



# Electrons in Atoms

**BIG Idea** The atoms of each element have a unique arrangement of electrons.

## 5.1 Light and Quantized Energy

**MAIN Idea** Light, a form of electromagnetic radiation, has characteristics of both a wave and a particle.

## 5.2 Quantum Theory and the Atom

**MAIN Idea** Wavelike properties of electrons help relate atomic emission spectra, energy states of atoms, and atomic orbitals.

## 5.3 Electron Configuration

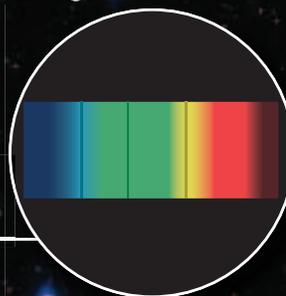
**MAIN Idea** A set of three rules can be used to determine electron arrangement in an atom.

### ChemFacts

- Scientists use stellar absorption spectra to identify a star's elemental composition and classify it into one of the several spectral types.
- The properties of absorption spectra are also related to a star's surface temperature.
- Stellar spectra revealed that stars are made out of the same elements as those found on Earth.
- There are around 600 dark lines in the Sun's absorption spectra.



Absorption spectrum of Betelgeuse



Absorption spectrum of Rigel

# Start-Up Activities

## LAUNCH Lab

### How do you know what is inside an atom?

Imagine that it is your birthday, and there is one wrapped present that is different from all the rest. Unlike the other gifts that you can open, you can only guess what is inside this package. In trying to determine the structure of the atom, early chemists had a similar experience. How good are your skills of observation and deduction?



#### Procedure

1. Read and complete the lab safety form.
2. Obtain a **wrapped box** from your instructor.
3. Using as many observation methods as you can, and without unwrapping or opening the box, try to determine what is inside the box.
4. Record the observations you make throughout this discovery process.

#### Analysis

1. **Describe** how you were able to determine characteristics such as the size, shape, and composition of the object in the box.
2. **Indicate** what senses you used to make your observations.
3. **Discuss** why it is hard to determine what type of object is in the box without opening it.

**Inquiry** After reading the chapter, design another investigation that illustrates the difficulties associated with the study of subatomic particles.

## FOLDABLES™ Study Organizer

**Electron Configuration** Make a Foldable to help you summarize the three rules that define how electrons are arranged in an atom.

- ▶ **STEP 1** Fold a sheet of paper in half lengthwise. Make the back edge about 2 cm longer than the front edge.



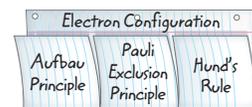
- ▶ **STEP 2** Fold into thirds.



- ▶ **STEP 3** Unfold and cut along the folds of the top flap to make three tabs.



- ▶ **STEP 4** Label the tabs as follows: *Aufbau Principle*, *Pauli Exclusion Principle*, and *Hund's Rule*.



**FOLDABLES** Use this Foldable with Sections 5.3. As you read this section, summarize each rule under the appropriate tab.

## Chemistry Online

Visit [glencoe.com](http://glencoe.com) to:

- ▶ study the entire chapter online
- ▶ explore **Concepts In Motion**
- ▶ take Self-Check quizzes
- ▶ use the Personal Tutor to work Example Problems step-by-step
- ▶ access Web Links for more information, projects, and activities
- ▶ find the Try at Home Lab, Observing Light's Wave Nature

## Section 5.1

### Objectives

- ▶ **Compare** the wave and particle natures 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.

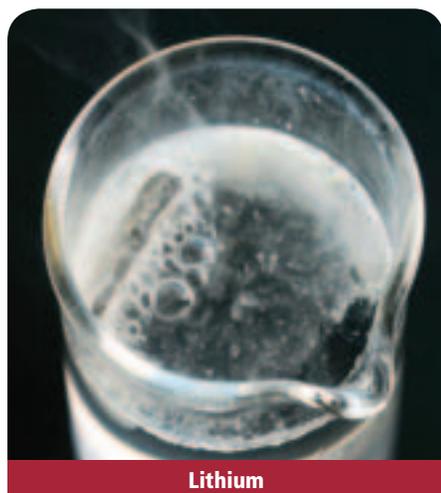
### Review Vocabulary

**radiation:** the rays and particles—alpha particles, beta particles, and gamma rays—that are emitted by radioactive material

### New Vocabulary

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

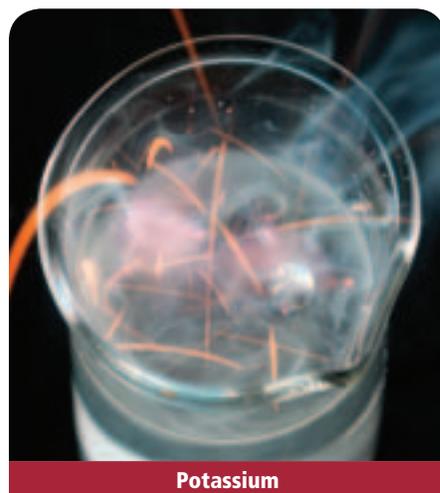
■ **Figure 5.1** Different elements can have similar reactions with water.



Lithium



Sodium



Potassium

## Light and Quantized Energy

**MAIN Idea** Light, a form of electromagnetic radiation, has characteristics of both a wave and a particle.

**Real-World Reading Link** Have you ever come inside on a cold day, headed for the kitchen, and popped a cold snack into the microwave oven? When the microwaves reached your snack, small packets of energy warmed it in practically no time at all.

### The Atom and Unanswered Questions

After discovering three subatomic particles in the early 1900s, scientists continued their quest to understand atomic structure and the arrangement of electrons within atoms.

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. 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. Rutherford's nuclear model did not begin to account for the differences and similarities in chemical behavior among the various elements.

For example, consider the elements lithium, sodium, and potassium, which are found in different periods on the periodic table but have similar chemical behaviors. All three elements appear metallic in nature, and their atoms react vigorously with water to liberate hydrogen gas. In fact, as shown in **Figure 5.1**, both sodium and potassium react so violently that the hydrogen gas can ignite and even explode.

In the early 1900s, scientists began to unravel the puzzle of chemical behavior. They 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. To understand this relationship and the nature of atomic structure, it will be helpful to first understand the nature of light.

## The Wave Nature of Light

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

**Characteristics of waves** All waves can be described by several characteristics, a few of which you might be familiar with from everyday experience. You might have seen concentric waves when dropping an object into water, as shown in **Figure 5.2a**.

The **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 ( $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$ ).

The **frequency** (represented by  $\nu$ , the Greek letter nu) is the number of waves that pass a given 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, (1/s) or ( $\text{s}^{-1}$ ). When expressed in this way, the term *waves* is understood. For example,  $652 \text{ Hz} = 652 \text{ waves/second} = 652/\text{s} = 652 \text{ s}^{-1}$ .

The **amplitude** of a wave is the wave's height from the origin to a crest, or from the origin to a trough, as illustrated in **Figure 5.2b**. Wavelength and frequency do not affect the amplitude of a wave.

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 ( $\nu$ ).

### Electromagnetic Wave Relationship

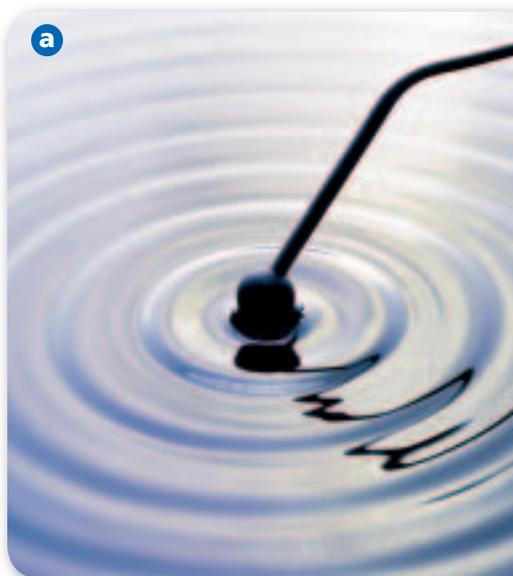
$$c = \lambda\nu$$

$c$  is the speed of light in a vacuum.

$\lambda$  is the wavelength.

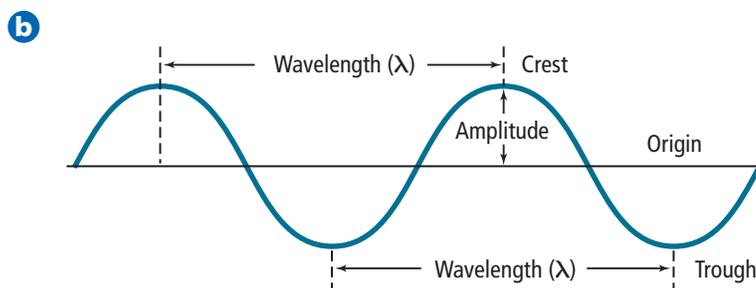
$\nu$  is the frequency.

The product of the frequency and the wavelength is equal to the speed of light in a vacuum.



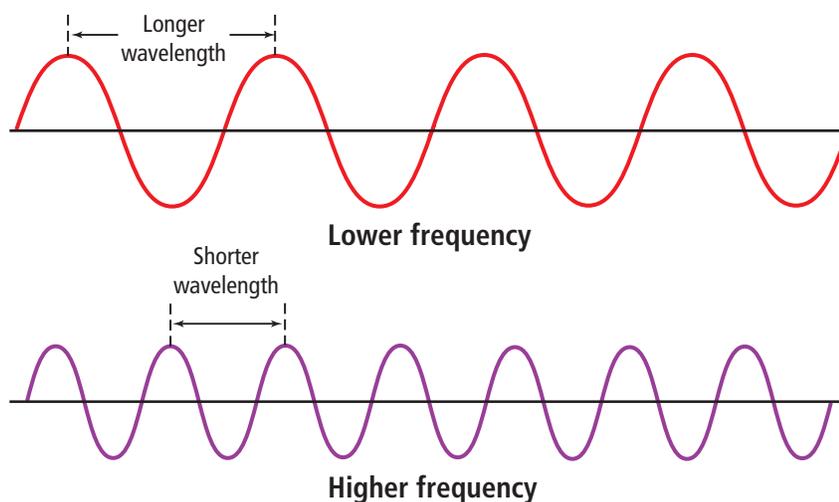
■ **Figure 5.2 a.** The concentric waves in the water show the characteristic properties of all waves. **b.** Amplitude, wavelength, and frequency are the main characteristics of waves.

**Identify** a crest, a trough, and one wavelength in the photo.



■ **Figure 5.3** These waves illustrate the relationship between wavelength and frequency. As wavelength increases, frequency decreases.

**Infer** Does frequency or wavelength affect amplitude?



Although the speed of all electromagnetic waves in a vacuum is the same, waves can have different wavelengths and frequencies. As you can see from the equation on the previous page, wavelength and frequency are inversely related; in other words, as one quantity increases, the other decreases. To better understand this relationship, examine the two waves illustrated in **Figure 5.3**. Although both waves travel at the speed of light, you can see that the red wave has a longer wavelength and lower frequency than the violet wave.

**Electromagnetic spectrum** Sunlight, which is one example of white light, contains a nearly continuous range of wavelengths and frequencies. White light passing through a prism separates into a continuous spectrum of colors similar to the spectrum in **Figure 5.4**. These are the colors of the visible spectrum. The spectrum is called continuous because each point of it corresponds to a unique wavelength and frequency. You might be familiar with the colors of the visible spectrum. 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 spectrum that arches across the sky.

■ **Figure 5.4** When white light passes through a prism, it is separated into a continuous spectrum of its different components—red, orange, yellow, green, blue, indigo, and violet light.



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, includes 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 bend varies with the wavelengths as they pass through the prism, resulting in the sequence of the colors red, orange, yellow, green, blue, indigo, and violet. In examining the energy of the radiation shown in **Figure 5.5**, 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.

Because all electromagnetic waves travel at the same speed in a given medium, you can use the formula  $c = \lambda\nu$  to calculate the wavelength or frequency of any wave.

