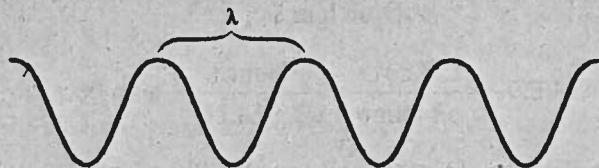


CHAPTER 11

Modern Atomic Theory

CHAPTER ANSWERS

1. positively; negatively
2. Rutherford's experiments determined that the atom had a nucleus containing positively charged particles called protons and neutral particles called neutrons. He established that the nucleus was very small compared to the overall size of the atom. He was not able to determine where the electrons were in the atom or what they were doing.
3. Electromagnetic radiation is radiant energy that travels through space with wavelike behavior at the speed of light.
4. The different forms of electromagnetic radiation are similar in that they all exhibit the same type of wave-like behavior and are propagated through space at the same speed (the speed of light). The types of electromagnetic radiation differ in their frequency (and wavelength) and in the resulting amount of energy carried per photon.
5. See also Figure 11.3 in the text.



6. The *speed* of electromagnetic radiation represents how fast a given wave moves through space. The *frequency* of electromagnetic radiation represents how many complete cycles of the wave pass a given point per second. These two concepts are not the same.
7. Light appears to have both wave and particle properties. It travels through space and carries energy in a regular manner as would be expected for wavelike behavior. Under other circumstances, such as when it is interacting with matter, however, it behaves as if it is a stream of tiny particles of energy.
8. more
9. The atoms in salts emit light of characteristic wavelengths because of the way their internal electron structures interact with the applied energy. Each element's atoms have a different electron structure, so each atom emits light differently from all other atoms.
10. exactly equal to
11. The ground state represents the lowest-energy state of the atom.
12. A photon having an energy corresponding to the energy difference between the two states is emitted by an atom in an excited state when it returns to its ground state.
13. The energy of a photon is given by $E = hc/\lambda$ and therefore short wavelength light carries more energy per photon.

14. higher; lower
15. The emission of light by excited atoms has been the key interconnection between the macroscopic world we can observe and measure and what is happening on a microscopic basis within an atom. Excited atoms emit light (which we can measure) because of changes in the microscopic structure of the atom. By studying the emissions of atoms we can trace back to what happened inside the atom.
16. When excited hydrogen atoms emit their excess energy, the photons of radiation emitted are always of exactly the same wavelength and energy. We consider this to mean that the hydrogen atom possesses only certain allowed energy states and that the photons emitted correspond to the atom changing from one of these allowed energy states to another of the allowed energy states. The energy of the photon emitted corresponds to the energy difference in the allowed states. If the hydrogen atom did not possess discrete energy levels, then we would expect the photons emitted to have random wavelengths and energies.
17. transitions of electrons
18. The energy of an emitted photon is *identical* to the energy change within the atom that gave rise to the emitted photon.
19. Hydrogen always emits light at exactly the same wavelengths that correspond to transitions of the electron between the fixed energy states within the atom.
20. When the hydrogen atoms are excited by the applied voltage, they absorb only photons of energy corresponding to the different energy states within the atom. When the hydrogen atoms relax, they emit photons corresponding to the differences between these same energy states, which correspond to light of only certain characteristic wavelengths.
21. Bohr envisioned electrons as moving in circular orbits corresponding to the various allowed energy levels. He suggested that the electron could jump to a different orbit by absorbing or emitting a photon of light with exactly the correct energy content (corresponding to the difference in energy between the orbits).
22. orbit
23. Bohr suggested that the electron could jump to a different orbit by absorbing or emitting a photon of light with exactly the correct energy content (corresponding to the difference in energy between the orbits).

As the energy levels of a given atom were fixed and definite, then the atom should always emit energy at the same discrete wavelengths.
24. Bohr's theory *explained* the experimentally *observed* line spectrum of hydrogen *exactly*. Bohr's theory was ultimately discarded because when attempts were made to extend the theory to atoms other than hydrogen, the calculated properties did *not* correspond closely to experimental measurements.
25. Schrödinger and de Broglie reasoned that, because light seems to have both wave and particle characteristics (it behaves simultaneously as a wave and as if it were a stream of particles), perhaps the electron might exhibit both of these characteristics. That is, although the electron behaves as a discrete particle, perhaps the properties of the electron in the atom could be treated as if they were wavelike.
26. An orbit represents a definite, exact circular pathway around the nucleus in which an electron can be found. An orbital represents a region of space in which there is a high probability of finding the electron.

