



# Electrons in Atoms

## What You'll Learn

- ▶ You will compare the wave and particle models of light.
- ▶ You will describe how the frequency of light emitted by an atom is a unique characteristic of that atom.
- ▶ You will compare and contrast the Bohr and quantum mechanical models of the atom.
- ▶ You will express the arrangements of electrons in atoms through orbital notations, electron configurations, and electron dot structures.

## Why It's Important

Why are some fireworks red, some white, and others blue? The key to understanding the chemical behavior of fireworks, and all matter, lies in understanding how electrons are arranged in atoms of each element.



Visit the Chemistry Web site at [chemistrymc.com](http://chemistrymc.com) to find links about electrons in atoms.

The colorful display from fireworks is due to changes in the electron configurations of atoms.





## DISCOVERY LAB



### Materials

a wrapped box from your instructor

### What's Inside?

It's your birthday, and there are many wrapped presents for you to open. Much of the fun is trying to figure out what's inside the package before you open it. In trying to determine the structure of the atom, chemists had a similar experience. How good are your skills of observation and deduction?

### Procedure

1. Obtain a wrapped box from your instructor.
2. Using as many observation methods as you can, and without unwrapping or opening the box, try to figure out what the object inside the box is.
3. Record the observations you make throughout this discovery process.

### Analysis

How were you able to determine things such as size, shape, number, and composition of the object in the box? What senses did you use to make your observations? Why is it hard to figure out what type of object is in the box without actually seeing it?

## Section

## 5.1

# Light and Quantized Energy

### Objectives

- **Compare** the wave and particle models of light.
- **Define** a quantum of energy and explain how it is related to an energy change of matter.
- **Contrast** continuous electromagnetic spectra and atomic emission spectra.

### Vocabulary

electromagnetic radiation  
wavelength  
frequency  
amplitude  
electromagnetic spectrum  
quantum  
Planck's constant  
photoelectric effect  
photon  
atomic emission spectrum

Although three subatomic particles had been discovered by the early-1900s, the quest to understand the atom and its structure had really just begun. That quest continues in this chapter, as scientists pursued an understanding of how electrons were arranged within atoms. Perform the **DISCOVERY LAB** on this page to better understand the difficulties scientists faced in researching the unseen atom.

## The Nuclear Atom and Unanswered Questions

As you learned in Chapter 4, Rutherford proposed that all of an atom's positive charge and virtually all of its mass are concentrated in a nucleus that is surrounded by fast-moving electrons. Although his nuclear model was a major scientific development, it lacked detail about how electrons occupy the space surrounding the nucleus. In this chapter, you will learn how electrons are arranged in an atom and how that arrangement plays a role in chemical behavior.

Many scientists in the early twentieth century found Rutherford's nuclear atomic model to be fundamentally incomplete. To physicists, the model did not explain how the atom's electrons are arranged in the space around the nucleus. Nor did it address the question of why the negatively charged electrons are not pulled into the atom's positively charged nucleus. Chemists found Rutherford's nuclear model lacking because it did not begin to account for the differences in chemical behavior among the various elements.



**Figure 5-1**

**a** Chlorine gas, shown here reacting vigorously with steel wool, reacts with many other atoms as well. **b** Argon gas fills the interior of this incandescent bulb. The nonreactive argon prevents the hot filament from oxidizing, thus extending the life of the bulb. **c** Solid potassium metal is submerged in oil to prevent it from reacting with air or water.

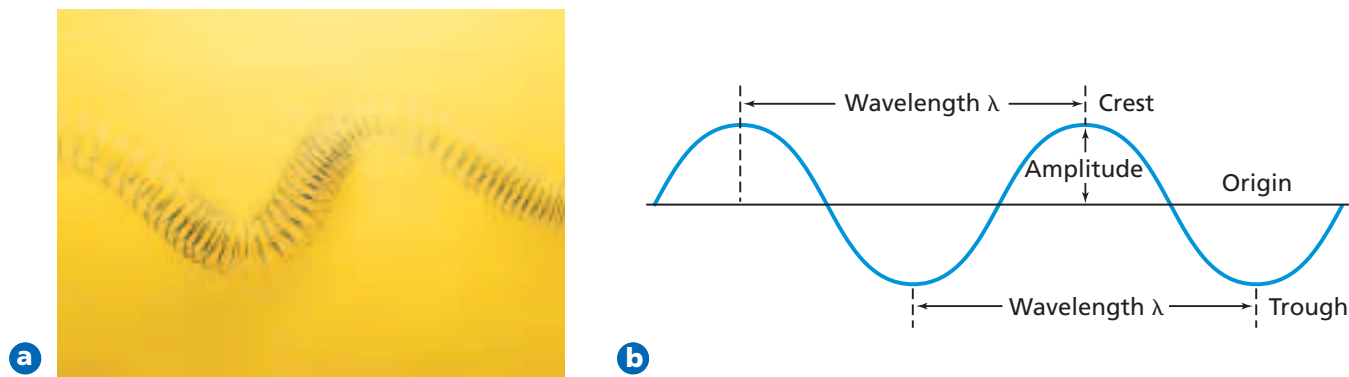
For example, consider the elements chlorine, argon, and potassium, which are found in consecutive order on the periodic table but have very different chemical behaviors. Atoms of chlorine, a yellow-green gas at room temperature, react readily with atoms of many other elements. **Figure 5-1a** shows chlorine atoms reacting with steel wool. The interaction of highly reactive chlorine atoms with the large surface area provided by the steel results in a vigorous reaction. Argon, which is used in the incandescent bulb shown in **Figure 5-1b**, also is a gas. Argon, however, is so unreactive that it is considered a noble gas. Potassium is a reactive metal at room temperature. In fact, as you can see in **Figure 5-1c**, because potassium is so reactive, it must be stored under kerosene or oil to prevent its atoms from reacting with the oxygen and water in the air. Rutherford's nuclear atomic model could not explain why atoms of these elements behave the way they do.

In the early 1900s, scientists began to unravel the puzzle of chemical behavior. They had observed that certain elements emitted visible light when heated in a flame. Analysis of the emitted light revealed that an element's chemical behavior is related to the arrangement of the electrons in its atoms. In order for you to better understand this relationship and the nature of atomic structure, it will be helpful for you to first understand the nature of light.

## Wave Nature of Light

**Electromagnetic radiation** is a form of energy that exhibits wavelike behavior as it travels through space. Visible light is a type of electromagnetic radiation. Other examples of electromagnetic radiation include visible light from the sun, microwaves that warm and cook your food, X rays that doctors and dentists use to examine bones and teeth, and waves that carry radio and television programs to your home.

All waves can be described by several characteristics, a few of which you may be familiar with from everyday experience. **Figure 5-2a** shows a standing wave created by rhythmically moving the free end of a spring toy. **Figure 5-2b** illustrates several primary characteristics of all waves, wavelength, frequency, amplitude, and speed. **Wavelength** (represented by  $\lambda$ , the Greek letter lambda) is the shortest distance between equivalent points on a continuous wave. For example, in **Figure 5-2b** the wavelength is measured from crest to crest or from trough to trough. Wavelength is usually expressed in meters, centimeters, or nanometers ( $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$ ). **Frequency** (represented by  $\nu$ , the Greek letter nu) is the number of waves that pass a given



**Figure 5-2**

**a** The standing wave produced with this spring toy displays properties that are characteristic of all waves. **b** The primary characteristics of waves are wavelength, frequency, amplitude, and speed. What is the wavelength of the wave in centimeters?

point per second. One hertz (Hz), the SI unit of frequency, equals one wave per second. In calculations, frequency is expressed with units of “waves per second,”  $\left(\frac{1}{s}\right)$  or  $(s^{-1})$ , where the term “waves” is understood. For example,

$$652 \text{ Hz} = 652 \text{ waves/second} = \frac{652}{s} = 652 \text{ s}^{-1}$$

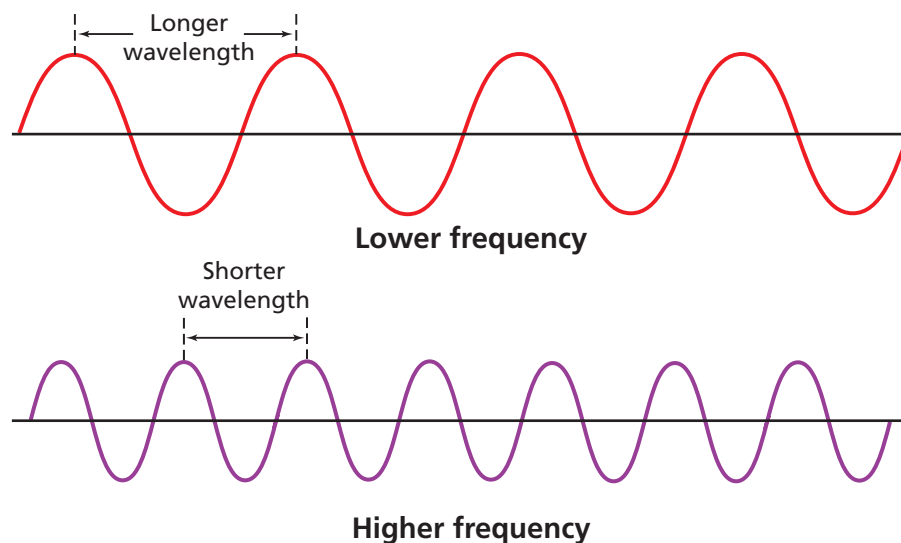
The **amplitude** of a wave is the wave’s height from the origin to a crest, or from the origin to a trough. To learn how lightwaves are able to form powerful laser beams, read the **How It Works** at the end of this chapter.

All electromagnetic waves, including visible light, travel at a speed of  $3.00 \times 10^8 \text{ m/s}$  in a vacuum. Because the speed of light is such an important and universal value, it is given its own symbol,  $c$ . The speed of light is the product of its wavelength ( $\lambda$ ) and its frequency ( $\nu$ ).

$$c = \lambda \nu$$

Although the speed of all electromagnetic waves is the same, waves may have different wavelengths and frequencies. As you can see from the equation above, wavelength and frequency are inversely related; in other words, as one quantity increases, the other decreases. To better understand this relationship, examine the red and violet light waves illustrated in **Figure 5-3**. Although both waves travel at the speed of light, you can see that red light has a longer wavelength and lower frequency than violet light.

Sunlight, which is one example of what is called white light, contains a continuous range of wavelengths and frequencies. Sunlight passing through a prism



**Figure 5-3**

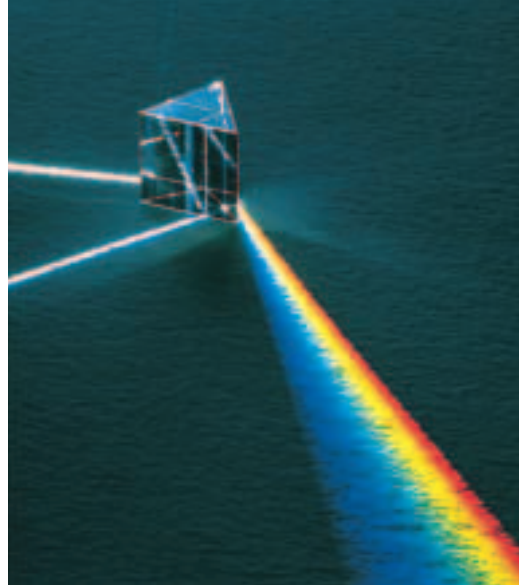
The inverse relationship between wavelength and frequency of electromagnetic waves can be seen in these red and violet waves. As wavelength increases, frequency decreases. Wavelength and frequency do not affect the amplitude of a wave. Which wave has the larger amplitude?





**Figure 5-4**

White light is separated into a continuous spectrum when it passes through a prism.



is separated into a continuous spectrum of colors. These are the colors of the visible spectrum. The spectrum is called continuous because there is no portion of it that does not correspond to a unique wavelength and frequency of light. You are already familiar with all of the colors of the visible spectrum from your everyday experiences. And if you have ever seen a rainbow, you have seen all of the visible colors at once. A rainbow is formed when tiny drops of water in the

air disperse the white light from the sun into its component colors, producing a continuous spectrum that arches across the sky.

