

## 1 FOCUS

## Objectives

- 5.3.1 Describe** the relationship between the wavelength and frequency of light.
- 5.3.2 Identify** the source of atomic emission spectra.
- 5.3.3 Explain** how the frequencies of emitted light are related to changes in electron energies.
- 5.3.4 Distinguish** between quantum mechanics and classical mechanics.

## Guide for Reading

## Build Vocabulary

L2

**Word Parts** The word *emission* comes from the Latin *emittere* meaning “to send out.” The word *spectrum* is the Latin word meaning “appearance.” An atomic emission spectrum allows the light sent out from an atom to appear or be seen. Ask, **What is an emissary?** (*a person sent out to represent another person*)

## Reading Strategy

L2

**Relate Text and Visuals** Refer students to Figure 5.9 and the text on this page. Then, have them write, in their own words, definitions of wavelength, frequency, and amplitude.

## 2 INSTRUCT

## Connecting to Your World

Ask, **What causes an element such as neon to emit light when heated by an electric current?** (*Energy absorbed by electrons is emitted as light.*)

## Light

## Use Visuals

L1

**Figure 5.9** Ask students to assume that both waves are traveling at the same speed. Ask, **Which wave would have more complete wave cycles passing a point in a given time interval?** (*The wave on the right.*) Ask, **Which wave has the higher frequency?** (*The wave on the right.*)

## Guide for Reading



## Key Concepts

- How are the wavelength and frequency of light related?
- What causes atomic emission spectra?
- How are the frequencies of light an atom emits related to changes of electron energies?
- How does quantum mechanics differ from classical mechanics?

## Vocabulary

amplitude  
wavelength  
frequency  
hertz  
electromagnetic radiation  
spectrum  
atomic emission spectrum  
ground state  
photons  
Heisenberg uncertainty principle

## Reading Strategy

## Monitoring Your Understanding

Before you read, preview the Key Concepts, the section heads, the vocabulary terms, and the visuals. List three things you expect to learn. After reading, state what you learned about each item you listed.

## Connecting to Your World

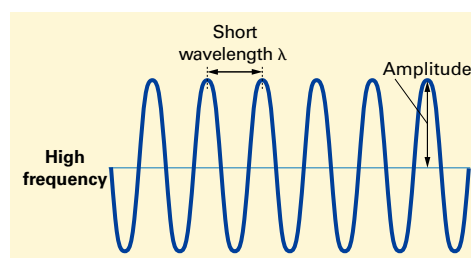
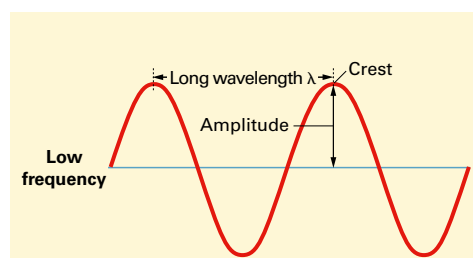
If you walk in the evening along a busy street lined with shops and theaters, you are likely to see neon advertising signs. The signs are formed from glass tubes bent in various shapes. An electric current passing through the gas in each glass tube makes the gas glow with its own characteristic color. In this section you will learn why each gas glows with a specific color of light.



## Light

The previous sections in this chapter introduced you to some ideas about how electrons in atoms are arranged in orbitals, each with a particular energy level. You also learned how to write electron configurations for atoms. In the remainder of this chapter, you will get a closer look into what led to the development of Schrödinger's equation and the quantum mechanical model of the atom.

Rather curiously, the quantum mechanical model grew out of the study of light. Isaac Newton (1642–1727) tried to explain what was known about the behavior of light by assuming that light consists of particles. By the year 1900, however, there was enough experimental evidence to convince scientists that light consists of waves. Figure 5.9 illustrates some of the properties of waves. As shown, each complete wave cycle starts at zero, increases to its highest value, passes through zero to reach its lowest value, and returns to zero again. The **amplitude** of a wave is the wave's height from zero to the crest, as shown in Figure 5.9. The **wavelength**, represented by  $\lambda$  (the Greek letter lambda), is the distance between the crests. The **frequency**, represented by  $\nu$  (the Greek letter nu), is the number of wave cycles to pass a given point per unit of time. The units of frequency are usually cycles per second. The SI unit of cycles per second is called a **hertz** (Hz). A hertz can also be expressed as a reciprocal second ( $s^{-1}$ ).



**Figure 5.9** The frequency and wavelength of light waves are inversely related. As the wavelength increases, the frequency decreases.

