Section 5.2 Quantum Theory and the Atom

You now know that the behavior of light can be explained only by a dual wave-particle model. Although this model was successful in accounting for several previously unexplainable phenomena, an understanding of the relationships among atomic structure, electrons, and atomic emission spectra still remained to be established.

Bohr Model of the Atom

Recall that hydrogen's atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are elements' atomic emission spectra discontinuous rather than continuous? Niels Bohr, a young Danish physicist working in Rutherford's laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. Impressively, Bohr's model also correctly predicted the frequencies of the lines in hydrogen's atomic emission spectrum.

Energy states of hydrogen Building on Planck's and Einstein's concepts of quantized energy (quantized means that only certain values are allowed), Bohr proposed that the hydrogen atom has only certain allowable energy states. The lowest allowable energy state of an atom is called its **ground state.** When an atom gains energy, it is said to be in an excited state. And although a hydrogen atom contains only a single electron, it is capable of having many different excited states.

Bohr went even further with his atomic model by relating the hydrogen atom's energy states to the motion of the electron within the atom. Bohr suggested that the single electron in a hydrogen atom moves around the nucleus in only certain allowed circular orbits. The smaller the electron's orbit, the lower the atom's energy state, or energy level. Conversely, the larger the electron's orbit, the higher the atom's energy state, or energy level. Bohr assigned a quantum number, n, to each orbit and even calculated the orbit's radius. For the first orbit, the one closest to the nucleus, $n = 1$ and the orbit radius is 0.0529 nm; for the second orbit, $n = 2$ and the orbit radius is 0.212 nm; and so on. Additional information about Bohr's description of hydrogen's allowable orbits and energy levels is given in **Table 5-1**.

Table 5-1

Bohr's Description of the Hydrogen Atom

Bohr atomic orbit	Quantum number	Orbit radius (nm)	Corresponding atomic energy level	Relative energy
First	$n = 1$	0.0529	1	E_1
Second	$n = 2$	0.212	2	$E_2 = 4E_1$
Third	$n = 3$	0.476	3	$E_3 = 9E_1$
Fourth	$n = 4$	0.846	4	$E_4 = 16E_1$
Fifth	$n = 5$	1.32	5	$E_5 = 25E_1$
Sixth	$n = 6$	1.90	6	$E_6 = 36E_1$
Seventh	$n = 7$	2.59	7	$E_7 = 49E_1$

Objectives

- **Compare** the Bohr and quantum mechanical models of the atom.
- **Explain** the impact of de Broglie's wave-particle duality and the Heisenberg uncertainty principle on the modern view of electrons in atoms.
- **Identify** the relationships among a hydrogen atom's energy levels, sublevels, and atomic orbitals.

Vocabulary

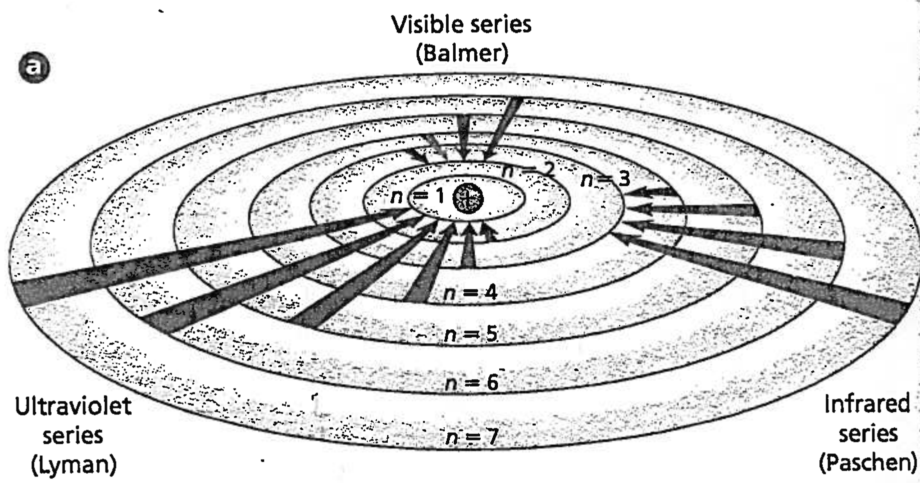
ground state
de Broglie equation
Heisenberg uncertainty principle
quantum mechanical model of the atom
atomic orbital
principal quantum number
principal energy level
energy sublevel

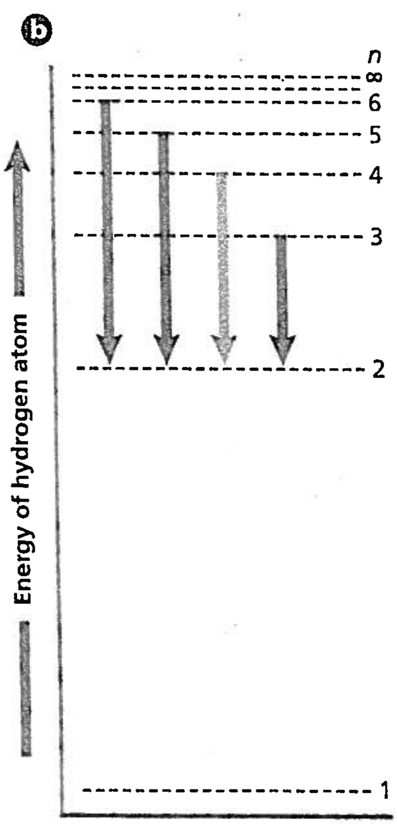
An explanation of hydrogen's line spectrum Bohr suggested that the hydrogen atom is in the ground state, also called the first energy level, when the electron is in the $n = 1$ orbit. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron moves to a higher-energy orbit such as the $n = 2$ orbit shown in **Figure 5-10a**. Such an electron transition raises the atom to an excited state. When the atom is in an excited state, the electron can drop from the higher-energy orbit to a lower-energy orbit. As a result of this transition, the atom emits a photon corresponding to the difference between the energy levels associated with the two orbits.

$$\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu$$

Figure 5-10

ⓐ When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon with a specific energy is emitted. Although hydrogen has spectral lines associated with higher energy levels, only the visible, ultraviolet, and infrared series of spectral lines are shown in this diagram. ⓑ The relative energies of the electron transitions responsible for hydrogen's four visible spectral lines are shown. Note how the energy levels become more closely spaced as n increases.





Note that because only certain atomic energies are possible, only certain frequencies of electromagnetic radiation can be emitted. You might compare hydrogen's seven atomic orbits to seven rungs on a ladder. A person can climb up or down the ladder only from rung to rung. Similarly, the hydrogen atom's electron can move only from one allowable orbit to another, and therefore, can emit or absorb only certain amounts of energy.

