Arrangement of Electrons in Atoms
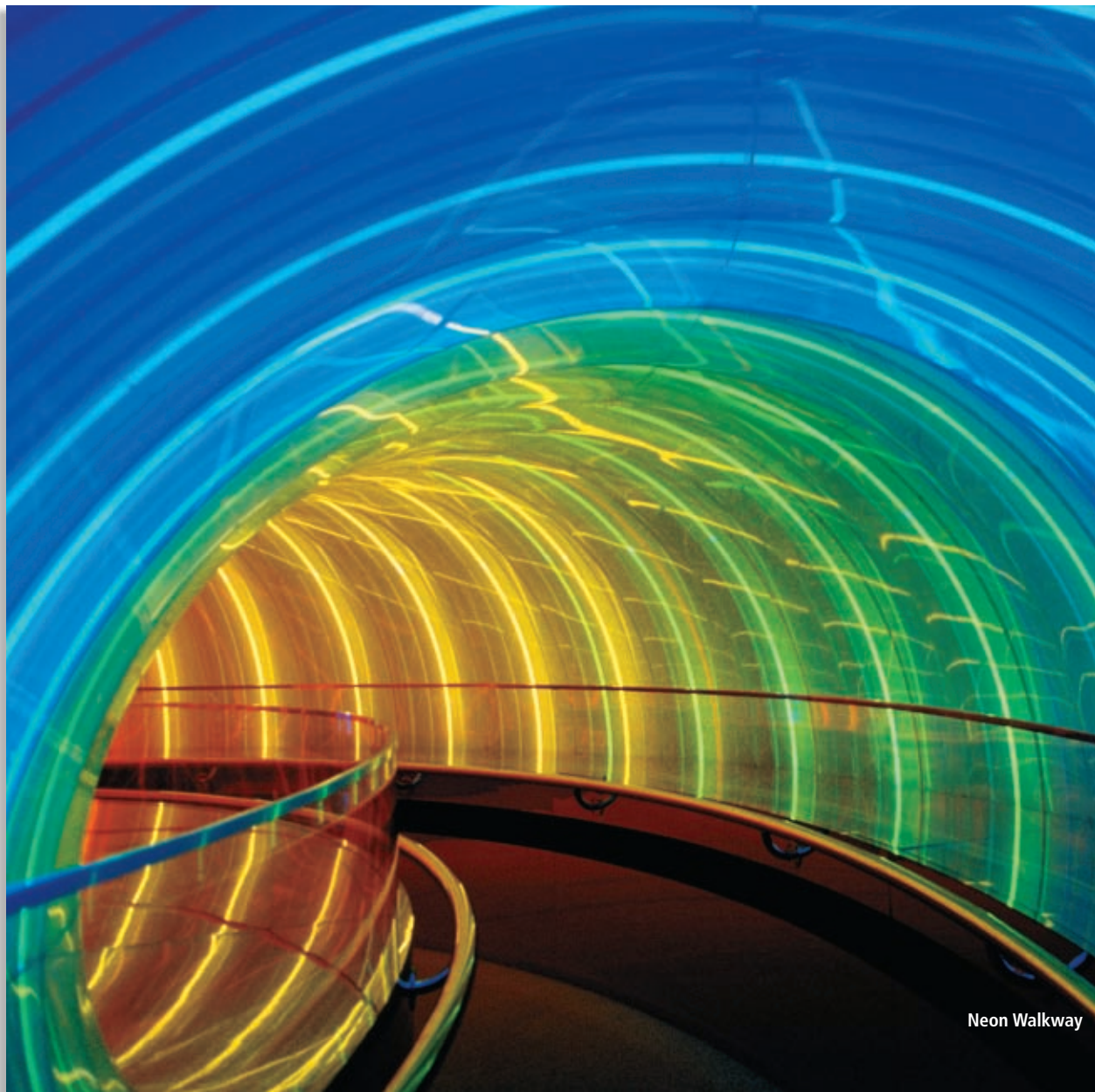
The emission of light is fundamentally related to the behavior of electrons.

Neon Walkway

The Development of a New Atomic Model

OBJECTIVES

- Explain the mathematical relationship among the speed, wavelength, and frequency of electromagnetic radiation.

- Discuss the dual wave-particle nature of light.

- Discuss the significance of the photoelectric effect and the line-emission spectrum of hydrogen to the development of the atomic model.

- Describe the Bohr model of the hydrogen atom.

The Rutherford model of the atom was an improvement over previous models, but it was incomplete. It did not explain how the atom's negatively charged electrons are distributed in the space surrounding its positively charged nucleus. After all, it was well known that oppositely charged particles attract each other. So what prevented the negative electrons from being drawn into the positive nucleus?

In the early twentieth century, a new atomic model evolved as a result of investigations into the absorption and emission of light by matter. The studies revealed a relationship between light and an atom's electrons. This new understanding led directly to a revolutionary view of the nature of energy, matter, and atomic structure.

Properties of Light

Before 1900, scientists thought light behaved solely as a wave. This belief changed when it was later discovered that light also has particle-like characteristics. Still, many of light's properties can be described in terms of waves. A quick review of these wavelike properties will help you understand the basic theory of light as it existed at the beginning of the twentieth century.

The Wave Description of Light

Visible light is a kind of **electromagnetic radiation,** *which is a form of energy that exhibits wavelike behavior as it travels through space.* Other kinds of electromagnetic radiation include X rays, ultraviolet and infrared light, microwaves, and radio waves. *Together, all the forms of electromagnetic radiation form the* **electromagnetic spectrum.** The electromagnetic spectrum is represented in **Figure 1** on the next page. All forms of electromagnetic radiation move at a constant speed of 3.00×10^8 meters per second (m/s) through a vacuum and at slightly slower speeds through matter. Because air is mostly empty space, the value of 3.00×10^8 m/s is also light's approximate speed through air.

The significant feature of wave motion is its repetitive nature, which can be characterized by the measurable properties of wavelength and frequency. **Wavelength** (λ) *is the distance between corresponding points on adjacent waves.* The unit for wavelength is a distance unit. Depending on the type of electromagnetic radiation, it may be expressed in meters,
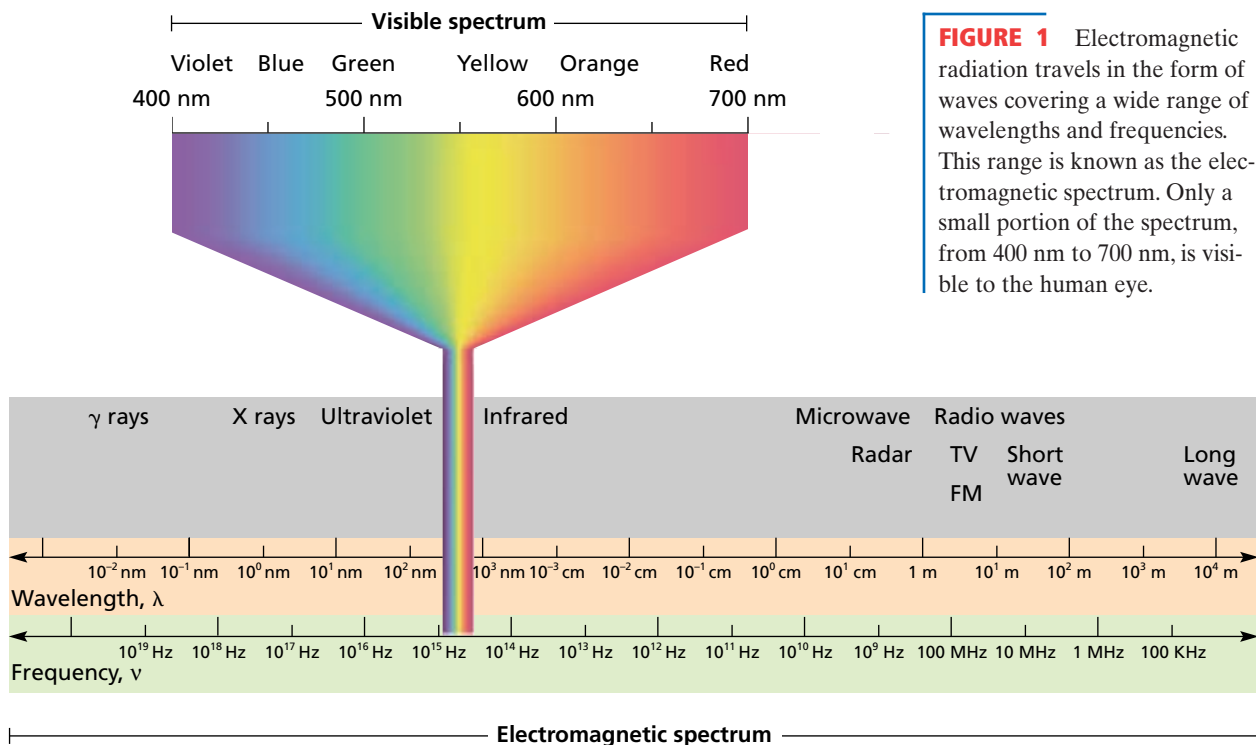
SCI LINKS.

www.scilinks.org
Topic: Electromagnetic Spectrum
Code: HC60482

Visible spectrum

Violet	Blue	Green	Yellow	Orange	Red
400 nm		500 nm		600 nm	700 nm

FIGURE 1 Electromagnetic radiation travels in the form of waves covering a wide range of wavelengths and frequencies. This range is known as the electromagnetic spectrum. Only a small portion of the spectrum, from 400 nm to 700 nm, is visible to the human eye.

γ rays   X rays   Ultraviolet   Infrared   Microwave   Radio waves

Radar   TV   Short wave   Long wave

FM

Wavelength, λ

10^{-2} nm 10^{-1} nm 10^0 nm 10^1 nm 10^2 nm 10^3 nm 10^{-3} cm 10^{-2} cm 10^{-1} cm 10^0 cm 10^1 cm 1 m 10^1 m 10^2 m 10^3 m 10^4 m

Frequency, ν

10^{19} Hz 10^{18} Hz 10^{17} Hz 10^{16} Hz 10^{15} Hz 10^{14} Hz 10^{13} Hz 10^{12} Hz 10^{11} Hz 10^{10} Hz 10^9 Hz 100 MHz 10 MHz 1 MHz 100 KHz

Electromagnetic spectrum

centimeters, or nanometers, as shown in **Figure 1. Frequency** (ν) *is defined as the number of waves that pass a given point in a specific time, usually one second.* Frequency is expressed in waves/second. One wave/second is called a hertz (Hz), named for Heinrich Hertz, who was a pioneer in the study of electromagnetic radiation. **Figure 2** illustrates the properties of wavelength and frequency for a familiar kind of wave, a wave on the surface of water. The wave in **Figure 2a** has a longer wavelength and a lower frequency than the wave in **Figure 2b.**
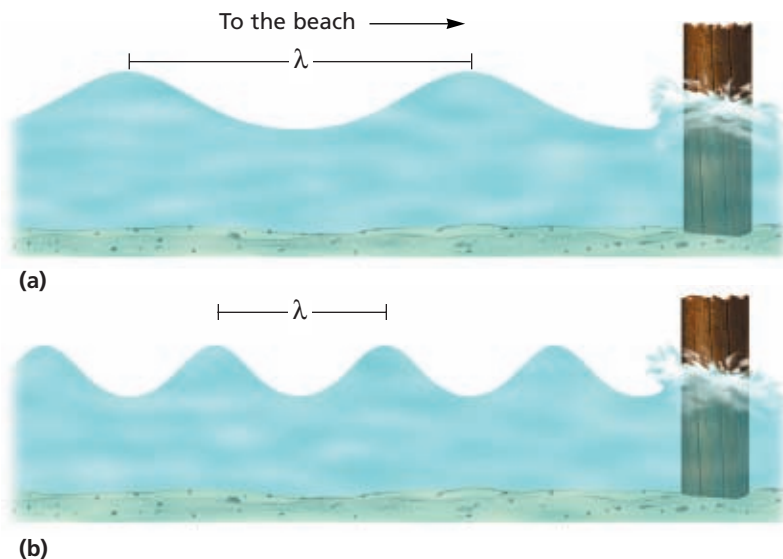
To the beach ⟶

λ

(a)

λ

(b)

FIGURE 2 The distance between any two corresponding points on one of these water waves, such as from crest to crest, is the wave's wavelength, λ. We can measure the wave's frequency, ν, by observing how often the water level rises and falls at a given point, such as at the post.

Frequency and wavelength are mathematically related to each other. For electromagnetic radiation, this relationship is written as follows.

$$c = \lambda\nu$$

In the equation, c is the speed of light (in m/s), λ is the wavelength of the electromagnetic wave (in m), and ν is the frequency of the electromagnetic wave (in s^{-1}). Because c is the same for all electromagnetic radiation, the product $\lambda\nu$ is a constant. Consequently, we know that λ is inversely proportional to ν. In other words, as the wavelength of light decreases, its frequency increases, and vice versa.

The Photoelectric Effect

In the early 1900s, scientists conducted two experiments involving interactions of light and matter that could not be explained by the wave theory of light. One experiment involved a phenomenon known as the photoelectric effect. *The* **photoelectric effect** *refers to the emission of electrons from a metal when light shines on the metal,* as illustrated in **Figure 3.**

The mystery of the photoelectric effect involved the frequency of the light striking the metal. For a given metal, no electrons were emitted if the light's frequency was below a certain minimum—regardless of the light's intensity. Light was known to be a form of energy, capable of knocking loose an electron from a metal. But the wave theory of light predicted that light of any frequency could supply enough energy to eject an electron. Scientists couldn't explain why the light had to be of a minimum frequency in order for the photoelectric effect to occur.

FIGURE 3 The photoelectric effect: electromagnetic radiation strikes the surface of the metal, ejecting electrons from the metal and causing an electric current.

The Particle Description of Light

The explanation of the photoelectric effect dates back to 1900, when German physicist Max Planck was studying the emission of light by hot objects. He proposed that a hot object does not emit electromagnetic energy continuously, as would be expected if the energy emitted were in the form of waves. Instead, Planck suggested that the object emits energy in small, specific packets called quanta. *A* **quantum** *of energy is the minimum quantity of energy that can be lost or gained by an atom.* Planck proposed the following relationship between a quantum of energy and the frequency of radiation.

$$E = h\nu$$

In the equation, E is the energy, in joules, of a quantum of radiation, ν is the frequency, in s^{-1}, of the radiation emitted, and h is a fundamental physical constant now known as Planck's constant; $h = 6.626 \times 10^{-34}$ J·s.

In 1905, Albert Einstein expanded on Planck's theory by introducing the radical idea that electromagnetic radiation has a dual wave-particle nature. While light exhibits many wavelike properties, it can also be
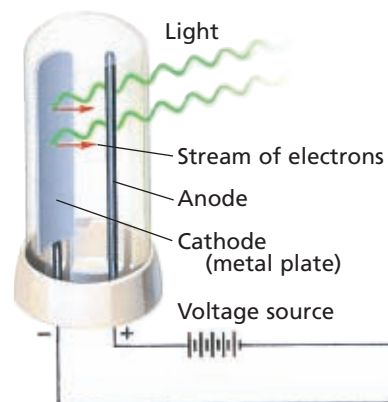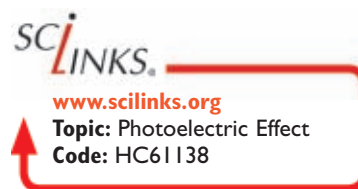
thought of as a stream of particles. Each particle of light carries a quantum of energy. Einstein called these particles photons. *A* **photon** *is a particle of electromagnetic radiation having zero mass and carrying a quantum of energy.* The energy of a particular photon depends on the frequency of the radiation.

$$E_{photon} = h\nu$$

Einstein explained the photoelectric effect by proposing that electromagnetic radiation is absorbed by matter only in whole numbers of photons. In order for an electron to be ejected from a metal surface, the electron must be struck by a single photon possessing at least the minimum energy required to knock the electron loose. According to the equation $E_{photon} = h\nu$, this minimum energy corresponds to a minimum frequency. If a photon's frequency is below the minimum, then the electron remains bound to the metal surface. Electrons in different metals are bound more or less tightly, so different metals require different minimum frequencies to exhibit the photoelectric effect.

The Hydrogen-Atom Emission-Line Spectrum

When current is passed through a gas at low pressure, the potential energy of the gas atoms increases. *The lowest energy state of an atom is its* **ground state.** *A state in which an atom has a higher potential energy than it has in its ground state is an* **excited state.** There are many possible excited states, each with a unique energy, but only one ground state energy for atoms of a given element. When an excited atom returns to its ground state or a lower energy excited state, it gives off the energy it gained in the form of electromagnetic radiation. The production of colored light in neon signs, as shown in **Figure 4,** is a familiar example of this process.

When investigators passed electric current through a vacuum tube containing hydrogen gas at low pressure, they observed the emission of a characteristic pinkish glow. *When a narrow beam of the emitted light was shined through a prism, it was separated into four specific colors of the visible spectrum*. The four bands of light were part of what is known as hydrogen's **emission-line spectrum.** The production of hydrogen's emission-line spectrum is illustrated in **Figure 5.** Additional series of lines were discovered in the ultraviolet and infrared regions of hydrogen's emission-line spectrum. The wavelengths of some of the spectral series are shown in **Figure 6.** They are known as the Lyman, Balmer, and Paschen series, after their discoverers.

Classical theory predicted that the hydrogen atoms would be excited by whatever amount of energy was added to them. Scientists had thus expected to observe *the emission of a continuous range of frequencies of electromagnetic radiation,* that is, a **continuous spectrum.** Why had the hydrogen atoms given off only specific frequencies of light? Attempts to explain this observation led to an entirely new atomic theory called *quantum theory.*

FIGURE 4 Excited neon atoms emit light when electrons in higher energy levels fall back to the ground state or to a lower-energy excited state.
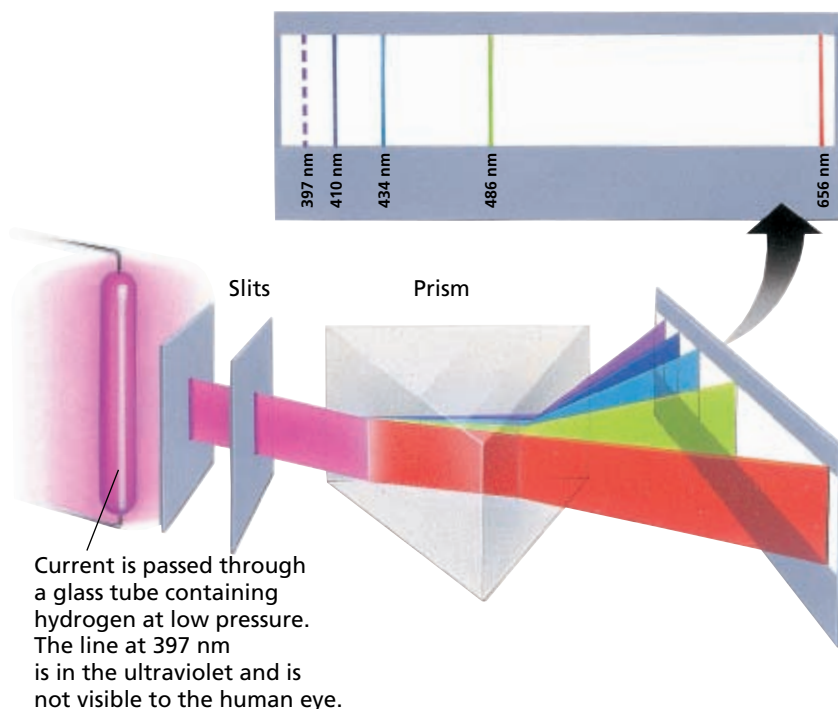
FIGURE 5 Excited hydrogen atoms emit a pinkish glow, as is shown in this diagram. When the visible portion of the emitted light is passed through a prism, it is separated into specific wavelengths that are part of hydrogen's emission-line spectrum. The line at 397 nm is in the ultraviolet and is not visible to the human eye.
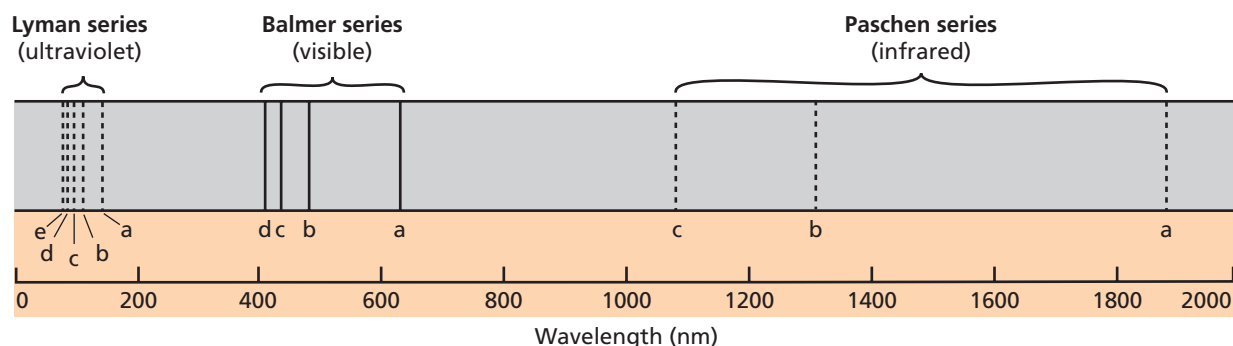
Current is passed through a glass tube containing hydrogen at low pressure. The line at 397 nm is in the ultraviolet and is not visible to the human eye.

FIGURE 6 A series of specific wavelengths of emitted light makes up hydrogen's emission-line spectrum. The letters below the lines label hydrogen's various energy-level transitions. Niels Bohr's model of the hydrogen atom provided an explanation for these transitions.

Whenever an excited hydrogen atom falls to its ground state or to a lower-energy excited state, it emits a photon of radiation. The energy of this photon ($E_{photon} = h\nu$) is equal to the difference in energy between the atom's initial state and its final state, as illustrated in **Figure 7.** The fact that hydrogen atoms emit only specific frequencies of light indicated that the energy differences between the atoms' energy states were fixed. This suggested that the electron of a hydrogen atom exists only in very specific energy states.

In the late nineteenth century, a mathematical formula that related the various wavelengths of hydrogen's emission-line spectrum was discovered. The challenge facing scientists was to provide a model of the hydrogen atom that accounted for this relationship.
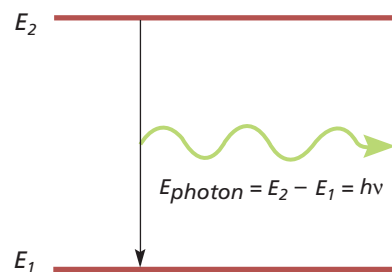

FIGURE 7 When an excited atom with energy E_2 falls back to energy E_1, it releases a photon that has energy $E_2 - E_1 = E_{photon} = h\nu$.

Bohr Model of the Hydrogen Atom

The puzzle of the hydrogen-atom spectrum was solved in 1913 by the Danish physicist Niels Bohr. He proposed a hydrogen-atom model that linked the atom's electron to photon emission. According to the model, the electron can circle the nucleus only in allowed paths, or *orbits*. When the electron is in one of these orbits, the atom has a definite, fixed energy. The electron—and therefore the hydrogen atom—is in its lowest energy state when it is in the orbit closest to the nucleus. This orbit is separated from the nucleus by a large empty space where the electron cannot exist. The energy of the electron is higher when the electron is in orbits that are successively farther from the nucleus.

The electron orbits, or atomic energy levels, in Bohr's model can be compared to the rungs of a ladder. When you are standing on a ladder, your feet are on one rung or another. The amount of potential energy that you possess corresponds to standing on the first rung, the second rung, and so forth. Your energy cannot correspond to standing between two rungs because you cannot stand in midair. In the same way, an electron can be in one orbit or another, but not in between.

How does Bohr's model of the hydrogen atom explain the observed spectral lines? While in a given orbit, the electron is neither gaining nor losing energy. It can, however, move to a higher-energy orbit by gaining an amount of energy equal to the difference in energy between the higher-energy orbit and the initial lower-energy orbit. When a hydrogen atom is in an excited state, its electron is in one of the higher-energy orbits. When the electron falls to a lower energy level, a photon is emitted, and the process is called *emission*. The photon's energy is equal to the energy difference between the initial higher energy level and the final lower energy level. Energy must be added to an atom in order to move an electron from a lower energy level to a higher energy level. This process is called *absorption*. Absorption and emission of radiation in Bohr's model of the hydrogen atom are illustrated in **Figure 8.** The energy of each absorbed or emitted photon corresponds to a particular frequency of emitted radiation, $E_{photon} = h\nu$.

Based on the different wavelengths of the hydrogen emission-line spectrum, Bohr calculated the allowed energy levels for the hydrogen

FIGURE 8 (a) Absorption and (b) emission of a photon by a hydrogen atom according to Bohr's model. The frequencies of light that can be absorbed and emitted are restricted because the electron can only be in orbits corresponding to the energies E_1, E_2, E_3, and so forth.

$E_{photon} = E_2 - E_1$

$E_{photon} = E_3 - E_1$
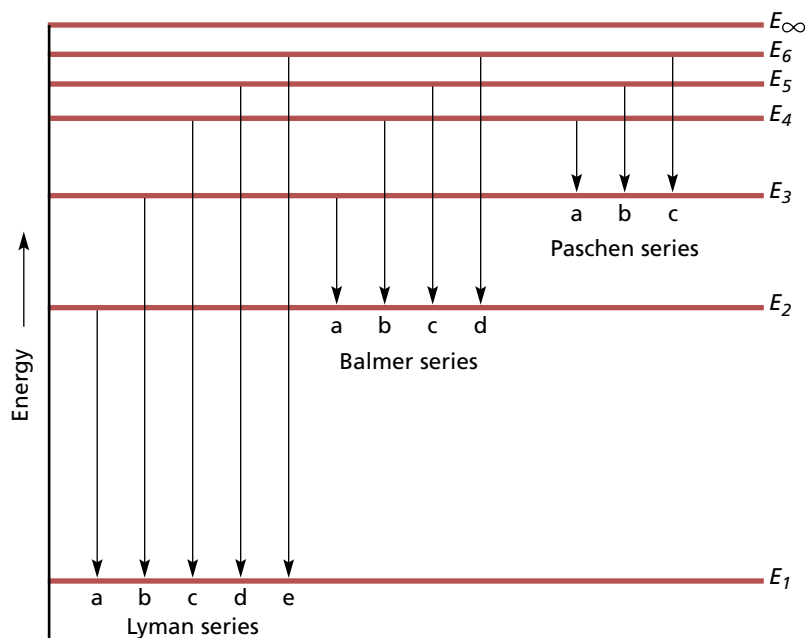
(a) Absorption

(b) Emission
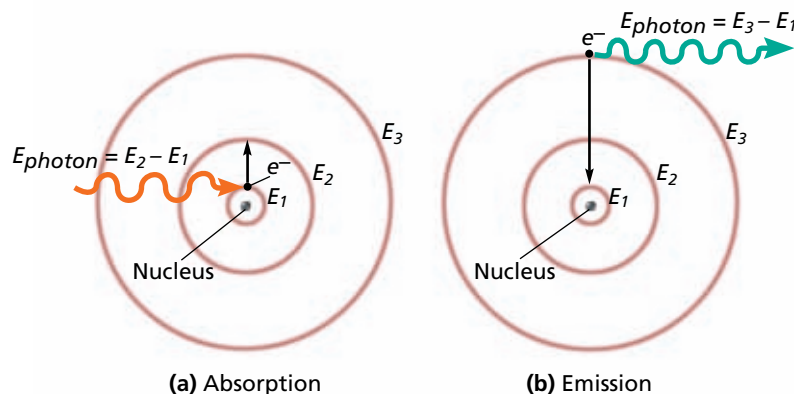
FIGURE 9 This energy-state diagram for a hydrogen atom shows some of the energy transitions for the Lyman, Balmer, and Paschen spectral series. Bohr's model of the atom accounted mathematically for the energy of each of the transitions shown.

atom. He then related the possible energy-level changes to the lines in the hydrogen emission-line spectrum. The five lines in the Lyman series, for example, were shown to be the result of electrons dropping from energy levels E_6, E_5, E_4, E_3, and E_2 to the ground-state energy level E_1.

Bohr's calculated values agreed with the experimentally observed values for the lines in each series. The origins of three of the series of lines in hydrogen's emission-line spectrum are shown in **Figure 9.**

Bohr's model of the hydrogen atom explained observed spectral lines so well that many scientists concluded that the model could be applied to all atoms. It was soon recognized, however, that Bohr's approach did not explain the spectra of atoms with more than one electron. Nor did Bohr's theory explain the chemical behavior of atoms.

SECTION REVIEW

1. What was the major shortcoming of Rutherford's model of the atom?

2. Write and label the equation that relates the speed, wavelength, and frequency of electromagnetic radiation.

3. Define the following:

   a. electromagnetic radiation   **b.** wavelength

   c. frequency   **d.** quantum   **e.** photon

4. What is meant by the dual wave-particle nature of light?

5. Describe the Bohr model of the hydrogen atom.

Critical Thinking

6. **INTERPRETING GRAPHICS** Use the diagram in **Figure 9** to answer the following:

   a. Characterize each of the following as absorption or emission: an electron moves from E_2 to E_1; an electron moves from E_1 to E_3; and an electron moves from E_6 to E_3.

   b. Which energy-level change above emits or absorbs the highest energy? the lowest energy?

The Quantum Model of the Atom

*O*BJECTIVES

- Discuss Louis de Broglie's role in the development of the quantum model of the atom.

- Compare and contrast the Bohr model and the quantum model of the atom.

- Explain how the Heisenberg uncertainty principle and the Schrödinger wave equation led to the idea of atomic orbitals.

- List the four quantum numbers and describe their significance.

- Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

To the scientists of the early twentieth century, Bohr's model of the hydrogen atom contradicted common sense. Why did hydrogen's electron exist around the nucleus only in certain allowed orbits with definite energies? Why couldn't the electron exist in a limitless number of orbits with slightly different energies? To explain why atomic energy states are quantized, scientists had to change the way they viewed the nature of the electron.

Electrons as Waves

The investigations into the photoelectric effect and hydrogen's emission-line spectrum revealed that light could behave as both a wave and a particle. Could electrons have a dual wave-particle nature as well? In 1924, the French scientist Louis de Broglie asked himself this very question. And the answer that he proposed led to a revolution in our basic understanding of matter.

De Broglie pointed out that in many ways the behavior of electrons in Bohr's quantized orbits was similar to the known behavior of waves. For example, scientists at the time knew that any wave confined to a space can have only certain frequencies. De Broglie suggested that electrons be considered waves confined to the space around an atomic nucleus. It followed that the electron waves could exist only at specific frequencies. And according to the relationship $E = h\nu$, these frequencies corresponded to specific energies—the quantized energies of Bohr's orbits.

Other aspects of de Broglie's hypothesis that electrons have wave-like properties were soon confirmed by experiments. Investigators demonstrated that electrons, like light waves, can be bent, or diffracted. *Diffraction* refers to the bending of a wave as it passes by the edge of an object or through a small opening. Diffraction experiments and other investigations also showed that electron beams, like waves, can interfere with each other. *Interference* occurs when waves overlap (see the Quick Lab in this section). This overlapping results in a reduction of energy in some areas and an increase of energy in others. The effects of diffraction and interference can be seen in **Figure 10.**
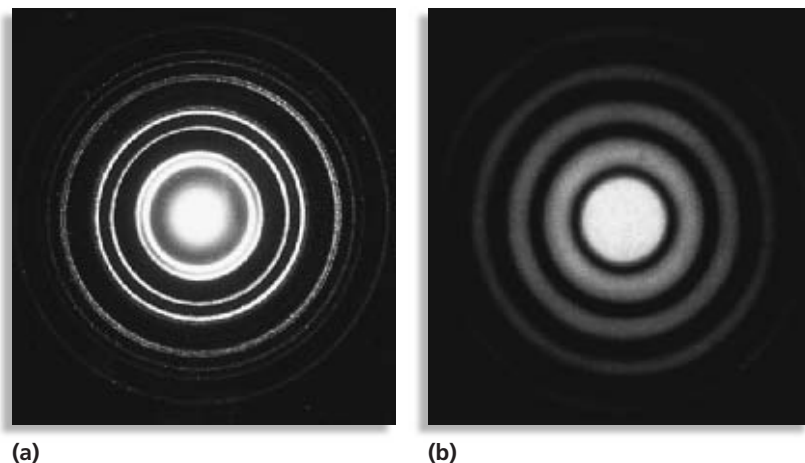
(a)

(b)

FIGURE 10 Diffraction patterns produced by (a) a beam of electrons passed through a substance and (b) a beam of visible light passed through a tiny aperture. Each pattern shows the results of bent waves that have interfered with each other. The bright areas correspond to areas of increased energy, while the dark areas correspond to areas of decreased energy.

The Heisenberg Uncertainty Principle

The idea of electrons having a dual wave-particle nature troubled scientists. If electrons are both particles and waves, then where are they in the atom? To answer this question, it is important to consider a proposal first made in 1927 by the German theoretical physicist Werner Heisenberg.

Heisenberg's idea involved the detection of electrons. Electrons are detected by their interaction with photons. Because photons have about the same energy as electrons, any attempt to locate a specific electron with a photon knocks the electron off its course. As a result, there is always a basic uncertainty in trying to locate an electron (or any other particle). *The* **Heisenberg uncertainty principle** *states that it is impossible to determine simultaneously both the position and velocity of an electron or any other particle*. Although it was difficult for scientists to accept this fact at the time, it has proven to be one of the fundamental principles of our present understanding of light and matter.

The Schrödinger Wave Equation

In 1926, the Austrian physicist Erwin Schrödinger used the hypothesis that electrons have a dual wave-particle nature to develop an equation that treated electrons in atoms as waves. Unlike Bohr's theory, which assumed quantization as a fact, quantization of electron energies was a natural outcome of Schrödinger's equation. Only waves of specific energies, and therefore frequencies, provided solutions to the equation. Together with the Heisenberg uncertainty principle, the Schrödinger wave equation laid the foundation for modern quantum theory. **Quantum theory** *describes mathematically the wave properties of electrons and other very small particles.*

QuickLAB

The Wave Nature of Light: Interference

Question

Does light show the wave property of interference when a beam of light is projected through a pinhole onto a screen?

Procedure

Record all your observations.

1. To make the pinhole screen, cut a 20 cm × 20 cm square from a manila folder. In the center of the square, cut a 2 cm square hole. Cut a 7 cm × 7 cm square of aluminum foil. Using a thumbtack, make a pinhole in the center of the foil square. Tape the aluminum foil over the 2 cm square hole, making sure the pinhole is centered as shown in the diagram.

2. Use white poster board to make a projection screen 35 cm × 35 cm.

3. In a dark room, center the light beam from a flashlight on the pinhole. Hold the flashlight about 1 cm from the pinhole. The pinhole screen should be about 50 cm from the projection screen, as shown in the diagram. Adjust the distance to form a sharp image on the projection screen.

Discussion

1. Did you observe interference patterns on the screen?

2. As a result of your observations, what do you conclude about the nature of light?

Materials

- scissors
- manila folders
- thumbtack
- masking tape
- aluminum foil
- white poster board or cardboard
- flashlight
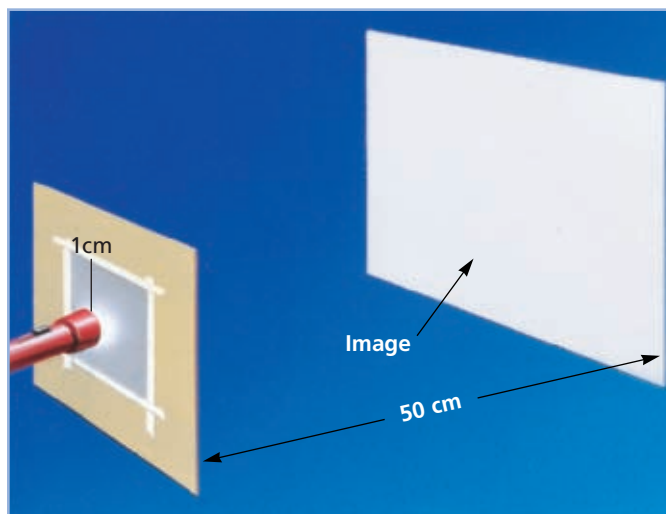
Solutions to the Schrödinger wave equation are known as wave functions. Based on the Heisenberg uncertainty principle, the early developers of quantum theory determined that wave functions give only the *probability* of finding an electron at a given place around the nucleus. Thus, electrons do not travel around the nucleus in neat orbits, as Bohr had postulated. Instead, they exist in certain regions called orbitals. *An* **orbital** *is a three-dimensional region around the nucleus that indicates the probable location of an electron.*

Figure 11 illustrates two ways of picturing one type of atomic orbital. As you will see later in this section, atomic orbitals have different shapes and sizes.
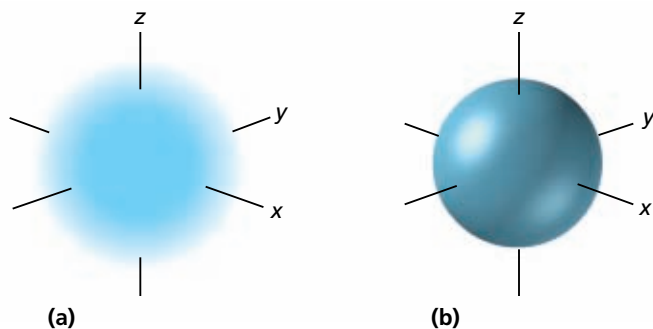
(a)   **(b)**

FIGURE 11 Two ways of showing a simple atomic orbital are presented. In (a) the probability of finding the electron is proportional to the density of the cloud. Shown in (b) is a surface within which the electron can be found a certain percentage of the time, conventionally 90%.

Atomic Orbitals and Quantum Numbers

In the Bohr atomic model, electrons of increasing energy occupy orbits farther and farther from the nucleus. According to the Schrödinger equation, electrons in atomic orbitals also have quantized energies. An electron's energy level is not the only characteristic of an orbital that is indicated by solving the Schrödinger equation.

In order to completely describe orbitals, scientists use quantum numbers. **Quantum numbers** *specify the properties of atomic orbitals and the properties of electrons in orbitals*. The first three quantum numbers result from solutions to the Schrödinger equation. They indicate the main energy level, the shape, and the orientation of an orbital. The fourth, the spin quantum number, describes a fundamental state of the electron that occupies the orbital. As you read the following descriptions of the quantum numbers, refer to the appropriate columns in **Table 2.**

Principal Quantum Number

The **principal quantum number,** *symbolized by n, indicates the main energy level occupied by the electron.* Values of n are positive integers only—1, 2, 3, and so on. As n increases, the electron's energy and its average distance from the nucleus increase (see **Figure 12**). For example, an electron for which $n = 1$ occupies the first, or lowest, main energy level and is located closest to the nucleus. As you will see, more than one electron can have the same n value. These electrons are sometimes said to be in the same electron *shell.* The total number of orbitals that exist in a given shell, or main energy level, is equal to n^2.

Angular Momentum Quantum Number

Except at the first main energy level, orbitals of different shapes—known as *sublevels*—exist for a given value of n. *The* **angular momentum quantum number,** *symbolized by l, indicates the shape of the orbital.* For a specific main energy level, the number of orbital shapes possible is equal to n. The values of l allowed are zero and all positive integers less
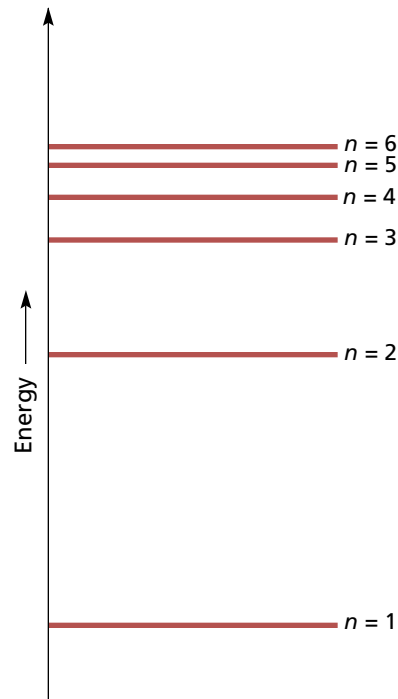

FIGURE 12 The main energy levels of an atom are represented by the principal quantum number, n.

TABLE 1 *Orbital Letter Designations According to Values of l*

l	Letter
0	s
1	p
2	d
3	f

than or equal to $n - 1$. For example, orbitals for which $n = 2$ can have one of two shapes corresponding to $l = 0$ and $l = 1$. Depending on its value of l, an orbital is assigned a letter, as shown in **Table 1.**

As shown in **Figure 13,** s orbitals are spherical, p orbitals have dumbbell shapes, and d orbitals are more complex. (The f orbital shapes are even more complex.) In the first energy level, $n = 1$, there is only one sublevel possible—an s orbital. As mentioned, the second energy level, $n = 2$, has two sublevels—the s and p orbitals. The third energy level, $n = 3$, has three sublevels—the s, p, and d orbitals. The fourth energy level, $n = 4$, has four sublevels—the s, p, d, and f orbitals. In an nth main energy level, there are n sublevels.

Each atomic orbital is designated by the principal quantum number followed by the letter of the sublevel. For example, the 1s sublevel is the s orbital in the first main energy level, while the 2p sublevel is the set of three p orbitals in the second main energy level. On the other hand, a 4d orbital is part of the d sublevel in the fourth main energy level. How would you designate the p sublevel in the third main energy level? How many other sublevels are in the third main energy level with this one?

Magnetic Quantum Number

Atomic orbitals can have the same shape but different orientations around the nucleus. *The* **magnetic quantum number,** *symbolized by m, indicates the orientation of an orbital around the nucleus.* Values of m are whole numbers, including zero, from $-l$ to $+l$. Because an s orbital is spherical and is centered around the nucleus, it has only one possible orientation. This orientation corresponds to a magnetic quantum num-

FIGURE 13 The orbitals s, p, and d have different shapes. Each of the orbitals shown occupies a different region of space around the nucleus.
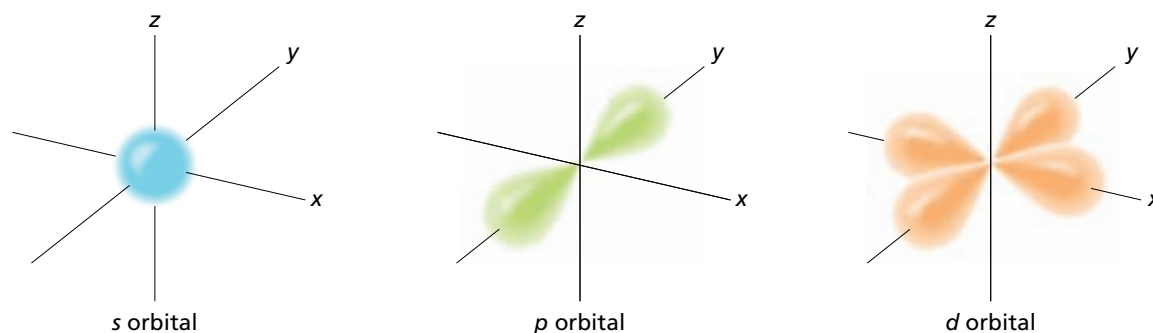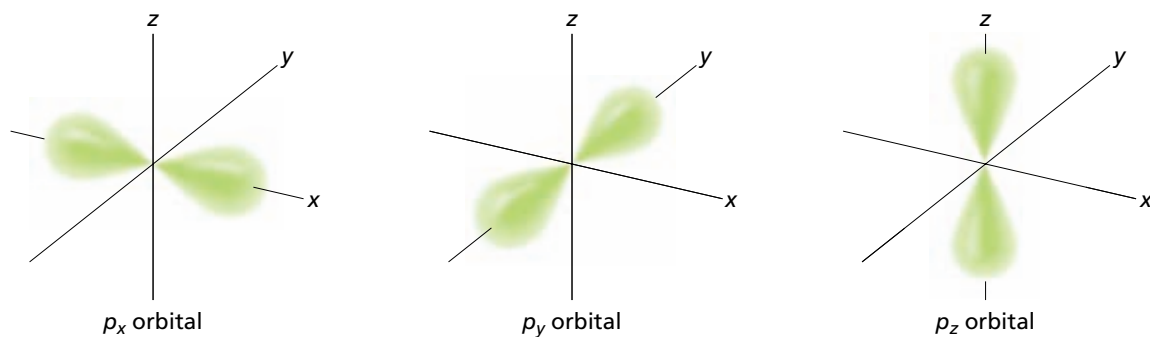

s orbital

p orbital

d orbital

p_x orbital

p_y orbital

p_z orbital

ber of $m = 0$. There is therefore only one *s* orbital in each *s* sublevel. As shown in **Figure 14,** the lobes of a *p* orbital extend along the *x, y,* or *z* axis of a three-dimensional coordinate system. There are therefore three *p* orbitals in each *p* sublevel, which are designated as p_x, p_y, and p_z orbitals. The three *p* orbitals occupy different regions of space and those regions are related to values of $m = -1$, $m = 0$, and $m = +1$.

There are five different *d* orbitals in each *d* sublevel (see **Figure 15**). The five different orientations, including one with a different shape, correspond to values of $m = -2$, $m = -1$, $m = 0$, $m = +1$, and $m = +2$. There are seven different *f* orbitals in each *f* sublevel.
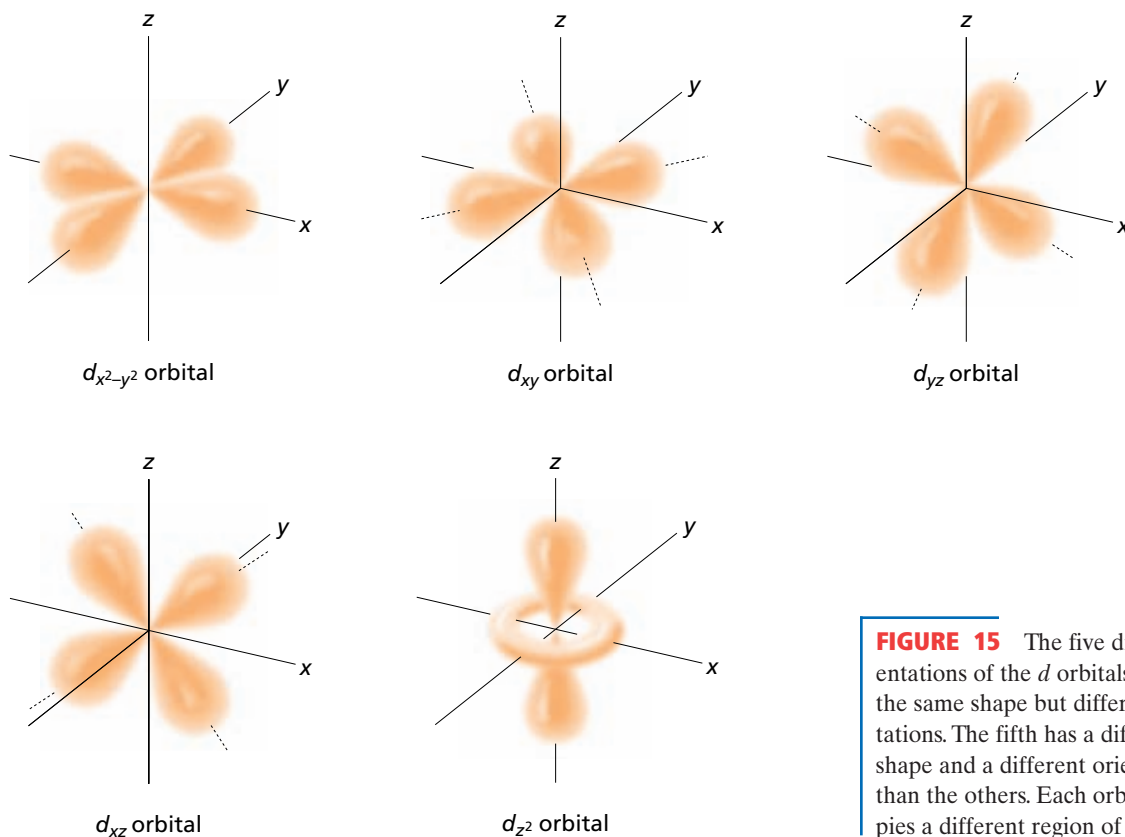

$d_{x^2-y^2}$ orbital

d_{xy} orbital

d_{yz} orbital

d_{xz} orbital

d_{z^2} orbital

TABLE 2 Quantum Number Relationships in Atomic Structure

Principal quantum number: main energy level (n)	Sublevels in main energy level (n sublevels)	Number of orbitals per sublevel	Number of orbitals per main energy level (n^2)	Number of electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	1	2	2
2	s	1	4	2	8
	p	3		6	
3	s	1	9	2	18
	p	3		6	
	d	5		10	
4	s	1	16	2	32
	p	3		6	
	d	5		10	
	f	7		14	

As you can see in **Table 2,** the total number of orbitals in a main energy level increases with the value of n. In fact, the number of orbitals at each main energy level equals the square of the principal quantum number, n^2. What is the total number of orbitals in the third energy level? Specify each of the sublevels using the orbital designations you've learned so far.

Spin Quantum Number

An electron in an orbital behaves in some ways like Earth spinning on an axis. The electron exists in one of two possible spin states, which creates a magnetic field. To account for the magnetic properties of the electron, theoreticians of the early twentieth century created the spin quantum number. *The **spin quantum number** has only two possible values—$(+\frac{1}{2}, -\frac{1}{2})$—which indicate the two fundamental spin states of an electron in an orbital.* A single orbital can hold a maximum of two electrons, but the two electrons must have opposite spin states.

SECTION REVIEW

1. Define the following:

    a. main energy levels

    b. quantum numbers

2. a. List the four quantum numbers.

    b. What general information about atomic orbitals is provided by the quantum numbers?

3. Describe briefly what specific information is given by each of the four quantum numbers.

Critical Thinking

4. INFERRING RELATIONSHIPS What are the possible values of the magnetic quantum number m for f orbitals? What is the maximum number of electrons that can exist in $4f$ orbitals?

Electron Configurations

OBJECTIVES

- List the total number of electrons needed to fully occupy each main energy level.

- State the Aufbau principle, the Pauli exclusion principle, and Hund's rule.

- Describe the electron configurations for the atoms of any element using orbital notation, electron-configuration notation, and, when appropriate, noble-gas notation.

The quantum model of the atom improves on the Bohr model because it describes the arrangements of electrons in atoms other than hydrogen. *The arrangement of electrons in an atom is known as the atom's* **electron configuration.** Because atoms of different elements have different numbers of electrons, a unique electron configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest-energy arrangement of the electrons for each element is called the element's *ground-state electron configuration.* A few simple rules, combined with the quantum number relationships discussed in Section 2, allow us to determine these ground-state electron configurations.

Rules Governing Electron Configurations

To build up electron configurations for the ground state of any particular atom, first the energy levels of the orbitals are determined. Then electrons are added to the orbitals, one by one, according to three basic rules. (Remember that real atoms are not built up by adding protons and electrons one at a time.)

The first rule shows the order in which electrons occupy orbitals. According to the **Aufbau principle,** *an electron occupies the lowest-energy orbital that can receive it.* **Figure 16** shows the atomic orbitals in order of increasing energy. The orbital with the lowest energy is the 1s orbital. In a ground-state hydrogen atom, the electron is in this orbital. The 2s orbital is the next highest in energy, then the 2p orbitals. Beginning with the third main energy level, $n = 3$, the energies of the sublevels in different main energy levels begin to overlap.

Note in the figure, for example, that the 4s sublevel is lower in energy than the 3d sublevel. Therefore, the 4s orbital is filled before any electrons enter the 3d orbitals. (Less energy is required for two electrons to pair up in the 4s orbital than for those two electrons to
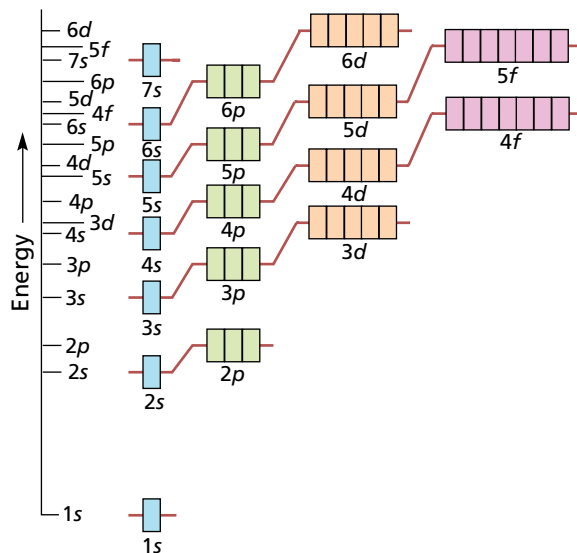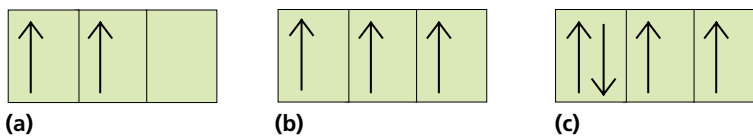
FIGURE 16 The order of increasing energy for atomic sublevels is shown on the vertical axis. Each individual box represents an orbital.