27. Schrödinger's mathematical treatment could describe only the movement of the electron through the atom in terms of the *probability* of finding the electron in given regions of space within the atom, but not at a particular point within the atom at a particular time. Any attempt to determine the exact position of the electron within an atom experimentally would, in fact, cause the electron to move from wherever it had been.
28. The firefly analogy is intended to demonstrate the concept of a probability map for electron density. In the wave mechanical model of the atom, we cannot say specifically where the electron is in the atom; we can only say where there is a high probability of finding the electron. The analogy is to imagine a time-exposure photograph of a firefly in a closed room. Most of the time, the firefly will be found near the center of the room.
29. Although the probability of finding the electron decreases at greater distances from the nucleus, the probability of finding it even at great distances from the nucleus never becomes exactly zero. The probability becomes less and less as you move away from the nucleus, similar to the way the atmosphere becomes thinner and thinner as you move away from the surface of the earth.
30. Pictures we draw to represent orbitals should only be interpreted as probability maps. They are not meant to represent that the electron moves only on the surface of, or within, the region drawn in the picture. As the mathematical probability of finding the electron never actually becomes zero on moving outward from the nucleus, scientists have decided that pictures of orbitals should represent a 90% probability that the electron will be found inside the region depicted in the drawing (for 100% probability, the orbital would have to encompass all space).
31. The $2s$ orbital is similar in shape to the $1s$ orbital, but is larger.
32. The p orbitals, in general, have two lobes and are sometimes described as having a "dumbbell" shape. The $2p$ and $3p$ orbitals are similar in shape, and in fact there are three equivalent $2p$ or $3p$ orbitals in the $2p$ or $3p$ subshell. The orbitals differ in size, mean distance from the nucleus, and energy.
33. farther from
34. $n = 1$
35. The other orbitals serve as the excited states of the hydrogen atom. When energy of the right frequency is applied to the hydrogen atom, the electron can move from its normal orbital (ground state) to one of the other orbitals (excited states). Later on, the electron can move back to its normal orbital and release the absorbed energy as light.
- 36.

Value of n	Possible Subshells
1	$1s$
2	$2s, 2p$
3	$3s, 3p, 3d$
4	$4s, 4p, 4d, 4f$

37. electron spin

38. Electrons have an intrinsic spin (they spin on their own axes). Geometrically, there are only two directions possible for spin (clockwise or counter-clockwise). This means only two electrons having opposite senses or directions of spin can occupy an orbital. This idea is called the Pauli Exclusion Principle.
39. The higher the value of the principal quantum number, n , the higher the energy of the principal energy level.
40. increases; As you move out from the nucleus, there is more space and room for more sublevels.
41. two
42. opposite
43. b, c, and d are possible; A is not correct because p subshells do not occur until $n = 2$.
44. Answer depends on student choices.
45. The $1s$ orbital is closest to the nucleus and lowest in energy, so it is always filled first.
46. When a hydrogen atom is in its ground state, the electron is found in the $1s$ orbital. The $1s$ orbital has the lowest energy of all the possible hydrogen orbitals.
47. Valence electrons are those in the outermost (highest) principal energy level of an atom. These electrons are especially important because they are at the "outside edge" of an atom and are those electrons that are "seen" by other atoms and can interact with the electrons of another atom in a chemical reaction.
48. The elements in a given vertical column of the periodic table have the same valence electron configuration. Having the same valence electron configuration causes the elements in a given group to have similar chemical properties.
- 49.
- $1s^2$
 - $1s^2 2s^2 2p^6$
 - $1s^2 2s^2 2p^6 3s^2 3p^6$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
- 50.
- C
 - P
 - Sc
 - Kr
- 51.
- $1s^2 2s^2 2p^6 3s^1$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
 - $1s^2 2s^2 2p^3$
 - $1s^2 2s^2$

52.

- a. Br
- b. Ca
- c. Cl
- d. O

53.

- a. $1s(\uparrow\downarrow)$
- b. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$
- c. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$
- d. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 5s(\uparrow\downarrow) 4d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$
 $5p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$

54.

- a. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow)(\uparrow)(\uparrow)$
- b. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow)(\uparrow)(\uparrow)$
- c. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$
- d. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$

55. For the representative elements (those filling s and p subshells), the group number gives the number of valence electrons.

