Chapter 28 The Atom

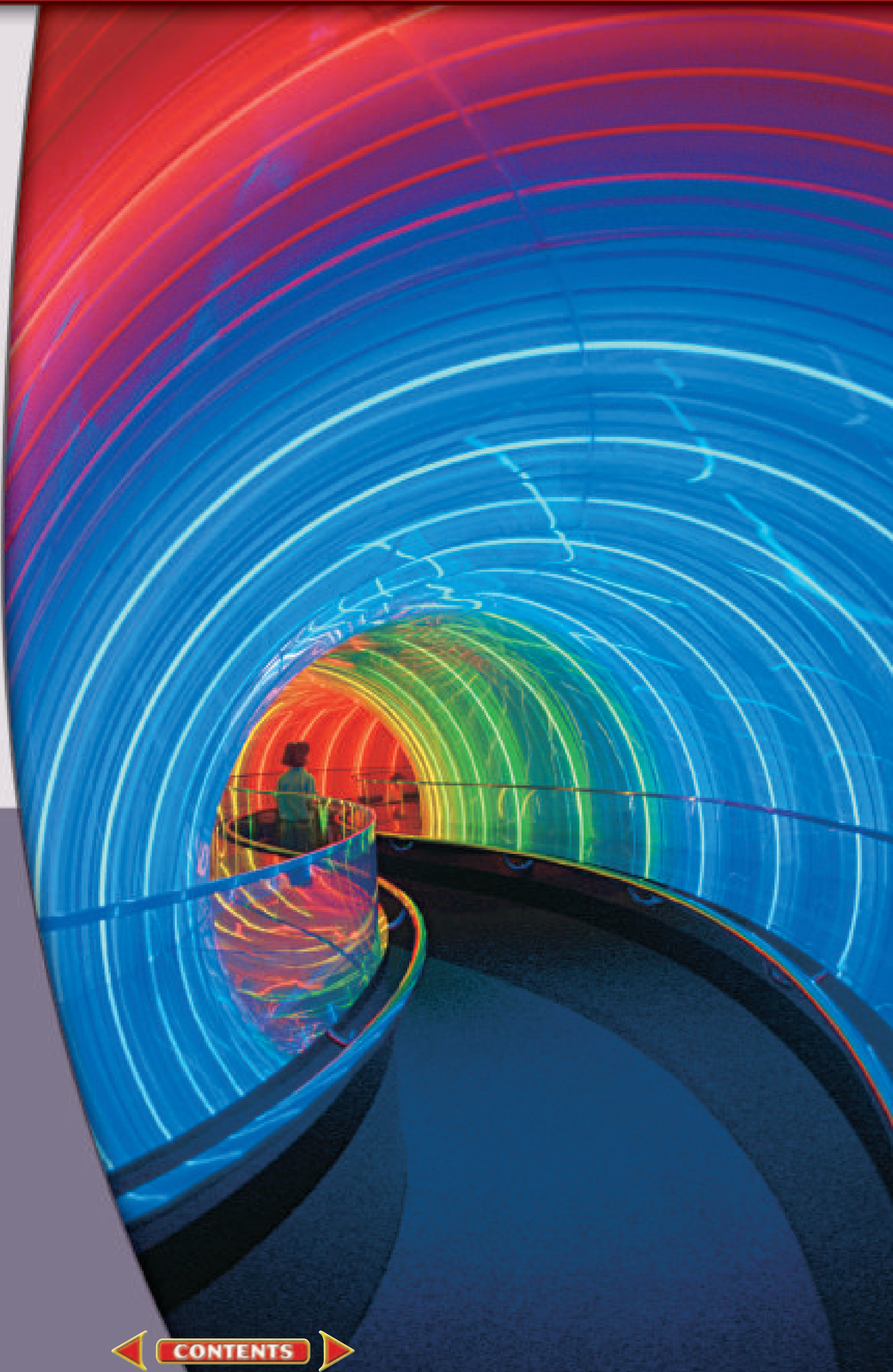
What You'll Learn

- You will learn about the discovery of the atom's composition.
- You will determine energies of the hydrogen atom.
- You will learn how quantum theory led to the modern atomic model.
- You will learn how lasers work and what their applications are.

Why It's Important

The quantum model of the atom and the transition of electrons between energy levels explain much of the observed behavior of all matter.

Emission Spectra Each of these gas-filled tubes emits a unique spectrum of colors. The bright light is emitted when electrons in the gas make transitions to lower energy states.