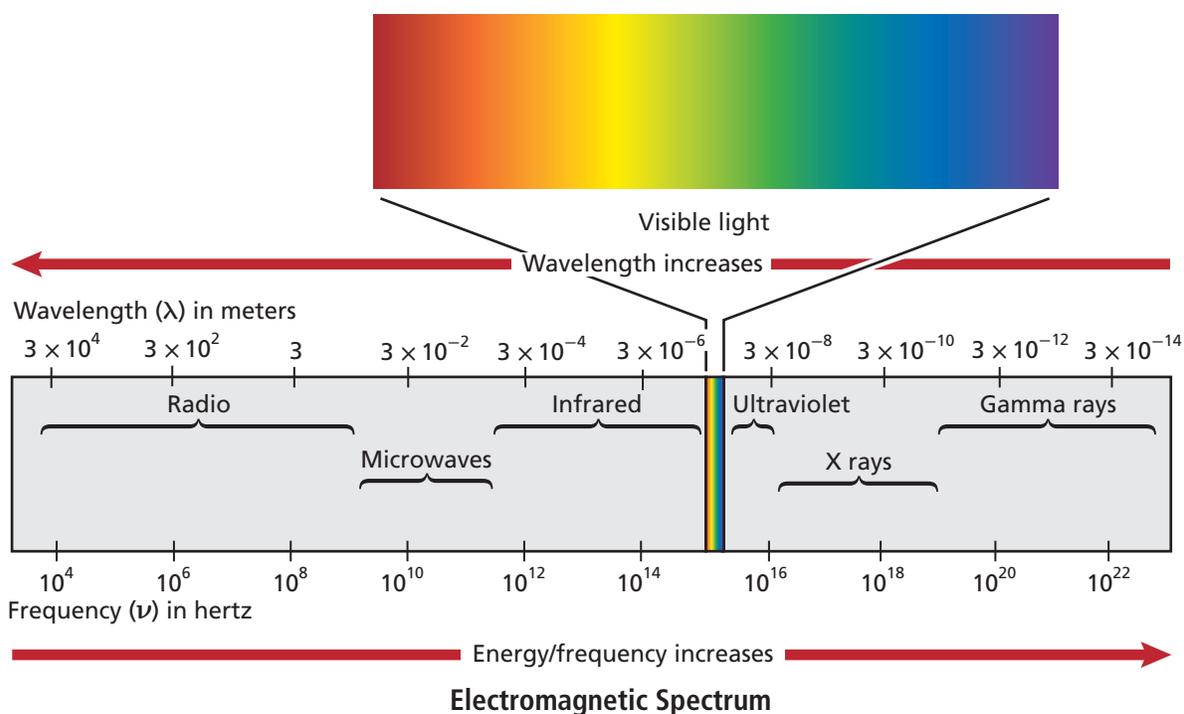
**✓ Reading Check State** the relationship between the energy and the frequency of electromagnetic radiation.

**Connection to Physics** Electromagnetic radiation from diverse origins constantly bombards us. In addition to the radiation from the Sun, human activities also produce radiation which include radio and TV signals, phone relay stations, lightbulbs, medical X-ray equipment, and particle accelerators. Natural sources on Earth, such as lightning, natural radioactivity, and even the glow of fireflies, also contribute. Our knowledge of the universe is based on electromagnetic radiation emitted by distant objects and detected with instruments on Earth.

## CAREERS IN CHEMISTRY

**Spectroscopist** Spectroscopy is the study of the spectra absorbed or emitted by matter. Because each element's spectrum is unique, it is like a fingerprint. Astrophysicists use spectroscopy to investigate what a star, such as the Sun, is made of. A star's absorption spectrum shows many dark lines, which allow spectroscopists to identify the elements present in the star. For more information on chemistry careers, visit [glencoe.com](http://glencoe.com).

■ **Figure 5.5** The electromagnetic spectrum covers a wide range of frequencies. The visible-light section of the spectrum is very narrow. As frequency and energy increase, wavelength decreases.



## EXAMPLE Problem 5.1

### Math Handbook

Solving Algebraic  
Equations  
pages 954–955

**Calculating Wavelength of an EM Wave** Microwaves are used to cook food and transmit information. What is the wavelength of a microwave that has 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 speeds, frequencies, and wavelengths 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

$$\nu = 3.44 \times 10^9 \text{ Hz}$$

$$c = 3.00 \times 10^8 \text{ m/s}$$

#### Unknown

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

### 2 Solve for the Unknown

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

$$c = \lambda\nu$$

State the electromagnetic wave relationship.

$$\lambda = c/\nu$$

Solve for  $\lambda$ .

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

Substitute  $c = 3.00 \times 10^8$  m/s and  $\nu = 3.44 \times 10^9$  Hz.

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 numbers and units.

$$\lambda = 8.72 \times 10^{-2} \text{ m}$$

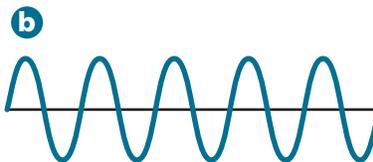
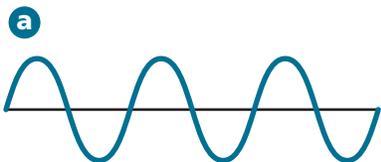
### 3 Evaluate the Answer

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**.

## PRACTICE Problems

Extra Practice Page 978 and [glencoe.com](http://glencoe.com)

- Objects get their colors from reflecting only certain wavelengths when hit with white light. Light reflected from a green leaf is found to have a wavelength of  $4.90 \times 10^{-7}$  m. What is the frequency of the light?
- X rays can penetrate body tissues and are widely used to diagnose and treat disorders of internal body structures. What is the frequency of an X ray with a wavelength of  $1.15 \times 10^{-10}$  m?
- After careful analysis, an electromagnetic wave is found to have a frequency of  $7.8 \times 10^6$  Hz. What is the speed of the wave?
- Challenge** While an FM radio station broadcasts at a frequency of 94.7 MHz, an AM station broadcasts at a frequency of 820 kHz. What are the wavelengths of the two broadcasts? Which of the drawings below corresponds to the FM station? To the AM station?



## The Particle Nature of Light

While considering light as a wave explains 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 light of a specific frequency shines on them. Scientists realized that a new model or a revision of the wave model of light was needed to address these phenomena.

**The quantum concept** When objects are heated, they emit glowing light. **Figure 5.6** illustrates this phenomenon with iron. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and turns orange, then bluish in color at even higher temperatures. As you will learn in later chapters, 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. In 1900, German physicist Max Planck (1858–1947) began searching for an explanation of this phenomenon as he studied the light emitted by heated objects. His study led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. A **quantum** is the minimum amount of energy that can be gained or lost by an atom.

 **Reading Check** Explain why the color of heated objects changes with their temperature.

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 think 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.



### VOCABULARY

#### ACADEMIC VOCABULARY

##### Phenomenon

an observable fact or event

*During rainstorms, electric currents often pass from the sky to Earth—a phenomenon we call lightning.*

■ **Figure 5.6** The wavelength of the light emitted by heated metal, such as the iron at left, depends on the temperature. At room temperature, iron is gray. When heated, it first turns red, then glowing orange.

**Explain** the relationship between the color and the temperature of the metal.

Planck proposed that the energy emitted by hot objects was quantized. He then went further and demonstrated mathematically that a relationship exists between the energy of a quantum and the frequency of the emitted radiation.

### Energy of a Quantum

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

$E_{\text{quantum}}$  represents energy.

$h$  is Planck's constant.

$\nu$  represents frequency.

The energy of a quantum is given by the product of Planck's constant and the frequency.

## Real-World Chemistry The Photoelectric Effect



**Solar energy** is sometimes used to power road signs. Photovoltaic cells use the photoelectric effect to convert the energy of light into electric energy.

**Planck's constant** has a value of  $6.626 \times 10^{-34}$  J·s, where J is the symbol for joule, the SI unit of energy. The equation shows that the energy of radiation increases as the radiation's frequency,  $\nu$ , increases.

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 whole numbers of blocks. Similarly, matter can have only certain amounts of energy—quantities of energy between these values do not exist.

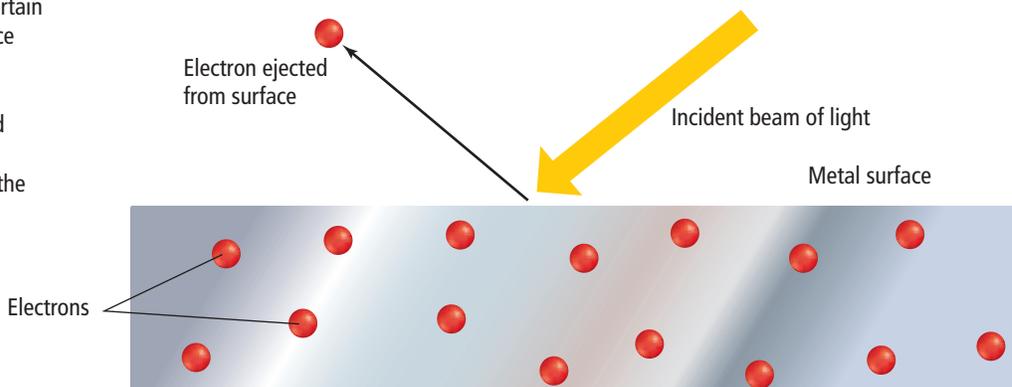
**The photoelectric effect** Scientists also knew that the wave model of light could not explain a phenomenon called the photoelectric effect. In the **photoelectric effect**, electrons, called photoelectrons, are emitted from a metal's surface when light of a certain frequency, or higher than a certain frequency, shines on the surface, as shown in **Figure 5.7**.

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. In reality, a metal will not eject photoelectrons below a specific frequency of incident light. For example, no matter how intensely or how long it shines, light with a frequency less than  $1.14 \times 10^{15}$  Hz does not eject photoelectrons from silver. But even dim light with a frequency equal to or greater than  $1.14 \times 10^{15}$  Hz ejects photoelectrons from silver.



**Reading Check** Describe the photoelectric effect.

■ **Figure 5.7** The photoelectric effect occurs when light of a certain frequency strikes a metal surface and ejects electrons. When the intensity of the light increases, the number of electrons ejected increases. When the frequency (energy) of the light increases, the energy of the ejected electrons increases.



**Light's dual nature** To explain the photoelectric effect, Albert Einstein proposed in 1905 that light has a dual nature. A beam of light has wavelike and particlelike properties. It can be thought of as a beam of bundles of energy called photons. A **photon** is a massless particle that carries a quantum of energy. Extending Planck's idea of quantized energy, Einstein calculated that a photon's energy depends on its frequency.

### Energy of a Photon

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

$E_{\text{photon}}$  represents energy.  
 $h$  is Planck's constant.  
 $\nu$  represents frequency.

The energy of a photon is given by the product of Planck's constant and the frequency.

Einstein also proposed that the energy of a photon must have a certain threshold value to cause the ejection of a photoelectron from the surface of the metal. Thus, even small numbers of photons with energy above the threshold value will cause the photoelectric effect. Einstein won the Nobel Prize in Physics in 1921 for this work.

### EXAMPLE Problem 5.2

**Calculate the Energy of a Photon** Every object gets its color by reflecting a certain portion of incident light. The color is determined by the wavelength of the reflected photons, thus by their energy. What is the energy of a photon from the violet portion of the Sun's light if it has a frequency of  $7.230 \times 10^{14} \text{ s}^{-1}$ ?

#### Math Handbook

Operations with  
Scientific Notation  
page 948

#### 1 Analyze the Problem

##### Known

$$\nu = 7.230 \times 10^{14} \text{ s}^{-1}$$

$$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$$

##### Unknown

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

#### 2 Solve for the Unknown

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

State the equation for the energy of a photon.

$$E_{\text{photon}} = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(7.230 \times 10^{14} \text{ s}^{-1})$$

Substitute  $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$  and  $\nu = 7.230 \times 10^{14} \text{ s}^{-1}$ .

$$E_{\text{photon}} = 4.791 \times 10^{-19} \text{ J}$$

Multiply and divide numbers and units.

#### 3 Evaluate the Answer

As expected, the energy of a single photon of light is extremely small. The unit is joules, an energy unit, and there are four significant figures.

### PRACTICE Problems

Extra Practice Page 978 and [glencoe.com](http://glencoe.com)

- Calculate the energy possessed by a single photon of each of the following types of electromagnetic radiation.
  - $6.32 \times 10^{20} \text{ s}^{-1}$
  - $9.50 \times 10^{13} \text{ Hz}$
  - $1.05 \times 10^{16} \text{ s}^{-1}$
- The blue color in some fireworks occurs when copper(I) chloride is heated to approximately 1500 K and emits blue light of wavelength  $4.50 \times 10^2 \text{ nm}$ . How much energy does one photon of this light carry?
- Challenge** The microwaves used to heat food have a wavelength of 0.125 m. What is the energy of one photon of the microwave radiation?

## MiniLab

### Identify Compounds

How do flame colors vary for different elements?

**Procedure** 

1. Read and complete the lab safety form.
2. Dip one of six **cotton swabs** into the **lithium chloride** solution. Put the swab into the flame of a **Bunsen burner**. Observe the color of the flame, and record it in your data table.
3. Repeat Step 2 for each of the metallic chloride solutions (**sodium chloride, potassium chloride, calcium chloride, and strontium chloride**). Record the color of each flame in your data table.
4. Compare your results to the flame tests shown in the Elements Handbook.
5. Repeat Step 2 using a sample of **unknown solution** obtained from your teacher. Record the color of the flame produced.
6. Dispose of the used cotton swabs as directed by your teacher.

