5.3 Physics and the Quantum Mechanical Model

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 λ (the Greek letter lambda), is the distance between the crests. The **frequency,** represented by ν (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}).

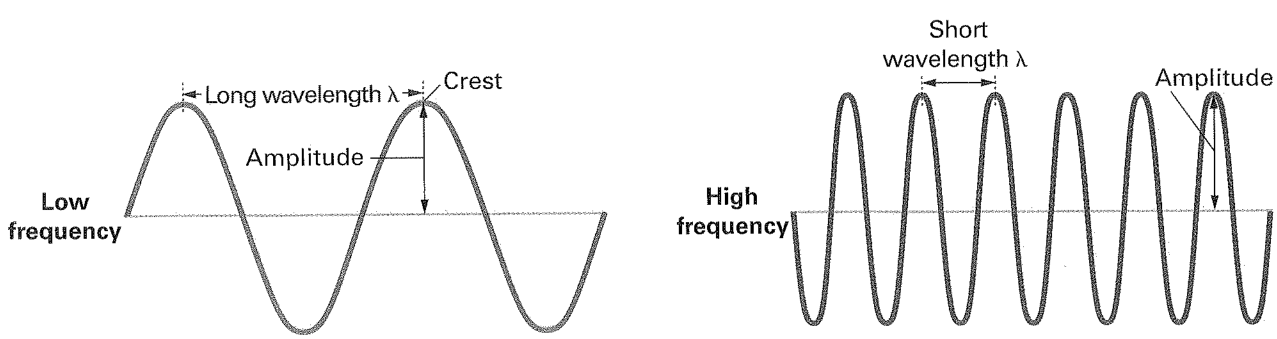
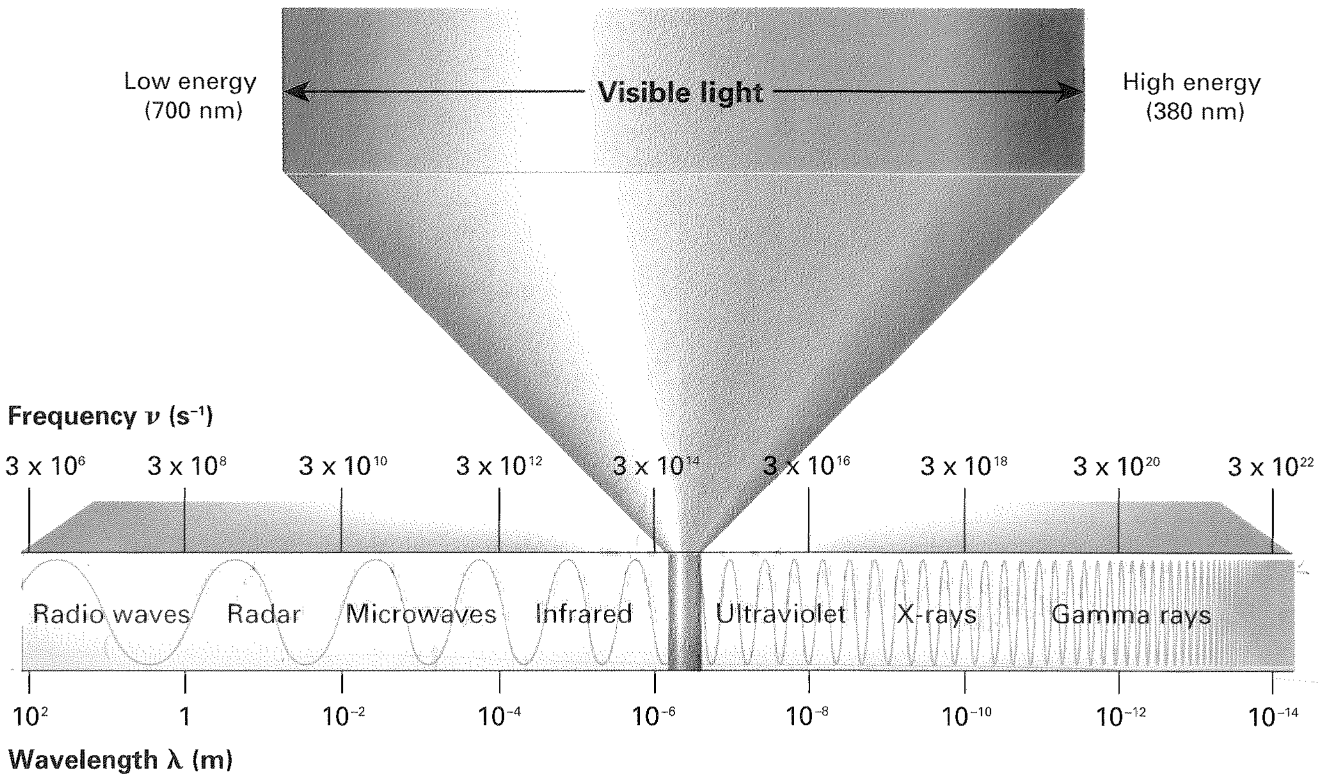