1s orbital

FIGURE 17 According to the Pauli exclusion principle, an orbital can hold two electrons of opposite spin states. In this electron configuration of a helium atom, each arrow represents one of the atom's two electrons. The direction of the arrow indicates the electron's spin state.

occupy a 3d orbital.) Once the 3d orbitals are fully occupied, which sublevel will be occupied next?

The second rule reflects the importance of the spin quantum number. According to the **Pauli exclusion principle,** *no two electrons in the same atom can have the same set of four quantum numbers.* The principal, angular momentum, and magnetic quantum numbers specify the energy, shape, and orientation of an orbital. The two values of the spin quantum number reflect the fact that for two electrons to occupy the same orbital, they must have opposite spin states (see **Figure 17**).

The third rule requires placing as many unpaired electrons as possible in separate orbitals in the same sublevel. In this way, electron-electron repulsion is minimized so that the electron arrangements have the lowest energy possible. According to **Hund's rule,** *orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin state.* Applying this rule shows, for example, that one electron will enter each of the three p orbitals in a main energy level before a second electron enters any of them. This is illustrated in **Figure 18.** What is the maximum number of unpaired electrons in a d sublevel?

FIGURE 18 The figure shows how (a) two, (b) three, and (c) four electrons fill the p sublevel of a given main energy level according to Hund's rule.

(a)   (b)   (c)

Representing Electron Configurations

Three methods, or notations, are used to indicate electron configurations. Two of these notations will be discussed in the next two sections for the first-period elements, hydrogen and helium. The third notation is used mostly with elements of the third period and higher. It will be discussed in the section on third-period elements.

In a ground-state hydrogen atom, the single electron is in the lowest-energy orbital, the 1s orbital. The electron can be in either one of its two spin states. Helium has two electrons, which are paired in the 1s orbital.

Orbital Notation

In orbital notation, an unoccupied orbital is represented by a line, ___, with the orbital's name written underneath the line. An orbital containing one electron is represented as _↑_ . An orbital containing two electrons is represented as _↑↓_ , showing the electrons paired and with opposite spin states. The lines are labeled with the principal quantum number and

sublevel letter. For example, the orbital notations for hydrogen and helium are written as follows.

$$H \; \underset{1s}{\uparrow} \qquad He \; \underset{1s}{\uparrow\downarrow}$$

Electron-Configuration Notation

Electron-configuration notation eliminates the lines and arrows of orbital notation. Instead, the number of electrons in a sublevel is shown by adding a superscript to the sublevel designation. The hydrogen configuration is represented by $1s^1$. The superscript indicates that one electron is present in hydrogen's $1s$ orbital. The helium configuration is represented by $1s^2$. Here the superscript indicates that there are two electrons in helium's $1s$ orbital.

SAMPLE PROBLEM A

For more help, go to the Math Tutor at the end of Chapter 5.

The electron configuration of boron is $1s^2 2s^2 2p^1$. How many electrons are present in an atom of boron? What is the atomic number for boron? Write the orbital notation for boron.

SOLUTION   The number of electrons in a boron atom is equal to the sum of the superscripts in its electron-configuration notation: $2 + 2 + 1 = 5$ electrons. The number of protons equals the number of electrons in a neutral atom. So we know that boron has 5 protons and thus has an atomic number of 5. To write the orbital notation, first draw the lines representing orbitals.

$$\overline{}_{1s} \quad \overline{}_{2s} \quad \underset{2p}{\overline{}\;\overline{}\;\overline{}}$$

Next, add arrows showing the electron locations. The first two electrons occupy $n = 1$ energy level and fill the $1s$ orbital.

$$\underset{1s}{\uparrow\downarrow} \quad \overline{}_{2s} \quad \underset{2p}{\overline{}\;\overline{}\;\overline{}}$$
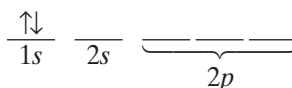
The next three electrons occupy the $n = 2$ main energy level. Two of these occupy the lower-energy $2s$ orbital. The third occupies a higher-energy p orbital.

$$\underset{1s}{\uparrow\downarrow} \quad \underset{2s}{\uparrow\downarrow} \quad \underset{2p}{\uparrow\;\overline{}\;\overline{}}$$
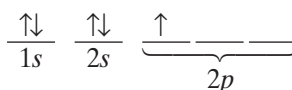
PRACTICE

Answers in Appendix E

1. The electron configuration of nitrogen is $1s^2 2s^2 2p^3$. How many electrons are present in a nitrogen atom? What is the atomic number of nitrogen? Write the orbital notation for nitrogen.

2. The electron configuration of fluorine is $1s^2 2s^2 2p^5$. What is the atomic number of fluorine? How many of its p orbitals are filled?

extension
Go to **go.hrw.com** for more practice problems that deal with electron configurations.

Keyword: HC6ARRX

HISTORICAL CHEMISTRY

The Noble Decade

By the late nineteenth century, the science of chemistry had begun to be organized. In 1860, the First International Congress of Chemistry established the field's first standards. And Dmitri Mendeleev's periodic table of elements gave chemists across the globe a systematic understanding of matter's building blocks. But many important findings—including the discovery of a family of rare, unreactive gases that were unlike any substances known at the time—were yet to come.

Cross-Disciplinary Correspondence

In 1888, the British physicist Lord Rayleigh encountered a small but significant discrepancy in the results of one of his experiments. In an effort to redetermine the atomic mass of nitrogen, he measured the densities of several samples of nitrogen gas. Each sample had been prepared by a different method. All samples that had been isolated from chemical reactions exhibited similar densities. But they were about one-tenth of a percent less dense than the nitrogen isolated from air, which at the time was believed to be a mixture of nitrogen, oxygen, water vapor, and carbon dioxide.

This excerpt from Lord Rayleigh's letter was originally published in Nature *magazine in 1892.*▼

SEPTEMBER 29, 1892   NATURE

LETTERS TO THE EDITOR.

[*The Editor does not hold himself responsible for opinions expressed by his correspondents. Neither can he undertake to return, or to correspond with the writers of, rejected manuscripts intended for this or any other part of* NATURE. *No notice is taken of anonymous communications.*]

Density of Nitrogen.

I AM much puzzled by some recent results as to the density of *nitrogen*, and shall be obliged if any of your chemical readers can offer suggestions as to the cause. According to two methods of preparation I obtain quite distinct values. The relative difference, amounting to about ¹⁄₁₀₀₀ part, is small in itself; but it lies entirely outside the errors of experiment, and can only be attributed to a variation in the character of the gas...

Is it possible that the difference is independent of impurity, the nitrogen itself being to some extent in a different (dissociated) state ?...

RAYLEIGH.

Terling Place, Witham, September 24.

Rayleigh was at a loss to explain his discovery. Finally, in 1892, he published a letter in *Nature* magazine to appeal to his colleagues for an explanation. A month later, he received a reply from a Scottish chemist named William Ramsay. Ramsay related that he too had been stumped by the density difference between chemical and atmospheric nitrogen. Rayleigh decided to report his findings to the Royal Society.

A Chemist's Approach

With Rayleigh's permission, Ramsay attempted to remove all known components from a sample of air and to analyze what, if anything, remained. Having removed water vapor, carbon dioxide, and oxygen from the air, Ramsay repeatedly passed the sample over hot magnesium. The nitrogen reacted with the magnesium to form solid magnesium nitride. As a result, all of the then-known components of air were removed. What remained was a minuscule portion of a mysterious gas.

Ramsay tried to cause the gas to react with chemically active substances, such as hydrogen, sodium, and caustic soda, but the gas remained unaltered. He decided to name this new atmospheric component *argon* (Greek for "inert" or "idle").

Periodic Problems

Rayleigh and Ramsay were sure that they had discovered a new element. But this created a problem. Their calculations indicated that argon had an atomic mass of about 40. However, as it appeared in 1894, the periodic table had no space for such an element. The elements with atomic masses closest to that of argon were chlorine and potassium. Unfortunately, the chemical properties of the families of each of these elements were completely dissimilar to those of the strange gas.

Ramsay contemplated argon's lack of reactivity. He knew that Mendeleev had created the periodic table on the basis of valence, or the number of atomic partners an element bonds with in forming a compound. Because Ramsay could not cause argon to form any compounds, he assigned it a valence of zero. And because the valence of the elements in
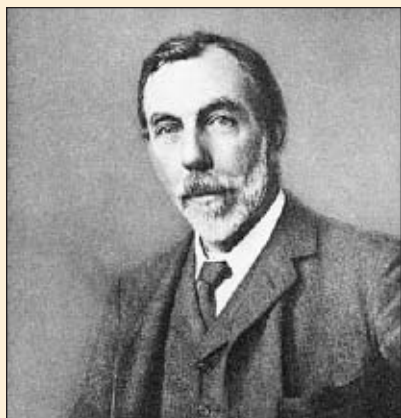
In 1893, Scottish chemist William Ramsay isolated a previously unknown component of the atmosphere.

Groups / Periods	III b	IV b	V b	VI b	VII b	VIII b	I b	II b	III a	IV a	V a	VI a	VII a	0	I a	II a
1															H	He
2															Li	Be
3									B	C	N	O	F	Ne	Na	Mg
4									Al	Si	P	S	Cl	Ar	K	Ca
5	Sc	Ti	V	Cr	Mn	Fe Co Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	Rb	Sr
6	Y	Zr	Nb	Mo	Tc	Ru Rh Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	Cs	Ba
7	La	Hf	Ta	W	Re	Os Ir Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	Fr	Ra
8	Ac															

Transition elements   Main-group elements

This version of the periodic table shows how it looked after the discovery of the noble gases. The placement of the Group 1 and 2 elements at the far right of the table shows clearly how the noble gases fit in between the chlorine family and the potassium family of elements. The 0 above the noble-gas family indicates the zero valency of the gases.

the families of both chlorine and potassium was one, perhaps argon fit in between them.

Ramsay's insight that argon merited a new spot between the halogen family and the alkali metal family on the periodic table was correct. And as Ramsay would soon confirm, his newly discovered gas was indeed one of a previously unknown family of elements.

New Neighbors

In 1895, Ramsay isolated a light, inert gas from a mineral called *cleveite.* Physical analysis revealed that the gas was the same as one that had been identified in the sun in 1868—helium. Helium was the second zero-valent element found on Earth, and its discovery made chemists aware that the periodic table had been missing a whole column of elements.

Over the next three years, Ramsay and his assistant, Morris Travers, identified three more inert gases present in the atmosphere: neon (Greek for "new"), krypton ("hidden"), and xenon ("stranger"). Finally in 1900, German chemist Friedrich Ernst Dorn discovered radon, the last of the new family of elements known today as the *noble gases.* For his discovery, Ramsay received the Nobel Prize in 1904.

Questions

1. What evidence led Ramsay to report that the mysterious gas was inert?

2. What property of argon caused Ramsay to propose a new column in the periodic table?

SC*L*INKS.

www.scilinks.org
Topic: William Ramsay
Code: HC61666

1s

2s 2p

3s 3p 3d
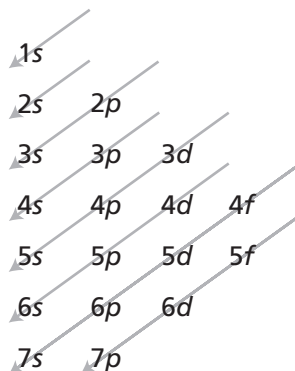
4s 4p 4d 4f

5s 5p 5d 5f

6s 6p 6d

7s 7p

FIGURE 19 Follow the diagonal arrows from the top to get the order in which atomic orbitals are filled according to the Aufbau principle.

Elements of the Second Period

In the first-period elements, hydrogen and helium, electrons occupy the orbital of the first main energy level. The ground-state configurations in **Table 3** illustrate how the Aufbau principle, the Pauli exclusion principle, and Hund's rule are applied to atoms of elements in the second period. **Figure 19** provides a pattern to help you remember the order in which orbitals are filled according to the Aufbau principle.

According to the Aufbau principle, after the $1s$ orbital is filled, the next electron occupies the s sublevel in the second main energy level. Thus, lithium, Li, has a configuration of $1s^2 2s^1$. The electron occupying the $2s$ level of a lithium atom is in the atom's highest, or outermost, occupied level. The *highest-occupied energy level* is the electron-containing main energy level with the highest principal quantum number. The two electrons in the $1s$ sublevel of lithium are no longer in the outermost main energy level. They have become *inner-shell electrons,* which are electrons that are not in the highest-occupied energy level.

The fourth electron in an atom of beryllium, Be, must complete the pair in the $2s$ sublevel because this sublevel is of lower energy than the $2p$ sublevel. With the $2s$ sublevel filled, the $2p$ sublevel, which has three vacant orbitals of equal energy, can be occupied. One of the three p orbitals is occupied by a single electron in an atom of boron, B. Two of the three p orbitals are occupied by unpaired electrons in an atom of carbon, C. And all three p orbitals are occupied by unpaired electrons in an atom of nitrogen, N. Hund's rule applies here, as is shown in the orbital notations in **Table 3.**

According to the Aufbau principle, the next electron must pair with another electron in one of the $2p$ orbitals rather than enter the third main energy level. The Pauli exclusion principle allows the electron to pair with

TABLE 3 *Electron Configurations of Atoms of Second-Period Elements Showing Two Notations*

Name	Symbol	1s	2s	2p			Electron-configuration notation
Lithium	Li	↑↓	↑	__	__	__	$1s^2 2s^1$
Beryllium	Be	↑↓	↑↓	__	__	__	$1s^2 2s^2$
Boron	B	↑↓	↑↓	↑	__	__	$1s^2 2s^2 2p^1$
Carbon	C	↑↓	↑↓	↑	↑	__	$1s^2 2s^2 2p^2$
Nitrogen	N	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
Oxygen	O	↑↓	↑↓	↑↓	↑	↑	$1s^2 2s^2 2p^4$
Fluorine	F	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^5$
Neon	Ne	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$

one of the electrons occupying the $2p$ orbitals as long as the spins of the paired electrons are opposite. Thus, atoms of oxygen, O, have the configuration $1s^2 2s^2 2p^4$. Oxygen's orbital notation is shown in **Table 3.**

Two $2p$ orbitals are filled in fluorine, F, and all three are filled in neon, Ne. Atoms such as those of neon, which have the s and p sublevels of their highest occupied level filled with eight electrons, are said to have an *octet* of electrons. Examine the periodic table inside the back cover of the text. Notice that neon is the last element in the second period.

Elements of the Third Period

After the outer octet is filled in neon, the next electron enters the s sublevel in the $n = 3$ main energy level. Thus, atoms of sodium, Na, have the configuration $1s^2 2s^2 2p^6 3s^1$. Compare the configuration of a sodium atom with that of an atom of neon in **Table 3.** Notice that the first 10 electrons in a sodium atom have the same configuration as a neon atom, $1s^2 2s^2 2p^6$. In fact, the first 10 electrons in an atom of each of the third-period elements have the same configuration as neon. This similarity allows us to use a shorthand notation for the electron configurations of the third-period elements.

Noble-Gas Notation

Neon is a member of the Group 18 elements. *The Group 18 elements (helium, neon, argon, krypton, xenon, and radon) are called the* **noble gases.** To simplify sodium's notation, the symbol for neon, enclosed in square brackets, is used to represent the complete neon configuration: $[Ne] = 1s^2 2s^2 2p^6$. This allows us to write sodium's electron configuration as $[Ne]3s^1$, which is called sodium's *noble-gas notation.* **Table 4** shows the electron configuration of each of the third-period elements using noble-gas notation.

TABLE 4 *Electron Configurations of Atoms of Third-Period Elements*

Name	Symbol	Atomic number	Number of electrons in sublevels					Noble-gas notation
			1s	2s	2p	3s	3p	
Sodium	Na	11	2	2	6	1		*$[Ne]3s^1$
Magnesium	Mg	12	2	2	6	2		$[Ne]3s^2$
Aluminum	Al	13	2	2	6	2	1	$[Ne]3s^2 3p^1$
Silicon	Si	14	2	2	6	2	2	$[Ne]3s^2 3p^2$
Phosphorus	P	15	2	2	6	2	3	$[Ne]3s^2 3p^3$
Sulfur	S	16	2	2	6	2	4	$[Ne]3s^2 3p^4$
Chlorine	Cl	17	2	2	6	2	5	$[Ne]3s^2 3p^5$
Argon	Ar	18	2	2	6	2	6	$[Ne]3s^2 3p^6$

*$[Ne] = 1s^2 2s^2 2p^6$

The last element in the third period is argon, Ar, which is a noble gas. As in neon, the highest-occupied energy level of argon has an octet of electrons, $[Ne]3s^23p^6$. In fact, each noble gas other than He has an electron octet in its highest energy level. *A noble-gas configuration* refers to *an outer main energy level occupied, in most cases, by eight electrons.*

Elements of the Fourth Period

The electron configurations of atoms in the fourth-period elements are shown in **Table 5.** The period begins by filling the $4s$ orbital, the empty orbital of lowest energy. Thus, the first element in the fourth period is potassium, K, which has the electron configuration $[Ar]4s^1$. The next element is calcium, Ca, which has the electron configuration $[Ar]4s^2$.

With the $4s$ sublevel filled, the $4p$ and $3d$ sublevels are the next available vacant orbitals. **Figure 16** shows that the $3d$ sublevel is lower in energy than the $4p$ sublevel. Therefore, the five $3d$ orbitals are next to

TABLE 5 *Electron Configuration of Atoms of Elements in the Fourth Period*

Name	Symbol	Atomic number	Number of electrons in sublevels above 2p					Noble-gas notation
			$3s$	$3p$	$3d$	$4s$	$4p$	
Potassium	K	19	2	6		1		*$[Ar]4s^1$
Calcium	Ca	20	2	6		2		$[Ar]4s^2$
Scandium	Sc	21	2	6	1	2		$[Ar]3d^14s^2$
Titanium	Ti	22	2	6	2	2		$[Ar]3d^24s^2$
Vanadium	V	23	2	6	3	2		$[Ar]3d^34s^2$
Chromium	Cr	24	2	6	5	1		$[Ar]3d^54s^1$
Manganese	Mn	25	2	6	5	2		$[Ar]3d^54s^2$
Iron	Fe	26	2	6	6	2		$[Ar]3d^64s^2$
Cobalt	Co	27	2	6	7	2		$[Ar]3d^74s^2$
Nickel	Ni	28	2	6	8	2		$[Ar]3d^84s^2$
Copper	Cu	29	2	6	10	1		$[Ar]3d^{10}4s^1$
Zinc	Zn	30	2	6	10	2		$[Ar]3d^{10}4s^2$
Gallium	Ga	31	2	6	10	2	1	$[Ar]3d^{10}4s^24p^1$
Germanium	Ge	32	2	6	10	2	2	$[Ar]3d^{10}4s^24p^2$
Arsenic	As	33	2	6	10	2	3	$[Ar]3d^{10}4s^24p^3$
Selenium	Se	34	2	6	10	2	4	$[Ar]3d^{10}4s^24p^4$
Bromine	Br	35	2	6	10	2	5	$[Ar]3d^{10}4s^24p^5$
Krypton	Kr	36	2	6	10	2	6	$[Ar]3d^{10}4s^24p^6$

*$[Ar] = 1s^22s^22p^63s^23p^6$

be filled. A total of 10 electrons can occupy the 3d orbitals. These are filled successively in the 10 elements from scandium (atomic number 21) to zinc (atomic number 30).

Scandium, Sc, has the electron configuration $[Ar]3d^14s^2$. Titanium, Ti, has the configuration $[Ar]3d^24s^2$. And vanadium, V, has the configuration $[Ar]3d^34s^2$. Up to this point, three electrons with the same spin have been added to three separate d orbitals, as required by Hund's rule.

Surprisingly, chromium, Cr, has the electron configuration $[Ar]3d^54s^1$. Not only did the added electron go into the fourth 3d orbital, but an electron also moved from the 4s orbital into the fifth 3d orbital, leaving the 4s orbital with a single electron. Chromium's electron configuration is contrary to what is expected according to the Aufbau principle. However, in reality the $[Ar]3d^54s^1$ configuration is of lower energy than a $[Ar]3d^44s^2$ configuration. For chromium, having six orbitals, all with unpaired electrons, is a more stable arrangement than having four unpaired electrons in the 3d orbitals and forcing two electrons to pair up in the 4s orbital. On the other hand, for tungsten, W, which is in the same group as chromium, having four electrons in the 5d orbitals and two electrons paired in the 6s orbital is the most stable arrrangement. Unfortunately, there is no simple explanation for such deviations from the expected order given in **Figure 19.**

Manganese, Mn, has the electron configuration $[Ar]3d^54s^2$. The added electron goes to the 4s orbital, completely filling this orbital while leaving the 3d orbitals still half-filled. Beginning with the next element, electrons continue to pair in the d orbitals. Thus, iron, Fe, has the configuration $[Ar]3d^64s^2$; cobalt, Co, has the configuration $[Ar]3d^74s^2$; and nickel, Ni, has the configuration $[Ar]3d^84s^2$. Next is copper, Cu, in which an electron moves from the 4s orbital to pair with the electron in the fifth 3d orbital. The result is an electron configuration of $[Ar]3d^{10}4s^1$—the lowest-energy configuration for Cu. As with Cr, there is no simple explanation for this deviation from the expected order.

In atoms of zinc, Zn, the 4s sublevel is filled to give the electron configuration $[Ar]3d^{10}4s^2$. In atoms of the next six elements, electrons add one by one to the three 4p orbitals. According to Hund's rule, one electron is added to each of the three 4p orbitals before electrons are paired in any 4p orbital.

Elements of the Fifth Period

In the 18 elements of the fifth period, sublevels fill in a similar manner as in elements of the fourth period. However, they start at the 5s orbital instead of the 4s. Successive electrons are added first to the 5s orbital, then to the 4d orbitals, and finally to the 5p orbitals. This can be seen in **Table 6.** There are occasional deviations from the predicted configurations here also. The deviations differ from those for fourth-period elements, but in each case the preferred configuration has the lowest possible energy.

TABLE 6 Electron Configurations of Atoms of Elements in the Fifth Period

Name	Symbol	Atomic number	Number of electrons in sublevels above 3d					Noble-gas notation
			4s	4p	4d	5s	5p	
Rubidium	Rb	37	2	6		1		*$[Kr]5s^1$
Strontium	Sr	38	2	6		2		$[Kr]5s^2$
Yttrium	Y	39	2	6	1	2		$[Kr]4d^15s^2$
Zirconium	Zr	40	2	6	2	2		$[Kr]4d^25s^2$
Niobium	Nb	41	2	6	4	1		$[Kr]4d^45s^1$
Molybdenum	Mo	42	2	6	5	1		$[Kr]4d^55s^1$
Technetium	Tc	43	2	6	6	1		$[Kr]4d^65s^1$
Ruthenium	Ru	44	2	6	7	1		$[Kr]4d^75s^1$
Rhodium	Rh	45	2	6	8	1		$[Kr]4d^85s^1$
Palladium	Pd	46	2	6	10			$[Kr]4d^{10}$
Silver	Ag	47	2	6	10	1		$[Kr]4d^{10}5s^1$
Cadmium	Cd	48	2	6	10	2		$[Kr]4d^{10}5s^2$
Indium	In	49	2	6	10	2	1	$[Kr]4d^{10}5s^25p^1$
Tin	Sn	50	2	6	10	2	2	$[Kr]4d^{10}5s^25p^2$
Antimony	Sb	51	2	6	10	2	3	$[Kr]4d^{10}5s^25p^3$
Tellurium	Te	52	2	6	10	2	4	$[Kr]4d^{10}5s^25p^4$
Iodine	I	53	2	6	10	2	5	$[Kr]4d^{10}5s^25p^5$
Xenon	Xe	54	2	6	10	2	6	$[Kr]4d^{10}5s^25p^6$

*$[Kr] = 1s^22s^22p^63s^23p^63d^{10}4s^24p^6$


SAMPLE PROBLEM B   *For more help, go to the Math Tutor at the end of Chapter 5.*

a. **Write both the complete electron-configuration notation and the noble-gas notation for iron, Fe.**

b. **How many electron-containing orbitals are in an atom of iron? How many of these orbitals are completely filled? How many unpaired electrons are there in an atom of iron? In which sublevel are the unpaired electrons located?**

SOLUTION   a. The complete electron-configuration notation of iron is $1s^22s^22p^63s^23p^63d^64s^2$. The periodic table inside the back cover of the text reveals that $1s^22s^22p^63s^23p^6$ is the electron configuration of the noble gas argon, Ar. Therefore, as shown in **Table 5**, iron's noble-gas notation is $[Ar]3d^64s^2$.

b. An iron atom has 15 orbitals that contain electrons. They consist of one 1s orbital, one 2s orbital, three 2p orbitals, one 3s orbital, three 3p orbitals, five 3d orbitals, and one 4s orbital. Eleven of these orbitals are filled, and there are four unpaired electrons. They are located in the 3d sublevel. The notation $3d^6$ represents
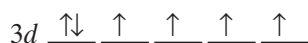
$$3d \;\; \underline{\uparrow\downarrow} \;\; \underline{\uparrow} \;\; \underline{\uparrow} \;\; \underline{\uparrow} \;\; \underline{\uparrow}$$

1. a. Write both the complete electron-configuration notation and the noble-gas notation for iodine, I. How many inner-shell electrons does an iodine atom contain?

   b. How many electron-containing orbitals are in an atom of iodine? How many of these orbitals are filled? How many unpaired electrons are there in an atom of iodine?

2. a. Write the noble-gas notation for tin, Sn. How many unpaired electrons are there in an atom of tin?

   b. How many electron-containing *d* orbitals are there in an atom of tin? Name the element in the fourth period whose atoms have the same number of electrons in their highest energy levels that tin's atoms do.

3. a. Write the complete electron configuration for the element with atomic number 25. You may use the diagram shown in **Figure 19.**

   b. Identify the element described in item 3a.

4. a. How many orbitals are completely filled in an atom of the element with atomic number 18? Write the complete electron configuration for this element.

   b. Identify the element described in item 4a.

Elements of the Sixth and Seventh Periods

The sixth period consists of 32 elements. It is much longer than the periods that precede it in the periodic table. To build up electron configurations for elements of this period, electrons are added first to the 6*s* orbital in cesium, Cs, and barium, Ba. Then, in lanthanum, La, an electron is added to the 5*d* orbital.

With the next element, cerium, Ce, the 4*f* orbitals begin to fill, giving cerium atoms a configuration of $[Xe]4f^15d^16s^2$. In the next 13 elements, the 4*f* orbitals are filled. Next the 5*d* orbitals are filled and the period is completed by filling the 6*p* orbitals. Because the 4*f* and the 5*d* orbitals are very close in energy, numerous deviations from the simple rules occur as these orbitals are filled. The electron configurations of the sixth-period elements can be found in the periodic table inside the back cover of the text.

The seventh period is incomplete and consists largely of synthetic elements, which will be discussed in Chapter 21.

a. **Write both the complete electron-configuration notation and the noble-gas notation for a rubidium atom.**

b. **Identify the elements in the second, third, and fourth periods that have the same number of highest-energy-level electrons as rubidium.**

SOLUTION   a. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$, $[Kr]5s^1$

b. Rubidium has one electron in its highest energy level (the fifth). The elements with the same outermost configuration are, in the second period, lithium, Li; in the third period, sodium, Na; and in the fourth period, potassium, K.

PRACTICE   *Answers in Appendix E*

1. a. Write both the complete electron-configuration notation and the noble-gas notation for a barium atom.

   b. Identify the elements in the second, third, fourth, and fifth periods that have the same number of highest-energy-level electrons as barium.

2. a. Write the noble-gas notation for a gold atom.

   b. Identify the elements in the sixth period that have one unpaired electron in their $6s$ sublevel.

SECTION REVIEW

1. a. What is an atom's electron configuration?
   b. What three principles guide the electron configuration of an atom?

2. What three methods are used to represent the arrangement of electrons in atoms?

3. What is an octet of electrons? Which elements contain an octet of electrons?

4. Write the complete electron-configuration notation, the noble-gas notation, and the orbital notation for the following elements:

   a. carbon   b. neon   c. sulfur

5. Identify the elements having the following electron configurations:

   a. $1s^2 2s^2 2p^6 3s^2 3p^3$

   b. $[Ar]4s^1$

   c. contains four electrons in its third and outer main energy level

   d. contains one set of paired and three unpaired electrons in its fourth and outer main energy level

Critical Thinking

6. **RELATING IDEAS** Write the electron configuration for the third-period elements Al, Si, P, S, and Cl. Is there a relationship between the group number of each element and the number of electrons in the outermost energy level?

The Development of a New Atomic Model

Vocabulary
electromagnetic radiation
electromagnetic spectrum
wavelength
frequency
photoelectric effect
quantum
photon
ground state
excited state
line-emission spectrum
continuous spectrum

- In the early twentieth century, light was determined to have a dual wave-particle nature.
- Quantum theory was developed to explain observations such as the photoelectric effect and the line-emission spectrum of hydrogen.
- Quantum theory states that electrons can exist only at specific atomic energy levels.
- When an electron moves from one main energy level to a main energy level of lower energy, a photon is emitted. The photon's energy equals the energy difference between the two levels.
- An electron in an atom can move from one main energy level to a higher main energy level only by absorbing an amount of energy exactly equal to the difference between the two levels.

The Quantum Model of the Atom

Vocabulary
Heisenberg uncertainty principle
quantum theory
orbital
quantum number
principal quantum number
angular momentum quantum number
magnetic quantum number
spin quantum number

- In the early twentieth century, electrons were determined to have a dual wave-particle nature.
- The Heisenberg uncertainty principle states that it is impossible to determine simultaneously the position and velocity of an electron or any other particle.
- Quantization of electron energies is a natural outcome of the Schrödinger wave equation, which describes the properties of an atom's electrons.
- An orbital, a three-dimensional region around the nucleus, shows the region in space where an electron is most likely to be found.
- The four quantum numbers that describe the properties of electrons in atomic orbitals are the principal quantum number, the angular momentum quantum number, the magnetic quantum number, and the spin quantum number.

Electron Configurations

Vocabulary
electron configuration
Aufbau principle
Pauli exclusion principle
Hund's rule
noble gas
noble-gas configuration

- The ground-state electron configuration of an atom can be written by using the Aufbau principle, Hund's rule, and the Pauli exclusion principle.
- Electron configurations can be depicted by using different types of notation. In this book, three types of notation are used: orbital notation, electron-configuration notation, and noble-gas notation.
- Electron configurations of some atoms, such as chromium, deviate from the predictions of the Aufbau principle, but the ground-state configuration that results is the configuration with the minimum possible energy.

The Development of a New Atomic Model

SECTION 1 REVIEW

1. a. List five examples of electromagnetic radiation.
   b. What is the speed of all forms of electromagnetic radiation in a vacuum?

2. Prepare a two-column table. List the properties of light that can best be explained by the wave theory in one column. List those best explained by the particle theory in the second column. You may want to consult a physics textbook for reference.

3. What are the frequency and wavelength ranges of visible light?

4. List the colors of light in the visible spectrum in order of increasing frequency.

5. In the early twentieth century, what two experiments involving light and matter could not be explained by the wave theory of light?

6. a. How are the wavelength and frequency of electromagnetic radiation related?
   b. How are the energy and frequency of electromagnetic radiation related?
   c. How are the energy and wavelength of electromagnetic radiation related?

7. Which theory of light—the wave or particle theory—best explains the following phenomena?
   a. the interference of light
   b. the photoelectric effect
   c. the emission of electromagnetic radiation by an excited atom

8. Distinguish between the ground state and an excited state of an atom.

9. According to Bohr's model of the hydrogen atom, how is hydrogen's emission spectrum produced?

PRACTICE PROBLEMS

10. Determine the frequency of light whose wavelength is 4.257×10^{-7} cm.

11. Determine the energy in joules of a photon whose frequency is 3.55×10^{17} Hz.

12. Using the two equations $E = h\nu$ and $c = \lambda\nu$, derive an equation expressing E in terms of h, c, and λ.

13. How long would it take a radio wave whose frequency is 7.25×10^5 Hz to travel from Mars to Earth if the distance between the two planets is approximately 8.00×10^7 km?

14. Cobalt-60 is an artificial radioisotope that is produced in a nuclear reactor and is used as a gamma-ray source in the treatment of certain types of cancer. If the wavelength of the gamma radiation from a cobalt-60 source is 1.00×10^{-3} nm, calculate the energy of a photon of this radiation.

The Quantum Model of the Atom

SECTION 2 REVIEW

15. Describe two major shortcomings of Bohr's model of the atom.

16. a. What is the principal quantum number?
    b. How is it symbolized?
    c. What are shells?
    d. How does n relate to the number of electrons allowed per main energy level?

17. a. What information is given by the angular momentum quantum number?
    b. What are sublevels, or subshells?

18. For each of the following values of n, indicate the numbers and types of sublevels possible for that main energy level. (Hint: See **Table 2.**)
    a. $n = 1$
    b. $n = 2$
    c. $n = 3$
    d. $n = 4$
    e. $n = 7$ (number only)

19. a. What information is given by the magnetic quantum number?
    b. How many orbital orientations are possible in each of the s, p, d, and f sublevels?
    c. Explain and illustrate the notation for distinguishing between the different p orbitals in a sublevel.

20. a. What is the relationship between n and the total number of orbitals in a main energy level?
   b. How many total orbitals are contained in the third main energy level? in the fifth?

21. a. What information is given by the spin quantum number?
   b. What are the possible values for this quantum number?

22. How many electrons could be contained in the following main energy levels with n equal to the number provided?
   a. 1
   b. 3
   c. 4
   d. 6
   e. 7

PRACTICE PROBLEMS

23. Sketch the shape of an s orbital and a p orbital.
24. How does a $2s$ orbital differ from a $1s$ orbital?
25. How do a $2p_x$ and a $2p_y$ orbital differ?

Electron Configurations

SECTION 3 REVIEW

26. a. In your own words, state the Aufbau principle.
   b. Explain the meaning of this principle in terms of an atom with many electrons.

27. a. In your own words, state Hund's rule.
   b. What is the basis for this rule?

28. a. In your own words, state the Pauli exclusion principle.
   b. What is the significance of the spin quantum number?

29. a. What is meant by the highest occupied energy level in an atom?
   b. What are inner-shell electrons?

30. Determine the highest occupied energy level in the following elements:
   a. He
   b. Be
   c. Al
   d. Ca
   e. Sn

31. Write the orbital notation for the following elements. (Hint: See Sample Problem A.)
   a. P
   b. B
   c. Na
   d. O

32. Write the electron-configuration notation for the element whose atoms contain the following number of electrons:
   a. 3
   b. 6
   c. 8
   d. 13

33. Given that the electron configuration for oxygen is $1s^2 2s^2 2p^4$, answer the following questions:
   a. How many electrons are in each oxygen atom?
   b. What is the atomic number of this element?
   c. Write the orbital notation for oxygen's electron configuration.
   d. How many unpaired electrons does oxygen have?
   e. What is the highest occupied energy level?
   f. How many inner-shell electrons does the atom contain?
   g. In which orbital(s) are these inner-shell electrons located?

34. a. What are the noble gases?
   b. What is a noble-gas configuration?
   c. How does noble-gas notation simplify writing an atom's electron configuration?

35. Write the noble-gas notation for the electron configuration of each of the elements below. (Hint: See Sample Problem B.)
   a. Cl
   b. Ca
   c. Se

36. a. What information is given by the noble-gas notation $[\text{Ne}]3s^2$?
   b. What element does this represent?

37. Write both the complete electron-configuration notation and the noble-gas notation for each of the elements below. (Hint: See Sample Problem C.)
   a. Na
   b. Sr
   c. P

38. Identify each of the following atoms on the basis of its electron configuration:
   a. $1s^2 2s^2 2p^1$
   b. $1s^2 2s^2 2p^5$
   c. $[Ne]3s^2$
   d. $[Ne]3s^2 3p^2$
   e. $[Ne]3s^2 3p^5$
   f. $[Ar]4s^1$
   g. $[Ar]3d^6 4s^2$

PRACTICE PROBLEMS

39. List the order in which orbitals generally fill, from the $1s$ to the $7p$ orbital.

40. Write the noble-gas notation for the electron configurations of each of the following elements:
   a. As   e. Sn
   b. Pb   f. Xe
   c. Lr   g. La
   d. Hg

41. How do the electron configurations of chromium and copper contradict the Aufbau principle?

MIXED REVIEW

42. a. Which has a longer wavelength: green light or yellow light?
   b. Which has a higher frequency: an X ray or a microwave?
   c. Which travels at a greater speed: ultraviolet light or infrared light?

43. Write both the complete electron-configuration and noble-gas notation for each of the following:
   a. Ar   b. Br   c. Al

44. Given the speed of light as 3.00×10^8 m/s, calculate the wavelength of the electromagnetic radiation whose frequency is 7.500×10^{12} Hz.

45. a. What is the electromagnetic spectrum?
   b. What units can be used to express wavelength?
   c. What unit is used to express frequencies of electromagnetic waves?

46. Given that the electron configuration for phosphorus is $1s^2 2s^2 2p^6 3s^2 3p^3$, answer the following questions:
   a. How many electrons are in each atom?
   b. What is the atomic number of this element?
   c. Write the orbital notation for this element.
   d. How many unpaired electrons does an atom of phosphorus have?
   e. What is its highest occupied energy level?
   f. How many inner-shell electrons does the atom contain?
   g. In which orbital(s) are these inner-shell electrons located?

47. What is the frequency of a radio wave whose energy is 1.55×10^{-24} J per photon?

48. Write the noble-gas notation for the electron configurations of each of the following elements:
   a. Hf   d. At
   b. Sc   e. Ac
   c. Fe   f. Zn

49. Describe the major similarities and differences between Schrödinger's model of the atom and the model proposed by Bohr.

50. When sodium is heated, a yellow spectral line whose energy is 3.37×10^{-19} J per photon is produced.
   a. What is the frequency of this light?
   b. What is the wavelength of this light?

51. a. What is an orbital?
   b. Describe an orbital in terms of an electron cloud.

CRITICAL THINKING

52. Inferring Relationships In the emission spectrum of hydrogen shown in **Figure 5,** each colored line is produced by the emission of photons with specific energies. Substances also produce absorption spectra when electromagnetic radiation passes through them. Certain wavelengths are absorbed. Using the diagram below, predict what the wavelengths of the absorption lines will be when white light (all of the colors of the visible spectrum) is passed through hydrogen gas.

300 nm   700 nm

Hydrogen absorption spectrum
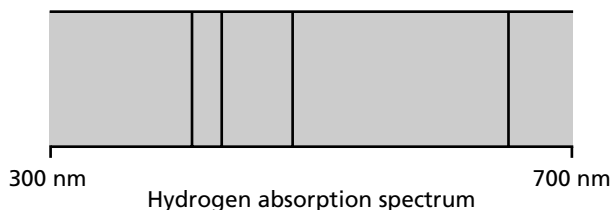
53. Applying Models In discussions of the photo-electric effect, the minimum energy needed to remove an electron from the metal is called the *threshold energy* and is a characteristic of the metal. For example, chromium, Cr, will emit electrons when the wavelength of the radiation is 284 nm or less. Calculate the threshold energy for chromium. (Hint: You will need to use the two equations that describe the relationships between wavelength, frequency, speed of light, and Planck's constant.)

54. Analyzing Information Four electrons in an atom have the four sets of quantum numbers given below. Which electrons are in the same orbital? Explain your answer.

a. 1, 0, 0, −__
b. 1, 0, 0, +__
c. 2, 1, 1, +__
d. 2, 1, 0, +__

55. Relating Ideas Which of the sets of quantum numbers below are possible? Which are impossible? Explain your choices.

a. 2, 2, 1, +__
b. 2, 0, 0, −__
c. 2, 0, 1, −__

USING THE HANDBOOK

56. Sections 1 and 2 of the *Elements Handbook* contain information on an analytical test and a technological application for Group 1 and 2 elements. The test and application are based on the emission of light from atoms. Review these sections to answer the following:

a. What analytical technique utilizes the emission of light from excited atoms?
b. What elements in Groups 1 and 2 can be identified by this technique?
c. What types of compounds are used to provide color in fireworks?
d. What wavelengths within the visible spectrum would most likely contain emission lines for barium?

RESEARCH & WRITING

57. Neon signs do not always contain neon gas. The various colored lights produced by the signs are due to the emission of a variety of low-pressure gases in different tubes. Research other kinds of gases used in neon signs, and list the colors that they emit.

58. Prepare a report about the photoelectric effect, and cite some of its practical uses. Explain the basic operation of each device or technique mentioned.

ALTERNATIVE ASSESSMENT

59. Performance A spectroscope is a device used to produce and analyze spectra. Construct a simple spectroscope, and determine the absorption spectra of several elemental gases. (Your teacher will provide you with the gas discharge tubes containing samples of different gases.)

extension

Graphing Calculator Calculating Quantum Number Relationships
Go to **go.hrw.com** for a graphing calculator exercise that asks you to calculate quantum number relationships.

Keyword: HC6ARRX

Math Tutor WEIGHTED AVERAGES AND ATOMIC MASS

You have learned that the mass of a proton is about 1 amu and that a neutron is only slightly heavier. Because atomic nuclei consist of whole numbers of protons and neutrons, you might expect that the atomic mass of an element would be very near a whole number. However, if you look at the periodic table, you will see that the atomic masses of many elements lie somewhere between whole numbers. In fact, the atomic masses listed on the table are *average* atomic masses. The atomic masses are averages because most elements occur in nature as a specific mixture of isotopes. For example, 75.76% of chlorine atoms have a mass of 34.969 amu, and 24.24% have a mass of 36.966 amu. If the isotopes were in a 1:1 ratio, you could simply add the masses of the two isotopes together and divide by 2. However, to account for the differing abundance of the isotopes, you must calculate a *weighted average.* For chlorine, the weighted average is 35.45 amu. The following two examples demonstrate how weighted averages are calculated.

SAMPLE 1

Naturally occurring silver consists of 51.839% Ag-107 (atomic mass 106.905 093) and 48.161% Ag-109 (atomic mass 108.904 756). What is the average atomic mass of silver?

To find average atomic mass, convert each percentage to a decimal equivalent and multiply by the atomic mass of the isotope.

$$0.518\ 39 \times 106.905\ 093\ \text{amu} = 55.419\ \text{amu}$$
$$\underline{0.481\ 61 \times 108.904\ 756\ \text{amu} = 52.450\ \text{amu}}$$
$$107.869\ \text{amu}$$

Adding the masses contributed by each isotope gives an average atomic mass of 107.869 amu. Note that this value for the average atomic mass of silver is very near the one given in the periodic table.

SAMPLE 2

Naturally occurring magnesium consists of 78.99% Mg-24 (atomic mass 23.985 042), 10.00% Mg-25 (atomic mass 24.985 837), and 11.01% Mg-26 (atomic mass 25.982 593). What is the average atomic mass of magnesium?

Again, convert each percentage to a decimal and multiply by the atomic mass of the isotope to get the mass contributed by each isotope.

$$0.7899 \times 23.985\ 042\ \text{amu} = 18.95\ \text{amu}$$
$$0.1000 \times 24.985\ 837\ \text{amu} = \ \ 2.499\ \text{amu}$$
$$\underline{0.1101 \times 25.982\ 593\ \text{amu} = \ \ 2.861\ \text{amu}}$$
$$24.31\ \text{amu}$$

Adding the masses contributed by each isotope gives an average atomic mass of 24.31 amu.

PRACTICE PROBLEMS

1. Rubidium occurs naturally as a mixture of two isotopes, 72.17% Rb-85 (atomic mass 84.911 792 amu) and 27.83% Rb-87 (atomic mass 86.909 186 amu). What is the average atomic mass of rubidium?

2. The element silicon occurs as a mixture of three isotopes: 92.22% Si-28, 4.69% Si-29, and 3.09% Si-30. The atomic masses of these three isotopes are as follows: Si-28 = 27.976 926 amu, Si-29 = 28.976 495 amu, and Si-30 = 29.973 770 amu.

Find the average atomic mass of silicon.

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Which of the following relationships is true?
   A. Higher-energy light has a higher frequency than lower-energy light does.
   B. Higher-energy light has a longer wavelength than lower-energy light does.
   C. Higher-energy light travels at a faster speed than lower-energy light does.
   D. Higher-frequency light travels at a slower speed than lower-energy light does.

2. The energy of a photon is greatest for
   A. visible light.
   B. ultraviolet light.
   C. infrared light.
   D. X-ray radiation.

3. What is the wavelength of radio waves that have a frequency of 88.5 MHz?
   A. 3.4 m   C. 0.30 m
   B. 8.9 nm   D. 300 nm

4. Which transition in an excited hydrogen atom will emit the longest wavelength of light?
   A. E_5 to E_1 C. E_3 to E_1
   B. E_4 to E_1 D. E_2 to E_1

5. Which of the following quantum numbers is often designated by the letters *s, p, d,* and *f* instead of by numbers?
   A. *n*   C. *m*
   B. *l*   D. *s*

6. Which quantum number is related to the shape of an orbital?
   A. *n*   C. *m*
   B. *l*   D. *s*

7. What is the maximum number of unpaired electrons that can be placed in a 3*p* sublevel?
   A. 1   C. 3
   B. 2   D. 4

8. What is the maximum number of electrons that can occupy a 3*s* orbital?
   A. 1   C. 6
   B. 2   D. 10

9. Which element has the noble-gas notation $[Kr]5s^2 4d^2$?
   A. Se   C. Zr
   B. Sr   D. Mo

SHORT ANSWER

10. When a calcium salt is heated in a flame, a photon of light with an energy of 3.2×10^{-19} J is emitted. On the basis of this fact and the table below, what color would be expected for the calcium flame?

Frequency, s^{-1}	7.1×10^{14}	6.4×10^{14}	5.7×10^{14}
Wavelength, nm	422	469	526
Color	violet	blue	green
Frequency, s^{-1}	5.2×10^{14}	4.8×10^{14}	4.3×10^{14}
Wavelength, nm	577	625	698
Color	yellow	orange	red

11. The electron configuration of sulfur is $1s^2 2s^2 2p^6 3s^2 3p^4$. Write the orbital notation for sulfur.

EXTENDED RESPONSE

12. Explain the reason for the hydrogen line-emission spectrum.

13. When blue light shines on potassium metal in a photocell, electrons are emitted. But when yellow light shines on the metal, no current is observed. Explain.

Test TIP If time permits, take short mental breaks during the test to improve your concentration.

MICRO-LAB

Flame Tests

OBJECTIVES

- *Identify* a set of flame-test color standards for selected metal ions.
- *Relate* the colors of a flame test to the behavior of excited electrons in a metal ion.
- *Identify* an unknown metal ion by using a flame test.
- *Demonstrate* proficiency in performing a flame test and in using a spectroscope.

MATERIALS

- 250 mL beaker
- Bunsen burner and related equipment
- cobalt glass plates
- crucible tongs
- distilled water
- flame-test wire
- glass test plate (or a microchemistry plate with wells)
- spectroscope
- 1.0 M HCl solution
- $CaCl_2$ solution
- K_2SO_4 solution
- Li_2SO_4 solution
- Na_2SO_4 solution
- $SrCl_2$ solution
- unknown solution

OPTIONAL EQUIPMENT

- wooden splints

BACKGROUND

The characteristic light emitted by an element is the basis for the chemical test known as a *flame test.*

To identify an unknown substance, you must first determine the characteristic colors produced by different elements. You will do this by performing a flame test on a variety of standard solutions of metal compounds. Then, you will perform a flame test with an unknown sample to see if it matches any of the standard solutions. The presence of even a speck of another substance can interfere with the identification of the true color of a particular type of atom, so be sure to keep your equipment very clean and perform multiple trials to check your work.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Prepare a data table in your lab notebook. Include rows for each of the solutions of metal compounds listed in the materials list and an unknown solution. The table should have three wide columns for the three trials you will perform with each substance. Each column should have room to record the colors and wavelengths of light. Be sure you have plenty of room to write your observations about each test.

2. Label a beaker *Waste.* Thoroughly clean and dry a well strip. Fill the first well one-fourth full with 1.0 M HCl on the plate. Clean the test wire by first dipping it in the HCl and then holding it in the colorless flame of the Bunsen burner. Repeat this procedure until the flame is not colored by the wire. When the wire is ready, rinse

the well with distilled water and collect the rinse water in the waste beaker.

3. Put 10 drops of each metal ion solution listed in the materials list in a row in each well of the well strip. Put a row of 1.0 M HCl drops on a glass plate across from the metal ion solutions. Record the positions of all of the chemicals placed in the wells. The wire will need to be cleaned thoroughly between each test solution with HCl to avoid contamination from the previous test.

PROCEDURE

1. Dip the wire into the $CaCl_2$ solution, and then hold it in the Bunsen burner flame. Observe the color of the flame, and record it in the data table. Repeat the procedure again, but this time look through the spectroscope to view the results. Record the wavelengths you see from the flame. Repeat each test three times. Clean the wire with the HCl as you did in Preparation step 2.

2. Repeat step 1 with the K_2SO_4 and with each of the remaining solutions in the well strip.

3. Test another drop of Na_2SO_4, but this time view the flame through two pieces of cobalt glass. Clean the wire, and repeat the test. Record in your data table the colors and wavelengths of the flames as they appear when viewed through the cobalt glass. Clean the wire and the well strip, and rinse the well strip with distilled water. Pour the rinse water into the waste beaker.

4. Put a drop of K_2SO_4 in a clean well. Add a drop of Na_2SO_4. Perform a flame test for the mixture. Observe the flame without the cobalt glass. Repeat the test again, but this time observe the flame through the cobalt glass. Record in your data table the colors and wavelengths of the flames. Clean the wire, and rinse the well strip with distilled water. Pour the rinse water into the waste beaker.

5. Obtain a sample of the unknown solution. Perform flame tests for it with and without the cobalt glass. Record your observations. Clean the wire, and rinse the well strip with distilled water. Pour the rinse water into the waste beaker.

CLEANUP AND DISPOSAL

6. Dispose of the contents of the waste beaker in the container designated by your teacher. Wash your hands thoroughly after cleaning up the area and equipment.

ANALYSIS AND INTERPRETATION

1. **Organizing Data:** Examine your data table, and create a summary of the flame test for each metal ion.

2. **Analyzing Data:** Account for any differences in the individual trials for the flame tests for the metals ions.

3. **Organizing Ideas:** Explain how viewing the flame through cobalt glass can make it easier to analyze the ions being tested.

4. **Relating Ideas:** For three of the metal ions tested, explain how the flame color you saw relates to the lines of color you saw when you looked through the spectroscope.

CONCLUSIONS

1. **Inferring Conclusions:** What metal ions are in the unknown solution?

2. **Evaluating Methods:** How would you characterize the flame test with respect to its sensitivity? What difficulties could there be when identifying ions by the flame test?

EXTENSIONS

1. **Inferring Conclusions:** A student performed flame tests on several unknowns and observed that they all were shades of red. What should the student do to correctly identify these substances? Explain your answer.

2. **Applying Ideas:** During a flood, the labels from three bottles of chemicals were lost. The three unlabeled bottles of white solids were known to contain the following: strontium nitrate, ammonium carbonate, and potassium sulfate. Explain how you could easily test the substances and relabel the three bottles. (Hint: Ammonium ions do not provide a distinctive flame color.)

The Periodic Law

Atoms of the heavier elements form when stars explode, such as when a supernova caused this nebula.

Rosette Nebula

History of the Periodic Table

OBJECTIVES

- Explain the roles of Mendeleev and Moseley in the development of the periodic table.

- Describe the modern periodic table.

- Explain how the periodic law can be used to predict the physical and chemical properties of elements.

- Describe how the elements belonging to a group of the periodic table are interrelated in terms of atomic number.

Imagine the confusion among chemists during the middle of the nineteenth century. By 1860, more than 60 elements had been discovered. Chemists had to learn the properties of these elements as well as those of the many compounds that they formed—a difficult task. And to make matters worse, there was no method for accurately determining an element's atomic mass or the number of atoms of an element in a particular chemical compound. Different chemists used different atomic masses for the same elements, resulting in different compositions being proposed for the same compounds. This made it nearly impossible for one chemist to understand the results of another.

In September 1860, a group of chemists assembled at the First International Congress of Chemists in Karlsruhe, Germany, to settle the issue of atomic mass as well as some other matters that were making communication difficult. At the Congress, Italian chemist Stanislao Cannizzaro presented a convincing method for accurately measuring the relative masses of atoms. Cannizzaro's method enabled chemists to agree on standard values for atomic mass and initiated a search for relationships between atomic mass and other properties of the elements.