#### Analysis

1. **Suggest** a reason why each compound produced a flame of a different color, even though they each contain chlorine.
2. **Explain** how an element's flame test might be related to its atomic emission spectrum.
3. **Infer** the identity of the unknown crystals. Explain your reasoning.

## Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? This process is 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 atoms return to their stable state by emitting light to release that energy. 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 corresponding to the frequencies of the radiation emitted by the atoms of neon. It is not a continuous range of colors, as in the visible spectrum of white light.

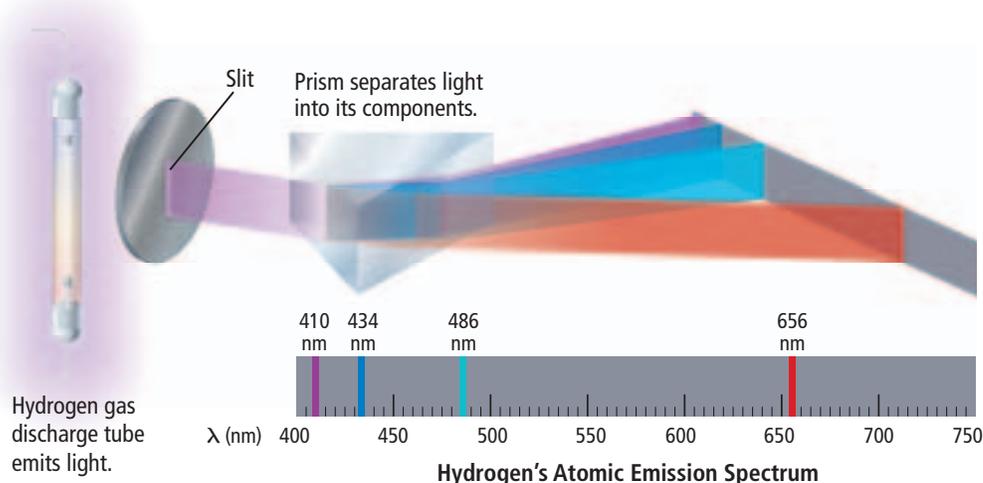
 **Reading Check** Explain how an emission spectrum is produced.

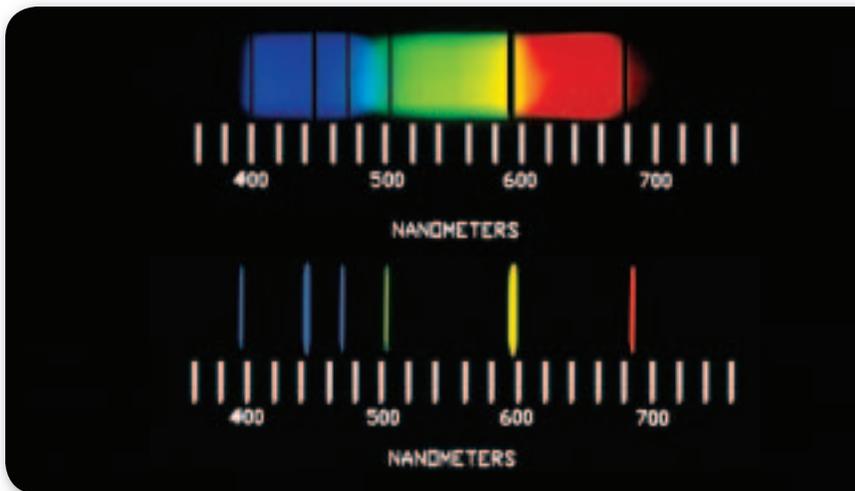
Each element's atomic emission spectrum is unique and can be used to identify an element or determine whether that element is part of an unknown compound. For example, when a platinum wire is dipped into a strontium nitrate solution and then inserted into a burner flame, the strontium atoms emit a characteristic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

**Figure 5.8** 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.

■ **Figure 5.8** The purple light emitted by hydrogen can be separated into its different components using a prism. Hydrogen has an atomic emission spectrum that comprises four lines of different wavelengths.

**Determine** Which line has the highest energy?





■ **Figure 5.9** The first spectrum is an absorption spectrum. It is composed of black lines on a continuous spectrum. The black lines correspond to certain frequencies absorbed by a given element, helium in this case. They can be matched to the colored lines present in helium's emission spectrum, shown below the absorption spectrum.

**Connection to Astronomy** 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 specific frequencies of light are emitted. Because those emitted frequencies are related to energy by the formula  $E_{\text{photon}} = h\nu$ , only photons with specific energies are emitted. This was not predicted by the laws of classical physics. Scientists had expected to observe the emission of a continuous series of colors as excited electrons lost energy. Elements absorb the same specific frequencies of light as the frequencies they emit, thus creating an absorption spectrum. In an absorption spectrum, the absorbed frequencies appear as black lines, as shown in **Figure 5.9**. By comparing the black lines to the emission spectrum of elements, scientists are able to determine the composition of the outer layers of stars.

## Section 5.1 Assessment

### Section Summary

- ▶ All waves are defined by their wavelengths, frequencies, amplitudes, and speeds.
- ▶ In a vacuum, all electromagnetic waves travel at the speed of light.
- ▶ All electromagnetic waves have both wave and particle properties.
- ▶ Matter emits and absorbs energy in quanta.
- ▶ White light produces a continuous spectrum. An element's emission spectrum consists of a series of lines of individual colors.

8. **MAIN Idea** Compare the dual nature of light.
9. **Describe** the phenomena that can be explained only by the particle model of light.
10. **Compare and contrast** continuous spectrum and emission spectrum.
11. **Assess** Employ quantum theory to assess the amount of energy that matter gains and loses.
12. **Discuss** the way in which Einstein utilized Planck's quantum concept to explain the photoelectric effect.
13. **Calculate** Heating 235 g of water from 22.6°C to 94.4°C in a microwave oven requires  $7.06 \times 10^4$  J of energy. If the microwave frequency is  $2.88 \times 10^{10} \text{ s}^{-1}$ , how many quanta are required to supply the  $7.06 \times 10^4$  J?
14. **Interpret Scientific Illustrations** Use **Figure 5.5** and your knowledge of electromagnetic radiation to match the numbered items with the lettered items. The numbered items may be used more than once or not at all.
 

a. longest wavelength	1. gamma ray
b. highest frequency	2. infrared wave
c. greatest energy	3. radio waves

## Section 5.2

### 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 current view of electrons in atoms.
- ▶ **Identify** the relationships among a hydrogen atom's energy levels, sublevels, and atomic orbitals.

### Review Vocabulary

**atom:** the smallest particle of an element that retains all the properties of that element; is composed of electrons, protons, and neutrons

### New Vocabulary

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

# Quantum Theory and the Atom

**MAIN Idea** Wavelike properties of electrons help relate atomic emission spectra, energy states of atoms, and atomic orbitals.

**Real-World Reading Link** Imagine climbing a ladder and trying to stand between the rungs. Unless you could stand on air, it would not work. When atoms are in various energy states, electrons behave in much the same way as a person climbing up the rungs of a ladder.

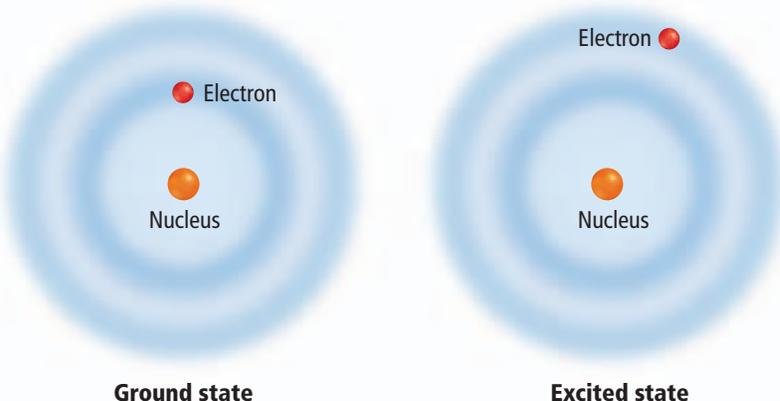
## Bohr's Model of the Atom

The dual wave-particle model of light accounted for several previously unexplainable phenomena, but scientists still did not understand the relationships among atomic structure, electrons, and atomic emission spectra. Recall that hydrogen's atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are the atomic emission spectra of elements discontinuous rather than continuous? Niels Bohr, a Danish physicist working in Rutherford's laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. 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, 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.

Bohr also related the hydrogen atom's energy states to the electron within the atom. He suggested that the 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. Thus, a hydrogen atom can have many different excited states, although it contains only one electron. Bohr's idea is illustrated in **Figure 5.10**.

■ **Figure 5.10** The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.



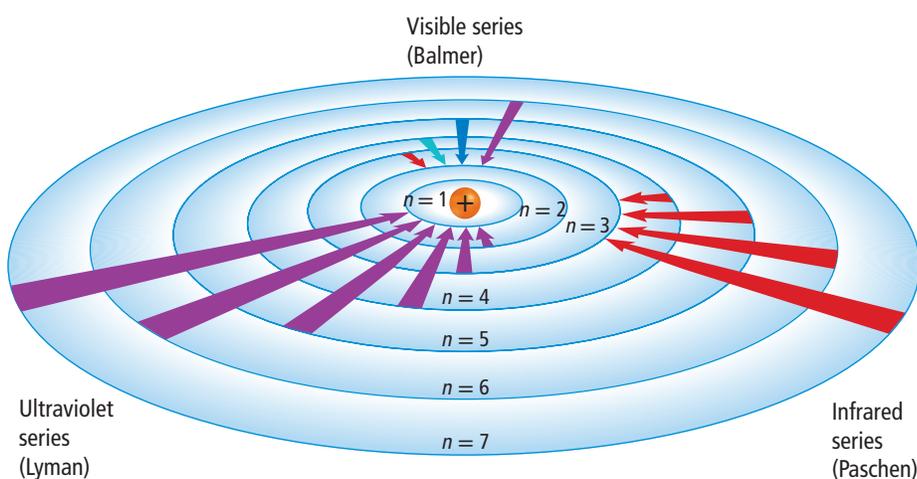
**Table 5.1 Bohr's Description of the Hydrogen Atom**

Bohr's 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	$E_7 = 49E_1$

In order to complete his calculations, Bohr assigned a number,  $n$ , called a **quantum number**, to each orbit. He also calculated the radius of each orbit. 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 allowed orbits and energy levels is given in **Table 5.1**.

**The hydrogen line spectrum** Bohr suggested that the hydrogen atom is in the ground state, also called the first energy level, when its single 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 5.11**. 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 energy difference between the two levels.

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



**Figure 5.11** When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to  $n = 1$ ,  $n = 2$ , and  $n = 3$ , respectively.



Concepts In Motion

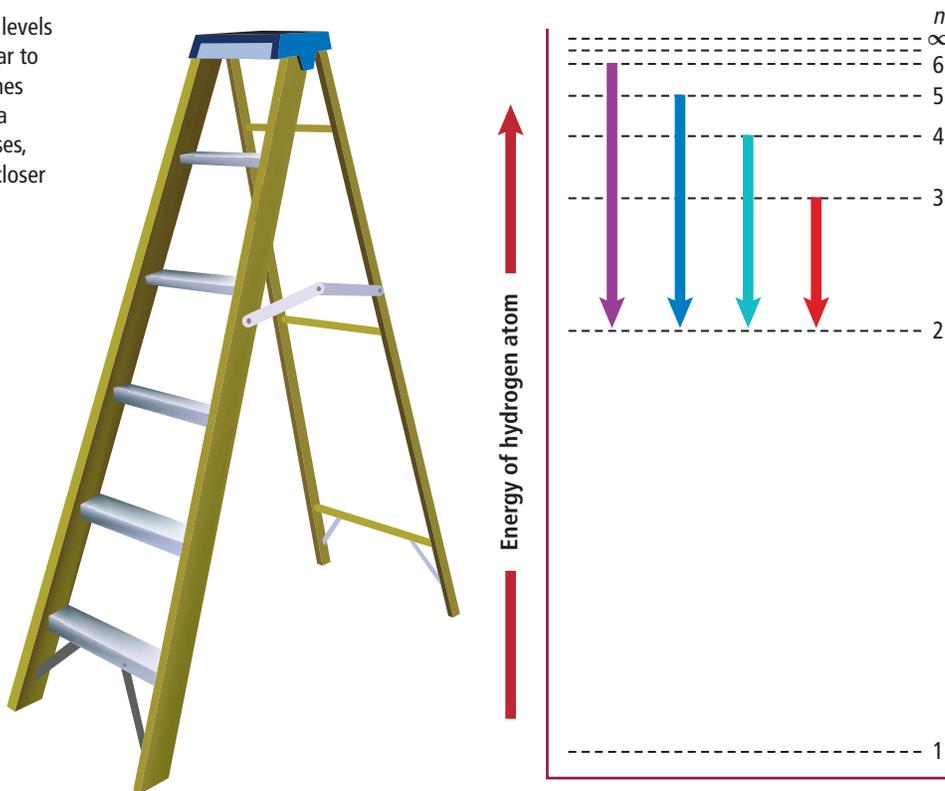
**Interactive Figure** To see an animation of the Balmer Series, visit [glencoe.com](http://glencoe.com).