The four electron transitions that account for visible lines in hydrogen's atomic emission spectrum are shown in **Figure 5-10b**. For example, electrons dropping from the third orbit to the second orbit cause the red line. Note that electron transitions from higher-energy orbits to the second orbit account for all of hydrogen's visible lines. This series of visible lines is called the Balmer series. Other electron transitions have been measured that are not visible, such as the Lyman series (ultraviolet) in which electrons drop into the $n = 1$ orbit and the Paschen series (infrared) in which electrons drop into the $n = 3$ orbit. **Figure 5-10b** also shows that unlike rungs on a ladder, the hydrogen atom's energy levels are not evenly spaced. You will be able to see in greater detail how Bohr's atomic model was able to account for hydrogen's line spectrum by doing the **problem-solving LAB** later in this chapter.

Bohr's model explained hydrogen's observed spectral lines remarkably well. Unfortunately, however, the model failed to explain the spectrum of any other element. Moreover, Bohr's model did not fully account for the chemical behavior of atoms. In fact, although Bohr's idea of quantized energy levels laid the groundwork for atomic models to come, later experiments demonstrated that the Bohr model was fundamentally incorrect. The movements of electrons in atoms are not completely understood even now; however, substantial evidence indicates that electrons do not move around the nucleus in circular orbits.

The Quantum Mechanical Model of the Atom

Scientists in the mid-1920s, by then convinced that the Bohr atomic model was incorrect, formulated new and innovative explanations of how electrons are arranged in atoms. In 1924, a young French graduate student in physics named Louis de Broglie (1892–1987) proposed an idea that eventually accounted for the fixed energy levels of Bohr's model.

Electrons as waves De Broglie had been thinking that Bohr's quantized electron orbits had characteristics similar to those of waves. For example, as **Figure 5-11b** shows, only multiples of half-wavelengths are possible on a plucked guitar string because the string is fixed at both ends. Similarly, de Broglie saw that only whole numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in **Figure 5-11c**. He also reflected on the fact that light—at one time thought to be strictly a wave phenomenon—has both wave and particle characteristics. These thoughts led de Broglie to pose a new question. If waves can have particlelike behavior, could the opposite also be true? That is, can particles of matter, including electrons, behave like waves?

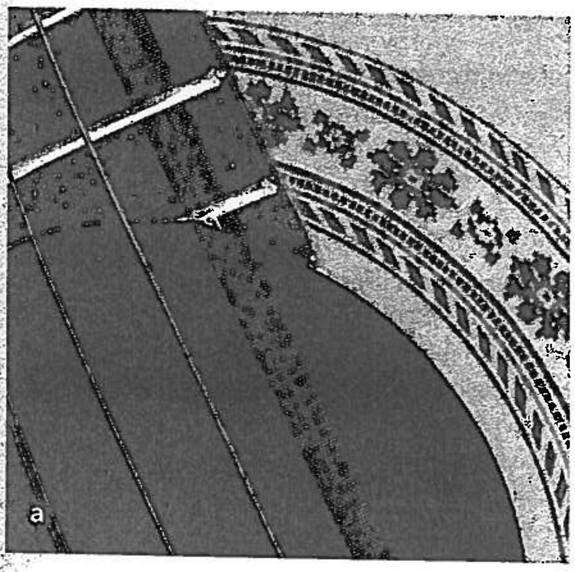
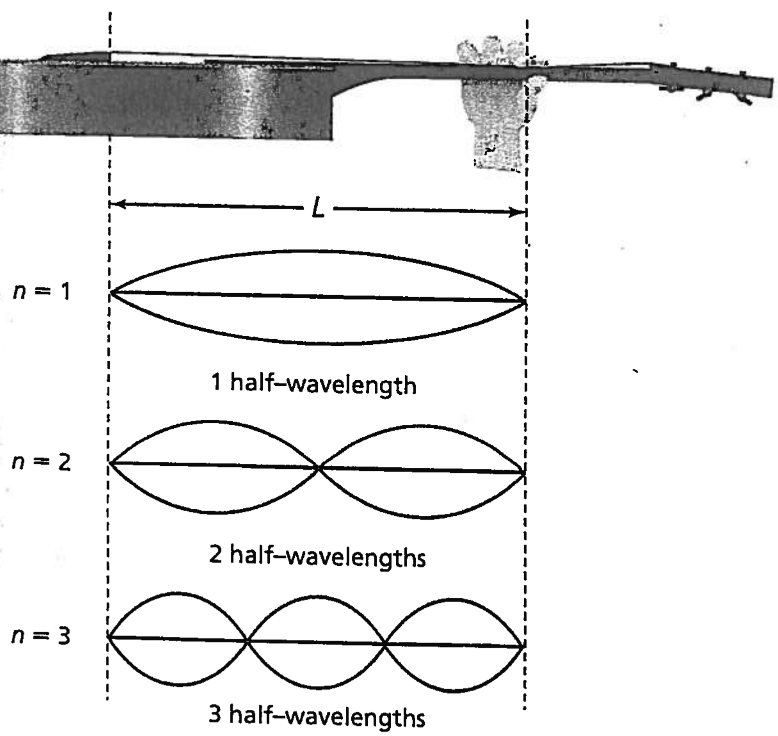
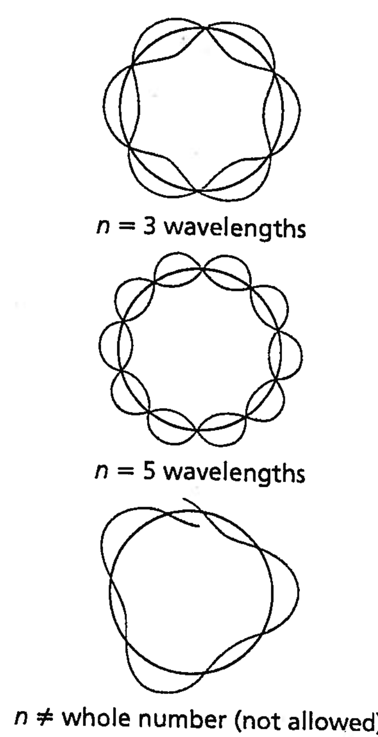


Figure 5-11

ⓐ A vibrating guitar string is constrained to vibrate between two fixed end points. ⓑ The possible vibrations of the guitar string are limited to multiples of half-wavelengths. Thus, the "quantum" of the guitar string is one-half wavelength. ⓒ The possible circular orbits of an electron are limited to whole numbers of complete wavelengths.



ⓑ Vibrating guitar string
Only multiples of half wavelengths allowed



ⓒ Orbiting electron
Only whole numbers of wavelengths allowed

Section 5.3 Electron Configurations

When you consider that atoms of the heaviest elements contain in excess of 100 electrons, that there are numerous principal energy levels and sublevels and their corresponding orbitals, and that each orbital may contain a maximum of two electrons, the idea of determining the arrangement of an atom's electrons seems daunting. Fortunately, the arrangement of electrons in atoms follows a few very specific rules. In this section, you'll learn these rules and their occasional exceptions.