Think About This ▶

Why are the colors of the lights different, and how could you identify what gases are used in each tube?

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LAUNCH Lab

How can identifying different spinning coins model types of atoms?

Question

When a quarter, a nickel, a penny, and a dime are spun on a tabletop, what characteristics allow you to identify the type of spinning coin?

Procedure

1. Hold a quarter up on its edge and flick it with your index finger to set it spinning. Note the appearance and sound of the spinning coin until it comes to a stop on the table.
2. Repeat step 1 three more times using a dime, a nickel, and a penny, respectively.
3. Have a classmate spin the coins, one at a time, in a random order. Observe each coin only after it is already spinning and then try to identify the type of coin that it is.
4. Repeat step 3, except this time, keep your eyes closed while trying to identify each of the spinning coins.

Analysis

How successful were you at identifying the individual coins when you were limited to listening to the sounds they made? What are the characteristics of a spinning coin that can be used to identify its type? What instruments might make the identification of the spinning coins easier?

Critical Thinking

Excited atoms of an element in a high-voltage gas-discharge tube dissipate energy by emitting light. How might the emitted light help you identify the type of atom in the discharge tube? What instruments might help you do this?

28.1 The Bohr Model of the Atom

By the end of the nineteenth century, most scientists agreed on the existence of atoms. J. J. Thomson's discovery of the electron provided convincing evidence that the atom was made up of even smaller, subatomic particles. Every atom tested by Thomson contained negatively charged electrons, and these electrons possessed very little mass. Because atoms were known to be much more massive than the mass accounted for by the electrons they contain, scientists began looking for the missing mass that must be part of each atom. What was the nature of this yet-to-be-discovered massive part of the atom? How was this mass distributed within the atom?

Moreover, atoms were known to be electrically neutral, yet, only negatively charged electrons had been identified within the atom. How were the negatively charged electrons arranged in the atom? What was the source of the atom's neutrality? Were positively charged particles also present in the atom? Knowing that their understanding of the atom was far from complete, scientists began searching for answers to numerous and challenging questions.

Objectives

- **Describe** the structure of the nuclear atom.
- **Compare and contrast** continuous spectra and line-emission spectra.
- **Solve** problems using orbital-radius and energy-level equations.

Vocabulary

alpha particles
nucleus
absorption spectrum
energy level
ground state
excited state
principal quantum number

CONTENTS

The Nuclear Model

Many questions faced the researchers investigating the nature of the atom. What caused the emission of light from atoms? How were the electrons distributed in the atom? Physicists and chemists from many countries searched for the solutions to this puzzle. The results not only provided knowledge about the structure of the atom, but also a totally new approach to both physics and chemistry. The history of the research into the nature of the atom is one of the most exciting stories of the twentieth century.

J. J. Thomson believed that a massive, positively charged substance filled the atom. He pictured the negatively charged electrons as being distributed throughout this positively charged substance like raisins in a muffin. Ernest Rutherford, along with laboratory collaborators Hans Geiger and Ernest Marsden, however, performed a series of experiments that showed the atom had a very different structure.

Rutherford's experiments made use of radioactive compounds that emitted penetrating rays. Some of these emissions had been found to be massive, positively charged particles that moved at high speeds. These particles, which were later named **alpha particles,** are represented by the symbol α. The α-particles in Rutherford's experiments could be detected by the small flashes of light that were emitted when the particles collided with a zinc-sulfide-coated screen.

As shown in **Figure 28-1,** Rutherford directed a beam of α-particles at an extremely thin sheet of gold foil. Rutherford was aware of Thomson's model of the atom, and he expected only minor deflections of the α-particles as they passed through the thin gold foil. He thought that the paths of the massive, high-speed α-particles would be only slightly altered as they passed through the evenly distributed positive charge making up each gold atom. The test results amazed him. While most of the α-particles passed through the gold foil either undeflected or only slightly deflected, a few of the particles were scattered through very large angles. Some were even deflected through angles larger than 90°. A diagram of these results is shown in **Figure 28-2.** Rutherford compared his amazement to that of firing a 15-inch cannon shell at tissue paper and then having the shell bounce back and hit him.