■ **Figure 5.12** Only certain energy levels are allowed. The energy levels are similar to the rungs of a ladder. The four visible lines correspond to electrons dropping from a higher  $n$  to the orbit  $n = 2$ . As  $n$  increases, the hydrogen atom's energy levels are closer to each other.



### Concepts in Motion

**Interactive Figure** To see an animation of electron transitions, visit [glencoe.com](http://glencoe.com).



Because only certain atomic energies are possible, only certain frequencies of electromagnetic radiation can be emitted. You might compare hydrogen's atomic energy states to 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, corresponding to the energy difference between the two orbits.

**Figure 5.12** shows that, unlike rungs on a ladder, however, the hydrogen atom's energy levels are not evenly spaced. **Figure 5.12** also illustrates the four electron transitions that account for visible lines in hydrogen's atomic emission spectrum, shown in **Figure 5.8**. Electron transitions from higher-energy orbits to the second orbit account for all of hydrogen's visible lines, which form 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.



**Reading Check** Explain why different colors of light result from electron behavior in the atom.

**The limits of Bohr's model** Bohr's model explained hydrogen's observed spectral lines. 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.

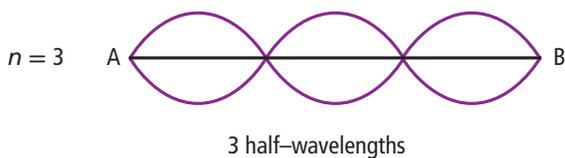
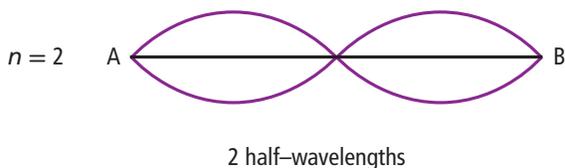
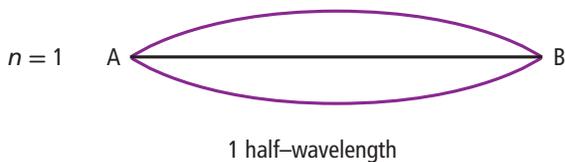
## The Quantum Mechanical Model of the Atom

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 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.

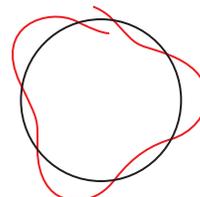
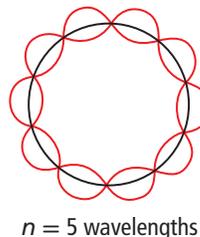
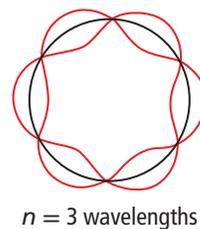
**Electrons as waves** De Broglie had been thinking that Bohr’s quantized electron orbits had characteristics similar to those of waves. For example, as **Figures 5.13a** and **5.13b** show, only multiples of half-wavelengths are possible on a plucked harp string because the string is fixed at both ends. Similarly, de Broglie saw that only odd numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in **Figure 5.13c**. 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 5.13 a.** The string on the harp vibrates between two fixed endpoints. **b.** The vibrations of a string between the two fixed endpoints labeled *A* and *B* are limited to multiples of half-wavelengths. **c.** Electrons on circular orbits can only have odd numbers of wavelengths.



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



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

The **de Broglie equation** predicts that all moving particles have wave characteristics. It also explains why it is impossible to notice the wavelength of a fast-moving car. 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. 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. De Broglie knew that if an electron has wavelike motion and is restricted to circular orbits of fixed radius, only certain wavelengths, frequencies, and energies are possible. Developing his idea, de Broglie derived the following equation.

### Particle Electromagnetic–Wave Relationship

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

$\lambda$  represents wavelength.  
 $h$  is Planck's constant.  
 $m$  represents mass of the particle.  
 $v$  represents frequency.

The wavelength of a particle is the ratio of Planck's constant, and the product of the particle's mass by its frequency.

## PROBLEM-SOLVING LAB

### Interpret Scientific Illustrations

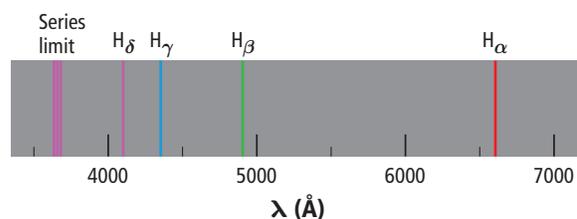
**What electron transitions account for the Balmer series?** Hydrogen's emission spectrum comprises three series of lines. Some wavelengths are ultraviolet (Lyman series) and infrared (Paschen series). Visible wavelengths comprise the Balmer series. The Bohr atomic model attributes these spectral lines to transitions from higher-energy states with electron orbits in which  $n = n_i$  to lower-energy states with smaller electron orbits in which  $n = n_f$ .

#### Analysis

The image at right illustrates some of the transitions in hydrogen's Balmer series. These Balmer lines are designated  $H_\alpha$  (6562 Å),  $H_\beta$  (4861 Å),  $H_\gamma$  (4340 Å), and  $H_\delta$  (4101 Å). Each wavelength ( $\lambda$ ) is related to an electron transition within a hydrogen atom by the following equation, in which  $1.09678 \times 10^7 \text{ m}^{-1}$  is known as the Rydberg constant.

$$\frac{1}{\lambda} = 1.09678 \times 10^7 \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \text{m}^{-1}$$

For hydrogen's Balmer series, electron orbit transitions occur from larger orbits to the  $n = 2$  orbit; that is,  $n_f = 2$ .



### Think Critically

- Calculate** the wavelengths for the following electron orbit transitions.
  - $n_i = 3$ ;  $n_f = 2$
  - $n_i = 4$ ;  $n_f = 2$
  - $n_i = 5$ ;  $n_f = 2$
  - $n_i = 6$ ;  $n_f = 2$
- Relate** the Balmer-series wavelengths you calculated in Question 1 to those determined experimentally. Allowing for experimental error and calculation uncertainty, do the wavelengths match? Explain your answer. One angstrom (Å) equals  $10^{-10}$  m.
- Apply** the formula  $E = hc/\lambda$  to determine the energy per quantum for each of the orbit transitions in Question 1.
- Extend** the Bohr model by calculating the wavelength and energy per quantum for the electron orbit transition for which  $n_f = 3$  and  $n_i = 5$ . This transition accounts for a spectral line in hydrogen's Paschen series.

**The Heisenberg uncertainty principle** 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) proved to have profound implications for atomic models.

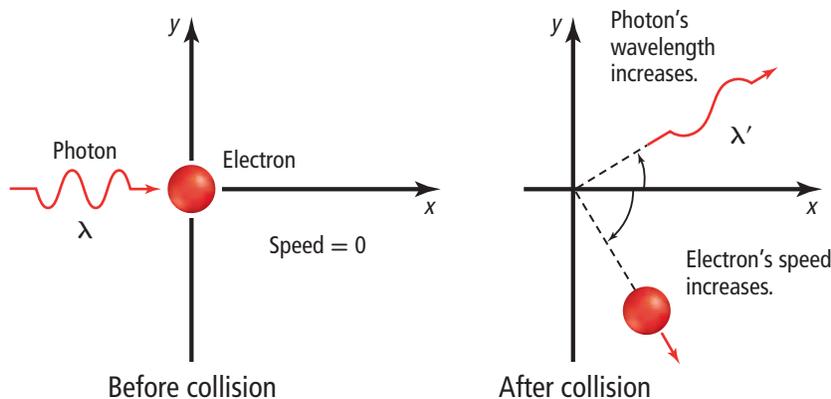
Heisenberg showed that it is impossible to take any measurement of an object without disturbing the object. Imagine trying to locate a hovering, helium-filled balloon in a darkened room. If you wave your hand about, you can locate the balloon's position when you touch it. However, when you touch the balloon, you transfer energy to it and change its position. You could also detect the balloon's position by turning on a flashlight. Using this method, photons of light reflected from the balloon would reach your eyes and reveal the balloon's location. Because the balloon is a macroscopic object, the effect of the rebounding photons on its position is very small and not observable.

Imagine trying to determine an electron's location by “bumping” it with a high-energy photon. 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 5.14**. 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.

 **Reading Check** Explain the Heisenberg uncertainty principle.

Although scientists of the time found Heisenberg's principle difficult to accept, it has been proven to describe the fundamental limitations of what can be observed. The interaction of a photon with a macroscopic 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 is not the case with an electron moving at  $6 \times 10^6$  m/s near an atomic nucleus. The uncertainty of the electron's position is at least  $10^{-9}$  m, about 10 times greater than the diameter of the entire atom.

The Heisenberg uncertainty principle also means that it is impossible to assign fixed paths for electrons like the circular orbits in Bohr's model. The only quantity that can be known is the probability for an electron to occupy a certain region around the nucleus.



■ **Figure 5.14** When a photon interacts with an electron at rest, both the velocity and the position of the electron are modified. This illustrates the Heisenberg uncertainty principle. It is impossible to know at the same time the position and the velocity of a particle.

**Explain** Why has the photon's energy changed?

**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. 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, the **quantum mechanical model of the atom**. Like Bohr's model, 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.



**Reading Check Compare and contrast** Bohr's model and the quantum mechanical model.

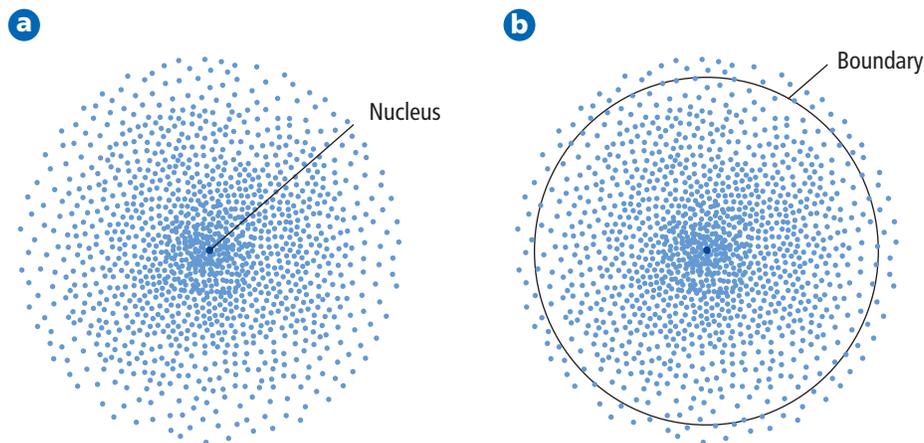
The Schrödinger wave equation is too complex to be considered here. However, each solution to the equation is known as a wave function, which is related to the probability of finding the electron within a particular volume of space around the nucleus. Recall from your study of mathematics that an event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

**Electron's probable location** The wave function predicts a three-dimensional region around the nucleus, called an **atomic orbital**, which describes the electron's probable location. An atomic orbital is like a fuzzy cloud in which the density at a given point is proportional to the probability of finding the electron at that point. **Figure 5.15a** illustrates the probability map that describes the electron in the atom's lowest energy state. The probability map can be thought of 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. The high density of dots near the nucleus indicates the electron's most probable location. However, because the cloud has no definite boundary, it is also possible that the electron might be found at a considerable distance from the nucleus.



**Reading Check Describe** where electrons are located in an atom.

■ **Figure 5.15** The density map represents the probability of finding an electron at a given position around the nucleus. **a.** The higher density of points near the nucleus shows that the electron is more likely to be found close to the nucleus. **b.** At any given time, there is a 90% probability of finding the electron within the circular region shown. This surface is sometimes chosen to represent the boundary of the atom. In this illustration, the circle corresponds to a projection of the 3-dimensional sphere that contains the electrons.