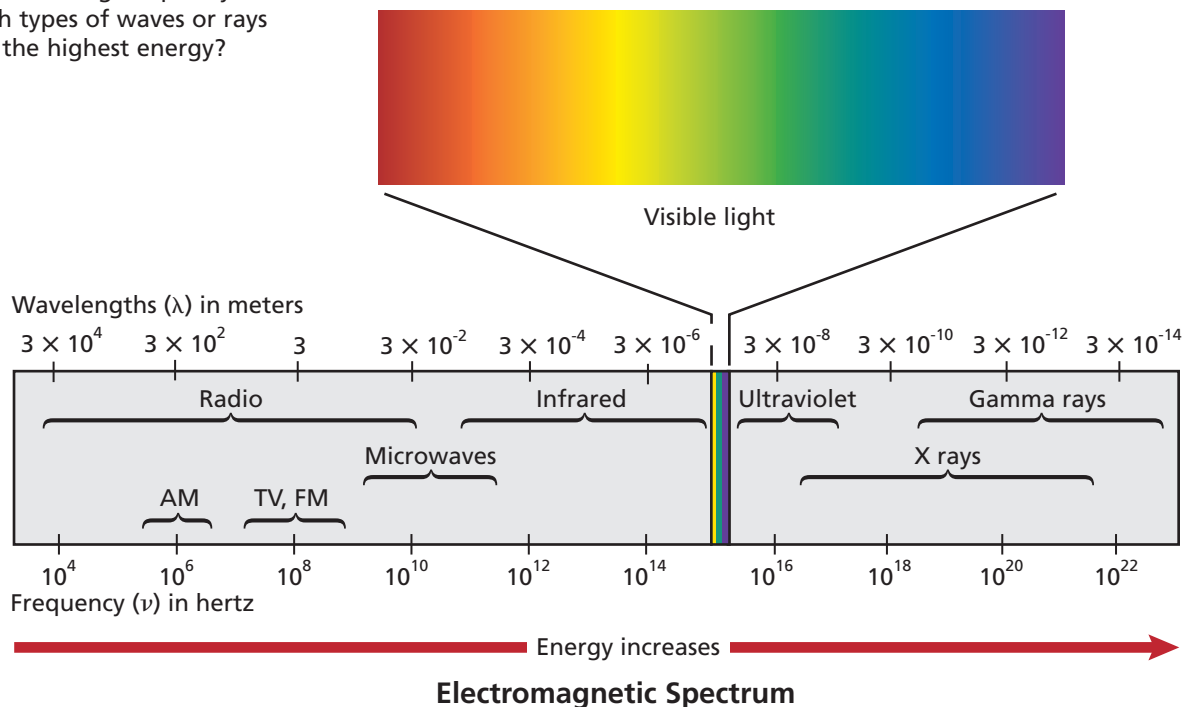
The visible spectrum of light shown in **Figure 5-4**, however, comprises only a small portion of the complete electromagnetic spectrum, which is illustrated in **Figure 5-5**. The **electromagnetic spectrum**, also called the EM spectrum, encompasses all forms of electromagnetic radiation, with the only differences in the types of radiation being their frequencies and wavelengths. Note in **Figure 5-4** that the short wavelengths bend more than long wavelengths as they pass through the prism, resulting in the sequence of colors red, orange, yellow, green, blue, indigo, and violet. This sequence can be remembered using the fictitious name Roy G. Biv as a memory aid. In examining the energy of the radiation shown in **Figure 5-5**, you should note that energy increases with increasing frequency. Thus, looking back at **Figure 5-3**, the violet light, with its greater frequency, has more energy than the red light. This relationship between frequency and energy will be explained in the next section.

**Try at Home LAB**

See page 954 in Appendix E for **Observing Light's Wave Nature**

**Figure 5-5**

The electromagnetic spectrum includes a wide range of wavelengths (and frequencies). Energy of the radiation increases with increasing frequency. Which types of waves or rays have the highest energy?





Because all electromagnetic waves travel at the same speed, you can use the formula  $c = \lambda\nu$  to calculate the wavelength or frequency of any wave. Example Problem 5-1 shows how this is done.

## EXAMPLE PROBLEM 5-1

### Calculating Wavelength of an EM Wave

Microwaves are used to transmit information. What is the wavelength of a microwave having a frequency of  $3.44 \times 10^9$  Hz?

#### 1. Analyze the Problem

You are given the frequency of a microwave. You also know that because microwaves are part of the electromagnetic spectrum, their speed, frequency, and wavelength are related by the formula  $c = \lambda\nu$ . The value of  $c$  is a known constant. First, solve the equation for wavelength, then substitute the known values and solve.

##### Known

$$\begin{aligned}\nu &= 3.44 \times 10^9 \text{ Hz} \\ c &= 3.00 \times 10^8 \text{ m/s}\end{aligned}$$

##### Unknown

$$\lambda = ? \text{ m}$$

#### 2. Solve for the Unknown

Solve the equation relating the speed, frequency, and wavelength of an electromagnetic wave for wavelength ( $\lambda$ ).

$$c = \lambda\nu$$

$$\lambda = c/\nu$$

Substitute  $c$  and the microwave's frequency,  $\nu$ , into the equation. Note that hertz is equivalent to  $1/\text{s}$  or  $\text{s}^{-1}$ .

$$\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \text{ s}^{-1}}$$

Divide the values to determine wavelength,  $\lambda$ , and cancel units as required.

$$\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \cancel{\text{s}^{-1}}} = 8.72 \times 10^{-2} \text{ m}$$

#### 3. Evaluate the Answer

The answer is correctly expressed in a unit of wavelength (m). Both of the known values in the problem are expressed with three significant figures, so the answer should have three significant figures, which it does. The value for the wavelength is within the wavelength range for microwaves shown in Figure 5-5.



Microwave relay antennas are used to transmit voice and data from one area to another without the use of wires or cables.

## PRACTICE PROBLEMS

1. What is the frequency of green light, which has a wavelength of  $4.90 \times 10^{-7}$  m?
2. An X ray has a wavelength of  $1.15 \times 10^{-10}$  m. What is its frequency?
3. What is the speed of an electromagnetic wave that has a frequency of  $7.8 \times 10^6$  Hz?
4. A popular radio station broadcasts with a frequency of 94.7 MHz. What is the wavelength of the broadcast? (1 MHz =  $10^6$  Hz)



For more practice with speed, frequency, and wavelength problems, go to **Supplemental Practice Problems** in Appendix A.



**Topic: Forensics**

To learn more about chemical analytical techniques that use light, visit the Chemistry Web site at [chemistrymc.com](http://chemistrymc.com)

**Activity:** Research the roles that infrared and ultraviolet light have played in forensic investigations. Share with the class a particular case that was solved with these tools.

## Particle Nature of Light

While considering light as a wave does explain much of its everyday behavior, it fails to adequately describe important aspects of light's interactions with matter. The wave model of light cannot explain why heated objects emit only certain frequencies of light at a given temperature, or why some metals emit electrons when colored light of a specific frequency shines on them. Obviously, a totally new model or a revision of the current model of light was needed to address these phenomena.

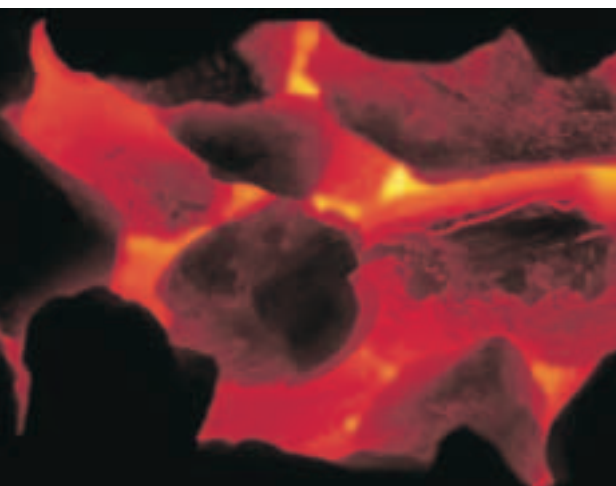
**The quantum concept** The glowing light emitted by the hot objects shown in **Figure 5-6** are examples of a phenomenon you have certainly seen. Iron provides another example of the phenomenon. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and appears bluish in color at even higher temperatures. As you will learn in greater detail later on in this course, the temperature of an object is a measure of the average kinetic energy of its particles. As the iron gets hotter it possesses a greater amount of energy, and emits different colors of light. These different colors correspond to different frequencies and wavelengths. The wave model could not explain the emission of these different wavelengths of light at different temperatures. In 1900, the German physicist Max Planck (1858–1947) began searching for an explanation as he studied the light emitted from heated objects. His study of the phenomenon led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. That is, a **quantum** is the minimum amount of energy that can be gained or lost by an atom.

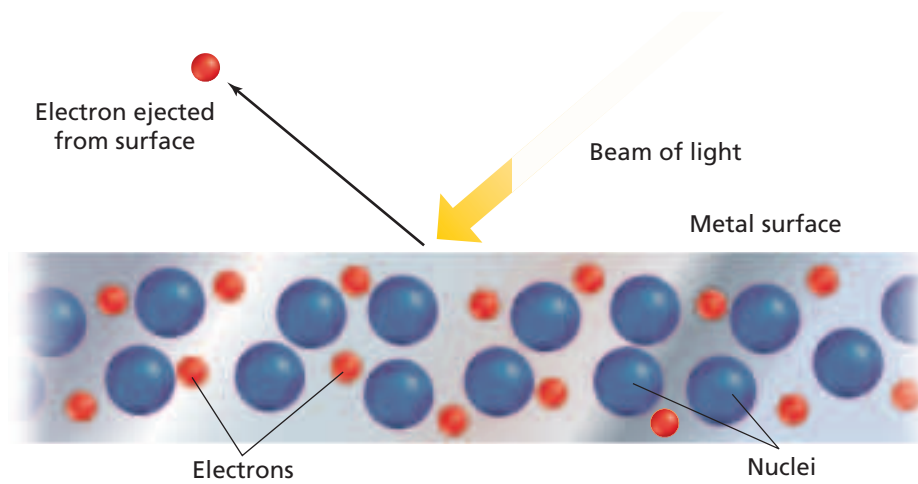
Planck and other physicists of the time thought the concept of quantized energy was revolutionary—and some found it disturbing. Prior experience had led scientists to believe that energy could be absorbed and emitted in continually varying quantities, with no minimum limit to the amount. For example, think about heating a cup of water in a microwave oven. It seems that you can add any amount of thermal energy to the water by regulating the power and duration of the microwaves. Actually, the water's temperature increases in infinitesimal steps as its molecules absorb quanta of energy. Because these steps are so small, the temperature seems to rise in a continuous, rather than a stepwise, manner.

The glowing objects shown in **Figure 5-6** are emitting light, which is a form of energy. Planck proposed that this emitted light energy was quantized.

**Figure 5-6**

These photos illustrate the phenomenon of heated objects emitting different frequencies of light. Matter, regardless of its form, can gain or lose energy only in small "quantized" amounts.





**Figure 5-7**

In the photoelectric effect, light of a certain minimum frequency (energy) ejects electrons from the surface of a metal. Increasing the intensity of the incident light results in more electrons being ejected. Increasing the frequency (energy) of the incident light causes the ejected electrons to travel faster.

He then went further and demonstrated mathematically that the energy of a quantum is related to the frequency of the emitted radiation by the equation

$$E_{\text{quantum}} = h\nu$$

where  $E$  is energy,  $h$  is Planck's constant, and  $\nu$  is frequency. **Planck's constant** has a value of  $6.626 \times 10^{-34} \text{ J} \cdot \text{s}$ , where J is the symbol for the joule, the SI unit of energy. Looking at the equation, you can see that the energy of radiation increases as the radiation's frequency,  $\nu$ , increases. This equation explains why the violet light in **Figure 5-3** has greater energy than the red light.

According to Planck's theory, for a given frequency,  $\nu$ , matter can emit or absorb energy only in whole-number multiples of  $h\nu$ ; that is,  $1h\nu$ ,  $2h\nu$ ,  $3h\nu$ , and so on. A useful analogy for this concept is that of a child building a wall of wooden blocks. The child can add to or take away height from the wall only in increments of a whole number of blocks. Partial blocks are not possible. Similarly, matter can have only certain amounts of energy—quantities of energy between these values do not exist.

**The photoelectric effect** Scientists knew that the wave model (still very popular in spite of Planck's proposal) could not explain a phenomenon called the photoelectric effect. In the **photoelectric effect**, electrons, called photoelectrons, are emitted from a metal's surface when light of a certain frequency shines on the surface, as shown in **Figure 5-7**. Perhaps you've taken advantage of the photoelectric effect by using a calculator, such as the one shown in **Figure 5-8**, that is powered by photoelectric cells. Photoelectric cells in these and many other devices convert the energy of incident light into electrical energy.