Mendeleev and Chemical Periodicity

When the Russian chemist Dmitri Mendeleev heard about the new atomic masses discussed at Karlsruhe, he decided to include the new values in a chemistry textbook that he was writing. In the book, Mendeleev hoped to organize the elements according to their properties. He went about this much as you might organize information for a research paper. He placed the name of each known element on a card, together with the atomic mass of the element and a list of its observed physical and chemical properties. He then arranged the cards according to various properties and looked for trends or patterns.

Mendeleev noticed that when the elements were arranged in order of increasing atomic mass, certain similarities in their chemical properties appeared at regular intervals. Such a repeating pattern is referred to as *periodic*. The second hand of a watch, for example, passes over any given mark at periodic, 60-second intervals. The circular waves created by a drop of water hitting a water surface are also periodic.
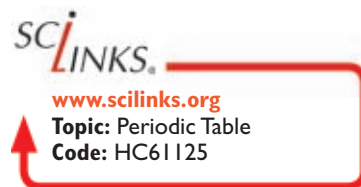
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Topic: Periodic Table
Code: HC61125

FIGURE 1 The regularly spaced water waves represent a simple periodic pattern.

FIGURE 2 In his first published periodic table, Mendeleev arranged the elements in vertical periods according to relative atomic mass. The atomic mass for each element is indicated by the number following the element's symbol. The unknown elements indicated by question marks at estimated atomic masses 45, 68, and 70 were later identified as scandium, Sc, gallium, Ga, and germanium, Ge.

Mendeleev created a table in which elements with similar properties were grouped together—a periodic table of the elements. His first periodic table, shown in **Figure 2,** was published in 1869. Note that Mendeleev placed iodine, I (atomic mass 127), after tellurium, Te (atomic mass 128). Although this contradicted the pattern of listing the elements in order of increasing atomic mass, it allowed Mendeleev to place tellurium in a group of elements with which it shares similar properties. Reading horizontally across Mendeleev's table, this group includes oxygen, O, sulfur, S, and selenium, Se. Iodine could also, then, be placed in the group it resembles chemically, which includes fluorine, F, chlorine, Cl, and bromine, Br.

Mendeleev's procedure left several empty spaces in his periodic table (see **Figure 2**). In 1871, the Russian chemist boldly predicted the existence and properties of the elements that would fill three of the spaces. By 1886, all three elements had been discovered. Today these elements are known as scandium, Sc, gallium, Ga, and germanium, Ge. Their properties are strikingly similar to those predicted by Mendeleev.

The success of Mendeleev's predictions persuaded most chemists to accept his periodic table and earned him credit as the discoverer of the periodic law. Two questions remained, however. (1) Why could most of the elements be arranged in the order of increasing atomic mass but a few could not? (2) What was the reason for chemical periodicity?

Moseley and the Periodic Law

The first question was not answered until more than 40 years after Mendeleev's first periodic table was published. In 1911, the English scientist Henry Moseley, who was working with Ernest Rutherford, examined the spectra of 38 different metals. When analyzing his data, Moseley discovered a previously unrecognized pattern. The elements in the periodic table fit into patterns better when they were arranged in increasing order according to nuclear charge, or the number of protons in the nucleus. Moseley's work led to both the modern definition of atomic number and the recognition that atomic number, not atomic mass, is the basis for the organization of the periodic table.

Moseley's discovery was consistent with Mendeleev's ordering of the periodic table by properties rather than strictly by atomic mass. For example, according to Moseley, tellurium, with an atomic number of 52, belongs before iodine, which has an atomic number of 53. Today, Mendeleev's principle of chemical periodicity is correctly stated in what is known as the **periodic law:** *The physical and chemical properties of the elements are periodic functions of their atomic numbers.* In other words, when the elements are arranged in order of increasing atomic number, elements with similar properties appear at regular intervals.

The Modern Periodic Table

The periodic table has undergone extensive change since Mendeleev's time (see **Figure 6** on pages 140–141). Chemists have discovered new elements and, in more recent years, synthesized new ones in the laboratory. Each of the more than 40 new elements, however, can be placed in a group of other elements with similar properties. *The* **periodic table** *is an arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same column, or group.*

The Noble Gases

Perhaps the most significant addition to the periodic table came with the discovery of the noble gases. In 1894, English physicist John William Strutt (Lord Rayleigh) and Scottish chemist Sir William Ramsay discovered argon, Ar, a gas in the atmosphere that had previously escaped notice because of its total lack of chemical reactivity. Back in 1868,
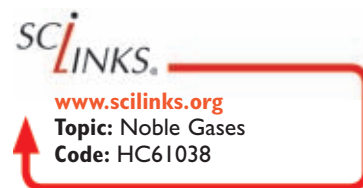
FIGURE 3 The noble gases, also known as the Group 18 elements, are all rather unreactive. As you will read, the reason for this low reactivity also accounts for the special place occupied by the noble gases in the periodic table.

another noble gas, helium, He, had been discovered as a component of the sun, based on the emission spectrum of sunlight. In 1895, Ramsay showed that helium also exists on Earth.

In order to fit argon and helium into the periodic table, Ramsay proposed a new group. He placed this group between the groups now known as Group 17 (the fluorine family) and Group 1 (the lithium family). In 1898, Ramsay discovered two more noble gases to place in his new group, krypton, Kr, and xenon, Xe. The final noble gas, radon, Rn, was discovered in 1900 by the German scientist Friedrich Ernst Dorn.

The Lanthanides

The next step in the development of the periodic table was completed in the early 1900s. It was then that the puzzling chemistry of the lanthanides was finally understood. *The* **lanthanides** *are the 14 elements with atomic numbers from 58 (cerium, Ce) to 71 (lutetium, Lu).* Because these elements are so similar in chemical and physical properties, the process of separating and identifying them was a tedious task that required the effort of many chemists.

The Actinides

Another major step in the development of the periodic table was the discovery of the actinides. *The* **actinides** *are the 14 elements with atomic numbers from 90 (thorium, Th) to 103 (lawrencium, Lr).* The lanthanides and actinides belong in Periods 6 and 7, respectively, of the periodic table, between the elements of Groups 3 and 4. To save space, the lanthanides and actinides are usually set off below the main portion of the periodic table, as shown in **Figure 6** on pages 140–141.

Periodicity

Periodicity with respect to atomic number can be observed in any group of elements in the periodic table. Consider the noble gases of Group 18. The first noble gas is helium, He. It has an atomic number of 2. The elements following helium in atomic number have completely different properties until the next noble gas, neon, Ne, which has an atomic number of 10, is reached. The remaining noble gases in order of increasing atomic number are argon (Ar, atomic number 18), krypton (Kr, atomic number 36), xenon (Xe, atomic number 54), and radon (Rn, atomic number 86). The differences in atomic number between successive noble gases are shown in **Figure 4.** Also shown in **Figure 4** are atomic-number differences between the elements of Group 1, which are all solid, silvery metals. As you can see, the differences in atomic number between the Group 1 metals follow the same pattern as the differences in atomic number between the noble gases.

Starting with the first member of Groups 13–17, a similar periodic pattern is repeated. The atomic number of each successive element is 8, 18, 18, and 32 higher than the atomic number of the element above it. In Section 2, you will see that the second mystery presented by Mendeleev's periodic table—the reason for periodicity—is explained by the arrangement of the electrons around the nucleus.

FIGURE 4 In each of Groups 1 and 18, the differences between the atomic numbers of successive elements are 8, 8, 18, 18, and 32, respectively. Groups 2 and 13–17 follow a similar pattern.

Quick LAB

Designing Your Own Periodic Table

Question

Can you design your own periodic table using information similar to that available to Mendeleev?

Procedure

1. Write down the information available for each element on separate index cards. The following information is appropriate: a letter of the alphabet (A, B, C, etc.) to identify each element; atomic mass; state; density; melting point; boiling point; and any other readily observable physical properties. Do not write the name of the element on the index card, but keep a separate list indicating the letters you have assigned to each element.

2. Organize the cards for the elements in a logical pattern as you think Mendeleev might have done.

Discussion

1. Keeping in mind that the information you have is similar to that available to Mendeleev in 1869, answer the following questions.
   a. Why are atomic masses given instead of atomic numbers?
   b. Can you identify each element by name?

2. How many groups of elements, or families, are in your periodic table? How many periods, or series, are in the table?

3. Predict the characteristics of any missing elements. When you have finished, check your work using your separate list of elements and a periodic table.

SECTION REVIEW

1. **a.** Who is credited with developing a method that led to the determination of standard relative atomic masses?

   b. Who discovered the periodic law?

   c. Who established atomic numbers as the basis for organizing the periodic table?

2. State the periodic law.

3. Name three sets of elements that have been added to the periodic table after Mendeleev's time.

4. How do the atomic numbers of the elements within each of Groups 1, 2, and 13–18 of the periodic table vary? (Refer to **Figure 4** as a guide.)

Critical Thinking

5. **RELATING IDEAS** Why are elements' atomic masses not in strict increasing order in the periodic table, even though the properties of the elements are similar? For example, by atomic mass, tellurium, Te, should be in group 17 and iodine, I, should be in Group 16, but grouping by properties has Te in Group 16 and I in Group 17.

- Describe the relationship between electrons in sublevels and the length of each period of the periodic table.

- Locate and name the four blocks of the periodic table. Explain the reasons for these names.

- Discuss the relationship between group configurations and group numbers.

- Describe the locations in the periodic table and the general properties of the alkali metals, the alkaline-earth metals, the halogens, and the noble gases.

Electron Configuration and the Periodic Table

The Group 18 elements of the periodic table (the noble gases) undergo few chemical reactions. This stability results from the gases' special electron configurations. Helium's highest occupied level, the $1s$ orbital, is completely filled with electrons. And the highest occupied levels of the other noble gases contain stable octets. Generally the electron configuration of an atom's highest occupied energy level governs the atom's chemical properties.

Periods and Blocks of the Periodic Table

While the elements are arranged vertically in the periodic table in groups that share similar chemical properties, they are also organized horizontally in rows, or *periods*. (As shown in **Figure 6,** there are a total of seven periods of elements in the modern periodic table.) As can be seen in **Table 1,** the length of each period is determined by the number of electrons that can occupy the sublevels being filled in that period.

TABLE 1 *Relationship Between Period Length and Sublevels Being Filled in the Periodic Table*
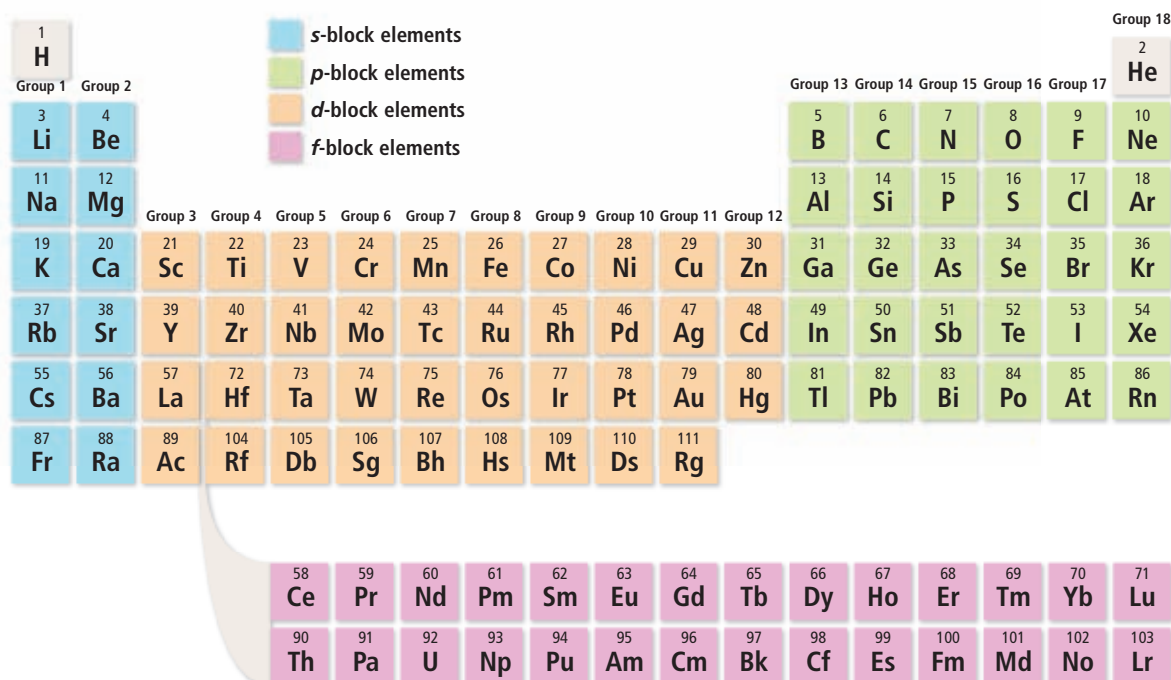
Period number	Number of elements in period	Sublevels in order of filling
1	2	$1s$
2	8	$2s\ 2p$
3	8	$3s\ 3p$
4	18	$4s\ 3d\ 4p$
5	18	$5s\ 4d\ 5p$
6	32	$6s\ 4f\ 5d\ 6p$
7	32	$7s\ 5f\ 6d\ 7p$

In the first period, the 1s sublevel is being filled. The 1s sublevel can hold a total of two electrons. Therefore, the first period consists of two elements—hydrogen and helium. In the second period, the 2s sublevel, which can hold two electrons, and the 2p sublevel, which can hold six electrons, are being filled. Consequently, the second period totals eight elements. Similarly, filling of the 3s and 3p sublevels accounts for the eight elements of the third period. Filling 3d and 4d sublevels in addition to the s and p sublevels adds 10 elements to both the fourth and fifth periods. Therefore, each of these periods totals 18 elements. Filling 4f sublevels in addition to s, p, and d sublevels adds 14 elements to the sixth period, which totals 32 elements. And as new elements are created, the 25 named elements in Period 7 could, in theory, be extended to 32.

The period of an element can be determined from the element's electron configuration. For example, arsenic, As, has the electron configuration $[Ar]3d^{10}4s^24p^3$. The 4 in $4p^3$ indicates that arsenic's highest occupied energy level is the fourth energy level. Arsenic is therefore in the fourth period in the periodic table. The period and electron configuration for each element can be found in the periodic table on pages 140–141.

Based on the electron configurations of the elements, the periodic table can be divided into four blocks, the s, p, d, and f blocks. This division is illustrated in **Figure 5.** The name of each block is determined by whether an s, p, d, or f sublevel is being filled in successive elements of that block.

FIGURE 5   Based on the electron configurations of the elements, the periodic table can be subdivided into four sublevel blocks.

Periodic Table of the Elements

Period

Period 1

1
H
Hydrogen
1.007 94
$1s^1$

Key:

6
C
Carbon
12.0107
$[He]2s^22p^2$

Atomic number — 6
Symbol — C
Name — Carbon
Average atomic mass — 12.0107
Electron configuration — $[He]2s^22p^2$

Group 1

3	11	19	37	55	87
Li	**Na**	**K**	**Rb**	**Cs**	**Fr**
Lithium	Sodium	Potassium	Rubidium	Cesium	Francium
6.941	22.989 769 28	39.0983	85.4678	132.905 4519	(223)
$[He]2s^1$	$[Ne]3s^1$	$[Ar]4s^1$	$[Kr]5s^1$	$[Xe]6s^1$	$[Rn]7s^1$

Group 2

4	12	20	38	56	88
Be	**Mg**	**Ca**	**Sr**	**Ba**	**Ra**
Beryllium	Magnesium	Calcium	Strontium	Barium	Radium
9.012 182	24.3050	40.078	87.62	137.327	(226)
$[He]2s^2$	$[Ne]3s^2$	$[Ar]4s^2$	$[Kr]5s^2$	$[Xe]6s^2$	$[Rn]7s^2$

Group 3

21	39	57	89
Sc	**Y**	**La**	**Ac**
Scandium	Yttrium	Lanthanum	Actinium
44.955 912	88.905 85	138.905 47	(227)
$[Ar]3d^14s^2$	$[Kr]4d^15s^2$	$[Xe]5d^16s^2$	$[Rn]6d^17s^2$

Group 4

22	40	72	104
Ti	**Zr**	**Hf**	**Rf**
Titanium	Zirconium	Hafnium	Rutherfordium
47.867	91.224	178.49	(261)
$[Ar]3d^24s^2$	$[Kr]4d^25s^2$	$[Xe]4f^{14}5d^26s^2$	$[Rn]5f^{14}6d^27s^2$

Group 5

23	41	73	105
V	**Nb**	**Ta**	**Db**
Vanadium	Niobium	Tantalum	Dubnium
50.9415	92.906 38	180.947 88	(262)
$[Ar]3d^34s^2$	$[Kr]4d^45s^1$	$[Xe]4f^{14}5d^36s^2$	$[Rn]5f^{14}6d^37s^2$

Group 6

24	42	74	106
Cr	**Mo**	**W**	**Sg**
Chromium	Molybdenum	Tungsten	Seaborgium
51.9961	95.94	183.84	(266)
$[Ar]3d^54s^1$	$[Kr]4d^55s^1$	$[Xe]4f^{14}5d^46s^2$	$[Rn]5f^{14}6d^47s^2$

Group 7

25	43	75	107
Mn	**Tc**	**Re**	**Bh**
Manganese	Technetium	Rhenium	Bohrium
54.938 045	(98)	186.207	(264)
$[Ar]3d^54s^2$	$[Kr]4d^65s^1$	$[Xe]4f^{14}5d^56s^2$	$[Rn]5f^{14}6d^57s^2$

Group 8

26	44	76	108
Fe	**Ru**	**Os**	**Hs**
Iron	Ruthenium	Osmium	Hassium
55.845	101.07	190.23	(277)
$[Ar]3d^64s^2$	$[Kr]4d^75s^1$	$[Xe]4f^{14}5d^66s^2$	$[Rn]5f^{14}6d^67s^2$

Group 9

27	45	77	109
Co	**Rh**	**Ir**	**Mt**
Cobalt	Rhodium	Iridium	Meitnerium
58.933 195	102.905 50	192.217	(268)
$[Ar]3d^74s^2$	$[Kr]4d^85s^1$	$[Xe]4f^{14}5d^76s^2$	$[Rn]5f^{14}6d^77s^2$

* The systematic names and symbols for elements greater than 111 will be used until the approval of trivial names by IUPAC.

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58	59	60	61	62
Ce	**Pr**	**Nd**	**Pm**	**Sm**
Cerium	Praseodymium	Neodymium	Promethium	Samarium
140.116	140.907 65	144.242	(145)	150.36
$[Xe]4f^15d^16s^2$	$[Xe]4f^36s^2$	$[Xe]4f^46s^2$	$[Xe]4f^56s^2$	$[Xe]4f^66s^2$

90	91	92	93	94
Th	**Pa**	**U**	**Np**	**Pu**
Thorium	Protactinium	Uranium	Neptunium	Plutonium
232.038 06	231.035 88	238.028 91	(237)	(244)
$[Rn]6d^27s^2$	$[Rn]5f^26d^17s^2$	$[Rn]5f^36d^17s^2$	$[Rn]5f^46d^17s^2$	$[Rn]5f^67s^2$

FIGURE 6 In the common periodic table, the elements are arranged in vertical groups and in horizontal periods.

Group 18

2		
He		
Helium		
4.002 602		
$1s^2$		

Group 13 · Group 14 · Group 15 · Group 16 · Group 17

5	6	7	8	9	10
B	**C**	**N**	**O**	**F**	**Ne**
Boron	Carbon	Nitrogen	Oxygen	Fluorine	Neon
10.811	12.0107	14.0067	15.9994	18.998 4032	20.1797
$[He]2s^22p^1$	$[He]2s^22p^2$	$[He]2s^22p^3$	$[He]2s^22p^4$	$[He]2s^22p^5$	$[He]2s^22p^6$

13	14	15	16	17	18
Al	**Si**	**P**	**S**	**Cl**	**Ar**
Aluminum	Silicon	Phosphorus	Sulfur	Chlorine	Argon
26.981 5386	28.0855	30.973 762	32.065	35.453	39.948
$[Ne]3s^23p^1$	$[Ne]3s^23p^2$	$[Ne]3s^23p^3$	$[Ne]3s^23p^4$	$[Ne]3s^23p^5$	$[Ne]3s^23p^6$

Group 10 · Group 11 · Group 12

28	29	30	31	32	33	34	35	36
Ni	**Cu**	**Zn**	**Ga**	**Ge**	**As**	**Se**	**Br**	**Kr**
Nickel	Copper	Zinc	Gallium	Germanium	Arsenic	Selenium	Bromine	Krypton
58.6934	63.546	65.409	69.723	72.64	74.921 60	78.96	79.904	83.798
$[Ar]3d^84s^2$	$[Ar]3d^{10}4s^1$	$[Ar]3d^{10}4s^2$	$[Ar]3d^{10}4s^24p^1$	$[Ar]3d^{10}4s^24p^2$	$[Ar]3d^{10}4s^24p^3$	$[Ar]3d^{10}4s^24p^4$	$[Ar]3d^{10}4s^24p^5$	$[Ar]3d^{10}4s^24p^6$

46	47	48	49	50	51	52	53	54
Pd	**Ag**	**Cd**	**In**	**Sn**	**Sb**	**Te**	**I**	**Xe**
Palladium	Silver	Cadmium	Indium	Tin	Antimony	Tellurium	Iodine	Xenon
106.42	107.8682	112.411	114.818	118.710	121.760	127.60	126.904 47	131.293
$[Kr]4d^{10}5s^0$	$[Kr]4d^{10}5s^1$	$[Kr]4d^{10}5s^2$	$[Kr]4d^{10}5s^25p^1$	$[Kr]4d^{10}5s^25p^2$	$[Kr]4d^{10}5s^25p^3$	$[Kr]4d^{10}5s^25p^4$	$[Kr]4d^{10}5s^25p^5$	$[Kr]4d^{10}5s^25p^6$

78	79	80	81	82	83	84	85	86
Pt	**Au**	**Hg**	**Tl**	**Pb**	**Bi**	**Po**	**At**	**Rn**
Platinum	Gold	Mercury	Thallium	Lead	Bismuth	Polonium	Astatine	Radon
195.084	196.966 569	200.59	204.3833	207.2	208.980 40	(209)	(210)	(222)
$[Xe]4f^{14}5d^96s^1$	$[Xe]4f^{14}5d^{10}6s^1$	$[Xe]4f^{14}5d^{10}6s^2$	$[Xe]4f^{14}5d^{10}6s^26p^1$	$[Xe]4f^{14}5d^{10}6s^26p^2$	$[Xe]4f^{14}5d^{10}6s^26p^3$	$[Xe]4f^{14}5d^{10}6s^26p^4$	$[Xe]4f^{14}5d^{10}6s^26p^5$	$[Xe]4f^{14}5d^{10}6s^26p^6$

110	111	112		114		116		
Ds	**Rg**	**Uub***		**Uuq***		**Uuh***		
Darmstadtium	Roentgenium	Ununbium		Ununquadium		Ununhexium		
(271)	(272)	(285)		(289)		(292)		
$[Rn]5f^{14}6d^97s^1$	$[Rn]5f^{14}6d^{10}7s^1$	$[Rn]5f^{14}6d^{10}7s^2$		$[Rn]5f^{14}6d^{10}7s^27p^2$		$[Rn]5f^{14}6d^{10}7s^27$		

The discoveries of elements with atomic numbers 112, 114, and 116 have been reported but not fully confirmed.

63	64	65	66	67	68	69	70	71
Eu	**Gd**	**Tb**	**Dy**	**Ho**	**Er**	**Tm**	**Yb**	**Lu**
Europium	Gadolinium	Terbium	Dysprosium	Holmium	Erbium	Thulium	Ytterbium	Lutetium
151.964	157.25	158.925 35	162.500	164.930 32	167.259	168.934 21	173.04	174.967
$[Xe]4f^76s^2$	$[Xe]4f^75d^16s^2$	$[Xe]4f^96s^2$	$[Xe]4f^{10}6s^2$	$[Xe]4f^{11}6s^2$	$[Xe]4f^{12}6s^2$	$[Xe]4f^{13}6s^2$	$[Xe]4f^{14}6s^2$	$[Xe]4f^{14}5d^16s^2$

95	96	97	98	99	100	101	102	103
Am	**Cm**	**Bk**	**Cf**	**Es**	**Fm**	**Md**	**No**	**Lr**
Americium	Curium	Berkelium	Californium	Einsteinium	Fermium	Mendelevium	Nobelium	Lawrencium
(243)	(247)	(247)	(251)	(252)	(257)	(258)	(259)	(262)
$[Rn]5f^77s^2$	$[Rn]5f^76d^17s^2$	$[Rn]5f^97s^2$	$[Rn]5f^{10}7s^2$	$[Rn]5f^{11}7s^2$	$[Rn]5f^{12}7s^2$	$[Rn]5f^{13}7s^2$	$[Rn]5f^{14}7s^2$	$[Rn]5f^{14}6d^17s^2$

The atomic masses listed in this table reflect the precision of current measurements. (Each value listed in parentheses is the mass number of that radioactive element's most stable or most common isotope.)
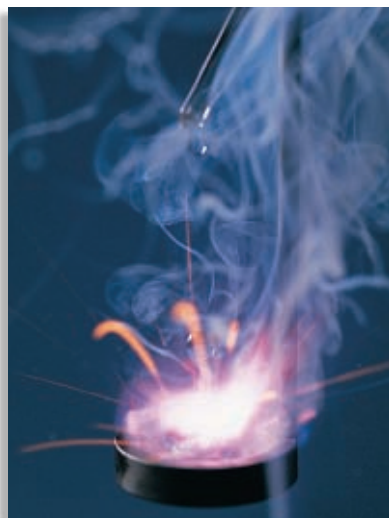
(a)
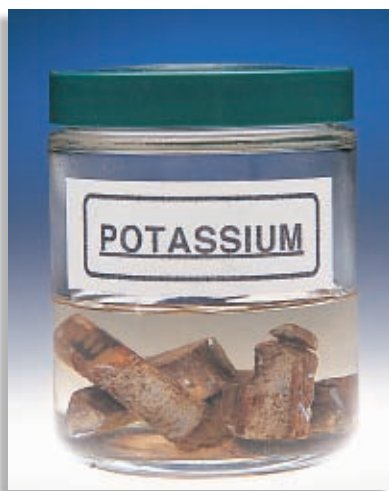
(b)

FIGURE 7 (a) Like other alkali metals, potassium reacts so strongly with water that (b) it must be stored in kerosene or oil to prevent it from reacting with moisture in the air.

The *s*-Block Elements: Groups 1 and 2

The elements of the *s* block are chemically reactive *metals*. The Group 1 metals are more reactive than those of Group 2. The outermost energy level in an atom of each Group 1 element contains a single *s* electron. For example, the configurations of lithium and sodium are $[He]2s^1$ and $[Ne]3s^1$, respectively. As you will learn in Section 3, the ease with which the single electron is lost helps to make the Group 1 metals extremely reactive. Using *n* for the number of the highest occupied energy level, the outer, or group, configurations of the Group 1 and 2 elements are written ns^1 and ns^2, respectively. For example, the configuration of Na is $[Ne]3s^1$, so the group configuration is written ns^1, where $n = 3$.

The elements of Group 1 of the periodic table (lithium, sodium, potassium, rubidium, cesium, and francium) are known as the **alkali metals.** In their pure state, all of the alkali metals have a silvery appearance and are soft enough to cut with a knife. However, because they are so reactive, alkali metals are not found in nature as free elements. They combine vigorously with most nonmetals. And they react strongly with water to produce hydrogen gas and aqueous solutions of substances known as alkalis. Because of their extreme reactivity with air or moisture, alkali metals are usually stored in kerosene. Proceeding down the column, the elements of Group 1 melt at successively lower temperatures.

The elements of Group 2 of the periodic table (beryllium, magnesium, calcium, strontium, barium, and radium) are called the **alkaline-earth metals.** Atoms of alkaline-earth metals contain a pair of electrons in their outermost *s* sublevel. Consequently, the group configuration for Group 2 is ns^2. The Group 2 metals are harder, denser, and stronger than the alkali metals. They also have higher melting points. Although they are less reactive than the alkali metals, the alkaline-earth metals are also too reactive to be found in nature as free elements.

Hydrogen and Helium

Before discussing the other blocks of the periodic table, let's consider two special cases in the classification of the elements—hydrogen and helium. Hydrogen has an electron configuration of $1s^1$, but despite the ns^1 configuration, it does not share the same properties as the elements

FIGURE 8 Calcium, an alkaline-earth metal, is too reactive to be found in nature in its pure state (a). Instead, it exists in compounds, such as in the minerals that make up marble (b).

(a)

(b)

of Group 1. Although it is located above the Group 1 elements in many periodic tables, hydrogen is a unique element, with properties that do not closely resemble those of any group.

Like the Group 2 elements, helium has an ns^2 group configuration. Yet it is part of Group 18. Because its highest occupied energy level is filled by two electrons, helium possesses special chemical stability, exhibiting the unreactive nature of a Group 18 element. By contrast, the Group 2 metals have no special stability; their highest occupied energy levels are not filled because each metal has an empty available p sublevel.

SAMPLE PROBLEM A

For more help, go to the Math Tutor at the end of this chapter.

a. Without looking at the periodic table, identify the group, period, and block in which the element that has the electron configuration [Xe]$6s^2$ is located.

b. Without looking at the periodic table, write the electron configuration for the Group 1 element in the third period. Is this element likely to be more reactive or less reactive than the element described in (a)?

SOLUTION

a. The element is in Group 2, as indicated by the group configuration of ns^2. It is in the sixth period, as indicated by the highest principal quantum number in its configuration, 6. The element is in the s block.

b. In a third-period element, the highest occupied energy level is the third main energy level, $n = 3$. The $1s$, $2s$, and $2p$ sublevels are completely filled (see **Table 1**). A Group 1 element has a group configuration of ns^1, which indicates a single electron in its highest s sublevel. Therefore, this element has the following configuration:

$$1s^2 2s^2 2p^6 3s^1 \quad \text{or} \quad [\text{Ne}]3s^1$$
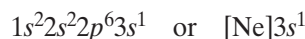
Because it is in Group 1 (the alkali metals), this element is likely to be more reactive than the element described in (a), which is in Group 2 (the alkaline-earth metals).

PRACTICE

Answers in Appendix E

1. Without looking at the periodic table, identify the group, period, and block in which the element that has the electron configuration [Kr]$5s^1$ is located.

2. **a.** Without looking at the periodic table, write the group configuration for the Group 2 elements.

   b. Without looking at the periodic table, write the complete electron configuration for the Group 2 element in the fourth period.

   c. Refer to **Figure 6** to identify the element described in (b). Then, write the element's noble-gas notation.
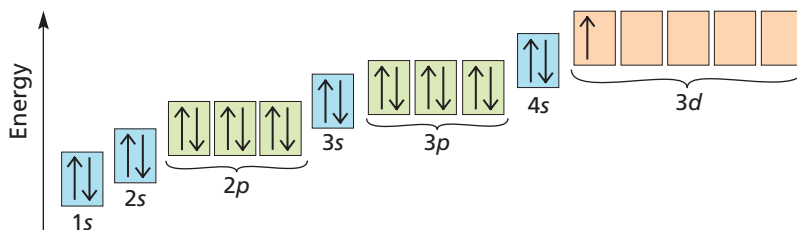
FIGURE 9

The *d*-Block Elements: Groups 3–12

For energy level n, there are n possible sublevels, so the d sublevel first appears when $n = 3$. This $3d$ sublevel is slightly higher in energy than the $4s$ sublevel, so these are filled in the order $4s3d$ (see **Figure 9**). This order of filling is also seen for higher values of n. Each d sublevel consists of five orbitals with a maximum of two electrons each, or up to 10 electrons possible in each d sublevel. In addition to the two ns electrons of Group 2, atoms of the Group 3 elements each have one electron in the d sublevel of the $(n-1)$ energy level. The group configuration for Group 3 is therefore $(n-1)d^1ns^2$. Atoms of the Group 12 elements have 10 electrons in the d sublevel plus two electrons in the ns sublevel. The group configuration for Group 12 is $(n-1)d^{10}ns^2$.

As you read in Chapter 4, some deviations from orderly d sublevel filling occur in Groups 4–11. As a result, elements in these d-block groups, unlike those in s-block and p-block groups, do not necessarily have identical outer electron configurations. For example, in Group 10, nickel, Ni, has the electron configuration $[Ar]3d^84s^2$. Palladium, Pd, has the configuration $[Kr]4d^{10}5s^0$. And platinum, Pt, has the configuration $[Xe]4f^{14}5d^96s^1$. Notice, however, that in each case the sum of the outer s and d electrons is equal to the group number.

The d-block elements are metals with typical metallic properties and are often referred to as **transition elements.** They are good conductors of electricity and have a high luster. They are typically less reactive than the alkali metals and the alkaline-earth metals. Some are so unreactive that they do not easily form compounds, existing in nature as free elements. Palladium, platinum, and gold are among the least reactive of all the elements. Some d-block elements are shown in **Figure 10.**

FIGURE 10 Mercury, tungsten, and vanadium are transition elements. Locate them in the d block of the periodic table on pages 140–141.

Mercury

Tungsten

Vanadium

Materials Scientist

Almost everything we use or wear is composed of materials. Materials can include metals, ceramics, polymers, semiconductors, and composites. Metals have been used for thousands of years. Metals are good conductors of heat and electricity. They are strong but deformable. Ceramics are typically compounds made up of both metals and nonmetals. Usually, ceramics are insulators and are resistant to high temperatures. Examples of products that contain ceramics are dishes, building materials, bone and tooth replacements, and high-speed communications equipment. Polymers are generally organic compounds and have very large molecular structures. Polymer products include toys, storage containers, paints, and biomedical items; plastics (polymers) are everywhere. The electrical properties of semiconductors are between those of conductors and insulators. Computers, calculators, and cell phones are just a few examples of products that use semiconductors. Composites, such as ceramics and polymers, are found in flooring, tiles, bicycles, space shuttles, and insulation.

Materials science includes the study of the processing, structure, properties, and performance of materials. Processing involves manufacturing the material from its raw components. Structure refers to the arrangement of the material's components from an atomic to a macro scale. Understanding the structure helps materials scientists develop new materials that have the desired properties. Typical properties of interest are mechanical, such as tensile strength or hardness; electrical, such as conductivity or resistivity; magnetic, such as magnetic susceptibility; optical, such as refractive index; thermal, such as heat capacity or thermal conductivity; and environmental, such as corrosion behavior. Performance testing and analysis ensures that the product has the desired properties. Because a property or a structural characterization of a material closely depends on the processing used, evaluation often improves the process used to make the material.

Careers in Materials Science

Materials scientists go by many different titles, such as process or production engineers, research scientists, metallurgists, polymer scientists, ceramic engineers, plant managers, and quality control engineers. Materials scientists are employed in industries such as transportation, electronics, aerospace, and biomedical engineering or work for research laboratories, the government, or universities. Materials scientists help determine what materials should be used in products based on desired properties, help synthesize new materials, and help produce these materials efficiently.

Because of the interdisciplinary nature of this field, many kinds of programs, disciplines, or majors can prepare a person for a career in materials

▲ Materials scientists study the microstructure of materials, testing how they behave under various conditions.

science. An undergraduate degree in materials science and engineering, chemistry, physics, or engineering can lead to a career in this field. Many materials scientists obtain a more advanced degree, such as a master's degree in materials science and engineering or a doctorate (Ph.D.) in chemistry, physics, or engineering.

The materials field is exciting and expanding. The ability to create materials to meet specific needs is just starting to be realized. With the developing field of materials science, many say we are living in the Materials Age.

Questions

1. Choose a product you might use in your daily life. Discuss the materials that compose the product and the properties that those materials give the product.

2. Using the Internet, find a profile of a materials scientist or engineer, and describe what that scientist or engineer does in her or his position.

For more help, go to the **Math Tutor** *at the end of this chapter.*

An element has the electron configuration [Kr]$4d^5 5s^1$. Without looking at the periodic table, identify the period, block, and group in which this element is located. Then, consult the periodic table to identify this element and the others in its group.

SOLUTION   The number of the highest occupied energy level is 5, so the element is in the fifth period. There are five electrons in the *d* sublevel, which means that it is incompletely filled. The *d* sublevel can hold 10 electrons. Therefore, the element is in the *d* block. For *d*-block elements, the number of electrons in the *ns* sublevel (1) plus the number of electrons in the $(n - 1)d$ sublevel (5) equals the group number, 6. This Group 6 element is molybdenum. The others in Group 6 are chromium, tungsten, and seaborgium.

PRACTICE   *Answers in Appendix E*

1. Without looking at the periodic table, identify the period, block, and group in which the element that has the electron configuration [Ar]$3d^8 4s^2$ is located.

2. Without looking at the periodic table, write the outer electron configuration for the Group 12 element in the fifth period.

extension
Go to **go.hrw.com** for more practice problems that ask you to use the electron configurations of elements to locate those elements in the periodic table.

Keyword: HC6PERX

The *p*-Block Elements: Groups 13–18

The *p*-block elements consist of all the elements of Groups 13–18 except helium. Electrons add to a *p* sublevel only after the *s* sublevel in the same energy level is filled. Therefore, atoms of all *p*-block elements contain two electrons in the *ns* sublevel. *The p-block elements together with the s-block elements are called the* **main-group elements.** For Group 13 elements, the added electron enters the *np* sublevel, giving a group configuration of $ns^2 np^1$. Atoms of Group 14 elements contain two electrons in the *p* sublevel, giving $ns^2 np^2$ for the group configuration. This pattern continues in Groups 15–18. In Group 18, the stable noble-gas configuration of $ns^2 np^6$ is reached. The relationships among group numbers and electron configurations for all the groups are summarized in **Table 2.**

For atoms of *p*-block elements, the total number of electrons in the highest occupied level is equal to the group number minus 10. For example, bromine is in Group 17. It has $17 - 10 = 7$ electrons in its highest energy level. Because atoms of *p*-block elements contain two electrons in the *ns* sublevel, we know that bromine has five electrons in its outer *p* sublevel. The electron configuration of bromine is [Ar]$3d^{10} 4s^2 4p^5$.

The properties of elements of the *p* block vary greatly. At its right-hand end, the *p* block includes all of the *nonmetal*s except hydrogen and helium. All six of the *metalloids* (boron, silicon, germanium, arsenic,

TABLE 2 Relationships Among Group Numbers, Blocks, and Electron Configurations

Group number	Group configuration	Block	Comments
1, 2	$ns^{1,2}$	s	One or two electrons in ns sublevel
3–12	$(n-1)d^{1-10}ns^{0-2}$	d	Sum of electrons in ns and $(n-1)d$ levels equals group number
13–18	ns^2np^{1-6}	p	Number of electrons in np sublevel equals group number minus 12

antimony, and tellurium) are also in the p block. At the left-hand side and bottom of the block, there are eight p-block metals. The locations of the nonmetals, metalloids, and metals in the p block are shown with distinctive colors in **Figure 6.**

The elements of Group 17 (fluorine, chlorine, bromine, iodine, and astatine) are known as the **halogens,** shown in **Figure 11.** The halogens are the most reactive nonmetals. They react vigorously with most metals to form examples of the type of compound known as salts. As you will see later, the reactivity of the halogens is based on the presence of seven electrons in their outer energy levels—one electron short of the stable noble-gas configuration. Fluorine and chlorine are gases at room temperature, bromine is a reddish liquid, and iodine is a dark purple solid. Astatine is a synthetic element prepared in only very small quantities. Most of its properties are estimated, although it is known to be a solid.

The metalloids, or semiconducting elements, are located between nonmetals and metals in the p block. They are mostly brittle solids with some properties of metals and some of nonmetals. The metalloid elements have electrical conductivity intermediate between that of metals, which are good conductors, and nonmetals, which are nonconductors.

The metals of the p block are generally harder and denser than the s-block alkaline-earth metals, but softer and less dense than the d-block metals. With the exception of bismuth, these metals are sufficiently reactive to be found in nature only in the form of compounds. Once obtained as free metals, however, they are stable in the presence of air.

FIGURE 11 Fluorine, chlorine, bromine, and iodine are members of Group 17 of the periodic table, also known as the halogens. Locate the halogens in the p block of the periodic table on pages 140–141.

Chlorine

Fluorine

Bromine

Iodine

For more help, go to the Math Tutor at the end of this chapter.

Without looking at the periodic table, write the outer electron configuration for the Group 14 element in the second period. Then, name the element, and identify it as a metal, nonmetal, or metalloid.

SOLUTION The group number is higher than 12, so the element is in the *p* block. The total number of electrons in the highest occupied *s* and *p* sublevels is therefore equal to the group number minus 10 (14 − 10 = 4). Two electrons are in the *s* sublevel, so two electrons must also be present in the 2*p* sublevel, which means that the outer electron configuration is $2s^2 2p^2$. The element is carbon, C, which is a nonmetal.

PRACTICE *Answers in Appendix E*

1. a. Without looking at the periodic table, write the outer electron configuration for the Group 17 element in the third period.

    b. Name the element described in (a), and identify it as a metal, nonmetal, or metalloid.

2. a. Without looking at the periodic table, identify the period, block, and group of an element that has the electron configuration $[Ar]3d^{10}4s^2 4p^3$.

    b. Name the element described in (a), and identify it as a metal, nonmetal, or metalloid.

extension
Go to **go.hrw.com** for more practice problems that ask you to classify elements and write their electron configurations based on their period and group.

Keyword: HC6PERX

The *f*-Block Elements: Lanthanides and Actinides

In the periodic table, the *f*-block elements are wedged between Groups 3 and 4 in the sixth and seventh periods. The position of these inner transition elements reflects the fact that they involve the filling of the 4*f* sublevel. With seven 4*f* orbitals to be filled with two electrons each, there are a total of 14 *f*-block elements between lanthanum, La, and hafnium, Hf, in the sixth period. The lanthanides are shiny metals similar in reactivity to the Group 2 alkaline-earth metals.

There are also 14 *f*-block elements, the actinides, between actinium, Ac, and element 104, Rf, in the seventh period. In these elements the 5*f* sublevel is being filled with 14 electrons. The actinides are all radioactive. The first four actinides (thorium, Th, through neptunium, Np) have been found naturally on Earth. The remaining actinides are known only as laboratory-made elements.

The electron configurations of atoms of four elements are written at the top of page 149. Name the block and group in which each of these elements is located in the periodic table. Then, use the periodic table on pages 140–141 to name each element. Identify each element as a metal, nonmetal, or metalloid.

Finally, describe whether each element has high reactivity or low reactivity.

a. $[Xe]4f^{14}5d^96s^1$

c. $[Ne]3s^23p^6$

b. $[Ne]3s^23p^5$

d. $[Xe]4f^66s^2$
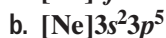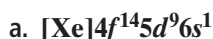
SOLUTION

a. The $4f$ sublevel is filled with 14 electrons. The $5d$ sublevel is partially filled with nine electrons. Therefore, this element is in the d block. The element is the transition metal platinum, Pt, which is in Group 10 and has a low reactivity.

b. The incompletely filled p sublevel shows that this element is in the p block. A total of seven electrons are in the ns and np sublevels, so this element is in Group 17, the halogens. The element is chlorine, Cl, and is highly reactive.

c. This element has a noble-gas configuration and thus is in Group 18 in the p block. The element is argon, Ar, which is an unreactive nonmetal and a noble gas.

d. The incomplete $4f$ sublevel shows that the element is in the f block and is a lanthanide. Group numbers are not assigned to the f block. The element is samarium, Sm. All of the lanthanides are reactive metals.

PRACTICE   *Answers in Appendix E*

1. For each of the following, identify the block, period, group, group name (where appropriate), element name, element type (metal, nonmetal, or metalloid), and relative reactivity (high or low):
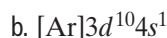
   a. $[He]2s^22p^5$

   b. $[Ar]3d^{10}4s^1$

SECTION REVIEW

1. Into what four blocks can the periodic table be divided to illustrate the relationship between the elements' electron configurations and their placement in the periodic table?

2. What name is given to each of the following groups of elements in the periodic table?

   a. Group 1   **c.** Groups 3–12   **e.** Group 18

   b. Group 2   **d.** Group 17

3. What are the relationships between group configuration and group number for elements in the $s, p,$ and d blocks?

4. Without looking at the periodic table, write the outer electron configuration for the Group 15 element in the fourth period.

5. Without looking at the periodic table, identify the period, block, and group of the element that has the electron configuration $[Ar]3d^74s^2$.

Critical Thinking

6. **APPLYING MODELS** Period 7 contains elements in the $s, p, d,$ and f blocks. Suppose that there were a Period 8 and it contained elements in the "g" block, where "g" had the angular momentum quantum number $\ell = 4$. If a hypothetical element in Period 8 had an atomic number of 120, into what group in the periodic table would the element fit, and what properties might it have (assuming it does not radioactively decay)?

Electron Configuration and Periodic Properties

OBJECTIVES

- Define *atomic* and *ionic radii*, *ionization energy, electron affinity,* and *electronegativity*.

- Compare the periodic trends of atomic radii, ionization energy, and electronegativity, and state the reasons for these variations.

- Define *valence electrons,* and state how many are present in atoms of each main-group element.

- Compare the atomic radii, ionization energies, and electronegativities of the *d*-block elements with those of the main-group elements.

So far, you have learned that the elements are arranged in the periodic table according to their atomic number and that there is a rough correlation between the arrangement of the elements and their electron configurations. In this section, the relationship between the periodic law and electron configurations will be further explored.

Atomic Radii

Ideally, the size of an atom is defined by the edge of its orbital. However, this boundary is fuzzy and varies under different conditions. Therefore, to estimate the size of an atom, the conditions under which the atom exists must be specified. One way to express an atom's radius is to measure the distance between the nuclei of two identical atoms that are chemically bonded together, then divide this distance by two. As illustrated in **Figure 12, atomic radius** *may be defined as one-half the distance between the nuclei of identical atoms that are bonded together*.

Period Trends

Figure 13 gives the atomic radii of the elements and **Figure 14** presents this information graphically. Note that there is a gradual decrease in atomic radii across the second period from lithium, Li, to neon, Ne. *The*

FIGURE 12   One method of determining atomic radius is to measure the distance between the nuclei of two identical atoms that are bonded together in an element or compound, then divide this distance by two. The atomic radius of a chlorine atom, for example, is 100 picometers (pm).

Chlorine nucleus

Atomic radius 99 pm

198 pm

Distance between nuclei

Chlorine nucleus
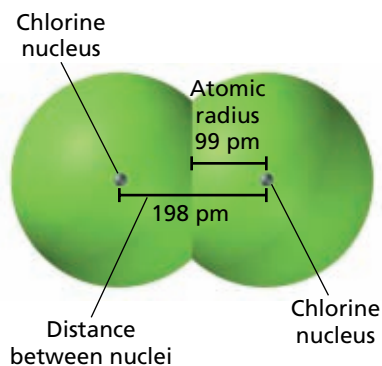
Periodic Table of Atomic Radii (pm)

Atomic symbol — C   Atomic number — 6

Relative atomic size — 77   Atomic radius

Period 1: H 1, 37; He 2, 31

Period 2: Group 1: Li 3, 152; Group 2: Be 4, 112; Group 13: B 5, 85; Group 14: C 6, 77; Group 15: N 7, 75; Group 16: O 8, 73; Group 17: F 9, 72; Group 18: Ne 10, 71

Period 3: Na 11, 186; Mg 12, 160; Al 13, 143; Si 14, 118; P 15, 110; S 16, 103; Cl 17, 100; Ar 18, 98

Period 4: K 19, 227; Ca 20, 197; Sc 21, 162 (Group 3); Ti 22, 147 (Group 4); V 23, 134 (Group 5); Cr 24, 128 (Group 6); Mn 25, 127 (Group 7); Fe 26, 126 (Group 8); Co 27, 125 (Group 9); Ni 28, 124 (Group 10); Cu 29, 128 (Group 11); Zn 30, 134 (Group 12); Ga 31, 135; Ge 32, 122; As 33, 120; Se 34, 119; Br 35, 114; Kr 36, 112

Period 5: Rb 37, 248; Sr 38, 215; Y 39, 180; Zr 40, 160; Nb 41, 146; Mo 42, 139; Tc 43, 136; Ru 44, 134; Rh 45, 134; Pd 46, 137; Ag 47, 144; Cd 48, 149; In 49, 167; Sn 50, 140; Sb 51, 140; Te 52, 142; I 53, 133; Xe 54, 131

Period 6: Cs 55, 265; Ba 56, 222; La 57, 183; Hf 72, 159; Ta 73, 146; W 74, 139; Re 75, 137; Os 76, 135; Ir 77, 136; Pt 78, 139; Au 79, 144; Hg 80, 151; Tl 81, 170; Pb 82, 175; Bi 83, 150; Po 84, 168; At 85, 140; Rn 86, 141

Period 7: Fr 87, 270; Ra 88, 220; Ac 89, 188; Rf 104, —; Db 105, —; Sg 106, —; Bh 107, —; Hs 108, —; Mt 109, —; Ds 110, —; Rg 111, —; 112; 113; 114; 115; 116; 117; 118

Lanthanide series: Ce 58, 182; Pr 59, 182; Nd 60, 181; Pm 61, 183; Sm 62, 180; Eu 63, 208; Gd 64, 180; Tb 65, 177; Dy 66, 178; Ho 67, 176; Er 68, 176; Tm 69, 176; Yb 70, —; Lu 71, 174

Actinide series: Th 90, 179; Pa 91, 163; U 92, 156; Np 93, 155; Pu 94, 159; Am 95, 173; Cm 96, 174; Bk 97, —; Cf 98, 186; Es 99, 186; Fm 100, —; Md 101, —; No 102, —; Lr 103, —

FIGURE 13 Atomic radii decrease from left to right across a period and increase down a group.

trend to smaller atoms across a period is caused by the increasing positive charge of the nucleus. As electrons add to s and p sublevels in the same main energy level, they are gradually pulled closer to the more highly charged nucleus. This increased pull results in a decrease in atomic radii. The attraction of the nucleus is somewhat offset by repulsion among the increased number of electrons in the same outer energy level. As a result, the difference in radii between neighboring atoms in each period grows smaller, as shown in **Figure 13.**

Group Trends

Examine the atomic radii of the Group 1 elements in **Figure 13.** Notice that the radii of the elements increase as you read down the group. As electrons occupy sublevels in successively higher main energy levels located farther from the nucleus, the sizes of the atoms increase. *In general, the atomic radii of the main-group elements increase down a group*.

Now examine the radii of the Group 13 elements. Although gallium, Ga, follows aluminum, Al, it has a slightly smaller atomic radius than does aluminum. This is because gallium, unlike aluminum, is preceded in its period by the 10 *d*-block elements. The expected increase in gallium's radius caused by the filling of the fourth main-energy level is outweighed by a shrinking of the electron cloud caused by a nuclear charge that is considerably higher than that of aluminum.

Atomic Radius vs. Atomic Number
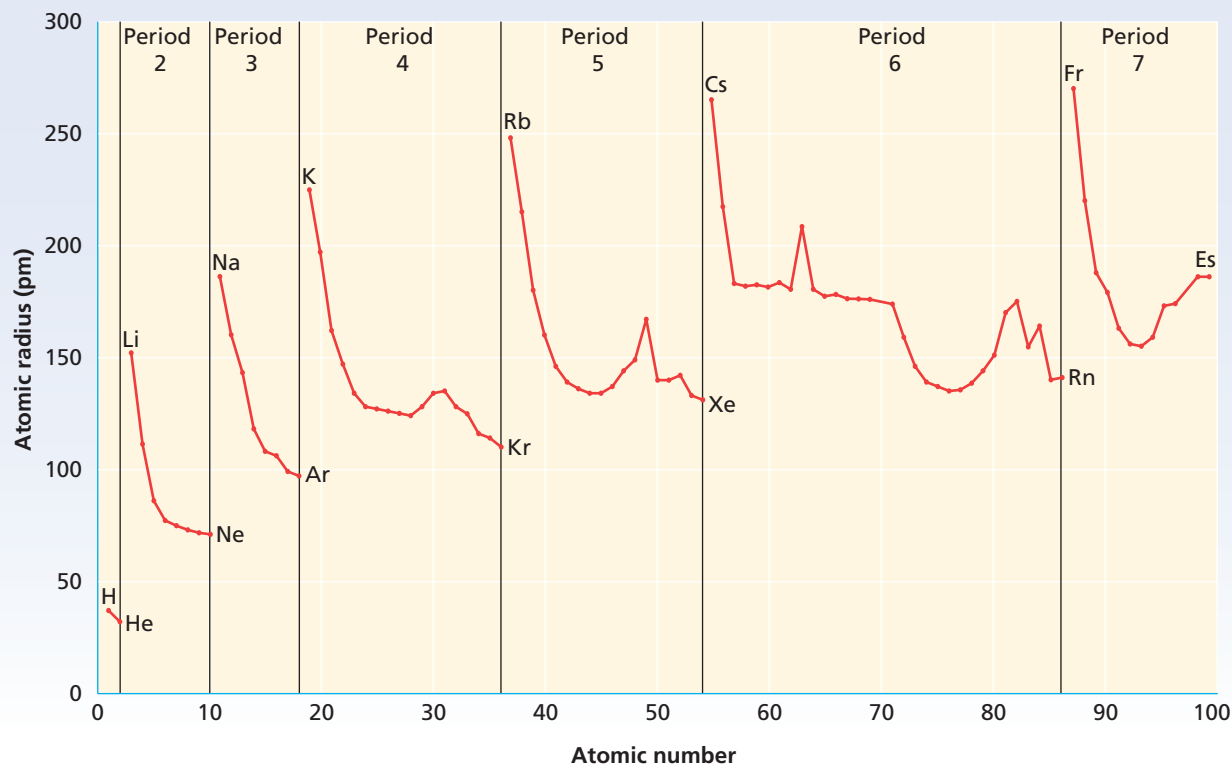
FIGURE 14 The plot of atomic radius versus atomic number shows period and group trends.