- a. one ($2s$)
- b. three ($2s, 2p$)
- c. eight ($3s, 3p$)
- d. five ($3s, 3p$)

56. For the representative elements (those filling s and p subshells), the group number gives the number of valence electrons.

- a. five
- b. seven
- c. one
- d. six

57. This belief is based on the *experimental properties* of K and Ca. The physical and chemical properties of K are like those of the other Group 1 elements; Ca's properties are similar to the other Group 2 elements.

58. The properties of Rb and Sr suggest that they are members of Groups 1 and 2, respectively, and so must be filling the $5s$ orbital. The $5s$ orbital is lower in energy (and fills before) the $4d$ orbitals.

- 59.
- $[\text{Kr}] 5s^2 4d^2$
 - $[\text{Ar}] 4s^2 3d^8$
 - $[\text{Ar}] 4s^2 3d^{10} 4p^3$
 - $[\text{Ne}] 3s^2 3p^2$
- 60.
- Mg
 - Rb
 - Sc
 - P
- 61.
- $[\text{Ar}] 4s^2 3d^1$
 - $[\text{Kr}] 5s^2 4d^1$
 - $[\text{Xe}] 6s^2 5d^1$
 - $[\text{Rn}] 7s^2 6d^1$
- 62.
- $[\text{Ne}] 3s^2 3p^3$
 - $[\text{Ne}] 3s^2 3p^5$
 - $[\text{Ne}] 3s^2$
 - $[\text{Ar}] 4s^2 3d^{10}$
- 63.
- one
 - five (configuration is $4s^1 3d^5$ rather than $4s^2 3d^4$)
 - ten
 - two
- 64.
- one
 - two
 - zero
 - ten
65. The *position* of the element (both in terms of vertical column and horizontal row) indicates which set of orbitals is being filled last.
- $5f$
 - $6p$
 - $4d$

- d. $4d$
66. The *position* of the element (both in terms of vertical column and horizontal row) indicates which set of orbitals is being filled last.
- $5f$
 - $5f$
 - $4f$
 - $6p$
- 67.
- $[\text{Xe}] 4f^{14} 5d^2 6s^2$
 - $[\text{Rn}] 7s^2$
 - $[\text{Kr}] 4d^{10} 5s^2 5p^3$
 - $[\text{Xe}] 4f^{14} 5d^{10} 6s^2 6p^2$
- 68.
- Group 7
 - Group 8
 - Group 4
 - Group 5
69. Some typical properties of metals are a lustrous appearance, the ability to be pounded into sheets (malleability) or pulled into wires (ductility), and the ability to conduct heat and electricity. Nonmetals typically have a non-shiny appearance, are brittle, and do not conduct heat or electricity well. There are exceptions to these general properties. For example, graphite (a form of the nonmetal carbon) conducts electricity well.
70. The metallic elements *lose* electrons and form *positive* ions (cations); the nonmetallic elements *gain* electrons and form *negative* ions (anions). Remember that the electron itself is *negatively* charged.
71. The Group 1 metals are all highly reactive, and all form $1+$ ions almost exclusively when they react. Physically, these metals are soft (they can be cut with a knife) and very low in density. Because of their high reactivity, these metals tend to be found with a coating of metal oxide that hides their metallic luster (This luster can be seen, however, if a fresh surface of the metal is exposed.).
72. All exist as *diatomic* molecules (F_2 , Cl_2 , Br_2 , I_2); all are *nonmetals*; all have relatively high electronegativities; all form $1-$ ions in reacting with metallic elements.
73. Of the Group 1 elements, francium would be expected to lose its valence electron most easily, because the valence electron is located farther from the nucleus than in any other Group 1 element.
74. Elements at the *left* of a period (horizontal row) lose electrons more readily. At the left of a period (given principal energy level), the nuclear charge is the smallest, and the electrons are least tightly held.