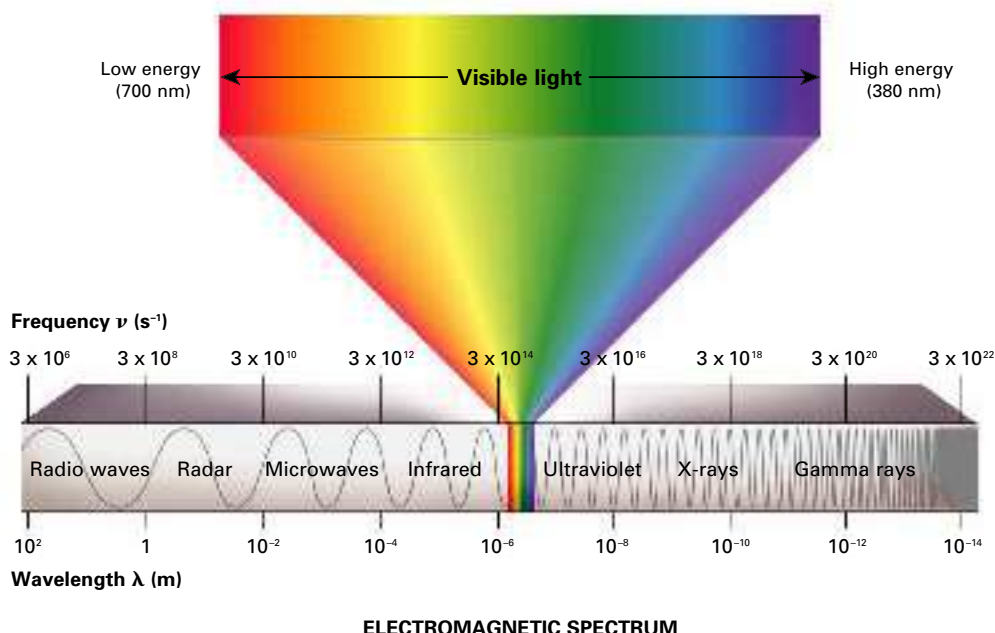
## Section Resources

## Print

- **Guided Reading and Study Workbook**, Section 5.3
- **Core Teaching Resources**, Section 5.3 Review, Interpreting Graphics
- **Laboratory Manual**, Labs 6–8
- **Small-Scale Chemistry Laboratory Manual**, Lab 8
- **Transparencies**, T61–T64

## Technology

- **Interactive Textbook with ChemASAP**, Simulations 3, 4, Problem-Solving 5.15, Animation 6, Assessment 5.3
- **Go Online**, Section 5.3
- **Virtual Chemistry Labs** 1, 7–10



The product of frequency and wavelength always equals a constant ( $c$ ), the speed of light:

$$c = \lambda \nu$$

**The wavelength and frequency of light are inversely proportional to each other.** As the wavelength of light increases, for example, the frequency decreases.

According to the wave model, light consists of electromagnetic waves. **Electromagnetic radiation** includes radio waves, microwaves, infrared waves, visible light, ultraviolet waves, X-rays, and gamma rays. All electromagnetic waves travel in a vacuum at a speed of  $2.998 \times 10^8$  m/s.

Sunlight consists of light with a continuous range of wavelengths and frequencies. As you can see from Figure 5.10, the color of light for each frequency found in sunlight depends on its frequency. When sunlight passes through a prism, the different frequencies separate into a **spectrum** of colors. A rainbow is an example of this phenomenon. Each tiny droplet of water acts as a prism to produce a spectrum. Each color blends into the next in the order red, orange, yellow, green, blue, and violet. In the visible spectrum, as shown in Figure 5.10, red light has the longest wavelength and the lowest frequency.

**Checkpoint** What color in the visible spectrum has the longest wavelength?

**Figure 5.10** The electromagnetic spectrum consists of radiation over a broad band of wavelengths. The visible light portion is very small. It is in the  $10^{-7}$  m wavelength range and  $10^{15}$  Hz ( $s^{-1}$ ) frequency range.

**Interpreting Diagrams** What types of nonvisible radiation have wavelengths close to those of red light? To those of blue light?

**Interactive Textbook**

**Simulation 3** Explore the properties of electromagnetic radiation.

with ChemASAP

## Black Box Discovery

L2

**Purpose** Students discover information about objects without seeing them.

**Materials** Several sealed boxes containing different common items such as rubber bands, golf, tennis, or table-tennis balls, cans of different sizes, and different shapes cut from blocks of wood or foam

**Procedure** Pass the boxes around the room and record student observations and guesses about what is inside each "black box." Allow students to move the boxes and listen to what happens. Point out that they are adding kinetic energy and observing how the hidden object responds. Point out that this method is analogous to the way scientists add energy to atoms and observe how they respond.

**Expected Outcome** If the items are carefully selected, students will be surprised at how much they can determine without opening the boxes. This activity could also be done as a group activity with groups reporting what they find.

## Use Visuals

L1

**Figure 5.10** Ask students to name some types of radiation that are not visible. (*infrared, ultraviolet, X-rays, gamma rays*) Ask, **Are any of these invisible rays more energetic than visible light, and if so, what evidence do you have?** (*UV causes suntan or sunburns, X-rays penetrate flesh, and gamma rays can penetrate bones.*) Ask, **What do all the rays in the electromagnetic spectrum have in common?** (*All travel at the speed of light.*)

## Facts and Figures

### Light as a Particle

In 1666, Sir Isaac Newton (1642–1727) carried out important experiments with light that led to the proposal that light consists of tiny particles called corpuscles. Twelve years later, a Dutch scientist, Christian Huygens, suggested a wave theory to explain the

properties of light. For more than 200 years, scientists argued about these seemingly contradictory theories. By the year 1900, most scientists had finally accepted that light could be described as a wave. Paradoxically, they soon had to accept that light could, at the same time, be described as a particle.

### Answers to...

**Figure 5.10** infrared, ultraviolet

**Checkpoint** red

## Section 5.3 (continued)

### CLASS Activity

#### Properties of Waves

L2

**Purpose** Students gain familiarity with wavelength, frequency, and amplitude.

**Procedure** On a sheet of paper, have students draw a horizontal line approximately 5 inches long. Then, using the line as a midpoint, ask them to draw two waves, superimposed, that have the same amplitudes but different frequencies. Have them draw another horizontal line and two waves with the same wavelength but different amplitudes.

