

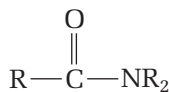
Answers to Questions in Appendices

Appendix B Supplemental Practice Problems

UNIT 1

Chapter 1

1. (a) ketone
(b) aldehyde
(c) cycloalkane
(d) ether
2. Amine: functional group is $-\text{NR}_2$, $-\text{NH}_2$, or $-\text{NHR}$;
Amide: functional group is as shown below, where R is H or an alkyl group. Primary and secondary amines and amides of low molar mass are soluble.



3. (a) alcohol
(b) hydrogen bonding, dipole-dipole
(c) soluble in water
4. (a) H-bonding, dipole-dipole, dispersion forces
(b) yes
(c) 2,3-hydroxy-1,4-butanedioic acid
5. (a) phenyl butanoate
(b) 1-ethoxy-1-methylpropane
(c) 4,5-dichlorocyclohexene
6. (a) $\text{F}-\text{C}\equiv\text{C}-\text{CFH}-\text{C}(\text{CH}_3)-\text{CH}_3$
(b)

$$\begin{array}{c} \text{H}_3\text{C} \quad \quad \text{H} \\ \quad \quad \quad \diagdown \quad \diagup \\ \quad \quad \quad \text{C}=\text{C} \\ \quad \quad \quad \diagup \quad \diagdown \\ \text{H} \quad \quad \quad \text{CH}_2-\text{CH}_2-\text{CH}_2-\text{C}(\text{CH}_3)_2-\text{CH}_3 \end{array}$$
- (c) $\text{C}_4\text{H}_9\text{CONH}_2$
7. (a) isopropyl-2,2-dimethylpentanoate
(b) N-ethyl-2-butanamine
(c) 4-heptanone
8. *t*-alcohol: $-\text{OH}$ functional group bonded to a C attached to 3 other carbons;
t-amine: N of amine functional group is bonded to 3 other alkyl groups
9. 1,2,3-trichlorobenzene, 1,2,4-trichlorobenzene, 1,3,4-trichlorobenzene, 1,3,5-trichlorobenzene
10. (a) heptanoic acid
(b) methoxyethane
(c) 2-propanone
11. (a) $\text{C}_{10}\text{H}_{18}$
(b) many possibilities: cyclic 2,4,6,8-tetramethylcyclohexene
non-cyclic 4,4-dimethyl-1,7-octadiene

12. (a) methyl propanoate
(b) methoxyethane
(c) 3-hexanol
(d) N-pentyl-N-ethylethanamide
(e) bromobenzene
13. (a) trans-2,3-dibromopentene
(b) methylcyclopentane
(c) 2-heptanol
(d) 1-chloro-1,3-dimethylbutane

Chapter 2

14. (a) 2-pentanol
(b) 1-pentanol
15. (a) 1-bromobutane + H_2O
(b) HCl + 2-chloro-2-methylpropane
(c) bromobenzene + HBr
(d) 2-pentanol + NaBr
16. (a) $\text{C}_{16}\text{H}_{33}-\text{HC}=\text{CH}_2$
(b) 2,3-dimethyl-2-butene
17. (a) 3-pentanone
(b) ethanoic acid
(c) 2-ethyl-2-methylbutanal
(d) cyclooctanone
18. (a) 1-propanol
(b) 3-hexanol
(c) ethane
(d) methane
19. (a) methyl heptanoate $\text{C}_6\text{H}_{13}\text{COOCH}_3$
(b) ethyl octanoate $\text{C}_7\text{H}_{15}\text{COOC}_2\text{H}_5$
(c) $\text{C}_2\text{H}_5\text{COO}-\text{CH}(\text{CH}_3)_2$
(d) 2-butylpentanoate $\text{C}_4\text{H}_9\text{COO}-\text{C}(\text{CH}_3)-\text{C}_2\text{H}_5$
20. (a) esterification, condensation
(b) hydrolysis
(c) only single bonds to carbons in the $\text{C}_{17}\text{H}_{35}$ portion of molecule
(d) 3 $\text{C}=\text{C}$ bonds in linolenic acid
(e) A high fat diet may lead to buildup of solid fat in tissue and in arteries (note that other factors affect this)
21. ether linkage, two benzene rings, one chlorinated at the ortho or 1,2 position and the other chlorinated at the meta or 1,3 position
22. alkenes, molecules with the same functional group at each end of the molecule, molecules having two different functional groups at opposite ends of the molecule
23. 2-methylpropene
24. condensation reaction
25. diamine, dicarboxylic acid