SAMPLE PROBLEM E

Of the elements magnesium, Mg, chlorine, Cl, sodium, Na, and phosphorus, P, which has the largest atomic radius? Explain your answer in terms of trends in the periodic table.

SOLUTION All of the elements are in the third period. Of the four, sodium has the lowest atomic number and is the first element in the period. Therefore, sodium has the largest atomic radius because atomic radii decrease across a period.

PRACTICE *Answers in Appendix E*

1. Which of the following elements has the largest atomic radius: Li, O, C, or F? Which has the smallest atomic radius?

2. Of the elements calcium, Ca, beryllium, Be, barium, Ba, and strontium, Sr, which has the largest atomic radius? Explain your answer in terms of trends in the periodic table.

3. Of the elements aluminum, Al, magnesium, Mg, silicon, Si, and sodium, Na, which has the smallest atomic radius? Explain your answer in terms of trends in the periodic table.

extension
Go to **go.hrw.com** for more practice problems that ask you to use periodic trends in atomic radius to predict relative sizes of atoms.

Keyword: HC6PERX

Ionization Energy

An electron can be removed from an atom if enough energy is supplied. Using A as a symbol for an atom of any element, the process can be expressed as follows.

$$A + energy \rightarrow A^+ + e^-$$
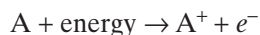
The A^+ represents an ion of element A with a single positive charge, referred to as a 1+ ion. *An* **ion** *is an atom or group of bonded atoms that has a positive or negative charge.* Sodium, for example, forms an Na^+ ion. *Any process that results in the formation of an ion is referred to as* **ionization.**

To compare the ease with which atoms of different elements give up electrons, chemists compare ionization energies. *The energy required to remove one electron from a neutral atom of an element is the* **ionization energy,** *IE (or first ionization energy, IE_1).* To avoid the influence of nearby atoms, measurements of ionization energies are made on isolated atoms in the gas phase. **Figure 15** gives the first ionization energies for the elements in kilojoules per mole (kJ/mol). **Figure 16** presents this information graphically.

FIGURE 15 In general, first ionization energies increase across a period and decrease down a group.

Periodic Table of Ionization Energies (kJ/mol)

First Ionization Energy vs. Atomic Number
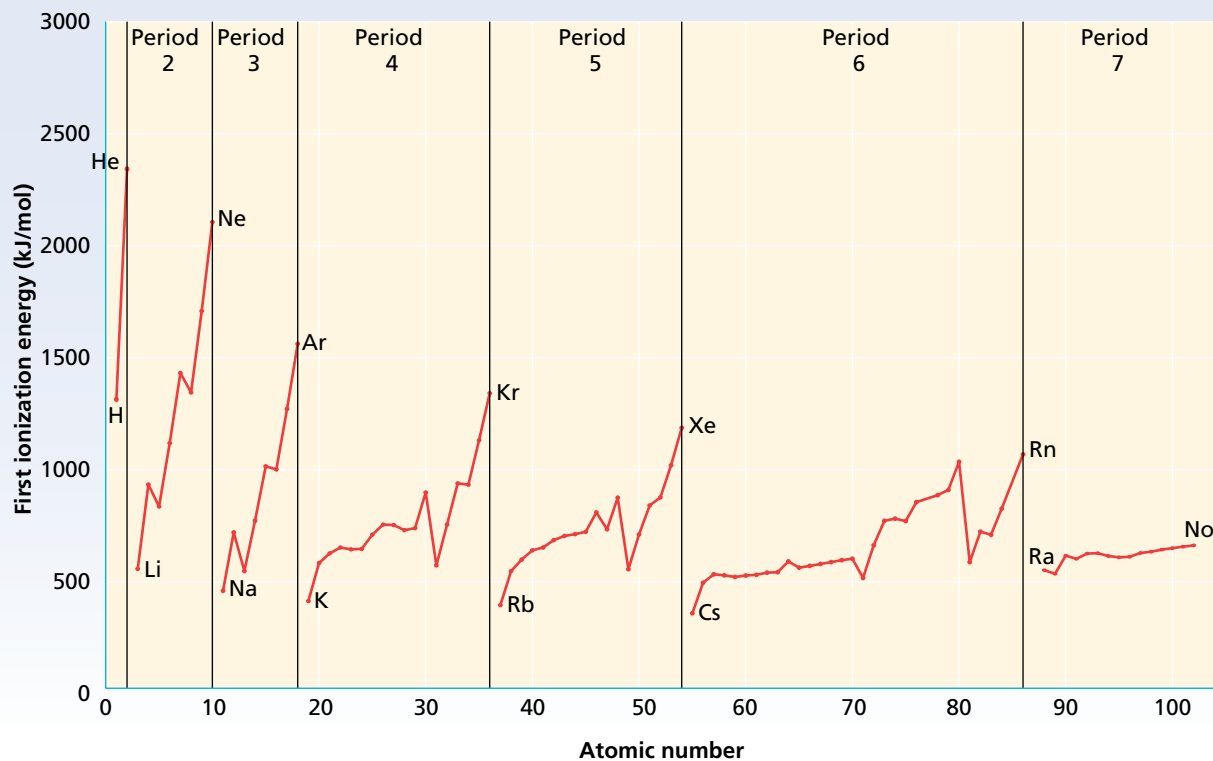
FIGURE 16 Plot of first ionization energy, IE_1, versus atomic number. As atomic number increases, both the period and the group trends become less pronounced.

Period Trends

In **Figures 15** and **16,** examine the ionization energies for the first and last elements in each period. You can see that the Group 1 metals have the lowest first ionization energies in their respective periods. Therefore, they lose electrons most easily. This ease of electron loss is a major reason for the high reactivity of the Group 1 (alkali) metals. The Group 18 elements, the noble gases, have the highest ionization energies. They do not lose electrons easily. The low reactivity of the noble gases is partly based on this difficulty of electron removal.

In general, ionization energies of the main-group elements increase across each period. This increase is caused by increasing nuclear charge. A higher charge more strongly attracts electrons in the same energy level. Increasing nuclear charge is responsible for both increasing ionization energy and decreasing radii across the periods. Note that, in general, nonmetals have higher ionization energies than metals do. In each period, the element of Group 1 has the lowest ionization energy and the element of Group 18 has the highest ionization energy.

Group Trends

Among the main-group elements, ionization energies generally decrease down the groups. Electrons removed from atoms of each succeeding element in a group are in higher energy levels, farther from the nucleus. Therefore, they are removed more easily. Also, as atomic number

increases going down a group, more electrons lie between the nucleus and the electrons in the highest occupied energy levels. This partially shields the outer electrons from the effect of the nuclear charge. Together, these influences overcome the attraction of the electrons to the increasing nuclear charge.

Removing Electrons from Positive Ions

With sufficient energy, electrons can be removed from positive ions as well as from neutral atoms. The energies for removal of additional electrons from an atom are referred to as *the second ionization energy* (IE_2), *third ionization energy* (IE_3), and so on.

Table 3 shows the first five ionization energies for the elements of the first, second, and third periods. You can see that the second ionization energy is always higher than the first, the third is always higher than the second, and so on. This is because as electrons are removed in successive ionizations, fewer electrons remain within the atom to shield the attractive force of the nucleus. *Thus, each successive electron removed from an ion feels an increasingly stronger effective nuclear charge (the nuclear charge minus the electron shielding)*.

The first ionization energies in **Table 3** show that removing a single electron from an atom of a Group 18 element is more difficult than removing an electron from atoms of other elements in the same period. This special stability of the noble-gas configuration also applies to ions that have noble-gas configurations. Notice in **Table 3** the large increases between the first and second ionization energies of lithium, Li, and between the second and third ionization energies of beryllium, Be. Even larger increases in ionization energy exist between the third and fourth

TABLE 3 *Ionization Energies (in kJ/mol) for Elements of Periods 1–3*

	Period 1		Period 2							
	H	He	Li	Be	B	C	N	O	F	Ne
IE_1	1312	2372	520	900	801	1086	1402	1314	1681	2081
IE_2		5250	7298	1757	2427	2353	2856	3388	3374	3952
IE_3			11 815	14 849	3660	4621	4578	5300	6050	6122
IE_4				21 007	25 026	6223	7475	7469	8408	9370
IE_5					32 827	37 830	9445	10 990	11 023	12 178
			Period 3							
			Na	Mg	Al	Si	P	S	Cl	Ar
IE_1			496	738	578	787	1012	1000	1251	1521
IE_2			4562	1451	1817	1577	1903	2251	2297	2666
IE_3			6912	7733	2745	3232	2912	3361	3822	3931
IE_4			9544	10 540	11 578	4356	4957	4564	5158	5771
IE_5			13 353	13 628	14 831	16 091	6274	7013	6540	7238

ionization energies of boron, B, and between the fourth and fifth ionization energies of carbon, C. In each case, the jump in ionization energy occurs when an ion assumes a noble-gas configuration. For example, the removal of one electron from a lithium atom ($[He]2s^1$) leaves the helium noble-gas configuration. The removal of four electrons from a carbon atom ($[He]2s^22p^2$) also leaves the helium configuration. A bigger table would show that this trend continues across the entire periodic system.

Consider two main-group elements, A and B. Element A has a first ionization energy of 419 kJ/mol. Element B has a first ionization energy of 1000 kJ/mol. Decide if each element is more likely to be in the *s* block or *p* block. Which element is more likely to form a positive ion?

SOLUTION   Element A has a very low ionization energy, which means that atoms of A lose electrons easily. Therefore, element A is most likely to be an *s*-block metal because ionization energies increase across the periods.

Element B has a very high ionization energy which means that atoms of B have difficulty losing electrons. Element B would most likely lie at the end of a period in the *p* block.

Element A is more likely to form a positive ion because it has a much lower ionization energy than element B does.

PRACTICE   *Answers in Appendix E*

1. Consider four hypothetical main-group elements, Q, R, T, and X, that have the outer electron configurations indicated below. Then, answer the questions that follow.

   $$Q: 3s^23p^5 \quad R: 3s^1 \quad T: 4d^{10}5s^25p^5 \quad X: 4d^{10}5s^25p^1$$
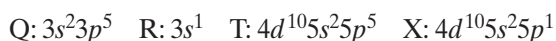
   a. Identify the block location of each hypothetical main-group element.

   b. Which of these elements are in the same period? Which are in the same group?

   c. Which element would you expect to have the highest first ionization energy? Which would have the lowest first ionization energy?

   d. Which element would you expect to have the highest second ionization energy?

   e. Which of the elements is most likely to form a 1+ ion?

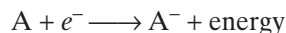# Electron Affinity

Neutral atoms can also acquire electrons. *The energy change that occurs when an electron is acquired by a neutral atom is called the atom's* **electron affinity.** Most atoms release energy when they acquire an electron.

$$A + e^- \longrightarrow A^- + \text{energy}$$

On the other hand, some atoms must be "forced" to gain an electron by the addition of energy.

$$A + e^- + \text{energy} \longrightarrow A^-$$

The quantity of energy absorbed would be represented by a positive number, but ions produced in this way are very unstable and hence the electron affinity for them is very difficult to determine. An ion produced in this way will be unstable and will lose the added electron spontaneously.

Figure 17 shows the electron affinity in kilojoules per mole for the elements. Positive electron affinities, because they are so difficult to determine with any accuracy, are denoted in **Figure 17** by "(0)." **Figure 18,** on the next page, presents these data graphically.

Period Trends

Among the elements of each period, the halogens (Group 17) gain electrons most readily. This is indicated in **Figure 17** by the large negative values of halogens' electron affinities. The ease with which halogen atoms gain electrons is a major reason for the high reactivities of the Group 17 elements. In general, as electrons add to the same p sublevel of atoms with increasing nuclear charge, electron affinities become more negative across each period within the p block. An exception to this trend occurs between Groups 14 and 15. Compare the electron affinities of carbon ($[He]2s^2 2p^2$) and nitrogen ($[He]2s^2 2p^3$). Adding an electron to a carbon atom gives a half-filled p sublevel. This occurs much more easily

FIGURE 17 The values listed in parentheses in this periodic table of electron affinities are approximate. Electron affinity is estimated to be −50 kJ/mol for each of the lanthanides and 0 kJ/mol for each of the actinides.

Periodic Table of Electron Affinities (kJ/mol)
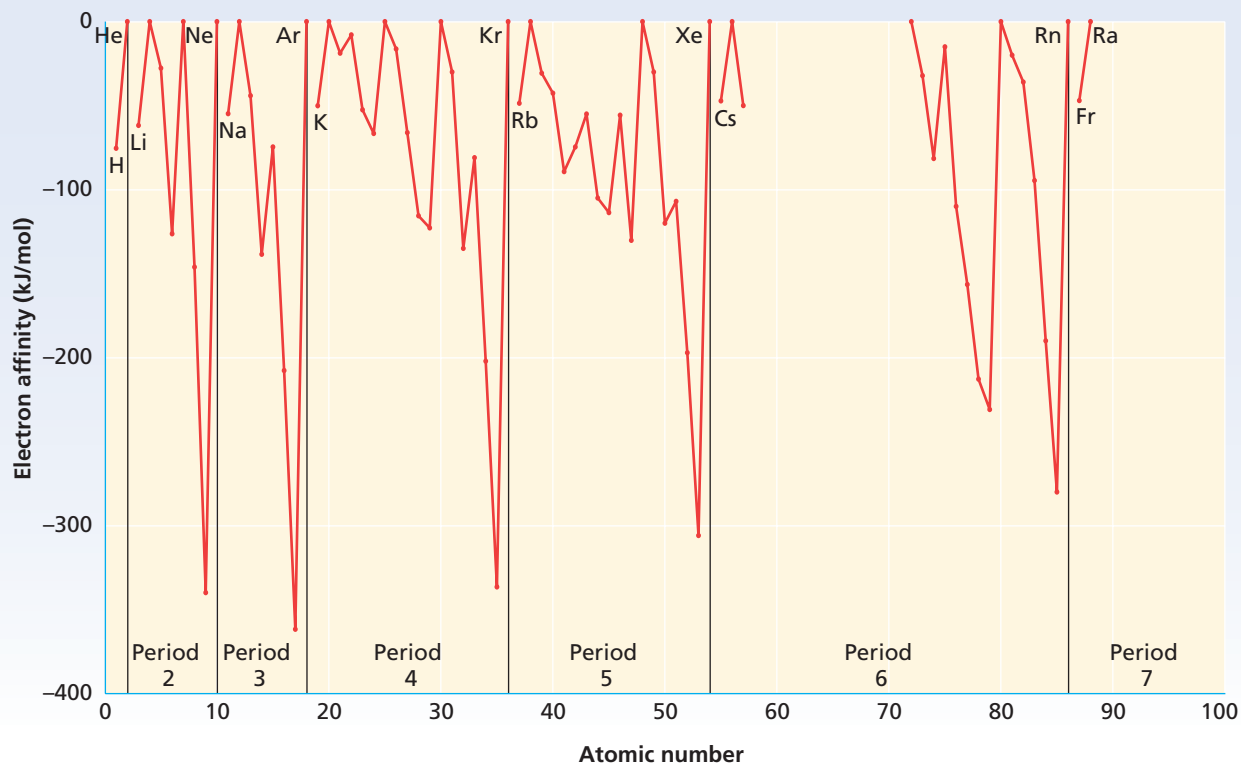
Electron Affinity vs. Atomic Number

FIGURE 18 The plot of electron affinity versus atomic number shows that most atoms release energy when they acquire an electron, as indicated by negative values.

than forcing an electron to pair with another electron in an orbital of the already half-filled p sublevel of a nitrogen atom.

Group Trends

Trends for electron affinities within groups are not as regular as trends for ionization energies. As a general rule, electrons add with greater difficulty down a group. This pattern is a result of two competing factors. The first is a slight increase in effective nuclear charge down a group, which increases electron affinities. The second is an increase in atomic radius down a group, which decreases electron affinities. In general, the size effect predominates. But there are exceptions, especially among the heavy transition metals, which tend to be the same size or even decrease in radius down a group.

Adding Electrons to Negative Ions

For an isolated ion in the gas phase, it is always more difficult to add a second electron to an already negatively charged ion. Therefore, second electron affinities are all positive. Certain p-block nonmetals tend to form negative ions that have noble gas configurations. The halogens do so by adding one electron. For example, chlorine has the configuration $[Ne]3s^23p^5$. An atom of chlorine achieves the configuration of the noble gas argon by adding an electron to form the ion Cl^- ($[Ne]3s^23p^6$). Adding another electron is so difficult that Cl^{2-} never occurs. Atoms of

Group 16 elements are present in many compounds as 2– ions. For example, oxygen ($[He]2s^2 2p^4$) achieves the configuration of the noble gas neon by adding two electrons to form the ion O^{2-}($[He]2s^2 2p^6$).

Ionic Radii

Figure 19 shows the radii of some of the most common ions of the elements. Positive and negative ions have specific names.

A positive ion is known as a **cation.** The formation of a cation by the loss of one or more electrons always leads to a decrease in atomic radius because the removal of the highest-energy-level electrons results in a smaller electron cloud. Also, the remaining electrons are drawn closer to the nucleus by its unbalanced positive charge.

A negative ion is known as an **anion.** The formation of an anion by the addition of one or more electrons always leads to an increase in atomic radius. This is because the total positive charge of the nucleus remains unchanged when an electron is added to an atom or an ion. So the electrons are not drawn to the nucleus as strongly as they were before the addition of the extra electron. The electron cloud also spreads out because of greater repulsion between the increased number of electrons.

Period Trends

Within each period of the periodic table, the metals at the left tend to form cations and the nonmetals at the upper right tend to form anions. Cationic radii decrease across a period because the electron cloud shrinks due to the increasing nuclear charge acting on the electrons in the same main energy level. Starting with Group 15, in which atoms assume stable noble-gas configurations by gaining three electrons,

FIGURE 19 The ionic radii of the ions most common in chemical compounds are shown. Cations are smaller and anions are larger than the atoms from which they are formed.

Periodic Table of Ionic Radii (pm)
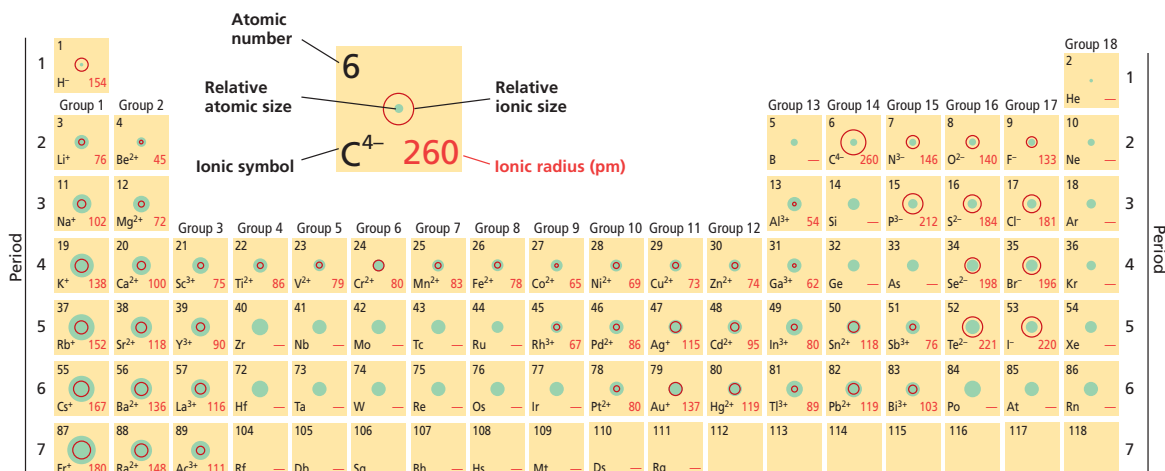
anions are more common than cations. Anionic radii decrease across each period for the elements in Groups 15–18. The reasons for this trend are the same as the reasons that cationic radii decrease from left to right across a period.

Group Trends

As they are in atoms, the outer electrons in both cations and anions are in higher energy levels as one reads down a group. Therefore, just as there is a gradual increase of atomic radii down a group, there is also a gradual increase of ionic radii.

Valence Electrons

Chemical compounds form because electrons are lost, gained, or shared between atoms. The electrons that interact in this manner are those in the highest energy levels. These are the electrons most subject to the influence of nearby atoms or ions. *The electrons available to be lost, gained, or shared in the formation of chemical compounds are referred to as* **valence electrons.** Valence electrons are often located in incompletely filled main-energy levels. For example, the electron lost from the $3s$ sublevel of Na to form Na^+ is a valence electron.

For main-group elements, the valence electrons are the electrons in the outermost s and p sublevels. The inner electrons are in filled energy levels and are held too tightly by the nucleus to be involved in compound formation. The Group 1 and Group 2 elements have one and two valence electrons, respectively, as shown in **Table 4.** The elements of Groups 13–18 have a number of valence electrons equal to the group number minus 10. In some cases, both the s and p sublevel valence electrons of the p-block elements are involved in compound formation. In other cases, only the electrons from the p sublevel are involved.

TABLE 4 *Valence Electrons in Main-Group Elements*

Group number	Group configuration	Number of valence electrons
1	ns^1	1
2	ns^2	2
13	ns^2p^1	3
14	ns^2p^2	4
15	ns^2p^3	5
16	ns^2p^4	6
17	ns^2p^5	7
18	ns^2p^6	8

Electronegativity

Valence electrons hold atoms together in chemical compounds. In many compounds, the negative charge of the valence electrons is concentrated closer to one atom than to another. This uneven concentration of charge has a significant effect on the chemical properties of a compound. It is therefore useful to have a measure of how strongly one atom attracts the electrons of another atom within a compound.

Linus Pauling, one of America's most famous chemists, devised a scale of numerical values reflecting the tendency of an atom to attract electrons. **Electronegativity** *is a measure of the ability of an atom in a chemical compound to attract electrons from another atom in the compound.* The most electronegative element, fluorine, is arbitrarily assigned an electronegativity value of four. Values for the other elements are then calculated in relation to this value.

Period Trends

As shown in **Figure 20,** *electronegativities tend to increase across each period, although there are exceptions.* The alkali and alkaline-earth metals are the least electronegative elements. In compounds, their atoms have a low attraction for electrons. Nitrogen, oxygen, and the halogens are the most electronegative elements. Their atoms attract electrons strongly in compounds. *Electronegativities tend to either decrease down a group or remain about the same.* The noble gases are unusual in that some of them

FIGURE 20 Shown are the electronegativities of the elements according to the Pauling scale. The most-electronegative elements are located in the upper right of the *p* block. The least-electronegative elements are located in the lower left of the *s* block.

Periodic Table of Electronegativities
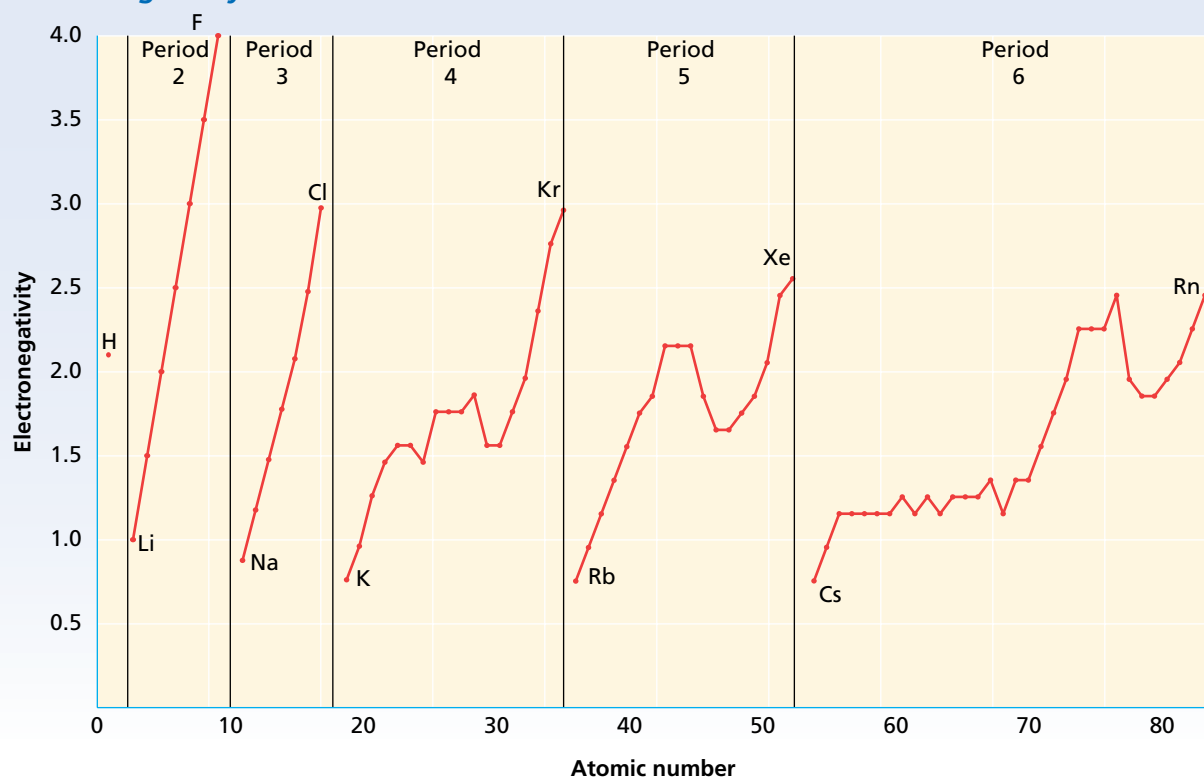
Electronegativity vs. Atomic Number

FIGURE 21 The plot shows electronegativity versus atomic number for Periods 1–6.

do not form compounds and therefore cannot be assigned electronegativities. When a noble gas does form a compound, its electronegativity is rather high, similar to the values for the halogens. The combination of the period and group trends in electronegativity results in the highest values belonging to the elements in the upper right of the periodic table. The lowest values belong to the elements in the lower left of the table. These trends are shown graphically in **Figure 21.**

SAMPLE PROBLEM G

Of the elements gallium, Ga, bromine, Br, and calcium, Ca, which has the highest electronegativity? Explain your answer in terms of periodic trends.

SOLUTION All of these elements are in the fourth period. Bromine has the highest atomic number and is farthest to the right in the period. Therefore, bromine should have the highest electronegativity because electronegativity increases across the periods.

PRACTICE *Answers in Appendix E*

1. Consider five hypothetical main-group elements, E, G, J, L, and M, that have the outer electron configurations shown at the top of the next page.

$$E = 2s^2 2p^5 \quad G = 4d^{10}5s^2 5p^5 \quad J = 2s^2 2p^2$$
$$L = 5d^{10}6s^2 6p^5 \quad M = 2s^2 2p^4$$
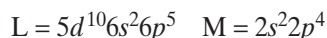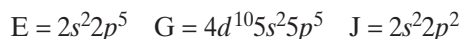
a. Identify the block location for each element. Then, determine which elements are in the same period and which are in the same group.

b. Which element would you expect to have the highest electron affinity? Which would you expect to form a 1– ion? Which should have the highest electronegativity?

c. Compare the ionic radius of the typical ion formed by the element G with the radius of the atom from which the ion was formed.

d. Which element(s) contain seven valence electrons?

Periodic Properties of the *d*- and *f*-Block Elements

The properties of the *d*-block elements (which are all metals) vary less and with less regularity than those of the main-group elements. This trend is indicated by the curves in **Figures 14** and **16,** which flatten where the *d*-block elements fall in the middle of Periods 4–6.

Recall that atoms of the *d*-block elements contain from zero to two electrons in the *s* orbital of their highest occupied energy level and one to ten electrons in the *d* sublevel of the next-lower energy level. Therefore, electrons in both the *ns* sublevel and the $(n - 1)d$ sublevel are available to interact with their surroundings. As a result, electrons in the incompletely filled *d* sublevels are responsible for many characteristic properties of the *d*-block elements.

Atomic Radii

The atomic radii of the *d*-block elements generally decrease across the periods. However, this decrease is less than that for the main-group elements because the electrons added to the $(n - 1)d$ sublevel shield the outer electrons from the nucleus. Also, note in **Figure 14** that the radii dip to a low and then increase slightly across each of the four periods that contain *d*-block elements. As the number of electrons in the *d* sublevel increases, the radii increase because of repulsion among the electrons.

In the sixth period, the *f*-block elements fall between lanthanum (Group 3) and hafnium (Group 4). Because of the increase in atomic number that occurs from lanthanum to hafnium, the atomic radius of hafnium is actually slightly less than that of zirconium, Zr, the element immediately above it. The radii of elements following hafnium in the sixth period vary with increasing atomic number in the usual manner.

Ionization Energy

As they do for the main-group elements, ionization energies of the d-block and f-block elements generally increase across the periods. In contrast to the decrease down the main groups, however, the first ionization energies of the d-block elements generally increase down each group. This is because the electrons available for ionization in the outer s sublevels are less shielded from the increasing nuclear charge by electrons in the incomplete $(n-1)d$ sublevels.

Ion Formation and Ionic Radii

Among all atoms of the d-block and f-block elements, electrons in the highest occupied sublevel are always removed first. For the d-block elements, this means that although newly added electrons occupy the d sublevels, the first electrons to be removed are those in the outermost s sublevels. For example, iron, Fe, has the electron configuration $[Ar]3d^64s^2$. First, it loses two $4s$ electrons to form Fe^{2+} ($[Ar]3d^6$). Fe^{2+} can then lose a $3d$ electron to form Fe^{3+} ($[Ar]3d^5$).

Most d-block elements commonly form 2+ ions in compounds. Some, such as iron and chromium, also commonly form 3+ ions. The Group 3 elements form only ions with a 3+ charge. Copper forms 1+ and 2+ ions, and silver usually forms only 1+ ions. As expected, the cations have smaller radii than the atoms do. Comparing 2+ ions across the periods shows a decrease in size that parallels the decrease in atomic radii.

Electronegativity

The d-block elements all have electronegativities between 1.1 and 2.54. Only the active metals of Groups 1 and 2 have lower electronegativities. The d-block elements also follow the general trend for electronegativity values to increase as radii decrease, and vice versa. The f-block elements all have similar electronegativities, which range from 1.1 to 1.5.

SECTION REVIEW

1. State the general period and group trends among main-group elements with respect to each of the following properties:

   a. atomic radii

   b. first ionization energy

   c. electron affinity

   d. ionic radii

   e. electronegativity

2. **a.** In general, how do the periodic properties of the d-block elements compare with those of the main-group elements?

   b. Explain the comparisons made in (a).

3. For each main-group element, what is the relationship between its group number and the number of valence electrons that the group members have?

Critical Thinking

4. **RELATING IDEAS** Graph the general trends (left to right and top to bottom) in the second ionization energy (IE_2) of an element as a function of its atomic number, over the range $Z = 1$–20. Label the minima and maxima on the graph with the appropriate element symbol.

CHAPTER HIGHLIGHTS

History of the Periodic Table

Vocabulary
periodic law
periodic table
lanthanide
actinide

- The periodic law states that the physical and chemical properties of the elements are periodic functions of their atomic numbers.
- The periodic table is an arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same column.
- The columns in the periodic table are referred to as groups.

Electron Configuration and the Periodic Table

Vocabulary
alkali metals
alkaline-earth metals
transition elements
main-group elements
halogens

- The rows in the periodic table are called periods.
- Many chemical properties of the elements can be explained by the configurations of the elements' outermost electrons.
- The noble gases exhibit unique chemical stability because their highest occupied levels have an octet of electrons, ns^2np^6 (with the exception of helium, whose stability arises from its highest occupied level being completely filled with two electrons, $1s^2$).
- Based on the electron configurations of the elements, the periodic table can be divided into four blocks: the s block, the p block, the d block, and the f block.

Electron Configuration and Periodic Properties

Vocabulary
atomic radius
ion
ionization
ionization energy
electron affinity
cation
anion
valence electrons
electronegativity

- The groups and periods of the periodic table display general trends in the following properties of the elements: electron affinity, electronegativity, ionization energy, atomic radius, and ionic radius.
- The electrons in an atom that are available to be lost, gained, or shared in the formation of chemical compounds are referred to as valence electrons.
- In determining the electron configuration of an ion, the order in which electrons are removed from the atom is the reverse of the order given by the atom's electron-configuration notation.

History of the Periodic Table

SECTION 1 REVIEW

1. Describe the contributions made by the following scientists to the development of the periodic table:
   a. Stanislao Cannizzaro
   b. Dmitri Mendeleev
   c. Henry Moseley
2. State the periodic law.
3. How is the periodic law demonstrated within the groups of the periodic table?

Electron Configuration and the Periodic Table

SECTION 2 REVIEW

4. a. How do the electron configurations within the same group of elements compare?
   b. Why are the noble gases relatively unreactive?
5. What determines the length of each period in the periodic table?
6. What is the relationship between the electron configuration of an element and the period in which that element appears in the periodic table?
7. a. What information is provided by the specific block location of an element?
   b. Identify, by number, the groups located within each of the four block areas.
8. a. Which elements are designated as the alkali metals?
   b. List four of their characteristic properties.
9. a. Which elements are designated as the alkaline-earth metals?
   b. How do their characteristic properties compare with those of the alkali metals?
10. a. Write the group configuration notation for each *d*-block group.
    b. How do the group numbers of those groups relate to the number of outer *s* and *d* electrons?

11. What name is sometimes used to refer to the entire set of *d*-block elements?
12. a. What types of elements make up the *p* block?
    b. How do the properties of the *p*-block metals compare with those of the metals in the *s* and *d* blocks?
13. a. Which elements are designated as the halogens?
    b. List three of their characteristic properties.
14. a. Which elements are metalloids?
    b. Describe their characteristic properties.
15. Which elements make up the *f* block in the periodic table?
16. a. What are the main-group elements?
    b. What trends can be observed across the various periods within the main-group elements?

PRACTICE PROBLEMS

17. Write the noble-gas notation for the electron configuration of each of the following elements, and indicate the period in which each belongs.
    a. Li   c. Cu   e. Sn
    b. O   d. Br
18. Without looking at the periodic table, identify the period, block, and group in which the elements with the following electron configurations are located. (Hint: See Sample Problem A.)
    a. $[Ne]3s^2 3p^4$
    b. $[Kr]4d^{10}5s^2 5p^2$
    c. $[Xe]4f^{14}5d^{10}6s^2 6p^5$
19. Based on the information given below, give the group, period, block, and identity of each element described. (Hint: See Sample Problem B.)
    a. $[He]2s^2$
    b. $[Ne]3s^1$
    c. $[Kr]5s^2$
    d. $[Ar]4s^2$
    e. $[Ar]3d^5 4s^1$
20. Without looking at the periodic table, write the expected outer electron configuration for each of the following elements. (Hint: See Sample Problem C.)
    a. Group 7, fourth period
    b. Group 3, fifth period
    c. Group 12, sixth period

21. Identify the block, period, group, group name (where appropriate), element name, element type, and relative reactivity for the elements with the following electron configurations. (Hint: See Sample Problem D.)
　a. $[Ne]3s^23p^1$
　b. $[Ar]3d^{10}4s^24p^6$
　c. $[Kr]4d^{10}5s^1$
　d. $[Xe]4f^15d^16s^2$

## *Electron Configuration and Periodic Properties*

SECTION 3 REVIEW

22. a. What is meant by atomic radius?
　b. What trend is observed among the atomic radii of main-group elements across a period?
　c. Explain this trend.
23. a. What trend is observed among the atomic radii of main-group elements down a group?
　b. Explain this trend.
24. Define each of the following terms:
　a. ion
　b. ionization
　c. first ionization energy
　d. second ionization energy
25. a. How do the first ionization energies of main-group elements vary across a period and down a group?
　b. Explain the basis for each trend.
26. a. What is electron affinity?
　b. What signs are associated with electron affinity values, and what is the significance of each sign?
27. a. Distinguish between a cation and an anion.
　b. How does the size of each compare with the size of the neutral atom from which it is formed?
28. a. What are valence electrons?
　b. Where are such electrons located?

29. For each of the following groups, indicate whether electrons are more likely to be lost or gained in compound formation and give the number of such electrons typically involved.
　a. Group 1　　　　d. Group 16
　b. Group 2　　　　e. Group 17
　c. Group 13　　　f. Group 18
30. a. What is electronegativity?
　b. Why is fluorine special in terms of electronegativity?
31. Identify the most- and least-electronegative groups of elements in the periodic table.

PRACTICE PROBLEMS

32. Of cesium, Cs, hafnium, Hf, and gold, Au, which element has the smallest atomic radius? Explain your answer in terms of trends in the periodic table. (Hint: see Sample Problem E.)
33. a. Distinguish between the first, second, and third ionization energies of an atom.
　b. How do the values of successive ionization energies compare?
　c. Why does this occur?
34. Without looking at the electron affinity table, arrange the following elements in order of *decreasing* electron affinities: C, O, Li, Na, Rb, and F.
35. a. Without looking at the ionization energy table, arrange the following elements in order of decreasing first ionization energies: Li, O, C, K, Ne, and F.
　b. Which of the elements listed in (a) would you expect to have the highest second ionization energy? Why?
36. a. Which of the following cations is least likely to form: Sr^{2+}, Al^{3+}, K^{2+}?
　b. Which of the following anions is least likely to form: I^-, Cl^-, O^{2-}?
37. Which element is the most electronegative among C, N, O, Br, and S? Which group does it belong to? (Hint: See Sample Problem G.)
38. The two ions K^+ and Ca^{2+} each have 18 electrons surrounding the nucleus. Which would you expect to have the smaller radius? Why?

MIXED REVIEW

39. Without looking at the periodic table, identify the period, block, and group in which each of the following elements is located.
  a. $[Rn]7s^1$
  b. $[Ar]3d^24s^2$
  c. $[Kr]4d^{10}5s^1$
  d. $[Xe]4f^{14}5d^96s^1$

40. a. Which elements are designated as the noble gases?
  b. What is the most significant property of these elements?

41. Which of the following does not have a noble-gas configuration: Na^+, Rb^+, O^{2-}, Br^- Ca^+, Al^{3+}, S^{2-}?

42. a. How many groups are in the periodic table?
  b. How many periods are in the periodic table?
  c. Which two blocks of the periodic table make up the main-group elements?

43. Write the noble-gas notation for the electron configuration of each of the following elements, and indicate the period and group in which each belongs.
  a. Mg
  b. P
  c. Sc
  d. Y

44. Use the periodic table to describe the chemical properties of the following elements:
  a. fluorine, F
  b. xenon, Xe
  c. sodium, Na
  d. gold, Au

45. For each element listed below, determine the charge of the ion that is most likely to be formed and the identity of the noble gas whose electron configuration is thus achieved.
  a. Li   e. Mg   i. Br
  b. Rb   f. Al   j. Ba
  c. O   g. P
  d. F   h. S

46. Describe some differences between the *s*-block metals and the *d*-block metals.

47. Why do the halogens readily form 1− ions?

48. Identify which trends in the diagrams below describe atomic radius, ionization energy, electron affinity, and electronegativity.

Increases

Increases

a.

Increases

Increases

b.

Increases

Increases

c.
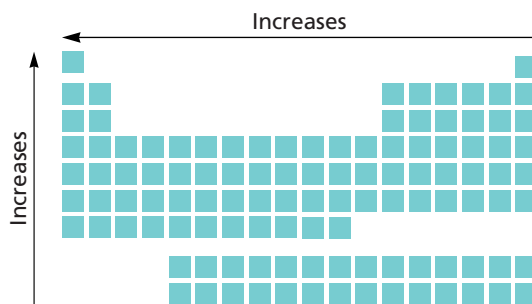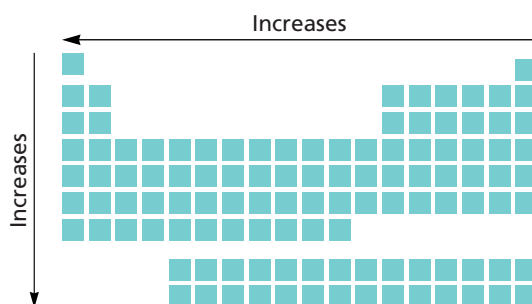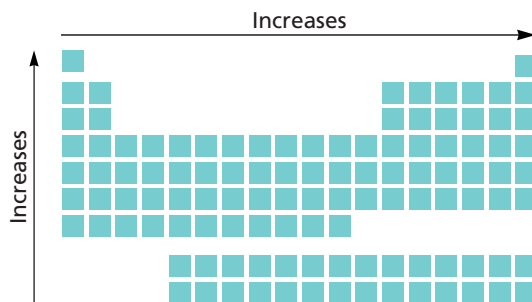
49. The electron configuration of argon differs from those of chlorine and potassium by one electron each. Compare the reactivity of these three elements.

CRITICAL THINKING

As a member on the newly-inhabited space station Alpha, you are given the task of organizing information on newly discovered elements as it comes in from the laboratory. To date, five elements have been discovered and have been assigned names and symbols from the Greek alphabet. An analysis of the new elements has yielded the following data:

Element name	Atomic no.	Atomic mass	Properties
Epsilon ε	23	47.33	nonmetal, very reactive, produces a salt when combined with a metal, gaseous state
Beta β	13	27.01	metal, very reactive, soft solid, low melting point
Gamma γ	12	25.35	nonmetal, gaseous element, extremely unreactive
Delta Δ	4	7.98	nonmetal, very abundant, forms compounds with most other elements
Lambda Λ	9	16.17	metal, solid state, good conductor, high luster, hard and dense

50. Applying Models Create a periodic table based on the properties of the five new elements.

51. Predicting Outcomes Using your newly created periodic table, predict the atomic number of an element with an atomic mass of 11.29 that has nonmetallic properties and is very reactive.

52. Predicting Outcomes Predict the atomic number of an element having an atomic mass of 15.02 that exhibits metallic properties but is softer than lambda and harder than beta.

53. Analyzing Information Analyze your periodic table for trends, and describe those trends.

USING THE HANDBOOK

54. Review the boiling point and melting point data in the tables of the *Elements Handbook.* Make a list of the elements that exist as liquids or gases at the boiling point of water, 100°C.

55. Because transition metals have vacant *d* orbitals, they form a greater variety of colored compounds than do the metals of Groups 1 and 2. Review the section of the *Elements Handbook* on transition metals (pages 798–807) and answer the following:
a. What colors are exhibited by chromium in its common oxidation states?
b. What gems contain chromium impurities?
c. What colors are often associated with the following metal ions: copper, cadmium, cobalt, zinc, and nickel?
d. What transition elements are considered noble metals? What are the characteristics of a noble metal?

RESEARCH & WRITING

56. Prepare a report tracing the evolution of the current periodic table since 1900. Cite the chemists involved and their major contributions.
57. Write a report describing the contributions of Glenn Seaborg toward the discovery of many of the actinide elements.

ALTERNATIVE ASSESSMENT

58. Construct your own periodic table or obtain a poster that shows related objects, such as fruits or vegetables, in periodic arrangement. Describe the organization of the table and the trends it illustrates. Use this table to make predictions about your subject matter.

extension

Graphing Calculator
Graphing Atomic Radii
Go to **go.hrw.com** for a graphing calculator exercise that asks you to graph atomic radii versus atomic number for the Period 2 elements.

Keyword: HC6PERX

Math Tutor WRITING ELECTRON CONFIGURATIONS

The arrangement of elements in the periodic table reflects the arrangement of electrons in an atom. Each period begins with an atom that has an electron in a new energy level and, with the exception of the first period, ends with an atom that has a filled set of *p* orbitals. To write the electron configuration of an element, you must fill the sublevels in order of increasing energy. If you follow the arrows in either of the two types of mnemonics shown below, you will get correct configurations for most elements.
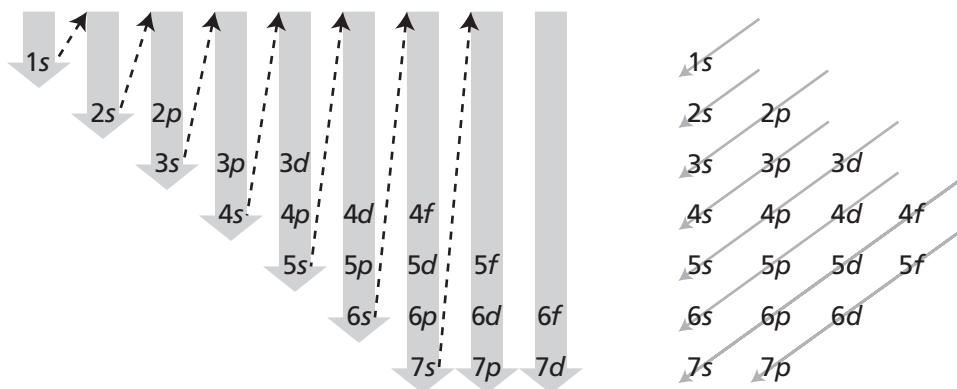

You also need to know how many orbitals are in each sublevel and that each orbital can contain two electrons of opposite spins. As shown in the following table, the sublevels *s, p, d,* and *f* have 1, 3, 5, and 7 available orbitals, respectively.

Sublevel	*s*	*p*	*d*	*f*
No. of orbitals	1	3	5	7
No. of electrons	2	6	10	14

SAMPLE

Write the full electron configuration for phosphorus.

The atomic number of phosphorus is 15, so a phosphorus atom has 15 protons and electrons. Assign each of the 15 electrons to the appropriate sublevels. The final sublevel can be unfilled and will contain the number of valence electrons.

$$1s^2 \quad 2s^2 \quad 2p^6 \quad 3s^2 \quad 3p^3$$
$$2e^- + 2e^- + 6e^- + 2e^- + 3e^- = 15e^-$$
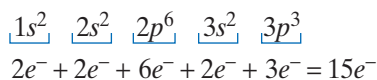
So, the full electron configuration of phosphorus is $1s^2 2s^2 2p^6 3s^2 3p^3$.

PRACTICE PROBLEMS

1. Write full electron configurations for the following elements.

  a. aluminum
  b. neon
  c. tin
  d. potassium

2. Use noble gas symbols to write shorthand electron configurations for the following elements.

  a. silicon
  b. rubidium
  c. antimony
  d. arsenic

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. In the modern periodic table, elements are arranged according to
   A. decreasing atomic mass.
   B. Mendeleev's original model.
   C. increasing atomic number.
   D. when they were discovered.

2. Group 17 elements, the halogens, are the most reactive of the nonmetal elements because they
   A. require only one electron to fill their outer energy level.
   B. have the highest ionization energies.
   C. have the largest atomic radii.
   D. are the farthest to the right in the periodic table.

3. The periodic law states that
   A. the chemical properties of elements can be grouped according to periodicity.
   B. the properties of the elements are functions of atomic mass.
   C. all elements in the same group have the same number of valence electrons.
   D. all elements with the same number of occupied energy levels must be in the same group.

4. As you move left to right across Period 3 from Mg to Cl, the energy needed to remove an electron from an atom
   A. generally increases.
   B. generally decreases.
   C. does not change.
   D. varies unpredictably.

5. Which of the following elements has the highest electronegativity?
   A. oxygen
   B. hydrogen
   C. fluorine
   D. carbon

6. The noble gases have
   A. high ionization energies.
   B. high electron affinities.
   C. large atomic radii.
   D. a tendency to form both cations and anions.

7. Which electron configuration is *not* correct?
   A. O^{2-} [He]$2s^2 2p^6$
   B. Mg^{2+} [He]$2s^2 2p^6$
   C. V^{3+} [Ar]$3d^2$
   D. Al^{3+} [Ar]$2s^2 2p^6$

8. Which two elements are more likely to have the same charge on their ions?
   A. Se and As
   B. Sn and Si
   C. Ca and Rb
   D. I and Xe

9. Using only the periodic table, choose the list that ranks the elements Sr, Te, Kr, Ru, and Cs in order of increasing ionization energy.
   A. Sr < Te < Ru < Cs < Kr
   B. Te < Ru < Sr < Cs < Kr
   C. Cs < Sr < Ru < Te < Kr
   D. Kr < Cs < Sr < Ru < Te

SHORT ANSWER

10. The second ionization energies for the elements S–Ti are listed in a scrambled order below. Assign the correct IE_2 value to each element. (Hint: S has IE_2 = 2251 kJ/mol, and Ti has IE_2 = 1310 kJ/mol.) Explain your reasoning.

    IE_2 values (kJ/mol): 2666, 2297, 3051, 1235, 2251, 1310, and 1145

11. What group most commonly forms 2– ions? Explain your reasoning.

EXTENDED RESPONSE

12. An ordered list of atomic radii for 14 consecutive elements is shown below. Without using **Figure 13** on page 151, make a graph of these atomic radii versus the element's atomic number. Explain your reasoning.

    Atomic radii (pm): 75, 73, 72, 71, 186, 160, 143, 118, 110, 103, 100, 98, 227, and 197

Test TIP If you are short on time, quickly scan the unanswered questions to see which might be easiest to answer.

The Mendeleev Lab of 1869

OBJECTIVES

- *Observe* the physical properties of common elements.

- *Observe* the properties and trends in the elements on the periodic table.

- *Draw conclusions* and *identify* unknown elements based on observed trends in properties.

MATERIALS

- blank periodic table
- elemental samples: Ar, C, Sn, and Pb
- note cards, 3 in. × 5 in.
- periodic table

BACKGROUND

Russian chemist Dmitri Mendeleev is generally credited with being the first chemist to observe that patterns emerge when the elements are arranged according to their properties. Mendeleev's arrangement of the elements was unique because he left blank spaces for elements that he claimed were undiscovered as of 1869. Mendeleev was so confident that he even predicted the properties of these undiscovered elements. His predictions were eventually proven to be quite accurate, and these new elements fill the spaces that originally were blank in his table. Use your knowledge of the periodic table to determine the identity of each of the nine unknown elements in this activity. These unknown elements are from the periodic table's groups that are listed below. Each of these groups contains at least one unknown element.

<div align="center">1   2   11   13   14   17   18</div>

None of the known elements serves as one of the nine unknown elements. No radioactive elements are used during this experiment. The relevant radioactive elements include Fr, Ra, At, and Rn. You may not use your textbook or other reference materials. You have been provided with enough information to determine each of the unknown elements.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Make a data table to record the properties of each unknown that you test.

2. Use the note cards to copy the information listed on each of the sample cards. If the word *observe* is listed, you will need to visually inspect the sample and then write the observation in the appropriate space.

PROCEDURE

1. Arrange the note cards of the known elements in a rough representation of the periodic table. In other words, all of the known elements from Group 1 should be arranged in the appropriate order. Arrange all of the other cards accordingly.

2. Inspect the properties of the unknowns to see where properties would best "fit" the trends of the elements of each group.

3. Assign the proper element name to each of the unknowns. Add the symbol for each one of the unknown elements to your data table.

CLEANUP AND DISPOSAL

4. Clean up your lab station, and return the leftover note cards and samples of the elements to your teacher. Do not pour any of the samples down the drain or place them in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all your work is finished.

ANALYSIS AND INTERPRETATION

1. **Organizing Ideas:** In what order did your group arrange the properties to determine the unknowns? Explain your reasoning. Would a different order have been better? If so, what is the better order and why?

2. **Evaluating Methods:** What properties were the most useful in sorting the unknowns? What properties were the least useful? Explain your answer.

CONCLUSIONS

1. **Interpreting Information:** Summarize your group's reasoning for the assignment of each unknown. Explain in a few sentences exactly how you predicted the identity of the nine unknown elements.

EXTENSIONS

1. **Predicting Outcomes:** Use only the data from your group's experiment to predict the properties of the not yet discovered element, which has an atomic number of 120 (assuming it does not radioactively decay).