### CHEMath

#### Algebraic Equations

Solving problems in chemistry often involves using algebraic equations such as the one in Sample Problem 5.1. Write the equation  $c = \lambda\nu$  on the board. Stress that the equation must be rearranged in order to solve for  $\lambda$  (lambda), that is,  $\lambda$  must be on one side of the equation by itself. Show students that they can do this by dividing both sides of the equation by  $\nu$  (nu).

#### Math

#### Handbook

For a math refresher and practice, direct students to algebraic equations, page R69.

### CHEMath

#### Algebraic Equations

An algebraic equation shows the relationship between two or more variables. Often, an equation must be solved for the unknown variable before substituting the known values into the equation and doing the arithmetic.

Most equations can be solved if you remember that you can carry out any mathematical operation, such as addition (+), subtraction (−), multiplication (×), or division (x/y or  $x \div y$ ), without destroying the equality, as long as you do it to both sides of the equation.

#### Math

#### Handbook

For help and practice solving algebraic equations, go to page R69.



**Problem-Solving 5.15** Solve Problem 15 with the help of an interactive guided tutorial.

with ChemASAP

**Figure 5.11** Sodium vapor lamps produce a yellow glow.

### SAMPLE PROBLEM 5.1

#### Calculating the Wavelength of Light

Calculate the wavelength of the yellow light emitted by the sodium lamp shown above if the frequency of the radiation is  $5.10 \times 10^{14}$  Hz ( $5.10 \times 10^{14}$ /s).

#### 1 Analyze List the knowns and the unknown.

##### Knowns

- frequency ( $\nu$ ) =  $5.10 \times 10^{14}$ /s
- $c = 2.998 \times 10^8$  m/s

##### Unknown

- wavelength ( $\lambda$ ) = ? m

#### 2 Calculate Solve for the unknown.

Solve the equation  $c = \lambda\nu$  for  $\lambda$ .

$$\lambda = \frac{c}{\nu}$$

Substitute the known values and solve.

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{5.10 \times 10^{14}/\text{s}} = 5.88 \times 10^{-7} \text{ m}$$

#### 3 Evaluate Does the result make sense?

The magnitude of the frequency is much larger than the numerical value of the speed of light, so the answer should be much less than 1. The answer should have three significant figures, because the original known value had three significant figures.

### Practice Problems

- What is the wavelength of radiation with a frequency of  $1.50 \times 10^{13}$  Hz? Does this radiation have a longer or shorter wavelength than red light?
- What is the frequency of radiation with a wavelength of  $5.00 \times 10^{-8}$  m? In what region of the electromagnetic spectrum is this radiation?

### Sample Problem 5.2

#### Answers

- $2.00 \times 10^{-5}$  m; longer wavelength than red light
- $6.00 \times 10^{15} \text{ s}^{-1}$ ; ultraviolet

#### Practice Problems Plus

L2

**What is the frequency of radiation which has a wavelength of  $7.00 \times 10^{-5}$  cm? ( $4.29 \times 10^{14} \text{ s}^{-1}$ ) In what region of the electromagnetic spectrum is this radiation? (infrared)**

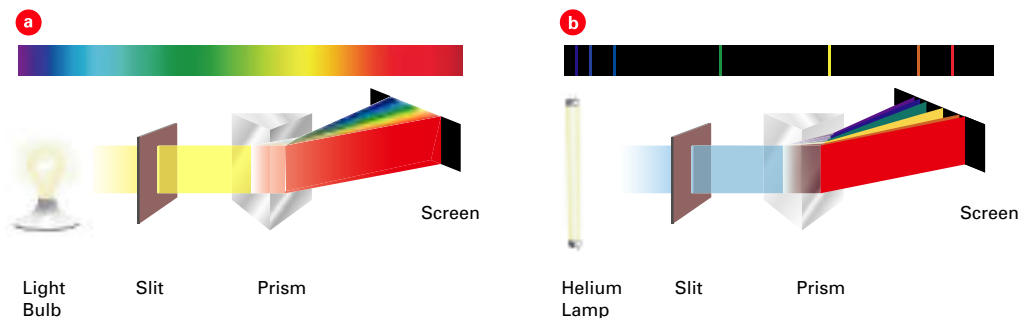
## Differentiated Instruction

### Special Needs

L1

When introducing the formulas and units in this section, be sure to review mathematical operations for multiplication and division using scientific notation. Also be sure to

show that  $1/10 = 0.10 = 10^{-1}$ . This will help students understand the units used for frequency: cycles per second = cycles/s =  $\text{s}^{-1}$ , or hertz (Hz).

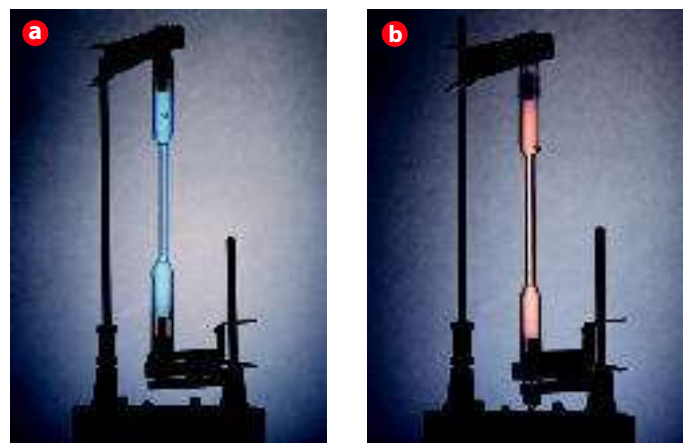


## Atomic Spectra

Passing an electric current through a gas in a neon tube energizes the electrons of the atoms of the gas, and causes them to emit light. **When atoms absorb energy, electrons move into higher energy levels. These electrons then lose energy by emitting light when they return to lower energy levels.** Figure 5.12a shows how ordinary light is made up of a mixture of all the wavelengths of light. However, the light emitted by atoms consists of a mixture of only specific frequencies. Each specific frequency of visible light emitted corresponds to a particular color. Therefore, when the light passes through the prism shown in Figure 5.12b, the frequencies of light emitted by an element separate into discrete lines to give the **atomic emission spectrum** of the element.