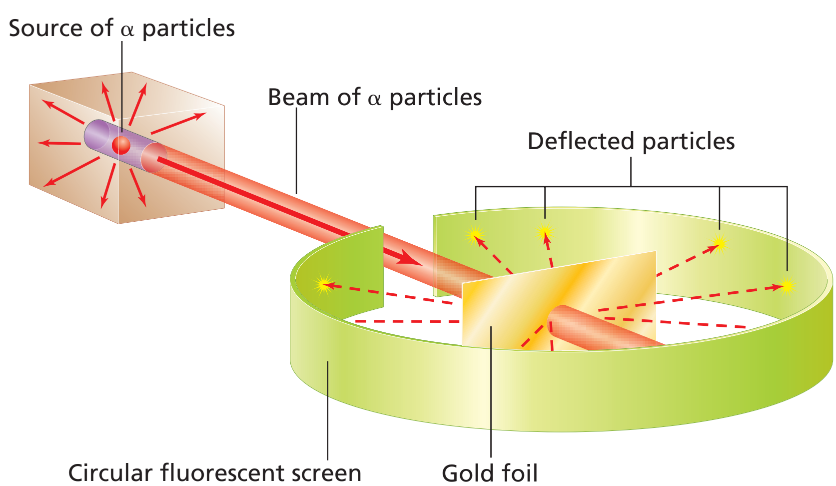
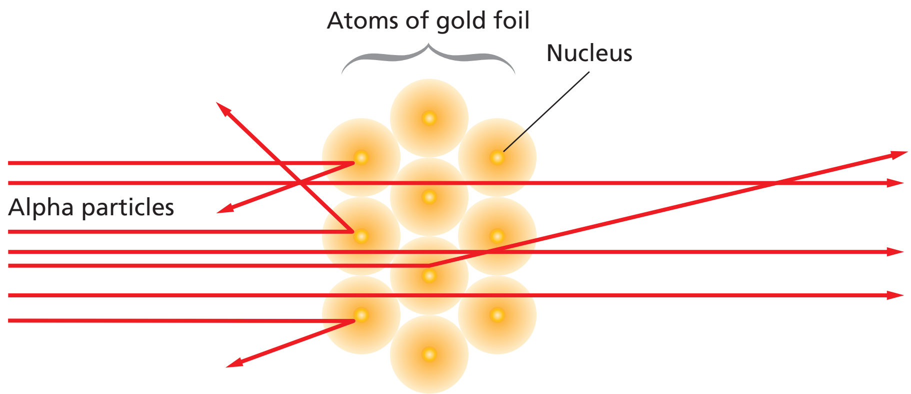


Figure 28-1 After bombarding metal foil with alpha particles, Rutherford's team concluded that most of the mass of the atom was concentrated in the nucleus.

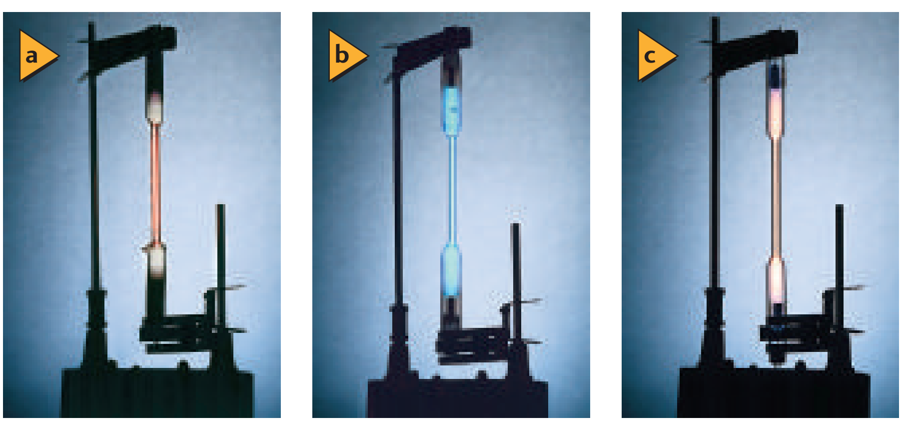


■ **Figure 28-2** Most of the alpha particles directed at the thin sheet of gold foil passed through it without deflection. One in 20,000, however, were deflected at large angles.