UNIT 2

Chapter 3

26. Compare a spectral analysis of the light from the object to atomic spectra of known elements.
27. $\text{Cs} < \text{Rb} < \text{Sr}$; because size is $\text{Cs} > \text{Rb} > \text{Sr}$
28. EA for He $< EA$ for Xe and IE_1 for He $> IE_1$ for Xe (that is, He gains and loses electrons less readily than Xe).
29. (a) 14 e
(b) too much energy is required
(c) Sn^{2+} is $[\text{Kr}]4d^{10}5s^2$; Sn^{4+} is $[\text{Kr}]4d^{10}$
(d) Neither is a noble gas electron configuration. Since both ions are known to be stable, their electron configurations must produce some distribution of charge within the atom that gives stability, though not likely as much as a noble gas configuration.
(e) Sn^{2+}
(f) Sn^{4+}
(g) too great a repulsion between the added electrons
30. In a many-electron atom, electron-electron repulsions and nucleus-electron attractions combine to cause a splitting of the energy level (as found in hydrogen) into sublevels. The energy of the atomic orbitals in hydrogen depends only on the principal quantum number, n ; in a many-electron atom the energy also depends upon the angular momentum quantum number l . At atomic number 20, the $3d$ sublevel energy $> 4s$ so electrons fill the $4s$ orbital first. At atomic number > 20 , the $3d$ sublevel energy $< 4s$ and electrons fill the $3d$ orbital, even though it is not the outermost orbital.
31. (a) $3s$
(b) $6f$
(c) $3d$
32. $7.80 \times 10^{-19} \text{ J}$
33. $[\text{Ar}]3d^{10}4s^24p^55s^1$
34. same number, shape, and capacity for electrons; different energy and size
35. (a) The negative sign indicates that the electron has an “energy debt” and must be given this much energy to become free from the attraction of the nucleus.
(b) $-5.45 \times 10^{-19} \text{ J}$
(c) same since in hydrogen, $2s$ and $2p$ have same energy
36. (a) Ba^{2+}
(b) N^{3-}
(c) Se^{2-}
(d) N^{1-}
37. (a) As, Sb, Sn
(b) Ba, Sr, Mg
(c) Te, I, Br

Chapter 4

38. (a) tetrahedral
(b) T-shaped
(c) see-saw

39. (a) polar
(b) polar
(c) non-polar (linear)
(d) non-polar (linear)
40. central N atom with a LP, one N=O and one N–O bond, bent, 2 resonance structures
41. PO_4^{3-} , tetrahedral, one P=O, and 3 P–O bonds;
 HPO_4^{2-} , tetrahedral, one P=O, 2 P–O, one P–OH bond
42. Both decrease because of greater E_K
43. Must know number of electrons (dispersion forces), polarity of bonds, shape of molecule (dipole-dipole), presence of H–F, H–O and H–N (hydrogen bonding) to compare relative strengths of intermolecular bonds.
44. CH_3F tetrahedral, no hydrogen bonding; NH_2F , pyramidal, hydrogen bonding; dispersion forces are equal since number of electrons is the same in each molecule
45. (a) ionic
(b) network
(c) metal
(d) molecular
46. Sodium halides: inter-ionic distance increases and strength of ionic bond decreases. Hydrogen halides: as number of electrons increases, dispersion forces increase
47. linear, $\text{I}-\text{I}-\text{I}^-$
48. tetrahedral
49. PbO_2 linear non-polar; PbF_4 tetrahedral, non-polar
50. bent, polar, 2 resonance structures