Each discrete line in an emission spectrum corresponds to one exact frequency of light emitted by the atom. Figure 5.12b shows the visible portion of the emission spectrum of helium.

The emission spectrum of each element is like a person's fingerprint. Just as no two people have the same fingerprints, no two elements have the same emission spectrum. In the same way that fingerprints identify people, atomic emission spectra are useful for identifying elements. Figure 5.13 shows the characteristic colors emitted by mercury and by nitrogen. Much of the knowledge about the composition of the universe comes from studying the atomic spectra of the stars, which are hot glowing bodies of gases.



**Figure 5.12** A prism separates light into the colors it contains.

**a** For white light this produces a rainbow of colors. **b** Light from a helium lamp produces discrete lines. **Identifying Which color has the highest frequency?**

**Figure 5.13** No two elements have the same emission spectrum. **a** Mercury vapor lamps produce a blue glow. **b** Nitrogen gas gives off a yellowish-orange light.

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## Atomic Spectra

### Use Visuals

L1

**Figures 5.12 and 5.13** Ask, **What is the difference between the spectrum of light from an ordinary light bulb and the spectrum from a helium lamp?** (The light bulb spectrum shows a continuous display of all the colors of visible light. The helium lamp shows discrete bands of only certain frequencies of light.) Explain that the frequencies correspond to transitions of electrons in the helium atom. Direct students to Figure 5.13 and tell them that high voltage is required to measure the emission spectrum of a diatomic element, such as nitrogen, because the molecules must first be dissociated into atoms. The separated atoms release the absorbed energy in discrete packets with definite frequencies. Each different frequency is observed as a line in the nitrogen or mercury vapor emission spectrum. Scientists who were trying to understand the inner architecture of atoms recognized a relationship between the emission spectrum of an atom and its atomic structure. Ask, **Why is hydrogen a good model for studying the correlation between emission spectrum and atomic structure?** (Hydrogen, with only one electron, has the simplest electron configuration and thus the simplest emission spectrum.)

### Discuss

L2

Remind students of the Inquiry Activity on page 126. Crushing the mint creates an unequal division of electrons: positive charge accumulates on one piece, negative charge on the other. When an electric spark jumps between the pieces, nitrogen atoms in the air absorb the energy, emitting it as ultra-violet light. The wintergreen flavor molecules absorb part of this energy and release it as visible blue-green light. All wintergreen mints should emit light. To see the light, students must allow enough time for their eyes to adapt to the dark room.

## Differentiated Instruction

### Gifted and Talented

L3

Students may welcome the opportunity to design a system for refracting light and projecting a rainbow on a wall or screen. Depending on their setup, they may need a prism and a bright light source.

### Answers to...

**Figure 5.12** violet

## Quick LAB

### Flame Tests

L2

**Objective** Students use flame tests to determine the identity of the cation in an unknown solution.

**Skills Focus** Observe, interpret, identify

**Prep Time** 30 minutes

**Materials** Bunsen burner, 6 small test tubes, test tube rack, tongs, 6 cotton swabs, 0.1M NaCl, 0.1M CaCl<sub>2</sub>, 0.1M LiCl, 0.1M CuCl<sub>2</sub>, 0.1M KCl, unknown solution

#### Advance Prep

| Solution               | Preparation                             |
|------------------------|-----------------------------------------|
| 0.1M NaCl              | 0.6 g NaCl in 100 mL water              |
| 0.5M CaCl <sub>2</sub> | 5.5 g CaCl <sub>2</sub> in 100 mL water |
| 0.1M LiCl              | 0.5 g LiCl in 100 mL water              |
| 0.1M CuCl <sub>2</sub> | 1.3 g CuCl <sub>2</sub> in 100 mL water |

**Class Time** 20 minutes

**Safety** Students should wear safety goggles and a lab apron and observe all safety procedures.

#### Expected Outcome

|                  |                    |
|------------------|--------------------|
| Na <sup>+</sup>  | bright yellow      |
| Ca <sup>2+</sup> | orange-red         |
| Li <sup>+</sup>  | bright crimson red |
| Cu <sup>2+</sup> | blue-green         |

## Quick LAB

### Flame Tests

#### Purpose

Use the flame test to determine the identity of the cation in an unknown solution based on its characteristic color in a flame.