The mystery of the photoelectric effect concerns the frequency, and therefore color, of the incident light. The wave model predicts that given enough time, even low-energy, low-frequency light would accumulate and supply enough energy to eject photoelectrons from a metal. However, a metal will not eject photoelectrons below a specific frequency of incident light. For example, no matter how intense or how long it shines, light with a frequency less than  $1.14 \times 10^{15} \text{ Hz}$  does not eject photoelectrons from silver. But even dim light having a frequency equal to or greater than  $1.14 \times 10^{15} \text{ Hz}$  causes the ejection of photoelectrons from silver.

In explaining the photoelectric effect, Albert Einstein proposed in 1905 that electromagnetic radiation has both wavelike and particle-like natures. That is, while a beam of light has many wavelike characteristics, it also can be thought of as a stream of tiny particles, or bundles of energy, called photons. Thus, a **photon** is a particle of electromagnetic radiation with no mass that carries a quantum of energy.

**Figure 5-8**

The direct conversion of sunlight into electrical energy is a viable power source for low-power consumption devices such as this calculator. The cost of photoelectric cells makes them impractical for large-scale power production.







Extending Planck's idea of quantized energy, Einstein calculated that a photon's energy depends on its frequency.

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

Further, Einstein proposed that the energy of a photon of light must have a certain minimum, or threshold, value to cause the ejection of a photoelectron. That is, for the photoelectric effect to occur a photon must possess, at a minimum, the energy required to free an electron from an atom of the metal. According to this theory, even small numbers of photons with energy above the threshold value will cause the photoelectric effect. Although Einstein was able to explain the photoelectric effect by giving electromagnetic radiation particlelike properties, it's important to note that a dual wave-particle model of light was required.

## EXAMPLE PROBLEM 5-2



Sunlight bathes Earth in white light—light composed of all of the visible colors of the electromagnetic spectrum.

### Calculating the Energy of a Photon

Tiny water drops in the air disperse the white light of the sun into a rainbow. What is the energy of a photon from the violet portion of the rainbow if it has a frequency of  $7.23 \times 10^{14} \text{ s}^{-1}$ ?

#### 1. Analyze the Problem

You are given the frequency of a photon of violet light. You also know that the energy of a photon is related to its frequency by the equation  $E_{\text{photon}} = h\nu$ . The value of  $h$ , Planck's constant, is known. By substituting the known values, the equation can be solved for the energy of a photon of violet light.

#### Known

$$\begin{aligned}\nu &= 7.23 \times 10^{14} \text{ s}^{-1} \\ h &= 6.626 \times 10^{-34} \text{ J} \cdot \text{s}\end{aligned}$$

#### Unknown

$$E_{\text{photon}} = ? \text{ J}$$

#### 2. Solve for the Unknown

Substitute the known values for frequency and Planck's constant into the equation relating energy of a photon and frequency.

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

$$E_{\text{photon}} = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(7.23 \times 10^{14} \text{ s}^{-1})$$

Multiply the known values and cancel units.

$$E_{\text{photon}} = (6.626 \times 10^{-34} \text{ J} \cdot \cancel{\text{s}})(7.23 \times 10^{14} \cancel{\text{ s}^{-1}}) = 4.79 \times 10^{-19} \text{ J}$$

The energy of one photon of violet light is  $4.79 \times 10^{-19} \text{ J}$ .

#### 3. Evaluate the Answer

The answer is correctly expressed in a unit of energy (J). The known value for frequency has three significant figures, and the answer also is expressed with three significant figures, as it should be. As expected, the energy of a single photon of light is extremely small.

## PRACTICE PROBLEMS

- What is the energy of each of the following types of radiation?
  - $6.32 \times 10^{20} \text{ s}^{-1}$
  - $9.50 \times 10^{13} \text{ Hz}$
  - $1.05 \times 10^{16} \text{ s}^{-1}$
- Use Figure 5-5 to determine the types of radiation described in problem 5.

Practice!

For more practice with photon energy problems, go to **Supplemental Practice Problems** in Appendix A.



## Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? The process illustrates another phenomenon that cannot be explained by the wave model of light. The light of the neon sign is produced by passing electricity through a tube filled with neon gas. Neon atoms in the tube absorb energy and become excited. These excited and unstable atoms then release energy by emitting light. If the light emitted by the neon is passed through a glass prism, neon's atomic emission spectrum is produced. The **atomic emission spectrum** of an element is the set of frequencies of the electromagnetic waves emitted by atoms of the element. Neon's atomic emission spectrum consists of several individual lines of color, not a continuous range of colors as seen in the visible spectrum.

Each element's atomic emission spectrum is unique and can be used to determine if that element is part of an unknown compound. For example, when a platinum wire is dipped into a strontium nitrate solution and then inserted into a burner flame, the strontium atoms emit a characteristic red color. You can perform a series of flame tests yourself by doing the **miniLAB** below.

**Figure 5-9** on the following page shows an illustration of the characteristic purple-pink glow produced by excited hydrogen atoms and the visible portion of hydrogen's emission spectrum responsible for producing the glow. Note how the line nature of hydrogen's atomic emission spectrum differs from that of a continuous spectrum. To gain firsthand experience with types of line spectra, you can perform the **CHEMLAB** at the end of this chapter.

### miniLAB

#### Flame Tests

**Classifying** When certain compounds are heated in a flame, they emit a distinctive color. The color of the emitted light can be used to identify the compound.

**Materials** Bunsen burner; cotton swabs (6); distilled water; crystals of lithium chloride, sodium chloride, potassium chloride, calcium chloride, strontium chloride, unknown

**Procedure**

1. Dip a cotton swab into the distilled water. Dip the moistened swab into the lithium chloride so that a few of the crystals stick to the cotton. Put the crystals on the swab into the flame of a Bunsen burner. Observe the color of the flame and record it in your data table.
2. Repeat step 1 for each of the metallic chlorides (sodium chloride, potassium chloride, calcium chloride, and strontium chloride). Be sure to record the color of each flame in your data table.
3. Obtain a sample of unknown crystals from your teacher. Repeat the procedure in step 1 using

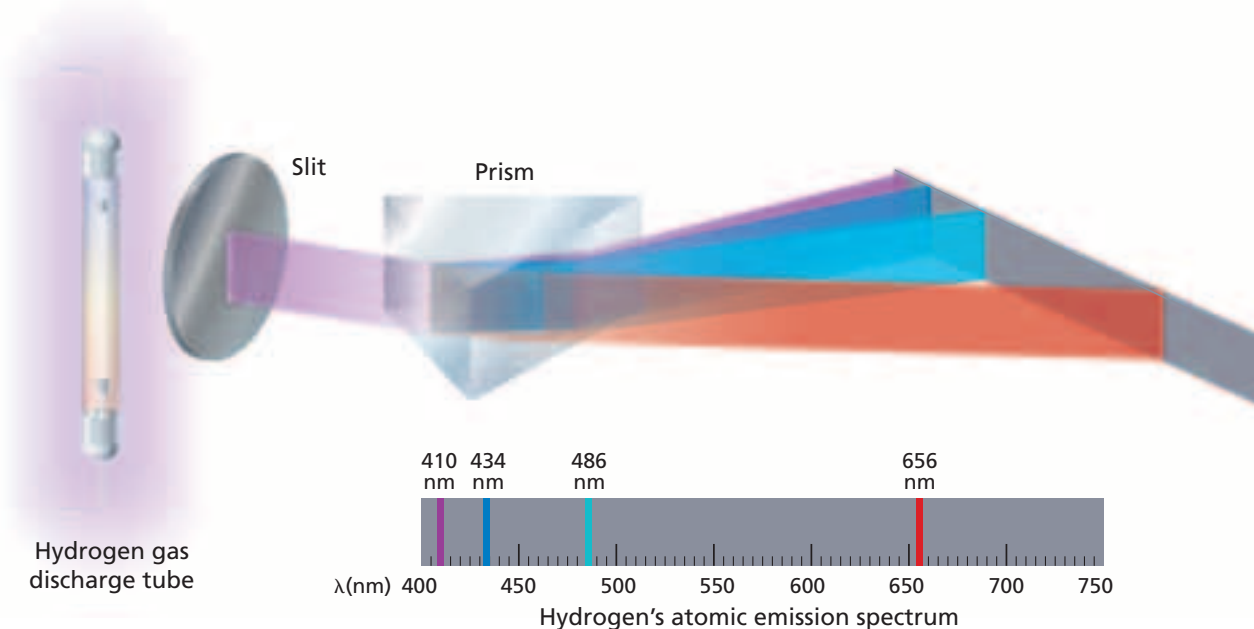
Flame Test Results	
Compound	Flame color
Lithium chloride	
Sodium chloride	
Potassium chloride	
Calcium chloride	
Strontium chloride	
Unknown	

the unknown crystals. Record the color of the flame produced by the unknown crystals in your data table. Dispose of used cotton swabs as directed by your teacher.

#### Analysis

1. Each of the known compounds tested contains chlorine, yet each compound produced a flame of a different color. Explain why this occurred.
2. How is the atomic emission spectrum of an element related to these flame tests?
3. What is the identity of the unknown crystals? Explain how you know.





**Figure 5-9**

The atomic emission spectrum of hydrogen consists of four distinct colored lines of different frequencies. This type of spectrum is also known as a line spectrum. Which line has the highest energy?

An atomic emission spectrum is characteristic of the element being examined and can be used to identify that element. The fact that only certain colors appear in an element's atomic emission spectrum means that only certain specific frequencies of light are emitted. And because those emitted frequencies of light are related to energy by the formula  $E_{\text{photon}} = h\nu$ , it can be concluded that only photons having certain specific energies are emitted. This conclusion was not predicted by the laws of classical physics known at that time. Scientists found atomic emission spectra puzzling because they had expected to observe the emission of a continuous series of colors and energies as excited electrons lost energy and spiraled toward the nucleus. In the next section, you will learn about the continuing development of atomic models, and how one of those models was able to account for the frequencies of the light emitted by excited atoms.

## Section 5.1 Assessment

- List the characteristic properties of all waves. At what speed do electromagnetic waves travel in a vacuum?
- Compare the wave and particle models of light. What phenomena can only be explained by the particle model?
- What is a quantum of energy? Explain how quanta of energy are involved in the amount of energy matter gains and loses.
- Explain the difference between the continuous spectrum of white light and the atomic emission spectrum of an element.
- Thinking Critically** Explain how Einstein utilized Planck's quantum concept in explaining the photoelectric effect.
- Interpreting Scientific Illustrations** Use **Figure 5-5** and your knowledge of light to match the numbered items on the right with the lettered items on the left. The numbered items may be used more than once or not at all.
  - longest wavelength
  - highest frequency
  - greatest energy
  - gamma rays
  - infrared waves
  - radio waves

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

## Bohr Model of the Atom

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

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

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

**Table 5-1**

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

## Objectives

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

## Vocabulary

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

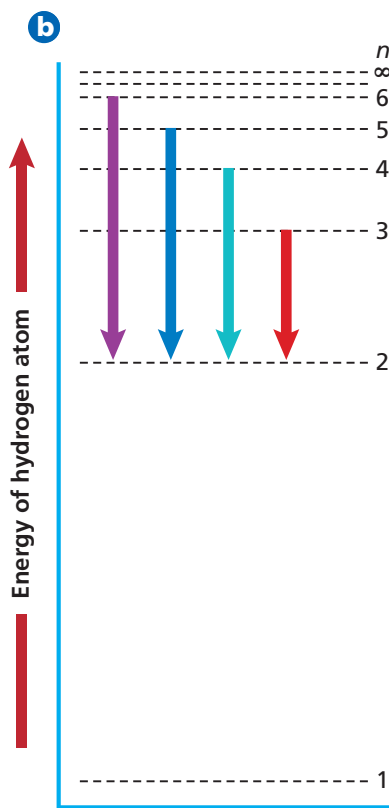
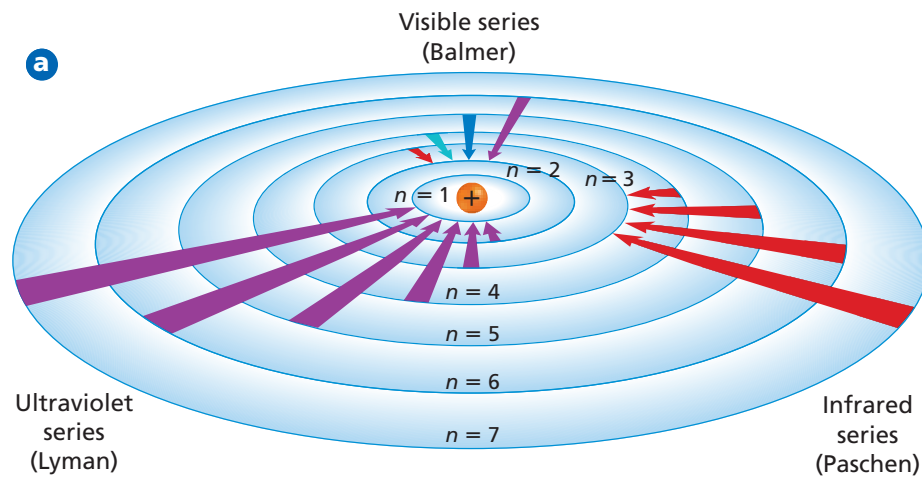


