 \times 10^{-27}$ kg
Planck's constant	h	$= 6.62606931 \times 10^{-34}$ J · s
Speed of light in vacuum	c	$= 2.99792458 \times 10^{8}$ m/s (exactly)

SI Unit Prefixes

a	f	p	n	μ	m	c	d	k	M	G	T	P	E
atto	femto	pico	nano	micro	milli	centi	deci	kilo	mega	giga	tera	peta	exa
10^{-18}	10^{-15}	10^{-12}	10^{-9}	10^{-6}	10^{-3}	10^{-2}	10^{-1}	10^{3}	10^{6}	10^{9}	10^{12}	10^{15}	10^{18}

Selected Key Equations

Density (1.6)

$$d = \frac{m}{V}$$

Solution Dilution (4.4)

$$M_1 V_1 = M_2 V_2$$

Ideal Gas Law (5.4)

$$PV = nRT$$

Dalton's Law (5.6)

$$P_{\text{total}} = P_a + P_b + P_c + \ldots$$

Mole Fraction (5.6)

$$\chi_a = \frac{n_a}{n_{\text{total}}}$$

Average Kinetic Energy (5.8)

$$KE_{\text{avg}} = \frac{3}{2}RT$$

Root Mean Square Velocity (5.8)

$$u_{\text{rms}} = \sqrt{\frac{3\,RT}{\mathcal{M}}}$$

Effusion (5.9)

$$\frac{\text{rate A}}{\text{rate B}} = \sqrt{\frac{\mathcal{M}_B}{\mathcal{M}_A}}$$

Van der Waals Equation (5.10)

$$\left[P + a\left(\frac{n}{V}\right)^2\right] \times [V - nb] = nRT$$

Kinetic Energy (6.2)

$$\text{KE} = \frac{1}{2}mv^2$$

Internal Energy (6.3)

$$\Delta E = q + w$$

Heat Capacity (6.4)

$$q = m \times C_s \times \Delta T$$

Pressure-Volume Work (6.4)

$$w = -P\,\Delta V$$

Change in Enthalpy (6.6)

$$\Delta H = \Delta E + P\,\Delta V$$

Standard Enthalpy of Reaction (6.9)

$$\Delta H^\circ_{\text{rxn}} = \sum n_p\,\Delta H^\circ_f(\text{products}) - \sum n_r\,\Delta H^\circ_f(\text{reactants})$$

Frequency and Wavelength (7.2)

$$\nu = \frac{c}{\lambda}$$

Energy of a Photon (7.2)

$$E = h\nu$$

$$E = \frac{hc}{\lambda}$$

De Broglie Relation (7.4)

$$\lambda = \frac{h}{m\nu}$$

Heisenberg's Uncertainty Principle (7.4)

$$\Delta x \times m\,\Delta v \geq \frac{h}{4\pi}$$

Energy of Hydrogen Atom Levels (7.5)

$$E_n = -2.18 \times 10^{-18}\,\text{J}\left(\frac{1}{n^2}\right) \quad (n = 1, 2, 3 \ldots)$$

Coulomb's Law (8.3)

$$E = \frac{1}{4\,\pi\varepsilon_o}\,\frac{q_1 q_2}{r}$$

Dipole Moment (9.6)

$$\mu = qr$$

Clausius–Clapeyron Equation (11.5)

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{RT} + \ln \beta$$

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

Henry's Law (12.4)

$$S_{\text{gas}} = k_H P_{\text{gas}}$$

Raoult's Law (12.6)

$$P_{\text{solution}} = \chi_{\text{solvent}} P^\circ_{\text{solvent}}$$

Freezing Point Depression (12.6)

$$\Delta T_f = m \times K_f$$

Boiling Point Elevation Constant (12.6)

$$\Delta T_b = m \times K_b$$

Osmotic Pressure (12.6)

$$\Pi = MRT$$

The Rate Law (13.3)

$$\text{Rate} = k[A]^n \quad \text{(single reactant)}$$

$$\text{Rate} = k[A]^m[B]^n \quad \text{(multiple reactants)}$$

Integrated Rate Laws and Half-Life (13.4)

Order	Integrated Rate Law	Half-Life Expression
0	$[A]_t = -kt + [A]_0$	$t_{1/2} = \frac{[A]_0}{2k}$
1	$\ln[A]_t = -kt + \ln[A]_0$	$t_{1/2} = \frac{0.693}{k}$
2	$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$	$t_{1/2} = \frac{1}{k[A]_0}$

Arrhenius Equation (13.5)

$$k = A\,e^{\frac{-E_a}{RT}}$$

$$\ln k = -\frac{E_a}{R}\left(\frac{1}{T}\right) + \ln A \quad \text{(linearized form)}$$

$$k = pz\,e^{\frac{-E_a}{RT}} \quad \text{(collision theory)}$$

K_c and K_p (14.4)

$$K_p = K_c(RT)^{\Delta n}$$

pH Scale (15.5)

$$\text{pH} = -\log[H_3O^+]$$

Henderson–Hasselbalch Equation (16.2)

$$\text{pH} = \text{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]}$$

Entropy (17.3)

$$S = k \ln W$$

Change in the Entropy of the Surroundings (17.4)

$$\Delta S_{\text{surr}} = \frac{-\Delta H_{\text{sys}}}{T}$$

Change in Gibbs Free Energy (17.5)

$$\Delta G = \Delta H - T\,\Delta S$$

The Change in Free Energy: Nonstandard Conditions (17.8)

$$\Delta G_{\text{rxn}} = \Delta G^\circ_{\text{rxn}} + RT \ln Q$$

$\Delta G^\circ_{\text{rxn}}$ and K (17.9)

$$\Delta G^\circ_{\text{rxn}} = -RT \ln K$$

Temperature Dependence of the Equilibrium Constant (17.9)

$$\ln K = -\frac{\Delta H^\circ_{\text{rxn}}}{R}\left(\frac{1}{T}\right) + \frac{\Delta S^\circ_{\text{rxn}}}{R}$$

ΔG° and E°_{cell} (18.5)

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

E°_{cell} and K (18.5)

$$E^\circ_{\text{cell}} = \frac{0.0592\text{ V}}{n}\log K$$

Nerst Equation (18.6)

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592\text{ V}}{n}\log Q$$

Einstein's Energy-Mass Equation (19.8)

$$E = mc^2$$

~StormRG~