**Density Maps**

## Hydrogen's Atomic Orbitals

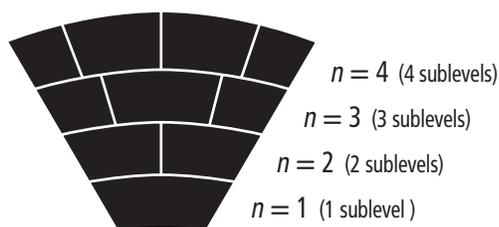
Because the boundary of an atomic orbital is fuzzy, the orbital does not have an exact defined size. To overcome the inherent uncertainty about the electron's location, chemists arbitrarily draw an orbital's surface to contain 90% of the electron's total probability distribution. This means that the probability of finding the electron within the boundary is 0.9 and the probability of finding it outside the boundary is 0.1. In other words, it is more likely to find the electron close to the nucleus and within the volume defined by the boundary, than to find it outside the volume. The circle shown in **Figure 5.15b** encloses 90% of the lowest-energy orbital of hydrogen.

**Principal quantum number** Recall that the Bohr atomic model assigns quantum numbers to electron orbits. Similarly, the quantum mechanical model assigns four quantum numbers to atomic orbitals. The first one is the **principal quantum number** ( $n$ ) and indicates the relative size and energy of atomic orbitals. As  $n$  increases, the orbital becomes larger, the electron spends more time farther from the nucleus, and the atom's energy increases. Therefore,  $n$  specifies the atom's major energy levels. Each major energy level is called a **principal energy level**. An atom's lowest principal energy level is assigned a principal quantum number of 1. When the hydrogen atom's single electron occupies an orbital with  $n = 1$ , the atom is in its ground state. Up to 7 energy levels have been detected for the hydrogen atom, giving  $n$  values ranging from 1 to 7.

**Energy sublevels** 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 5.16**. 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.

 **Reading Check** Explain the relationship between energy levels and sublevels.

■ **Figure 5.16** Energy levels can be thought of as rows of seats in a theater. The rows that are higher up and farther from the stage contain more seats. Similarly, energy levels related to orbitals farther from the nucleus contain more sublevels.

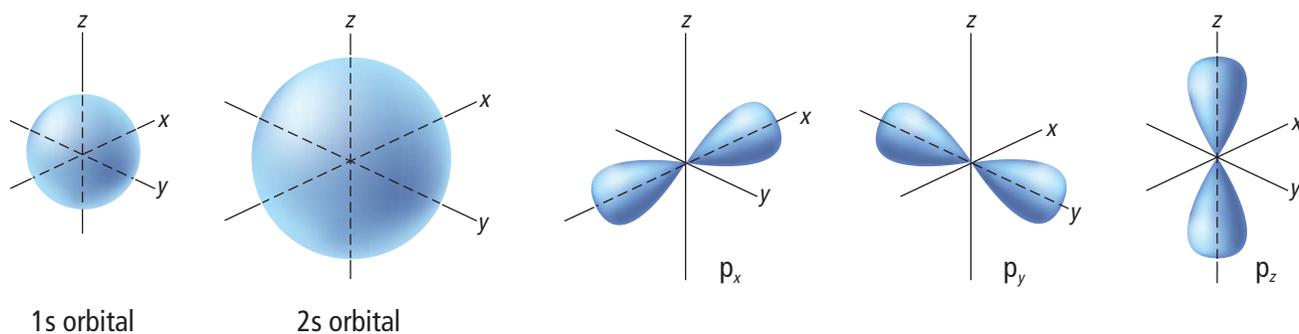


**Shapes of orbitals** Sublevels are labeled  $s$ ,  $p$ ,  $d$ , or  $f$  according to the shapes of the atom's orbitals. All  $s$  orbitals are spherical, and all  $p$  orbitals are dumbbell-shaped; however, not all  $d$  or  $f$  orbitals have the same shape. Each orbital can contain, at most, two electrons. The single sublevel in principal energy level 1 corresponds to a spherical orbital called the  $1s$  orbital. The two sublevels in principal energy level 2 are designated  $2s$  and  $2p$ . The  $2s$  sublevel corresponds to the  $2s$  orbital, which is spherical like the  $1s$  orbital but larger in size, as shown in **Figure 5.17a**. The  $2p$  sublevel corresponds to three dumbbell-shaped  $p$  orbitals 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 5.17b**. Each of the  $p$  orbitals related to an energy sublevel has the same energy.

 **Reading Check** Describe the shapes of  $s$  and  $p$  orbitals.

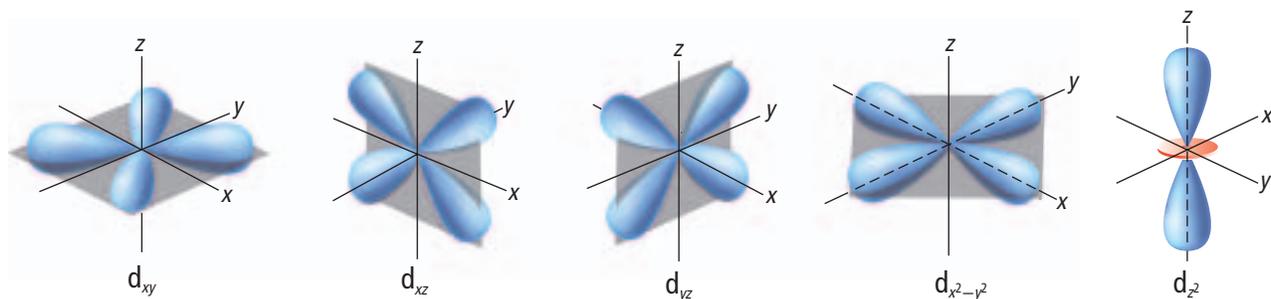
Principal energy level 3 consists of three sublevels designated  $3s$ ,  $3p$ , and  $3d$ . Each  $d$  sublevel relates to five orbitals of equal energy. Four of the  $d$  orbitals have identical shapes but different orientations along the  $x$ ,  $y$ , and  $z$  coordinate axes. However, the fifth orbital,  $d_{z^2}$ , is shaped and oriented differently than the other four. The shapes and orientations of the five  $d$  orbitals are illustrated in **Figure 5.17c**. The fourth principal energy level ( $n = 4$ ) contains a fourth sublevel, called the  $4f$  sublevel, which relates to seven  $f$  orbitals of equal energy. The  $f$  orbitals have complex, multilobed shapes.

■ **Figure 5.17** Each energy sublevel relates to orbitals of different shape.



**a.** All  $s$  orbitals are spherical, and their size increases with increasing principal quantum number.

**b.** The three  $p$  orbitals are dumbshell-shaped and are oriented along the three perpendicular  $x$ ,  $y$ , and  $z$  axes.



**c.** Four of the five  $d$  orbitals have the same shape but lie in different planes. 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	1	4
	p	3	
3	s	1	9
	p	3	
	d	5	
4	s	1	16
	p	3	
	d	5	
	f	7	

Hydrogen's first four principal energy levels, sublevels, and related atomic orbitals are summarized in **Table 5.2**. Note that the number of orbitals related to each sublevel is always an odd number, and that the maximum number of orbitals related to each principal energy level equals  $n^2$ .

At any given time, the electron in a hydrogen atom can occupy just one orbital. 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, when the atom gains a quantum of energy, the electron is excited to one of the unoccupied orbitals. Depending on the amount of energy available, the electron can move to the 2s orbital, to one of the three 2p orbitals, or to any other orbital that is vacant.

## Section 5.2 Assessment

### Section Summary

- ▶ Bohr's atomic model attributes hydrogen's emission spectrum to electrons dropping from higher-energy to lower-energy orbits.
- ▶ The de Broglie equation relates a particle's wavelength to its mass, its velocity, and Planck's constant.
- ▶ The quantum mechanical model assumes that electrons have wave properties.
- ▶ Electrons occupy three-dimensional regions of space called atomic orbitals.

15. **MAIN Idea** Explain the reason, according to Bohr's atomic model, why atomic emission spectra contain only certain frequencies of light.
16. **Differentiate** between the wavelength of visible light and the wavelength of a moving soccer ball.
17. **Enumerate** the sublevels contained in the hydrogen atom's first four energy levels. What orbitals are related to each s sublevel and each p sublevel?
18. **Explain** why the location of an electron in an atom is uncertain using the Heisenberg uncertainty principle and de Broglie's wave-particle duality. How is the location of electrons in atoms defined?
19. **Calculate** Use the information in **Table 5.1** to calculate how many times larger the hydrogen atom's seventh Bohr radius is than its first Bohr radius.
20. **Compare and contrast** Bohr's model and the quantum mechanical model of the atom.



**Table 5.3****Features of the Aufbau Diagram**

Feature	Example
All orbitals related to an energy sublevel are of equal energy.	All three 2p orbitals are of equal energy.
In a multi-electron atom, the energy sublevels within a principal energy level have different energies.	The three 2p orbitals are of higher energy than the 2s orbital.
In order of increasing energy, the sequence of energy sublevels within a principal energy level is s, p, d, and f.	If $n = 4$ , then the sequence of energy sublevels is 4s, 4p, 4d, and 4f.
Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level.	The orbital related to the atom's 4s sublevel has a lower energy than the five orbitals related to the 3d sublevel.

**Table 5.3** summarizes several features of the aufbau diagram.

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

**The Pauli exclusion principle** Electrons in orbitals can be represented by arrows in boxes. Each electron has an associated spin, similar to the way a top spins on its point. 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, and an arrow pointing down  $\downarrow$  represents the electron spinning in the opposite direction. 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.

The **Pauli exclusion principle** states that a maximum of two electrons can occupy a single atomic orbital, but only if the electrons have opposite spins. Austrian physicist Wolfgang Pauli (1900–1958) proposed this principle after observing atoms in excited states. An atomic orbital containing paired electrons with opposite spins is written as  $\uparrow\downarrow$ . Because each orbital can contain, at most, two electrons, the maximum number of electrons related to each principal energy level equals  $2n^2$ .

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

- $\uparrow \square \square$
- $\uparrow \uparrow \square$
- $\uparrow \uparrow \uparrow$
- $\uparrow\downarrow \uparrow \uparrow$
- $\uparrow\downarrow \uparrow\downarrow \uparrow$
- $\uparrow\downarrow \uparrow\downarrow \uparrow\downarrow$

 **Reading Check State** the three rules that define how electrons are arranged in atoms.

**VOCABULARY****WORD ORIGIN****Aufbau**

comes from the German word *aufbauen*, which means to *configure* or *arrange*.

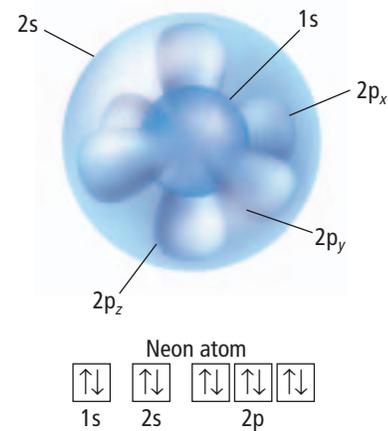
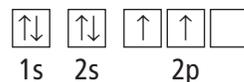
**FOLDABLES**

Incorporate information from this section into your Foldable.

## Electron Arrangement

You can represent an atom's electron configuration using one of two convenient methods: orbital diagrams or electron configuration notation.

**Orbital diagrams** As mentioned earlier, electrons in orbitals can be represented by arrows in boxes. Each box is labeled with the principal quantum number and sublevel associated with the orbital. For example, the orbital diagram for a ground-state carbon atom, which contains two electrons in the 1s orbital, two electrons in the 2s orbital, and one electron in two of three separate 2p orbitals, is shown below.



■ **Figure 5.19** The 1s, 2s, and 2p orbitals of a neon atom overlap. **Determine** how many electrons a neon atom has.

**Electron configuration notation** The electron configuration notation 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  $1s^2 2s^2 2p^2$ . Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in **Table 5.4**. **Figure 5.19** illustrates how the 1s, 2s,  $2p_x$ ,  $2p_y$ , and  $2p_z$  orbitals of the neon atom, previously illustrated in **Figure 5.17**, overlap.

Concepts in Motion

**Interactive Table** Explore electron configurations and orbital diagrams at [glencoe.com](http://glencoe.com).



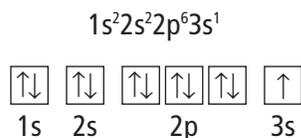
**Table 5.4**

### Electron Configurations and Orbital Diagrams for Elements 1–10

Element	Atomic Number	Orbital Diagram					Electron Configuration Notation
		1s	2s	$2p_x$	$2p_y$	$2p_z$	
Hydrogen	1	↑					$1s^1$
Helium	2	↑↓					$1s^2$
Lithium	3	↑↓	↑				$1s^2 2s^1$
Beryllium	4	↑↓	↑↓				$1s^2 2s^2$
Boron	5	↑↓	↑↓	↑			$1s^2 2s^2 2p^1$
Carbon	6	↑↓	↑↓	↑	↑		$1s^2 2s^2 2p^2$
Nitrogen	7	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
Oxygen	8	↑↓	↑↓	↑↓	↑	↑	$1s^2 2s^2 2p^4$
Fluorine	9	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^5$
Neon	10	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$

Note that the electron configuration notation does not usually show the orbital distributions of electrons related to a sublevel. It is 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 as follows.