**An explanation of hydrogen's line spectrum** Bohr suggested that the hydrogen atom is in the ground state, also called the first energy level, when the electron is in the  $n = 1$  orbit. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron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

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

**Figure 5-10**

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



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

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

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



## The Quantum Mechanical Model of the Atom

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

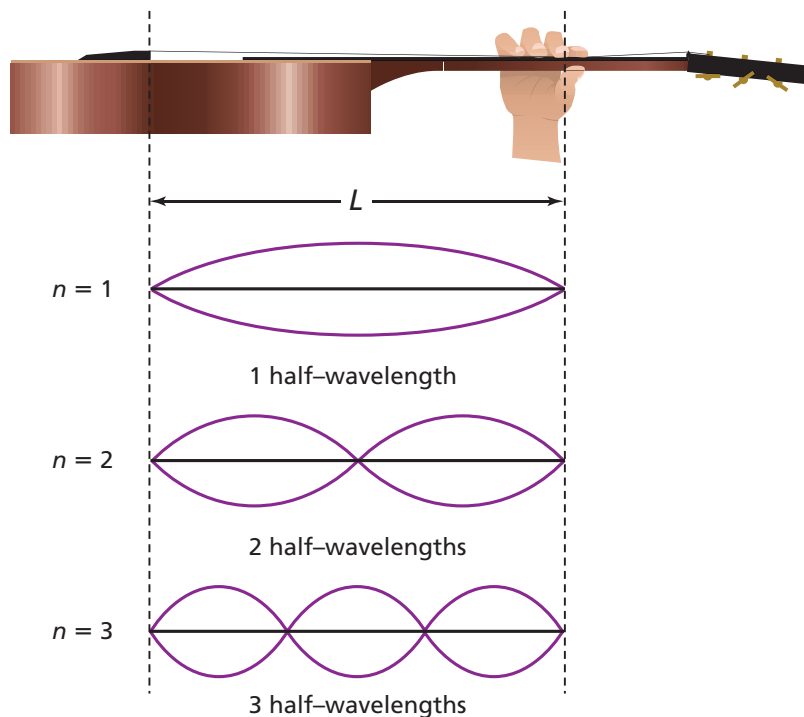
**Electrons as waves** De Broglie had been thinking that Bohr’s quantized electron orbits had characteristics similar to those of waves. For example, as **Figure 5-11b** shows, only multiples of half-wavelengths are possible on a plucked guitar string because the string is fixed at both ends. Similarly, de Broglie saw that only whole numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in **Figure 5-11c**. He also reflected on



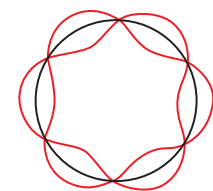
the fact that light—at one time thought to be strictly a wave phenomenon—has both wave and particle characteristics. These thoughts led de Broglie to pose a new question. If waves can have particlelike behavior, could the opposite also be true? That is, can particles of matter, including electrons, behave like waves?

**Figure 5-11**

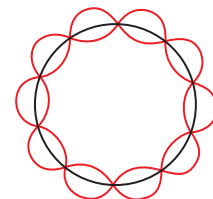
**a** A vibrating guitar string is constrained to vibrate between two fixed end points. **b** The possible vibrations of the guitar string are limited to multiples of half-wavelengths. Thus, the “quantum” of the guitar string is one-half wavelength. **c** The possible circular orbits of an electron are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



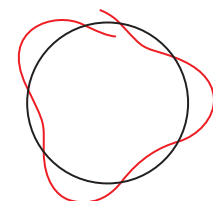
**b** Vibrating guitar string  
Only multiples of half wavelengths allowed



$n = 3$  wavelengths



$n = 5$  wavelengths



$n \neq$  whole number (not allowed)

**c** Orbiting electron  
Only whole numbers of wavelengths allowed



## problem-solving LAB

### How was Bohr's atomic model able to explain the line spectrum of hydrogen?

**Using Models** Niels Bohr proposed that electrons must occupy specific, quantized energy levels in an atom. He derived the following equations for hydrogen's electron orbit energies ( $E_n$ ) and radii ( $r_n$ ).

$$r_n = (0.529 \times 10^{-10} \text{ m})n^2$$

$$E_n = -(2.18 \times 10^{-18} \text{ J})/n^2$$

Where  $n$  = quantum number (1, 2, 3...).

#### Analysis

Using the orbit radii equation, calculate hydrogen's first seven electron orbit radii and then construct a scale model of those orbits. Use a compass and a metric ruler to draw your scale model on two sheets of paper that have been taped together. (Use caution when handling sharp objects.) Using the orbit energy equation, calculate the energy of each electron orbit and record the values on your model.

#### Thinking Critically

1. What scale did you use to plot the orbits? How is the energy of each orbit related to its radius?
2. Draw a set of arrows for electron jumps that end at each energy level (quantum number). For example, draw a set of arrows for all transitions that end at  $n = 1$ , a set of arrows for all transitions that end at  $n = 2$ , and so on, up to  $n = 7$ .
3. Calculate the energy released for each of the jumps in step 2, and record the values on your model. The energy released is equal to the difference in the energies of each level.
4. Each set of arrows in step 2 represents a spectral emission series. Label five of the series, from greatest energy change to least energy change, as the Lyman, Balmer, Paschen, Brackett, and Pfund series.
5. Use the energy values in step 3 to calculate the frequency of each photon emitted in each series. Record the frequencies on your model.
6. Using the electromagnetic spectrum as a guide, identify in which range (visible, ultraviolet, infrared, etc.) each series falls.

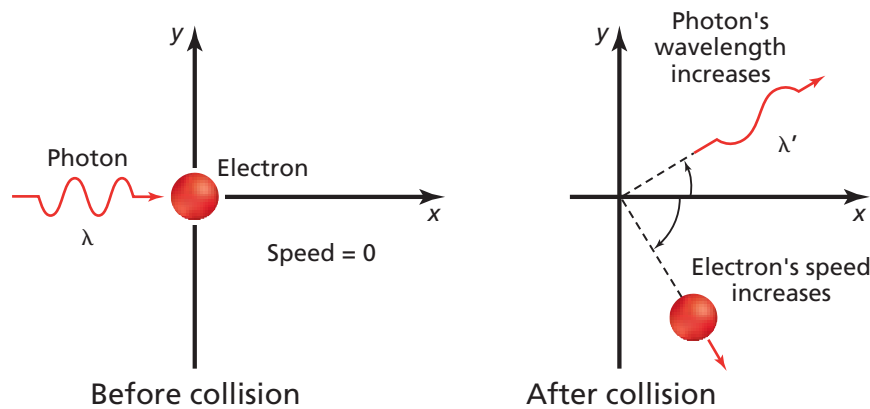
In considering this question, de Broglie knew that if an electron has wave-like motion and is restricted to circular orbits of fixed radius, the electron is allowed only certain possible wavelengths, frequencies, and energies. Developing his idea, de Broglie derived an equation for the wavelength ( $\lambda$ ) of a particle of mass ( $m$ ) moving at velocity ( $v$ ).

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

The **de Broglie equation** predicts that all moving particles have wave characteristics. Why, then, you may be wondering, haven't you noticed the wavelength of a fast-moving automobile? Using de Broglie's equation provides an answer. An automobile moving at 25 m/s and having a mass of 910 kg has a wavelength of  $2.9 \times 10^{-38}$  m—a wavelength far too small to be seen or detected, even with the most sensitive scientific instrument. By comparison, an electron moving at the same speed has the easily measured wavelength of  $2.9 \times 10^{-5}$  m. Subsequent experiments have proven that electrons and other moving particles do indeed have wave characteristics.

Step by step, scientists such as Rutherford, Bohr, and de Broglie had been unraveling the mysteries of the atom. However, a conclusion reached by the German theoretical physicist Werner Heisenberg (1901–1976), a contemporary of de Broglie, proved to have profound implications for atomic models.





**Figure 5-12**

A photon that strikes an electron at rest alters the position and velocity of the electron. This collision illustrates the Heisenberg uncertainty principle: It is impossible to simultaneously know both the position and velocity of a particle. Note that after the collision, the photon's wavelength is longer. How has the photon's energy changed?

## The Heisenberg Uncertainty Principle

Heisenberg concluded that it is impossible to make any measurement on an object without disturbing the object—at least a little. Imagine trying to locate a hovering, helium-filled balloon in a completely darkened room. When you wave your hand about, you'll locate the balloon's position when you touch it. However, when you touch the balloon, even gently, you transfer energy to it and change its position. Of course, you could also detect the balloon's position by turning on a flashlight. Using this method, photons of light that reflect from the balloon reach your eyes and reveal the balloon's location. Because the balloon is much more massive than the photons, the rebounding photons have virtually no effect on the balloon's position.

Can photons of light help determine the position of an electron in an atom? As a thought experiment, imagine trying to determine the electron's location by “bumping” it with a high-energy photon of electromagnetic radiation. Unfortunately, because such a photon has about the same energy as an electron, the interaction between the two particles changes both the wavelength of the photon and the position and velocity of the electron, as shown in **Figure 5-12**. In other words, the act of observing the electron produces a significant, unavoidable uncertainty in the position and motion of the electron. Heisenberg's analysis of interactions such as those between photons and electrons led him to his historic conclusion. The **Heisenberg uncertainty principle** states that it is fundamentally impossible to know precisely both the velocity and position of a particle at the same time.

Although scientists of the time found Heisenberg's principle difficult to accept, it has been proven to describe the fundamental limitations on what can be observed. How important is the Heisenberg uncertainty principle? The interaction of a photon with an object such as a helium-filled balloon has so little effect on the balloon that the uncertainty in its position is too small to measure. But that's not the case with an electron moving at  $6 \times 10^6$  m/s near an atomic nucleus. The uncertainty in the electron's position is at least  $10^{-9}$  m, about ten times greater than the diameter of the entire atom!

**The Schrödinger wave equation** In 1926, Austrian physicist Erwin Schrödinger (1887–1961) furthered the wave-particle theory proposed by de Broglie. Schrödinger derived an equation that treated the hydrogen atom's electron as a wave. Remarkably, Schrödinger's new model for the hydrogen atom seemed to apply equally well to atoms of other elements—an area in which Bohr's model failed. The atomic model in which electrons are treated as waves is called the wave mechanical model of the atom or, more commonly, the **quantum mechanical model of the atom**. Like Bohr's model,

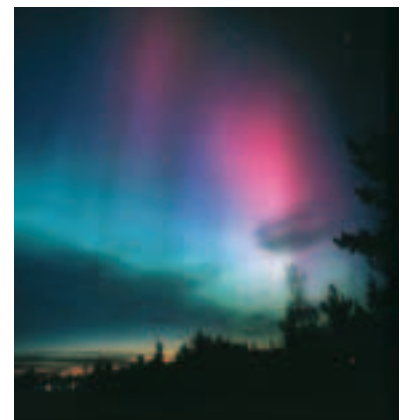
## Physics

### CONNECTION

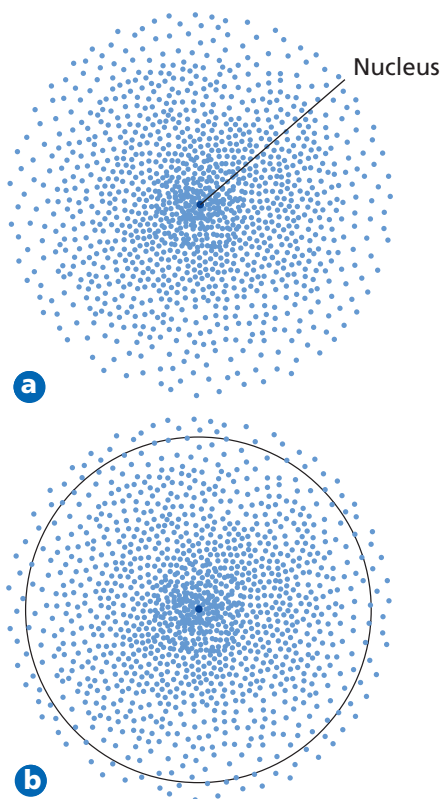
People travel thousands of miles to see the aurora borealis (the northern lights) and the aurora australis (the southern lights). Once incorrectly believed to be reflections from the polar ice fields, the auroras occur 100 to 1000 km above Earth.