Chemical Bonding

*In nature, most atoms are joined
to other atoms by chemical bonds.*
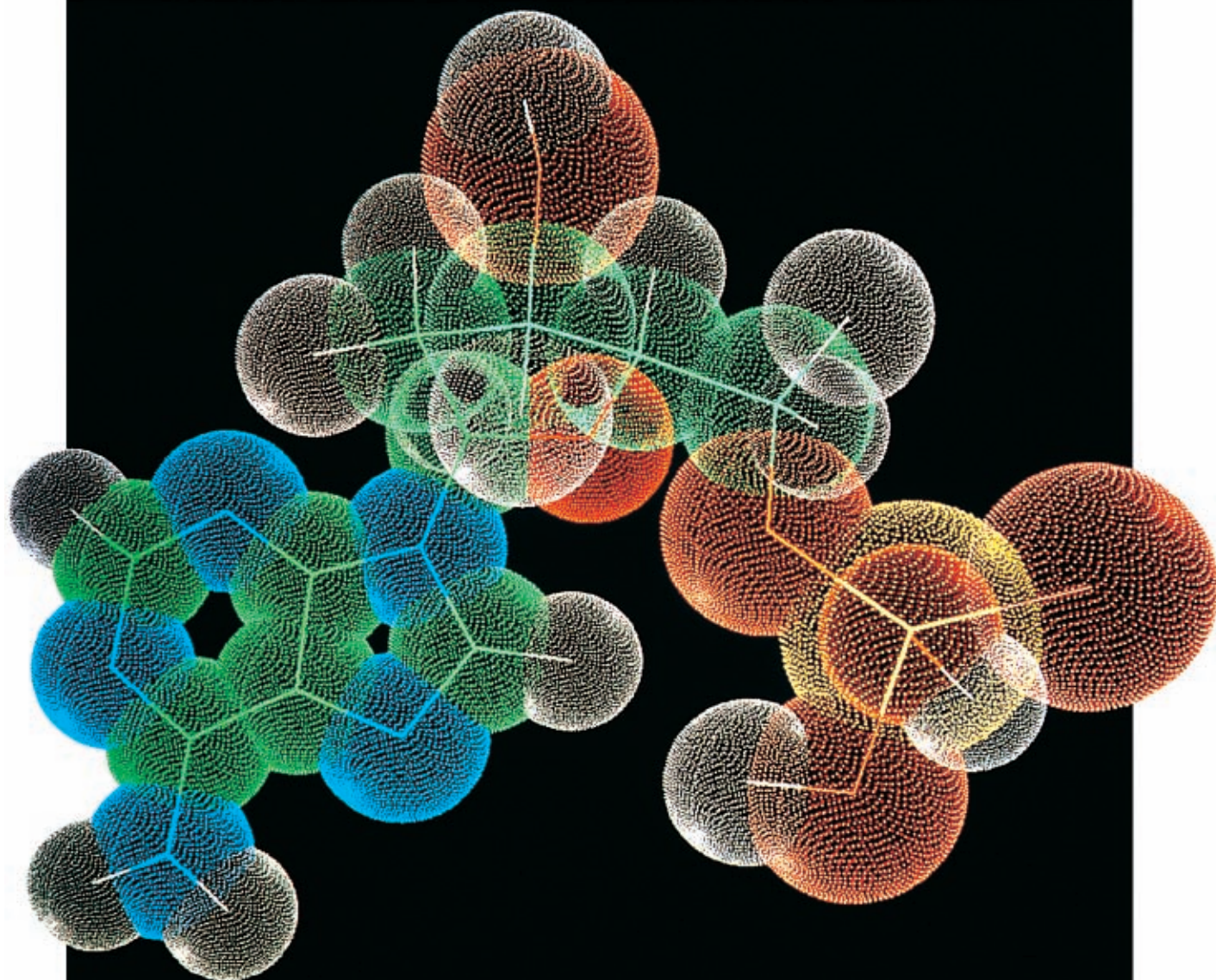
**Computer Graphic Representation of a
Molecule of Adenosine Monophosphate**

Introduction to Chemical Bonding

OBJECTIVES

- Define *chemical bond.*

- Explain why most atoms form chemical bonds.

- Describe ionic and covalent bonding.

- Explain why most chemical bonding is neither purely ionic nor purely covalent.

- Classify bonding type according to electronegativity differences.

Atoms seldom exist as independent particles in nature. The oxygen you breathe, the water you drink, and nearly all other substances consists of combinations of atoms that are held together by chemical bonds. *A* **chemical bond** *is a mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together.*

Why are most atoms chemically bonded to each other? As independent particles, most atoms are at relatively high potential energy. Nature, however, favors arrangements in which potential energy is minimized. This means that most atoms are less stable existing by themselves than when they are combined. By bonding with each other, atoms decrease in potential energy, thereby creating more stable arrangements of matter.
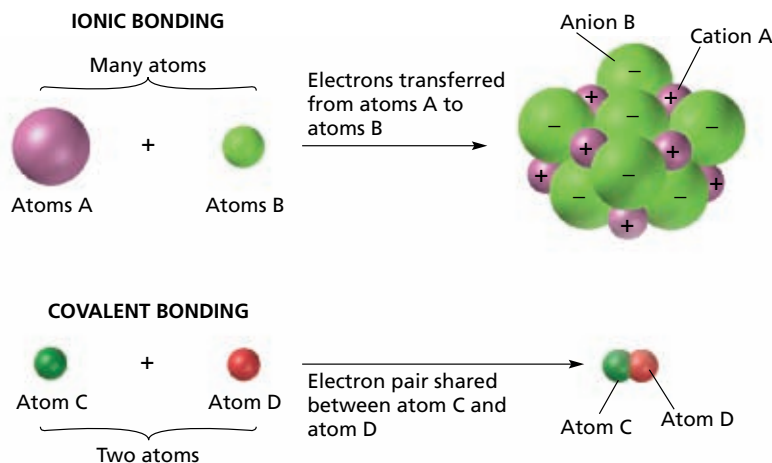
Types of Chemical Bonding

When atoms bond, their valence electrons are redistributed in ways that make the atoms more stable. The way in which the electrons are redistributed determines the type of bonding. In Chapter 5, you read that main-group metals tend to lose electrons to form positive ions, or cations, and nonmetals tend to gain electrons to form negative ions, or anions. *Chemical bonding that results from the electrical attraction between cations and anions is called* **ionic bonding.** In purely ionic bonding, atoms completely give up electrons to other atoms, as illustrated in **Figure 1** on the next page. In contrast to atoms joined by ionic bonding, atoms joined by covalent bonding share electrons. **Covalent bonding** *results from the sharing of electron pairs between two atoms* (see **Figure 1**). In a purely covalent bond, the shared electrons are "owned" equally by the two bonded atoms.

Ionic or Covalent?

Bonding between atoms of different elements is rarely purely ionic or purely covalent. It usually falls somewhere between these two extremes, depending on how strongly the atoms of each element attract electrons. Recall that electronegativity is a measure of an atom's ability to attract electrons. The degree to which bonding between atoms of two elements is ionic or covalent can be estimated by calculating the difference in

FIGURE 1 In ionic bonding, many atoms transfer electrons. The resulting positive and negative ions combine due to mutual electrical attraction. In covalent bonding, atoms share electron pairs to form independent molecules.

IONIC BONDING

Many atoms

Atoms A + Atoms B

Electrons transferred from atoms A to atoms B →

Anion B   Cation A

COVALENT BONDING

Atom C + Atom D

Two atoms

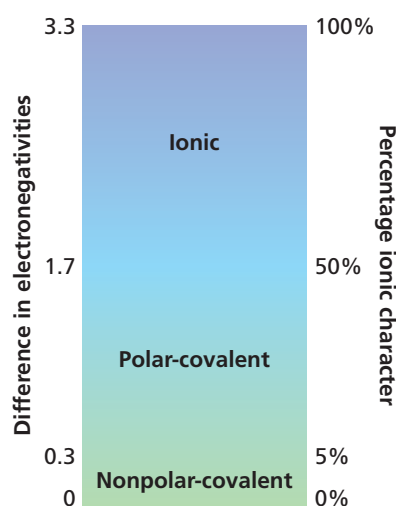Electron pair shared between atom C and atom D →

Atom C   Atom D

FIGURE 2 Differences in electronegativities reflect the character of bonding between elements. The electronegativity of the less-electronegative element is subtracted from that of the more-electronegative element. The greater the electronegativity difference, the more ionic is the bonding.

the elements' electronegativities (see **Figure 2**). For example, the electronegativity difference between fluorine, F, and cesium, Cs, is 4.0 – 0.7 = 3.3. (See **Figure 20** on page 161 for a periodic table of electronegativity values.) So, according to **Figure 2,** cesium-fluorine bonding is ionic. Fluorine atoms, which are highly electronegative, gain valence electrons, causing the atoms to become anions. Cesium atoms, which are less electronegative, lose valence electrons, causing the atoms to become cations.

Bonding between atoms with an electronegativity difference of 1.7 or less has an ionic character of 50% or less. These compounds are typically classified as covalent. Bonding between two atoms of the same element is completely covalent. Hydrogen, for example, exists in nature not as isolated atoms, but as pairs of atoms held together by covalent bonds. The hydrogen-hydrogen bond is a **nonpolar-covalent bond,** *a covalent bond in which the bonding electrons are shared equally by the bonded atoms, resulting in a balanced distribution of electrical charge.* Bonds having 0% to 5% ionic character, corresponding to electronegativity differences of roughly 0 to 0.3, are generally considered nonpolar-covalent bonds. In bonds with significantly different electronegativities, the electrons are more strongly attracted by the more-electronegative atom. *Such bonds are* **polar,** *meaning that they have an uneven distribution of charge.* Covalent bonds having 5% to 50% ionic character, corresponding to electronegativity differences of 0.3 to 1.7, are classified as polar. *A* **polar-covalent bond** *is a covalent bond in which the bonded atoms have an unequal attraction for the shared electrons.*

Nonpolar- and polar-covalent bonds are compared in **Figure 3,** which illustrates the electron density distribution in hydrogen-hydrogen and hydrogen-chlorine bonds. The electronegativity difference between chlorine and hydrogen is 3.0 – 2.1 = 0.9, indicating a polar-covalent bond. The electrons in this bond are closer to the more-electronegative chlorine atom than to the hydrogen atom, as indicated in **Figure 3b.** Consequently, the chlorine end of the bond has a partial negative charge, indicated by the symbol δ–. The hydrogen end of the bond then has an equal partial positive charge, δ+.

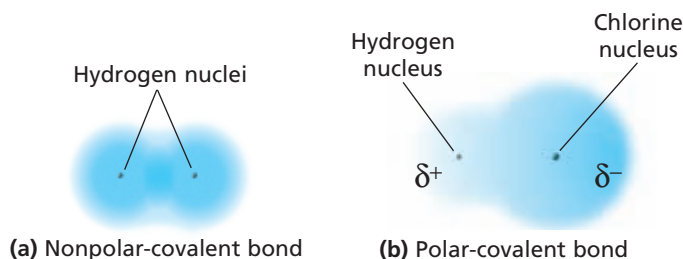(a) Nonpolar-covalent bond (b) Polar-covalent bond

FIGURE 3 Comparison of the electron density in (a) a nonpolar, hydrogen-hydrogen bond and (b) a polar, hydrogen-chlorine bond. Because chlorine is more electronegative than hydrogen, the electron density in the hydrogen-chlorine bond is greater around the chlorine atom.

SAMPLE PROBLEM A

Use electronegativity differences and Figure 2 to classify bonding between sulfur, S, and the following elements: hydrogen, H; cesium, Cs; and chlorine, Cl. In each pair, which atom will be more negative?

SOLUTION From **Figure 20** on page 161, we know that the electronegativity of sulfur is 2.5. The electronegativities of hydrogen, cesium, and chlorine are 2.1, 0.7, and 3.0, respectively. In each pair, the atom with the larger electronegativity will be the more-negative atom.

Bonding between sulfur and	Electronegativity difference	Bond type	More-negative atom
hydrogen	2.5 − 2.1 = 0.4	polar-covalent	sulfur
cesium	2.5 − 0.7 = 1.8	ionic	sulfur
chlorine	3.0 − 2.5 = 0.5	polar-covalent	chlorine

PRACTICE *Answers in Appendix E*

Use electronegativity differences and **Figure 2** to classify bonding between chlorine, Cl, and the following elements: calcium, Ca; oxygen, O; and bromine, Br. Indicate the more-negative atom in each pair.

SECTION REVIEW

1. What is the main distinction between ionic and covalent bonding?

2. How is electronegativity used in determining the ionic or covalent character of the bonding between two elements?

3. What type of bonding would be expected between the following atoms?
   a. Li and F
   b. Cu and S
   c. I and Br

4. List the three pairs of atoms referred to in the previous question in order of increasing ionic character of the bonding between them.

Critical Thinking

5. **INTERPRETING CONCEPTS** Compare the following two pairs of atoms: Cu and Cl; I and Cl.
   a. Which pair would have a bond with a greater percent ionic character?
   b. In which pair would Cl have the greater negative charge?

6. **INFERRING RELATIONSHIPS** The isolated K atom is larger than the isolated Br atom.
   a. What type of bond is expected between K and Br?
   b. Which ion in the compound KBr is larger?

Covalent Bonding and Molecular Compounds

Many chemical compounds, including most of the chemicals that are in living things and are produced by living things, are composed of molecules. *A* **molecule** *is a neutral group of atoms that are held together by covalent bonds.* A single molecule of a chemical compound is an individual unit capable of existing on its own. It may consist of two or more atoms of the same element, as in oxygen, or of two or more different atoms, as in water or sugar (see **Figure 4** below). *A chemical compound whose simplest units are molecules is called a* **molecular compound.**

The composition of a compound is given by its chemical formula. *A* **chemical formula** *indicates the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts.* The chemical formula of a molecular compound is referred to as a molecular formula. *A* **molecular formula** *shows the types and numbers of atoms combined in a single molecule of a molecular compound.* The molecular formula for water, for example, is H_2O, which reflects the fact that a single water molecule consists of one oxygen atom joined by separate covalent bonds to two hydrogen atoms. A molecule of oxygen, O_2, is an example of a diatomic molecule. *A diatomic molecule* is a molecule containing only two atoms.

(a) Water molecule, H_2O

(b) Oxygen molecule, O_2

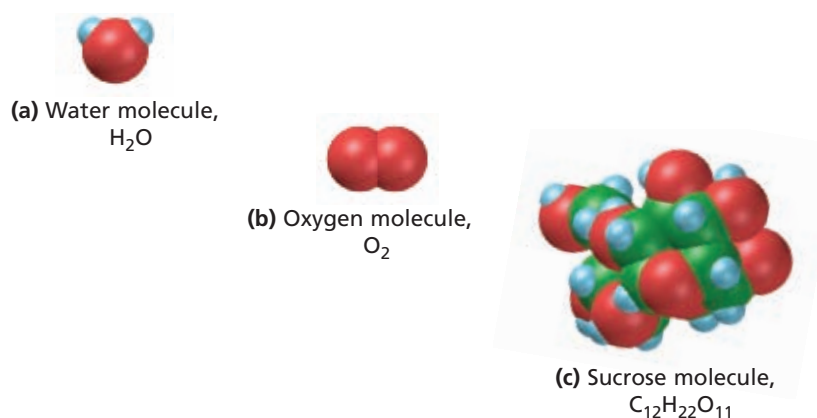**(c)** Sucrose molecule, $C_{12}H_{22}O_{11}$

FIGURE 4 The models for (a) water, (b) oxygen, and (c) sucrose, or table sugar, represent a few examples of the many molecular compounds in and around us. Atoms within molecules may form one or more covalent bonds.
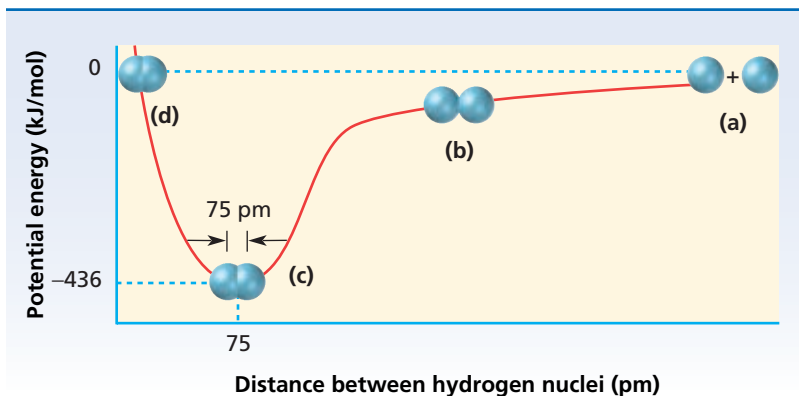
Distance between hydrogen nuclei (pm)

Formation of a Covalent Bond

As you read in Section 1, nature favors chemical bonding because most atoms have lower potential energy when they are bonded to other atoms than they have as they are independent particles. In the case of covalent bond formation, this idea is illustrated by a simple example, the formation of a hydrogen-hydrogen bond.

Picture two isolated hydrogen atoms separated by a distance large enough to prevent them from influencing each other. At this distance, the overall potential energy of the atoms is arbitrarily set at zero, as shown in part (a) of **Figure 5.**

Now consider what happens if the hydrogen atoms approach each other. Each atom has a nucleus containing a single positively charged proton. The nucleus of each atom is surrounded by a negatively charged electron in a spherical 1s orbital. As the atoms near each other, their charged particles begin to interact. As shown in **Figure 6,** the approaching nuclei and electrons are *attracted* to each other, which corresponds to a *decrease* in the total potential energy of the atoms. At the same time, the two nuclei *repel* each other and the two electrons *repel* each other, which results in an *increase* in potential energy.

The relative strength of attraction and repulsion between the charged particles depends on the distance separating the atoms. When the atoms first "sense" each other, the electron-proton attraction is stronger than the electron-electron and proton-proton repulsions. Thus, the atoms are drawn to each other and their potential energy is lowered, as shown in part (b) of **Figure 5.**

The attractive force continues to dominate and the total potential energy continues to decrease until, eventually, a distance is reached at which the repulsion between the like charges equals the attraction of the opposite charges. This is shown in part (c) of **Figure 5.** At this point, which is represented by the bottom of the valley in the curve, potential energy is at a minimum and a stable hydrogen molecule forms. A closer approach of the atoms, shown in part (d) of **Figure 5,** results in a sharp rise in potential energy as repulsion becomes increasingly greater than attraction.
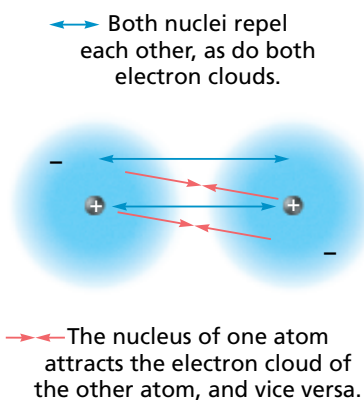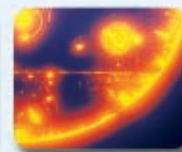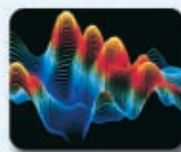
Both nuclei repel each other, as do both electron clouds.

The nucleus of one atom attracts the electron cloud of the other atom, and vice versa.

Chemistry in Action

Ultrasonic Toxic-Waste Destroyer

Paints, pesticides, solvents, and sulfides are just a few components of the 3 million tons of toxic waste that flow out of U.S. factories every year. Some of this waste ends up in groundwater and contaminates our streams and drinking water.

Eliminating hazardous waste is a constant challenge. Unfortunately, today's disposal methods often damage the environment as much as they help it. Incinerators burning certain waste, for example, produce dioxins, one of the most dangerous class of toxins known to man.

Finding new methods to destroy toxic waste is a puzzle. Michael Hoffmann, a professor of environmental chemistry at the California Institute of Technology, thinks that part of the solution lies in sound-wave technology.

According to Hoffmann, cavitation is the key to eliminating certain chemical wastes from polluted water. Cavitation occurs when the pressure in water is made to fluctuate from slightly above to slightly below normal, causing bubbles. The bubbles are unstable and collapse, creating tiny areas of extremely high pressure and heat. The pressure inside a collapsing bubble can be 1000 times greater than normal, and the temperature reaches about 5000°C—just a bit cooler than the surface of the sun. These conditions are harsh enough to combust most toxic-waste compounds in the water, breaking them down into harmless components.

Hoffmann has employed a device that uses ultrasound—sound waves at frequencies just above the range of human hearing—to create cavitation in polluted water. As water flows between two panels that generate ultrasound at different frequencies, the ultrasonic waves generated by one panel form cavitation bubbles. An instant later, the ultrasound produced by the other panel collapses the bubbles. The intense pressure and heat generated break down toxic compounds into innocuous substances, such as carbon dioxide, chloride ions, and hydrogen ions.

"With ultrasound," says Hoffmann, "we can harness frequencies . . . of about 16 kilohertz up to 1 megahertz, and different . . . compounds are destroyed more readily at one frequency versus another . . . applying a particular frequency range, we can destroy a very broad range of chemical compounds."

The device destroys simple toxins in a few minutes and other toxins in several hours. To be destroyed completely, some compounds must form intermediate chemicals first and then be treated again. To be sure the waste is totally removed, scientists use sophisticated tracking methods to trace what happens to every single molecule of the toxin.

The ultrasound toxic-waste destroyer treats about 10% of all types of waste, eliminating both organic and inorganic compounds, such as hydrogen cyanide, TNT, and many pesticides. While the device cannot destroy complex mixtures of compounds, such as those found in raw sewage, it does have many advantages over current technologies. Aside from having no harmful environmental side effects, ultrasonic waste destruction is cheaper and simpler than the process of combustion.

Questions

1. How does Dr. Hoffmann's ultrasound device benefit society?

2. Briefly explain why the bulk temperature of the water remains low (at room temperature).

SC*i*LINKS.

www.scilinks.org
Topic: Ultrasound
Code: HC61576

Characteristics of the Covalent Bond

In **Figure 5,** the bottom of the valley in the curve represents the balance between attraction and repulsion in a stable covalent bond. At this point, the electrons of each hydrogen atom of the hydrogen molecule are shared between the nuclei. As shown below in **Figure 7,** the molecule's electrons can be pictured as occupying overlapping orbitals, moving about freely in either orbital.

The bonded atoms vibrate a bit, but as long as their potential energy remains close to the minimum, they are covalently bonded to each other. The distance between two bonded atoms at their minimum potential energy, that is, the average distance between two bonded atoms, is the *bond length.* The bond length of a hydrogen-hydrogen bond is 75 pm.

In forming a covalent bond, the hydrogen atoms release energy as they change from isolated individual atoms to parts of a molecule. The amount of energy released equals the difference between the potential energy at the zero level (separated atoms) and that at the bottom of the valley (bonded atoms) in **Figure 5.** The same amount of energy must be added to separate the bonded atoms. **Bond energy** *is the energy required to break a chemical bond and form neutral isolated atoms.* Scientists usually report bond energies in kilojoules per mole (kJ/mol), which indicates the energy required to break one mole of bonds in isolated molecules. For example, 436 kJ of energy is needed to break the hydrogen-hydrogen bonds in one mole of hydrogen molecules and form two moles of separated hydrogen atoms.

The energy relationships described here for the formation of a hydrogen-hydrogen bond apply generally to all covalent bonds. However, bond lengths and bond energies vary with the types of atoms that have combined. Even the energy of a bond between the same two types of atoms varies somewhat, depending on what other bonds the atoms have formed. These facts should be considered when examining the data in **Table 1** on the next page. The first three columns in the table list bonds, bond lengths, and bond energies of atoms in specific diatomic molecules. The last three columns give average values of specified bonds in many different compounds.

Electron clouds

Region of orbital overlap

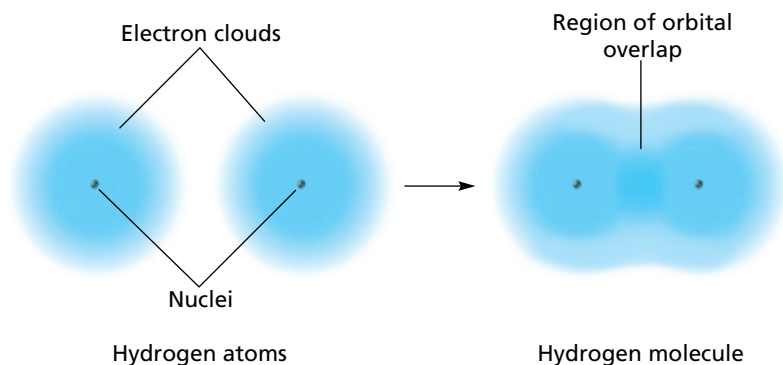Nuclei

Hydrogen atoms

Hydrogen molecule

FIGURE 7 The orbitals of the hydrogen atoms in a hydrogen molecule overlap, allowing each electron to feel the attraction of both nuclei. The result is an increase in electron density between the nuclei.

TABLE 1 Bond Lengths and Bond Energies for Selected Covalent Bonds

Bond	Average bond length (pm)	Average bond energy (kJ/mol)	Bond	Average bond length (pm)	Average bond energy (kJ/mol)
H−H	75	436	C−C	154	346
F−F	142	159	C−N	147	305
Cl−Cl	199	243	C−O	143	358
Br−Br	229	193	C−H	109	418
I−I	266	151	C−Cl	177	327
H−F	92	569	C−Br	194	285
H−Cl	127	432	N−N	145	163
H−Br	141	366	N−H	101	386
H−I	161	299	O−H	96	459

All individual hydrogen atoms contain a single, unpaired electron in a $1s$ atomic orbital. When two hydrogen atoms form a molecule, they share electrons in a covalent bond. As **Figure 8** shows, sharing electrons allows each atom to have the stable electron configuration of helium, $1s^2$. This tendency for atoms to achieve noble-gas configurations by bonding covalently extends beyond the simple case of a hydrogen molecule.

FIGURE 8   By sharing electrons in overlapping orbitals, each hydrogen atom in a hydrogen molecule experiences the effect of a stable $1s^2$ configuration.

Bonding electron pair in overlapping orbitals

Hydrogen atoms → Hydrogen molecule
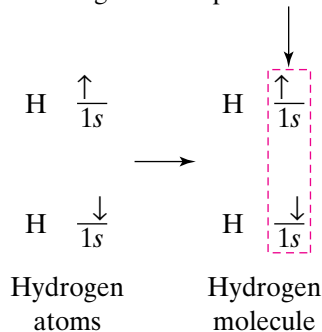
The Octet Rule

Unlike other atoms, the noble-gas atoms exist independently in nature. They possess a minimum of energy existing on their own because of the special stability of their electron configurations. This stability results from the fact that, with the exception of helium and its two electrons in a completely filled outer shell, the noble-gas atoms' outer s and p orbitals are completely filled by a total of eight electrons. Other main-group atoms can effectively fill their outermost s and p orbitals with electrons by sharing electrons through covalent bonding. Such bond

formation follows the *octet rule:* Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet of electrons in its highest occupied energy level.

Let's examine how the bonding in a fluorine molecule illustrates the octet rule. An independent fluorine atom has seven electrons in its highest energy level ($[He]2s^22p^5$). Like hydrogen atoms, fluorine atoms bond covalently with each other to form diatomic molecules, F_2. When two fluorine atoms bond, each atom shares one of its valence electrons with its partner. The shared electron pair effectively fills each atom's outermost energy level with an octet of electrons, as illustrated in **Figure 9a. Figure 9b** shows another example of the octet rule, in which the chlorine atom in a molecule of hydrogen chloride, HCl, achieves an outermost octet by sharing an electron pair with an atom of hydrogen.

(a)

Fluorine atoms → Fluorine molecule

Bonding electron pair in overlapping orbitals

(b)

Hydrogen and chlorine atoms → Hydrogen chloride molecule

Bonding electron pair in overlapping orbitals
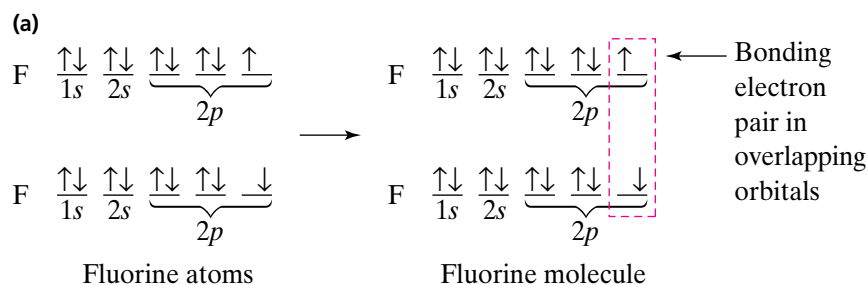
FIGURE 9 (a) By sharing valence electrons in overlapping orbitals, each atom in a fluorine molecule feels the effect of neon's stable configuration, $[He]2s^22p^6$. (b) In a hydrogen chloride molecule, the hydrogen atom effectively fills its 1s orbital with two electrons, while the chlorine atom experiences the stability of an outermost octet of electrons.
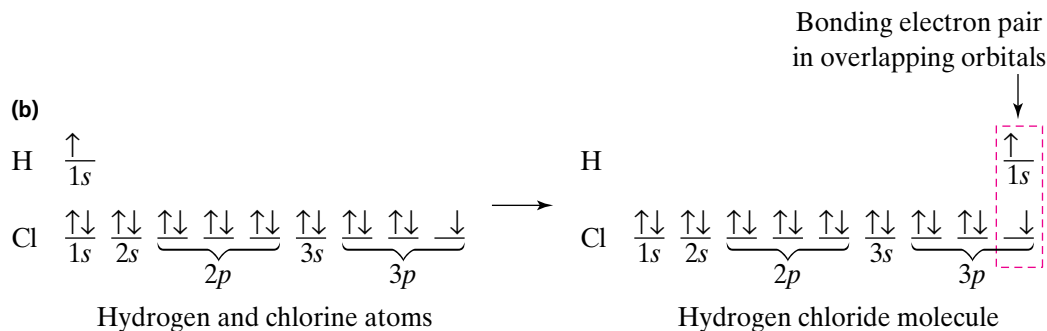
Exceptions to the Octet Rule

Most main-group elements tend to form covalent bonds according to the octet rule. However, there are exceptions. As you have seen, hydrogen forms bonds in which it is surrounded by only two electrons. Boron, B, has just three valence electrons ($[He]2s^22p^1$). Because electron pairs are shared in covalent bonds, boron tends to form bonds in which it is surrounded by six electrons. In boron trifluoride, BF_3, for example, the boron atom is surrounded by its own three valence electrons plus one from each of the three fluorine atoms bonded to it. Other elements can be surrounded by *more* than eight electrons when they combine with the highly electronegative elements fluorine, oxygen, and chlorine. In these cases of *expanded valence,* bonding involves electrons in *d* orbitals as well as in *s* and *p* orbitals. Examples of compounds that have an expanded valence include PF_5 and SF_6, as shown in **Table 5.**

Number of valence electrons	Electron-dot notation	Example
1	X·	Na·
2	·X·	·Mg·
3	·X:	·B·
4	·X:	·C·
5	·X:	·N:
6	:X:	:O:
7	:X:	:F:
8	:X:	:Ne:

FIGURE 10 To write an element's electron-dot notation, determine the element's number of valence electrons. Then place a corresponding number of dots around the element's symbol, as shown.

Electron-Dot Notation

Covalent bond formation usually involves only the electrons in an atom's outermost energy levels, or the atom's valence electrons. To keep track of these electrons, it is helpful to use electron-dot notation. **Electron-dot notation** *is an electron-configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol.* The inner-shell electrons are not shown. For example, the electron-dot notation for a fluorine atom (electron configuration $[He]2s^2 2p^5$) may be written as follows.

$$:\overset{..}{\underset{..}{F}}:$$

In general, an element's number of valence electrons can be determined by adding the superscripts of the element's noble-gas notation. In this book, the electron-dot notations for elements with 1–8 valence electrons are written as shown in **Figure 10.**

a. **Write the electron-dot notation for hydrogen.**
b. **Write the electron-dot notation for nitrogen.**

SOLUTION   a. A hydrogen atom has only one occupied energy level, the $n = 1$ level, which contains a single electron. Therefore, the electron-dot notation for hydrogen is written as follows.

$$H·$$

b. The group notation for nitrogen's family of elements is $ns^2 np^3$, which indicates that nitrogen has five valence electrons. Therefore, the electron-dot notation for nitrogen is written as follows.

$$·\overset{.}{\underset{.}{N}}:$$

Lewis Structures

Electron-dot notation can also be used to represent molecules. For example, a hydrogen molecule, H_2, is represented by combining the notations of two individual hydrogen atoms, as follows.

$$H:H$$

The pair of dots represents the shared electron pair of the hydrogen-hydrogen covalent bond. For a molecule of fluorine, F_2, the electron-dot notations of two fluorine atoms are combined.

$$\ddot{\underset{..}{F}}:\ddot{\underset{..}{F}}:$$

Here also the pair of dots between the two symbols represents the shared pair of a covalent bond. In addition, each fluorine atom is surrounded by three pairs of electrons that are not shared in bonds. An *unshared pair,* also called a *lone pair,* is a pair of electrons that is not involved in bonding and that belongs exclusively to one atom.

The pair of dots representing a shared pair of electrons in a covalent bond is often replaced by a long dash. According to this convention, hydrogen and fluorine molecules are represented as follows.

$$H-H \qquad :\ddot{\underset{..}{F}}-\ddot{\underset{..}{F}}:$$
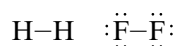
These representations are all **Lewis structures,** *formulas in which atomic symbols represent nuclei and inner-shell electrons, dot-pairs or dashes between two atomic symbols represent electron pairs in covalent bonds, and dots adjacent to only one atomic symbol represent unshared electrons.* It is common to write Lewis structures that show only the electrons that are shared, using dashes to represent the bonds. *A* **structural formula** *indicates the kind, number, arrangement, and bonds but not the unshared pairs of the atoms in a molecule.* For example, $F-F$ and $H-Cl$ are structural formulas.

The Lewis structures (and therefore the structural formulas) for many molecules can be drawn if one knows the composition of the molecule and which atoms are bonded to each other. The following sample problem illustrates the basic steps for writing Lewis structures. The molecule described in this problem contains bonds with single shared electron pairs. A single covalent bond, or a **single bond,** *is a covalent bond in which one pair of electrons is shared between two atoms.*

SAMPLE PROBLEM C   *For more help, go to the* **Math Tutor** *at the end of this chapter.*

Draw the Lewis structure of iodomethane, CH_3I.

SOLUTION
1. *Determine the type and number of atoms in the molecule.*
   The formula shows one carbon atom, one iodine atom, and three hydrogen atoms.

2. *Write the electron-dot notation for each type of atom in the molecule.*
   Carbon is from Group 14 and has four valence electrons. Iodine is from Group 17 and has seven valence electrons. Hydrogen has one valence electron.
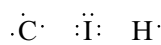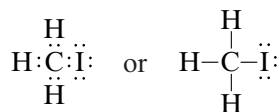
$$\cdot\dot{\underset{.}{C}}\cdot \qquad :\ddot{\underset{..}{I}}: \qquad H\cdot$$

3. *Determine the total number of valence electrons available in the atoms to be combined.*

$$
\begin{array}{lll}
\text{C} & 1 \times 4e^- = & 4e^- \\
\text{I} & 1 \times 7e^- = & 7e^- \\
\text{3H} & 3 \times 1e^- = & \underline{3e^-} \\
& & 14e^-
\end{array}
$$

4. *Arrange the atoms to form a skeleton structure for the molecule. If carbon is present, it is the central atom. Otherwise, the least-electronegative atom is central (except for hydrogen, which is never central). Then connect the atoms by electron-pair bonds.*

$$
\begin{array}{c}
\text{H} \\
\text{H} : \ddot{\text{C}} : \text{I} \\
\text{H}
\end{array}
$$

5. *Add unshared pairs of electrons to each nonmetal atom (except hydrogen) such that each is surrounded by eight electrons.*

$$
\begin{array}{ccc}
\text{H} & & \text{H} \\
\text{H} : \ddot{\text{C}} : \ddot{\text{I}} : & \text{or} & \text{H} - \underset{|}{\overset{|}{\text{C}}} - \ddot{\text{I}} : \\
\text{H} & & \text{H}
\end{array}
$$

6. *Count the electrons in the structure to be sure that the number of valence electrons used equals the number available. Be sure the central atom and other atoms besides hydrogen have an octet.*

There are eight electrons in the four covalent bonds and six electrons in the three unshared pairs, giving the correct total of 14 valence electrons.

PRACTICE   *Answers in Appendix E*

1. Draw the Lewis structure of ammonia, NH_3.

2. Draw the Lewis structure for hydrogen sulfide, H_2S.

3. Draw the Lewis structure for silane, SiH_4.

4. Draw the Lewis structure for phosphorus trifluoride, PF_3.

Multiple Covalent Bonds

Atoms of some elements, especially carbon, nitrogen, and oxygen, can share more than one electron pair. A double covalent bond, or simply a *double bond,* is a covalent bond in which two pairs of electrons are shared between two atoms. A double bond is shown either by two side-by-side pairs of dots or by two parallel dashes. All four electrons in a double bond "belong" to both atoms. In ethene, C_2H_4, for example, two electron pairs are simultaneously shared by two carbon atoms.

$$
\begin{array}{ccc}
\overset{\text{H}}{\underset{\text{H}}{\diagup}}\text{C} :: \text{C}\overset{\diagup \text{H}}{\diagdown_{\text{H}}} & \text{or} & \overset{\text{H}}{\underset{\text{H}}{\diagup}}\text{C} = \text{C}\overset{\diagup \text{H}}{\diagdown_{\text{H}}}
\end{array}
$$
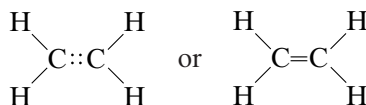
A triple covalent bond, or simply a *triple bond,* is a covalent bond in which three pairs of electrons are shared between two atoms. For example, elemental nitrogen, N_2, like hydrogen and the halogens, normally exists as diatomic molecules. In this case, however, each nitrogen atom, which has five valence electrons, acquires three electrons to complete an octet by sharing three pairs of electrons with its partner. This is illustrated in the Lewis structure and the formula structure for N_2, as shown below.

$$:N:::N: \quad \text{or} \quad :N{\equiv}N:$$

Figure 11 represents nitrogen's triple bond through orbital notation. Like the single bonds in hydrogen and halogen molecules, the triple bond in nitrogen molecules is nonpolar.

Carbon forms a number of compounds containing triple bonds. For example, the compound ethyne, C_2H_2, contains a carbon-carbon triple bond.

$$H:C:::C:H \quad \text{or} \quad H{-}C{\equiv}C{-}H$$

Double and triple bonds are referred to as **multiple bonds,** *or* multiple covalent bonds. Double bonds in general have greater bond energies and are shorter than single bonds. Triple bonds are even stronger and shorter. **Table 2** compares average bond lengths and bond energies for some single, double, and triple bonds.

In writing Lewis structures for molecules that contain carbon, nitrogen, or oxygen, one must remember that multiple bonds between pairs of these atoms are possible. (A hydrogen atom, on the other hand, has only one electron and therefore always forms a single covalent bond.) The need for a multiple bond becomes obvious if there are not enough valence electrons to complete octets by adding unshared pairs. Sample Problem D on the next page shows how to deal with this situation.
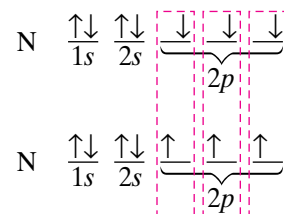
Nitrogen molecule

FIGURE 11 In a molecule of nitrogen, N_2, each nitrogen atom is surrounded by six shared electrons plus one unshared pair of electrons. Thus, each nitrogen atom follows the octet rule in forming a triple covalent bond.

TABLE 2 *Bond Lengths and Bond Energies for Single and Multiple Covalent Bonds*

| Bond	Average bond length (pm)	Average bond energy (kJ/mol)	Bond	Average bond length (pm)	Average bond energy (kJ/mol)
C−C	154	346	C−O	143	358
C=C	134	612	C=O	120	732
C≡C	120	835	C≡O	113	1072
C−N	147	305	N−N	145	163
C=N	132	615	N=N	125	418
C≡N	116	887	N≡N	110	945

Draw the Lewis structure for methanal, CH$_2$O, which is also known as formaldehyde.

SOLUTION

1. *Determine the number of atoms of each element present in the molecule.*
   The formula shows one carbon atom, two hydrogen atoms, and one oxygen atom.

2. *Write the electron-dot notation for each type of atom.*
   Carbon is from Group 14 and has four valence electrons. Oxygen, which is in Group 16, has six valence electrons. Hydrogen has only one electron.
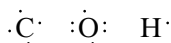
   $$\cdot\dot{\underset{\cdot}{C}}\cdot \quad :\dot{\underset{\cdot}{O}}: \quad H\cdot$$

3. *Determine the total number of valence electrons available in the atoms to be combined.*

   $$\begin{array}{lll}
   \mathbf{C} & 1 \times 4e^- = & 4e^- \\
   \mathbf{O} & 1 \times 6e^- = & 6e^- \\
   \mathbf{2H} & 2 \times 1e^- = & \underline{2e^-} \\
   & & 12e^-
   \end{array}$$

4. *Arrange the atoms to form a skeleton structure for the molecule, and connect the atoms by electron-pair bonds.*

   $$\begin{array}{c}H \\ H\!:\!C\!:\!O\end{array}$$

5. *Add unshared pairs of electrons to each nonmetal atom (except hydrogen) such that each is surrounded by eight electrons.*

   $$\begin{array}{c}H \\ H\!:\!\dot{\underset{\cdot\cdot}{C}}\!:\!\dot{\underset{\cdot\cdot}{O}}:\end{array}$$
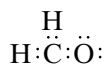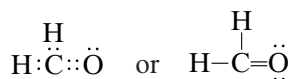
6a. *Count the electrons in the Lewis structure to be sure that the number of valence electrons used equals the number available.*
   The structure above has six electrons in covalent bonds and eight electrons in four lone pairs, for a total of 14 electrons. The structure has two valence electrons too many.

6b. *If too many electrons have been used, subtract one or more lone pairs until the total number of valence electrons is correct. Then move one or more lone electron pairs to existing bonds between non-hydrogen atoms until the outer shells of all atoms are completely filled.*
   Subtract the lone pair of electrons from the carbon atom. Then move one lone pair of electrons from the oxygen to the bond between carbon and oxygen to form a double bond.

   $$\begin{array}{c}H \\ H\!:\!\ddot{C}\!::\!\dot{\underset{\cdot\cdot}{O}}\end{array} \quad \text{or} \quad \begin{array}{c}H \\ | \\ H\!-\!C\!=\!\dot{\underset{\cdot\cdot}{O}}\end{array}$$

   There are eight electrons in covalent bonds and four electrons in lone pairs, for a total of 12 valence electrons.

PRACTICE

Answers in Appendix E

1. Draw the Lewis structure for carbon dioxide, CO$_2$.

2. Draw the Lewis structure for hydrogen cyanide, which contains one hydrogen atom, one carbon atom, and one nitrogen atom.
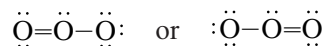
Resonance Structures

Some molecules and ions cannot be represented adequately by a single Lewis structure. One such molecule is ozone, O_3, which can be represented by either of the following Lewis structures.

$$\ddot{O}=\ddot{O}-\ddot{O}: \quad \text{or} \quad :\ddot{O}-\ddot{O}=\ddot{O}$$

Notice that each structure indicates that the ozone molecule has two types of O—O bonds, one single and one double. Chemists once speculated that ozone split its time existing as one of these two structures, constantly alternating, or "resonating," from one to the other. Experiments, however, revealed that the oxygen-oxygen bonds in ozone are identical. Therefore, scientists now say that ozone has a single structure that is the average of these two structures. Together the structures are referred to as *resonance structures* or *resonance hybrids*. **Resonance** *refers to bonding in molecules or ions that cannot be correctly represented by a single Lewis structure.* To indicate resonance, a double-headed arrow is placed between a molecule's resonance structures.

$$\ddot{O}=\ddot{O}-\ddot{O}: \longleftrightarrow :\ddot{O}-\ddot{O}=\ddot{O}$$
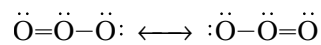
Covalent-Network Bonding

All the covalent compounds that you have read about so far are molecular. They consist of many identical molecules held together by forces acting between the molecules. (You will read more about these intermolecular forces in Section 5.) There are many covalently bonded compounds that do not contain individual molecules, but instead can be pictured as continuous, three-dimensional networks of bonded atoms. You will read more about covalently bonded networks in Chapter 7.

SECTION REVIEW

1. Define the following:
   a. bond length   b. bond energy
2. State the octet rule.
3. How many pairs of electrons are shared in the following types of covalent bonds?
   a. a single bond
   b. a double bond
   c. a triple bond

4. Draw the Lewis structures for the following molecules:
   a. IBr   d. $SiCl_4$
   b. CH_3Br e. OF_2
   c. C_2HCl

Critical Thinking

5. **APPLYING MODELS** Compare the molecules H_2NNH_2 and HNNH. Which molecule has the stronger N-N bond?

Ionic Bonding and Ionic Compounds

OBJECTIVES

● Compare and contrast a chemical formula for a molecular compound with one for an ionic compound.

● Discuss the arrangements of ions in crystals.

● Define *lattice energy* and explain its significance.

● List and compare the distinctive properties of ionic and molecular compounds.

● Write the Lewis structure for a polyatomic ion given the identity of the atoms combined and other appropriate information.

Most of the rocks and minerals that make up Earth's crust consist of positive and negative ions held together by ionic bonding. A familiar example of an ionically bonded compound is sodium chloride, or common table salt, which is found in nature as rock salt. A sodium ion, Na^+, has a charge of 1+. A chloride ion, Cl^-, has a charge of 1−. There is an electrical force of attraction between oppositely charged ions. In sodium chloride, these ions combine in a one-to-one ratio—Na^+Cl^-—so that each positive charge is balanced by a negative charge. The chemical formula for sodium chloride is usually written simply as NaCl.

An **ionic compound** *is composed of positive and negative ions that are combined so that the numbers of positive and negative charges are equal.* Most ionic compounds exist as crystalline solids (see **Figure 12**). A crystal of any ionic compound is a three-dimensional network of positive and negative ions mutually attracted to one another. As a result, in contrast to a molecular compound, an ionic compound is not composed of independent, neutral units that can be isolated and examined. The chemical formula of an ionic compound merely represents the simplest ratio of the compound's combined ions that gives electrical neutrality.

The chemical formula of an ionic compound shows the ratio of the ions present in a sample of any size. *A* **formula unit** *is the simplest collection of atoms from which an ionic compound's formula can be established.* For example, one formula unit of sodium chloride, NaCl, is one sodium cation plus one chloride anion. (In the naming of a monatomic anion, the ending of the element's name is replaced with *-ide.* See Chapter 7 for more details.)

The ratio of ions in a formula unit depends on the charges of the ions combined. For example, to achieve electrical neutrality in the ionic compound calcium fluoride, two fluoride anions, F^-, each with a charge of 1−, must balance the 2+ charge of each calcium cation, Ca^{2+}. Therefore, the formula of calcium fluoride is CaF_2.

FIGURE 12 Like most ionic compounds, sodium chloride is a crystalline solid.
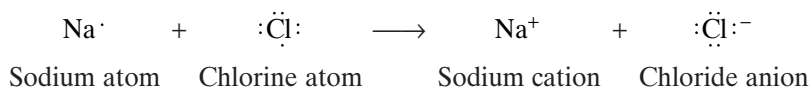
Formation of Ionic Compounds

Electron-dot notation can be used to demonstrate the changes that take place in ionic bonding. Ionic compounds do not ordinarily form by the combination of isolated ions, but consider for a moment a sodium
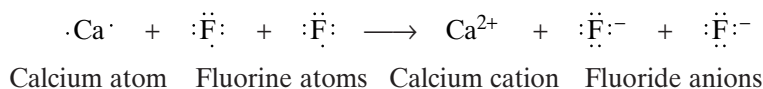
atom and a chlorine atom approaching each other. The two atoms are neutral and have one and seven valence electrons, respectively.

$$\text{Na}\cdot \qquad\qquad\qquad :\ddot{\text{Cl}}:$$

Sodium atom   Chlorine atom

We have already seen that atoms of sodium and the other alkali metals readily lose one electron to form cations. And we have seen that atoms of chlorine and the other halogens readily gain one electron to form anions. The combination of sodium and chlorine atoms to produce one formula unit of sodium chloride can thus be represented as follows.

$$\text{Na}\cdot \quad + \quad :\ddot{\text{Cl}}: \quad \longrightarrow \quad \text{Na}^+ \quad + \quad :\ddot{\text{Cl}}:^-$$

Sodium atom   Chlorine atom   Sodium cation   Chloride anion

The transfer of an electron from the sodium atom to the chlorine atom transforms each atom into an ion with a noble-gas configuration. In the combination of calcium with fluorine, two fluorine atoms are needed to accept the two valence electrons given up by one calcium atom.

$$\cdot\text{Ca}\cdot \quad + \quad :\ddot{\text{F}}: \quad + \quad :\ddot{\text{F}}: \quad \longrightarrow \quad \text{Ca}^{2+} \quad + \quad :\ddot{\text{F}}:^- \quad + \quad :\ddot{\text{F}}:^-$$

Calcium atom   Fluorine atoms   Calcium cation   Fluoride anions

Characteristics of Ionic Bonding

Recall that nature favors arrangements in which potential energy is minimized. In an ionic crystal, ions minimize their potential energy by combining in an orderly arrangement known as a *crystal lattice* (see **Figure 13**). The attractive forces at work within an ionic crystal include those between oppositely charged ions and those between the nuclei and electrons of adjacent ions. The repulsive forces include those between like-charged ions and those between electrons of adjacent ions. The distances between ions and their arrangement in a crystal represent a balance among all these forces. Sodium chloride's crystal structure is shown in **Figure 14** below.
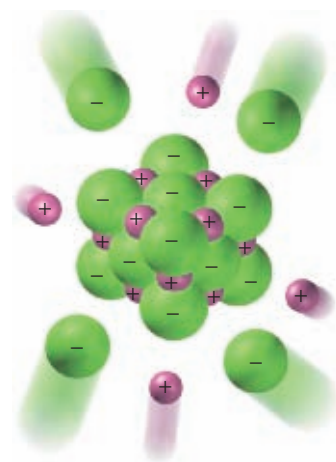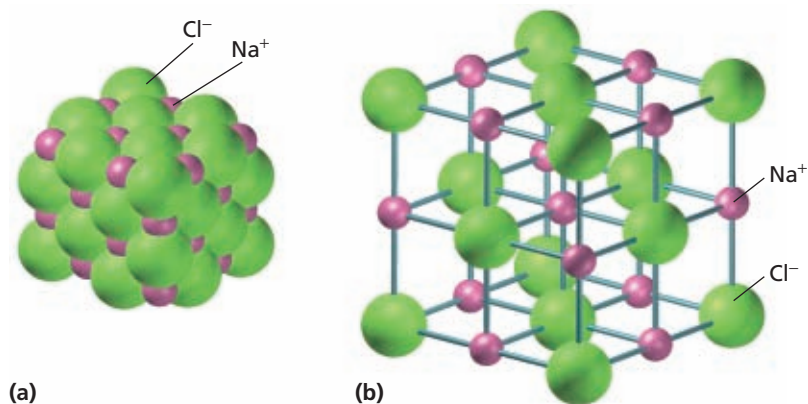
FIGURE 13 The ions in an ionic compound lower their potential energy by forming an orderly, three-dimensional array in which the positive and negative charges are balanced. The electrical forces of attraction between oppositely charged ions extend over long distances, causing a large decrease in potential energy.

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Code: HC60815

(a)   (b)

FIGURE 14 Two models of the crystal structure of sodium chloride are shown. (a) To illustrate the ions' actual arrangement, the sodium and chloride ions are shown with their electron clouds just touching. (b) In an expanded view, the distances between ions have been exaggerated in order to clarify the positioning of the ions in the structure.

FIGURE 15 The figure shows the ions that most closely surround a chloride anion and a sodium cation within the crystal structure of NaCl. The structure is composed such that (a) six Na^+ ions surround each Cl^- ion. At the same time, (b) six Cl^- ions surround each Na^+ ion (which cannot be seen but whose location is indicated by the dashed outline).
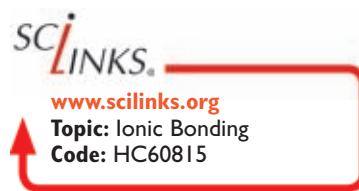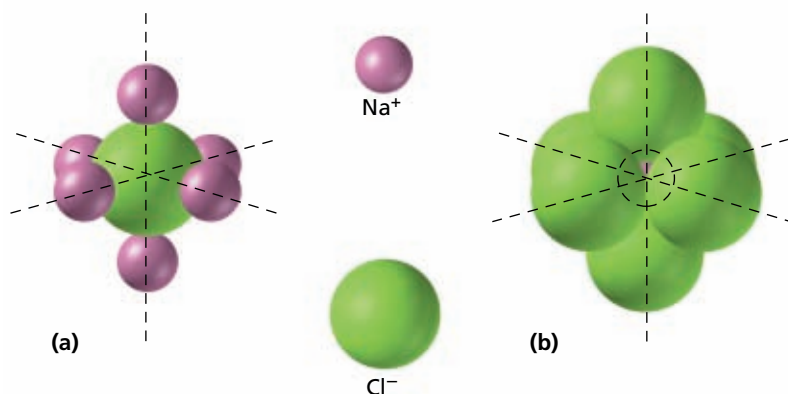
Figure 15 shows the crystal structure of sodium chloride in greater detail. Within the arrangement, each sodium cation is surrounded by six chloride anions. At the same time, each chloride anion is surrounded by six sodium cations. Attraction between the adjacent oppositely charged ions is much stronger than repulsion by other ions of the same charge, which are farther away.