Figure 5.9 The frequency and wavelength of light waves are inversely related. As the wavelength increases, the frequency decreases.



ELECTROMAGNETIC SPECTRUM

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?***

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 2.998×10^8 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?***

Interactive Textbook

Simulation 3 Explore the properties of electromagnetic radiation.

with ChemASAP

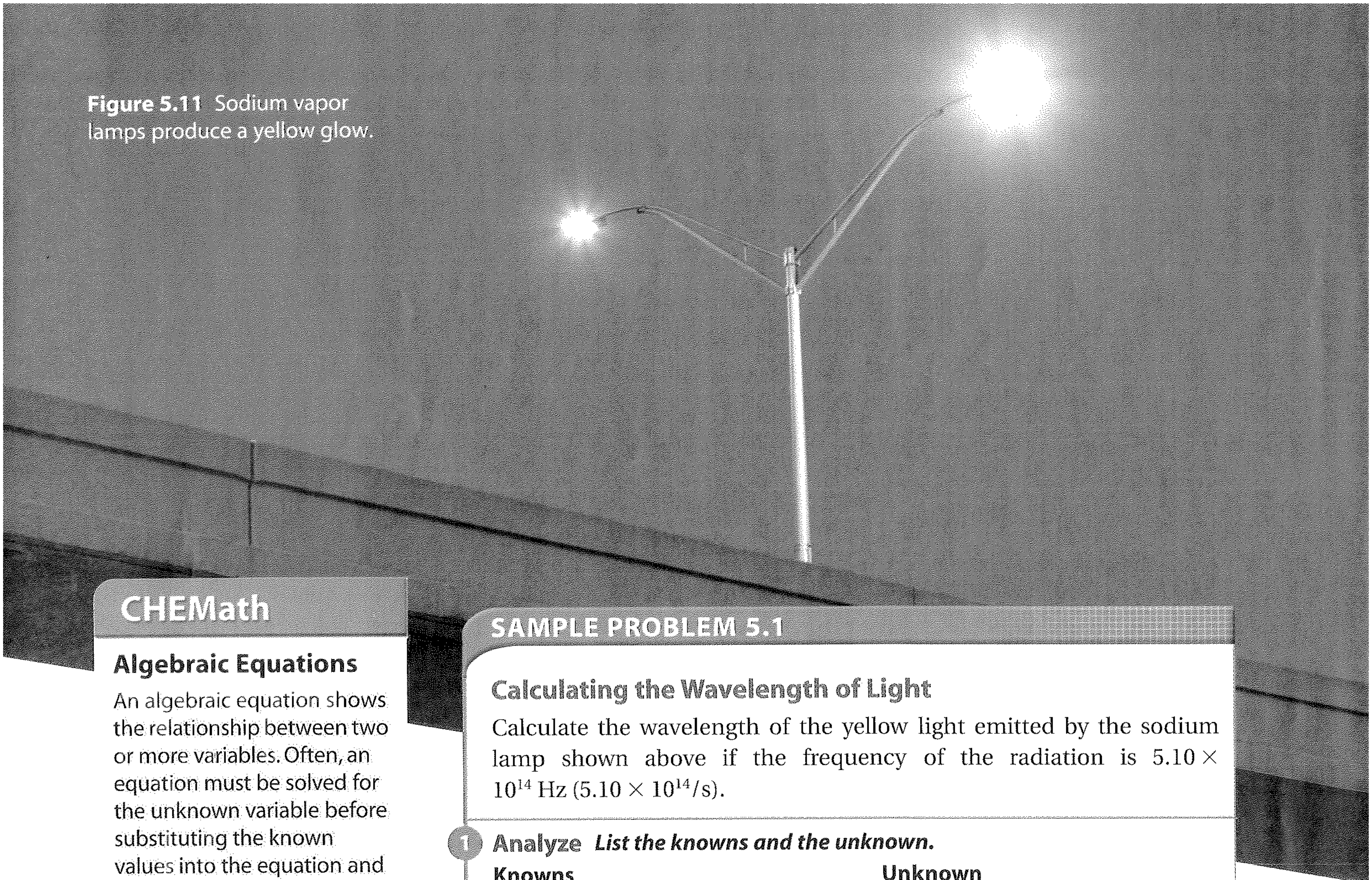


Figure 5.11 Sodium vapor lamps produce a yellow glow.

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.

Interactive Textbook

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

with ChemASAP

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 5.10×10^{14} Hz (5.10×10^{14}/s).

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

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