High-energy electrons and positive ions in the solar wind speed away from the sun at more than one million kilometers per hour. These particles become trapped in Earth's magnetic field and follow along Earth's magnetic field lines.

The electrons interact with and transfer energy to oxygen and nitrogen atoms in the upper atmosphere. The color of the aurora depends on altitude and which atoms become excited. Oxygen emits green light up to about 250 km and red light above 250 km; nitrogen emits blue light up to about 100 km and purple/violet at higher altitudes.







**Figure 5-13**

- a** This electron density diagram for a hydrogen atom represents the likelihood of finding an electron at a particular point in the atom. The greater the density of dots, the greater the likelihood of finding hydrogen's electron. **b** The boundary of an atom is defined as the volume that encloses a 90% probability of containing its electrons.

the quantum mechanical model limits an electron's energy to certain values. However, unlike Bohr's model, the quantum mechanical model makes no attempt to describe the electron's path around the nucleus.

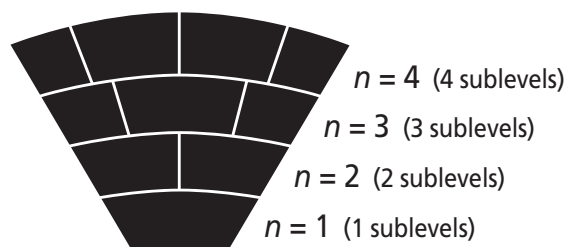
The Schrödinger wave equation is too complex to be considered here. However, each solution to the equation is known as a wave function. And most importantly, the wave function is related to the probability of finding the electron within a particular volume of space around the nucleus. Recall from your study of math that an event having a high probability is more likely to occur than one having a low probability.

What does the wave function predict about the electron's location in an atom? A three-dimensional region around the nucleus called an **atomic orbital** describes the electron's probable location. You can picture an atomic orbital as a fuzzy cloud in which the density of the cloud at a given point is proportional to the probability of finding the electron at that point. **Figure 5-13a** illustrates the probability map, or orbital, that describes the hydrogen electron in its lowest energy state. It might be helpful to think of the probability map as a time-exposure photograph of the electron moving around the nucleus, in which each dot represents the electron's location at an instant in time. Because the dots are so numerous near the positive nucleus, they seem to form a dense cloud that is indicative of the electron's most probable location. However, because the cloud has no definite boundary, it also is possible that the electron might be found at a considerable distance from the nucleus.

## Hydrogen's Atomic Orbitals

Because the boundary of an atomic orbital is fuzzy, the orbital does not have an exactly defined size. To overcome the inherent uncertainty about the electron's location, chemists arbitrarily draw an orbital's surface to contain 90% of the electron's total probability distribution. In other words, the electron spends 90% of the time within the volume defined by the surface, and 10% of the time somewhere outside the surface. The spherical surface shown in **Figure 5-13b** encloses 90% of the lowest-energy orbital of hydrogen.

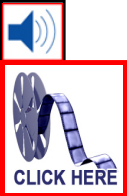
Recall that the Bohr atomic model assigns quantum numbers to electron orbits. In a similar manner, the quantum mechanical model assigns **principal quantum numbers** ( $n$ ) that indicate the relative sizes and energies of atomic



**Figure 5-14**

Energy sublevels can be thought of as a section of seats in a theater. The rows that are higher up and farther from the stage contain more seats, just as energy levels that are farther from the nucleus contain more sublevels.



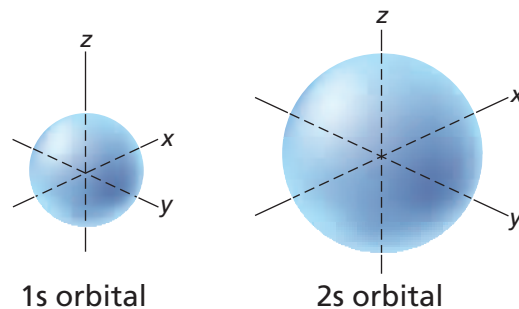


orbitals. That is, as  $n$  increases, the orbital becomes larger, the electron spends more time farther from the nucleus, and the atom's energy level increases. Therefore,  $n$  specifies the atom's major energy levels, called **principal energy levels**. An atom's lowest principal energy level is assigned a principal quantum number of one. When the hydrogen atom's single electron occupies an orbital with  $n = 1$ , the atom is in its ground state. Up to seven energy levels have been detected for the hydrogen atom, giving  $n$  values ranging from 1 to 7.

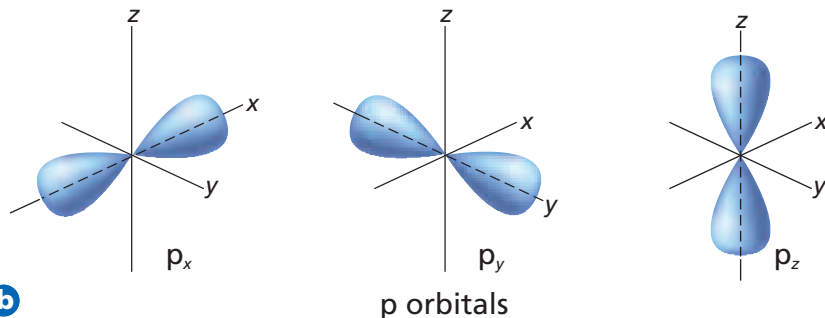
Principal energy levels contain **energy sublevels**. Principal energy level 1 consists of a single sublevel, principal energy level 2 consists of two sublevels, principal energy level 3 consists of three sublevels, and so on. To better understand the relationship between the atom's energy levels and sublevels, picture the seats in a wedge-shaped section of a theater, as shown in **Figure 5-14**. As you move away from the stage, the rows become higher and contain more seats. Similarly, the number of energy sublevels in a principal energy level increases as  $n$  increases.

Sublevels are labeled s, p, d, or f according to the shapes of the atom's orbitals. All s orbitals are spherical and all p orbitals are dumbbell shaped; however, not all d or f orbitals have the same shape. Each orbital may contain at most two electrons. The single sublevel in principal energy level 1 consists of a spherical orbital called the 1s orbital. The two sublevels in principal energy level 2 are designated 2s and 2p. The 2s sublevel consists of the 2s orbital, which is spherical like the 1s orbital but larger in size. See **Figure 5-15a**. The 2p sublevel consists of three dumbbell-shaped p orbitals of equal energy designated  $2p_x$ ,  $2p_y$ , and  $2p_z$ . The subscripts  $x$ ,  $y$ , and  $z$  merely designate the orientations of p orbitals along the  $x$ ,  $y$ , and  $z$  coordinate axes, as shown in **Figure 5-15b**.

Principal energy level 3 consists of three sublevels designated 3s, 3p, and 3d. Each d sublevel consists of five orbitals of equal energy. Four d orbitals have identical shapes but different orientations. However, the fifth,  $d_{z^2}$  orbital is shaped and oriented differently from the other four. The shapes and orientations of the five d orbitals are illustrated in **Figure 5-16**. The fourth principal energy level ( $n = 4$ ) contains a fourth sublevel, called the 4f sublevel, which consists of seven f orbitals of equal energy.



**a**



**b**

**Figure 5-15**

Atomic orbitals represent the electron probability clouds of an atom's electrons. **a** The spherical 1s and 2s orbitals are shown here. All s orbitals are spherical in shape and increase in size with increasing principal quantum number. **b** The three dumbbell-shaped p orbitals are oriented along the three perpendicular  $x$ ,  $y$ , and  $z$  axes. Each of the p orbitals related to an energy sublevel has equal energy.

**Figure 5-16**

Four of five equal-energy d orbitals have the same shape. Notice how the  $d_{xy}$  orbital lies in the plane formed by the  $x$  and  $y$  axes, the  $d_{xz}$  orbital lies in the plane formed by the  $x$  and  $z$  axes, and so on. The  $d_{z^2}$  orbital has its own unique shape.

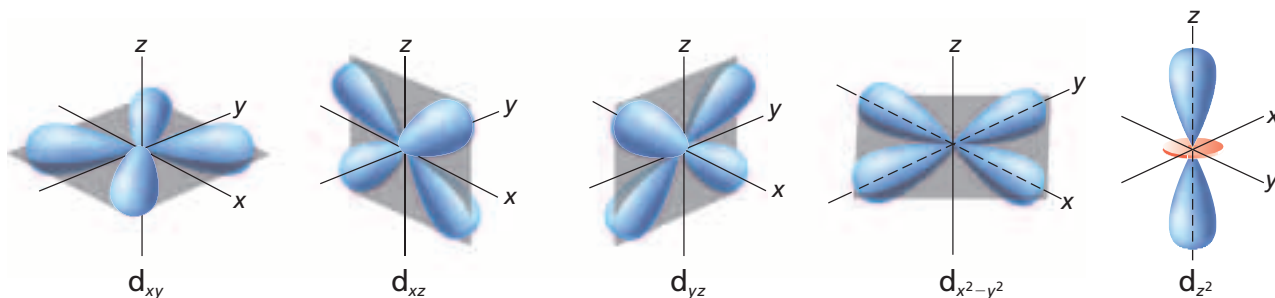




Table 5-2

Hydrogen's First Four Principal Energy Levels			
Principal quantum number ( $n$ )	Sublevels (types of orbitals) present	Number of orbitals related to sublevel	Total number of orbitals related to principal energy level ( $n^2$ )
1	s	1	1
2	s p	1 3	4
3	s p d	1 3 5	9
4	s p d f	1 3 5 7	16

Hydrogen's first four principal energy levels, sublevels, and related atomic orbitals are summarized in **Table 5-2**. Note that the maximum number of orbitals related to each principal energy level equals  $n^2$ . Because each orbital may contain at most two electrons, the maximum number of electrons related to each principal energy level equals  $2n^2$ .

Given the fact that a hydrogen atom contains only one electron, you might wonder how the atom can have so many energy levels, sublevels, and related atomic orbitals. At any given time, the atom's electron can occupy just one orbital. So you can think of the other orbitals as unoccupied spaces—spaces available should the atom's energy increase or decrease. For example, when the hydrogen atom is in the ground state, the electron occupies the 1s orbital. However, the atom may gain a quantum of energy that excites the electron to the 2s orbital, to one of the three 2p orbitals, or to another vacant orbital.

You have learned a lot about electrons and quantized energy in this section: how Bohr's orbits explained the hydrogen atom's quantized energy states; how de Broglie's insight led scientists to think of electrons as both particles and waves; and how Schrödinger's wave equation predicted the existence of atomic orbitals containing electrons. In the next section, you'll learn how the electrons are arranged in atomic orbitals of atoms having more than one electron.

## Section 5.2 Assessment

13. According to the Bohr atomic model, why do atomic emission spectra contain only certain frequencies of light?
14. Why is the wavelength of a moving soccer ball not detectable to the naked eye?
15. What sublevels are contained in the hydrogen atom's first four energy levels? What orbitals are related to each s sublevel and each p sublevel?
16. **Thinking Critically** Use de Broglie's wave-particle duality and the Heisenberg uncertainty principle to explain why the location of an electron in an atom is uncertain.
17. **Comparing and Contrasting** Compare and contrast the Bohr model and quantum mechanical model of the atom.