#### Materials

- Bunsen burner
- 6 small test tubes
- test tube rack
- tongs
- 6 cotton swabs
- 0.1M NaCl
- 0.1M CaCl<sub>2</sub>
- 0.1M LiCl
- 0.1M CuCl<sub>2</sub>
- 0.1M KCl
- unknown solution

#### Procedure



1. Make a two-column data table. Label the columns Cation and Flame Color. Enter the cation's name for each salt solution in the first column.
2. Label each of 5 test tubes with the name of a salt solution; label the sixth tube Unknown. Add 1 mL of each salt solution to the appropriately labeled test tube.
3. Dip one of the cotton ends of a cotton swab into the sodium chloride solution and then hold it briefly in the burner flame. Record the color of the flame. Do not leave the swab in the flame too long or the plastic will melt.
4. Repeat Step 3 for each of the remaining salt solutions using a new cotton swab each time.
5. Perform a flame test with the unknown solution. Note the color of the flame.



#### Analyze and Conclude

1. Identify the cation in the unknown.
2. Each salt solution produces a unique color. Would you expect this based on the modern view of the atom? Explain.
3. Some commercially available fireplace logs burn with a red and/or green flame. On the basis of your data, what elements could be responsible for these colored flames?
4. Aerial fireworks contain gunpowder and chemicals that produce colors. What element would you include to produce crimson red? Yellow?

### An Explanation of Atomic Spectra

Atomic line spectra were known before Bohr proposed his model of the hydrogen atom. However, Bohr's model not only explained why the emission spectrum of hydrogen consists of specific frequencies of light. It also predicted specific values of these frequencies that agreed with experiment.

In the Bohr model, the lone electron in the hydrogen atom can have only certain specific energies. When the electron has its lowest possible energy, the atom is in its **ground state**. In the ground state, the principal quantum number ( $n$ ) is 1. Excitation of the electron by absorbing energy raises the atom from the ground state to an excited state with  $n = 2, 3, 4, 5$ , or 6, and so forth. A quantum of energy in the form of light is emitted when the electron drops back to a lower energy level. The emission occurs in a single abrupt step, called an electronic transition. Bohr already knew from earlier work that this quantum of energy  $E$  is related to the frequency  $\nu$  of the emitted light by the equation  $E = h \times \nu$ , where  $h$  is equal to  $6.626 \times 10^{-34}$  J·s.  **The light emitted by an electron moving from a higher to a lower energy level has a frequency directly proportional to the energy change of the electron.** Therefore each transition produces a line of a specific frequency in the spectrum.

#### Analyze and Conclude

1. Answer depends on the unknown.
2. Yes; because the composition of every atom is different, the amount of energy needed for transitions between energy sub-levels is different. These energy differences are reflected in the different line spectra.

3. copper (green), lithium (red), or calcium (orange-red)
4. red, lithium; yellow, sodium

#### For Enrichment

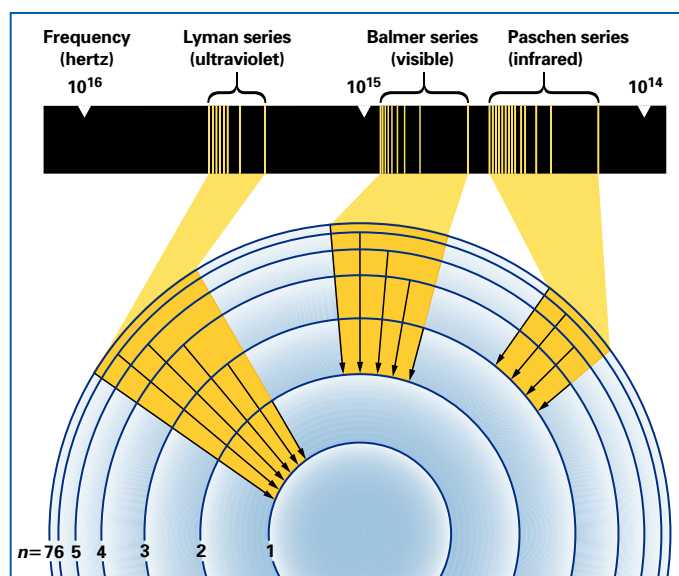
L3

Mix two of the unknown solutions. Ask students to identify the two metal ions.

Figure 5.14 shows the explanation for the three groups of lines observed in the emission spectrum of hydrogen atoms. The lines at the ultraviolet end of the hydrogen spectrum are the Lyman series. These match expected values for the emission due to the transition of electrons from higher energy levels to the lowest energy level,  $n = 1$ . The lines in the visible spectrum are the Balmer series. These lines result from transitions from higher energy levels to  $n = 2$ . This generally involves a smaller change in electron energy than transitions to  $n = 1$ . Transitions to  $n = 3$  from higher energy levels produce the Paschen series. The energy changes of the electron, and therefore the frequencies of emitted light, are generally smaller still. The lines are in the infrared range. Spectral lines for the transitions from higher energy levels to  $n = 4$  and  $n = 5$  also exist. Note that the spectral lines in each group become more closely spaced at increased values of  $n$  because the energy levels become closer together. There is an upper limit to the frequency of emitted light for each set of lines. The upper limit exists because an electron with enough energy completely escapes the atom.

Bohr's theory of the atom was only partially satisfactory. It explained the emission spectrum of hydrogen but not the emission spectra of atoms with more than one electron. Moreover, it was of no help in understanding how atoms bond to form molecules. Eventually a new and better model, the quantum mechanical model, displaced the Bohr model of the atom. The quantum mechanical model is based on the description of the motion of material objects as waves.

**Checkpoint** What is the name of the series of visible lines in the hydrogen spectrum?



**Figure 5.14** The three groups of lines in the hydrogen spectrum correspond to the transition of electrons from higher energy levels to lower energy levels. The Lyman series corresponds to the transition to the  $n = 1$  energy level. The Balmer series corresponds to the transition to the  $n = 2$  energy level. The Paschen series corresponds to the transition to the  $n = 3$  energy level.