Unknown
- wavelength (λ) = ? m

2 Calculate ***Solve for the unknown.***

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

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

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

14. What is the wavelength of radiation with a frequency of 1.50×10^{13} Hz? Does this radiation have a longer or shorter wavelength than red light?

15. What is the frequency of radiation with a wavelength of 5.00×10^{-8} m? In what region of the electromagnetic spectrum is this radiation?

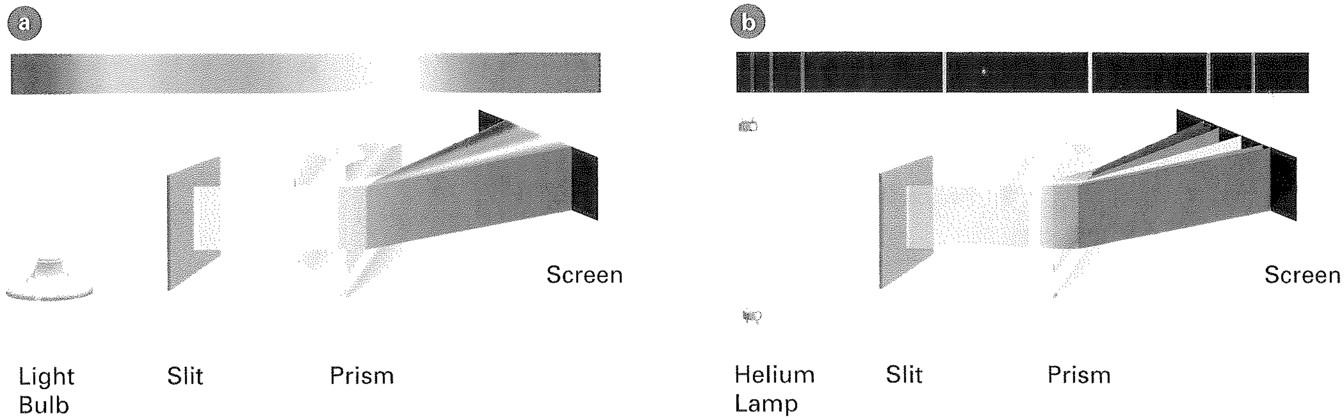


Figure 5.12 A prism separates light into the colors it contains. (a) For white light this produces a rainbow of colors. (b) Light from a helium lamp produces discrete lines. **Identifying** ***Which color has the highest frequency?***

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.