**Noble-gas notation** is a method of representing electron configurations of noble gases. Noble gases are the elements in the last column of the periodic table. They have eight electrons in their outermost orbital and they are unusually stable. You will learn more about noble gases in Chapter 6. The noble-gas notation uses 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  $[\text{Ne}]3s^1$ . 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 additional orbitals being filled. The complete and abbreviated (using noble-gas notation) electron configurations of the period 3 elements are shown in **Table 5.5**.

**Reading Check Explain** how to write the noble-gas notation for an element. What is the noble-gas notation for calcium?

Table 5.5		Electron Configurations for Elements 11–18	
Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[\text{Ne}]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[\text{Ne}]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[\text{Ne}]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[\text{Ne}]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[\text{Ne}]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[\text{Ne}]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[\text{Ne}]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[\text{Ne}]3s^2 3p^6$ or [Ar]



**Personal Tutor** For an online tutorial on electron configuration notation, visit [glencoe.com](http://glencoe.com).

## VOCABULARY

### SCIENCE USAGE V. COMMON USAGE

#### Period

**Science usage:** a horizontal row of elements in the current periodic table

*There are seven periods in the current periodic table.*

**Common usage:** an interval of time determined by some recurring phenomenon

*The period of Earth's orbit is one year.*

### Concepts in Motion

**Interactive Figure** To see an animation of electron configurations, visit [glencoe.com](http://glencoe.com).

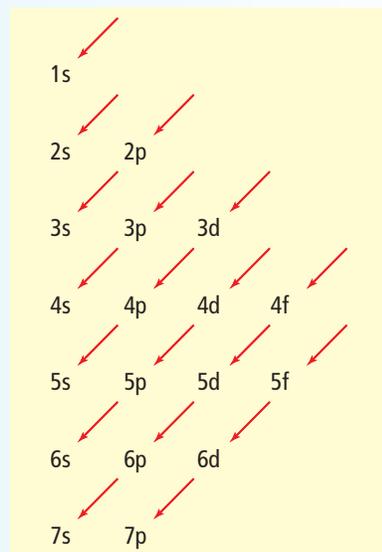
**Exceptions to predicted configurations** You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium,  $[\text{Ar}]4s^23d^4$ , and copper,  $[\text{Ar}]4s^23d^9$ , would be incorrect. The correct configurations for these two elements are  $[\text{Ar}]4s^13d^5$  for chromium and  $[\text{Ar}]4s^13d^{10}$  for copper. The electron configurations for these two elements, as well as those of several other elements, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

## Problem-Solving Strategy

### Filling Atomic Orbitals

By drawing a sublevel diagram and following the arrows, you can write the ground-state electron configuration for any chemical element.

1. Sketch the sublevel diagram on a blank piece of paper.
2. Determine the number of electrons in one atom of the element for which you are writing the electron configuration. The number of electrons in a neutral atom equals the element's atomic number.
3. Starting with 1s, write the aufbau sequence of atomic orbitals by following the diagonal arrows from the top of the sublevel diagram to the bottom. When you complete one line of arrows, move to the right, to the beginning of the next line of arrows. As you proceed, add superscripts indicating the numbers of electrons in each set of atomic orbitals. Continue only until you have sufficient atomic orbitals to accommodate the total number of electrons in one atom of the element.
4. Apply noble-gas notation.



The sublevel diagram shows the order in which the orbitals are usually filled.

### Apply the Strategy

**Write** the ground-state electron configuration for zirconium.

## PRACTICE Problems

Extra Practice Page 978 and [glencoe.com](http://glencoe.com)

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

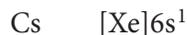
a. bromine (Br)	c. antimony (Sb)	e. terbium (Tb)
b. strontium (Sr)	d. rhenium (Re)	f. titanium (Ti)
22. A chlorine atom in its ground state has a total of seven electrons in orbitals related to the atom's third energy level. How many of the seven electrons occupy p orbitals? How many of the 17 electrons in a chlorine atom occupy p orbitals?
23. When a sulfur atom reacts with other atoms, electrons in orbitals related to the atom's third energy level are involved. How many such electrons does a sulfur atom have?
24. An element has the ground-state electron configuration  $[\text{Kr}]5s^24d^{10}5p^1$ . It is part of some semiconductors and used in various alloys. What element is it?
25. **Challenge** In its ground state, an atom of an element has two electrons in all orbitals related to the atom's highest energy level for which  $n = 6$ . Using noble-gas notation, write the electron configuration for this element, and identify the element.

## 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 just one valence electron, the 6s electron shown in cesium's electron configuration.



**Electron-dot structures** Because valence electrons are involved in forming chemical bonds, chemists often represent them visually using a simple shorthand method, called electron-dot structure. 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 all of the atom's valence electrons. 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 5.6**.

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	$1s^22s^22p^4$	: O ·
Fluorine	9	$1s^22s^22p^5$	: F ·
Neon	10	$1s^22s^22p^6$	: Ne :

Concepts in Motion

**Interactive Table** Explore electron-dot structures at [glencoe.com](http://glencoe.com).



## EXAMPLE Problem 5.3

**Electron-Dot Structures** Some toothpastes contain stannous fluoride, a compound of tin and fluorine. What is tin's electron-dot structure?

### 1 Analyze the Problem

Consult the periodic table to determine the total number of electrons in a tin atom. Write out tin's electron configuration, and determine its number of valence electrons. Then use the rules for electron-dot structures to draw the electron-dot structure for tin.

### 2 Solve for the Unknown

Tin has an atomic number of 50. Thus, a tin atom has 50 electrons.



Write out tin's electron configuration using noble-gas notation. The closest noble gas is Kr.

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 the four valence electrons around tin's chemical symbol (Sn) to show tin's electron-dot structure.  $\cdot \ddot{\text{S}}\text{n} \cdot$

### 3 Evaluate the Answer

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

## PRACTICE Problems

Extra Practice Page 978 and [glencoe.com](http://glencoe.com)

26. Draw electron-dot structures for atoms of the following elements.
- a. magnesium                      b. thallium                      c. xenon
27. An atom of an element has a total of 13 electrons. What is the element, and how many electrons are shown in its electron-dot structure?
28. **Challenge** An element exists in the gaseous state at room temperature and normal atmospheric pressure and is known to be one of the following: hydrogen, helium, nitrogen, oxygen, fluorine, chlorine, or neon. Identify the element based on the electron-dot structure at right.



## Section 5.3 Assessment

### Section Summary

- ▶ The arrangement of electrons in an atom is called the atom's electron configuration.
- ▶ Electron configurations are defined by the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- ▶ An element's valence electrons determine the chemical properties of the element.
- ▶ Electron configurations can be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

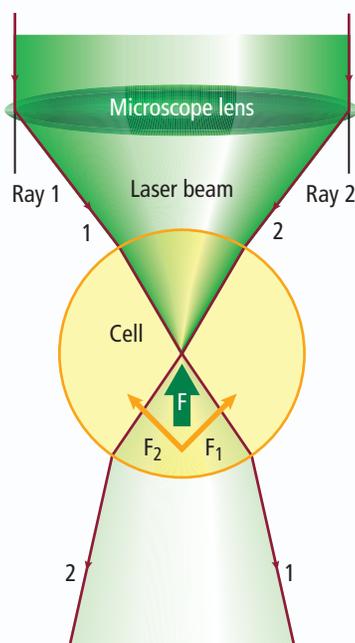
29. **MAIN Idea Apply** the Pauli exclusion principle, the aufbau principle, and Hund's rule to write out the electron configuration and draw the orbital diagram for each of the following elements.
- a. silicon                      b. fluorine                      c. calcium                      d. krypton
30. **Define** *valence electron*.
31. **Illustrate** and describe the sequence in which ten electrons occupy the five orbitals related to an atom's d sublevel.
32. **Extend** the aufbau sequence through an element that has not yet been identified, but whose atoms would completely fill 7p orbitals. How many electrons such an atom would have? Write its electron configuration using noble-gas notation for the previous noble gas, radon.
33. **Interpret Scientific Illustrations** Which is the correct electron-dot structure for an atom of selenium? Explain.
- a.  $\cdot \ddot{\text{S}}\text{e} :$                       b.  $\cdot \ddot{\text{S}}\text{e} \cdot$                       c.  $\cdot \ddot{\text{S}}\text{e} \cdot$                       d.  $\cdot \ddot{\text{S}} \cdot$

## Tiny Tweezers

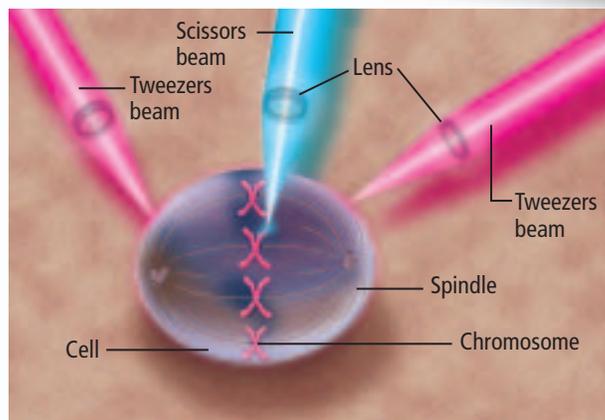
Peering through a microscope, a cell biologist can grasp a single cell with a pair of “tweezers.” But these are not the kind of tweezers you might find in a medicine cabinet. These tweezers are made from two laser beams and can hold very tiny things such as cells and even individual atoms.

You might have heard that lasers can be used to cut things. Laser “scissors” are used in some surgeries. But surprisingly, lasers can also trap living cells and other microscopic objects in their beams without damaging them. How can beams of light hold things in place?

**Gripping with light** When light rays pass through a cell, they change direction slightly. This is similar to how light rays bend when passing through water in an aquarium. When light rays are bent, they exert a force. Large objects, such as aquariums, are too massive to be affected by this minuscule force, but tiny cells respond to the force. If the light rays are positioned in just the right way, they can hold a small object in place, as shown in **Figure 1**.



**Figure 1** As the laser passes through the cell, the beam is bent. The cell feels a force in the opposite direction of bending and so is held within the beam of light.



**Figure 2** Organelles found within living cells are accessible to the smallest lasers.

**Lasers and cancer** So what use do scientists have for these tiny tweezers? One group of scientists is using them to study cell organelles. They are studying the forces exerted by mitotic spindles—the grouping of microtubules that coordinates cell division. The spindles guide replicated chromosomes to opposite sides of the cell—a key role in cell division. However, scientists do not know exactly how the spindles perform this function.

Tiny laser scissors have been used to cut off pieces of chromosomes during cell division. Laser tweezers were then used to move the pieces around the cell and the spindles, as shown in **Figure 2**. Knowing the force with which the tweezers grasp the chromosomes, scientists can measure the opposing force exerted by the spindles. Scientists hope that learning how spindles function during cell division will help them learn more about diseases related to cell division, such as cancer—a disease in which cells divide uncontrollably.

### WRITING in Chemistry

**Laser Light** Lasers can be found in a wide variety of everyday settings. Research the different types of lasers you might encounter daily, and find out what kind of light each laser uses. Summarize the results of your research in a journal. For more information about how laser light is made, visit [glencoe.com](http://glencoe.com).

# CHEMLAB

## ANALYZE LINE SPECTRA

**Background:** Emission spectra are produced when excited atoms return to a more stable state by emitting radiation of specific wavelengths. When white light passes through a sample, atoms in the sample absorb specific wavelengths. This produces dark lines in the continuous spectrum of white light and is called an absorption spectrum.

**Question:** *What absorption and emission spectra do various substances produce?*

### Materials

ring stand with clamp  
40-W tubular lightbulb  
light socket with grounded power cord  
275-mL polystyrene culture flask  
Flinn C-Spectra® or similar diffraction grating  
red, green, blue, and yellow food coloring  
set of colored pencils  
spectrum tubes (hydrogen, neon, and sodium)  
spectrum-tube power supply (3)

### Safety Precautions

**WARNING:** *Use care around the spectrum-tube power supplies. Spectrum tubes will become hot when used.*

### Procedure

1. Read and complete the lab safety form.
2. Use a Flinn C-Spectra® or similar diffraction grating to view an incandescent lightbulb. Draw the observed spectrum using colored pencils.
3. Use the Flinn C-Spectra® to view the emission spectra from tubes of gaseous hydrogen, neon, and sodium. Use colored pencils to draw the observed spectra.
4. Fill a 275-mL culture flask with about 100 mL of water. Add two or three drops of red food coloring to the water. Shake the solution.
5. Repeat Step 4 for the green, blue, and yellow food coloring.
6. Set up the 40-W lightbulb so that it is near eye level. Place the flask with red food coloring about 8 cm from the lightbulb so that you are able to see light from the bulb above the solution and light from the bulb projecting through the solution.