Objectives

- **Apply** the Pauli exclusion principle, the aufbau principle, and Hund's rule to write electron configurations using orbital diagrams and electron configuration notation.
- **Define** valence electrons and draw electron-dot structures representing an atom's valence electrons.

Vocabulary

electron configuration
aufbau principle
Pauli exclusion principle
Hund's rule
valence electron
electron-dot structure

Ground-State Electron Configurations

The arrangement of electrons in an atom is called the atom's **electron configuration.** Because low-energy systems are more stable than high-energy systems, electrons in an atom tend to assume the arrangement that gives the atom the lowest possible energy. The most stable, lowest-energy arrangement of the electrons in atoms of each element is called the element's ground-state electron configuration. Three rules, or principles—the aufbau principle, the Pauli exclusion principle, and Hund's rule—define how electrons can be arranged in an atom's orbitals.

The aufbau principle The **aufbau principle** states that each electron occupies the lowest energy orbital available. Therefore, your first step in determining an element's ground-state electron configuration is learning the sequence of atomic orbitals from lowest energy to highest energy. This sequence, known as an aufbau diagram, is shown in **Figure 5-17.** In the diagram, each box represents an atomic orbital. Several features of the aufbau diagram stand out.

- *All orbitals related to an energy sublevel are of equal energy.* For example, all three 2p orbitals are of equal energy.
- *In a multi-electron atom, the energy sublevels within a principal energy level have different energies.* For example, the three 2p orbitals are of higher energy than the 2s orbital.

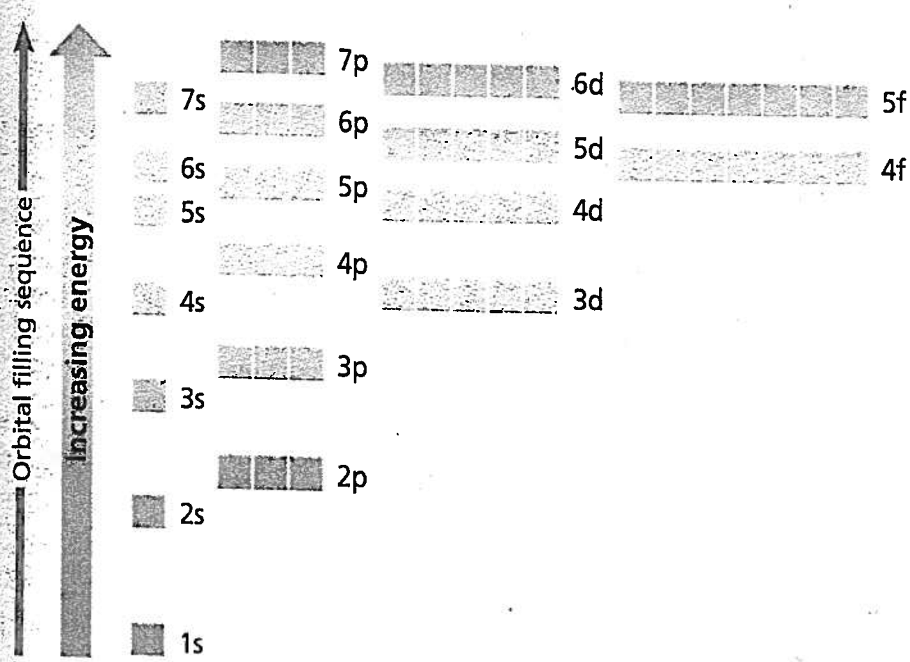


Figure 5-17

The aufbau diagram shows the energy of each sublevel. Each box on the diagram represents an atomic orbital. Does the 3d or 4s sublevel have greater energy?

Careers Using Chemistry

Spectroscopist

Are you interested in the composition of the materials around you? Do you wonder what stars are made of? Then consider a career as a spectroscopist.

Spectroscopy is the analysis of the characteristic spectra emitted by matter. Spectroscopists perform chemical analyses as part of many research laboratory projects, for quality control in industrial settings, and as part of forensics investigations for law enforcement agencies.

- *In order of increasing energy, the sequence of energy sublevels within a principal energy level is s, p, d, and f.*
- *Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level.* For example, the orbital related to the atom's 4s sublevel has a lower energy than the five orbitals related to the 3d sublevel.

Although the aufbau principle describes the sequence in which orbitals are filled with electrons, it's important to know that atoms are not actually built up electron by electron.

The Pauli exclusion principle Each electron in an atom has an associated spin, similar to the way a top spins on its axis. Like the top, the electron is able to spin in only one of two directions. An arrow pointing up (↑) represents the electron spinning in one direction, an arrow pointing down (↓) represents the electron spinning in the opposite direction. The **Pauli exclusion principle** states that a maximum of two electrons may occupy a single atomic orbital, but only if the electrons have opposite spins. Austrian physicist Wolfgang Pauli proposed this principle after observing atoms in excited states. An atomic orbital containing paired electrons with opposite spins is written as ↑↓.

Hund's rule The fact that negatively charged electrons repel each other has an important impact on the distribution of electrons in equal-energy orbitals. **Hund's rule** states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. For example, let the boxes below represent the 2p orbitals. One electron enters each of the three 2p orbitals before a second electron enters any of the orbitals. The sequence in which six electrons occupy three p orbitals is shown below.

1.	2.	3.
[↑][ ][ ]	[↑][↑][ ]	[↑][↑][↑]
4.	**5.**	**6.**
[↑↓][↑][↑]	[↑↓][↑↓][↑]	[↑↓][↑↓][↑↓]

Orbital Diagrams and Electron Configuration Notations

You can represent an atom's electron configuration using two convenient methods. One method is called an orbital diagram. An orbital diagram includes a box for each of the atom's orbitals. An empty box [ ] represents an unoccupied orbital; a box containing a single up arrow [↑] represents an orbital with one electron; and a box containing both up and down arrows [↑↓] represents a filled orbital. Each box is labeled with the principal quantum number and sublevel associated with the orbital. For example, the orbital diagram for a ground-state carbon atom, which contains two electrons in the 1s orbital, two electrons in the 2s orbital, and 1 electron in two of three separate 2p orbitals, is shown below.