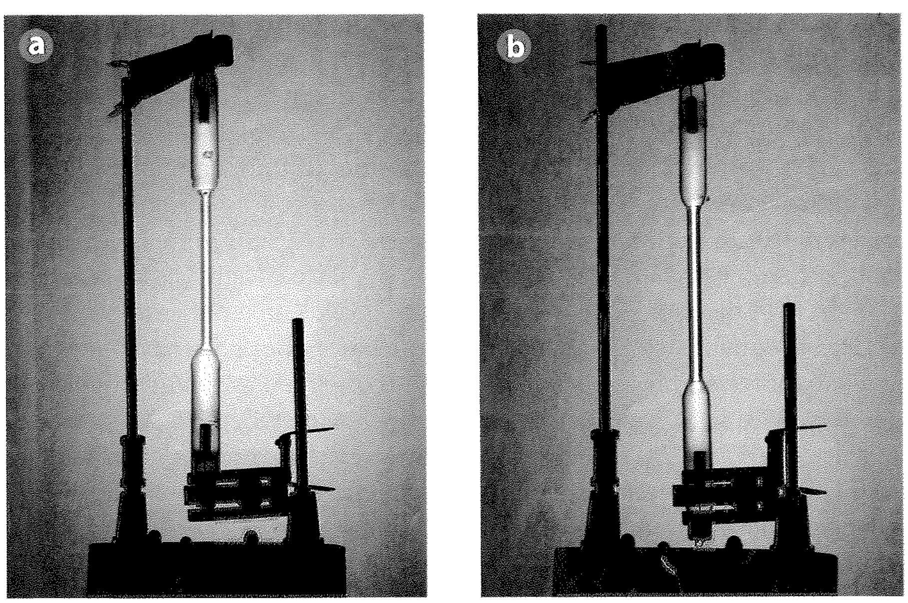


Figure 5.13 No two elements have the same emission spectrum. (a) Mercury vapor lamps produce a blue glow. (b) Nitrogen gas gives off a yellowish-orange light.

Quick LAB

Flame Tests

Purpose

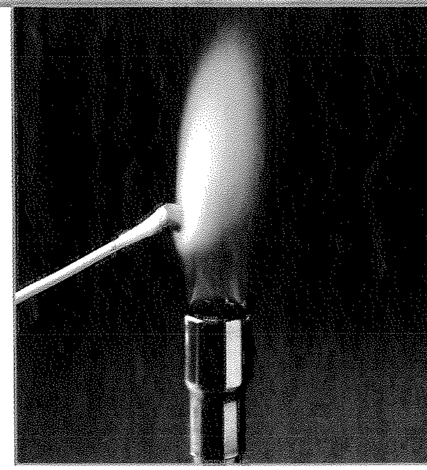
Use the flame test to determine the identity of the cation in an unknown solution based on its characteristic color in a flame.

Materials

- Bunsen burner
- 6 small test tubes
- test tube rack
- tongs
- 6 cotton swabs
- 0.1*M* NaCl
- 0.1*M* $CaCl_2$
- 0.1*M* LiCl
- 0.1*M* $CuCl_2$
- 0.1*M* 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 ν of the emitted light by the equation $E = h \times \nu$, where h is equal to 6.626×10^{-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.

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?***

Interactive Textbook

Animation 6 Learn about atomic emission spectra and how neon lights work.

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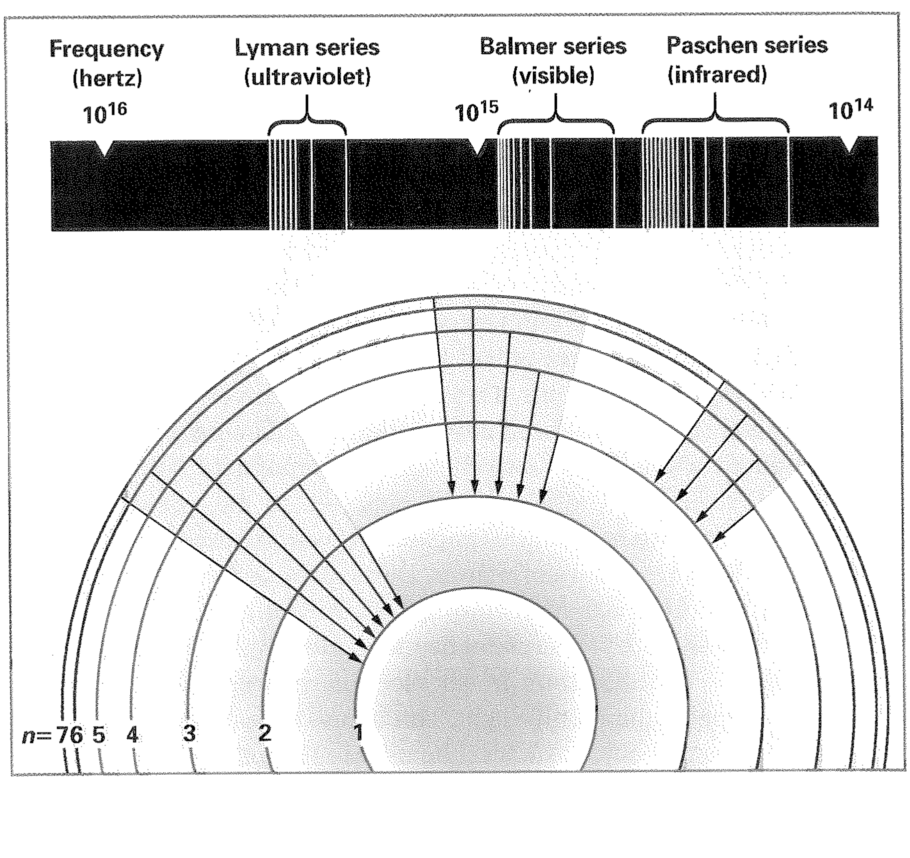