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## An Explanation of Atomic Spectra

### Use Visuals

L1

**Figure 5.14** Pose this question to students: **Suppose an electron, in its ground state at energy level one, absorbs enough energy to jump to level two. What type of radiation will it emit when it returns to the ground state?** (UV radiation) Ask, **If you observed a hydrogen gas discharge tube through a diffraction grating, could you see the line corresponding to this emission?** (No; the human eye cannot detect radiation in the UV range.) **Which series of lines could you detect?** (the Balmer series, which has frequencies in the visible region of the spectrum) **Compare the energy of the Paschen and Balmer series.** (The Paschen series has lower energy.) Ask, **What do you notice about the spacing of the energy levels from  $n = 1$  to  $n = 7$ ?** (The levels are not evenly spaced. The lines get closer as the distance from the nucleus increases.) Note that each set of lines is a converging series that approaches the ionization energy of hydrogen.

## Facts and Figures

### The Discovery of Helium

Sometimes discoveries in one area of science, such as chemistry, become important in solving problems in another, such as astronomy. In 1868, Pierre Janssen and Joseph Norman Lockyer discovered an emission spectrum from gases on the surface of the sun that did not match any known element on Earth. In 1895, William Ramsay dis-

covered the existence of helium on Earth. The emission spectrum of helium was found to be identical to that of the unknown gas observed by Janssen and Lockyer almost thirty years earlier. Thus, by combining two discoveries from two different fields of science, a new discovery was made: Earth and the stars have some elements in common.

### Answers to...

**Checkpoint** the Balmer series

## Quantum Mechanics

### Discuss

L2

Tell students that, according to quantum mechanics, scientists cannot observe an object without changing the object in the process. The change produced by making an observation is the basis for the uncertainty principle. Present the following analogy: Suppose you want to find out if a car is in a long tunnel. In quantum mechanics, the only sort of experiment you could do would be to send another car into the tunnel and listen for a crash. Although it is possible to detect the presence of a car with this method, it is obvious that the car will be changed by the crash.

### Relate

L2

The wave-particle behavior of light was difficult for scientists trained in classical physics to accept. However, it became difficult to dispute because it provided an explanation for a previously mysterious phenomenon called the photoelectric effect. In the photoelectric effect, it was shown that metals eject electrons called photoelectrons when light shines on them. But not just any frequency of light will cause the photoelectric effect. Red light ( $\nu = 4.3 \times 10^{14} \text{ s}^{-1}$  to  $4.6 \times 10^{14} \text{ s}^{-1}$ ), for example, will not cause the ejection of photoelectrons from potassium, no matter how intense the light or for how long the light radiates the metal surface. But even a weak yellow light ( $\nu = 5.1 \times 10^{14} \text{ s}^{-1}$  to  $5.2 \times 10^{14} \text{ s}^{-1}$ ) shining on potassium causes the emission of photoelectrons. Thus, photons of red light do not have sufficient energy to cause the ejection of electrons, but photons of yellow light do.



**Figure 5.15** An electron microscope can produce sharp images of a very small object, such as this mite, because of the small wavelength of a moving electron compared with that of light.

## Quantum Mechanics

In 1905, Albert Einstein, then a patent examiner in Bern, Switzerland, returned to Newton's idea of particles of light. Einstein successfully explained experimental data by proposing that light could be described as quanta of energy. The quanta behave as if they were particles. Light quanta are called **photons**. Although the wave nature of light was well known, the dual wave-particle behavior of light was difficult for scientists trained in classical physics to accept.

In 1924, Louis de Broglie (1892–1987), a French graduate student, asked an important question: Given that light behaves as waves and particles, can particles of matter behave as waves? De Broglie referred to the wavelike behavior of particles as matter waves. His reasoning led him to a mathematical expression for the wavelength of a moving particle. The proposal that matter moves in a wavelike way would not have been accepted unless experiments confirmed its validity. Only three years later, experiments by Clinton Davisson and Lester Germer at Bell Labs in New Jersey did just that. The two scientists had been studying the bombardment of metals with beams of electrons. They noticed that the electrons reflected from the metal surface produced curious patterns. The patterns were like those obtained when X-rays (which are electromagnetic waves) reflect from metal surfaces. The electrons—believed to be particles—were reflected as if they were waves! De Broglie was awarded the Nobel Prize for his work on the wave nature of matter. Davisson also received the Nobel Prize for his experiments demonstrating the wave nature of electrons.

Today, the wavelike properties of beams of electrons are useful in magnifying objects. The electrons in an electron microscope have much smaller wavelengths than visible light. This allows a much clearer enlarged image of a very small object, such as the mite in Figure 5.15, than is possible with an ordinary microscope.

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Download a worksheet on the **Photoelectric Effect** for students to complete, and find additional teacher support from NSTA SciLinks.