	1s	2s	2p
C	[↑↓]	[↑↓]	[↑][↑][ ]

Table 5-3

Electron Configurations and Orbital Diagrams for Elements in the First Two Periods

Element	Atomic number	Orbital diagram 1s 2s $2p_x$ $2p_y$ $2p_z$	Electron configuration notation
Hydrogen	1	[↑]	$1s^1$
Helium	2	[↑↓]	$1s^2$
Lithium	3	[↑↓] [↑]	$1s^22s^1$
Beryllium	4	[↑↓] [↑↓]	$1s^22s^2$
Boron	5	[↑↓] [↑↓] [↑]	$1s^22s^22p^1$
Carbon	6	[↑↓] [↑↓] [↑] [↑]	$1s^22s^22p^2$
Nitrogen	7	[↑↓] [↑↓] [↑] [↑] [↑]	$1s^22s^22p^3$
Oxygen	8	[↑↓] [↑↓] [↑↓] [↑] [↑]	$1s^22s^22p^4$
Fluorine	9	[↑↓] [↑↓] [↑↓] [↑↓] [↑]	$1s^22s^22p^5$
Neon	10	[↑↓] [↑↓] [↑↓] [↑↓] [↑↓]	$1s^22s^22p^6$

Recall that the number of electrons in an atom equals the number of protons, which is designated by the element's atomic number. Carbon, which has an atomic number of six, has six electrons in its configuration.

Another shorthand method for describing the arrangement of electrons in an element's atoms is called electron configuration notation. This method designates the principal energy level and energy sublevel associated with each of the atom's orbitals and includes a superscript representing the number of electrons in the orbital. For example, the electron configuration notation of a ground-state carbon atom is written $1s^22s^22p^2$. Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in **Table 5-3**. To help you visualize the relative sizes and orientations of atomic orbitals, the filled 1s, 2s, $2p_x$, $2p_y$, and $2p_z$ orbitals of the neon atom are illustrated in **Figure 5-18**.

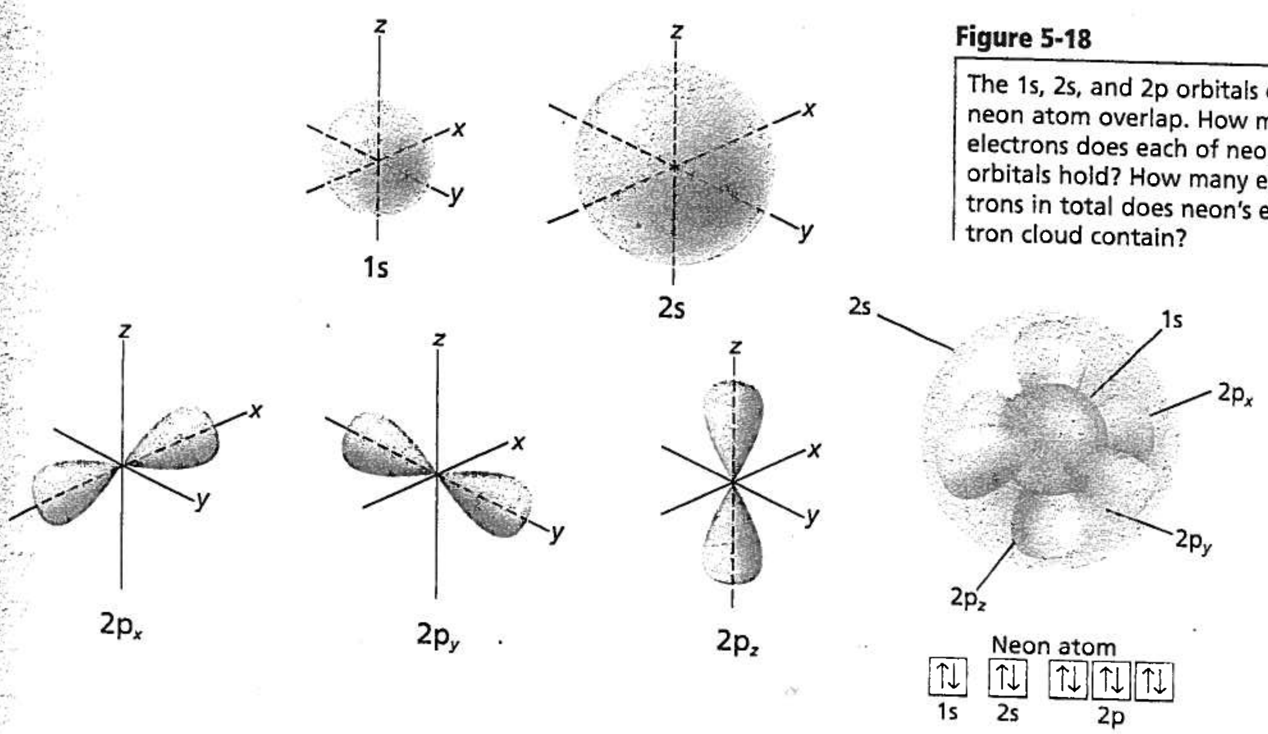


Figure 5-18

The 1s, 2s, and 2p orbitals of a neon atom overlap. How many electrons does each of neon's orbitals hold? How many electrons in total does neon's electron cloud contain?

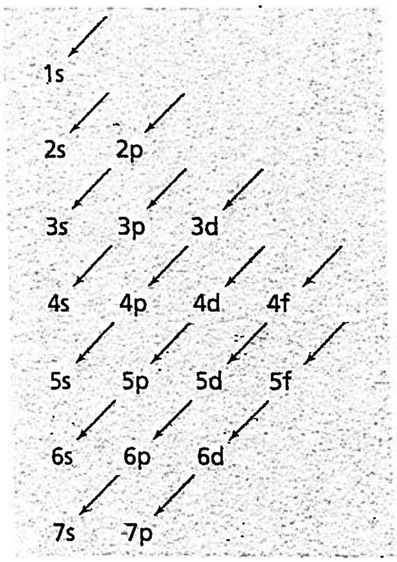


Figure 5-19

This sublevel diagram shows the order in which the orbitals are usually filled. The proper sequence for the first seven orbitals is 1s, 2s, 2p, 3s, 3p, 4s, and 3d. Which is filled first, the 5s or the 4p orbital?

Note that electron configuration notation usually does not show the orbital distributions of electrons related to a sublevel. It's understood that a designation such as nitrogen's $2p^3$ represents the orbital occupancy $2p_x^{1}2p_y^{1}2p_z^{1}$.