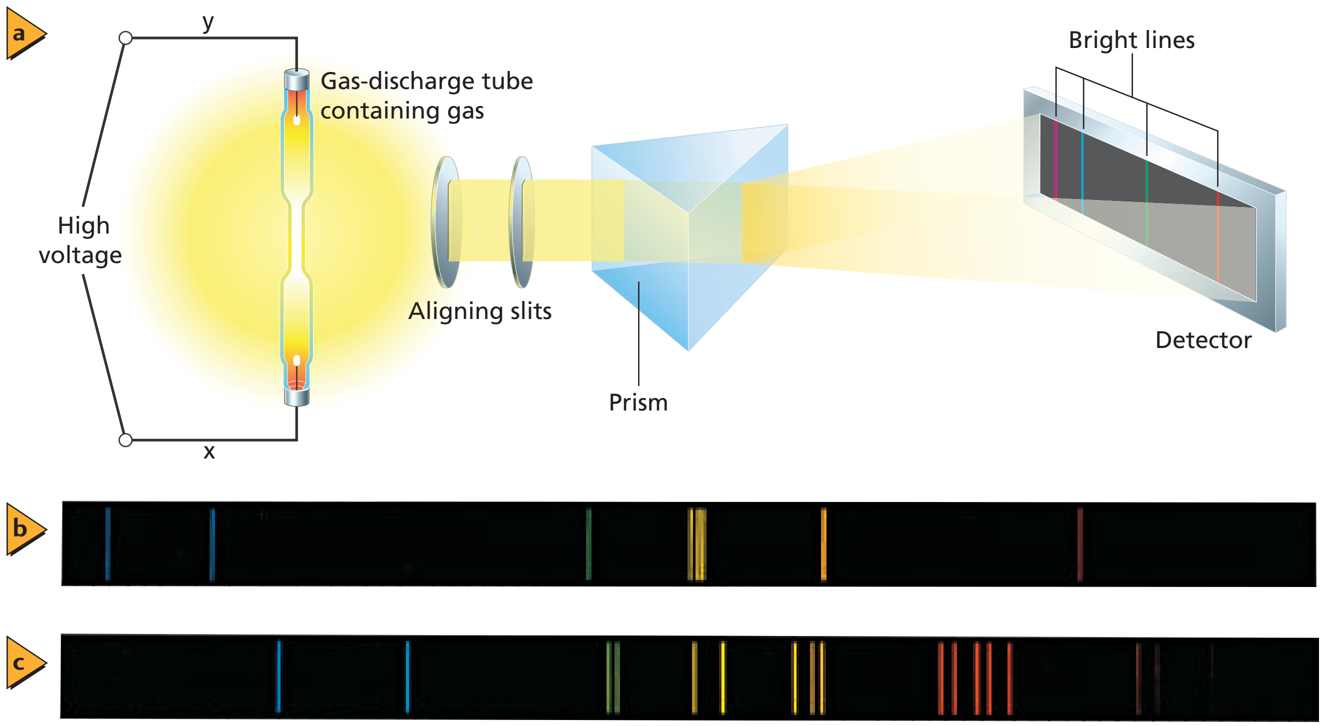
Using Coulomb's force law and Newton's laws of motion, Rutherford concluded that the results could be explained only if all of the atom's positive charge were concentrated in a tiny, massive central core, now called the **nucleus.** Therefore, Rutherford's model of the atom is called the nuclear model. Researchers have since determined that all the positive charge and more than 99.9 percent of the mass of the atom are contained in its nucleus. The electrons, which do not contribute a significant amount of mass to the atom, are distributed outside of and far away from the nucleus. Thus, the space occupied by the electrons defines the overall size, or diameter, of the atom. Because the diameter of the atom is about 10,000 times larger than the diameter of the nucleus, the atom mostly is made up of empty space.

Emission spectra How are the electrons arranged around the nucleus of the atom? One of the clues that scientists used to answer this question came from the study of the light emitted by atoms. Recall from the previous chapter that the set of electromagnetic wavelengths emitted by an atom is called the atom's emission spectrum.

As shown in **Figure 28-3,** atoms of a gaseous sample can be made to emit light in a gas-discharge tube. You probably are familiar with the colorful neon signs used by some businesses. These signs work on the same principles as gas-discharge tubes do. A gas-discharge tube consists of a low-pressure gas contained within a glass tube that has metal electrodes attached to each end. The gas glows when high voltage is applied across the tube. What interested scientists the most about this phenomenon was the fact that each different gas glowed with a different, unique color. The characteristic glows emitted by several gases are shown in Figure 28-3.



■ **Figure 28-3** When high voltage is applied to a gas, the gas emits light, producing a unique glow. Hydrogen gas glows magenta **(a),** mercury glows bright blue **(b),** and nitrogen glows rose-orange **(c).**



■ **Figure 28-4** A prism spectroscope can be used to observe emission spectra **(a).** The emission spectra of mercury **(b)** and barium **(c)** show characteristic lines.