## Facts and Figures

### Planck's Constant

Max Planck (1858–1947) showed that energy could be absorbed or emitted by a body only in quanta whose energy is given by the equation  $E = h \times \nu$ , where  $h$  is a proportionality constant called Planck's constant. Planck's proposal was revolutionary. Everyday experience had led people to believe that no limitation existed for the smallness of

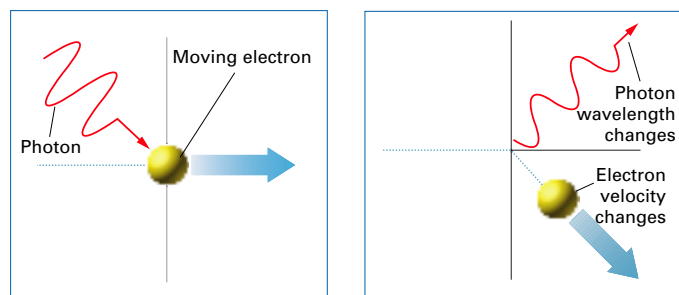
permissible energy changes in a system. For example, appearances would lead you to believe that thermal energy may be continuously supplied to heat liquid water to any temperature between  $0^{\circ}\text{C}$  and  $100^{\circ}\text{C}$ . Actually, the water temperature increases by infinitesimally small steps that occur as individual molecules absorb quanta of energy.

De Broglie's equation predicts that all moving objects have wavelike behavior. Why are you unable to observe the effects of this wavelike motion for ordinary objects like baseballs and trains? The answer is that the mass of the object must be very small in order for its wavelength to be large enough to observe. For example, a 50-gram golf ball traveling at 40 m/s (about 90 mi/h) has a wavelength of only  $3 \times 10^{-34}$  m, which is much too small to detect experimentally. On the other hand, an electron has a mass of only  $9.11 \times 10^{-28}$  g. If it were moving at a velocity of 40 m/s, it would have a wavelength of  $2 \times 10^{-5}$  m, which is comparable to infrared radiation and is readily measured. The newer theory is called quantum mechanics; the older theory is called classical mechanics.  **Classical mechanics adequately describes the motions of bodies much larger than atoms, while quantum mechanics describes the motions of subatomic particles and atoms as waves.**

German physicist Werner Heisenberg examined another feature of quantum mechanics that is absent in classical mechanics. The **Heisenberg uncertainty principle** states that it is impossible to know exactly both the velocity and the position of a particle at the same time. This limitation is critical in dealing with small particles such as electrons. The Heisenberg uncertainty principle does not matter, however, for ordinary-sized objects such as cars or airplanes.

To understand this principle, consider how you determine the location of an object. To locate a set of keys in a dark room, for example, you can use a flashlight. You see the keys when the light bounces off them and strikes your eyes. Likewise, to locate an electron, you might strike it with a photon of light as shown in Figure 5.16. In contrast to the keys, the electron has such a small mass that striking it with a photon affects its motion in a way that cannot be predicted precisely. So the very act of measuring the position of the electron changes its velocity, and makes its velocity uncertain.

The discovery of matter waves paved the way for Schrödinger's quantum mechanical description of electrons in atoms. Schrödinger's theory leads to the concept of electron orbitals and configurations, and it includes the wavelike motion of matter and the uncertainty principle.



#### Before collision

A photon strikes an electron during an attempt to observe the electron's position.

#### After collision

The impact changes the electron's velocity, making it uncertain.

### Interactive Textbook

**Simulation 4** Simulate the photoelectric effect. Observe the results as a function of radiation frequency and intensity.

with **ChemASAP**

**Figure 5.16** The Heisenberg uncertainty principle states that it is impossible to know exactly both the velocity and the position of a particle at the same time.

## CLASS Activity

### Atomic Scientists

L2

Numerous physicists contributed to modern theories of light and atoms. Dalton, Becquerel, Roentgen, Rutherford, Thomson, Millikan, Planck, Einstein, de Broglie, Heisenberg, Schrödinger, Bohr, Chadwick, Balmer, Huygens, Newton, and Foucault all made contributions in one of the most productive periods of scientific discovery.

**Purpose** Students acquire a deeper understanding and appreciation for the achievements of these scientists.

**Procedure** Have students work in teams of two or three. Assign scientists to each group and ask students to prepare a brief oral discussion highlighting how the work of the scientists contributed to the understanding of light or the atom. Encourage students to consider the work of other scientists they encounter in their research. Contributions at the periphery can help students understand the interactive and collaborative nature of scientific research.

**Expected Outcome** Class presentations allow students to share information.

## 3 ASSESS

### Evaluate Understanding

L2

Have students draw and label wave diagrams that illustrate the relationship between wavelength and frequency. Ask students to explain the origin of the lines in the Lyman, Balmer, and Paschen series in the atomic emission spectrum of hydrogen. (Each line arises from light emitted by an atom as electrons move from a higher energy level to a lower one. The Lyman series is in the ultraviolet region of the spectrum, the Balmer is in the visible, and the Paschen is in the infrared.)

## Facts and Figures

### A Scientific Controversy

Albert Einstein and Niels Bohr had different views of quantum mechanics. Although Einstein accepted quantum mechanical theory as the best explanation available at the time, he was convinced the theory was incomplete. He believed in the existence of "hidden

variables" which, if known, would eliminate problems such as wave/particle duality and uncertainty. Bohr, on the other hand, urged physicists to accept quantum mechanics without qualification.

## Careers in Chemistry

### Spectroscopist

Astronomers use spectroscopy to analyze light from the stars. Atomic absorption spectroscopists study how atoms absorb radiation. X-ray spectroscopists use X-rays to determine the crystal structure of substances. Have students explore these and other fields of study that use spectroscopy for gathering data.

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Have students research chemistry-related careers in the library or on the Internet. Students can then construct a table that describes the nature of the work, educational and training requirements, employment outlook, working conditions, and other necessary information.

### Reteach

L1

Use the illustrations in this section as a basis for review. Starting with Figure 5.9 and continuing to Figure 5.16, question students about what is being illustrated and what is significant about it.