For sodium, the first ten electrons occupy 1s, 2s, and 2p orbitals. Then, according to the aufbau sequence, the eleventh electron occupies the 3s orbital. The electron configuration notation and orbital diagram for sodium are written

Na $1s^22s^22p^63s^1$	[↑↓]	[↑↓]	[↑↓][↑↓][↑↓]	[↑]
	1s	2s	2p	3s

Noble-gas notation is a method of representing electron configurations of noble gases using bracketed symbols. For example, [He] represents the electron configuration for helium, $1s^2$, and [Ne] represents the electron configuration for neon, $1s^22s^22p^6$. Compare the electron configuration for neon with sodium's configuration above. Note that the inner-level configuration for sodium is identical to the electron configuration for neon. Using noble-gas notation, sodium's electron configuration can be shortened to the form $[Ne]3s^1$. The electron configuration for an element can be represented using the noble-gas notation for the noble gas in the previous period and the electron configuration for the energy level being filled. The complete and abbreviated (using noble-gas notation) electron configurations of the period 3 elements are shown in **Table 5-4.**

When writing electron configurations, you may refer to a convenient memory aid called a sublevel diagram, which is shown in **Figure 5-19**. Note that following the direction of the arrows in the sublevel diagram produces the sublevel sequence shown in the aufbau diagram of **Figure 5-17.**

Exceptions to predicted configurations You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, $[Ar]4s^23d^4$, and copper, $[Ar]4s^23d^9$, would prove to be incorrect. The correct configurations for these two elements are:

Cr $[Ar]4s^13d^5$ Cu $[Ar]4s^13d^{10}$

The electron configurations for these two elements, as well as those of several elements in other periods, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

Table 5-4

Electron Configurations for Elements in Period Three

Element	Atomic number	Complete electron configuration	Electron configuration using noble-gas notation
Sodium	11	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Magnesium	12	$1s^22s^22p^63s^2$	$[Ne]3s^2$
Aluminum	13	$1s^22s^22p^63s^23p^1$	$[Ne]3s^23p^1$
Silicon	14	$1s^22s^22p^63s^23p^2$	$[Ne]3s^23p^2$
Phosphorus	15	$1s^22s^22p^63s^23p^3$	$[Ne]3s^23p^3$
Sulfur	16	$1s^22s^22p^63s^23p^4$	$[Ne]3s^23p^4$
Chlorine	17	$1s^22s^22p^63s^23p^5$	$[Ne]3s^23p^5$
Argon	18	$1s^22s^22p^63s^23p^6$	$[Ne]3s^23p^6$ or [Ar]

EXAMPLE PROBLEM 5-3

Writing Electron Configurations

Germanium (Ge), a semiconducting element, is commonly used in the manufacture of computer chips. What is the ground-state electron configuration for an atom of germanium?

1. Analyze the Problem

You are given the semiconducting element, germanium (Ge). Consult the periodic table to determine germanium's atomic number, which also is equal to its number of electrons. Also note the atomic number of the noble gas element that precedes germanium in the table. Determine the number of additional electrons a germanium atom has compared to the nearest preceding noble gas, and then write out germanium's electron configuration.

2. Solve for the Unknown

From the periodic table, germanium's atomic number is determined to be 32. Thus, a germanium atom contains 32 electrons. The noble gas preceding germanium is argon (Ar), which has an atomic number of 18. Represent germanium's first 18 electrons using the chemical symbol for argon written inside brackets.

[Ar]

The remaining 14 electrons of germanium's configuration need to be written out. Because argon is a noble gas in the third period of the periodic table, it has completely filled 3s and 3p orbitals. Thus, the remaining 14 electrons fill the 4s, 3d, and 4p orbitals in order.

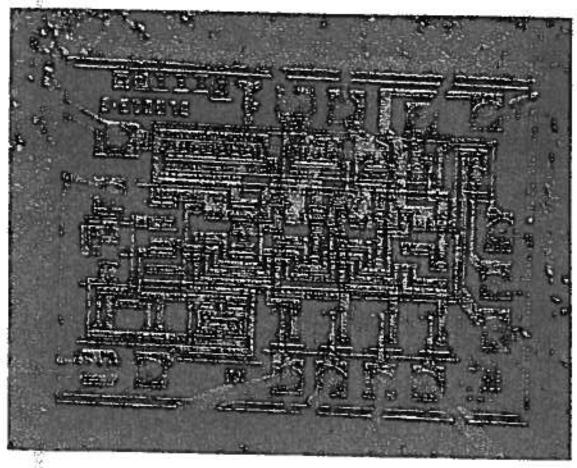
$[Ar]4s^?3d^?4p^?$

Using the maximum number of electrons that can fill each orbital, write out the electron configuration.

$[Ar]4s^2 3d^{10} 4p^2$

3. Evaluate the Answer

All 32 electrons in a germanium atom have been accounted for. The correct preceding noble gas (Ar) has been used in the notation, and the order of orbital filling for the fourth period is correct (4s, 3d, 4p).



Atoms of boron and arsenic are inserted into germanium's crystal structure in order to produce a semiconducting material that can be used to manufacture computer chips.

PRACTICE PROBLEMS

18. Write ground-state electron configurations for the following elements.

a. bromine (Br) 35
b. strontium (Sr) 38
c. antimony (Sb) 51
d. rhenium (Re)
e. terbium (Tb)
f. titanium (Ti)

19. How many electrons are in orbitals related to the third energy level of a sulfur atom?

20. How many electrons occupy p orbitals in a chlorine atom?

21. What element has the following ground-state electron configuration? $[Kr]5s^2 4d^{10} 5p^1$

22. What element has the following ground-state electron configuration? $[Xe]6s^2$

Practice!

For more practice with electron configuration problems, go to Supplemental Practice Problems in Appendix A.

Valence Electrons

Only certain electrons, called valence electrons, determine the chemical properties of an element. **Valence electrons** are defined as electrons in the atom's outermost orbitals—generally those orbitals associated with the atom's highest principal energy level. For example, a sulfur atom contains 16 electrons, only six of which occupy the outermost 3s and 3p orbitals, as shown by sulfur's electron configuration. Sulfur has six valence electrons.

$$\mathrm{S}\ [\mathrm{Ne}]3s^2 3p^4$$

Similarly, although a cesium atom contains 55 electrons, it has but one valence electron, the 6s electron shown in cesium's electron configuration.

$$\mathrm{Cs}\ [\mathrm{Xe}]6s^1$$

Francium, which belongs to the same group as cesium, also has a single valence electron.

$$\mathrm{Fr}\ [\mathrm{Rn}]7s^1$$

Electron-dot structures Because valence electrons are involved in forming chemical bonds, chemists often represent them visually using a simple shorthand method. An atom's **electron-dot structure** consists of the element's symbol, which represents the atomic nucleus and inner-level electrons, surrounded by dots representing the atom's valence electrons. The American chemist G. N. Lewis (1875–1946), devised the method while teaching a college chemistry class in 1902.

In writing an atom's electron-dot structure, dots representing valence electrons are placed one at a time on the four sides of the symbol (they may be placed in any sequence) and then paired up until all are used. The ground-state electron configurations and electron-dot structures for the elements in the second period are shown in **Table 5-5**.

