CHAPTER 5

Assessment

Reviewing Content

5.1 Models of the Atom

22. What was inadequate about Rutherford's model of the atom? Which subatomic particles did Thomson include in the plum-pudding model of the atom?
23. What did Bohr assume about the motion of electrons?
24. Describe Rutherford's model of the atom and compare it with the model proposed by his student Niels Bohr.
25. What is the significance of the boundary of an electron cloud?
26. What is an atomic orbital?
27. How many orbitals are in the $2p$ sublevel?
28. Sketch $1s$, $2s$, and $2p$ orbitals using the same scale for each.
29. How many sublevels are contained in each of these principal energy levels?
   a. $n = 1$ **b.** $n = 2$ **c.** $n = 3$ **d.** $n = 4$

5.2 Electron Arrangement in Atoms

30. How many electrons are in the highest occupied energy level of these atoms?
   a. barium **b.** sodium
   c. aluminum **d.** oxygen
31. What are the three rules that govern the filling of atomic orbitals by electrons?
32. Write electron configurations for the elements that are identified only by these atomic numbers.
   a. 15 **b.** 12 **c.** 9 **d.** 18
33. What is meant by $3p^3$?
34. Give electron configurations for atoms of these elements:
   a. Na **b.** S **c.** Mg **d.** Ne **e.** K
35. Which of these orbital designations are invalid?
   a. $4s$ **b.** $3f$ **c.** $2d$ **d.** $3d$
36. What is the maximum number of electrons that can go into each of the following sublevels?
   a. $2s$ **b.** $3p$ **c.** $4s$ **d.** $3d$
   e. $4p$ **f.** $5s$ **g.** $4f$ **h.** $5p$
37. Arrange the following sublevels in order of increasing energy:
   $3d$, $2s$, $4s$, $3p$.
38. How many electrons are in the second energy level of an atom of each element?
   a. chlorine **b.** phosphorus **c.** potassium
39. Write electron configurations for atoms of these elements.
   a. selenium **b.** vanadium
   c. nickel **d.** calcium

5.3 Physics and the Quantum Mechanical Model

40. List the colors of the visible spectrum in order of increasing wavelength.
41. What is meant by the frequency of a wave? What are the units of frequency? Describe the relationship between frequency and wavelength.
42. Use a diagram to illustrate each term for a wave.
   a. wavelength
   b. amplitude
   c. cycle
43. Explain the difference between the energy lost or gained by an atom according to the laws of classical physics and according to the quantum model of an atom.
44. How are ultraviolet radiation and microwave radiation the same? How are they different?
45. Consider the following regions of the electromagnetic spectrum: (i) ultraviolet, (ii) X-ray, (iii) visible, (iv) infrared, (v) radio wave, (vi) microwave.
   a. Use Figure 5.10 to arrange them in order of decreasing wavelength.
   b. How does this order differ from that of decreasing frequency?
46. List one way in which each of the radiations listed in Question 45 is used.
47. What happens when a hydrogen atom absorbs a quantum of energy?
48. When white light is viewed through sodium vapor in a spectroscope, the spectrum is continuous except for a dark line at 589 nm. How can you explain this observation?
49. The transition of electrons from higher energy levels to the $n = 2$ energy level results in the emission of light from hydrogen atoms. In what part of the spectrum is the emitted light, and what is the name given to this transition series?

Assessment *continued*

Understanding Concepts

50. Give the symbol for the atom that corresponds to each electron configuration.
a. $1s^22s^22p^63s^23p^6$
b. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^75s^1$
c. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^75s^25p^65d^16s^2$

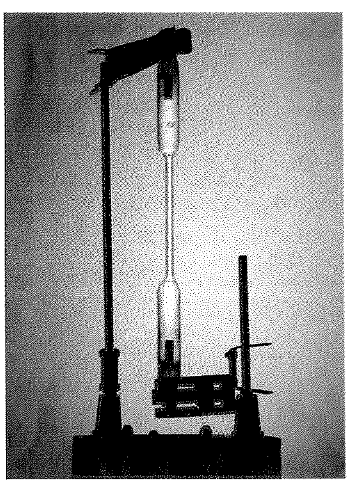
51. Write the electron configuration for an arsenic atom. Calculate the total number of electrons in each energy level and state which energy levels are not full.

52. How many paired electrons are there in an atom of each element?
a. helium **b.** boron
c. sodium **d.** oxygen

53. An atom of an element has two electrons in the first energy level and five electrons in the second energy level. Write the electron configuration for this atom and name the element. How many unpaired electrons does an atom of this element have?