75. The nonmetallic elements are clustered at the upper right side of the periodic table. These elements are effective at pulling electrons from metallic elements for several reasons. First, these elements have little tendency to lose electrons themselves (they have high ionization energies). Second, the atoms of these elements tend to be small in size, which means that electrons can be pulled in strongly since they can get closer to the nucleus. Finally, if these atoms gain electrons, they can approach the electronic configuration of the following noble gas elements (See Chapter 12 for why the electronic configuration of the noble gases is desirable for other atoms to attain.).
76. The elements of a given period (horizontal row) have valence electrons in the same principal energy level. Nuclear charge, however, increases across a period going from left to right. Atoms at the left side have smaller nuclear charges and hold onto their valence electrons less tightly.
77. Of the elements in Group 2, radium would be expected to have the largest atoms because the valence electrons in Ra are in a higher energy orbital and farther from the nucleus than in any other Group 2 element.
78. The *nuclear charge* increases from left to right within a period, pulling progressively more tightly on the valence electrons.
79. For most elements, the chemical activity is reflected in the ease with which the element gains or loses electrons
- Li (The less reactive metals are further up in a group.)
 - At (The less reactive nonmetals are at the bottom of a group.)
 - Be (The less reactive metals are further up in a group.)
 - Po (The less reactive elements are at the bottom of a group.)
80. Ionization energies decrease in going from top to bottom within a vertical group; ionization energies increase in going from left to right within a horizontal period.
- Li
 - Ca
 - Cl
 - S
81. Atomic size increases in going from top to bottom within a vertical group; atomic size decreases in going from left to right within a horizontal period.
- $Xe < Sn < Sr < Rb$
 - $He < Kr < Xe < Rn$
 - $At < Pb < Ba < Cs$
82. Atomic size increases in going from top to bottom within a vertical group; atomic size decreases in going from left to right within a horizontal period.
- Na
 - S
 - N
 - F

83. Blue light has a shorter wavelength than red light. Therefore, when transferring energy, more waves reach a given point per second.

84. speed of light

85. visible

86. photons

87. ground

88. quantized

89. orbits

90. orbital

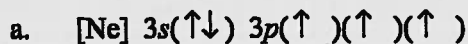
91. valence

92. transition metal

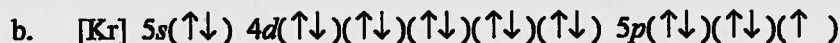
93. frequency

94. spins

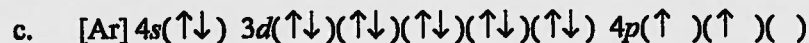
95.



P is expected to be paramagnetic; three unpaired 3p electrons

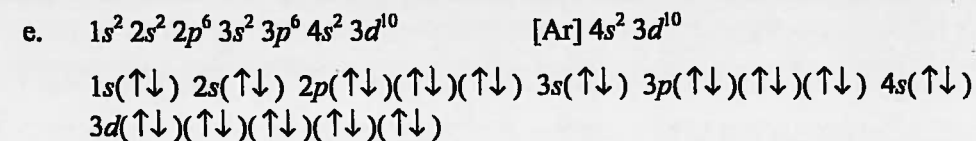
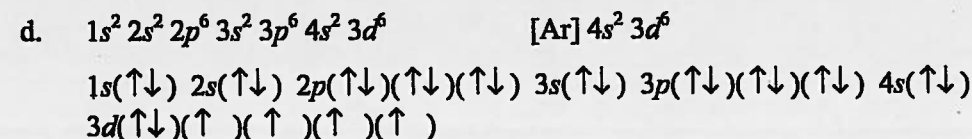
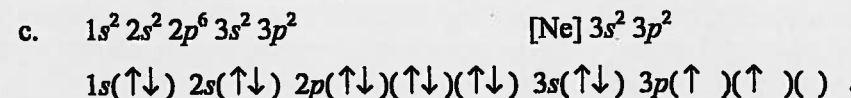
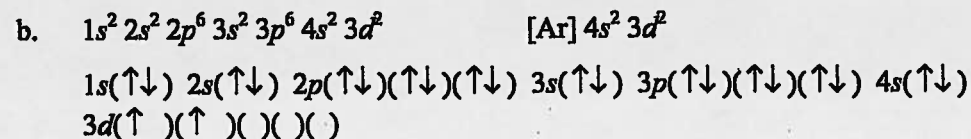
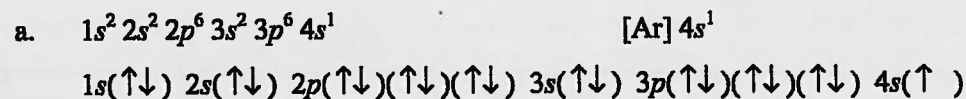


I is expected to be paramagnetic; one unpaired 5p electron



Ge is expected to be paramagnetic; two unpaired 4p electrons

96.



97.