Table 5-5

Electron-Dot Structures for Elements in Period Two

Element	Atomic number	Electron configuration	Electron-dot structure
Lithium	3	$1s^2 2s^1$	$\mathrm{Li}\cdot$
Beryllium	4	$1s^2 2s^2$	$\cdot\mathrm{Be}\cdot$
Boron	5	$1s^2 2s^2 2p^1$	$\cdot\dot{\mathrm{B}}\cdot$
Carbon	6	$1s^2 2s^2 2p^2$	$\cdot\underset{\cdot}{\dot{\mathrm{C}}}\cdot$
Nitrogen	7	$1s^2 2s^2 2p^3$	$\cdot\underset{\cdot}{\ddot{\mathrm{N}}}\cdot$
Oxygen	8	$1s^2 2s^2 2p^4$	$:\underset{\cdot}{\ddot{\mathrm{O}}}\cdot$
Fluorine	9	$1s^2 2s^2 2p^5$	$:\underset{\cdot\cdot}{\ddot{\mathrm{F}}}\cdot$
Neon	10	$1s^2 2s^2 2p^6$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ne}}}:$

EXAMPLE PROBLEM 5-4

Writing Electron-Dot Structures

Some sheet glass is manufactured using a process that makes use of molten tin. What is tin's electron-dot structure?

1. Analyze the Problem

You are given the element tin (Sn). Consult the periodic table to determine the total number of electrons an atom of tin has. Write out tin's electron configuration and determine the number of valence electrons it has. Then use the number of valence electrons and the rules for electron-dot structures to draw the electron-dot structure for tin.

2. Solve for the Unknown

From the periodic table, tin is found to have an atomic number of 50. Thus, a tin atom has 50 electrons. Write out the noble-gas form of tin's electron configuration.

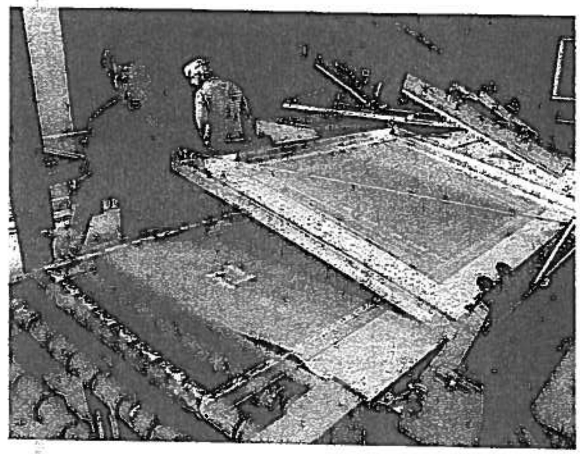
$[Kr]5s^2 4d^{10} 5p^2$

The two 5s and the two 5p electrons (the electrons in the orbitals related to the atom's highest principal energy level) represent tin's four valence electrons. Draw tin's electron-dot structure by representing its four valence electrons with dots, arranged one at a time, around the four sides of tin's chemical symbol (Sn).

$\cdot\overset{\cdot}{\underset{\cdot}{\mathrm{Sn}}}\cdot$

3. Evaluate the Answer

The correct symbol for tin (Sn) has been used, and the rules for drawing electron-dot structures have been correctly applied.



Flat-surfaced window glass may be manufactured by floating molten glass on top of molten tin.

PRACTICE PROBLEMS

23. Draw electron-dot structures for atoms of the following elements.
   - a. magnesium
   - b. sulfur
   - c. bromine
   - d. rubidium
   - e. thallium
   - f. xenon

Practice!

For more practice with electron-dot structure problems, go to Supplemental Practice Problems in Appendix A.

Section 5.3 Assessment

24. State the aufbau principle in your own words.
25. Apply the Pauli exclusion principle, the aufbau principle, and Hund's rule to write out the electron configuration and draw the orbital diagram for each of the following elements.
   - a. silicon
   - b. fluorine
   - c. calcium
   - d. krypton
26. What is a valence electron? Draw the electron-dot structures for the elements in problem 25.
27. **Thinking Critically** Use Hund's rule and orbital diagrams to describe the sequence in which ten electrons occupy the five orbitals related to an atom's d sublevel.
28. **Interpreting Scientific Illustrations** Which of the following is the correct electron-dot structure for an atom of selenium? Explain.

   a. $\cdot\overset{..}{\underset{..}{\mathrm{Se}}}:$ b. $\cdot\overset{...}{\underset{\cdot}{\mathrm{Se}}}\cdot$ c. $\cdot\overset{..}{\underset{..}{\mathrm{Se}}}\cdot$ d. $\cdot\overset{..}{\underset{..}{\mathrm{S}}}\cdot$

62. Light is said to have a dual wave-particle nature. What does this statement mean? (5.3)

63. Describe the difference between a quantum and a photon. (5.3)

64. How many electrons are shown in the electron-dot structures of the following elements? (5.3)

a. carbon
b. iodine
c. calcium
d. gallium

Mastering Problems

Wavelength, Frequency, Speed, and Energy (5.1)

65. What is the wavelength of electromagnetic radiation having a frequency of 5.00×10^{12} Hz? What kind of electromagnetic radiation is this?

66. What is the frequency of electromagnetic radiation having a wavelength of 3.33×10^{-8} m? What type of electromagnetic radiation is this?

67. The laser in a compact disc (CD) player uses light with a wavelength of 780 nm. What is the frequency of this light?

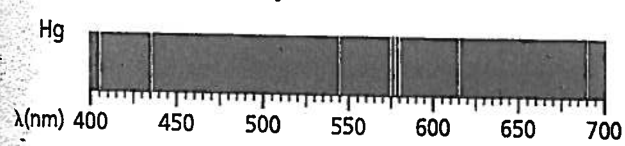
68. What is the speed of an electromagnetic wave having a frequency of 1.33×10^{17} Hz and a wavelength of 2.25 nm?

69. Use **Figure 5-5** to determine each of the following types of radiation.

a. radiation with a frequency of $8.6 \times 10^{11}\ s^{-1}$
b. radiation with a wavelength 4.2 nm
c. radiation with a frequency of 5.6 MHz
d. radiation that travels at a speed of 3.00×10^{8} m/s

70. What is the energy of a photon of red light having a frequency of 4.48×10^{14} Hz?

71. Mercury's atomic emission spectrum is shown below. Estimate the wavelength of the orange line. What is its frequency? What is the energy of an orange photon emitted by the mercury atom?



72. What is the energy of an ultraviolet photon having a wavelength of 1.18×10^{-8} m?

73. A photon has an energy of 2.93×10^{-25} J. What is its frequency? What type of electromagnetic radiation is the photon?

74. A photon has an energy of 1.10×10^{-13} J. What is the photon's wavelength? What type of electromagnetic radiation is it?

75. How long does it take a radio signal from the Voyager spacecraft to reach Earth if the distance between Voyager and Earth is 2.72×10^{9} km?

76. If your favorite FM radio station broadcasts at a frequency of 104.5 MHz, what is the wavelength of the station's signal in meters? What is the energy of a photon of the station's electromagnetic signal?