The three-dimensional arrangements of ions and the strengths of attraction between them vary with the sizes and charges of the ions and the numbers of ions of different charges. For example, in calcium fluoride, there are two anions for each cation. Each calcium cation is surrounded by eight fluoride anions. At the same time, each fluoride ion is surrounded by four calcium cations, as shown in **Figure 16.**

To compare bond strengths in ionic compounds, chemists compare the amounts of energy released when separated ions in a gas come together to form a crystalline solid. **Lattice energy** *is the energy released when one mole of an ionic crystalline compound is formed from gaseous ions.* Lattice energy values for a few common ionic compounds are shown in **Table 3.** The negative energy values indicate that energy is *released* when the crystals are formed.

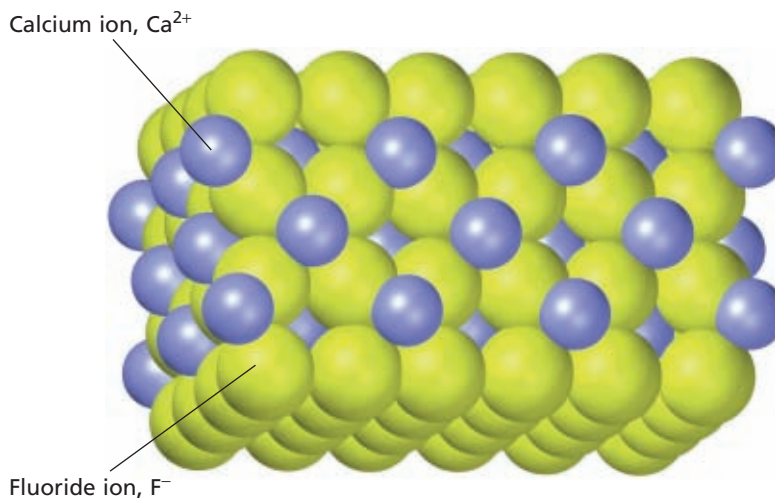Calcium ion, Ca^{2+}

Fluoride ion, F^-

FIGURE 16 In the crystal structure of calcium fluoride, CaF_2, each calcium cation is surrounded by eight fluoride anions and each fluoride ion is surrounded by four calcium cations. This is the closest possible packing of the ions in which the positive and negative charges are balanced.

A Comparison of Ionic and Molecular Compounds

The force that holds ions together in ionic compounds is a very strong over-all attraction between positive and negative charges. In a molecular compound, the covalent bonds of the atoms making up each molecule are also strong. But the forces of attraction *between* molecules are much weaker than the forces among formula units in ionic bonding. This difference in the strength of attraction between the basic units of molecular and ionic compounds gives rise to different properties in the two types of compounds.

The melting point, boiling point, and hardness of a compound depend on how strongly its basic units are attracted to each other. Because the forces of attraction between individual molecules are not very strong, many molecular compounds melt at low temperatures. In fact, many molecular compounds are already completely gaseous at room temperature. In contrast, the ions in ionic compounds are held together by strong attractive forces, so ionic compounds generally have higher melting and boiling points than do molecular compounds.

Ionic compounds are hard but brittle. Why? In an ionic crystal, even a slight shift of one row of ions relative to another causes a large buildup of repulsive forces, as shown in **Figure 17.** These forces make it difficult for one layer to move relative to another, causing ionic compounds to be hard. If one layer is moved, however, the repulsive forces make the layers part completely, causing ionic compounds to be brittle.

In the solid state, the ions cannot move, so the compounds are not electrical conductors. In the molten state, ionic compounds are electrical conductors because the ions can move freely to carry electrical current. Many ionic compounds can dissolve in water. When they dissolve, their ions separate from each other and become surrounded by water molecules. These ions are free to move through the solution, so such solutions are electrical conductors. Other ionic compounds do not dissolve in water, however, because the attractions between the water molecules and the ions cannot overcome the attractions between the ions.
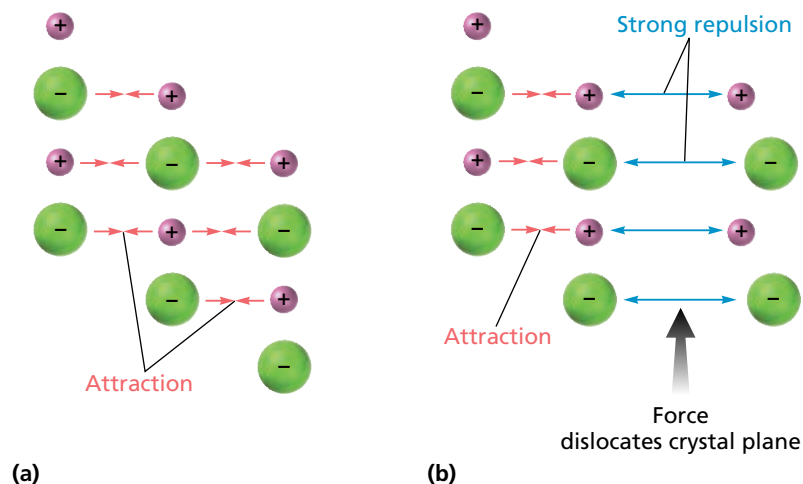
TABLE 3 *Lattice Energies of Some Common Ionic Compounds*

Compound	Lattice energy (kJ/mol)
NaCl	−787.5
NaBr	−751.4
CaF$_2$	−2634.7
LiCl	−861.3
LiF	−1032
MgO	−3760
KCl	−715

FIGURE 17 (a) The attraction between positive and negative ions in a crystalline ionic compound causes layers of ions to resist motion. (b) When struck with sufficient force, the layers shift so that ions of the same charge approach each other, causing repulsion. As a result, the crystal shatters along the planes.

Polyatomic Ions

Certain atoms bond covalently with each other to form a group of atoms that has both molecular and ionic characteristics. *A charged group of covalently bonded atoms is known as a* **polyatomic ion.** Polyatomic ions combine with ions of opposite charge to form ionic compounds. The charge of a polyatomic ion results from an excess of electrons (negative charge) or a shortage of electrons (positive charge). For example, an ammonium ion, a common positively charged polyatomic ion, contains one nitrogen atom and four hydrogen atoms and has a single positive charge. Its formula is NH_4^+, sometimes written as $[NH_4]^+$ to show that the group of atoms *as a whole* has a charge of 1+. The seven protons in the nitrogen atom plus the four protons in the four hydrogen atoms give the ammonium ion a total positive charge of 11+. An independent nitrogen atom has seven electrons, and four independent hydrogen atoms have a total of four electrons. When these atoms combine to form an ammonium ion, one of their electrons is lost, giving the polyatomic ion a total negative charge of 10−.

Lewis structures for the ammonium ion and some common negative polyatomic ions—the nitrate, sulfate, and phosphate ions—are shown below. To find the Lewis structure for a polyatomic ion, follow the steps of Sample Problem D, with the following exception. If the ion is negatively charged, add to the total number of valence electrons a number of electrons corresponding to the ion's negative charge. If the ion is positively charged, subtract from the total number of valence electrons a number of electrons corresponding to the ion's positive charge.

Ammonium ion Nitrate ion Sulfate ion Phosphate ion
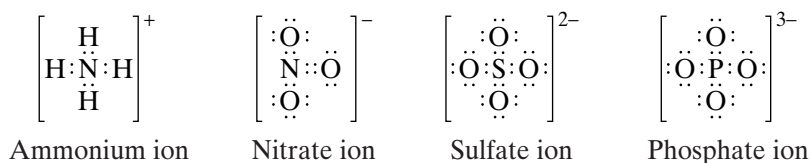
SECTION REVIEW

1. Give two examples of an ionic compound.

2. Use electron-dot notation to demonstrate the formation of ionic compounds involving the following:
   a. Li and Cl
   b. Ca and I

3. Distinguish between ionic and molecular compounds in terms of the basic units that each is composed of.

4. Compound B has lower melting and boiling points than compound A. At the same temperature, compound B vaporizes faster than compound A. If one of these compounds is ionic and the other is molecular, which would you expect to be molecular? ionic? Explain your reasoning.

Critical Thinking

5. **ANALYZING DATA** The melting points for the compounds Li_2S, Rb_2S, and K_2S are 900°C, 530°C, and 840°C, respectively. List these three compounds in order of increasing lattice energy.

Metallic Bonding

OBJECTIVES

- Describe the electron-sea model of metallic bonding, and explain why metals are good electrical conductors.

- Explain why metal surfaces are shiny.

- Explain why metals are malleable and ductile but ionic-crystalline compounds are not.

Chemical bonding is different in metals than it is in ionic, molecular, or covalent-network compounds. This difference is reflected in the unique properties of metals. They are excellent electrical conductors in the solid state—much better conductors than even molten ionic compounds. This property is due to the highly mobile valence electrons of the atoms that make up a metal. Such mobility is not possible in molecular compounds, in which valence electrons are localized in electron-pair bonds between neutral atoms. Nor is it possible in solid ionic compounds, in which electrons are bound to individual ions that are held in place in crystal structures.

The Metallic-Bond Model

The highest energy levels of most metal atoms are occupied by very few electrons. In *s*-block metals, for example, one or two valence electrons occupy the outermost orbital, and all three outermost *p* orbitals, which can hold a total of six electrons, are vacant. In addition to completely vacant outer *p* orbitals, *d*-block metals also possess many vacant *d* orbitals in the energy level just below their highest energy level.

Within a metal, the vacant orbitals in the atoms' outer energy levels overlap. This overlapping of orbitals allows the outer electrons of the atoms to roam freely throughout the entire metal. The electrons are *delocalized,* which means that they do not belong to any one atom but move freely about the metal's network of empty atomic orbitals. These mobile electrons form a *sea of electrons* around the metal atoms, which are packed together in a crystal lattice (see **Figure 18**). *The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons is called* **metallic bonding.**

Metallic Properties

The freedom of electrons to move in a network of metal atoms accounts for the high electrical and thermal conductivity characteristic of all metals. In addition, metals are both strong absorbers and reflectors of light. Because they contain many orbitals separated by extremely small energy differences, metals can absorb a wide range of light frequencies. This absorption of light results in the excitation of the metal atoms' electrons to higher energy levels. However, in metals the electrons immediately fall back down to lower levels, emitting energy in the form of light at a frequency similar to the absorbed frequency. This re-radiated (or reflected) light is responsible for the metallic appearance or luster of metal surfaces.

SCILINKS.

www.scilinks.org
Topic: Metallic Bonding
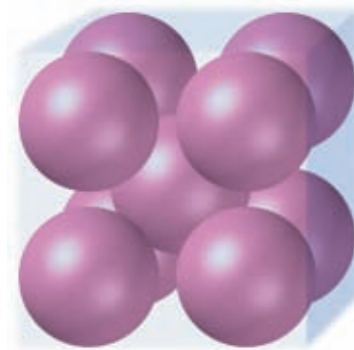**Code:** HC60944

FIGURE 18 The model shows a portion of the crystal structure of solid sodium. The atoms are arranged so that each sodium atom is surrounded by eight other sodium atoms. The atoms are relatively fixed in position, while the electrons are free to move throughout the crystal, forming an electron sea.

Most metals are also easy to form into desired shapes. Two important properties related to this characteristic are malleability and ductility. **Malleability** *is the ability of a substance to be hammered or beaten into thin sheets.* **Ductility** *is the ability of a substance to be drawn, pulled, or extruded through a small opening to produce a wire.* The malleability and ductility of metals are possible because metallic bonding is the same in all directions throughout the solid. When struck, one plane of atoms in a metal can slide past another without encountering resistance or breaking bonds. By contrast, recall from Section 3 that shifting the layers of an ionic crystal causes the bonds to break and the crystal to shatter.

Metallic Bond Strength

Metallic bond strength varies with the nuclear charge of the metal atoms and the number of electrons in the metal's electron sea. Both of these factors are reflected in a metal's *enthalpy of vaporization*. The amount of energy as heat required to vaporize the metal is a measure of the strength of the bonds that hold the metal together. The enthalpy of vaporization is defined as the amount of energy absorbed as heat when a specified amount of a substance vaporizes at constant pressure. Some enthalpies of vaporization for metals are given in **Table 4.**

FIGURE 19 Unlike ionic crystalline compounds, most metals are malleable. This property allows iron, for example, to be shaped into useful tools.

TABLE 4 *Enthalpies of Vaporization of Some Metals (kJ/mol)*

Period	Element		
Second	Li 147	Be 297	
Third	Na 97	Mg 128	Al 294
Fourth	K 77	Ca 155	Sc 333
Fifth	Rb 76	Sr 137	Y 365
Sixth	Cs 64	Ba 140	La 402

SECTION REVIEW

1. Describe the electron-sea model of metallic bonding.

2. What is the relationship between metallic bond strength and enthalpy of vaporization?

3. Explain why most metals are malleable and ductile but ionic crystals are not.

Critical Thinking

4. **ORGANIZING IDEAS** Explain why metals are good electrical conductors.

Molecular Geometry

OBJECTIVES

- Explain VSEPR theory.

- Predict the shapes of molecules or polyatomic ions using VSEPR theory.

- Explain how the shapes of molecules are accounted for by hybridization theory.

- Describe dipole-dipole forces, hydrogen bonding, induced dipoles, and London dispersion forces and their effects on properties such as boiling and melting points.

- Explain what determines molecular polarity.

The properties of molecules depend not only on the bonding of atoms but also on molecular geometry—the three-dimensional arrangement of a molecule's atoms in space. The polarity of each bond, along with the geometry of the molecule, determines *molecular polarity,* or the uneven distribution of molecular charge. As you will read, molecular polarity strongly influences the forces that act *between* molecules in liquids and solids.

A chemical formula reveals little information about a molecule's geometry. After performing many tests designed to reveal the shapes of various molecules, chemists developed two different, equally successful theories to explain certain aspects of their findings. One theory accounts for molecular bond angles. The other is used to describe the orbitals that contain the valence electrons of a molecule's atoms.

VSEPR Theory

As shown in **Figure 20,** diatomic molecules, like those of hydrogen, H_2, and hydrogen chloride, HCl, must be linear because they consist of only two atoms. To predict the geometries of more-complicated molecules, one must consider the locations of all electron pairs surrounding the bonded atoms. This is the basis of VSEPR theory.

The abbreviation VSEPR stands for "valence-shell, electron-pair repulsion," referring to the repulsion between pairs of valence electrons of the atoms in a molecule. **VSEPR theory** *states that repulsion between the sets of valence-level electrons surrounding an atom causes these sets to be oriented as far apart as possible.* How does the assumption that electrons in molecules repel each other account for molecular shapes? For now let us consider only molecules with no unshared valence electron pairs on the central atom.

Let's examine the simple molecule BeF_2. The beryllium atom forms a covalent bond with each fluorine atom and does not follow the octet rule. It is surrounded by only the two electron pairs that it shares with the fluorine atoms.

$$\ddot{\ddot{F}}\!:\!Be\!:\!\ddot{\ddot{F}}\!:$$
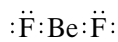
According to VSEPR theory, the shared pairs will be as far away from each other as possible. As shown in **Figure 21a** on the next page, the distance between electron pairs is maximized if the bonds to fluorine are

(a) Hydrogen, H_2

(b) Hydrogen chloride, HCl

FIGURE 20 Ball-and-stick models illustrate the linearity of diatomic molecules. (a) A hydrogen molecule is represented by two identical balls (the hydrogen atoms) joined by a solid bar (the covalent bond). (b) A hydrogen chloride molecule is composed of dissimilar atoms, but it is still linear.
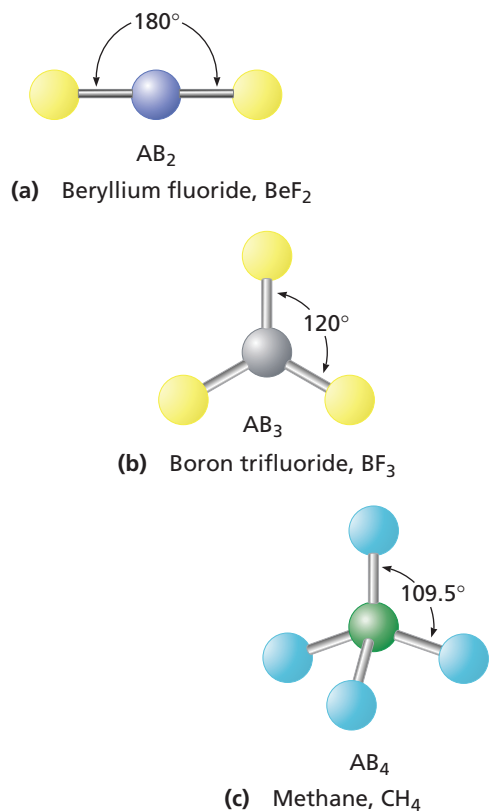
180°

AB₂

(a) Beryllium fluoride, BeF₂

120°

AB₃

(b) Boron trifluoride, BF₃

109.5°

AB₄

(c) Methane, CH₄

FIGURE 21 Ball-and-stick models show the shapes of (a) AB₂, (b) AB₃, and (c) AB₄ molecules according to VSEPR theory.

on opposite sides of the beryllium atom, 180° apart. Thus, all three atoms lie on a straight line. The molecule is linear.

If we represent the central atom in a molecule by the letter A and we represent the atoms bonded to the central atom by the letter B, then according to VSEPR theory, BeF_2 is an example of an AB_2 molecule, which is linear. Can you determine what an AB_3 molecule looks like? The three A—B bonds stay farthest apart by pointing to the corners of an equilateral triangle, giving 120° angles between the bonds. This trigonal-planar geometry is shown in **Figure 21b** for the AB_3 molecule boron trifluoride, BF_3.

The central atoms in AB_4 molecules follow the octet rule by sharing four electron pairs with B atoms. The distance between electron pairs is maximized if each A—B bond points to one of four corners of a tetrahedron. This geometry is shown in **Figure 21c** for the AB_4 molecule methane, CH_4. The same figure shows that in a tetrahedral molecule, each of the bond angles formed by the A atom and any two of the B atoms is equal to 109.5°.

The shapes of various molecules are summarized in **Table 5.** B can represent a single type of atom, a group of identical atoms, or a group of different atoms on the same molecule. The shape of the molecule will still be based on the forms given in the table. However, different sizes of B groups distort the bond angles, making some bond angles larger or smaller than those given in the table.

SAMPLE PROBLEM E

Use VSEPR theory to predict the molecular geometry of boron trichloride, BCl₃.

SOLUTION First write the Lewis structure for BCl_3. Boron is in Group 13 and has three valence electrons.

$$\cdot \dot{B} \cdot$$

Chlorine is in Group 17, so each chlorine atom has seven valence electrons.

$$:\ddot{\underset{..}{Cl}}:$$

The total number of available valence electrons is therefore $24e^-$ ($3e^-$ from boron and $21e^-$ from chlorine). The following Lewis structure uses all $24e^-$.

$$\begin{array}{c} :\ddot{\underset{..}{Cl}}: \\ :\ddot{\underset{..}{Cl}}:B:\ddot{\underset{..}{Cl}}: \end{array}$$
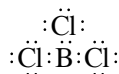
This molecule is an exception to the octet rule because in this case B forms only three bonds. Boron trichloride is an AB_3 type of molecule. Therefore, according to VSEPR theory, it should have trigonal-planar geometry.

1. Use VSEPR theory to predict the molecular geometry of the following molecules:

   a. HI   **b.** CBr_4 **c.** CH_2Cl_2

VSEPR Theory and Unshared Electron Pairs

Ammonia, NH_3, and water, H_2O, are examples of molecules in which the central atom has both shared and unshared electron pairs (see **Table 5** on the next page for their Lewis structures). How does VSEPR theory account for the geometries of these molecules?

The Lewis structure of ammonia shows that in addition to the three electron pairs it shares with the three hydrogen atoms, the central nitrogen atom has one unshared pair of electrons.

$$H\!:\!\overset{\cdot\cdot}{N}\!:\!H$$
$$\overset{|}{H}$$

VSEPR theory postulates that the lone pair occupies space around the nitrogen atom just as the bonding pairs do. Thus, as in an AB_4 molecule, the electron pairs maximize their separation by assuming the four corners of a tetrahedron. Lone pairs do occupy space, but our description of the observed shape of a molecule refers to the *positions of atoms only*. Consequently, as shown in **Figure 22a,** the molecular geometry of an ammonia molecule is that of a pyramid with a triangular base. The general VSEPR formula for molecules such as ammonia is AB_3E, where E represents the unshared electron pair.

A water molecule has two unshared electron pairs. It is an AB_2E_2 molecule. Here, the oxygen atom is at the center of a tetrahedron, with two corners occupied by hydrogen atoms and two by the unshared pairs **(Figure 22b)**. Again, VSEPR theory states that the lone pairs occupy space around the central atom but that the actual shape of the molecule is determined by the positions of the atoms only. In the case of water, this results in a "bent," or angular, molecule.
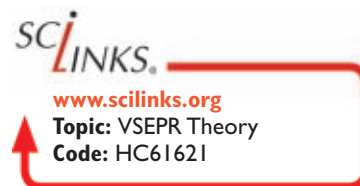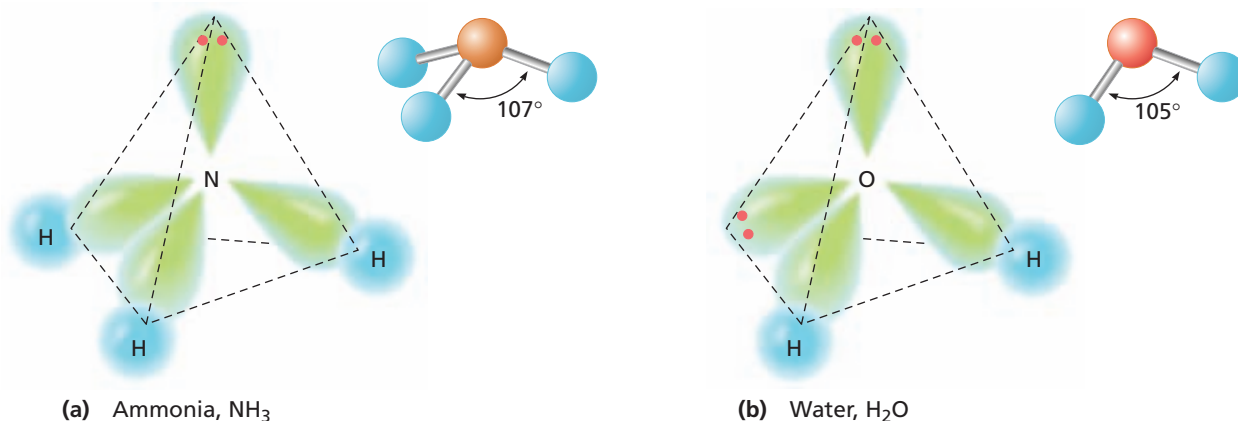
FIGURE 22 The locations of bonds and unshared electrons are shown for molecules of (a) ammonia and (b) water. Although unshared electrons occupy space around the central atoms, the shapes of the molecules depend only on the position of the molecules' atoms, as clearly shown by the ball-and-stick models.
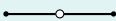

(a) Ammonia, NH_3

(b) Water, H_2O
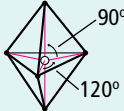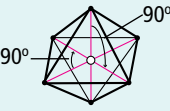
In **Figure 22b,** note that the bond angles in ammonia and water are somewhat less than the 109.5° bond angles of a perfectly tetrahedral molecule. These angles are smaller because the unshared electron pairs repel electrons more strongly than do bonding electron pairs.

Table 5 also includes an example of an AB_2E type molecule. This type of molecule results when a central atom forms two bonds and retains one unshared electron pair.

Finally, in VSEPR theory, double and triple bonds are treated in the same way as single bonds. And polyatomic ions are treated similarly to molecules. (Remember to consider *all* of the electron pairs present in any ion or molecule.) Thus, Lewis structures and **Table 5** can be used together to predict the shapes of polyatomic ions as well as molecules with double or triple bonds.

TABLE 5   *VSEPR Theory and Molecular Geometry*

	Molecular shape	Atoms bonded to central atom	Lone pairs of electrons	Type of molecule	Formula example	Lewis structure
Linear		2	0	AB_2	BeF_2	
Trigonal-planar		3	0	AB_3	BF_3	
Bent or Angular		2	1	AB_2E	ONF	
Tetrahedral		4	0	AB_4	CH_4	
Trigonal-pyramidal		3	1	AB_3E	NH_3	
Bent or Angular		2	2	AB_2E_2	H_2O	
Trigonal-bipyramidal		5	0	AB_5	PCl_5	
Octahedral		6	0	AB_6	SF_6	

a. Use VSEPR theory to predict the shape of a molecule of carbon dioxide, CO_2.

b. Use VSEPR theory to predict the shape of a chlorate ion, ClO_3^-.

SOLUTION **a.** The Lewis structure of carbon dioxide shows two carbon-oxygen double bonds and no unshared electron pairs on the carbon atom. To simplify the molecule's Lewis structure, we represent the covalent bonds with lines instead of dots.

$$\ddot{O}=C=\ddot{O}$$
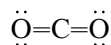
This is an AB_2 molecule, which is linear.

b. The Lewis structure of a chlorate ion shows three oxygen atoms and an unshared pair of electrons surrounding a central chlorine atom. Again, lines are used to represent the covalent bonds.

$$\left[\begin{array}{c} \ddot{Cl} \\ \ddot{O}\ :\!\ddot{O}\!:\ \ddot{O}\!: \end{array} \right]^{-}$$
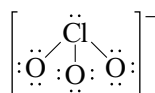
The chlorate ion is an AB_3E type. It has trigonal-pyramidal geometry, with the three oxygen atoms at the base of the pyramid and the chlorine atom at the top.

PRACTICE *Answers in Appendix E*

1. Use VSEPR theory to predict the molecular geometries of the molecules whose Lewis structures are given below.
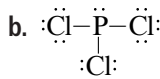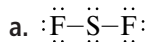
a. $:\!\ddot{F}\!-\!\ddot{S}\!-\!\ddot{F}\!:$ **b.** $:\!\ddot{Cl}\!-\!P\!-\!\ddot{Cl}\!:$
$:\!\ddot{Cl}\!:$

Hybridization

VSEPR theory is useful for explaining the shapes of molecules. However, it does not reveal the relationship between a molecule's geometry and the orbitals occupied by its bonding electrons. To explain how the orbitals of an atom become rearranged when the atom forms covalent bonds, a different model is used. This model is called **hybridization,** *which is the mixing of two or more atomic orbitals of similar energies on the same atom to produce new hybrid atomic orbitals of equal energies.*

Methane, CH_4, provides a good example of how hybridization is used to explain the geometry of molecular orbitals. The orbital notation for a carbon atom shows that it has four valence electrons, two in the $2s$ orbital and two in $2p$ orbitals.

$$C \quad \underset{1s}{\uparrow\downarrow} \ \underset{2s}{\uparrow\downarrow} \ \underset{\underbrace{\qquad\qquad}_{2p}}{\uparrow \ \ \uparrow \ \ \_\_}$$
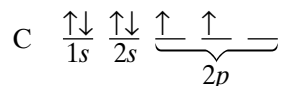
We know from experiments that a methane molecule has tetrahedral geometry. How does carbon form four equivalent, tetrahedrally arranged covalent bonds by orbital overlap with four other atoms?

Two of carbon's valence electrons occupy the 2s orbital, and two occupy the 2p orbitals. Recall that the 2s orbital and the 2p orbitals have different shapes. To achieve four equivalent bonds, carbon's 2s and three 2p orbitals *hybridize* to form four new, identical orbitals called sp^3 orbitals. The superscript 3 indicates that three p orbitals were included in the hybridization; the superscript 1 on the s is understood. The sp^3 orbitals all have the same energy, which is greater than that of the 2s orbital but less than that of the 2p orbitals, as shown in **Figure 23.**

FIGURE 23 The sp^3 hybridization of carbon's outer orbitals combines one s and three p orbitals to form four sp^3 hybrid orbitals. Whenever hybridization occurs, the resulting hybrid orbitals are at an energy level between the levels of the orbitals that have combined.

Carbon's orbitals before hybridization

Carbon's orbitals after sp^3 hybridization
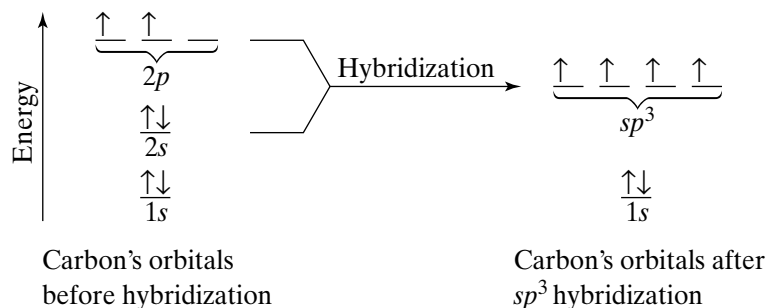
Hybrid orbitals *are orbitals of equal energy produced by the combination of two or more orbitals on the same atom.* The number of hybrid orbitals produced equals the number of orbitals that have combined. Bonding with carbon sp^3 orbitals is illustrated in **Figure 24a** for a molecule of methane.

Hybridization also explains the bonding and geometry of many molecules formed by Group 15 and 16 elements. The sp^3 hybridization of a nitrogen atom ($[He]2s^22p^3$) yields four hybrid orbitals—one orbital containing a pair of electrons and three orbitals that each contain an unpaired electron. Each unpaired electron is capable of forming a single bond, as shown for ammonia in **Figure 24b.** Similarly, two of the four sp^3 hybrid orbitals on an oxygen atom ($[He]2s^22p^4$) are occupied by two electron pairs and two are occupied by unpaired electrons. Each unpaired electron can form a single bond, as shown for water in **Figure 24c.**

FIGURE 24 Bonds formed by the overlap of the 1s orbitals of hydrogen atoms and the sp^3 orbitals of (a) carbon, (b) nitrogen, and (c) oxygen. For the sake of clarity, only the hybrid orbitals of the central atoms are shown.
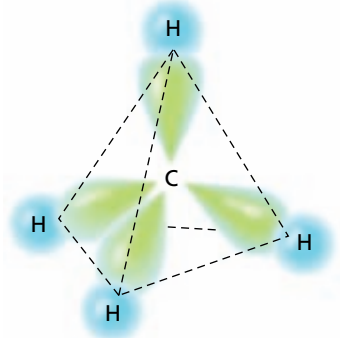
(a) Methane, CH_4
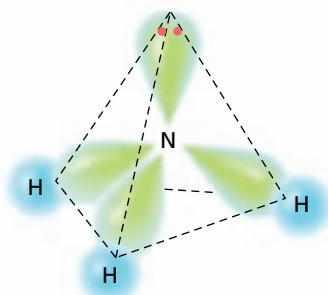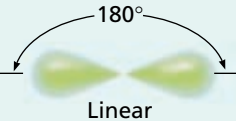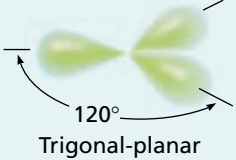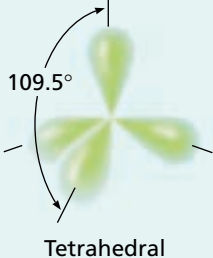
(b) Ammonia, NH_3
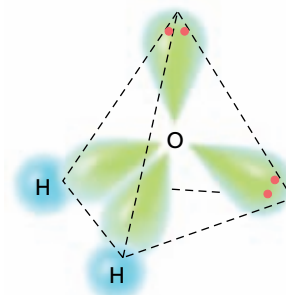
(c) Water, H_2O

TABLE 6 Geometry of Hybrid Orbitals

Atomic orbitals	Type of hybridization	Number of hybrid orbitals	Geometry
s, p	sp	2	180° — Linear
s, p, p	sp²	3	120° — Trigonal-planar
s, p, p, p	sp³	4	109.5° — Tetrahedral

The linear geometry of molecules such as beryllium fluoride, BeF_2, (see **Table 5**) is made possible by hybridization involving the *s* orbital and one available empty *p* orbital to yield *sp* hybrid orbitals. The trigonal-planar geometry of molecules such as boron fluoride, BF_3, is made possible by hybridization involving the *s* orbital, one singly occupied *p* orbital, and one empty *p* orbital to yield sp^2 hybrid orbitals. The geometries of sp, sp^2, and sp^3 hybrid orbitals are summarized in **Table 6.**

Intermolecular Forces

As a liquid is heated, the kinetic energy of its particles increases. At the boiling point, the energy is sufficient to overcome the force of attraction between the liquid's particles. The particles pull away from each other and enter the gas phase. Boiling point is therefore a good measure of the force of attraction between particles of a liquid. The higher the boiling point, the stronger the forces between particles.

The forces of attraction between molecules are known as *intermolecular forces*. Intermolecular forces vary in strength but are generally weaker than bonds that join atoms in molecules, ions in ionic compounds, or metal atoms in solid metals. Compare the boiling points of the metals and ionic compounds in **Table 7** on the next page with those of the molecular substances listed. Note that the values for ionic compounds and metals are much higher than those for molecular substances.

TABLE 7 Boiling Points and Bonding Types

Bonding type	Substance	bp (1 atm, °C)
Nonpolar-covalent (molecular)	H_2	−253
	O_2	−183
	Cl_2	−34
	Br_2	59
	CH_4	−164
	CCl_4	77
	C_6H_6	80
Polar-covalent (molecular)	PH_3	−88
	NH_3	−33
	H_2S	−61
	H_2O	100
	HF	20
	HCl	−85
	ICl	97
Ionic	NaCl	1413
	MgF_2	2239
Metallic	Cu	2567
	Fe	2750
	W	5660

Molecular Polarity and Dipole-Dipole Forces

The strongest intermolecular forces exist between polar molecules. Polar molecules act as tiny dipoles because of their uneven charge distribution. *A* **dipole** *is created by equal but opposite charges that are separated by a short distance.* The direction of a dipole is from the dipole's positive pole to its negative pole. A dipole is represented by an arrow with a head pointing toward the negative pole and a crossed tail situated at the positive pole. The dipole created by a hydrogen chloride molecule, which has its negative end at the more electronegative chlorine atom, is indicated as follows.

$$\overset{+\longrightarrow}{H-Cl}$$

The negative region in one polar molecule attracts the positive region in adjacent molecules, and so on throughout a liquid or solid. The forces of attraction between polar molecules are known as *dipole-dipole forces*. These forces are short-range forces, acting only between nearby molecules. The effect of dipole-dipole forces is reflected, for example, by the significant difference between the boiling points of iodine chloride, I−Cl, and bromine, Br−Br. The boiling point of polar iodine chloride is 97°C, whereas that of nonpolar bromine is only 59°C. The dipole-dipole forces responsible for the relatively high boiling point of ICl are illustrated schematically in **Figure 25.**

The polarity of diatomic molecules such as ICl is determined by just one bond. For molecules containing more than two atoms, molecular polarity depends on both the polarity and the orientation of each bond. A molecule of water, for example, has two hydrogen-oxygen bonds in which the more-electronegative oxygen atom is the negative pole of each bond. Because the molecule is bent, the polarities of these two bonds combine to make the molecule highly polar, as shown in **Figure 26.** An ammonia molecule is also highly polar because the dipoles of the three nitrogen-hydrogen bonds are additive, combining to create a net molecular dipole. In some molecules, individual bond dipoles cancel one another, causing the resulting molecular polarity to be zero. Carbon dioxide and carbon tetrachloride are molecules of this type.

A polar molecule can *induce* a dipole in a nonpolar molecule by temporarily attracting its electrons. The result is a short-range intermolecular force that is somewhat weaker than the dipole-dipole force. The force of an induced dipole accounts for the solubility of nonpolar O_2 in water. The positive pole of a water molecule attracts the outer electrons
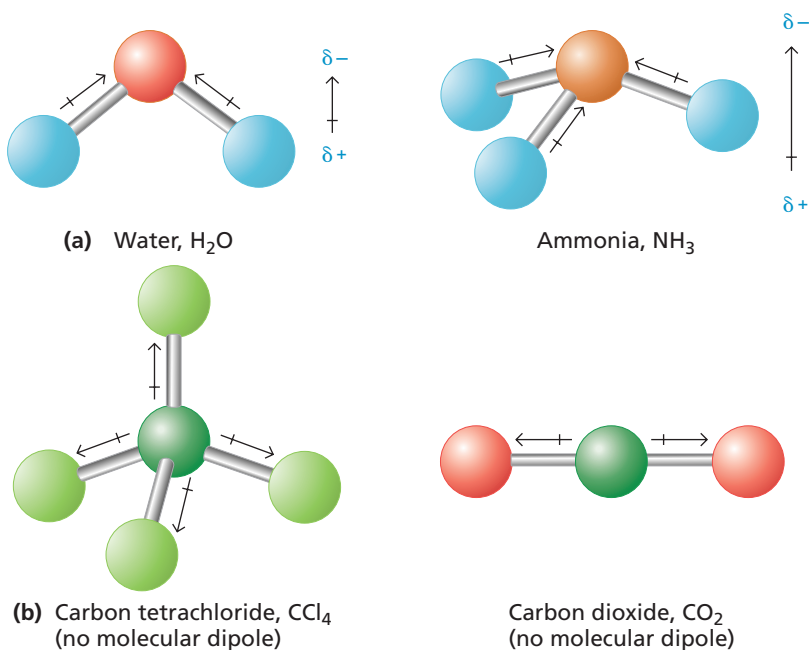
(a) Water, H_2O

Ammonia, NH_3

(b) Carbon tetrachloride, CCl_4 (no molecular dipole)

Carbon dioxide, CO_2 (no molecular dipole)

Water, H₂O   Oxygen, O₂
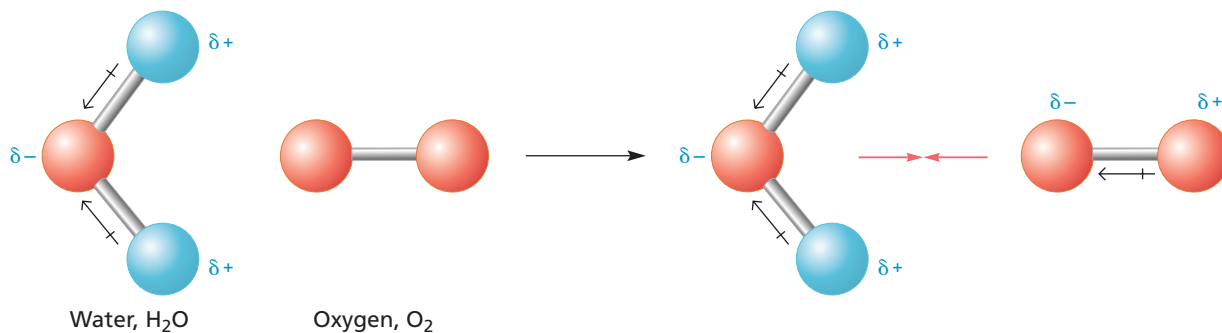
FIGURE 27 Dipole-induced dipole interaction. The positive pole of a water molecule causes a temporary change in the electron distribution of an oxygen molecule. The negative pole induced in the oxygen molecule is then attracted to the positive pole of the water molecule.

of an adjacent oxygen molecule. The oxygen molecule, then, has an induced negative pole on the side toward the water molecule and an induced positive pole on the opposite side. The result is an attraction to the water molecule, as shown in **Figure 27.**

Hydrogen Bonding

Some hydrogen-containing compounds, such as hydrogen fluoride (HF), water (H_2O), and ammonia (NH_3), have unusually high boiling points. This is explained by the presence of a particularly strong type of dipole-dipole force. In compounds containing H−F, H−O, or H−N bonds, the large electronegativity differences between hydrogen atoms and fluorine, oxygen, or nitrogen atoms make the bonds connecting them highly polar. This gives the hydrogen atom a positive charge that is almost half as large as that of a proton. Moreover, the small size of the hydrogen atom allows the atom to come very close to an unshared pair of electrons on an adjacent molecule. *The intermolecular force in which a hydrogen atom that is bonded to a highly electronegative atom is attracted to an unshared pair of electrons of an electronegative atom in a nearby molecule is known as* **hydrogen bonding.**

Hydrogen bonds are usually represented by dotted lines connecting the hydrogen-bonded hydrogen to the unshared electron pair of the electronegative atom to which it is attracted, as illustrated for water in **Figure 28.** The effect of hydrogen bonding can be seen by comparing the boiling points in **Table 7.** Look at phosphine, PH_3, compared with hydrogen-bonded ammonia, NH_3. How does hydrogen sulfide, H_2S, compare with hydrogen-bonded water, H_2O?

FIGURE 28 Space-filling models illustrate hydrogen bonding between water molecules. The dotted lines indicate the attraction between electronegative oxygen atoms and electropositive hydrogen atoms of neighboring molecules.
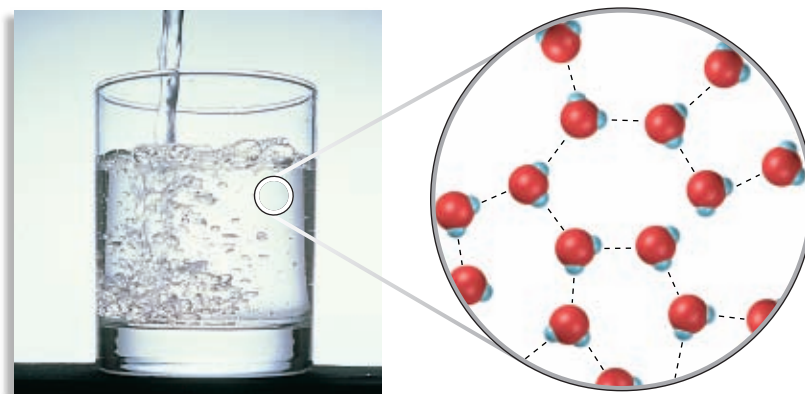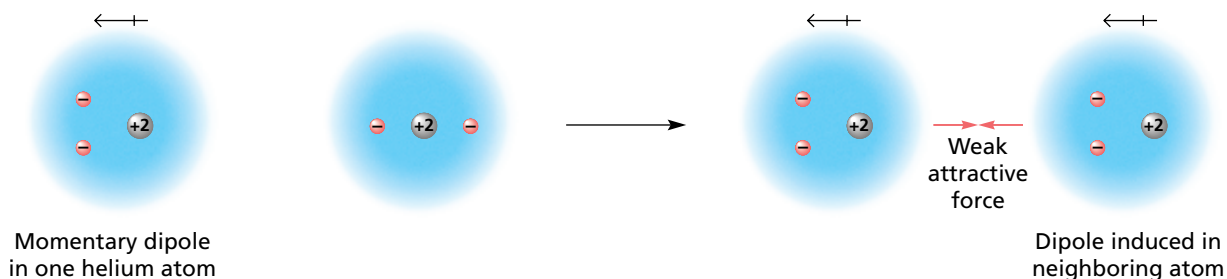
Momentary dipole
in one helium atom

Weak
attractive
force

Dipole induced in
neighboring atom

London Dispersion Forces

Even noble-gas atoms and molecules that are nonpolar experience a weak intermolecular attraction. In any atom or molecule—polar or nonpolar—the electrons are in continuous motion. As a result, at any instant the electron distribution may be slightly uneven. The momentary, uneven charge creates a positive pole in one part of the atom or molecule and a negative pole in another. This temporary dipole can then induce a dipole in an adjacent atom or molecule. The two are held together for an instant by the weak attraction between the temporary dipoles, as illustrated in **Figure 29.** *The intermolecular attractions resulting from the constant motion of electrons and the creation of instantaneous dipoles are called* **London dispersion forces,** after Fritz London, who first proposed their existence in 1930.

London forces act between all atoms and molecules. But they are the *only* intermolecular forces acting among noble-gas atoms and nonpolar molecules. This fact is reflected in the low boiling points of the noble gases and nonpolar molecular compounds listed in **Table 7.** Because London forces are dependent on the motion of electrons, their strength increases with the number of electrons in the interacting atoms or molecules. In other words, London forces increase with increasing atomic or molar mass. This trend can be seen by comparing the boiling points of the gases helium, He, and argon, Ar; hydrogen, H_2, and oxygen, O_2; and chlorine, Cl_2, and bromine, Br_2.

SECTION REVIEW

1. What two theories can be used to predict molecular geometry?

2. Draw the Lewis structure and predict the molecular geometry of the following molecules:

   a. SO_2 **b.** CI_4 **c.** BCl_3

3. What factors affect the geometry of a molecule?

4. Explain what is meant by sp^3 hybridization.

5. What type of intermolecular force contributes to the high boiling point of water? Explain.

Critical Thinking

6. **INFERRING RELATIONSHIPS** What experimental property directly correlates with the strength of the intermolecular forces? Briefly explain your answer.

Introduction to Chemical Bonding

Vocabulary

chemical bond
ionic bonding
covalent bonding
nonpolar-covalent bond
polar
polar-covalent bond

- Most atoms are chemically bonded to other atoms. The three major types of chemical bonding are ionic, covalent, and metallic.
- In general, atoms of metals bond ionically with atoms of nonmetals, atoms of metals bond metallically with each other, and atoms of nonmetals bond covalently with each other.

Covalent Bonding and Molecular Compounds

Vocabulary

molecule	Lewis structure
molecular compound	structural formula
chemical formula	single bond
molecular formula	multiple bond
bond energy	resonance
electron-dot notation	

- Atoms in molecules are joined by covalent bonds. In a covalent bond, two atoms share one or more pairs of electrons.
- The octet rule states that many chemical compounds tend to form bonds so that each atom shares or has eight electrons in its highest occupied energy level.
- Bonding within many molecules and ions can be indicated by a Lewis structure. Molecules or ions that cannot be correctly represented by a single Lewis structure are represented by resonance structures.

Ionic Bonding and Ionic Compounds

Vocabulary

ionic compound
formula unit
lattice energy
polyatomic ion

- An ionic compound is a three-dimensional network of positive and negative ions mutually attracted to one another.
- Ionic compounds tend to be harder and more brittle and to have higher boiling points than materials containing only covalently bonded atoms.

Metallic Bonding

Vocabulary

metallic bonding
malleability
ductility

- The "electron sea" formed in metallic bonding gives metals their properties of high electrical and thermal conductivity, malleability, ductility, and luster.

Molecular Geometry

Vocabulary

VSEPR theory
hybridization
hybrid orbitals
dipole
hydrogen bonding
London dispersion forces

- VSEPR theory is used to predict the shapes of molecules based on the fact that electron pairs strongly repel each other.
- Hybridization theory is used to predict the shapes of molecules based on the fact that orbitals within an atom can mix to form orbitals of equal energy.
- Intermolecular forces include dipole-dipole forces and London dispersion forces. Hydrogen bonding is a special case of dipole-dipole forces.

Introduction to Chemical Bonding

SECTION 1 REVIEW

1. What is a chemical bond?
2. Identify and define the three major types of chemical bonding.
3. What is the relationship between electronegativity and the ionic character of a chemical bond?
4. a. What is the meaning of the term polar, as applied to chemical bonding?
   b. Distinguish between polar-covalent and nonpolar-covalent bonds.
5. In general, what determines whether atoms will form chemical bonds?

PRACTICE PROBLEMS

6. Determine the electronegativity difference, the probable bond type, and the more-electronegative atom with respect to bonds formed between the following pairs of atoms. (Hint: See Sample Problem A.)
   a. H and I
   b. S and O
   c. K and Br
   d. Si and Cl
   e. K and Cl
   f. Se and S
   g. C and H
7. List the bonding pairs described in item 6 in order of increasing covalent character.
8. Use orbital notation to illustrate the bonding in each of the following molecules:
   a. chlorine, Cl_2
   b. oxygen, O_2
   c. hydrogen fluoride, HF
9. The lattice energy of sodium chloride, NaCl, is −787.5 kJ/mol. The lattice energy of potassium chloride, KCl, is −715 kJ/mol. In which compound is the bonding between ions stronger? Why?

Covalent Bonding and Molecular Compounds

SECTION 2 REVIEW

10. What is a molecule?
11. a. What determines bond length?
    b. In general, how are bond energies and bond lengths related?
12. Describe the general location of the electrons in a covalent bond.
13. As applied to covalent bonding, what is meant by an unshared or lone pair of electrons?
14. Describe the octet rule in terms of noble-gas configurations and potential energy.
15. Determine the number of valence electrons in an atom of each of the following elements:
    a. H
    b. F
    c. Mg
    d. O
    e. Al
    f. N
    g. C
16. When drawing Lewis structures, which atom is usually the central atom?
17. Distinguish between single, double, and triple covalent bonds by defining each and providing an illustration of each type.
18. In writing Lewis structures, how is the need for multiple bonds generally determined?

PRACTICE PROBLEMS

19. Use electron-dot notation to illustrate the number of valence electrons present in one atom of each of the following elements. (Hint: See Sample Problem B.)
    a. Li
    b. Ca
    c. Cl
    d. O
    e. C
    f. P
    g. Al
    h. S

20. Use electron-dot structures to demonstrate the formation of ionic compounds involving the following elements:
    a. Na and S
    b. Ca and O
    c. Al and S

21. Draw Lewis structures for each of the following molecules. (Hint: See Sample Problem D.)
    a. contains one C and four F atoms
    b. contains two H and one Se atom
    c. contains one N and three I atoms
    d. contains one Si and four Br atoms
    e. contains one C, one Cl, and three H atoms

22. Determine the type of hybrid orbitals formed by the boron atom in a molecule of boron fluoride, BF_3.

23. Draw Lewis structures for each of the following molecules. Show resonance structures, if they exist.
    a. O_2
    b. N_2
    c. CO
    d. SO_2

24. Draw Lewis structures for each of the following polyatomic ions. Show resonance structures, if they exist.
    a. OH^-
    b. $H_3C_2O_2^-$
    c. BrO_3^-

Ionic Bonding and Ionic Compounds

SECTION 3 REVIEW

25. a. What is an ionic compound?
    b. In what form do most ionic compounds occur?

26. a. What is a formula unit?
    b. What are the components of one formula unit of CaF_2?

27. a. What is lattice energy?
    b. In general, what is the relationship between lattice energy and the strength of ionic bonding?

28. a. In general, how do ionic and molecular compounds compare in terms of melting points, boiling points, and ease of vaporization?
    b. What accounts for the observed differences in the properties of ionic and molecular compounds?
    c. Cite three physical properties of ionic compounds.

29. a. What is a polyatomic ion?
    b. Give two examples of polyatomic ions.
    c. In what form do such ions often occur in nature?

Metallic Bonding

SECTION 4 REVIEW

30. a. How do the properties of metals differ from those of both ionic and molecular compounds?
    b. What specific property of metals accounts for their unusual electrical conductivity?

31. What properties of metals contribute to their tendency to form metallic bonds?

32. a. What is metallic bonding?
    b. How can the strength of metallic bonding be measured?

Molecular Geometry

SECTION 5 REVIEW

33. a. How is the VSEPR theory used to classify molecules?
    b. What molecular geometry would be expected for F_2 and HF?

34. According to the VSEPR theory, what molecular geometries are associated with the following types of molecules?
    a. AB_2
    b. AB_3
    c. AB_4
    d. AB_5
    e. AB_6

35. Describe the role of each of the following in predicting molecular geometries:
    a. unshared electron pairs
    b. double bonds

36. a. What are hybrid orbitals?
   b. What determines the number of hybrid orbitals produced by the hybridization of an atom?
37. a. What are intermolecular forces?
   b. In general, how do these forces compare in strength with those in ionic and metallic bonding?
   c. What types of molecules have the strongest intermolecular forces?
38. What is the relationship between electronegativity and the polarity of a chemical bond?
39. a. What are dipole-dipole forces?
   b. What determines the polarity of a molecule?
40. a. What is meant by an induced dipole?
   b. What is the everyday importance of this type of intermolecular force?
41. a. What is hydrogen bonding?
   b. What accounts for its extraordinary strength?
42. What are London dispersion forces?

PRACTICE PROBLEMS

43. According to the VSEPR theory, what molecular geometries are associated with the following types of molecules?
   a. AB_3E
   b. AB_2E_2
   c. AB_2E
44. Use hybridization to explain the bonding in methane, CH_4.
45. For each of the following polar molecules, indicate the direction of the resulting dipole:
   a. H$-$F
   b. H$-$Cl
   c. H$-$Br
   d. H$-$I
46. Determine whether each of the following bonds would be polar or nonpolar:
   a. H$-$H
   b. H$-$O
   c. H$-$F
   d. Br$-$Br
   e. H$-$Cl
   f. H$-$N

47. On the basis of individual bond polarity and orientation, determine whether each of the following molecules would be polar or nonpolar:
   a. H_2O
   b. I_2
   c. CF_4
   d. NH_3
   e. CO_2
48. Draw a Lewis structure for each of the following molecules, and then use the VSEPR theory to predict the molecular geometry of each:
   a. SCl_2
   b. PI_3
   c. Cl_2O
   d. NH_2Cl
   e. $SiCl_3Br$
   f. ONCl
49. Draw a Lewis structure for each of the following polyatomic ions, and then use VSEPR theory to determine the geometry of each:
   a. NO_3^-
   b. NH_4^+
   c. SO_4^{2-}
   d. ClO_2^-

MIXED REVIEW

50. Arrange the following pairs from strongest to weakest attraction:
   a. polar molecule and polar molecule
   b. nonpolar molecule and nonpolar molecule
   c. polar molecule and ion
   d. ion and ion
51. Determine the geometry of the following molecules:
   a. CCl_4
   b. $BeCl_2$
   c. PH_3
52. What types of atoms tend to form the following types of bonding?
   a. ionic
   b. covalent
   c. metallic

53. What happens to the energy level and stability of two bonded atoms when they are separated and become individual atoms?