UNIT 3

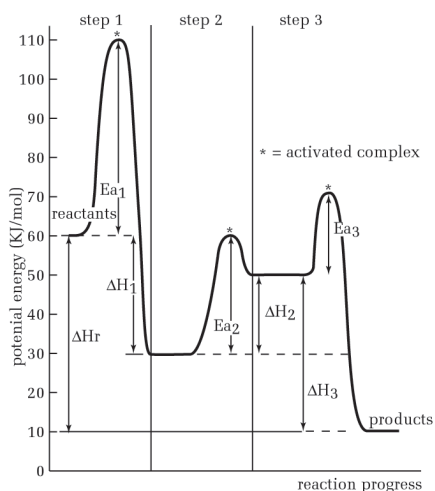
Chapter 5

51. $\Delta H_f^\circ: 2\text{C}_{\text{graphite}} + 3\text{H}_2(\text{g}) + 1/2\text{O}_2(\text{g}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{l})$
 $\Delta H_{\text{comb}}: \text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$
52. (a) same
(b) glass
(c) metal is a better conductor of heat and will feel hotter
53. (a) $0.55 \text{ J/g}^\circ\text{C}$
(b) heat loss to surroundings, oil remains on the metal, metal not at same final temperature as the water
54. (a) $1/2\text{H}_2(\text{g}) + 1/2\text{I}_2(\text{g}) + 25.9 \text{ kJ} \rightarrow \text{HI}(\text{g})$
(b) HI
(c) more energy needed to break bonds in $\text{H}_2(\text{g})$ and $\text{I}_2(\text{g})$ than is liberated when H–I bonds form.
55. (a) -1531.2 kJ
(b) -67.8 kJ
(c) 906.0 kJ
56. 55.2 kJ
57. 23.9 kJ
58. $+86 \text{ kJ}$
59. (a) $\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) + 337 \text{ kJ}$
(b) 24.6 kg
60. -96.7 kJ

61. -521 kJ/mol
 62. (a) $\text{C}_7\text{H}_{8(l)} + 9 \text{O}_{2(g)} \rightarrow 7 \text{CO}_{2(g)} + 4 \text{H}_2\text{O}_{(g)}$
 (b) $+182.3 \text{ kJ/mol}$
 63. 73.3%
 64. (a) $1.56 \text{ mol/min } [\text{O}_2]$; $0.960 \text{ mol/min } [\text{CO}_2]$
 (b) 1.08 min

Chapter 6

65. (a) $1/2$
 (b) 4
 66. (a) 2
 (b) quickly
 (c) 2
 (d) no
 67. (a) 2nd order wrt $[\text{ClO}_2]$, 1st order wrt $[\text{OH}^-]$
 (b) 3rd order
 (c) $\text{rate} = k[\text{ClO}_2]^2[\text{OH}^-]$
 (d) $k = 120 \text{ L}^2/\text{mol}^2\cdot\text{s}$
 68. (a) $\text{rate} = k[\text{B}]$
 (b) $k = 0.16 \text{ s}^{-1}$
 (c) first order, trimolecular
 69. (a) 1st order
 (b) 12 h
 (c) $k = 1.6 \times 10^{-5} \text{ s}^{-1}$
 70. $4.2 \times 10^{-9} \text{ mol/L}\cdot\text{s}$
 71. (a) more molecules have energy $> E_a$
 (b) $40^\circ\text{C} = 4$ times, $80^\circ\text{C} = 64$ times
 72. surface area increases from 6 cm^2 to 10 cm^2 , rate increases by $10/6$, time decreases by $6/10 = 2.4$ min
 73. $\text{H}_2(g) + 2\text{NO}(g) \rightarrow \text{H}_2\text{O}(g) + \text{N}_2\text{O}(g)$; rate law of this elementary rate determining step is same as overall rate law equation
 74. (a) $\text{CH}_3\text{I}(g)$, $\text{HI}(g)$
 (b) step 1
 (c) catalyst
 (d) $\text{CH}_3\text{CHO}(g) \rightarrow \text{CH}_4(g) + \text{CO}(g)$
 75. $\Delta H_{\text{rxn}} = -50 \text{ kJ}$