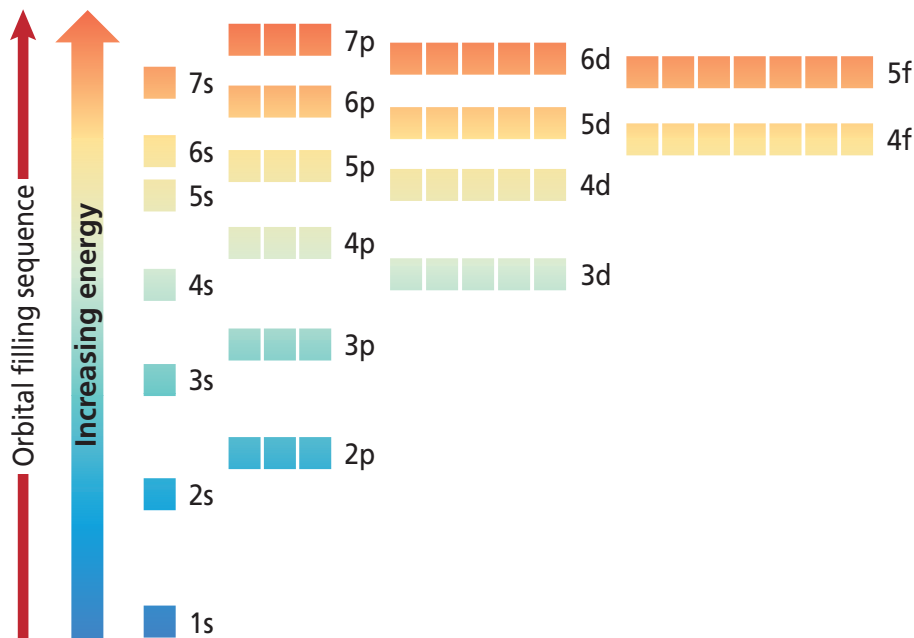
When you consider that atoms of the heaviest elements contain in excess of 100 electrons, that there are numerous principal energy levels and sublevels and their corresponding orbitals, and that each orbital may contain a maximum of two electrons, the idea of determining the arrangement of an atom's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

## Ground-State Electron Configurations

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

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

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



## Objectives

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

## Vocabulary

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

**Figure 5-17**

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



## Careers Using Chemistry

### Spectroscopist

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

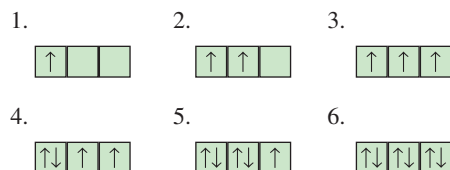
Spectroscopy is the analysis of the characteristic spectra emitted by matter. Spectroscopists perform chemical analyses as part of many research laboratory projects, for quality control in industrial settings, and as part of forensics investigations for law enforcement agencies.

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

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

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

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



## Orbital Diagrams and Electron Configuration Notations

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

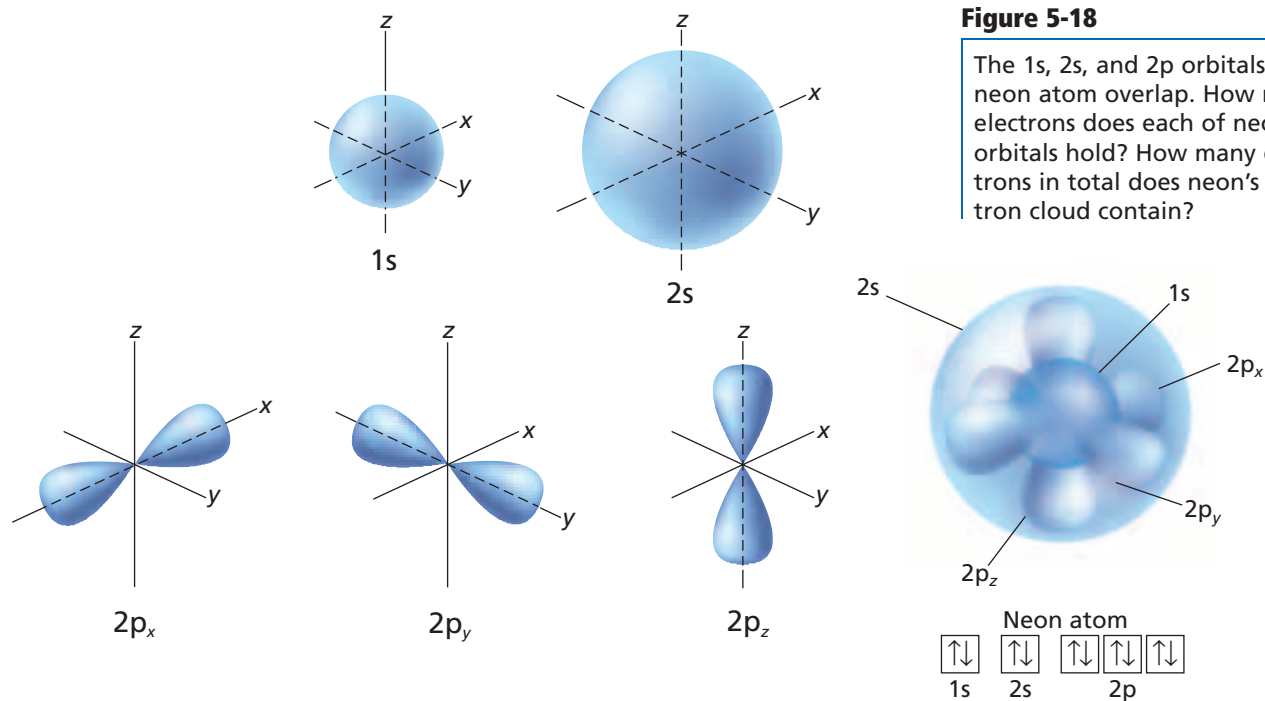


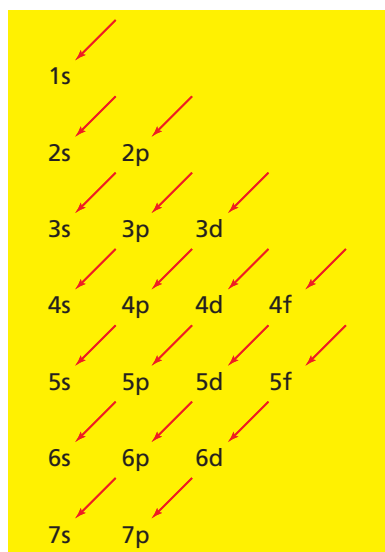
**Table 5-3**

Electron Configurations and Orbital Diagrams for Elements in the First Two Periods					
Element	Atomic number	Orbital diagram			Electron configuration notation
		1s	2s	2p <sub>x</sub> 2p <sub>y</sub> 2p <sub>z</sub>	
Hydrogen	1	↑			1s <sup>1</sup>
Helium	2	↑↓			1s <sup>2</sup>
Lithium	3	↑↓	↑		1s <sup>2</sup> 2s <sup>1</sup>
Beryllium	4	↑↓	↑↓		1s <sup>2</sup> 2s <sup>2</sup>
Boron	5	↑↓	↑↓	↑	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>1</sup>
Carbon	6	↑↓	↑↓	↑ ↑	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>2</sup>
Nitrogen	7	↑↓	↑↓	↑ ↑ ↑	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>3</sup>
Oxygen	8	↑↓	↑↓	↑↓ ↑ ↑	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>4</sup>
Fluorine	9	↑↓	↑↓	↑↓ ↑↓ ↑	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>5</sup>
Neon	10	↑↓	↑↓	↑↓ ↑↓ ↑↓	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup>

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

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



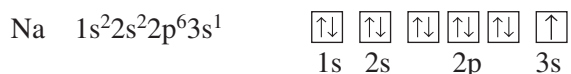


**Figure 5-19**

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

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

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



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

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

**Exceptions to predicted configurations** You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, [Ar] $4s^2 3d^4$ , and copper, [Ar] $4s^2 3d^9$ , would prove to be incorrect. The correct configurations for these two elements are:



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

**Table 5-4**

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

## EXAMPLE PROBLEM 5-3

### Writing Electron Configurations

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

#### 1. Analyze the Problem

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

#### 2. Solve for the Unknown

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

[Ar]

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

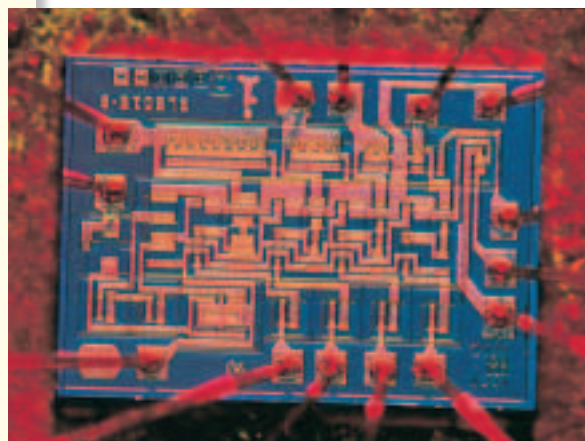
[Ar]4s<sup>2</sup>3d<sup>7</sup>4p<sup>2</sup>

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

[Ar]4s<sup>2</sup>3d<sup>10</sup>4p<sup>2</sup>

#### 3. Evaluate the Answer

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



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

## PRACTICE PROBLEMS

18. Write ground-state electron configurations for the following elements.
  - a. bromine (Br)
  - b. strontium (Sr)
  - c. antimony (Sb)
  - d. rhenium (Re)
  - e. terbium (Tb)
  - f. titanium (Ti)
19. How many electrons are in orbitals related to the third energy level of a sulfur atom?
20. How many electrons occupy p orbitals in a chlorine atom?
21. What element has the following ground-state electron configuration? [Kr]5s<sup>2</sup>4d<sup>10</sup>5p<sup>1</sup>
22. What element has the following ground-state electron configuration? [Xe]6s<sup>2</sup>



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





## Valence Electrons

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



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



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



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

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

**Table 5-5**

Electron-Dot Structures for Elements in Period Two			
Element	Atomic number	Electron configuration	Electron-dot structure
Lithium	3	$1s^22s^1$	Li·
Beryllium	4	$1s^22s^2$	·Be·
Boron	5	$1s^22s^22p^1$	·B·
Carbon	6	$1s^22s^22p^2$	·C·
Nitrogen	7	$1s^22s^22p^3$	·N·
Oxygen	8	$1s^22s^22p^4$	:Ö·
Fluorine	9	$1s^22s^22p^5$	:F·
Neon	10	$1s^22s^22p^6$	:Ne:



## EXAMPLE PROBLEM 5-4

### Writing Electron-Dot Structures

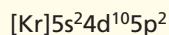
Some sheet glass is manufactured using a process that makes use of molten tin. What is tin's electron-dot structure?

#### 1. Analyze the Problem

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

#### 2. Solve for the Unknown

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

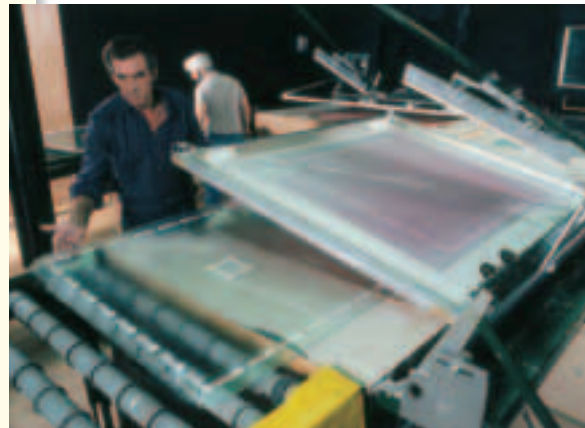


The two 5s and the two 5p electrons (the electrons in the orbitals related to the atom's highest principal energy level) represent tin's four valence electrons. Draw tin's electron-dot structure by representing its four valence electrons with dots, arranged one at a time, around the four sides of tin's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



#### 3. Evaluate the Answer

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



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

## PRACTICE PROBLEMS

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

- |              |             |
|--------------|-------------|
| a. magnesium | d. rubidium |
| b. sulfur    | e. thallium |
| c. bromine   | f. xenon    |



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

## Section 5.3 Assessment

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



## Line Spectra

You know that sunlight is made up of a continuous spectrum of colors that combine to form white light. You also have learned that atoms of gases can emit visible light of characteristic wavelengths when excited by electricity. The color you see is the sum of all of the emitted wavelengths. In this experiment, you will use a diffraction grating to separate these wavelengths into emission line spectra.

You also will investigate another type of line spectrum—the absorption spectrum. The color of each solution you observe is due to the reflection or transmission of unabsorbed wavelengths of light. When white light passes through a sample and then a diffraction grating, dark lines show up on the continuous spectrum of white light. These lines correspond to the wavelengths of the photons absorbed by the solution.

### Problem

What absorption and emission spectra do various substances produce?

### Objectives

- **Observe** emission spectra of several gases.
- **Observe** the absorption spectra of various solutions.
- **Analyze** patterns of absorption and emission spectra.

### Materials

(For each group)  
ring stand with clamp  
40-W tubular light bulb  
light socket with power cord  
275-mL polystyrene culture flask (4)  
Flinn C-Spectra® or similar diffraction grating

food coloring (red, green, blue, and yellow)  
set of colored pencils  
book

(For entire class)  
spectrum tubes (hydrogen, neon, and mercury)  
spectrum tube power supplies (3)

### Safety Precautions



- Always wear safety goggles and a lab apron.
- Use care around the spectrum tube power supplies.
- Spectrum tubes will get hot when used.