When the light emitted by the gas is passed through a prism or a diffraction grating, the emission spectrum of an atom is obtained. An emission spectrum can be studied in greater detail using an instrument called a spectroscope. As shown in **Figure 28-4a,** light in a prism spectroscope passes through a slit and is then dispersed as it travels through a prism. A lens system (not shown in the diagram) focuses the dispersed light so that it can be viewed, or recorded on a photographic plate or an electronic detector. The spectroscope forms an image of the slit at a different position for each wavelength.

The spectrum of a hot body, or incandescent solid, such as the filament in a lightbulb, is a continuous band of colors from red through violet. The spectrum of a gas, however, is a series of distinct lines of different colors. The bright-line emission spectra for mercury gas and barium gas are shown in **Figure 28-4b** and **Figure 28-4c,** respectively. Each colored line corresponds to a particular wavelength of light emitted by the atoms of that gas.

An emission spectrum is also a useful analytic tool, as it can be used to identify an unknown sample of gas. When the unknown gas is placed in a gas-discharge tube, it can be made to emit light. The emitted light consists of wavelengths that are uniquely characteristic of the atoms of that gas. Thus, the unknown gas can be identified by comparing its wavelengths with the wavelengths present in the spectra of known samples.

An emission spectrum also can be used to analyze a mixture of gases. When the emission spectrum of a combination of elements is photographed, an analysis of the lines on the photograph can indicate the identities and the relative concentrations of the elements present. If the material being examined contains a large amount of any particular element, the lines for that element are more intense on the photograph than those of the other elements of lesser quantities. Through a comparison of the line intensities, the percentage composition of the material can be determined.

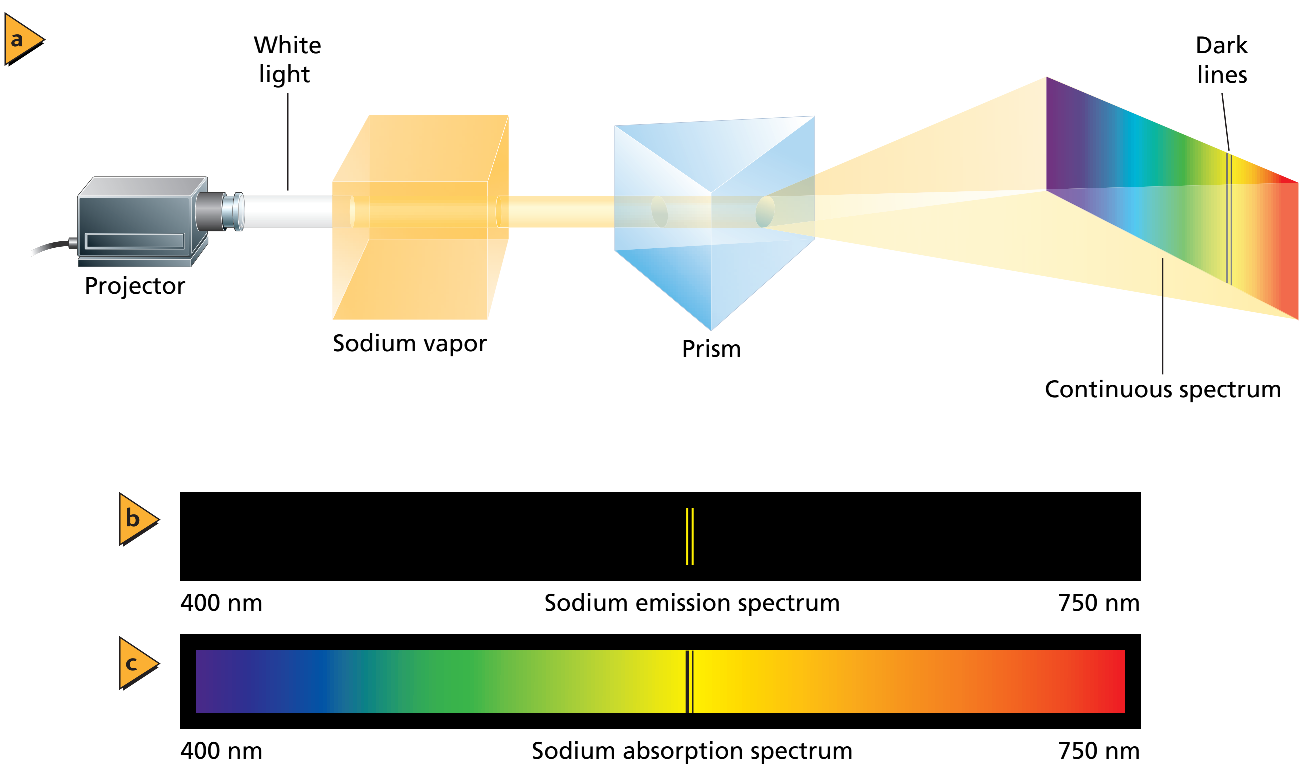
■ **Figure 28-5** Fraunhofer lines appear in the absorption spectrum of the Sun. There are many lines, some of which are faint and others that are very dark, depending on abundances of the elements in the Sun.