54. Suppose your favorite AM radio station broadcasts at a frequency of 1150 kHz. What is the wavelength, in centimeters, of the radiation from the station?

55. A mercury lamp, such as the one below, emits radiation with a wavelength of 4.36×10^{-7} m.



a. What is the wavelength of this radiation in centimeters?
b. In what region of the electromagnetic spectrum is this radiation?
c. Calculate the frequency of this radiation.

56. Sodium vapor lamps are used to illuminate streets and highways. The very bright light emitted by these lamps is actually due to two closely spaced emission lines in the visible region of the electromagnetic spectrum. One of these lines has a wavelength of 5.890×10^{-7} m, and the other line has a wavelength of 5.896×10^{-7} m.
a. What are the wavelengths of these radiations in centimeters?
b. Calculate the frequencies of these radiations.
c. In what region of the visible spectrum do these lines appear?

57. Give the symbol and name of the elements that correspond to these configurations of an atom.
a. $1s^22s^22p^63s^1$
b. $1s^22s^22p^3$
c. $1s^22s^22p^63s^23p^2$
d. $1s^22s^22p^4$
e. $1s^22s^22p^63s^23p^64s^1$
f. $1s^22s^22p^63s^23p^63d^24s^2$

58. Describe how the wavelength changes if the frequency of a wave is multiplied by 1.5.

59. State the Heisenberg uncertainty principle.

60. What is the maximum number of electrons that can be found in any orbital of an atom?

61. Pieces of energy are known as
a. isotopes **b.** particles
c. quanta **d.** line spectra

62. The lowest sublevel in each principal energy level is represented by the symbol
a. f **b.** p **c.** s **d.** d

63. Which electron transition results in the emission of energy?
a. $3p$ to $3s$
b. $3p$ to $4p$
c. $2s$ to $2p$
d. $1s$ to $2s$

64. Which is the ground state configuration of a magnesium atom?
a. $1s^22s^22p^63s^2$
b. $1s^22s^22p^63s^1$
c. $1s^22s^23s^22p^6$
d. $1s^22s^22p^43s^2$

Critical Thinking

65. Explain the difference between an orbit in the Bohr model and an orbital in the quantum mechanical model of the atom.

66. Traditional cooking methods make use of infrared radiation (heat). Microwave radiation cooks food faster. Could radio waves be used for cooking? Explain.

67. Think about the currently accepted models of the atom and of light. In what ways do these models seem strange to you? Why are these models not exact or definite?

68. Orbital diagrams for the ground states of two elements are shown below. Each diagram shows something that is incorrect. Identify the error in each diagram and then draw the correct diagram.

a. Nitrogen $1s$ $2s$ $2p$

[↑↓] [↑↓] [↑↓][↑][ ]

b. Magnesium $1s$ $2s$ $2p$ $3s$

[↑↓] [↑↓] [↑↓][↑↓][↑↓] [ ]

69. Picture two hydrogen atoms. The electron in the first hydrogen atom is in the $n = 1$ level. The electron in the second atom is in the $n = 4$ level.

a. Which atom has the ground state electron configuration?

b. Which atom can emit electromagnetic radiation?

c. In which atom is the electron in a larger orbital?

d. Which atom has the lower energy?

70. Identify the elements whose electrically neutral atoms have the following electron configurations.

a. $1s^2\ 2s^22p^5$

b. $1s^2\ 2s^22p^63s^23p^64s^23d^{10}4p^2$

c. $1s^2\ 2s^22p^63s^23p^63d^34s^2$

71. Which of the following is the ground state of an atom? Which is its excited state? Which is an impossible electron configuration? Identify the element and briefly explain your choices.

a. $1s^2\ 2s^22p^63s^23p^65p^1$

b. $1s^2\ 2s^22p^63s^23p^64s^1$

c. $1s^2\ 2s^22p^63s^23p^7$

72. Why do electrons occupy equal energy orbitals singly before beginning to pair up?

Concept Challenge

73. The energy of a photon is related to its frequency and its wavelength.

Energy of photon (J)	Frequency (s^{-1})	Wavelength (cm)
3.45×10^{-21}	a. ______	5.77×10^{-3}
2.92×10^{-20}	b. ______	6.82×10^{-4}
6.29×10^{-20}	c. ______	3.16×10^{-4}
1.13×10^{-19}	d. ______	1.76×10^{-4}
1.46×10^{-19}	e. ______	1.36×10^{-4}
3.11×10^{-19}	f. ______	6.38×10^{-5}

a. Complete the table above.

b. Plot the energy of the photon (y-axis) versus the frequency (x-axis).

c. Determine the slope of the line.

d. What is the significance of this slope?