54. Draw the three resonance structures for sulfur trioxide, SO_3.

55. a. How do ionic and covalent bonding differ?
b. How does an ionic compound differ from a molecular compound?
c. How does an ionic compound differ from a metal?

56. Write the electron-dot notation for each of the following elements:
a. He   d. P
b. Cl   e. B
c. O

57. Write the structural formula for methanol, CH_3OH.

58. How many K^+ and S^{2-} ions would be in one formula unit of the ionic compound formed by these ions?

59. Explain metallic bonding in terms of the sparsely populated outermost orbitals of metal atoms.

60. Explain the role of molecular geometry in determining molecular polarity.

61. How does the energy level of a hybrid orbital compare with the energy levels of the orbitals it was formed from?

62. Aluminum's enthalpy of vaporization is 284 kJ/mol. Beryllium's enthalpy of vaporization is 224 kJ/mol. In which element is the bonding stronger between atoms?

63. Determine the electronegativity difference, the probable bonding type, and the more-electronegative atom for each of the following pairs of atoms:
a. Zn and O   c. S and Cl
b. Br and I

64. Draw the Lewis structure for each of the following molecules:
a. PCl_3 c. CH_3NH_2
b. CCl_2F_2

65. Write the Lewis structure for $BeCl_2$. (Hint: Beryllium atoms do not follow the octet rule.)

66. Draw a Lewis structure for each of the following polyatomic ions and determine their geometries:
a. NO_2^- c. NH_4^+
b. NO_3^-

67. Why do most atoms tend to chemically bonded to other atoms?

CRITICAL THINKING

68. Inferring Relationships The length of a bond varies depending on the type of bond formed. Predict and compare the lengths of the carbon-carbon bonds in the following molecules. Explain your answer. (Hint: See Table 2.)

$$
\begin{array}{ccc}
\text{H--C--C--H} & \text{H--C=C--H} & \text{H--C}\equiv\text{C--H} \\
C_2H_6 & C_2H_4 & C_2H_2
\end{array}
$$

69. Why does F generally form covalent bonds with great polarity?

70. Explain what is wrong with the following Lewis structures, and then correct each one.
a. $H-H-\ddot{\underset{..}{S}}:$

b. $H-\overset{\overset{:O:}{\|}}{C}=\ddot{\underset{..}{O}}-H$

c. $:\ddot{\underset{..}{Cl}} \quad \overset{\overset{:\ddot{Cl}:}{\|}}{\underset{}{N}} \quad \ddot{\underset{..}{Cl}}:$

71. Ionic compounds tend to have higher boiling points than covalent substances do. Both ammonia, NH_3, and methane, CH_4, are covalent compounds, yet the boiling point of ammonia is 130°C higher than that of methane. What might account for this large difference?

USING THE HANDBOOK

72. Figure 18 shows a model for a body-centered cubic crystal. Review the Properties tables for all of the metals in the *Elements Handbook*. What metals exist in body-centered cubic structures?

73. Group 14 of the *Elements Handbook* contains a discussion of semiconductors and the band theory of metals. How does this model explain the electrical conductivity of metals?

RESEARCH & WRITING

74. Prepare a report on the work of Linus Pauling.
   a. Discuss his work on the nature of the chemical bond.
   b. Linus Pauling was an advocate of the use of vitamin C as a preventative for colds. Evaluate Pauling's claims. Determine if there is any scientific evidence that indicates whether vitamin C helps prevent colds.

75. Covalently bonded solids, such as silicon, an element used in computer components, are harder than pure metals. Research theories that explain the hardness of covalently bonded solids and their usefulness in the computer industry. Present your findings to the class.

76. Natural rubber consists of long chains of carbon and hydrogen atoms covalently bonded together. When Goodyear accidentally dropped a mixture of sulfur and rubber on a hot stove, the energy from the stove joined these chains together to make vulcanized rubber. Vulcan was the Roman god of fire. The carbon-hydrogen chains in vulcanized rubber are held together by two sulfur atoms that form covalent bonds between the chains. These covalent bonds are commonly called disulfide bridges. Explore other molecules that have such disulfide bridges. Present your findings to the class.

77. Searching for the perfect artificial sweetener— great taste with no Calories—has been the focus of chemical research for some time. Molecules such as sucralose, aspartamine, and saccharine owe their sweetness to their size and shape. One theory holds that any sweetener must have three sites that fit into the proper taste buds on the tongue. This theory is appropriately known as the "triangle theory." Research artificial sweeteners to develop a model to show how the triangle theory operates.

ALTERNATIVE ASSESSMENT

78. Devise a set of criteria that will allow you to classify the following substances as ionic or nonionic: $CaCO_3$, Cu, H_2O, NaBr, and C (graphite). Show your criteria to your instructor.

79. **Performance Assessment**   Identify 10 common substances in and around your home, and indicate whether you would expect these substances to contain ionic, covalent, or metallic bonds.

extension

Graphing Calculator
Classifying Bond Type
Go to **go.hrw.com** for a graphing calculator exercise that asks you to classify bonding type based on electronegativities of atoms.

Keyword: HC6BNDX

Drawing Lewis dot structures can help you understand how valence electrons participate in bonding. Dots are placed around the symbol of an element to represent the element's valence electrons. For example, carbon has four valence electrons, and its Lewis dot structure is usually written as $\cdot\overset{\displaystyle\cdot}{C}\cdot$. An atom of fluorine has seven valence electrons. Fluorine's Lewis dot structure can be written as $:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot}{F}}:$. When Lewis structures for covalently bonded atoms are written, the dots may be placed as needed to show the electrons shared in each bond. Most atoms bond in a way that gives them a stable octet of s and p electrons in the highest energy level. So, whenever possible, dots should be arranged in a way that represents a stable octet around each atom.
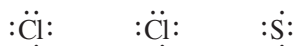
Problem-Solving *TIPS*

- Hydrogen is an exception to the octet rule because hydrogen has only one electron and becomes stable with two electrons.
- Some elements, such as boron, can bond without achieving an octet because they have three or fewer electrons to share.

SAMPLE

Draw the Lewis dot structure for a molecule of sulfur dichloride, SCl_2.
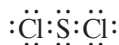
First, write the electron dot notation for each atom.

$:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot}{Cl}}:$ $:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot}{Cl}}:$ $:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot\cdot}{S}}:$

Next, determine the total number of valence electrons in the atoms.

S	$1 \times 6e^-$	=	$6e^-$
2Cl	$2 \times 7e^-$	=	$14e^-$
	Total e^-	=	$20e^-$

Arrange the atoms to form a skeleton structure for the molecule, and place electron pairs between atoms to represent covalent bonds. You can predict the arrangement of atoms by figuring out how many covalent bonds each atom must form in order to achieve a stable octet. Each chlorine atom, which has 7 valence electrons, must form a single covalent bond. Sulfur, which has 6 valence electrons, must form two covalent bonds. The only possible structure is Cl—S—Cl.

Finally, insert dots representing the remaining electrons (16 in this case), in order to give each atom an octet.

$:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot\cdot}{Cl}}:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot\cdot}{S}}:\overset{\displaystyle\cdot\cdot}{\underset{\displaystyle\cdot\cdot}{Cl}}:$

PRACTICE PROBLEMS

1. Draw the electron dot notations for a silicon atom and a strontium atom.

2. Draw Lewis structures for hydrogen sulfide, H_2S, and for formic acid, HCO_2H.

Answer the following items on a separate piece of paper.

1. A chemical bond results from the mutual attraction of the nuclei for
   A. electrons.
   B. neutrons.
   C. protons.
   D. dipoles.

2. A polar covalent bond is likely to form between two atoms that
   A. are similar in electronegativity.
   B. are of similar size.
   C. differ in electronegativity.
   D. have the same number of electrons.

3. The Lewis structure of HCN contains
   A. one double bond and one single bond.
   B. one triple bond and one single bond.
   C. two single bonds.
   D. two double bonds.

4. According to VSEPR theory, the molecular geometry for CH_3^+ is
   A. tetrahedral.
   B. trigonal-pyramidal.
   C. bent or angular.
   D. None of the above

5. Which molecule contains a double bond?
   A. $COCl_2$
   B. C_2H_6
   C. CF_4
   D. SF_2

6. Which molecule is polar?
   A. CCl_4
   B. CO_2
   C. SO_3
   D. none of these

7. What is the hybridization of the carbon atoms in C_2H_2?
   A. sp
   B. sp^2
   C. sp^3
   D. The carbon atoms do not hybridize in C_2H_2.

8. Which of the following compounds is predicted to have the highest boiling point?
   A. HCl
   B. CH_3COOH (Note: The two oxygen atoms bond to the carbon.)
   C. Cl_2
   D. SO_2

9. An unknown substance is an excellent electrical conductor in the solid state and is malleable. What type of chemical bonding does this substance exhibit?
   A. ionic bonding
   B. molecular bonding
   C. metallic bonding
   D. cannot determine from the information given

SHORT ANSWER

10. What does the hybridization model help explain?

11. Explain why ionic crystals are brittle.

EXTENDED RESPONSE

12. Naphthalene, $C_{10}H_8$, is a nonpolar molecule and has a boiling point of 218°C. Acetic acid, CH_3CO_2H, is a polar molecule and has a boiling point of 118°C. Which substance has the stronger intermolecular forces? Briefly explain your answer.

13. Describe and explain the potential energy changes that occur during the formation of a covalent bond.

Test TIP When several questions refer to the same graph, table, drawing, or passage, answer the questions you are sure of first.

Types of Bonding in Solids

OBJECTIVES

- *Observe* the physical properties of different solids.

- *Relate* knowledge of these properties to the type of bonding in each solid.

- *Identify* the type of bonding in an unknown solid.

MATERIALS

- beakers, 50 mL (6)
- Bunsen burner
- copper wire
- deionized water
- evaporating dishes or crucibles (6)
- graduated cylinder, 10 mL
- aluminum shot
- LED conductivity tester
- silicon dioxide (sand)
- sodium chloride (NaCl)
- spatula
- sucrose
- test tubes, small, with solid rubber stoppers (6)
- test-tube rack
- tongs
- unknown substance
- wire gauze, support stand, iron ring, and clay triangle

BACKGROUND

The purpose of this experiment is to relate certain properties of solids to the type of bonding the solids have. These observable properties depend on the type of bonding that holds the molecules, atoms, or ions together in each solid. Depending on the type of bonding, solids may be described as ionic, molecular, metallic, or covalent network solids. The properties to be studied are relative melting point, solubility in aqueous solution, and electrical conductivity.

SAFETY

For review of safety, please see the **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Make a data table in which to record the results of melting, water solubility, solid conductivity, aqueous solution conductivity, and type of bonding of each substance tested.

PROCEDURE

1. Place 1 g samples of each substance into separate evaporating dishes.

2. Touch the electrodes of the conductivity tester to each solid. After each test, rinse with distilled water and carefully dry the electrodes. Note which substances conducted electricity.

3. Place one evaporating dish on a triangle, and heat with a Bunsen burner. As soon as a solid melts, remove the flame.

4. Repeat this procedure for every substance. Do not heat any substance for more than 5 min. There may be some substances that will not melt.

5. Note which substances melted and how long the substances took to melt.

6. Place five test tubes in the test-tube rack. Place 0.5 g of each solid into its own individual test tube. Add 5 mL of deionized water to each test tube. Stopper and shake each test tube in an attempt to dissolve the solid.

7. Note which substances dissolved in the water.

8. Place the solutions or mixtures into separate 50 mL beakers, and immerse the electrodes of the conductivity tester. Rinse the electrodes with the solvent (deionized water) before and after each test. Note which substances conduct electricity.

CLEANUP AND DISPOSAL

9. Dispose of solids and solutions in containers designated by your teacher.

10. Clean all equipment and return it to its proper place.

11. Wash your hands thoroughly after cleaning up your area and equipment.

ANALYSIS AND INTERPRETATION

1. **Analyzing Methods:** Why did you rinse the electrodes before each conductivity test?

2. **Analyzing Methods:** Why did you use deionized water in making the solutions?

3. **Organizing Data:** List the results that each type of bonding should show.

CONCLUSIONS

1. **Inferring Conclusions:** What type of bonding describes each substance? Explain your reasoning.

2. **Inferring Conclusions:** Comparing the properties of your unknown solid with the properties of the known solids, determine the type of bonding present in your unknown solid.

EXTENSIONS

1. **Evaluating Methods:** Is it possible, for a specific type of bonding, for these properties to vary from what was observed in this experiment? If so, give an example of such a variance.

2. **Applying Conclusions:** Think about diamond. What would you predict to be the results of this experiment performed on diamond, and what would you predict the bond type to be?

Chemical Formulas and Chemical Compounds

Chemists use chemical names and formulas to describe the atomic composition of compounds.

Chemicals and Pigments
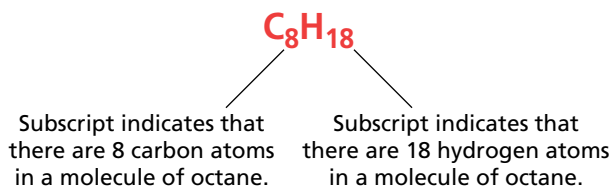
Chemical Names and Formulas

OBJECTIVES

- Explain the significance of a chemical formula.

- Determine the formula of an ionic compound formed between two given ions.

- Name an ionic compound given its formula.

- Using prefixes, name a binary molecular compound from its formula.

- Write the formula of a binary molecular compound given its name.

The total number of natural and synthetic chemical compounds runs in the millions. For some of these substances, certain common names remain in everyday use. For example, calcium carbonate is better known as limestone, and sodium chloride is usually referred to simply as table salt. And everyone recognizes dihydrogen monoxide by its popular name, water.

Unfortunately, common names usually give no information about chemical composition. To describe the atomic makeup of compounds, chemists use systematic methods for naming compounds and for writing chemical formulas. In this chapter, you will be introduced to some of the rules used to identify simple chemical compounds.

Significance of a Chemical Formula

Recall that a chemical formula indicates the relative number of atoms of each kind in a chemical compound. For a molecular compound, the chemical formula reveals the number of atoms of each element contained in a single molecule of the compound, as shown below for the hydrocarbon octane. (*Hydrocarbons* are molecular compounds composed solely of carbon and hydrogen.)

$$C_8H_{18}$$

Subscript indicates that there are 8 carbon atoms in a molecule of octane.

Subscript indicates that there are 18 hydrogen atoms in a molecule of octane.
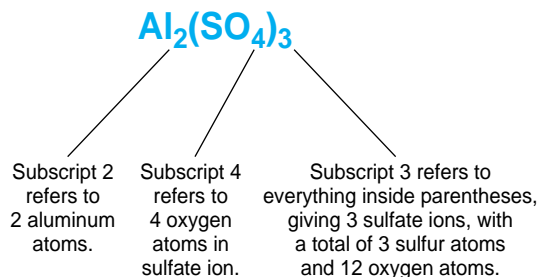
Unlike a molecular compound, an ionic compound consists of a lattice of positive and negative ions held together by mutual attraction. The chemical formula for an ionic compound represents one formula unit—the simplest ratio of the compound's positive ions (cations) and its negative ions (anions). The chemical formula for aluminum sulfate, an ionic compound consisting of aluminum cations and polyatomic sulfate anions, is written as shown on the next page.

$$Al_2(SO_4)_3$$

| Subscript 2 refers to 2 aluminum atoms. | Subscript 4 refers to 4 oxygen atoms in sulfate ion. | Subscript 3 refers to everything inside parentheses, giving 3 sulfate ions, with a total of 3 sulfur atoms and 12 oxygen atoms. |

Note how the parentheses are used. They surround the polyatomic anion to identify it as a unit. The subscript 3 refers to the entire unit. Notice also that there is no subscript written next to the symbol for sulfur. When there is no subscript written next to an atom's symbol, the value of the subscript is understood to be 1.

Monatomic Ions

By gaining or losing electrons, many main-group elements form ions with noble-gas configurations. For example, Group 1 metals lose one electron to give 1+ cations, such as Na^+. Group 2 metals lose two electrons to give 2+ cations, such as Mg^{2+}. *Ions formed from a single atom are known as* **monatomic ions.** The nonmetals of Groups 15, 16, and 17 gain electrons to form anions. For example, in ionic compounds nitrogen forms the 3– anion, N^{3-}. The three added electrons plus the five outermost electrons in nitrogen atoms give a completed outermost octet. Similarly, the Group 16 elements oxygen and sulfur form 2– anions, and the Group 17 halogens form 1– anions.

Not all main-group elements readily form ions, however. Rather than gain or lose electrons, atoms of carbon and silicon form covalent bonds in which they share electrons with other atoms. Other elements tend to form ions that do not have noble-gas configurations. For instance, it is difficult for the Group 14 metals tin and lead to lose four electrons to achieve a noble-gas configuration. Instead, they tend to lose the two electrons in their outer *p* orbitals but retain the two electrons in their outer *s* orbitals to form 2+ cations. (Tin and lead can also form molecular compounds in which all four valence electrons are involved in covalent bonding.)

Elements from the *d*-block form 2+, 3+, or, in a few cases, 1+ or 4+ cations. Many *d*-block elements form two ions of different charges. For example, copper forms 1+ and 2+ cations. Iron and chromium each form 2+ cations as well as 3+ cations. And vanadium forms 2+, 3+, and 4+ cations.

Naming Monatomic Ions

Monatomic cations are identified simply by the element's name, as illustrated by the examples at left. Naming monatomic anions is slightly more

Examples of Cations

K^+

Potassium cation

Mg^{2+}

Magnesium cation

complicated. First, the ending of the element's name is dropped. Then the ending -*ide* is added to the root name, as illustrated by the examples at right.

The names and symbols of the common monatomic cations and anions are organized according to their charges in **Table 1.** The names of many of the ions in the table include Roman numerals. These numerals are part of the *Stock system* of naming chemical ions and elements. You will read more about the Stock system and other systems of naming chemicals later in this chapter.

Examples of Anions

Element	Anion
F	**F$^-$**
Fluor*ine*	Fluor*ide* anion
N	**N^{3-}**
Nit*rogen*	Nit*ride* anion

TABLE 1 *Some Common Monatomic Ions*

Main-group elements

1+		2+		3+		
lithium	Li$^+$	beryllium	Be^{2+}	aluminum	Al^{3+}	
sodium	Na$^+$	magnesium	Mg^{2+}			
potassium	K$^+$	calcium	Ca^{2+}			
rubidium	Rb$^+$	strontium	Sr^{2+}			
cesium	Cs$^+$	barium	Ba^{2+}			

1−		2−		3−		
fluoride	F$^-$	oxide	O^{2-}	nitride	N^{3-}	
chloride	Cl$^-$	sulfide	S^{2-}	phosphide	P^{3-}	
bromide	Br$^-$					
iodide	I$^-$					

d-Block elements and others with multiple ions

| 1+		2+		3+		4+	
copper(I)	Cu$^+$	vanadium(II)	V^{2+}	vanadium(III)	V^{3+}	vanadium(IV)	V^{4+}
silver	Ag$^+$	chromium(II)	Cr^{2+}	chromium(III)	Cr^{3+}	tin(IV)	Sn^{4+}
		manganese(II)	Mn^{2+}	iron(III)	Fe^{3+}	lead(IV)	Pb^{4+}
		iron(II)	Fe^{2+}	cobalt(III)	Co^{3+}		
		cobalt(II)	Co^{2+}				
		nickel(II)	Ni^{2+}				
		copper(II)	Cu^{2+}				
		zinc	Zn^{2+}				
		cadmium	Cd^{2+}				
		tin(II)	Sn^{2+}				
		mercury(II)	Hg^{2+}				
		lead(II)	Pb^{2+}				

Binary Ionic Compounds

Compounds composed of two elements are known as **binary compounds.** In a binary ionic compound, the total numbers of positive charges and negative charges must be equal. Therefore, the formula for such a compound can be written given the identities of the compound's ions. For example, magnesium and bromine combine to form the ionic compound magnesium bromide. Magnesium, a Group 2 metal, forms the Mg^{2+} cation. Note that the $^{2+}$ in Mg^{2+} is written as a superscript. Bromine, a halogen, forms the Br^- anion when combined with a metal. In each formula unit of magnesium bromide, two Br^- anions are required to balance the 2+ charge of the Mg^{2+} cation. The compound's formula must therefore indicate one Mg^{2+} cation and two Br^- anions. The symbol for the cation is written first.

*Ions combined:* Mg^{2+}, Br^-, Br^- *Chemical formula:* $MgBr_2$

Note that the $_2$ in Br_2 is written as a subscript. The charges of the ions are not included in the formula. This is usually the case when writing formulas for binary ionic compounds.

As an aid to determining subscripts in formulas for ionic compounds, the positive and negative charges can be "crossed over." Crossing over is a method of balancing the charges between ions in an ionic compound. For example, the formula for the compound formed by the aluminum ion, Al^{3+}, and the oxide ion, O^{2-}, is determined as follows.

1. *Write the symbols for the ions side by side. Write the cation first.*

$$Al^{3+} \ O^{2-}$$
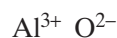
2. *Cross over the charges by using the absolute value of each ion's charge as the subscript for the other ion.*
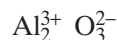
$$Al^{3+}_2 \ O^{2-}_3$$

3. *Check the subscripts and divide them by their largest common factor to give the smallest possible whole-number ratio of ions. Then write the formula.*

Multiplying the charge by the subscript shows that the charge on two Al^{3+} cations $(2 \times 3+ = 6+)$ equals the charge on three O^{2-} anions $(3 \times 2- = 6-)$. The largest common factor of the subscripts is 1. The correct formula is therefore written as follows.

$$Al_2O_3$$

Naming Binary Ionic Compounds

The **nomenclature,** *or naming system,* of binary ionic compounds involves combining the names of the compound's positive and negative ions. The name of the cation is given first, followed by the name of

the anion. For most simple ionic compounds, the ratio of the ions is not indicated in the compound's name because it is understood based on the relative charges of the compound's ions. The naming of a simple binary ionic compound is illustrated below.
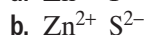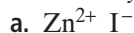
$$Al_2O_3$$

Name of cation   Name of anion

aluminum   **oxide**
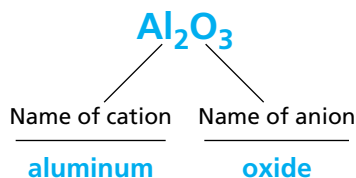
SAMPLE PROBLEM A

Write the formulas for the binary ionic compounds formed between the following elements:
a. zinc and iodine   **b. zinc and sulfur**

SOLUTION   *Write the symbols for the ions side by side. Write the cation first.*
a. Zn^{2+} I^-
b. Zn^{2+} S^{2-}

Cross over the charges to give subscripts.
a. Zn_1^{2+} I_2^-
b. Zn_2^{2+} S_2^{2-}
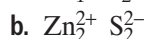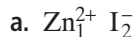
Check the subscripts and divide them by their largest common factor to give the smallest possible whole-number ratio of ions. Then write the formula.

a. The subscripts are mathematically correct because they give equal total charges of $1 \times 2+ = 2+$ and $2 \times 1- = 2-$. The largest common factor of the subscripts is 1.
The smallest possible whole-number ratio of ions in the compound is therefore 1:2.
The subscript 1 is not written, so the formula is ZnI_2.

b. The subscripts are mathematically correct because they give equal total charges of $2 \times 2+ = 4+$ and $2 \times 2- = 4-$. The largest common factor of the subscripts is 2.
The smallest whole-number ratio of ions in the compound is therefore 1:1. The correct formula is ZnS.

PRACTICE

Answers in Appendix E

1. Write formulas for the binary ionic compounds formed between the following elements:
   a. potassium and iodine   **d.** aluminum and sulfur
   b. magnesium and chlorine **e.** aluminum and nitrogen
   c. sodium and sulfur

2. Name the binary ionic compounds indicated by the following formulas:
   a. $AgCl$ **e.** BaO
   b. ZnO **f.** $CaCl_2$
   c. $CaBr_2$
   d. SrF_2

The Stock System of Nomenclature

Some elements, such as iron, form two or more cations with different charges. To distinguish the ions formed by such elements, scientists use the Stock system of nomenclature. This system uses a Roman numeral to indicate an ion's charge. The numeral is enclosed in parentheses and placed *immediately* after the metal name.

$$Fe^{2+} \qquad\qquad\qquad Fe^{3+}$$
$$iron(II) \qquad\qquad\qquad iron(III)$$
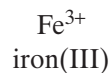
Names of metals that commonly form only one cation do *not* include a Roman numeral.

$$Na^+ \qquad\qquad Ba^{2+} \qquad\qquad Al^{3+}$$
$$sodium \qquad\qquad barium \qquad\qquad aluminum$$
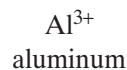
There is no element that commonly forms more than one monatomic anion.

Naming a binary ionic compound according to the Stock system is illustrated below.

$$CuCl_2$$

Name of cation	+	Roman numeral indicating charge	Name of anion
copper(II)			chloride
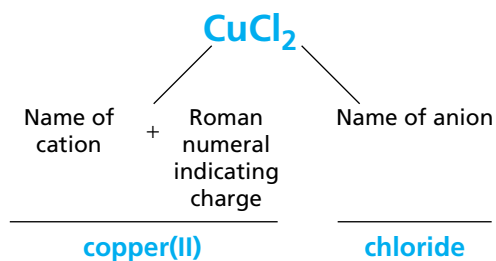
FIGURE 1 Different cations of the same metal form different compounds even when they combine with the same anion. Compare (a) lead(IV) oxide, PbO_2, with (b) lead(II) oxide, PbO.

Write the formula and give the name for the compound formed by the ions Cr^{3+} and F^-.

SOLUTION *Write the symbols for the ions side by side. Write the cation first.*

$$Cr^{3+} \quad F^-$$
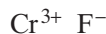
Cross over the charges to give subscripts.

$$Cr^{3+}_1 \quad F^-_3$$
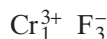
Check the subscripts and write the formula.
The subscripts are correct because they give charges of $1 \times 3+ = 3+$ and $3 \times 1- = 3-$. The largest common factor of the subscripts is 1, so the smallest whole-number ratio of the ions is 1:3. The formula is therefore CrF_3. As **Table 1** shows, chromium forms more than one ion. Therefore, the name of the 3+ chromium ion must be followed by a Roman numeral indicating its charge. The compound's name is chromium(III) fluoride.

PRACTICE *Answers in Appendix E*

1. Write the formula and give the name for the compounds formed between the following ions:
   a. Cu^{2+} and Br^- **d.** Hg^{2+} and S^{2-}
   b. Fe^{2+} and O^{2-} **e.** Sn^{2+} and F^-
   c. Pb^{2+} and Cl^- **f.** Fe^{3+} and O^{2-}

2. Give the names for the following compounds:
   a. CuO **c.** SnI_4
   b. CoF_3 **d.** FeS

extension
Go to **go.hrw.com** for more practice problems that ask you to write formulas for ionic compounds.

**Keyword: HC6FRMX**

Compounds Containing Polyatomic Ions

Table 2 on the next page lists some common polyatomic ions. Most are negatively charged and most are **oxyanions**—*polyatomic ions that contain oxygen.* Some elements can combine with oxygen to form more than one type of oxyanion. For example, nitrogen can form NO_3^- or NO_2^-. The name given a compound containing such an oxyanion depends on the number of oxygen atoms in the oxyanion. The name of the ion with the greater number of oxygen atoms ends in *-ate*. The name of the ion with the smaller number of oxygen atoms ends in *-ite*.

$$NO_2^- \qquad\qquad NO_3^-$$
$$\text{nitrite} \qquad\qquad \text{nitrate}$$

Sometimes, an element can form more than two types of oxyanions. In this case, an anion that has one fewer oxygen atom than the *-ite* anion

TABLE 2 Some Polyatomic Ions

1+		2+			
ammonium	NH_4^+	dimercury*	Hg_2^{2+}		

1−		2−		3−	
acetate	CH_3COO^-	carbonate	CO_3^{2-}	arsenate	AsO_4^{3-}
bromate	BrO_3^-	chromate	CrO_4^{2-}	phosphate	PO_4^{3-}
chlorate	ClO_3^-	dichromate	$Cr_2O_7^{2-}$		
chlorite	ClO_2^-	hydrogen phosphate	HPO_4^{2-}		
cyanide	CN^-	oxalate	$C_2O_4^{2-}$		
dihydrogen phosphate	$H_2PO_4^-$	peroxide	O_2^{2-}		
hydrogen carbonate (bicarbonate)	HCO_3^-	sulfate	SO_4^{2-}		
hydrogen sulfate	HSO_4^-	sulfite	SO_3^{2-}		
hydroxide	OH^-				
hypochlorite	ClO^-				
nitrate	NO_3^-				
nitrite	NO_2^-				
perchlorate	ClO_4^-				
permanganate	MnO_4^-				

*The mercury(I) cation exists as two Hg^+ ions joined together by a covalent bond and is written as Hg_2^{2+}.

has is given the prefix *hypo-*. An anion that has one more oxygen atom than the *-ate* anion has is given the prefix *per-*. This nomenclature is illustrated by the four oxyanions formed by chlorine.

ClO^-	ClO_2^-	ClO_3^-	ClO_4^-
hypochlorite	chlorite	chlorate	perchlorate

Compounds containing polyatomic ions are named in the same manner as binary ionic compounds. The name of the cation is given first, followed by the name of the anion. For example, the two compounds formed with silver by the nitrate and nitrite anions are named *silver nitrate,* $AgNO_3$, and *silver nitrite,* $AgNO_2$, respectively. When multiples of a polyatomic ion are present in a compound, the formula for the polyatomic ion is enclosed in parentheses, as shown on page 220 for aluminum sulfate, $Al_2(SO_4)_3$. The formula indicates that an aluminum sulfate formula unit has two aluminum cations and three sulfate anions.

Write the formula for tin(IV) sulfate.

SOLUTION   *Write the symbols for the ions side by side. Write the cation first.*
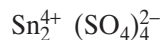
$$Sn^{4+} \quad SO_4^{2-}$$
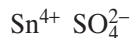
Cross over the charges to give subscripts. Add parentheses around the polyatomic ion if necessary.

$$Sn_2^{4+} \quad (SO_4)_4^{2-}$$

Check the subscripts and write the formula.

The total positive charge is $2 \times 4+ = 8+$. The total negative charge is $4 \times 2- = 8-$. The charges are equal. The largest common factor of the subscripts is 2, so the smallest whole-number ratio of ions in the compound is 1:2. The correct formula is therefore $Sn(SO_4)_2$.

PRACTICE   *Answers in Appendix E*

1. Write formulas for the following ionic compounds:
   a. sodium iodide
   b. calcium chloride
   c. potassium sulfide
   d. lithium nitrate
   e. copper(II) sulfate
   f. sodium carbonate
   g. calcium nitrite
   h. potassium perchlorate

2. Give the names for the following compounds:
   a. Ag_2O
   b. $Ca(OH)_2$
   c. $KClO_3$
   d. NH_4OH
   e. $Fe_2(CrO_4)_3$
   f. $KClO$

Naming Binary Molecular Compounds

Unlike ionic compounds, molecular compounds are composed of individual covalently bonded units, or molecules. Chemists use two nomenclature systems to name binary molecules. The newer system is the Stock system for naming molecular compounds, which requires an understanding of oxidation numbers. This system will be discussed in Section 2.

The old system of naming molecular compounds is based on the use of prefixes. For example, the molecular compound CCl_4 is named carbon *tetra*chloride. The prefix *tetra-* indicates that four chloride atoms are present in a single molecule of the compound. The two oxides of carbon, CO and CO_2, are named carbon *mon*oxide and carbon *di*oxide, respectively.
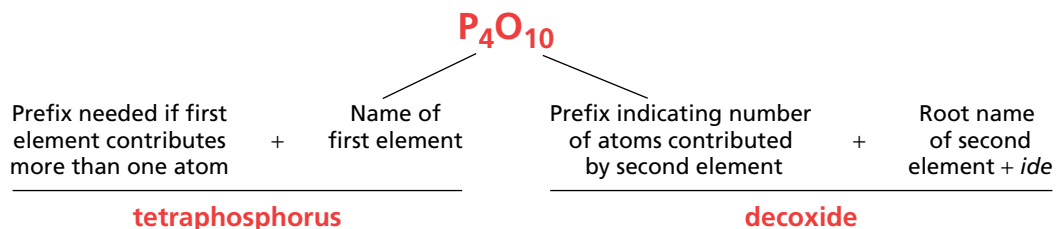
TABLE 3 *Numerical Prefixes*	
Number	**Prefix**
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

In these names, the prefix *mon-* (from *mono-*) indicates one oxygen atom, and the prefix *di-* indicates two oxygen atoms. The prefixes used to specify the number of atoms in a molecule are listed in **Table 3.**

The rules for the prefix system of nomenclature of binary molecular compounds are as follows.

1. The element that has the smaller group number is usually given first. If both elements are in the same group, the element whose period number is greater is given first. The element is given a prefix only if it contributes more than one atom to a molecule of the compound.

2. The second element is named by combining (a) a prefix indicating the number of atoms contributed by the element, (b) the root of the name of the element, and (c) the ending *-ide*. With few exceptions, the ending *-ide* indicates that a compound contains only two elements.

3. The *o* or *a* at the end of a prefix is usually dropped when the word following the prefix begins with another vowel. For example, one would write *monoxide* and *pentoxide* instead of *mono-oxide* and *penta-oxide*.

The prefix system is illustrated below.

$$P_4O_{10}$$

Prefix needed if first element contributes more than one atom	+	Name of first element	Prefix indicating number of atoms contributed by second element	+	Root name of second element + *ide*
tetraphosphorus			**decoxide**		

In general, the order of nonmetals in binary compound names and formulas is C, P, N, H, S, I, Br, Cl, O, and F.

TABLE 4 Binary Compounds of Nitrogen and Oxygen	
Formula	Prefix-system name
N_2O	dinitrogen monoxide
NO	nitrogen monoxide
NO_2	nitrogen dioxide
N_2O_3	dinitrogen trioxide
N_2O_4	dinitrogen tetroxide
N_2O_5	dinitrogen pentoxide

The prefix system is illustrated further in **Table 4,** which lists the names of the six oxides of nitrogen. Note the application of rule 1, for example, in the name *nitrogen dioxide* for NO_2. No prefix is needed with *nitrogen* because only one atom of nitrogen is present in a molecule of NO_2. On the other hand, according to rule 2, the prefix *di-* in *dioxide* is needed to indicate the presence of two atoms of oxygen. Take a moment to review the prefixes in the other names in **Table 4.**

SAMPLE PROBLEM D

a. Give the name for As_2O_5.
b. Write the formula for oxygen difluoride.

SOLUTION

a. A molecule of the compound contains two arsenic atoms, so the first word in the name is *diarsenic.* The five oxygen atoms are indicated by adding the prefix *pent-* to the word *oxide.* The complete name is *diarsenic pentoxide.*

b. The first symbol in the formula is that for oxygen. Oxygen is first in the name because it is less electronegative than fluorine. Since there is no prefix, there must be only one oxygen atom. The prefix *di-* in *difluoride* shows that there are two fluorine atoms in the molecule. The formula is OF_2.

PRACTICE

Answers in Appendix E

1. Name the following binary molecular compounds:
   a. SO_3
   b. ICl_3
   c. PBr_5

2. Write formulas for the following compounds:
   a. carbon tetraiodide
   b. phosphorus trichloride
   c. dinitrogen trioxide

Covalent-Network Compounds

As you read in Chapter 6, some covalent compounds do not consist of individual molecules. Instead, each atom is joined to all its neighbors in a covalently bonded, three-dimensional network. There are no distinct units in these compounds, just as there are no such units in ionic compounds. The subscripts in a formula for a covalent-network compound indicate the smallest whole-number ratio of the atoms in the compound. Naming such compounds is similar to naming molecular compounds. Some common examples are given below.

SiC	SiO_2	Si_3N_4
silicon carbide	silicon dioxide	trisilicon tetranitride

Acids and Salts

An *acid* is a distinct type of molecular compound about which you will read in detail in Chapter 14. Most acids used in the laboratory can be classified as either binary acids or oxyacids. *Binary acids* are acids that consist of two elements, usually hydrogen and one of the halogens—fluorine, chlorine, bromine, iodine. *Oxyacids* are acids that contain hydrogen, oxygen, and a third element (usually a nonmetal).

Acids were first recognized as a specific class of compounds based on their properties in solutions of water. Consequently, in chemical nomenclature, the term *acid* usually refers to a solution in water of one of these special compounds rather than to the compound itself. For example, *hydrochloric acid* refers to a water solution of the molecular compound hydrogen chloride, HCl. Some common binary and oxyacids are listed in **Table 5.**

Many polyatomic ions are produced by the loss of hydrogen ions from oxyacids. A few examples of the relationship between oxyacids and oxyanions are shown below.

sulfuric acid	H_2SO_4	sulfate	SO_4^{2-}
nitric acid	HNO_3	nitrate	NO_3^{-}
phosphoric acid	H_3PO_4	phosphate	PO_4^{3-}

TABLE 5 *Common Binary Acids and Oxyacids*					
HF	hydrofluoric acid	HNO_2	nitrous acid	$HClO$	hypochlorous acid
HCl	hydrochloric acid	HNO_3	nitric acid	$HClO_2$	chlorous acid
HBr	hydrobromic acid	H_2SO_3	sulfurous acid	$HClO_3$	chloric acid
HI	hydriodic acid	H_2SO_4	sulfuric acid	$HClO_4$	perchloric acid
H_3PO_4	phosphoric acid	CH_3COOH	acetic acid	H_2CO_3	carbonic acid

An ionic compound composed of a cation and the anion from an acid is often referred to as a **salt.** Table salt, NaCl, contains the anion from hydrochloric acid. Calcium sulfate, $CaSO_4$, is a salt containing an anion from sulfuric acid. Some salts contain anions in which one or more hydrogen atoms from the acid are retained. Such anions are named by adding the word *hydrogen* or the prefix *bi-* to the anion name. The best known such anion comes from carbonic acid, H_2CO_3.

$$HCO_3^-$$
hydrogen carbonate ion
bicarbonate ion
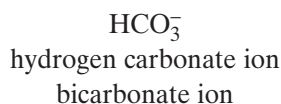
SECTION REVIEW

1. What is the significance of a chemical formula?

2. Write formulas for the compounds formed between the following:

   a. aluminum and bromine

   b. sodium and oxygen

   c. magnesium and iodine

   d. Pb^{2+} and O^{2-}

   e. Sn^{2+} and I^-

   f. Fe^{3+} and S^{2-}

   g. Cu^{2+} and NO_3^-

   h. NH_4^+ and SO_4^{2-}

3. Name the following compounds by using the Stock system:

   a. NaI   **c.** CaO   **e.** CuBr

   b. MgS   **d.** K_2S **f.** $FeCl_2$

4. Write formulas for each of the following compounds:

   a. sodium hydroxide   **e.** carbon diselenide

   b. lead(II) nitrate   **f.** acetic acid

   c. iron(II) sulfate   **g.** chloric acid

   d. diphosphorus trioxide   **h.** sulfurous acid

Critical Thinking

5. **RELATING IDEAS** Draw the Lewis structure, give the name, and predict VSEPR geometry of SCl_2.

Oxidation Numbers

OBJECTIVES

- List the rules for assigning oxidation numbers.

- Give the oxidation number for each element in the formula of a chemical compound.

- Name binary molecular compounds using oxidation numbers and the Stock system.

The charges on the ions composing an ionic compound reflect the electron distribution of the compound. *In order to indicate the general distribution of electrons among the bonded atoms in a molecular compound or a polyatomic ion,* **oxidation numbers,** *also called* **oxidation states,** *are assigned to the atoms composing the compound or ion.* Unlike ionic charges, oxidation numbers do not have an exact physical meaning. In fact, in some cases they are quite arbitrary. However, oxidation numbers are useful in naming compounds, in writing formulas, and in balancing chemical equations. And, as will be discussed in Chapter 19, they are helpful in studying certain types of chemical reactions.

Assigning Oxidation Numbers

As a general rule in assigning oxidation numbers, shared electrons are assumed to belong to the more electronegative atom in each bond. More specific rules for determining oxidation numbers are provided by the following guidelines.

1. The atoms in a pure element have an oxidation number of zero. For example, the atoms in pure sodium, Na, oxygen, O_2, phosphorus, P_4, and sulfur, S_8, all have oxidation numbers of zero.
2. The more-electronegative element in a binary molecular compound is assigned the number equal to the negative charge it would have as an anion. The less-electronegative atom is assigned the number equal to the positive charge it would have as a cation.
3. Fluorine has an oxidation number of −1 in all of its compounds because it is the most electronegative element.
4. Oxygen has an oxidation number of −2 in almost all compounds. Exceptions include when it is in peroxides, such as H_2O_2, in which its oxidation number is −1, and when it is in compounds with fluorine, such as OF_2, in which its oxidation number is +2.
5. Hydrogen has an oxidation number of +1 in all compounds containing elements that are more electronegative than it; it has an oxidation number of −1 in compounds with metals.
6. The algebraic sum of the oxidation numbers of all atoms in a neutral compound is equal to zero.
7. The algebraic sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the charge of the ion.
8. Although rules 1 through 7 apply to covalently bonded atoms, oxidation numbers can also be assigned to atoms in ionic compounds.

9. A monatomic ion has an oxidation number equal to the charge of the ion. For example, the ions Na^+, Ca^{2+}, and Cl^- have oxidation numbers of +1, +2, and –1, respectively.

Let's examine the assignment of oxidation numbers to the atoms in two molecular compounds, hydrogen fluoride, HF, and water, H_2O. Both compounds have polar-covalent bonds. In HF, the fluorine atom should have an oxidation number of –1 (see Rule 3 on the previous page). Rule 5 tells us that hydrogen should have an oxidation number of +1. This makes sense because fluorine is more electronegative than hydrogen, and in a polar-covalent bond, shared electrons are assumed to belong to the more-electronegative element. For water, Rules 4 and 5 tell us that the oxidation number of oxygen should be –2 and the oxidation number of each hydrogen atom should be +1. Again, oxygen is more electronegative than hydrogen, so the shared electrons are assumed to belong to oxygen.

Because the sum of the oxidation numbers of the atoms in a compound must satisfy Rule 6 or 7 of the guidelines on the previous page, it is often possible to assign oxidation numbers when they are not known. This is illustrated in Sample Problem E.

<div style="background:#c00;color:#fff;padding:4px;font-weight:bold">SAMPLE PROBLEM E</div>

Assign oxidation numbers to each atom in the following compounds or ions:

a. UF_6

b. H_2SO_4

c. ClO_3^-

SOLUTION

a. Start by placing known oxidation numbers above the appropriate elements. From the guidelines, we know that fluorine always has an oxidation number of –1.

$$\overset{-1}{U}F_6$$

Multiply known oxidation numbers by the appropriate number of atoms and place the totals underneath the corresponding elements. There are six fluorine atoms, $6 \times -1 = -6$.

$$\overset{-1}{U}F_6 \\ {\scriptstyle -6}$$

The compound UF_6 is molecular. According to the guidelines, the sum of the oxidation numbers must equal zero. The total of positive oxidation numbers is therefore +6.

$$\overset{-1}{U}F_6 \\ {\scriptstyle +6 \; -6}$$

Divide the total calculated oxidation number by the appropriate number of atoms. There is only one uranium atom in the molecule, so it must have an oxidation number of +6.

$$\overset{+6\ -1}{UF_6}$$
$$\underset{+6\ -6}{}$$

b. Oxygen and sulfur are each more electronegative than hydrogen, so hydrogen has an oxidation number of +1. Oxygen is not combined with a halogen, nor is H_2SO_4 a peroxide. Therefore, the oxidation number of oxygen is –2. Place these known oxidation numbers above the appropriate symbols. Place the total of the oxidation numbers underneath.

$$\overset{+1\ \ -2}{H_2SO_4}$$
$$\underset{+2\ \ -8}{}$$

The sum of the oxidation numbers must equal zero, and there is only one sulfur atom in each molecule of H_2SO_4. Because $(+2) + (-8) = -6$, the oxidation number of each sulfur atom must be +6.

c. To assign oxidation numbers to the elements in ClO_3^-, proceed as in parts (a) and (b). Remember, however, that the total of the oxidation numbers should equal the overall charge of the anion, 1–. The oxidation number of a single oxygen atom in the ion is –2. The total oxidation number due to the three oxygen atoms is –6. For the chlorate ion to have a 1– charge, chlorine must be assigned an oxidation number of +5.

$$\overset{+5\ -2}{ClO_3^-}$$
$$\underset{+5\ -6}{}$$

PRACTICE   *Answers in Appendix E*

1. Assign oxidation numbers to each atom in the following compounds or ions:

   a. HCl
   b. CF_4
   c. PCl_3
   d. SO_2

   e. HNO_3
   f. KH
   g. P_4O_{10}

   h. $HClO_3$
   i. N_2O_5
   j. $GeCl_2$

Using Oxidation Numbers for Formulas and Names

As shown in **Table 6,** many nonmetals can have more than one oxidation number. (A more extensive list of oxidation numbers is given in Appendix Table A-15.) These numbers can sometimes be used in the same manner as ionic charges to determine formulas. Suppose, for example, you want to know the formula of a binary compound formed between sulfur and oxygen. From the common +4 and +6 oxidation states of sulfur, you could expect that sulfur might form SO_2 or SO_3. Both are known compounds. Of course, a formula must represent facts. Oxidation numbers alone cannot be used to prove the existence of a compound.

| TABLE 6	Common Oxidation Numbers of Some Nonmetals That Have Variable Oxidation States*		
Group 14	carbon	$-4, +2, +4$	
Group 15	nitrogen	$-3, +1, +2, +3, +4, +5$	
	phosphorus	$-3, +3, +5$	
Group 16	sulfur	$-2, +4, +6$	
Group 17	chlorine	$-1, +1, +3, +5, +7$	
	bromine	$-1, +1, +3, +5, +7$	
	iodine	$-1, +1, +3, +5, +7$	

*In addition to the values shown, atoms of each element in its pure state are assigned an oxidation number of zero.

In Section 1 we introduced the use of Roman numerals to denote ionic charges in the Stock system of naming ionic compounds. The Stock system is actually based on oxidation numbers, and it can be used as an alternative to the prefix system for naming binary molecular compounds. In the prefix system, for example, SO_2 and SO_3 are named sulfur dioxide and sulfur trioxide, respectively. Their names according to the Stock system are sulfur(IV) oxide and sulfur(VI) oxide. The international body that governs nomenclature has endorsed the Stock system, which is more practical for complicated compounds. Prefix-based names and Stock-system names are still used interchangeably for many simple compounds, however. A few additional examples of names in both systems are given below.

	Prefix system	Stock system
PCl_3	phosphorus trichloride	phosphorus(III) chloride
PCl_5	phosphorus pentachloride	phosphorus(V) chloride
N_2O	dinitrogen monoxide	nitrogen(I) oxide
NO	nitrogen monoxide	nitrogen(II) oxide
PbO_2	lead dioxide	lead(IV) oxide
Mo_2O_3	dimolybdenum trioxide	molybdenum(III) oxide

SECTION REVIEW

1. Assign oxidation numbers to each atom in the following compounds or ions:

   a. HF
   b. CI_4
   c. H_2O
   d. PI_3
   e. CS_2
   f. Na_2O_2
   g. H_2CO_3
   h. NO_2^-
   i. SO_4^{2-}
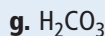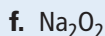
2. Name each of the following binary molecular compounds according to the Stock system:

   a. CI_4
   b. SO_3
   c. As_2S_3
   d. NCl_3
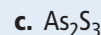
Critical Thinking

3. **DRAWING CONCLUSIONS** Determine the oxidation numbers for iron oxide, Fe_3O_4. (Recall that oxidation numbers are integers.)

Chemistry in Action
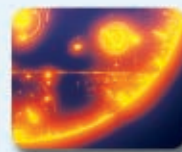
Mass Spectrometry: Identifying Molecules

Tests for locating oil deposits in the ground and detecting dioxins in our food supply are commonly performed today. These tests can be performed by using a technique known as *mass spectrometry.* Mass spectrometry is now used in many fields, such as medicine, chemistry, forensic science, and astronomy.

What is mass spectrometry? It is the most accurate technique available to measure the mass of an individual molecule or atom. Knowing the molecular mass is an essential part of identifying an unknown compound and determining the structure of a molecule of the compound. As the diagram of a mass spectrometer shows, the molecules in a gaseous sample are converted into ions. The ions then are separated and sorted according to their mass-to-charge ratio by a combination of electric and magnetic fields. The fields cause the ions' trajectories to change based on the ions' masses and charges. Then, the sorted ions are detected, and a mass spectrum is obtained. The mass spectrum is a graph of relative intensity (related to the number of ions detected) versus mass-to-charge ratio. Mass spectrometry uses a very small sample size (10^{-12} g) to obtain results.

The resulting spectrum is like a puzzle. It contains numerous peaks that correspond to fragments of the initial molecule. The largest peak (parent peak) corresponds to the molecular mass of the molecular ion. By analyzing the peaks, scientists can determine the identity and structure of a compound. Computers are used to help interpret the spectrum and identify the molecule by using online spectral database libraries.
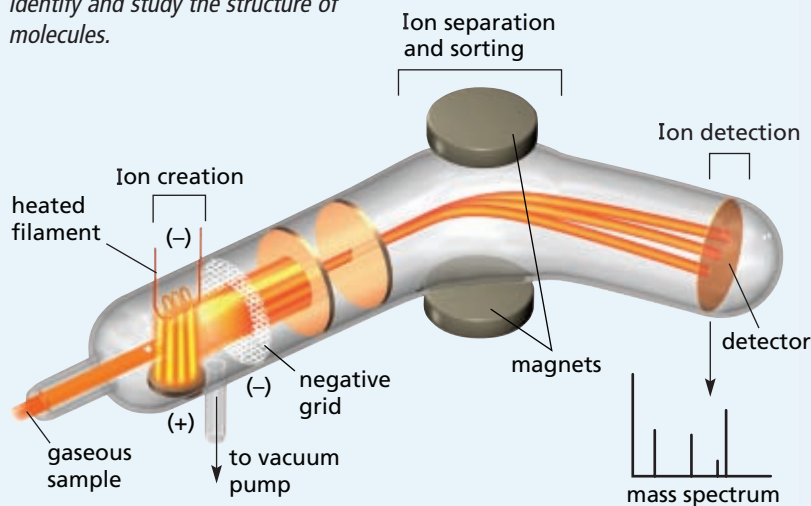
Mass spectrometry has been an essential tool for scientists since its invention in the early 1900s. But its use was limited to small molecules from which ion creation was easy. Large biological molecules could not be studied because they would break down or decompose during conventional ion-formation techniques. In the late 1980s, two groups developed ion-formation methods that are used today in commercial mass spectrometers. John Fenn (Virginia Commonwealth University) developed electrospray ionization

mass spectrometry, and Koichi Tanaka (Shimadzu Corporation, Japan) developed matrix-assisted laser desorption ionization (MALDI) mass spectrometry. In 2002, both scientists received the Nobel Prize in chemistry for their work. Their methods opened the field to the study of large molecules, allowing scientists to use mass spectrometry to study the structure of macromolecules such as nucleic acids and steroids and to identify the sequence of proteins.

Questions

1. Why is it necessary to convert the sample into ions in the mass spectrometer?

2. How have recent developments in mass spectrometry contributed to the field of medicine?

▼ *Scientists use mass spectrometers to identify and study the structure of molecules.*

Ion separation and sorting

Ion detection

Ion creation

heated filament

(–)

(–) negative grid

(+)

gaseous sample

to vacuum pump

magnets

detector

mass spectrum

Using Chemical Formulas

As you have seen, a chemical formula indicates the elements as well as the relative number of atoms or ions of each element present in a compound. Chemical formulas also allow chemists to calculate a number of characteristic values for a given compound. In this section, you will learn how to use chemical formulas to calculate the *formula mass,* the *molar mass,* and the *percentage composition* by mass of a compound.