7. With the room lights darkened, view the light using the Flinn C-Spectra®. The top spectrum viewed will be a continuous spectrum from the white lightbulb. The bottom spectrum will be the absorption spectrum of the red solution. Use colored pencils to make a drawing of the absorption spectra you observe.
8. Repeat Steps 6 and 7 using the green, blue, and yellow solutions.
9. **Cleanup and Disposal** Turn off the light and spectrum-tube power supplies. Wait several minutes for the lightbulb and spectrum tubes to cool. Dispose of the liquids and store the lightbulb and spectrum tubes as directed by your teacher.

### Analyze and Conclude

1. **Think Critically** How can the single electron in a hydrogen atom produce all of the lines found in its emission spectrum?
2. **Predict** How can you predict the absorption spectrum of a solution by looking at its color?
3. **Apply** How can spectra be used to identify the presence of specific elements in a substance?
4. **Error Analysis** Name a potential source of error in this experiment. Choose one of the elements you observed, and research its absorption spectrum. Compare your findings with the results of your experiment.

### INQUIRY EXTENSION

**Hypothesize** What would happen if you mixed more than one color of food coloring with water and repeated the experiment? Design an experiment to test your hypothesis.

**BIG Idea** The atoms of each element have a unique arrangement of electrons.

### Section 5.1 Light and Quantized Energy

**MAIN Idea** Light, a form of electromagnetic radiation, has characteristics of both a wave and a particle.

#### Vocabulary

- amplitude (p. 137)
- atomic emission spectrum (p. 144)
- electromagnetic radiation (p. 137)
- electromagnetic spectrum (p. 139)
- frequency (p. 137)
- photoelectric effect (p. 142)
- photon (p. 143)
- Planck's constant (p. 142)
- quantum (p. 141)
- wavelength (p. 137)

#### Key Concepts

- All waves are defined by their wavelengths, frequencies, amplitudes, and speeds.

$$c = \lambda\nu$$

- In a vacuum, all electromagnetic waves travel at the speed of light.
- All electromagnetic waves have both wave and particle properties.
- Matter emits and absorbs energy in quanta.

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

- White light produces a continuous spectrum. An element's emission spectrum consists of a series of lines of individual colors.

### Section 5.2 Quantum Theory and the Atom

**MAIN Idea** Wavelike properties of electrons help relate atomic emission spectra, energy states of atoms, and atomic orbitals.

#### Vocabulary

- atomic orbital (p. 152)
- de Broglie equation (p. 150)
- energy sublevel (p. 153)
- ground state (p. 146)
- Heisenberg uncertainty principle (p. 151)
- principal energy level (p. 153)
- principal quantum number (p. 153)
- quantum mechanical model of the atom (p. 152)
- quantum number (p. 147)

#### Key Concepts

- Bohr's atomic model attributes hydrogen's emission spectrum to electrons dropping from higher-energy to lower-energy orbits.

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

- The de Broglie equation relates a particle's wavelength to its mass, its velocity, and Planck's constant.

$$\lambda = h / m\nu$$

- The quantum mechanical model of the atom assumes that electrons have wave properties.
- Electrons occupy three-dimensional regions of space called atomic orbitals.

### Section 5.3 Electron Configuration

**MAIN Idea** A set of three rules can be used to determine electron arrangement in an atom.

#### Vocabulary

- aufbau principle (p. 156)
- electron configuration (p. 156)
- electron-dot structure (p. 161)
- Hund's rule (p. 157)
- Pauli exclusion principle (p. 157)
- valence electron (p. 161)

#### Key Concepts

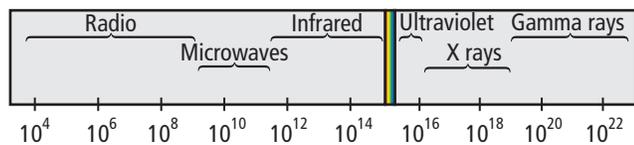
- The arrangement of electrons in an atom is called the atom's electron configuration.
- Electron configurations are defined by the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- An element's valence electrons determine the chemical properties of the element.
- Electron configurations can be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

## Section 5.1

## Mastering Concepts

34. Define the following terms.
  - a. frequency
  - b. wavelength
  - c. quantum
  - d. ground state
35. Arrange the following types of electromagnetic radiation in order of increasing wavelength.
  - a. ultraviolet light
  - b. microwaves
  - c. radio waves
  - d. X rays
36. A gamma ray has a frequency of  $2.88 \times 10^{21}$  Hz. What does this mean?
37. What is the photoelectric effect?
38. **Neon Sign** How does light emitted from a neon sign differ from sunlight?
39. Explain Planck's quantum concept as it relates to energy lost or gained by matter.
40. How did Einstein explain the photoelectric effect?
41. **Rainbow** What are two differences between the red and green electromagnetic waves in a rainbow?
42. **Temperature** What happens to the light emitted by a heated, glowing object as its temperature increases?
43. What are three deficiencies of the wave model of light related to light's interaction with matter?
44. How are radio waves and ultraviolet waves similar? How are they different?

## Mastering Problems

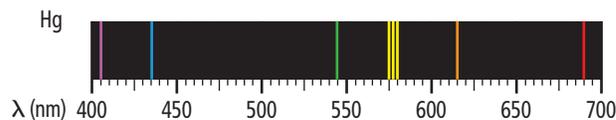
Frequency ( $\nu$ ) in hertz

Electromagnetic Spectrum

■ Figure 5.20

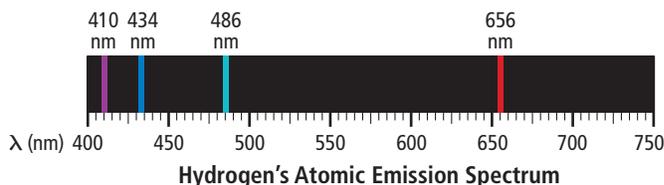
45. **Radiation** Use Figure 5.20 to determine the following types of radiation.
  - a. radiation with a frequency of  $8.6 \times 10^{11} \text{ s}^{-1}$
  - b. radiation with a wavelength of 4.2 nm
  - c. radiation with a frequency of 5.6 MHz
  - d. radiation that travels at a speed of  $3.00 \times 10^8 \text{ m/s}$
46. What is the wavelength of electromagnetic radiation with a frequency of  $5.00 \times 10^{12} \text{ Hz}$ ? What kind of electromagnetic radiation is this?
47. What is the frequency of electromagnetic radiation with a wavelength of  $3.33 \times 10^{-8} \text{ m}$ ? What type of electromagnetic radiation is this?

48. What is the speed of an electromagnetic wave with a frequency of  $1.33 \times 10^{17} \text{ Hz}$  and a wavelength of 2.25 nm?
49. What is the energy of a photon of red light that has a frequency of  $4.48 \times 10^{14} \text{ Hz}$ ?



■ Figure 5.21

50. **Mercury** Mercury's atomic emission spectrum is shown in Figure 5.21. Estimate the wavelength of the orange line. What is its frequency? What is the energy of a photon corresponding to the orange line emitted by the mercury atom?
51. What is the energy of an ultraviolet photon that has a wavelength of  $1.18 \times 10^{-8} \text{ m}$ ?
52. A photon has an energy of  $2.93 \times 10^{-25} \text{ J}$ . What is its frequency? What type of electromagnetic radiation is the photon?
53. A photon has an energy of  $1.10 \times 10^{-13} \text{ J}$ . What is the photon's wavelength? What type of electromagnetic radiation is it?
54. **Spacecraft** How long does it take a radio signal from the *Voyager* spacecraft to reach Earth if the distance between *Voyager* and Earth is  $2.72 \times 10^9 \text{ km}$ ?
55. **Radio Waves** 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?
56. **Platinum** What minimum frequency of light is needed to eject a photoelectron from atoms of platinum, which require at least  $9.08 \times 10^{-19} \text{ J/photon}$ ?
57. **Eye Surgery** The argon fluoride (ArF) laser used in some refractive eye surgeries emits electromagnetic radiation of 193.3 nm wavelength. What is the frequency of the ArF laser's radiation? What is the energy of a single quantum of the radiation?



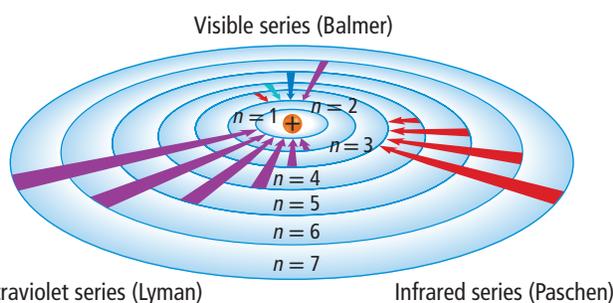
■ Figure 5.22

58. **Hydrogen** One line in hydrogen's emission spectrum has a wavelength of 486 nm. Examine Figure 5.22 to determine the line's color. What is the line's frequency?

## Section 5.2

## Mastering Concepts

- According to the Bohr model, how do electrons move in atoms?
- What does  $n$  designate in Bohr's atomic model?
- What is the difference between an atom's ground state and an excited state?
- 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?
- What is an atomic orbital?
- What does  $n$  represent in the quantum mechanical model of the atom?



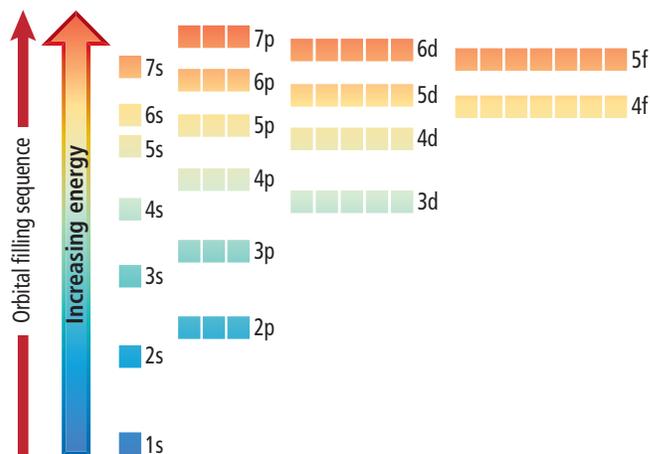
■ Figure 5.23

- Electron Transition** According to the Bohr model shown in Figure 5.23, what type of electron-orbit transitions produce the ultraviolet lines in hydrogen's Lyman series?
- How many energy sublevels are contained in each of the hydrogen atom's first three energy levels?
- What atomic orbitals are related to a d sublevel?
- What do the sublevel designations s, p, d, and f specify with respect to the atom's orbitals?
- How are the five orbitals related to an atom's d sublevel designated?
- What is the maximum number of electrons an orbital can contain?
- Describe the relative orientations of the orbitals related to an atom's 2p sublevel.
- How many electrons can be contained in all the orbitals related to an argon atom's third energy level?
- How does the quantum mechanical model of the atom describe the paths of an atom's electrons?
- Macroscopic Objects** Why do we not notice the wavelengths of moving objects such as automobiles?
- Why is it impossible to know precisely the velocity and position of an electron at the same time?

## Section 5.3

## Mastering Concepts

- In what sequence do electrons fill the atomic orbitals related to a sublevel?



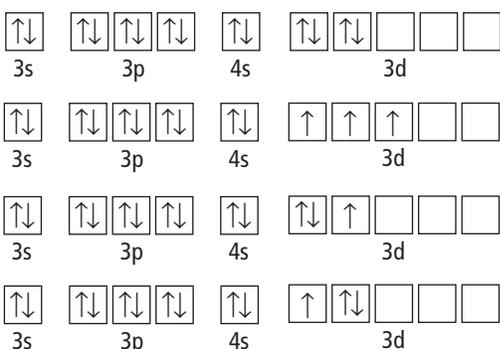
■ Figure 5.24

- Rubidium** Using Figure 5.24, explain why one electron in a rubidium atom occupies a 5s orbital rather than a 4d or 4f orbital.
- What are valence electrons? How many of a magnesium atom's 12 electrons are valence electrons?
- Light is said to have a dual wave-particle nature. What does this statement mean?
- Describe the difference between a quantum and a photon.
- How many electrons are shown in each element's electron-dot structure?
  - carbon
  - iodine
  - calcium
  - gallium
- When writing the electron configuration notation for an atom, what three principles or rules should you follow?
- Write the electron configuration and draw the orbital notation for atoms of oxygen and sulfur.