Electron Configurations (5.3)

77. List the aufbau sequence of orbitals from 1s to 7p.

78. Write orbital notations and complete electron configurations for atoms of the following elements.

a. beryllium
b. aluminum
c. nitrogen
d. sodium

79. Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.

a. Mn
b. Kr
c. P
d. Zn
e. Zr
f. W
g. Pb
h. Ra
i. Sm
j. Bk

80. What elements are represented by each of the following electron configurations?

a. $1s^2 2s^2 2p^5$
b. $[Ar]4s^2$
c. $[Xe]6s^2 4f^4$
d. $[Kr]5s^2 4d^{10} 5p^4$
e. $[Rn]7s^2 5f^{13}$
f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

81. Draw electron-dot structures for atoms of each of the following elements.

a. carbon
b. arsenic
c. polonium
d. potassium
e. barium

82. An atom of arsenic has how many electron-containing orbitals? How many of the orbitals are completely filled? How many of the orbitals are associated with the atom's $n = 4$ principal energy level?

CHAPTER 5 ASSESSMENT

Mixed Review

Sharpen your problem-solving skills by answering the following.

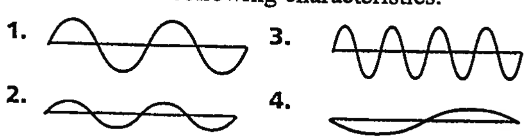
83. What is the frequency of electromagnetic radiation having a wavelength of 1.00 m?

84. What is the maximum number of electrons that can be contained in an atom's orbitals having the following principal quantum numbers?

a. 3 **b.** 4 **c.** 6 **d.** 7

85. What is the wavelength of light with a frequency of 5.77×10^{14} Hz?

86. Using the waves shown below, identify the wave or waves with the following characteristics.



a. longest wavelength
b. greatest frequency
c. largest amplitude
d. shortest wavelength

87. How many orientations are possible for the orbitals related to each of the following sublevels?

a. s **b.** p **c.** d **d.** f

88. Describe the electrons in an atom of nickel in the ground state using the electron configuration notation and the noble-gas notation.

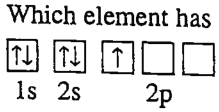
89. Which of the following elements have two electrons in their electron-dot structures: hydrogen, helium, lithium, aluminum, calcium, cobalt, bromine, krypton, and barium?

90. In Bohr's atomic model, what electron orbit transition produces the blue-green line in hydrogen's atomic emission spectrum?

91. A zinc atom contains a total of 18 electrons in its 3s, 3p, and 3d orbitals. Why does its electron-dot structure show only two dots?

92. An X-ray photon has an energy of 3.01×10^{-18} J. What is its frequency and wavelength?

93. Which element has the following orbital diagram?



94. Which element has the ground-state electron configuration represented by the noble-gas notation $[Rn]7s^1$?

95. How many photons of infrared radiation having a frequency of 4.88×10^{13} Hz are required to provide an energy of 1.00 J?

Thinking Critically

96. Comparing and Contrasting Briefly discuss the difference between an orbit in Bohr's model of the atom and an orbital in the quantum mechanical view of the atom.

97. Applying Concepts Scientists use atomic emission spectra to determine the elements in materials of unknown composition. Explain what makes this method possible.

98. Using Numbers It takes 8.17×10^{-19} J of energy to remove one electron from a gold surface. What is the maximum wavelength of light capable of causing this effect?

99. Drawing a Conclusion The elements aluminum, silicon, gallium, germanium, arsenic, selenium are all used in making various types of semiconductor devices. Write electron configurations and electron-dot structures for atoms of each of these elements. What similarities among the elements' electron configurations do you notice?

Writing in Chemistry

100. In order to make "neon" signs emit a variety of colors, manufacturers often fill the signs with gases other than neon. Research the use of gases in neon signs and specify the colors produced by the gases.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

101. Round 20.561 20 g to three significant figures. (Chapter 2)

102. Identify each of the following as either chemical or physical properties of the substance. (Chapter 3)

a. mercury is a liquid at room temperature
b. sucrose is a white, crystalline solid
c. iron rusts when exposed to moist air
d. paper burns when ignited

103. Identify each of the following as a pure substance or a mixture. (Chapter 3)

a. distilled water
b. orange juice with pulp
c. smog
d. diamond
e. milk
f. copper metal

104. An atom of gadolinium has an atomic number of 64 and a mass number of 153. How many electrons, protons, and neutrons does it contain? (Chapter 4)

EXAMPLE PROBLEM 6-2

Interpreting Trends in Atomic Radii

Which has the largest atomic radius: carbon (C), fluorine (F), beryllium (Be), or lithium (Li)? Do not use **Figure 6-11** to answer the question. Explain your answer in terms of trends in atomic radii.

1 Analyze the Problem

You are given four elements. First, determine the groups and periods the elements occupy. Then apply the general trends in atomic radii to determine which has the largest atomic radius.

2 Solve for the Unknown

From the periodic table, all the elements are found to be in period 2.

Ordering the elements from left-to-right across the period yields:

Li, Be, C, F

Applying the trend of decreasing radii across a period means that lithium, the first element in period 2, has the largest radius.

3 Evaluating the Answer

The group trend in atomic radii has been correctly applied. Checking radii values from **Figure 6-11** verifies the answer.

PRACTICE PROBLEMS

Answer the following questions using your knowledge of group and period trends in atomic radii. Do not use the atomic radii values in **Figure 6-11** to answer the questions.

16. Which has the largest radius: magnesium (Mg), silicon (Si), sulfur (S), or sodium (Na)? The smallest?
17. Which has the largest radius: helium (He), xenon (Xe), or argon (Ar)? The smallest?
18. Can you determine which of two unknown elements has the larger radius if the only known information is that the atomic number of one of the elements is 20 greater than the other?

Practice!

For more practice with periodic trend problems, go to Supplemental Practice Problems in Appendix A.

Ionic Radius

Atoms can gain or lose one or more electrons to form ions. Because electrons are negatively charged, atoms that gain or lose electrons acquire a net charge. Thus, an **ion** is an atom or a bonded group of atoms that has a positive or negative charge. You'll learn about ions in detail in Chapter 8, but for now, let's look at how the formation of an ion affects the size of an atom.

When atoms lose electrons and form positively charged ions, they always become smaller. For example, as shown in **Figure 6-13a** on the next page a sodium atom with a radius of 186 pm shrinks to a radius of 95 pm when it forms a positive sodium ion. The reason for the decrease in size is twofold. The electron lost from the atom will always be a valence electron. The loss of a valence electron may leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now fewer number of remaining electrons decreases, allowing them to be pulled closer to the nucleus.