Figure 5.14 The three groups of lines in the hydrogen spectrum correspond to the transitions of electrons from higher energy levels to lower energy levels. The Lyman series corresponds to the transitions to the $n = 1$ energy level. The Balmer series corresponds to the transitions to the $n = 2$ energy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

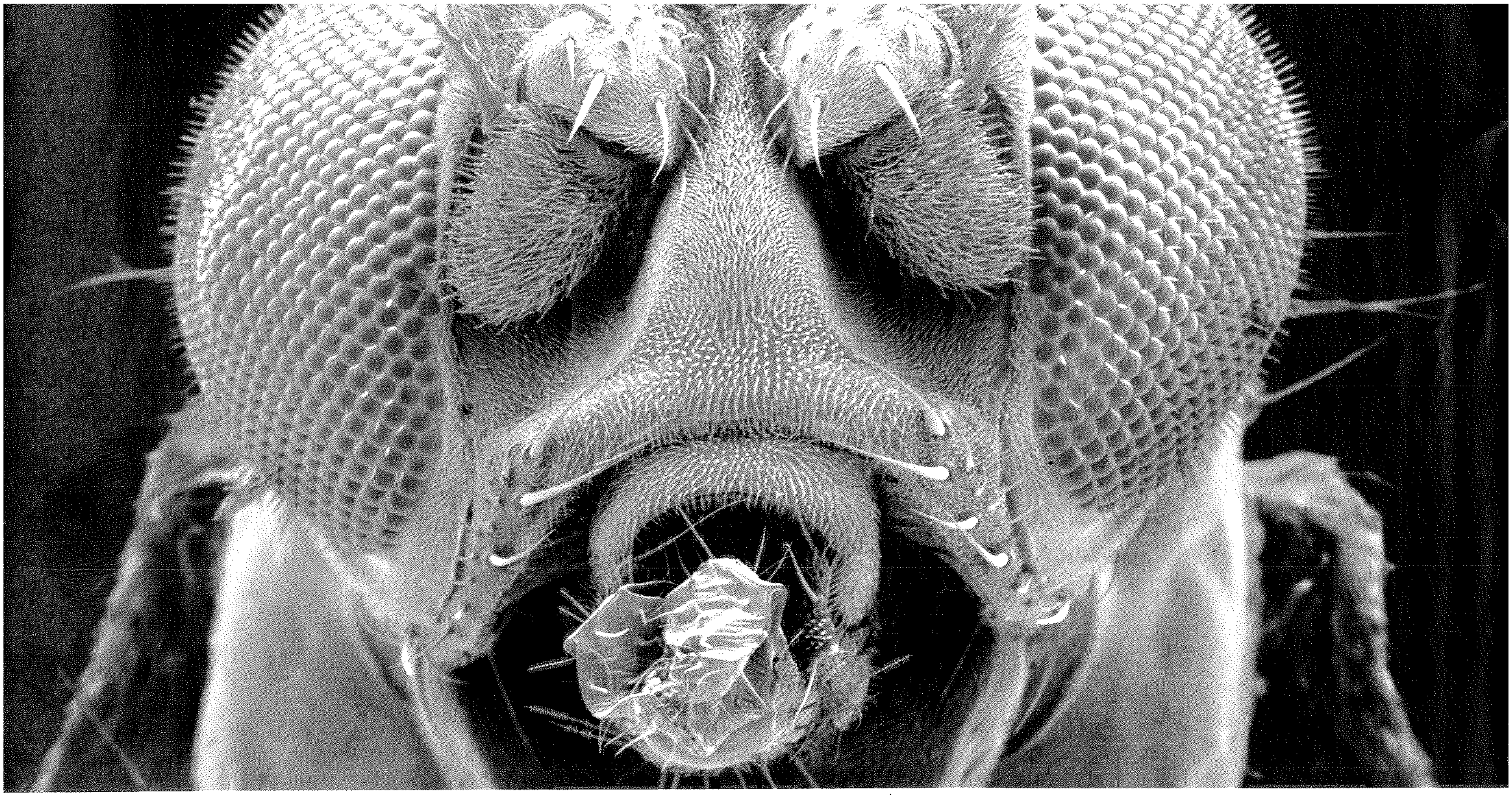

Figure 5.15 An electron microscope can produce sharp images of a very small object, such as this mite, because of the small wavelength of a moving electron compared with that of light.

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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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 3×10^{-34} m, which is much too small to detect experimentally. On the other hand, an electron has a mass of only 9.11×10^{-28} g. If it were moving at a velocity of 40 m/s, it would have a wavelength of 2×10^{-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.

Interactive Textbook

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