## Mastering Problems

- List the aufbau sequence of orbitals from 1s to 7p.
- Write each element's orbital notation and complete electron configuration.
  - beryllium
  - aluminum
  - nitrogen
  - sodium
- Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.
  - Kr
  - P
  - Zr
  - Pb

87. What element is represented by each electron configuration?
- $1s^2 2s^2 2p^5$
  - $[\text{Ar}] 4s^2$
  - $[\text{Xe}] 6s^2 4f^4$
  - $[\text{Kr}] 5s^2 4d^{10} 5p^4$
  - $[\text{Rn}] 7s^2 5f^{13}$
  - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
88. Which electron configuration notation describes an atom in an excited state?
- $[\text{Ar}] 4s^2 3d^{10} 4p^2$
  - $[\text{Ne}] 3s^2 3p^5$
  - $[\text{Kr}] 5s^2 4d^1$
  - $[\text{Ar}] 4s^2 3d^8 4p^1$



■ Figure 5.25

89. Which orbital diagram in **Figure 5.25** is incorrect for an atom in its ground state?
90. Draw an electron-dot structure for an atom of each element.
- carbon
  - arsenic
  - polonium
  - potassium
  - barium
91. **Arsenic** 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?

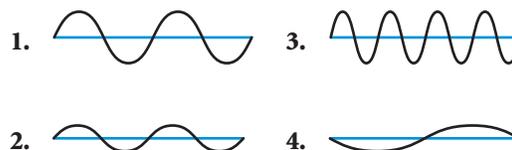


■ Figure 5.26

92. Which element could have the ground-state electron-dot notation shown in **Figure 5.26**?
- manganese
  - antimony
  - calcium
  - samarium
93. For an atom of tin in the ground state, write the electron configuration using noble-gas notation, and draw its electron-dot structure.

## Mixed Review

94. What is the maximum number of electrons that can be contained in an atom's orbitals having the following principal quantum numbers?
- 3
  - 4
  - 6
  - 7
95. What is the wavelength of light with a frequency of  $5.77 \times 10^{14}$  Hz?

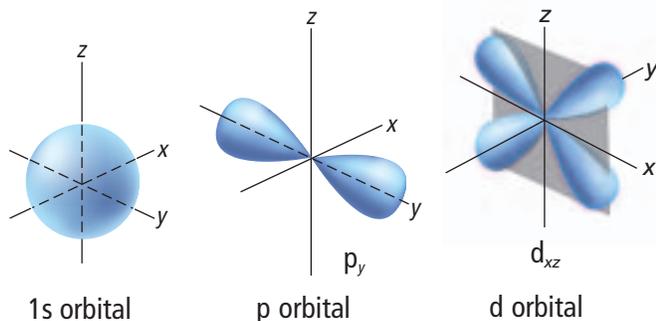


■ Figure 5.27

96. **Waves** Using the waves shown in **Figure 5.27**, identify the wave or waves with the following characteristics.
- longest wavelength
  - greatest frequency
  - largest amplitude
  - shortest wavelength
97. How many orientations are possible for the orbitals related to each sublevels?
- s
  - p
  - d
  - f
98. Which elements have only two electrons in their electron-dot structures: hydrogen, helium, lithium, aluminum, calcium, cobalt, bromine, krypton, or barium?
99. In Bohr's atomic model, what electron-orbit transition produces the blue-green line in hydrogen's atomic emission spectrum?
100. **Zinc** 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?
101. **X Ray** An X-ray photon has an energy of  $3.01 \times 10^{-18}$  J. What is its frequency and wavelength?
102. Which element has the ground-state electron configuration represented by the noble-gas notation  $[\text{Rn}] 7s^1$ ?
103. How did Bohr explain atomic emission spectra?
104. **Infrared Radiation** How many photons of infrared radiation with a frequency of  $4.88 \times 10^{13}$  Hz are required to provide an energy of 1.00 J?
105. Light travels slower in water than it does in air; however, its frequency remains the same. How does the wavelength of light change as it travels from air to water?
106. According to the quantum mechanical model of the atom, what happens when an atom absorbs a quantum of energy?

## Think Critically

- 107. Compare and Contrast** 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.
- 108. Calculate** 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?



■ Figure 5.28

- 109. Describe** the shapes of the atomic orbitals shown in Figure 5.28. Specify their orientations and relate each orbital to a particular type of energy sublevel.
- 110. Infer** Suppose that you live in a universe in which the Pauli exclusion principle states that a maximum of three, rather than two, electrons can occupy a single atomic orbital. Evaluate and explain the new chemical properties of the elements lithium and phosphorus.

## Challenge Problem

- 111. Hydrogen Atom** The hydrogen atom's energy is  $-6.05 \times 10^{-20}$  J when the electron is in the  $n = 6$  orbit and  $-2.18 \times 10^{-18}$  J when the electron is in the  $n = 1$ . Calculate the wavelength of the photon emitted when the electron drops from the  $n = 6$  orbit to the  $n = 1$  orbit. Use the following values:  $h = 6.626 \times 10^{-34}$  J·s and  $c = 3.00 \times 10^8$  m/s.

## Cumulative Review

- 112.** Round 20.56120 g to three significant figures. (Chapter 2)
- 113.** Identify whether each statement describes a chemical property or a physical property. (Chapter 3)
- Mercury is a liquid at room temperature.
  - Sucrose is a white, crystalline solid.
  - Iron rusts when exposed to moist air.
  - Paper burns when ignited.
- 114.** 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)

## Additional Assessment

### WRITING in Chemistry

- 115. Neon Signs** To make neon signs emit different colors, manufacturers often fill the signs with gases other than neon. Write an essay about the use of gases in neon signs and the colors produced by the gases.
- 116. Rutherford's Model** Imagine that you are a scientist in the early twentieth century, and you have just learned the details of a new, nuclear model of the atom proposed by the prominent English physicist Ernest Rutherford. After analyzing the model, you discern what you believe to be important limitations. Write a letter to Rutherford in which you express your concerns regarding his model. Use diagrams and examples of specific elements to help you make your point.

### DBQ Document-Based Questions

**Sodium Vapor** When sodium metal is vaporized in a gas-discharge lamp, two closely spaced, bright yellow-orange lines are produced. Because sodium vapor lamps are electrically efficient, they are used widely for outdoor lighting, such as streetlights and security lighting.

Figure 5.29 shows the emission spectrum of sodium metal. The entire visible spectrum is shown for comparison.

Data obtained from: Volland, W. March 2005. *Spectroscopy: Element Identification and Emission Spectra*.



■ Figure 5.29

- 117.** Differentiate between the two spectra shown above.
- 118.** Sodium's two bright lines have wavelengths of 588.9590 nm and 589.9524 nm. What is the ground-state electron configuration notation for sodium, and how does sodium's electron configuration relate to the lines?
- 119.** Calculate the energies of photons related to the two lines using the relationships expressed in the following equations.

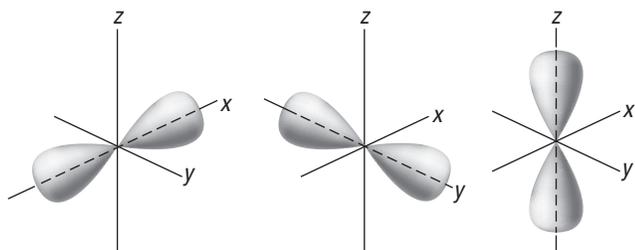
$$E_{\text{photon}} = h\nu; c = \lambda\nu; E = hc/\lambda$$

# Cumulative Standardized Test Practice

## Multiple Choice

- 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.)
  - $8.90 \times 10^{-22} \text{ s}^{-1}$
  - $3.75 \times 10^{12} \text{ s}^{-1}$
  - $8.01 \times 10^{-5} \text{ s}^{-1}$
  - $1.12 \times 10^{21} \text{ s}^{-1}$
- Which is the electron-dot structure for indium?
  - $\cdot \text{In}$
  - $\cdot \text{In} \cdot$
  - $\cdot \dot{\text{In}} \cdot$
  - $\cdot \ddot{\text{In}} \cdot$

Use the figure below to answer Questions 3 and 4.



- To which sublevel do all of these orbitals belong?
  - s
  - p
  - d
  - f
- How many electrons total can reside in this sublevel?
  - 2
  - 3
  - 6
  - 8
- What is the maximum theoretical number of electrons related to the fifth principal energy level of an atom?
  - 10
  - 20
  - 25
  - 50

Use the periodic table and the table below to answer Questions 6 to 8.

**Electron Configurations for Selected Transition Metals**

Element	Symbol	Atomic Number	Electron Configuration
Vanadium	V	23	[Ar]4s <sup>2</sup> 3d <sup>3</sup>
Yttrium	Y	39	[Kr]5s <sup>2</sup> 4d <sup>1</sup>
			[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>6</sup>
Scandium	Sc	21	[Ar]4s <sup>2</sup> 3d <sup>1</sup>
Cadmium	Cd	48	

- Using noble-gas notation, what is the ground-state electron configuration of Cd?
  - [Kr]4d<sup>10</sup>4f<sup>2</sup>
  - [Ar]4s<sup>2</sup>3d<sup>10</sup>
  - [Kr]5s<sup>2</sup>4d<sup>10</sup>
  - [Xe]5s<sup>2</sup>4d<sup>10</sup>
- What is the element that has the ground-state electron configuration [Xe]6s<sup>2</sup>4f<sup>14</sup>5d<sup>6</sup>?
  - La
  - Ti
  - W
  - Os
- What is the complete electron configuration of a scandium atom?
  - 1s<sup>2</sup>2s<sup>2</sup>2p<sup>6</sup>3s<sup>2</sup>3p<sup>6</sup>4s<sup>2</sup>3d<sup>1</sup>
  - 1s<sup>2</sup>2s<sup>2</sup>2p<sup>7</sup>3s<sup>2</sup>3p<sup>7</sup>4s<sup>2</sup>3d<sup>1</sup>
  - 1s<sup>2</sup>2s<sup>2</sup>2p<sup>5</sup>3s<sup>2</sup>3p<sup>5</sup>4s<sup>2</sup>3d<sup>1</sup>
  - 1s<sup>2</sup>2s<sup>1</sup>2p<sup>7</sup>3s<sup>1</sup>3p<sup>7</sup>4s<sup>2</sup>3d<sup>1</sup>
- Which is NOT evidence that a chemical change has occurred?
  - The properties of the substances involved in the reaction have changed.
  - An odor is produced.
  - The composition of the substances involved in the reaction have changed.
  - The total mass of all substances involved has changed.

## Short Answer

Use the data below to answer Questions 10 to 13.

Temperature of Water with Heating	
Time (s)	Temperature (°C)
0	16.3
30	19.7
60	24.2
90	27.8
120	32.0
150	35.3
180	39.6
210	43.3
240	48.1

- Make a graph showing temperature versus time.
- Is the heating of this sample of water a linear process? Explain how you can tell.
- Use your graph to find the approximate rate of heating in degrees per second. What is this value in degrees per minute?
- Show the equation to convert the temperature at 180 s from degrees Celsius to Kelvin and to degrees Fahrenheit.

## Extended Response

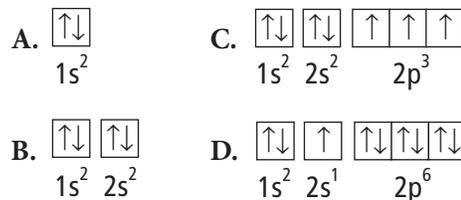
- Compare the information provided in an electron-dot structure with the information in an electron configuration.
- Explain why  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^2$  is not the correct electron configuration for germanium (Ge). Write the correct electron configuration for Germanium.

### NEED EXTRA HELP?

If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19
Review Section . . .	5.1	5.3	5.2	5.2	5.2	5.3	5.3	5.3	3.2	2.4	2.4	2.4	2.1	5.3	5.3	5.3	5.3	2.3	3.3

## SAT Subject Test: Chemistry

Use the diagram below to answer Questions 17 and 18.



- Which shows an orbital diagram that violates the aufbau principle?
  - A
  - B
  - C
  - D
  - none
- Which shows the orbital diagram for the element beryllium?
  - A
  - B
  - C
  - D
  - none
- A student performs an experiment to measure the boiling point of pentane and measures it at  $37.2^\circ\text{C}$ . The literature reports this value as  $36.1^\circ\text{C}$ . What is the student's percent error?
  - 97.0%
  - 2.95%
  - 1.1%
  - 15.5%
  - 3.05%
- Which method of separating components of a mixture depends on the different boiling points of the components of the mixture?
  - chromatography
  - filtration
  - crystallization
  - distillation
  - sublimation