When atoms gain electrons and form negatively charged ions, they always become larger, as shown in **Figure 6-13b.** The addition of an electron to an

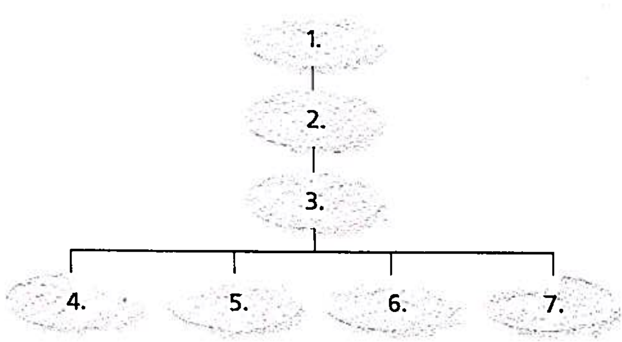
CHAPTER 6 ASSESSMENT

CHEMISTRY Online

Go to the Chemistry Web site at **science.glencoe.com** *or use the* **Chemistry CD-ROM** *for additional Chapter 6 Assessment.*

Concept Mapping

25. Complete the concept map using the following terms: electronegativity, electron configuration, periodic trends, ionic radius, atomic radius, ionization energy, and periodic table.



Mastering Concepts

26. Explain how Mendeleev's periodic table was in error. How was this error fixed? (6.1)

27. Explain the contribution of Newlands's law of octaves to the development of the modern periodic table. (6.1)

28. German chemist Lothar Meyer and Russian chemist Dmitri Mendeleev both proposed similar periodic tables in 1869. Why is Mendeleev generally given credit for the periodic table? (6.1)

29. How was Mendeleev's periodic table organized? (6.1)

30. What is the periodic law? (6.1)

31. Identify each of the following as a metal, nonmetal, or metalloid. (6.1)

a. oxygen
b. barium
c. germanium
d. iron
e. neon
f. praseodymium

32. Describe the general characteristics of metals. (6.1)

33. Match each numbered item on the right with the lettered item that it is related to on the left. (6.1)

a. alkali metals	**1.** group 8A
b. halogens	**2.** group 1A
c. alkaline earth metals	**3.** group 2A
d. noble gases	**4.** group 7A

34. Identify each of the elements in problem 31 as a representative element or a transition element. (6.1)

35. Sketch a simplified periodic table and use labels to identify the alkali metals, alkaline earth metals, transition metals, inner transition metals, noble gases, and halogens. (6.1)

36. A shiny solid element also is ductile. What side of the periodic table is it likely to be found? (6.1)

37. What are the general properties of a metalloid? List three metalloid elements. (6.1)

38. What is the purpose of the heavy stair-step line on the periodic table? (6.1)

39. Describe the two types of numbering used to identify groups on the periodic table. (6.1)

40. Give the chemical symbol of each of the following elements. (6.1)

a. the two elements that are liquids at room temperature
b. the noble gas with the greatest atomic mass
c. any metal from group 4A
d. any inner transition metal

41. Why do the elements chlorine and iodine have similar chemical properties? (6.2)

42. How are the numbers of valence electrons of the group A elements related to the group number? (6.2)

43. How is the energy level of an atom's valence electrons related to the period it is in on the periodic table? (6.2)

44. How many valence electrons do each of the noble gases have? (6.2)

45. What are the four blocks of the periodic table? (6.2)

46. In general, what electron configuration has the greatest stability? (6.2)

47. Determine the group, period, and block in which each of the following elements is located on the periodic table. (6.2)

a. $[Kr]5s^2 4d^1$
b. $[Ar]4s^2 3d^{10} 4p^3$
c. $[He]2s^2 2p^6$
d. $[Ne]3s^2 3p^1$

48. Categorize each of the elements in problem 47 as a representative element or a transition metal. (6.2)

49. Explain how an atom's valence electron configuration determines its place on the periodic table. (6.2)

50. Write the electron configuration for the element fitting each of the following descriptions. (6.2)

a. the metal in group 5A
b. the halogen in period 3
c. the alkali metal in period 2
d. the transition metal that is a liquid at room temperature

51. Explain why the radius of an atom cannot be measured directly. (6.3)

52. Given any two elements within a group, is the element with the larger atomic number likely to have a larger or smaller atomic radius than the other element? (6.2)

53. Which elements are characterized as having their d orbitals fill with electrons as you move left-to-right across a period? (6.2)

54. Explain why is it harder to remove an inner shell electron than a valence electron from an atom. (6.3)

55. An element forms a negative ion when ionized. On what side of the periodic table is the element located? Explain. (6.3)

56. Of the elements magnesium, calcium, and barium, which forms the ion with the largest radius? The smallest? What periodic trend explains this? (6.3)

57. What is ionization energy? (6.3)

58. Explain why each successive ionization of an electron requires a greater amount of energy. (6.3)

59. Which group has the highest ionization energies? Explain why. (6.3)

60. Define an ion. (6.3)

61. How does the ionic radius of a nonmetal compare with its atomic radius? Explain why the change in radius occurs. (6.3)

62. Explain why atomic radii decrease as you move left-to-right across a period. (6.3)

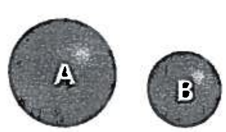
63. Which element in each pair has the larger ionization energy? (6.3)

a. Li, N
b. Kr, Ne
c. Cs, Li

64. Explain the octet rule. (6.3)

65. Use the illustration of spheres A and B to answer each of the following questions. Explain your reasoning for each answer. (6.3)

a. If A is an ion and B is an atom of the same element, is the ion a positive or negative ion?



b. If A and B represent the atomic radii of two elements in the same period, what is their correct order (left-to-right)?

c. If A and B represent the ionic radii of two elements in the same group, what is their correct order (top-to-bottom)?

66. How many valence electrons do elements in each of the following groups have? (6.3)

a. group 8A
b. group 3A
c. group 1A

67. Na^+ and Mg^{2+} ions each have ten electrons surrounding their nuclei. Which ion would you expect to have the larger radius? Why? (6.3)

Mixed Review

Sharpen your problem-solving skills by answering the following.

68. Match each numbered item on the right with the lettered item that it is related to on the left.

a. group A elements	**1.** periods
b. columns	**2.** representative elements
c. group B elements	**3.** groups
d. rows	**4.** transition elements

69. Which element in each pair is more electronegative?

a. K, As
b. N, Sb
c. Sr, Be

70. Explain why the s-block of the periodic table is two groups wide, the p-block is six groups wide, and the d-block is ten groups wide.

71. Arrange the elements oxygen, sulfur, tellurium, and selenium in order of increasing atomic radii. Is your order an example of a group trend or a period trend?

72. Identify the elements with the following valence electron configurations.

a. $5s^1$
b. $4s^2 3d^2$
c. $3s^2$
d. $4s^2 4p^3$

73. Which of the following is not a reason why atomic radii increase as you move down a group?
a. shielding of inner electrons
b. valence electrons in larger orbitals
c. increased charge in the nucleus

74. Explain why there are no p-block elements in the first period of the periodic table.

75. Identify each of the following as an alkali metal, alkaline earth metal, transition metal, or inner transition metal.

a. cesium
b. zirconium
c. gold
d. ytterbium
e. uranium
f. francium

76. An element is a brittle solid that does not conduct electricity well. Is the element a metal, nonmetal, or metalloid?