### Pre-Lab

1. Read the entire CHEMLAB.
2. Explain how electrons in an element's atoms produce an emission spectrum.
3. Distinguish among a continuous spectrum, an emission spectrum, and an absorption spectrum.
4. Prepare your data tables.

#### Drawings of Emission Spectra

Hydrogen	
Neon	
Mercury	

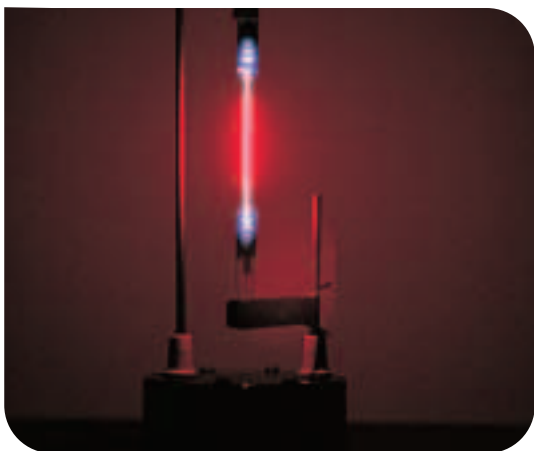
#### Drawings of Absorption Spectra

Red	
Green	
Blue	
Yellow	

### Procedure

1. Use a Flinn C-Spectra® to view an incandescent light bulb. What do you observe? Draw the spectrum using colored pencils.

- Use the Flinn C-Spectra<sup>®</sup> to view the emission spectra from tubes of gaseous hydrogen, neon, and mercury. Use colored pencils to make drawings in the data table of the spectra observed.



- Fill a 275-mL culture flask with about 100-mL water. Add 2 or 3 drops of red food coloring to the water. Shake the solution.
- Repeat step 3 for the green, blue, and yellow food coloring. **CAUTION:** *Be sure to thoroughly dry your hands before handling electrical equipment.*
- Set up the 40-W light bulb so that it is near eye level. Place the flask with red food coloring about 8 cm from the light bulb. Use a book or some other object to act as a stage to put the flask on. You should be able to see light from the bulb above the solution and light from the bulb projecting through the solution.



- With the room lights darkened, view the light using the Flinn C-Spectra<sup>®</sup>. The top spectrum viewed will be a continuous spectrum of the white light bulb. The bottom spectrum will be the absorption spectrum of the red solution. The black areas of the absorption spectrum represent the colors absorbed by the red food coloring in the solution. Use colored pencils to make a drawing in the data table of the absorption spectra you observed.
- Repeat steps 5 and 6 using the green, blue, and yellow colored solutions.

### Cleanup and Disposal

- Turn off the light socket and spectrum tube power supplies.
- Wait several minutes to allow the incandescent light bulb and the spectrum tubes to cool.
- Follow your teacher's instructions on how to dispose of the liquids and how to store the light bulb and spectrum tubes.

### Analyze and Conclude

- Thinking Critically** How can the existence of spectra help to prove that energy levels in atoms exist?
- Thinking Critically** How can the single electron in a hydrogen atom produce all of the lines found in its emission spectrum?
- Predicting** How can you predict the absorption spectrum of a solution by looking at its color?
- Thinking Critically** How can spectra be used to identify the presence of specific elements in a substance?

### Real-World Chemistry

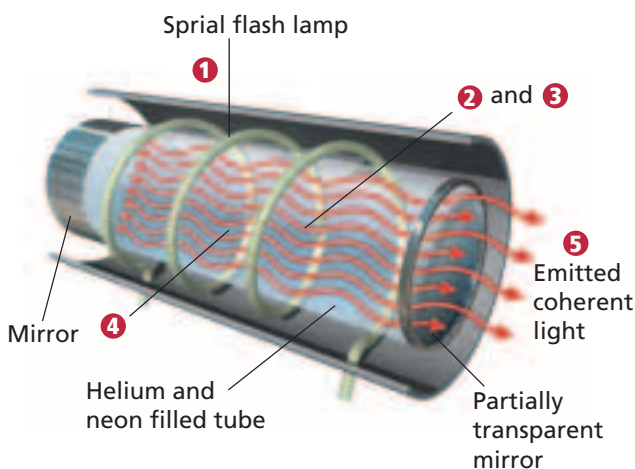
- How can absorption and emission spectra be used by the Hubble space telescope to study the structures of stars or other objects found in deep space?
- The absorption spectrum of chlorophyll *a* indicates strong absorption of red and blue wavelengths. Explain why leaves appear green.

# How It Works

## Lasers

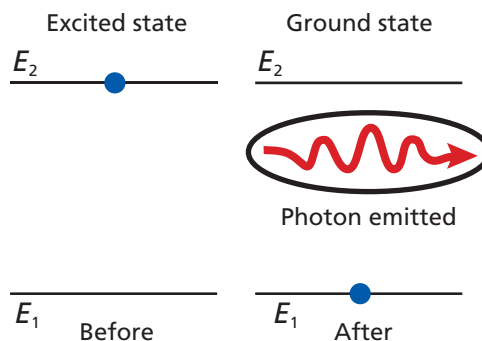
A laser is a device that produces a beam of intense light of a specific wavelength (color). Unlike light from a flashlight, laser light is coherent; that is, it does not spread out as it travels through space. The precise nature of lasers led to their use in pointing and aiming devices, CD players, optical fiber data transmission, and surgery.

- 1 The spiral-wound high-intensity lamp flashes, supplying energy to the helium-neon gas mixture inside the tube. The atoms of the gas absorb the light energy and are raised to an excited energy state.

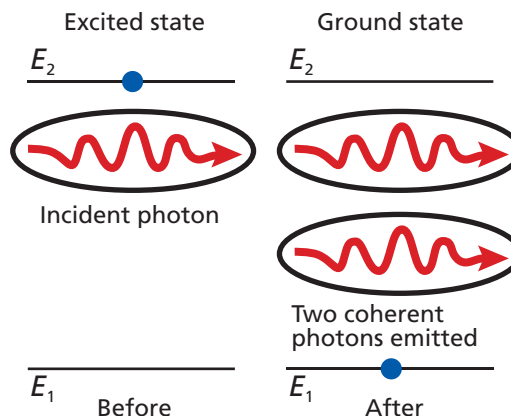


- 4 Photons traveling parallel to the tube are reflected back through the tube by the flat mirrors located at each end. The photons strike additional excited atoms and cause more photons to be released. The intensity of the light in the tube builds.
- 5 Some of the laser's coherent light passes through the partially transparent mirror at one end of the tube and exits the laser. These photons make up the light emitted by the laser.

- 2 The excited atoms begin returning to the ground state, emitting photons in the process. These initial photons travel in all directions.



- 3 The emitted photons hit other excited atoms, causing them to release additional photons. These additional photons are the same wavelength as the photons that struck the excited atoms, and they are coherent (their waves are in sync because they are identical in wavelength and direction).



### Thinking Critically

**1. Inferring** How does the material used in the laser affect the type of light emitted?

**2. Relating Cause and Effect** Why is one mirror partially transparent?

## Summary

### 5.1 Light and Quantized Energy

- All waves can be described by their wavelength, frequency, amplitude, and speed.
- Light is an electromagnetic wave. In a vacuum, all electromagnetic waves travel at a speed of  $3.00 \times 10^8$  m/s.
- All electromagnetic waves may be described as both waves and particles. Particles of light are called photons.
- Energy is emitted and absorbed by matter in quanta.
- In contrast to the continuous spectrum produced by white light, an element's atomic emission spectrum consists of a series of fine lines of individual colors.

### 5.2 Quantum Theory and the Atom

- According to the Bohr model of the atom, hydrogen's atomic emission spectrum results from electrons dropping from higher-energy atomic orbits to lower-energy atomic orbits.
- The de Broglie equation predicts that all moving particles have wave characteristics and relates each particle's wavelength to its mass, its velocity, and Planck's constant.

- The quantum mechanical model of the atom is based on the assumption that electrons are waves.
- The Heisenberg uncertainty principle states that it is not possible to know precisely the velocity and the position of a particle at the same time.
- Electrons occupy three-dimensional regions of space called atomic orbitals. There are four types of orbitals, denoted by the letters s, p, d, and f.

### 5.3 Electron Configurations

- The arrangement of electrons in an atom is called the atom's electron configuration. Electron configurations are prescribed by three rules: the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- Electrons related to the atom's highest principal energy level are referred to as valence electrons. Valence electrons determine the chemical properties of an element.
- Electron configurations may be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

## Key Equations and Relationships

- EM Wave relationship:  $c = \lambda\nu$  (p. 119)
- Energy of a quantum:  $E_{\text{quantum}} = h\nu$  (p. 123)
- Energy of a photon:  $E_{\text{photon}} = h\nu$  (p. 124)
- Energy change of an electron:  $\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}}$   
 $\Delta E = E_{\text{photon}} = h\nu$  (p. 128)
- de Broglie's equation:  $\lambda = \frac{h}{mv}$  (p. 130)

## Vocabulary

- amplitude (p. 119)
- atomic emission spectrum (p. 125)
- atomic orbital (p. 132)
- aufbau principle (p. 135)
- de Broglie equation (p. 130)
- electromagnetic radiation (p. 118)
- electromagnetic spectrum (p. 120)
- electron configuration (p. 135)
- electron-dot structure (p. 140)
- energy sublevel (p. 133)
- frequency (p. 118)
- ground state (p. 127)
- Heisenberg uncertainty principle (p. 131)
- Hund's rule (p. 136)
- Pauli exclusion principle (p. 136)
- photoelectric effect (p. 123)
- photon (p. 123)
- Planck's constant (p. 123)
- principal energy level (p. 133)
- principal quantum number (p. 132)
- quantum (p. 122)
- quantum mechanical model of the atom (p. 131)
- valence electron (p. 140)
- wavelength (p. 118)

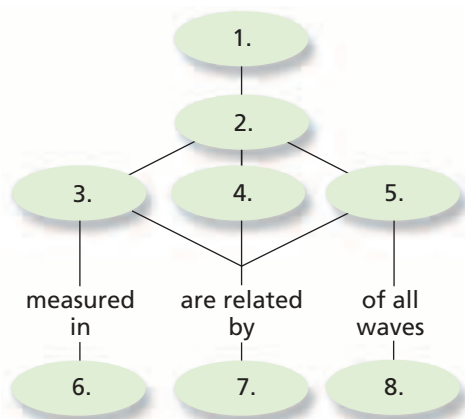




Go to the Chemistry Web site at [chemistrymc.com](http://chemistrymc.com) for additional Chapter 5 Assessment.

## Concept Mapping

29. Complete the concept map using the following terms: speed,  $c = \lambda\nu$ , electromagnetic waves, wavelength, characteristic properties, frequency,  $c$ , and hertz.