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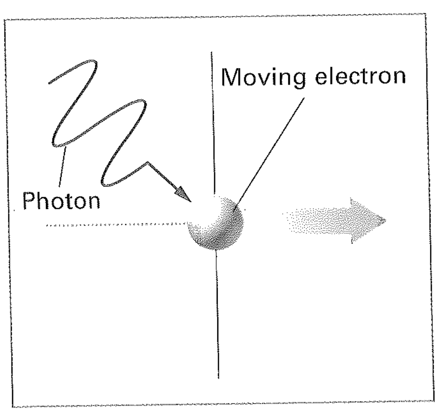
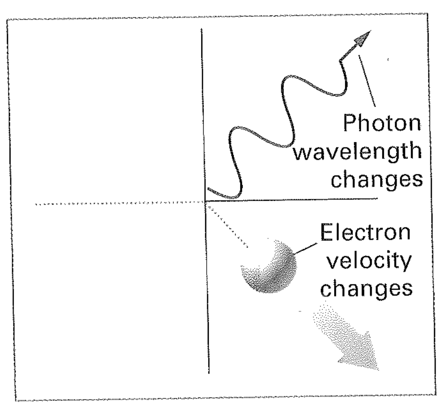


Before collision
A photon strikes an electron during an attempt to observe the electron's position.



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

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.

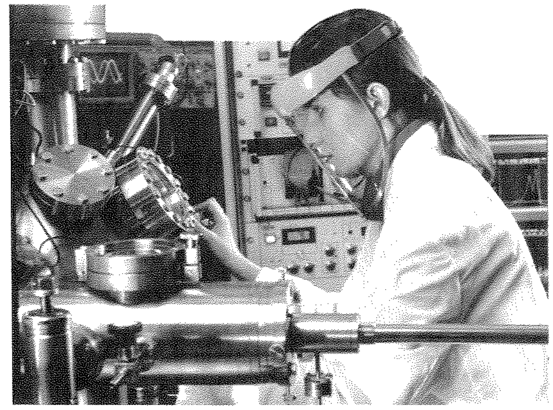
Careers *in* Chemistry

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

16. **Key Concept** How are wavelength and frequency of light related?
17. **Key Concept** Describe the cause of atomic emission spectrum of an element.
18. **Key Concept** How is the change in electron energy related to the frequency of light emitted in atomic transitions?
19. **Key Concept** How does quantum mechanics differ from classical mechanics?
20. 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?
21. Arrange the following in order of decreasing wavelength.
    a. infrared radiation from a heat lamp
    b. dental X-rays
    c. 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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