### Elements Handbook

Students will find that strontium produces red light and barium produces yellow-green light. Students should recognize that these colors, when separated by a spectroscope, are the colors of the identifying lines for strontium and barium and arise from electrons moving from higher energy levels to lower energy levels.

### Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 5.3.

with ChemASAP

## Spectroscopist

If you like the idea of finding the chemical content of unknown materials in chemical research, police investigations, and studies of distant stars, you might consider a career as a spectroscopist. Spectroscopy is the recording and analysis of the wavelengths of electromagnetic radiation emitted by samples of materials. Optical emission spectroscopy uses emission lines from atomic transitions in a heated sample of material. Spectroscopists observe emission lines from the sample by using an electronic detector and recording its output. The recorded data gives the wavelength and the intensity of each emission line. The characteristic pattern of wavelengths and intensities is

the emission spectrum of the sample.

Spectroscopists use spectrometers, densitometers, and other measuring instruments to collect data. They analyze the densitometer or spectrometer readings to find the ratio of various elements in the sample. They calculate the relative concentrations of substances in the sample by comparing with data for known concentrations. They also use their mathematical skills in statistics to calculate a numerical value indicating the reliability of each analysis.

Spectroscopists usually have an advanced degree in chemistry, along with skills in mathematics and in using scientific equipment.



Spectroscopist

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## 5.3 Section Assessment

- Key Concept** How are wavelength and frequency of light related?
- Key Concept** Describe the cause of atomic emission spectrum of an element.
- Key Concept** How is the change in electron energy related to the frequency of light emitted in atomic transitions?
- Key Concept** How does quantum mechanics differ from classical mechanics?
- The lines at the ultraviolet end of the hydrogen spectrum are known as the Lyman series. Which electron transitions within an atom are responsible for these lines?
- Arrange the following in order of decreasing wavelength.
  - infrared radiation from a heat lamp
  - dental X-rays
  - signal from a shortwave radio station

### Elements Handbook

**Color and Transitions** Look at the photographs of flame tests on page R11 of the Elements Handbook. List the colors emitted from strontium compounds and from barium compounds when heated in a flame, and explain how electron transitions account for the specific colors being emitted.

### Interactive Textbook

**Assessment 5.3** Test yourself on the concepts in Section 5.3.

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## Section 5.3 Assessment

- Frequency and wavelength of light are inversely proportional to each other.
- Electrons in atoms absorb energy as they move to higher energy levels, then lose the energy by emitting it as light as they drop back.
- The light emitted in an electronic transition from a higher to a lower energy level has a frequency that is directly proportional to the energy change of the electron.
- Quantum mechanics describes the motions of atoms and subatomic particles; classical mechanics describes the motions of larger bodies.
- electron transitions from higher levels to  $n = 1$
- c, a, b

## Lasers at Work

A laser produces an intense beam of light with the unusual property that the waves contributing to it have crests that coincide. The properties of laser light make it useful for many different purposes, including reading CD-ROMs and DVDs in electronic equipment and scanning bar codes at cash registers. Some industries use high-intensity lasers to cut metal.

Other lasers are used in surgery. **Inferring Explain** *how the properties of a laser beam make it useful for the kind of eye surgery shown.*

**1** A surgeon prepares for laser eye surgery to reshape the cornea, or clear outermost part of the eye, so that the patient no longer needs eyeglasses to see clearly.

**2** The surgeon uses a tool called a microkeratome to make a partial cut in the cornea.

**3** The doctor lifts and folds back the cornea.

**4** A laser removes some tissue to change the shape of the cornea and correct the patient's vision.



**Flashlight beam vs. laser beam** Light from a flashlight travels in different directions, covering a wide area. Light rays from a laser travel parallel to one another, resulting in a narrow beam of intense light in a specific direction.



## Lasers at Work

Have students note the narrow beam of the laser in the photo compared with the flashlight beam. Explain that the intensity and narrowness of the laser beam allows the surgeon to make a precise surgical cut and forestall bleeding. Laser is an acronym for *light amplification by the stimulated emission of radiation*.

Encourage students to find out how coherent light—light with one wavelength—is produced. (*When a photon of a particular frequency strikes an atom, the photon stimulates the atom to emit a photon of the same frequency. The original photon is unaffected. Now two photons of the same frequency can collide with other atoms and stimulate them to produce identical photons. The process goes on building up an aggregate of photons all with the same frequency.*) Students can investigate what substances are stimulated to create laser light and what wavelengths correspond to each substance. (*Some typical lasers are nitrogen,  $\lambda = 337$  nm; helium-cadmium,  $\lambda = 441.6$  nm; argon,  $\lambda = 476.5$  and  $488.0$  nm; krypton,  $\lambda = 534.5$  nm; ruby,  $\lambda = 694.3$  nm; neon,  $\lambda = 632.8$  nm; gallium arsenide,  $\lambda = 840$ – $930$  nm, depending on temperature; neodymium,  $\lambda = 1040$  nm; carbon dioxide,  $\lambda = 10,600$  nm.)*

Some students could research the technology of reading CD-ROMs and DVDs with lasers and the use of lasers in fiber optics and holograms. Others may prefer to find out more about how nearsightedness is cured by laser surgery. Students could research the history of laser development, a “hot” topic of research in the 1950s, and find out what is the magic behind holograms.

### Answers to...

The laser beam is a narrow beam of intense light that can be aimed in a specific direction.