UNIT 4
Chapter 7

76. (a) no change all components
 (b) $\text{S}_{8(s)}$ no change, O_2 decrease, SO_2 increase
 (c) no change all components
 (d) $\text{S}_{8(s)}$ no change, O_2 increase, SO_2 increase
 77. (a) 5.0
 (b) 1.5
 (c) $\text{CO}_{2(g)} + \text{H}_2(g) + \text{heat} \leftrightarrow \text{H}_2\text{O}(g) + \text{CO}(g)$
 78. 1.30 g
 79. coefficients in balanced equation appear in expression for K_c
 80. $[\text{I}_2] = 0.180 \text{ mol/L}$, $[\text{I}] = 2.62 \times 10^{-3} \text{ mol/L}$
 81. (a) increase in CO_2 , shift left; increase in temperature, shift right
 (b) not a closed system in usual sense, effects will not occur
 82. entropy decreasing; ΔH is negative, ΔG negative at low temperatures

Chapter 8

83. $[\text{H}^+] = 4.79 \times 10^{-8} \text{ mol/L}$; for aqueous solution, $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$
 84. $[\text{Fe}(\text{OH})_2]^{2+}_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow [\text{Fe}(\text{OH})_2\text{OH}]^+_{(aq)} + \text{H}_3\text{O}^+_{(aq)}$
 85. for acid properties use salt of a weak base and a strong acid, NH_4Cl ; for basic properties, use salt of a strong base and a weak acid, Na_2CO_3
 86. (a) $\text{HSO}_4^-_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{H}_3\text{O}^+_{(aq)} + \text{SO}_4^{2-}_{(aq)}$
 (b) $\text{pH} = 1.40$
 87. $\text{pH} = 11.6$
 88. CH_3COO^- , 5.6×10^{-10}
 89. (a) $\text{HN}_3_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{H}_3\text{O}^+_{(aq)} + \text{N}_3^-_{(aq)}$
 (b) N_3^- (c) 2.78, 5.34
 90. 10.5
 91. 0.200 mol/L
 92. conjugate acid H_2PO_4^- , conjugate base PO_4^{3-} ; depends on value of K_a of ion with which it reacts

Chapter 9

93. Ag_2CrO_4 : $6.54 \times 10^{-5} \text{ mol/L}$; AgCl : $1.33 \times 10^{-5} \text{ mol/L}$
 94. (a) an ion from another source is present in the equilibrium
 (b) $[\text{OH}^-]$ term is squared and has a greater affect
 95. $9.24 \times 10^{-5} \text{ g}$
 96. 1.05×10^{-6}
 (b) larger because solubility is higher at higher temperature
 97. $Q_{\text{sp}} = 1.6 \times 10^{-8}$, precipitate
 98. do not want radioactive material absorbed
 99. 0.099 mol
 100. (a) yes, CaF_2 has low solubility
 (b) $1.95 \times 10^{-6} \text{ mol/L}$