74. The average distance between Earth and Mars is about 2.08×10^8 km. How long does it take to transmit television pictures from the Mariner spacecraft to Earth from Mars?

75. Bohr's atomic theory can be used to calculate the energy required to remove an electron from an orbit of a hydrogen atom or an ion containing only one electron. This is the ionization energy for that atom or ion. The formula for determining the ionization energy (E) is

$$E = Z^2 \times \frac{k}{n^2}$$

where Z is the atomic number, k is 2.18×10^{-18} J, and n is the energy level. What is the energy required to eject an electron from a hydrogen atom when the electron is in the ground state ($n = 1$)? In the second energy level? How much energy is required to eject a ground state electron from the species Li^{+}?

76. The energy (E) of a photon absorbed or emitted by a body is proportional to its frequency (ν).

$$E = h \times \nu$$

The constant h equals 6.63×10^{-34} J·s. What is the energy of a photon of microwave radiation with a frequency of $3.20 \times 10^{11}\ s^{-1}$?

Assessment *continued*

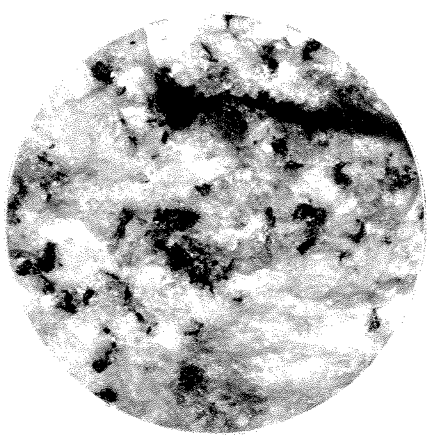
Cumulative Review

77. Classify each of the following as homogeneous or heterogeneous (*Chapter 2*)
a. a page of this textbook
b. a banana split
c. the water in bottled water

78. Hamburger undergoes a chemical change when cooked on a grill. All chemical changes are subject to the law of conservation of mass. Yet a cooked hamburger will invariably weigh less than the uncooked meat patty. Explain. (*Chapter 2*)

79. Homogeneous mixtures and compounds are both composed of two or more elements. How do you distinguish between a homogeneous mixture and a compound? (*Chapter 2*)

80. The photo shows a magnified view of a small piece of granite. Is granite a substance or a mixture? (*Chapter 2*)



81. The diameter of a carbon atom is 77 pm. Express this measurement in μm. (*Chapter 3*)

82. A silver bar has a mass of 368 g. What is the volume, in cm^3, of the bar? The density of silver is 19.5 g/cm^3. (*Chapter 3*)

83. Which has more mass, a 28.0-cm^3 piece of lead or a 16.0–cm^3 piece of gold? The density of lead is 11.4 g/cm^3; the density of gold is 19.3 g/cm^3. (*Chapter 3*)

84. Express the following measurements in scientific notation. (*Chapter 3*)
a. 0.000 039 kg **b.** 784 L
c. 0.0830 g **d.** 9 700 000 ng

85. Which of these quantities or relationships are exact? (*Chapter 3*)
a. 10 cm = 1 dm
b. 9 baseball players on the field
c. a diamond has a mass of 12.4 g
d. the temperature is 21 °C

86. A one-kilogram steel bar is brought to the moon. How are its mass and its weight each affected by this change in location? Explain. (*Chapter 3*)

87. When a piece of copper with a mass of 36.4 g is placed into a graduated cylinder containing 20.00 mL of water, the water level rises to 24.08 mL. What is the density of copper? (*Chapter 3*)

88. The density of gold is 19.3 g/cm^3. What is the mass, in grams, of a cube of gold 2.00 cm on each edge? In kilograms? (*Chapter 3*)

89. A balloon filled with helium will rise upward when released. What does this show about the relative density of helium and air? Explain. (*Chapter 3*)

90. Give the number of protons and electrons in each of the following. (*Chapter 4*)
a. Cs **b.** Ag
c. Cd **d.** Se

91. Explain the difference between the accuracy of a measurement and the precision of a measurement.

92. Which of these was an essential part of Dalton's atomic model? (*Chapter 4*)
a. indivisible atoms
b. electrons
c. atomic nuclei
d. neutrons

93. How do neon-20 and neon-21 differ from each other? (*Chapter 4*)

94. The mass of an atom should be very nearly the sum of the masses of its protons and neutrons. The mass of a proton and the mass of a neutron are each very close to 1 amu. Why is the atomic mass of chlorine, 35.453 amu, so far from a whole number? (*Chapter 4*)