## Mastering Concepts

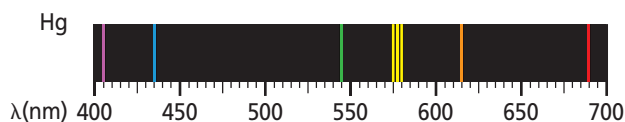
30. Define the following terms.
- |                     |                       |
|---------------------|-----------------------|
| a. frequency (5.1)  | c. quantum (5.1)      |
| b. wavelength (5.1) | d. ground state (5.2) |
31. Why did scientists consider Rutherford's nuclear model of the atom incomplete? (5.1)
32. Name one type of electromagnetic radiation. (5.1)
33. Explain how the gaseous neon atoms in a neon sign emit light. (5.1)
34. What is a photon? (5.1)
35. What is the photoelectric effect? (5.1)
36. Explain Planck's quantum concept as it relates to energy lost or gained by matter. (5.1)
37. How did Einstein explain the previously unexplainable photoelectric effect? (5.1)
38. Arrange the following types of electromagnetic radiation in order of increasing wavelength. (5.1)
- |                      |                |
|----------------------|----------------|
| a. ultraviolet light | c. radio waves |
| b. microwaves        | d. X rays      |
39. What is the difference between an atom's ground state and an excited state? (5.2)
40. According to the Bohr model, how do electrons move in atoms? (5.2)
41. What does  $n$  designate in Bohr's atomic model? (5.2)
42. Why are you unaware of the wavelengths of moving objects such as automobiles and tennis balls? (5.2)
43. What is the name of the atomic model in which electrons are treated as waves? Who first wrote the electron wave equations that led to this model? (5.2)
44. What is an atomic orbital? (5.2)
45. What is the probability that an electron will be found within an atomic orbital? (5.2)
46. What does  $n$  represent in the quantum mechanical model of the atom? (5.2)
47. How many energy sublevels are contained in each of the hydrogen atom's first three energy levels? (5.2)
48. What atomic orbitals are related to a p sublevel? To a d sublevel? (5.2)
49. Which of the following atomic orbital designations are incorrect? (5.2)
- |       |       |       |       |
|-------|-------|-------|-------|
| a. 7f | b. 3f | c. 2d | d. 6p |
|-------|-------|-------|-------|
50. What do the sublevel designations s, p, d, and f specify with respect to the atom's orbitals? (5.2)
51. What do subscripts such as  $y$  and  $xz$  tell you about atomic orbitals? (5.2)
52. What is the maximum number of electrons an orbital may contain? (5.2)
53. Why is it impossible to know precisely the velocity and position of an electron at the same time? (5.2)
54. What shortcomings caused scientists to finally reject Bohr's model of the atom? (5.2)
55. Describe de Broglie's revolutionary concept involving the characteristics of moving particles. (5.2)
56. How is an orbital's principal quantum number related to the atom's major energy levels? (5.2)
57. Explain the meaning of the aufbau principle as it applies to atoms with many electrons. (5.3)
58. In what sequence do electrons fill the atomic orbitals related to a sublevel? (5.3)
59. Why must the two arrows within a single block of an orbital diagram be written in opposite (up and down) directions? (5.3)
60. How does noble-gas notation shorten the process of writing an element's electron configuration? (5.3)
61. What are valence electrons? How many of a magnesium atom's 12 electrons are valence electrons? (5.3)

- 62.** Light is said to have a dual wave-particle nature. What does this statement mean? (5.3)
- 63.** Describe the difference between a quantum and a photon. (5.3)
- 64.** How many electrons are shown in the electron-dot structures of the following elements? (5.3)
- |           |            |
|-----------|------------|
| a. carbon | c. calcium |
| b. iodine | d. gallium |
- 74.** A photon has an energy of  $1.10 \times 10^{-13}$  J. What is the photon's wavelength? What type of electromagnetic radiation is it?
- 75.** How long does it take a radio signal from the Voyager spacecraft to reach Earth if the distance between Voyager and Earth is  $2.72 \times 10^9$  km?
- 76.** If your favorite FM radio station broadcasts at a frequency of 104.5 MHz, what is the wavelength of the station's signal in meters? What is the energy of a photon of the station's electromagnetic signal?

## Mastering Problems

### Wavelength, Frequency, Speed, and Energy (5.1)

- 65.** What is the wavelength of electromagnetic radiation having a frequency of  $5.00 \times 10^{12}$  Hz? What kind of electromagnetic radiation is this?
- 66.** What is the frequency of electromagnetic radiation having a wavelength of  $3.33 \times 10^{-8}$  m? What type of electromagnetic radiation is this?
- 67.** The laser in a compact disc (CD) player uses light with a wavelength of 780 nm. What is the frequency of this light?
- 68.** What is the speed of an electromagnetic wave having a frequency of  $1.33 \times 10^{17}$  Hz and a wavelength of 2.25 nm?
- 69.** Use **Figure 5-5** to determine each of the following types of radiation.
- |  |
|--|
| a. radiation with a frequency of $8.6 \times 10^{11} \text{ s}^{-1}$ |
| b. radiation with a wavelength 4.2 nm                                |
| c. radiation with a frequency of 5.6 MHz                             |
| d. radiation that travels at a speed of $3.00 \times 10^8$ m/s       |
- 70.** What is the energy of a photon of red light having a frequency of  $4.48 \times 10^{14}$  Hz?
- 71.** Mercury's atomic emission spectrum is shown below. Estimate the wavelength of the orange line. What is its frequency? What is the energy of an orange photon emitted by the mercury atom?



- 72.** What is the energy of an ultraviolet photon having a wavelength of  $1.18 \times 10^{-8}$  m?
- 73.** A photon has an energy of  $2.93 \times 10^{-25}$  J. What is its frequency? What type of electromagnetic radiation is the photon?


### Electron Configurations (5.3)


- 77.** List the aufbau sequence of orbitals from 1s to 7p.
- 78.** Write orbital notations and complete electron configurations for atoms of the following elements.
- |              |
|--------------|
| a. beryllium |
| b. aluminum  |
| c. nitrogen  |
| d. sodium    |
- 79.** Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.
- |       |       |
|-------|-------|
| a. Mn | f. W  |
| b. Kr | g. Pb |
| c. P  | h. Ra |
| d. Zn | i. Sm |
| e. Zr | j. Bk |
- 80.** What elements are represented by each of the following electron configurations?
- |   |
|---|
| a. $1s^2 2s^2 2p^5$                             |
| b. $[\text{Ar}] 4s^2$                           |
| c. $[\text{Xe}] 6s^2 4f^4$                      |
| d. $[\text{Kr}] 5s^2 4d^{10} 5p^4$              |
| e. $[\text{Rn}] 7s^2 5f^{13}$                   |
| f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ |
- 81.** Draw electron-dot structures for atoms of each of the following elements.
- |              |
|--------------|
| a. carbon    |
| b. arsenic   |
| c. polonium  |
| d. potassium |
| e. barium    |
- 82.** An atom of arsenic has how many electron-containing orbitals? How many of the orbitals are completely filled? How many of the orbitals are associated with the atom's  $n = 4$  principal energy level?

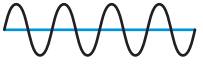



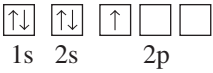
**Mixed Review**

Sharpen your problem-solving skills by answering the following.

- 83.** What is the frequency of electromagnetic radiation having a wavelength of 1.00 m?
- 84.** What is the maximum number of electrons that can be contained in an atom's orbitals having the following principal quantum numbers?  
**a.** 3      **b.** 4      **c.** 6      **d.** 7
- 85.** What is the wavelength of light with a frequency of  $5.77 \times 10^{14}$  Hz?
- 86.** Using the waves shown below, identify the wave or waves with the following characteristics.
- 1.** 

**2.** 

**3.** 

**4.** 
- a.** longest wavelength    **c.** largest amplitude  
**b.** greatest frequency    **d.** shortest wavelength
- 87.** How many orientations are possible for the orbitals related to each of the following sublevels?  
**a.** s      **b.** p      **c.** d      **d.** f
- 88.** Describe the electrons in an atom of nickel in the ground state using the electron configuration notation and the noble-gas notation.
- 89.** Which of the following elements have two electrons in their electron-dot structures: hydrogen, helium, lithium, aluminum, calcium, cobalt, bromine, krypton, and barium?
- 90.** In Bohr's atomic model, what electron orbit transition produces the blue-green line in hydrogen's atomic emission spectrum?
- 91.** A zinc atom contains a total of 18 electrons in its 3s, 3p, and 3d orbitals. Why does its electron-dot structure show only two dots?
- 92.** An X-ray photon has an energy of  $3.01 \times 10^{-18}$  J. What is its frequency and wavelength?
- 93.** Which element has the following orbital diagram?  
  
 1s   2s      2p
- 94.** Which element has the ground-state electron configuration represented by the noble-gas notation  $[\text{Rn}]7s^1$ ?
- 95.** How many photons of infrared radiation having a frequency of  $4.88 \times 10^{13}$  Hz are required to provide an energy of 1.00 J?

**Thinking Critically**

- 96. Comparing and Contrasting** Briefly discuss the difference between an orbit in Bohr's model of the atom and an orbital in the quantum mechanical view of the atom.
- 97. Applying Concepts** Scientists use atomic emission spectra to determine the elements in materials of unknown composition. Explain what makes this method possible.
- 98. Using Numbers** It takes  $8.17 \times 10^{-19}$  J of energy to remove one electron from a gold surface. What is the maximum wavelength of light capable of causing this effect?
- 99. Drawing a Conclusion** The elements aluminum, silicon, gallium, germanium, arsenic, selenium are all used in making various types of semiconductor devices. Write electron configurations and electron-dot structures for atoms of each of these elements. What similarities among the elements' electron configurations do you notice?

**Writing in Chemistry**

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

**Cumulative Review**

Refresh your understanding of previous chapters by answering the following.

- 101.** Round 20.561 20 g to three significant figures. (Chapter 2)
- 102.** Identify each of the following as either chemical or physical properties of the substance. (Chapter 3)  
**a.** mercury is a liquid at room temperature  
**b.** sucrose is a white, crystalline solid  
**c.** iron rusts when exposed to moist air  
**d.** paper burns when ignited
- 103.** Identify each of the following as a pure substance or a mixture. (Chapter 3)  
**a.** distilled water                      **d.** diamond  
**b.** orange juice with pulp          **e.** milk  
**c.** smog                                      **f.** copper metal
- 104.** An atom of gadolinium has an atomic number of 64 and a mass number of 153. How many electrons, protons, and neutrons does it contain? (Chapter 4)

# STANDARDIZED TEST PRACTICE

## CHAPTER 5

Use these questions and the test-taking tip to prepare for your standardized test.

- Cosmic rays are high-energy radiation from outer space. What is the frequency of a cosmic ray that has a wavelength of  $2.67 \times 10^{-13}$  m when it reaches Earth? (The speed of light is  $3.00 \times 10^8$  m/s.)
  - $8.90 \times 10^{-22} \text{ s}^{-1}$
  - $3.75 \times 10^{12} \text{ s}^{-1}$
  - $8.01 \times 10^{-5} \text{ s}^{-1}$
  - $1.12 \times 10^{21} \text{ s}^{-1}$
- Wavelengths of light shorter than about  $4.00 \times 10^{-7}$  m are not visible to the human eye. What is the energy of a photon of ultraviolet light having a frequency of  $5.45 \times 10^{16} \text{ s}^{-1}$ ? (Planck's constant is  $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ .)
  - $3.61 \times 10^{-17} \text{ J}$
  - $1.22 \times 10^{-50} \text{ J}$
  - $8.23 \times 10^{49} \text{ J}$
  - $3.81 \times 10^{-24} \text{ J}$

**Interpreting Charts** Use the periodic table and the chart below to answer questions 3–6.

Electron Configurations for Selected Transition Metals			
Element	Symbol	Atomic number	Electron configuration
Vanadium	V	23	$[\text{Ar}]4s^23d^3$
Yttrium	Y	39	$[\text{Kr}]5s^24d^1$
_____	___	___	$[\text{Xe}]6s^24f^{14}5d^6$
Scandium	Sc	21	$[\text{Ar}]4s^23d^1$
Cadmium	Cd	48	_____

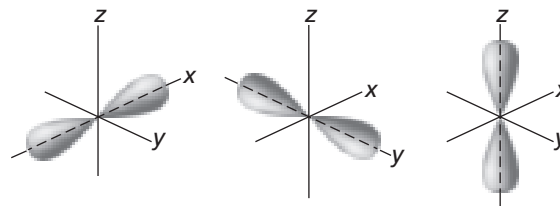
- Using noble-gas notation, what is the ground-state electron configuration of Cd?
  - $[\text{Kr}]4d^{10}4f^2$
  - $[\text{Ar}]4s^23d^{10}$
  - $[\text{Kr}]5s^24d^{10}$
  - $[\text{Xe}]5s^24d^{10}$
- What is the element that has the ground-state electron configuration  $[\text{Xe}]6s^24f^{14}5d^6$ ?
  - La
  - Ti
  - W
  - Os
- What is the complete electron configuration of a scandium atom?
  - $1s^22s^22p^63s^23p^64s^23d^1$
  - $1s^22s^22p^73s^23p^74s^23d^1$
  - $1s^22s^22p^53s^23p^54s^23d^1$
  - $1s^22s^12p^73s^13p^74s^23d^1$

- Which of the following is the correct orbital diagram for the third and fourth principal energy levels of vanadium?
  - 
  - 
  - 
  -

- Which of the following orbitals has the highest energy?
  - 4f
  - 5p
  - 6s
  - 3d

- What is the electron-dot structure for indium?
  - $\cdot\text{In}$
  - $\cdot\text{In}\cdot$
  - $\cdot\text{In}\cdot$
  - $\cdot\text{In}\cdot$

- The picture below shows all of the orbitals related to one type of sublevel. The type of sublevel to which these orbitals belong is \_\_\_\_\_.



- s
  - p
  - d
  - f
- What is the maximum number of electrons related to the fifth principal energy level of an atom?
    - 10
    - 20
    - 25
    - 50

### TEST-TAKING TIP

**Do Some Reconnaissance** Find out what the conditions will be for taking the test. Is it timed or untimed? Can you eat a snack at the break? Can you use a calculator or other tools? Will those tools be provided? Will mathematical constants be given? Know these things in advance so that you can practice taking tests under the same conditions.