Formula Masses

In Chapter 3, we saw that hydrogen atoms have an average atomic mass of 1.007 94 amu and that oxygen atoms have an average atomic mass of 15.9994 amu. Like individual atoms, molecules, formula units, or ions have characteristic average masses. For example, we know from the chemical formula H_2O that a single water molecule is composed of exactly two hydrogen atoms and one oxygen atom. The mass of a water molecule is found by adding the masses of the three atoms in the molecule. (In the calculation, the average atomic masses have been rounded to two decimal places.)

$$\text{average atomic mass of H: 1.01 amu}$$
$$\text{average atomic mass of O: 16.00 amu}$$

$$2 \text{ H atoms} \times \frac{1.01 \text{ amu}}{\text{H atom}} = 2.02 \text{ amu}$$

$$1 \text{ O atom} \times \frac{16.00 \text{ amu}}{\text{O atom}} = 16.00 \text{ amu}$$

$$\text{average mass of } H_2O \text{ molecule} = 18.02 \text{ amu}$$

The mass of a water molecule can be correctly referred to as a *molecular mass.* The mass of one NaCl formula unit, on the other hand, is not a molecular mass because NaCl is an ionic compound. The mass of *any* unit represented by a chemical formula, whether the unit is a molecule, a formula unit, or an ion, is known as the formula mass. *The* **formula mass** *of any molecule, formula unit, or ion is the sum of the average atomic masses of all atoms represented in its formula.*

SC*L*INKS.

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Topic: Photochemical Reaction
Code: HC61137

The procedure illustrated for calculating the formula mass of a water molecule can be used to calculate the mass of any unit represented by a chemical formula. In each of the problems that follow, the atomic masses from the periodic table in the back of the book have been rounded to two decimal places.

Find the formula mass of potassium chlorate, $KClO_3$.

SOLUTION The mass of a formula unit of $KClO_3$ is found by summing the masses of one K atom, one Cl atom, and three O atoms. The required atomic masses can be found in the periodic table in the back of the book. In the calculation, each atomic mass has been rounded to two decimal places.

$$1 \; \text{K atom} \times \frac{39.10 \; \text{amu}}{\text{K atom}} = 39.10 \; \text{amu}$$

$$1 \; \text{Cl atom} \times \frac{35.45 \; \text{amu}}{\text{Cl atom}} = 35.45 \; \text{amu}$$

$$3 \; \text{O atoms} \times \frac{16.00 \; \text{amu}}{\text{O atom}} = 48.00 \; \text{amu}$$

$$\text{formula mass of } KClO_3 = 122.55 \; \text{amu}$$

PRACTICE   *Answers in Appendix E*

1. Find the formula mass of each of the following:
   a. H_2SO_4
   b. $Ca(NO_3)_2$
   c. PO_4^{3-}
   d. $MgCl_2$

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate formula mass.

Keyword: HC6FRMX

Molar Masses

In Chapter 3 you learned that the molar mass of a substance is equal to the mass in grams of one mole, or approximately 6.022×10^{23} particles, of the substance. For example, the molar mass of pure calcium, Ca, is 40.08 g/mol because one mole of calcium atoms has a mass of 40.08 g.

The molar mass of a compound is calculated by summing the masses of the elements present in a mole of the molecules or formula units that make up the compound. For example, one mole of water molecules contains exactly two moles of H atoms and one mole of O atoms. Rounded to two decimal places, a mole of hydrogen atoms has a

mass of 1.01 g, and a mole of oxygen atoms has a mass of 16.00 g. The molar mass of water is calculated as follows.

$$2 \text{ mol H} \times \frac{1.01 \text{ g H}}{\text{mol H}} = 2.02 \text{ g H}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 16.00 \text{ g O}$$

molar mass of H_2O = 18.02 g/mol

Figure 3 shows a mole of water as well as a mole of several other substances.

You may have noticed that *a compound's molar mass is numerically equal to its formula mass*. For instance, in Sample Problem F the formula mass of $KClO_3$ was found to be 122.55 amu. Therefore, because molar mass is numerically equal to formula mass, we know that the molar mass of $KClO_3$ is 122.55 g/mol.
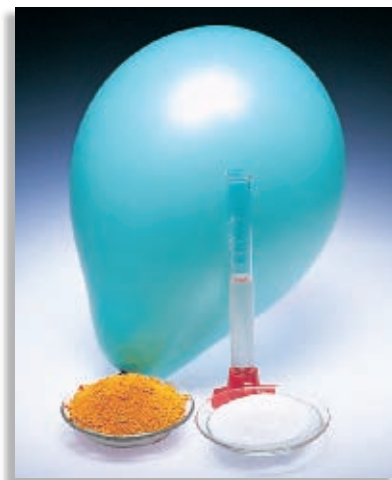

FIGURE 3 Every compound has a characteristic molar mass. Shown here are one mole each of nitrogen (in balloon), water (in graduated cylinder), cadmium sulfide, CdS (yellow substance), and sodium chloride, NaCl (white substance).

SAMPLE PROBLEM G

What is the molar mass of barium nitrate, $Ba(NO_3)_2$?

SOLUTION One mole of barium nitrate contains exactly one mole of Ba^{2+} ions and two moles of NO_3^- ions. The two moles of NO_3^- ions contain two moles of N atoms and six moles of O atoms. Therefore, the molar mass of $Ba(NO_3)_2$ is calculated as follows.

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{\text{mol Ba}} = 137.33 \text{ g Ba}$$

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{\text{mol N}} = 28.02 \text{ g N}$$

$$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 96.00 \text{ g O}$$

molar mass of $Ba(NO_3)_2$ = 261.35 g/mol

PRACTICE   *Answers in Appendix E*

1. How many moles of atoms of each element are there in one mole of the following compounds?
   a. Al_2S_3
   b. $NaNO_3$
   c. $Ba(OH)_2$

2. Find the molar mass of each of the compounds listed in item 1.
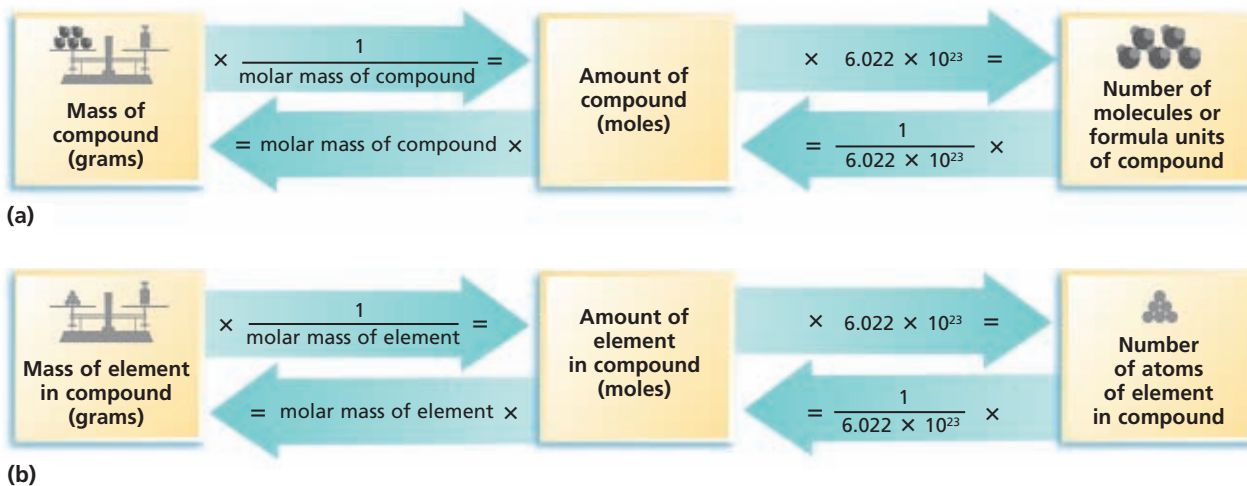
(a)

(b)

FIGURE 4 (a) The diagram shows the relationships between mass in grams, amount in moles, and number of molecules or atoms for a given compound. (b) Similar relationships exist for an element within a compound.

Molar Mass as a Conversion Factor

The molar mass of a compound can be used as a conversion factor to relate an amount in moles to a mass in grams for a given substance. Recall that molar mass usually has the units of grams per mole. To convert a known amount of a compound in moles to a mass in grams, multiply the amount in moles by the molar mass.

$$\text{amount in moles} \times \text{molar mass (g/mol)} = \text{mass in grams}$$

Conversions of this type for elements and compounds are summarized above in **Figure 4.**

SAMPLE PROBLEM H

What is the mass in grams of 2.50 mol of oxygen gas?

SOLUTION

1 ANALYZE **Given:** 2.50 mol O_2
Unknown: mass of O_2 in grams

2 PLAN moles $O_2 \longrightarrow$ grams O_2
To convert amount of O_2 in moles to mass of O_2 in grams, multiply by the molar mass of O_2.

$$\text{amount of } O_2 \text{ (mol)} \times \text{molar mass of } O_2 \text{ (g/mol)} = \text{mass of } O_2 \text{ (g)}$$

3 **COMPUTE** First the molar mass of O_2 must be calculated.

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 32.00 \text{ g (mass of one mole of } O_2)$$

The molar mass of O_2 is therefore 32.00 g/mol. Now do the calculation shown in step 2.

$$2.50 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{\text{mol } O_2} = 80.0 \text{ g } O_2$$

4 **EVALUATE** The answer is correctly given to three significant figures and is close to an estimated value of 75 g (2.50 mol $\times$ 30 g/mol).

To convert a known mass of a compound in grams to an amount in moles, the mass must be divided by the molar mass. Or you can invert the molar mass and multiply so that units are easily canceled.

$$\text{mass in grams} \times \frac{1}{\text{molar mass (g/mol)}} = \text{amount in moles}$$

SAMPLE PROBLEM I

Ibuprofen, $C_{13}H_{18}O_2$, is the active ingredient in many nonprescription pain relievers. Its molar mass is 206.31 g/mol.
a. If the tablets in a bottle contain a total of 33 g of ibuprofen, how many moles of ibuprofen are in the bottle?
b. How many molecules of ibuprofen are in the bottle?
c. What is the total mass in grams of carbon in 33 g of ibuprofen?

SOLUTION

1 **ANALYZE** **Given:** 33 g of $C_{13}H_{18}O_2$, molar mass 206.31 g/mol
Unknown: **a.** moles $C_{13}H_{18}O_2$
      b. molecules $C_{13}H_{18}O_2$
      c. total mass of C

2 **PLAN** **a.** grams $\longrightarrow$ moles
To convert mass of ibuprofen in grams to amount of ibuprofen in moles, multiply by the inverted molar mass of $C_{13}H_{18}O_2$.

$$\text{g } C_{13}H_{18}O_2 \times \frac{1 \text{ mol } C_{13}H_{18}O_2}{206.31 \text{ g } C_{13}H_{18}O_2} = \text{mol } C_{13}H_{18}O_2$$

b. moles $\longrightarrow$ molecules
To find the number of molecules of ibuprofen, multiply amount of $C_{13}H_{18}O_2$ in moles by Avogadro's number.

$$\text{mol } C_{13}H_{18}O_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = \text{molecules } C_{13}H_{18}O_2$$

c. moles $C_{13}H_{18}O_2 \longrightarrow$ moles C $\longrightarrow$ grams C

To find the mass of carbon present in the ibuprofen, the two conversion factors needed are the amount of carbon in moles per mole of $C_{13}H_{18}O_2$ and the molar mass of carbon.

$$\text{mol } C_{13}H_{18}O_2 \times \frac{13 \text{ mol C}}{\text{mol } C_{13}H_{18}O_2} \times \frac{12.01 \text{ g C}}{\text{mol C}} = \text{g C}$$

3 COMPUTE

a. $33 \text{ g } \cancel{C_{13}H_{18}O_2} \times \dfrac{1 \text{ mol } C_{13}H_{18}O_2}{206.31 \text{ g } \cancel{C_{13}H_{18}O_2}} = 0.16 \text{ mol } C_{13}H_{18}O_2$

b. $0.16 \cancel{\text{ mol }} C_{13}H_{18}O_2 \times \dfrac{6.022 \times 10^{23} \text{ molecules}}{\cancel{\text{mol}}} = 9.6 \times 10^{22} \text{ molecules } C_{13}H_{18}O_2$

c. $0.16 \cancel{\text{ mol } C_{13}H_{18}O_2} \times \dfrac{13 \cancel{\text{ mol C}}}{\cancel{\text{mol } C_{13}H_{18}O_2}} \times \dfrac{12.01 \text{ g C}}{\cancel{\text{mol C}}} = 25 \text{ g C}$

The bottle contains 0.16 mol of ibuprofen, which is 9.6×10^{22} molecules of ibuprofen. The sample of ibuprofen contains 25 g of carbon.

4 EVALUATE   Checking each step shows that the arithmetic is correct, significant figures have been used correctly, and units have canceled as desired.

PRACTICE

Answers in Appendix E

1. How many moles of compound are there in the following?
   a. 6.60 g $(NH_4)_2SO_4$
   b. 4.5 kg $Ca(OH)_2$

2. How many molecules are there in the following?
   a. 25.0 g H_2SO_4
   b. 125 g of sugar, $C_{12}H_{22}O_{11}$

3. What is the mass in grams of 6.25 mol of copper(II) nitrate?

Percentage Composition

It is often useful to know the percentage by mass of a particular element in a chemical compound. For example, suppose the compound potassium chlorate, $KClO_3$, were to be used as a source of oxygen. It would be helpful to know the percentage of oxygen in the compound. To find the mass percentage of an element in a compound, one can divide the mass of the element in a sample of the compound by the total mass of the sample, then multiply this value by 100.

$$\frac{\text{mass of element in sample of compound}}{\text{mass of sample of compound}} \times 100 = \frac{\% \text{ element in}}{\text{compound}}$$

The mass percentage of an element in a compound is the same regardless of the sample's size. Therefore, a simpler way to calculate the percentage of an element in a compound is to determine how many grams of the element are present in one mole of the compound. Then divide this value by the molar mass of the compound and multiply by 100.

$$\frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100 = \begin{array}{c}\text{\% element in}\\ \text{compound}\end{array}$$

The percentage by mass of each element in a compound is known as the **percentage composition** *of the compound.*

Find the percentage composition of copper(I) sulfide, Cu_2S.

SOLUTION

1 ANALYZE   **Given:** formula, Cu_2S
Unknown: percentage composition of Cu_2S

2 PLAN   formula $\longrightarrow$ molar mass $\longrightarrow$ mass percentage of each element

The molar mass of the compound must be found. Then the mass of each element present in one mole of the compound is used to calculate the mass percentage of each element.

3 COMPUTE

$$2 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{\text{mol Cu}} = 127.1 \text{ g Cu}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{\text{mol S}} = 32.07 \text{ g S}$$

$$\text{molar mass of } Cu_2S = 159.2 \text{ g}$$

$$\frac{127.1 \text{ g Cu}}{159.2 \text{ g } Cu_2S} \times 100 = 79.85\% \text{ Cu}$$

$$\frac{32.07 \text{ g S}}{159.2 \text{ g } Cu_2S} \times 100 = 20.15\% \text{ S}$$

4 EVALUATE   A good check is to see if the results add up to about 100%. (Because of rounding, the total may not always be exactly 100%.)

As some salts crystallize from a water solution, they bind water molecules in their crystal structure. Sodium carbonate forms such a *hydrate,* in which 10 water molecules are present for every formula unit of sodium carbonate. Find the mass percentage of water in sodium carbonate decahydrate, $Na_2CO_3 \cdot 10H_2O$, which has a molar mass of 286.19 g/mol.

SOLUTION

1 ANALYZE

Given: chemical formula, $Na_2CO_3 \cdot 10H_2O$
molar mass of $Na_2CO_3 \cdot 10H_2O$
Unknown: mass percentage of H_2O

2 PLAN

chemical formula $\longrightarrow$ mass H_2O per mole of $Na_2CO_3 \cdot 10H_2O \longrightarrow$ % water

The mass of water per mole of sodium carbonate decahydrate must first be found. This value is then divided by the mass of one mole of $Na_2CO_3 \cdot 10H_2O$.

3 COMPUTE

One mole of $Na_2CO_3 \cdot 10H_2O$ contains 10 mol H_2O. Recall from page 239 that the molar mass of H_2O is 18.02 g/mol. The mass of 10 mol H_2O is calculated as follows.

$$10 \text{ mol } H_2O \times \frac{18.02 \text{ g } H_2O}{\text{mol } H_2O} = 180.2 \text{ g } H_2O$$

mass of H_2O per mole of $Na_2CO_3 \cdot 10H_2O = 180.2$ g

The molar mass of $Na_2CO_3 \cdot 10H_2O$ is 286.19 g/mol, so we know that 1 mol of the hydrate has a mass of 286.19g. The mass percentage of 10 mol H_2O in 1 mol $Na_2CO_3 \cdot 10H_2O$ can now be calculated.

$$\text{mass percentage of } H_2O \text{ in } Na_2CO_3 \cdot 10H_2O = \frac{180.2 \text{ g } H_2O}{286.19 \text{ g } Na_2CO_3 \cdot 10H_2O} \times 100 = 62.97\% \ H_2O$$

4 EVALUATE

Checking shows that the arithmetic is correct and that units cancel as desired.

PRACTICE

Answers in Appendix E

1. Find the percentage compositions of the following:
   a. $PbCl_2$ **b.** $Ba(NO_3)_2$

2. Find the mass percentage of water in $ZnSO_4 \cdot 7H_2O$.

3. Magnesium hydroxide is 54.87% oxygen by mass. How many grams of oxygen are in 175 g of the compound? How many moles of oxygen is this?

SECTION REVIEW

1. Determine both the formula mass and molar mass of ammonium carbonate, $(NH_4)_2CO_3$.

2. How many moles of atoms of each element are there in one mole of $(NH_4)_2CO_3$?

3. What is the mass in grams of 3.25 mol $Fe_2(SO_4)_3$?

4. How many molecules of aspirin, $C_9H_8O_4$, are there in a 100.0 mg tablet of aspirin?

5. Calculate the percentage composition of $(NH_4)_2CO_3$.

Critical Thinking

6. **RELATING IDEAS** A sample of hydrated copper(II) sulfate ($CuSO_4 \cdot nH_2O$) is heated to 150°C and produces 103.74 g anhydrous copper(II) sulfate and 58.55 g water. How many moles of water molecules are present in 1.0 mol of hydrated copper(II) sulfate?

Determining Chemical Formulas

OBJECTIVES

- Define *empirical formula,* and explain how the term applies to ionic and molecular compounds.

- Determine an empirical formula from either a percentage or a mass composition.

- Explain the relationship between the empirical formula and the molecular formula of a given compound.

- Determine a molecular formula from an empirical formula.

When a new substance is synthesized or is discovered, it is analyzed quantitatively to reveal its percentage composition. From these data, the empirical formula is then determined. *An **empirical formula** consists of the symbols for the elements combined in a compound, with subscripts showing the smallest whole-number mole ratio of the different atoms in the compound.* For an ionic compound, the formula unit is usually the compound's empirical formula. For a molecular compound, however, the empirical formula does not necessarily indicate the actual numbers of atoms present in each molecule. For example, the empirical formula of the gas diborane is BH_3, but the molecular formula is B_2H_6. In this case, the number of atoms given by the molecular formula corresponds to the empirical ratio multiplied by two.

Calculation of Empirical Formulas

To determine a compound's empirical formula from its percentage composition, begin by converting percentage composition to a mass composition. Assume that you have a 100.0 g sample of the compound. Then calculate the amount of each element in the sample. For example, the percentage composition of diborane is 78.1% B and 21.9% H. Therefore, 100.0 g of diborane contains 78.1 g of B and 21.9 g of H.

Next, the mass composition of each element is converted to a composition in moles by dividing by the appropriate molar mass.

$$78.1 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g B}} = 7.22 \text{ mol B}$$

$$21.9 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 21.7 \text{ mol H}$$

These values give a mole ratio of 7.22 mol B to 21.7 mol H. However, this is not a ratio of smallest whole numbers. To find such a ratio, divide each number of moles by the smallest number in the existing ratio.

$$\frac{7.22 \text{ mol B}}{7.22} : \frac{21.7 \text{ mol H}}{7.22} = 1 \text{ mol B} : 3.01 \text{ mol H}$$

Because of rounding or experimental error, a compound's mole ratio sometimes consists of numbers close to whole numbers instead of exact whole numbers. In this case, the differences from whole numbers may be ignored and the nearest whole number taken. Thus, diborane contains atoms in the ratio 1 B:3 H. The compound's empirical formula is BH_3.

Sometimes mass composition is known instead of percentage composition. To determine the empirical formula in this case, convert mass composition to composition in moles. Then calculate the smallest whole-number mole ratio of atoms. This process is shown in Sample Problem M.

Quantitative analysis shows that a compound contains 32.38% sodium, 22.65% sulfur, and 44.99% oxygen. Find the empirical formula of this compound.

SOLUTION

1 **ANALYZE**   **Given:** percentage composition: 32.38% Na, 22.65% S, and 44.99% O
Unknown: empirical formula

2 **PLAN**   percentage composition ⟶ mass composition ⟶ composition in moles
⟶ smallest whole-number mole ratio of atoms

3 **COMPUTE**   Mass composition (mass of each element in 100.0 g sample): 32.38 g Na, 22.65 g S, 44.99 g O

Composition in moles: $32.38 \text{ g Na} \times \dfrac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 1.408 \text{ mol Na}$

$22.65 \text{ g S} \times \dfrac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.7063 \text{ mol S}$

$44.99 \text{ g O} \times \dfrac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.812 \text{ mol O}$

Smallest whole-number mole ratio of atoms:
The compound contains atoms in the ratio 1.408 mol Na:0.7063 mol S:2.812 mol O. To find the smallest whole-number mole ratio, divide each value by the smallest number in the ratio.

$$\frac{1.408 \text{ mol Na}}{0.7063} : \frac{0.7063 \text{ mol S}}{0.7063} : \frac{2.812 \text{ mol O}}{0.7063} = 1.993 \text{ mol Na:1 mol S:3.981 mol O}$$

Rounding each number in the ratio to the nearest whole number yields a mole ratio of 2 mol Na:1 mol S:4 mol O. The empirical formula of the compound is Na_2SO_4.

4 **EVALUATE**   Calculating the percentage composition of the compound based on the empirical formula determined in the problem reveals a percentage composition of 32.37% Na, 22.58% S, and 45.05% O. These values agree reasonably well with the given percentage composition.

For more help, go to the Math Tutor at the end of Chapter 22.

Analysis of a 10.150 g sample of a compound known to contain only phosphorus and oxygen indicates a phosphorus content of 4.433 g. What is the empirical formula of this compound?

SOLUTION

1 ANALYZE

Given: sample mass = 10.150 g
         phosphorus mass = 4.433 g
Unknown: empirical formula

2 PLAN

Mass composition $\longrightarrow$ composition in moles $\longrightarrow$ smallest whole-number ratio of atoms

3 COMPUTE

The mass of oxygen is found by subtracting the phosphorus mass from the sample mass.

$$\text{sample mass} - \text{phosphorus mass} = 10.150 \text{ g} - 4.433 \text{ g} = 5.717 \text{ g}$$

Mass composition: 4.433 g P, 5.717 g O

Composition in moles:

$$4.433 \text{ g P} \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.1431 \text{ mol P}$$

$$5.717 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.3573 \text{ mol O}$$

Smallest whole-number mole ratio of atoms:

$$\frac{0.1431 \text{ mol P}}{0.1431} : \frac{0.3573 \text{ mol O}}{0.1431}$$

$$1 \text{ mol P} : 2.497 \text{ mol O}$$

The number of O atoms is not close to a whole number. But if we multiply each number in the ratio by 2, then the number of O atoms becomes 4.994 mol, which is close to 5 mol. The simplest whole-number mole ratio of P atoms to O atoms is 2:5. The compound's empirical formula is P_2O_5.

4 EVALUATE

The arithmetic is correct, significant figures have been used correctly, and units cancel as desired. The formula is reasonable because +5 is a common oxidation state of phosphorus.

PRACTICE

Answers in Appendix E

1. A compound is found to contain 63.52% iron and 36.48% sulfur. Find its empirical formula.

2. Find the empirical formula of a compound found to contain 26.56% potassium, 35.41% chromium, and the remainder oxygen.

3. Analysis of 20.0 g of a compound containing only calcium and bromine indicates that 4.00 g of calcium are present. What is the empirical formula of the compound formed?

extension
Go to **go.hrw.com** for more practice problems that ask you to determine empirical formulas.

Keyword: HC6FRMX

Calculation of Molecular Formulas

Remember that the *empirical formula* contains the smallest possible whole numbers that describe the atomic ratio. The *molecular formula* is the actual formula of a molecular compound. An empirical formula may or may not be a correct molecular formula. For example, diborane's empirical formula is BH_3. Any multiple of BH_3, such as B_2H_6, B_3H_9, B_4H_{12}, and so on, represents the same ratio of B atoms to H atoms. The molecular compounds ethene, C_2H_4, and cyclopropane, C_3H_6, also share an identical atomic ratio (2 H:1 C), yet they are very different substances. How is the correct formula of a molecular compound found from an empirical formula?

The relationship between a compound's empirical formula and its molecular formula can be written as follows.

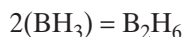$$x(\text{empirical formula}) = \text{molecular formula}$$

The number represented by x is a whole-number multiple indicating the factor by which the subscripts in the empirical formula must be multiplied to obtain the molecular formula. (The value of x is sometimes 1.) The formula masses have a similar relationship.

$$x(\text{empirical formula mass}) = \text{molecular formula mass}$$

To determine the molecular formula of a compound, you must know the compound's formula mass. For example, experimentation shows the formula mass of diborane to be 27.67 amu. The formula mass for the empirical formula, BH_3, is 13.84 amu. Dividing the experimental formula mass by the empirical formula mass gives the value of x for diborane.

$$x = \frac{27.67 \text{ amu}}{13.84 \text{ amu}} = 2.000$$

The molecular formula of diborane is therefore B_2H_6.

$$2(BH_3) = B_2H_6$$

Recall that a compound's molecular formula mass is numerically equal to its molar mass, so a compound's molecular formula can also be found given the compound's empirical formula and its molar mass.

SAMPLE PROBLEM N   *For more help, go to the* **Math Tutor** *at the end of Chapter 22.*

In Sample Problem M, the empirical formula of a compound of phosphorus and oxygen was found to be P_2O_5. Experimentation shows that the molar mass of this compound is 283.89 g/mol. What is the compound's molecular formula?

SOLUTION

1 *ANALYZE*   **Given:** empirical formula
Unknown: molecular formula

2 PLAN

$$x(\text{empirical formula}) = \text{molecular formula}$$

$$x = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$$

3 COMPUTE

Molecular formula mass is numerically equal to molar mass. Thus, changing the g/mol unit of the compound's molar mass to amu yields the compound's molecular formula mass.

$$\text{molecular molar mass} = 283.89 \text{ g/mol}$$
$$\text{molecular formula mass} = 283.89 \text{ amu}$$

The empirical formula mass is found by adding the masses of each of the atoms indicated in the empirical formula.

$$\text{mass of phosphorus atom} = 30.97 \text{ amu}$$
$$\text{mass of oxygen atom} = 16.00 \text{ amu}$$
$$\text{empirical formula mass of } P_2O_5 = 2 \times 30.97 \text{ amu} + 5 \times 16.00 \text{ amu} = 141.94 \text{ amu}$$

Dividing the experimental formula mass by the empirical formula mass gives the value of x. The formula mass is numerically equal to the molar mass.

$$x = \frac{283.89 \text{ amu}}{141.94 \text{ amu}} = 2.0001$$

The compound's molecular formula is therefore P_4O_{10}.

$$2 \times (P_2O_5) = P_4O_{10}$$
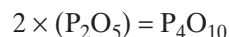
4 EVALUATE

Checking the arithmetic shows that it is correct.

PRACTICE

Answers in Appendix E

1. Determine the molecular formula of the compound with an empirical formula of CH and a formula mass of 78.110 amu.

2. A sample of a compound with a formula mass of 34.00 amu is found to consist of 0.44 g H and 6.92 g O. Find its molecular formula.

extension
Go to **go.hrw.com** for more practice problems that ask you to determine molecular formulas.

Keyword: HC6FRMX

SECTION REVIEW

1. A compound contains 36.48% Na, 25.41% S, and 38.11% O. Find its empirical formula.

2. Find the empirical formula of a compound that contains 53.70% iron and 46.30% sulfur.

3. Analysis of a compound indicates that it contains 1.04 g K, 0.70 g Cr, and 0.86 g O. Find its empirical formula.

4. If 4.04 g of N combine with 11.46 g O to produce a compound with a formula mass of 108.0 amu, what is the molecular formula of this compound?

Critical Thinking

5. **RELATING IDEAS** A compound containing sodium, chlorine, and oxygen is 25.42% sodium by mass. A 3.25 g sample gives 4.33×10^{22} atoms of oxygen. What is the empirical formula?

CHAPTER HIGHLIGHTS

Chemical Names and Formulas

Vocabulary
monatomic ion
binary compound
nomenclature
oxyanion
salt

- A positive monatomic ion is identified simply by the name of the appropriate element. A negative monatomic ion is named by dropping parts of the ending of the element's name and adding *-ide* to the root.
- The charge of each ion in an ionic compound may be used to determine the simplest chemical formula for the compound.
- Binary compounds are composed of two elements.
- Binary ionic compounds are named by combining the names of the positive and negative ions.
- The old system of naming binary molecular compounds uses prefixes. The new system, known as the *Stock system,* uses oxidation numbers.

Oxidation Numbers

Vocabulary
oxidation number
oxidation state

- Oxidation numbers are useful in naming compounds, in writing formulas, and in balancing chemical equations.
- Compounds containing elements that have more than one oxidation state are named by using the Stock system.
- Stock-system names and prefix-system names are used interchangeably for many molecular compounds.
- Oxidation numbers of each element in a compound may be used to determine the compound's simplest chemical formula.
- By knowing oxidation numbers, we can name compounds without knowing whether they are ionic or molecular.

Using Chemical Formulas

Vocabulary
formula mass
percentage composition

- Formula mass, molar mass, and percentage composition can be calculated from the chemical formula for a compound.
- The percentage composition of a compound is the percentage by mass of each element in the compound.
- Molar mass is used as a conversion factor between amount in moles and mass in grams of a given compound or element.

Determining Chemical Formulas

Vocabulary
empirical formula

- An empirical formula shows the simplest whole-number ratio of atoms in a given compound.
- Empirical formulas indicate how many atoms of each element are combined in the simplest unit of a chemical compound.
- A molecular formula can be found from the empirical formula if the molar mass is measured.

Chemical Names and Formulas

SECTION 1 REVIEW

1. a. What are monatomic ions?
 b. Give three examples of monatomic ions.

2. How does the chemical formula for the nitrite ion differ from the chemical formula for the nitrate ion?

3. Using only the periodic table, write the symbol of the ion most typically formed by each of the following elements:
 a. K d. Cl
 b. Ca e. Ba
 c. S f. Br

4. Write the formula for and indicate the charge on each of the following ions:
 a. sodium ion
 b. aluminum ion
 c. chloride ion
 d. nitride ion
 e. iron(II) ion
 f. iron(III) ion

5. Name each of the following monatomic ions:
 a. K^+ d. Cl^-
 b. Mg^{2+} e. O^{2-}
 c. Al^{3+} f. Ca^{2+}

6. Write formulas for the binary ionic compounds formed between the following elements. (Hint: See Sample Problem A.)
 a. sodium and iodine
 b. calcium and sulfur
 c. zinc and chlorine
 d. barium and fluorine
 e. lithium and oxygen

7. Give the name of each of the following binary ionic compounds. (Hint: See Sample Problem B.)
 a. KCl c. Li_2O
 b. $CaBr_2$ d. $MgCl_2$

8. Write the formulas for and give the names of the compounds formed by the following ions:
 a. Cr^{2+} and F^-
 b. Ni^{2+} and O^{2-}
 c. Fe^{3+} and O^{2-}

9. What determines the order in which the component elements of binary molecular compounds are written?

10. Name the following binary molecular compounds according to the prefix system. (Hint: See Sample Problem D.)
 a. CO_2 d. SeF_6
 b. CCl_4 e. As_2O_5
 c. PCl_5

11. Write formulas for each of the following binary molecular compounds. (Hint: See Sample Problem D.)
 a. carbon tetrabromide
 b. silicon dioxide
 c. tetraphosphorus decoxide
 d. diarsenic trisulfide

12. Distinguish between binary acids and oxyacids, and give two examples of each.

13. a. What is a salt?
 b. Give two examples of salts.

14. Name each of the following acids:
 a. HF d. H_2SO_4
 b. HBr e. H_3PO_4
 c. HNO_3

15. Give the molecular formula for each of the following acids:
 a. sulfurous acid
 b. chloric acid
 c. hydrochloric acid
 d. hypochlorous acid
 e. perchloric acid
 f. carbonic acid
 g. acetic acid

PRACTICE PROBLEMS

16. Write formulas for each of the following compounds:
 a. sodium fluoride
 b. calcium oxide
 c. potassium sulfide
 d. magnesium chloride
 e. aluminum bromide
 f. lithium nitride
 g. iron(II) oxide

17. Name each of the following ions:
   a. NH_4^+
   b. ClO_3^-
   c. OH^-
   d. SO_4^{2-}
   e. NO_3^-
   f. CO_3^{2-}
   g. PO_4^{3-}
   h. CH_3COO^-
   i. HCO_3^-
   j. CrO_4^{2-}

18. Write the formula and charge for each of the following ions:
   a. ammonium ion
   b. acetate ion
   c. hydroxide ion
   d. carbonate ion
   e. sulfate ion
   f. phosphate ion
   g. copper(II) ion
   h. tin(II) ion
   i. iron(III) ion
   j. copper(I) ion
   k. mercury(I) ion
   l. mercury(II) ion

Oxidation Numbers

SECTION 2 REVIEW

19. Name each of the following ions according to the Stock system:
   a. Fe^{2+}
   b. Fe^{3+}
   c. Pb^{2+}
   d. Pb^{4+}
   e. Sn^{2+}
   f. Sn^{4+}

20. Name each of the binary molecular compounds in item 11 by using the Stock system.

21. Write formulas for each of the following compounds:
   a. phosphorus(III) iodide
   b. sulfur(II) chloride
   c. carbon(IV) sulfide
   d. nitrogen(V) oxide

22. a. What are oxidation numbers?
   b. What useful functions do oxidation numbers serve?

PRACTICE PROBLEMS

23. Name each of the following ionic compounds by using the Stock system:
   a. NaCl
   b. KF
   c. CaS
   d. $Co(NO_3)_2$
   e. $FePO_4$
   f. Hg_2SO_4
   g. $Hg_3(PO_4)_2$

24. Assign oxidation numbers to each atom in the following compounds. (Hint: See Sample Problem E.)
   a. HI
   b. PBr_3
   c. GeS_2
   d. KH
   e. As_2O_5
   f. H_3PO_4

25. Assign oxidation numbers to each atom in the following ions. (Hint: See Sample Problem E.)
   a. NO_3^-
   b. ClO_4^-
   c. PO_4^{3-}
   d. $Cr_2O_7^{2-}$
   e. CO_3^{2-}

Using Chemical Formulas

SECTION 3 REVIEW

26. a. Define *formula mass*.
   b. In what unit is formula mass expressed?

27. What is meant by the molar mass of a compound?

PRACTICE PROBLEMS

28. Determine the formula mass of each of the following compounds or ions. (Hint: See Sample Problem F.)
   a. glucose, $C_6H_{12}O_6$
   b. calcium acetate, $Ca(CH_3COO)_2$
   c. the ammonium ion, NH_4^+
   d. the chlorate ion, ClO_3^-

29. Determine the number of moles of each type of monatomic or polyatomic ion in one mole of the following compounds. For each polyatomic ion, determine the number of moles of each atom present in one mole of the ion.
   a. KNO_3
   b. Na_2SO_4
   c. $Ca(OH)_2$
   d. $(NH_4)_2SO_3$
   e. $Ca_3(PO_4)_2$
   f. $Al_2(CrO_4)_3$

30. Determine the molar mass of each compound listed in item 29. (Hint: See Sample Problem G.)

31. Determine the number of moles of compound in each of the following samples. (Hint: See Sample Problem I.)
a. 4.50 g H_2O
b. 471.6 g $Ba(OH)_2$
c. 129.68 g $Fe_3(PO_4)_2$

32. Determine the percentage composition of each of the following compounds. (Hint: See Sample Problem J.)
a. NaCl
b. $AgNO_3$
c. $Mg(OH)_2$

33. Determine the percentage by mass of water in the hydrate $CuSO_4 \cdot 5H_2O$. (Hint: See Sample Problem K.)

Determining Chemical Formulas

SECTION 4 REVIEW

34. What three types of information are used to find an empirical formula from percentage composition data?

35. What is the relationship between the empirical formula and the molecular formula of a compound?

PRACTICE PROBLEMS

36. Determine the empirical formula of a compound containing 63.50% silver, 8.25% nitrogen, and 28.25% oxygen. (Hint: See Sample Problem L.)

37. Determine the empirical formula of a compound found to contain 52.11% carbon, 13.14% hydrogen, and 34.75% oxygen.

38. What is the molecular formula of the molecule that has an empirical formula of CH_2O and a molar mass of 120.12 g/mol?

39. A compound with a formula mass of 42.08 amu is found to be 85.64% carbon and 14.36% hydrogen by mass. Find its molecular formula.

MIXED REVIEW

40. Chemical analysis shows that citric acid contains 37.51% C, 4.20% H, and 58.29% O. What is the empirical formula for citric acid?

41. Name each of the following compounds by using the Stock system:
a. LiBr
b. $Sn(NO_3)_2$
c. $FeCl_2$
d. MgO
e. KOH
f. Fe_2O_3
g. $AgNO_3$
h. $Fe(OH)_2$
i. CrF_2

42. What is the mass in grams of each of the following samples?
a. 1.000 mol NaCl
b. 2.000 mol H_2O
c. 3.500 mol $Ca(OH)_2$
d. 0.625 mol $Ba(NO_3)_2$

43. Determine the formula mass and molar mass of each of the following compounds:
a. XeF_4
b. $C_{12}H_{24}O_6$
c. Hg_2I_2
d. CuCN

44. Write the chemical formulas for the following compounds:
a. aluminum fluoride
b. magnesium oxide
c. vanadium(V) oxide
d. cobalt(II) sulfide
e. strontium bromide
f. sulfur trioxide

45. How many atoms of each element are contained in a single formula unit of iron(III) formate, $Fe(CHO_2)_3 \cdot H_2O$? What percentage by mass of the compound is water?

46. Name each of the following acids, and assign oxidation numbers to the atoms in each:
a. HNO_2
b. H_2SO_3
c. H_2CO_3
d. HI

47. Determine the percentage composition of the following compounds:
a. NaClO
b. H_2SO_3
c. C_2H_5COOH
d. $BeCl_2$

CHEMICAL FORMULAS AND CHEMICAL COMPOUNDS **253**

48. Name each of the following binary compounds:
a. MgI_2
b. NaF
c. CS_2
d. N_2O_4
e. SO_2
f. PBr_3
g. $CaCl_2$
h. AgI

49. Assign oxidation numbers to each atom in the following molecules and ions:
a. CO_2
b. NH_4^+
c. MnO_4^-
d. $S_2O_3^{2-}$
e. H_2O_2
f. P_4O_{10}
g. OF_2

50. A 175.0 g sample of a compound contains 56.15 g C, 9.43 g H, 74.81 g O, 13.11 g N, and 21.49 g Na. What is the compound's empirical formula?

CRITICAL THINKING

51. Analyzing Information Sulfur trioxide is produced in the atmosphere through a reaction of sulfur dioxide and oxygen. Sulfur dioxide is a primary air pollutant. Analyze the formula for sulfur trioxide. Then, list all of the chemical information from the analysis that you can.

52. Analyzing Data In the laboratory, a sample of pure nickel was placed in a clean, dry, weighed crucible. The crucible was heated so that the nickel would react with the oxygen in the air. After the reaction appeared complete, the crucible was allowed to cool and the mass was determined. The crucible was reheated and allowed to cool. Its mass was then determined again to be certain that the reaction was complete. The following data were collected:

Mass of crucible	= 30.02 g
Mass of nickel and crucible	= 31.07 g
Mass of nickel oxide and crucible	= 31.36 g

Determine the following information based on the data given above:

Mass of nickel	=
Mass of nickel oxide	=
Mass of oxygen	=

Based on your calculations, what is the empirical formula for the nickel oxide?

USING THE HANDBOOK

53. Review the common reactions of Group 1 metals in the *Elements Handbook,* and answer the following questions:
a. Some of the Group 1 metals react with oxygen to form superoxides. Write the formulas for these compounds.
b. What is the charge on each cation for the formulas that you wrote in (a)?
c. How does the charge on the anion vary for oxides, peroxides, and superoxides?

54. Review the common reactions of Group 2 metals in the *Elements Handbook,* and answer the following questions:
a. Some of the Group 2 metals react with oxygen to form oxides. Write the formulas for these compounds.
b. Some of the Group 2 metals react with oxygen to form peroxides. Write the formulas for these compounds.
c. Some of the Group 2 metals react with nitrogen to form nitrides. Write the formulas for these compounds.
d. Most Group 2 elements form hydrides. What is hydrogen's oxidation state in these compounds?

55. Review the analytical tests for transition metals in the *Elements Handbook,* and answer the following questions:
a. Determine the oxidation state of each metal in the precipitates shown for cadmium, zinc, and lead.
b. Determine the oxidation state of each metal in the complex ions shown for iron, manganese, and cobalt.
c. The copper compound shown is called a *coordination compound.* The ammonia shown in the formula exists as molecules that do not have a charge. Determine copper's oxidation state in this compound.

56. Review the common reactions of Group 15 elements in the *Elements Handbook*, and answer the following questions:
   a. Write formulas for each of the oxides listed for the Group 15 elements.
   b. Determine nitrogen's oxidation state in the oxides listed in (a).

RESEARCH & WRITING

57. Nomenclature Biologists who name newly discovered organisms use a system that is structured very much like the one used by chemists in naming compounds. The system used by biologists is called the *Linnaeus system,* after its creator, Carolus Linnaeus. Research this system in a biology textbook, and then note similarities and differences between the Linnaeus system and chemical nomenclature.

58. Common Chemicals Find out the systematic chemical name and write the chemical formula for each of the following common compounds:
   a. baking soda   d. limestone
   b. milk of magnesia   e. lye
   c. Epsom salts   f. wood alcohol

ALTERNATIVE ASSESSMENT

59. Performance Assessment Your teacher will supply you with a note card that has one of the following formulas on it: $NaCH_3COO \cdot 3H_2O$, $MgCl_2 \cdot 6H_2O$, $LiC_2H_3O_2 \cdot 2H_2O$, or $MgSO_4 \cdot 7H_2O$. Design an experiment to determine the percentage of water by mass in the hydrated salt assigned to you. Be sure to explain what steps you will take to ensure that the salt is completely dry. If your teacher approves your design, obtain the salt and perform the experiment. What percentage of water does the salt contain?

60. Both ammonia, NH_3, and ammonium nitrate, NH_4NO_3, are used in fertilizers as a source of nitrogen. Which compound has the higher percentage of nitrogen? Research the physical properties of both compounds, and find out how each compound is manufactured and used. Explain why each compound has its own particular application. (Consider factors such as the cost of raw ingredients, the ease of manufacture, and shipping costs.)

extension

Graphing Calculator Calculating Molar Mass

Go to **go.hrw.com** for a graphing calculator exercise that asks you to calculate the molar mass of a compound.

Keyword: HC6FRMX

Math Tutor CALCULATING PERCENTAGE COMPOSITION

Chemists can analyze an unknown substance by determining its percentage composition by mass. Percentage composition is determined by finding the mass of each element in a sample of the substance as a percentage of the mass of the whole sample. The results of this analysis can then be compared with the percentage composition of known compounds to determine the probable identity of the unknown substance. Once you know a compound's formula, you can determine its percentage composition by mass.

SAMPLE 1

Determine the percentage composition of potassium chlorate, $KClO_3$.

First, calculate the molar mass of $KClO_3$. The formula shows you that one mole of $KClO_3$ consists of 1 mol K atoms, 1 mol Cl atoms, and 3 mol O atoms. Thus, the molar mass of $KClO_3$ is molar mass K + molar mass Cl + 3(molar mass O) = 39.10 g K + 35.45 g Cl + 3(16.00 g O).

$$molar\ mass\ KClO_3 = 122.55\ g$$

The percentage composition of $KClO_3$ is determined by calculating the percentage of the total molar mass contributed by each element.

$$\frac{mass\ of\ element\ in\ 1\ mol\ of\ compound}{molar\ mass\ of\ compound} \times 100$$

$$= \%\ element\ in\ compound$$

$$\%\ K\ in\ KClO_3 = \frac{39.10\ g\ K}{122.55\ g\ KClO_3} \times 100 = 31.91\%$$

$$\%\ Cl\ in\ KClO_3 = \frac{35.45\ g\ Cl}{122.55\ g\ KClO_3} \times 100 = 28.93\%$$

$$\%\ O\ in\ KClO_3 = \frac{48.00\ g\ O}{122.55\ g\ KClO_3} \times 100 = 39.17\%$$

SAMPLE 2

Determine the percentage of nitrogen in ammonium sulfate, $(NH_4)_2SO_4$.

Even though you want to find the percentage of only one element, you must calculate the molar mass of $(NH_4)_2SO_4$. To do that, examine the formula to find the number of moles of each element in the compound. The two ammonium groups, indicated by $(NH_4)_2$, contain 2 mol N and 8 mol H per mole of $(NH_4)_2SO_4$. The sulfate group, SO_4, contains 1 mol S and 4 mol O per mole of $(NH_4)_2SO_4$.

$$
\begin{aligned}
2\ mol\ N &= 2 \times 14.01\ g = & 28.02\ g \\
8\ mol\ H &= 8 \times 1.01\ g = & 8.08\ g \\
1\ mol\ S &= 1 \times 32.07 = & 32.07\ g \\
4\ mol\ O &= 4 \times 16.00 = & 64.00\ g \\
\end{aligned}
$$
$$molar\ mass\ (NH_4)_2SO_4 = \overline{132.17\ g}$$

Now, you can determine the percentage of nitrogen in the compound as follows.

$$\%\ N\ in\ (NH_4)_2SO_4 = \frac{28.02\ g\ N}{132.17\ g\ (NH_4)_2SO_4}$$

$$\times 100 = 21.20\%$$

PRACTICE PROBLEMS

1. What is the percentage composition of sodium carbonate, Na_2CO_3?

2. What is the percentage of iodine in zinc iodate, $Zn(IO_3)_2$?

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Which of the following compounds does not contain a polyatomic ion?
   A. sodium carbonate
   B. sodium sulfate
   C. sodium sulfite
   D. sodium sulfide

2. The correct formula for ammonium phosphate is
   A. $(NH_4)_3PO_4$.
   B. $(NH_4)_2PO_4$.
   C. NH_4PO_4.
   D. $NH_4(PO_4)_2$.

3. When writing the formula for a compound that contains a polyatomic ion,
   A. write the anion's formula first.
   B. use superscripts to show the number of polyatomic ions present.
   C. use parentheses if the number of polyatomic ions is greater than 1.
   D. always place the polyatomic ion in parentheses.

4. The correct name for NH_4CH_3COO is
   A. ammonium carbonate.
   B. ammonium hydroxide.
   C. ammonium acetate.
   D. ammonium nitrate.

5. Which of the following is the correct formula for iron(III) sulfate?
   A. Fe_3SO_4
   B. $Fe_3(SO_4)_2$
   C. $Fe_2(SO_4)_3$
   D. $3FeSO_4$

6. The molecular formula for acetylene is C_2H_2. The molecular formula for benzene is C_6H_6. The empirical formula for both is
   A. CH.
   B. C_2H_2.
   C. C_6H_6.
   D. $(CH)_2$.

7. Which of the following shows the percentage composition of H_2SO_4?
   A. 2.5% H, 39.1% S, 58.5% O
   B. 2.1% H, 32.7% S, 65.2% O
   C. 28.6% H, 14.3% S, 57.1% O
   D. 33.3% H, 16.7% S, 50% O

8. Which of the following compounds has the highest percentage of oxygen?
   A. CH_4O
   B. CO_2
   C. H_2O
   D. Na_2CO_3

9. The empirical formula for a compound that is 1.2% H, 42.0% Cl, and 56.8% O is
   A. HClO.
   B. $HClO_2$.
   C. $HClO_3$.
   D. $HClO_4$.

SHORT ANSWER

10. When a new substance is synthesized or is discovered experimentally, the substance is analyzed quantitatively. What information is obtained from this typical analysis, and how is this information used?

11. An oxide of selenium is 28.8% O. Find the empirical formula. Assuming that the empirical formula is also the molecular formula, name the oxide.

EXTENDED RESPONSE

12. What is an empirical formula, and how does it differ from a molecular formula?

13. What are Stock system names based on?

Test TIP Whenever possible, highlight or underline numbers or words critical to answering a question correctly.

Determining the Empirical Formula of Magnesium Oxide

OBJECTIVES

- *Measure* the mass of magnesium oxide.

- *Perform* a synthesis reaction by using gravimetric techniques.

- *Determine* the empirical formula of magnesium oxide.

- *Calculate* the class average and standard deviation for moles of oxygen used.

MATERIALS

- **15 cm magnesium ribbon, 2**

- **25 mL beaker**

- **Bunsen burner assembly**

- **clay triangle**

- **crucible and lid, metal or ceramic**

- **crucible tongs**

- **distilled water**

- **eyedropper or micropipet**

- **ring stand**

FIGURE A

BACKGROUND

This gravimetric analysis involves the combustion of magnesium metal in air to synthesize magnesium oxide. The mass of the product is greater than the mass of magnesium used because oxygen reacts with the magnesium metal. As in all gravimetric analyses, success depends on attaining a product yield near 100%. Therefore, the product will be heated and cooled and have its mass measured until two of these mass measurements are within 0.02% of one another. When the masses of the reactant and product have been carefully measured, the amount of oxygen used in the reaction can be calculated. The ratio of oxygen to magnesium can then be established, and the empirical formula of magnesium oxide can be determined.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Copy the following data table in your lab notebook.

DATA TABLE		
	Trial 1	Trial 2
1. Mass of crucible, lid, and metal (g)		
2. Mass of crucible, lid, and product (g)		
3. Mass of crucible and lid (g)		

PROCEDURE

1. Construct a setup for heating a crucible as shown in Figure A and as demonstrated in the Pre-Laboratory Procedure "Gravimetric Analysis."

2. Heat the crucible and lid for 5 min to burn off any impurities.

3. Cool the crucible and lid to room temperature. Measure their combined mass, and record the measurement on line 3 of your data table.

   NOTE: Handle the crucible and lid with crucible tongs at all times during this experiment. Such handling prevents burns and the transfer of dirt and oil from your hands to the crucible and lid.

4. Polish a 15 cm strip of magnesium with steel wool. The magnesium should be shiny. Cut the strip into small pieces to make the reaction proceed faster, and place the pieces in the crucible.

5. Cover the crucible with the lid, and measure the mass of the crucible, lid, and metal. Record the measurement on line 1 of your data table.

6. Use tongs to replace the crucible on the clay triangle. Heat the covered crucible gently. Lift the lid occasionally to allow air in.

   CAUTION: Do not look directly at the burning magnesium metal. The brightness of the light can blind you.

7. When the magnesium appears to be fully reacted, partially remove the crucible lid and continue heating for 1 min.

8. Remove the burner from under the crucible. After the crucible has cooled, use an eyedropper to carefully add a few drops of water to decompose any nitrides that may have formed.

   CAUTION: Use care when adding water. Using too much water can cause the crucible to crack.

9. Cover the crucible completely. Replace the burner under the crucible, and continue heating for about 30 to 60 s.

10. Turn off the burner. Cool the crucible, lid, and contents to room temperature. Measure the mass of the crucible, lid, and product. Record the measurement in the margin of your data table.

11. Replace the crucible, lid, and contents on the clay triangle, and reheat for another 2 min. Cool to room temperature, and remeasure the mass of the crucible, lid, and contents. Compare this mass measurement with the measurement obtained in step 10. If the new mass is ±0.02% of the mass in step 10, record the new mass on line 2 of your data table and go on to step 12. If not, your reaction is still incomplete, and you should repeat step 11.

12. Clean the crucible, and repeat steps 2–11 with a second strip of magnesium ribbon. Record your measurements under Trial 2 in your data table.

CLEANUP AND DISPOSAL

13. Put the solid magnesium oxide in the designated waste container. Return any unused magnesium ribbon to your teacher. Clean your equipment and lab station. Thoroughly wash your hands after completing the lab session and cleanup.

ANALYSIS AND INTERPRETATION

1. **Applying Ideas:** Calculate the mass of the magnesium metal and the mass of the product.

2. **Evaluating Data:** Determine the mass of the oxygen consumed.

3. **Applying Ideas:** Calculate the number of moles of magnesium and the number of moles of oxygen in the product.

CONCLUSIONS

1. **Inferring Relationships:** Determine the empirical formula for magnesium oxide, Mg_xO_y.