- a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
 $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow)() () ()$
 [Ar] $4s^2 3d^1$
- b. $1s^2 2s^2 2p^6 3s^2 3p^3$
 $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow)(\uparrow)(\uparrow)$
 [Ne] $3s^2 3p^3$
- c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
 $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$
 [Kr] is itself a noble gas
- d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$
 $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 5s(\uparrow\downarrow)$
 [Kr] $5s^2$
- e. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
 $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow)$
 $3d(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$
 [Ar] $4s^2 3d^{10}$

98.

- a. ns^2
 b. $ns^2 np^5$
 c. $ns^2 np^4$
 d. ns^1
 e. $ns^2 np^4$

99.

- a. four (two if the d electrons are not counted.)
 b. seven
 c. two
 d. seven (two if the d electrons are not counted.)

100.

a. $\lambda = \frac{h}{mv}$

$$\lambda = \frac{6.63 \times 10^{-34} \text{ J s}}{(9.1 \times 10^{-31} \text{ kg})(0.90 \times (3.00 \times 10^8 \text{ ms}^{-1}))}$$

$$\lambda = 2.7 \times 10^{-12} \text{ m (0.0027 nm)}$$

b. $4.4 \times 10^{-34} \text{ m}$

c. $2 \times 10^{-35} \text{ m}$

The wavelengths for the ball and the person are *infinitesimally small* whereas the wavelength for the electron is nearly the same order of magnitude as the diameter of a typical atom.

101. $3.00 \times 10^8 \text{ m/sec}$

102. Light is emitted from the hydrogen atom only at certain fixed wavelengths. If the energy levels of hydrogen were *continuous*, a hydrogen atom would emit energy at all possible wavelengths.

103. As an electron moves to a higher-number principal energy level, the electron's mean distance from the nucleus increases, thereby decreasing the attractive force between the electron and the nucleus.

104. The third principal energy level of hydrogen is divided into *three* sublevels (*3s*, *3p*, and *3d*); there is a *single* *3s* orbital; there is a set of *three* *3p* orbitals; there is a set of *five* *3d* orbitals. See Figures 11.24–11.28 for the shapes of these orbitals.

105. Orbitals in the *1s* and *2s* subshells can contain only two electrons.

106.

- a. incorrect; the $n = 1$ shell has only the *1s* subshell
- b. correct
- c. incorrect; the $n = 3$ shell has only *3s*, *3p*, and *3d* subshells
- d. correct
- e. correct
- f. correct

107. Electrons are negatively charged particles that repel each other. By placing the three electrons in separate *2p* orbitals (oriented at 90° to each other in space), the repulsion among the electrons is minimized. The electrons also can have the same spin if they are in separate orbitals, but this is not discussed at length in the text.

108.

- a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
- b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$
- c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$
- d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

109.

- a. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow\downarrow) 3d(\uparrow)(\uparrow)(\uparrow)(\uparrow)(\uparrow)$
- b. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow)(\uparrow)$
- c. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 3s(\uparrow\downarrow) 3p(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) 4s(\uparrow)$
- d. $1s(\uparrow\downarrow) 2s(\uparrow\downarrow) 2p(\uparrow)(\uparrow)(\uparrow)$

110.

- a. five ($2s, 2p$)
- b. seven ($3s, 3p$)
- c. one ($3s$)
- d. three ($3s, 3p$)

111. transition metals

112.

- a. $[\text{Kr}] 5s^2 4d^2$
- b. $[\text{Kr}] 5s^2 4d^{10} 5p^5$
- c. $[\text{Ar}] 4s^2 3d^{10} 4p^2$
- d. $[\text{Xe}] 6s^1$

113.

- a. $[\text{Ar}] 4s^2 3d^2$
- b. $[\text{Ar}] 4s^2 3d^{10} 4p^4$
- c. $[\text{Kr}] 5s^2 4d^{10} 5p^3$
- d. $[\text{Kr}] 5s^2$

114.

- a. Se
- b. Se
- c. Rb
- d. V

115.

- a. $[\text{Ar}] 4s^2 3d^8$
- b. $[\text{Kr}] 5s^2 4d^3$ (actually $[\text{Kr}] 5s^1 4d^4$ for reasons beyond text)
- c. $[\text{Xe}] 6s^2 4f^{14} 5d^2$
- d. $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^5$

116. metals, low; nonmetals, high

117.

- a. B and Al are both very reactive
- b. Na
- c. F

118. Atomic size increases in going from top to bottom within a vertical group; atomic size decreases in going from left to right within a horizontal period.

- a. Ca
- b. P
- c. K