UNIT 5**Chapter 10****101.** (c), (d)**102. (a)** left side: S₈(0), Na(+1) S(+4) O(-2); right side: Na(+1), S(+2), O(-2)**(b)** S in Na₂SO₃**(c)** S₈**(d)** Na⁺**103. (a)** -2; 0; +4; +2; +5/2**(b)** +3; +5; +3; +3; +3**104.** N₂O₅(+5), N₂O₄(+4), NO₂(+4), N₂O₃(+3), NO(+2), N₂O(+1)**105.**(c) and (e) are true**106.(a)** S changes oxidation number from +4 in SF₄ to +2 in SCl₂ and Cl changes from -1 in BCl₃ to 0 in Cl₂**(b)** S in SF₄**(c)** In assigning oxidation numbers in covalent molecules the bond is treated as ionic; both "F⁻" and "Cl⁻" exist on the left side but it is the Cl that undergoes oxidation.**107.** In < Ga < Mn < Np**108. (a)** 3Ti³⁺ + RuCl₅²⁻ + 6OH⁻ → Ru + 3TiO²⁺ + 5Cl⁻ + 3H₂O**(b)** 2ClO₂ + 2OH⁻ → ClO₂⁻ + ClO₃⁻ + H₂O**109. (a)** 2NO₂ + 2OH⁻ → NO₂⁻ + NO₃⁻ + H₂O**(b)** CrO₄⁻ + 2HSnO₂⁻ → CrO₂⁻ + 2HSnO₃⁻**(c)** 8Al + 3NO₃⁻ + 5OH⁻ + 18H₂O → 8Al(OH)₄⁻ + 3NH₃**110. (a)** 5ClO₃⁻ + 3I₂ + 3H₂O → 6IO₃⁻ + 5Cl⁻ + 6H⁺**(b)** 5C₂H₄ + 12MnO₄⁻ + 36H⁺ → 10CO₂ + 12Mn²⁺ + 28H₂O**(c)** Cu + SO₄²⁻ + 4H⁺ → Cu²⁺ + SO₂ + 2H₂O**111. (a)** 3Se + 4NO₃⁻ + 4H⁺ → 3SeO₂ + 4NO + 2H₂O**(b)** 6Ag + Cr₂O₇²⁻ + 14H⁺ → 6Ag⁺ + 2Cr³⁺ + 7H₂O**112. (a)** 3P₄ + 20NO₃⁻ + 20H⁺ + 8H₂O → 12H₃PO₄ + 20NO**(b)** MnO₂ + NO₂⁻ + 2H⁺ → NO₃⁻ + Mn²⁺ + H₂O**(c)** TeO₃²⁻ + 2N₂O₄ + 2OH⁻ → Te + 4NO₃⁻ + H₂O**(d)** 4MnO₄⁻ + 3N₂H₄ → 4MnO₂ + 3N₂ + 4H₂O + 4OH⁻**(e)** S₂O₃²⁻ + 4OCl⁻ + 2OH⁻ → 2SO₄²⁻ + 4Cl⁻ + H₂O**(f)** Br₂ + SO₂ + 2H₂O → 2Br⁻ + SO₄²⁻ + 4H⁺**(g)** PbO₂ + 4Cl⁻ + 4H⁺ → PbCl₂ + Cl₂ + 2H₂O**(h)** 2MnO₄⁻ + 5H₂O₂ + 6H⁺ → 2Mn²⁺ + 5O₂ + 8H₂O**113. (a)** 5CH₃OH + 4MnO₄⁻ + 12H⁺ → 5HCOOH + 4Mn²⁺ + 11H₂O**(b)** 40 e**(c)** 3.33 L**Chapter 11****114. (a)** Pt**(b)** Al to Pt**(c)** Ce⁴⁺**(d)** decrease**115. (a)** MnO₄⁻ since the half-cell potential is more positive than for Cr₂O₇²⁻.**(b)** Fe and Ag**116.** -2.50 V**117. (a)** Sn²⁺ + Fe → Fe²⁺ + Sn**(b)** 0.310 V**(c)** Sn**(d)** 0.704 g**118.** H₂ and O₂**119.** No, since the combination of these two half reactions gives a positive E^ocell**120.** Yes, since E^ocell = 0.14 V**121.** 1448 C**122. (a)** 0.131 V**(b)** -0.767 V**(c)** -2.786 V**(d)** 0.793 V**123.** 3.56 g**124. (a)** anode 2Cl⁻ - 2e → Cl_{2(g)}; cathode Al³⁺ + 3e → Al_(l)**(b)** No, because the standard half-cell potentials are determined for ions in aqueous solution at a concentration of 1.0 mol/L**125.** Fuel cell: continuous flow of fuel and removal of product; can operate as long as fuel is supplied; voltage 1.2 V; electrodes are inert and do not react in the process; non-polluting.

Dry cell: finite amount of fuel (electrode) to react; voltage 1.2 V - 1.5 V; electrodes are used up; must be disposed of when used up.