Absorption spectra In 1814, Josef von Fraunhofer observed the presence of several dark lines in the spectrum of sunlight. These dark lines, now called Fraunhofer lines, are seen in **Figure 28-5.** He reasoned that as sunlight passed through the gaseous atmosphere surrounding the Sun, the gases absorbed certain characteristic wavelengths. These absorbed wavelengths produced the dark lines in the observed spectrum. The set of wavelengths absorbed by a gas is the **absorption spectrum** of the gas. The composition of the Sun's atmosphere was determined by comparing the missing lines in the observed spectrum with the known emission spectra of various elements. The compositions of many other stars have been determined using this technique.

Astronomy Connection

You can observe an absorption spectrum by passing white light through a gas sample and a spectroscope, as shown in **Figure 28-6a.** Because the gas absorbs specific wavelengths, the normally continuous spectrum of the white light has dark lines in it after passing through the gas. For a gas, the bright lines of the emission spectrum and the dark lines of the absorption spectrum often occur at the same wavelengths, as shown in **Figure 28-6b** and **Figure 28-6c,** respectively. Thus, cool, gaseous elements absorb the same wavelengths that they emit when excited. As you might expect, the composition of a gas can be determined from the wavelengths of the dark lines of the gas's absorption spectrum.

■ **Figure 28-6** This apparatus is used to produce the absorption spectrum of sodium **(a).** The emission spectrum of sodium consists of several distinct lines **(b),** whereas the absorption spectrum of sodium is nearly continuous **(c).**



Images were created by Dr. Donald Mickey, University of Hawaii, Institute for Astronomy, from National Solar Observatory spectra data.

Spectroscopy Both emission and absorption spectra are valuable scientific tools. As a result of the elements' characteristic spectra, scientists are able to analyze, identify, and quantify unknown materials by observing the spectra that they emit or absorb. The emission and absorption spectra of elements are important in industry as well as in scientific research. For example, steel mills reprocess large quantities of scrap iron of varying compositions. The exact composition of a sample of scrap iron can be determined in minutes by spectrographic analysis. The composition of the steel can then be adjusted to suit commercial specifications. Aluminum, zinc, and other metal-processing plants employ the same method.

The study of spectra is a branch of science known as spectroscopy. Spectroscopists are employed throughout research and industrial communities. Spectroscopy has proven to be an effective tool for analyzing materials on Earth, and it is the only currently available tool for studying the composition of stars over the vast expanse of space.

The Bohr Model of the Atom

In the nineteenth century, many physicists tried to use atomic spectra to determine the structure of the atom. Hydrogen was studied extensively because it is the lightest element and has the simplest spectrum. The visible spectrum of hydrogen consists of four lines: red, green, blue, and violet, as shown in **Figure 28-7.** Any theory that explained the structure of the atom would have to account for these wavelengths and support the nuclear model. However, the nuclear model as proposed by Rutherford was not without its problems. Rutherford had suggested that electrons orbit the nucleus much like the planets orbit the Sun. There was, however, a serious flaw in this planetary model.

Problems with the planetary model An electron in an orbit constantly is accelerated toward the nucleus. As you learned in Chapter 26, accelerating electrons radiate energy by emitting electromagnetic waves. At the rate that an orbiting electron would lose energy, it should spiral into the nucleus within 10^{-9} s. This, however, must not be happening because atoms are known to be stable. Thus, the planetary model was not consistent with the laws of electromagnetism. In addition, the planetary model predicted that the accelerating electrons would radiate energy at all wavelengths. However, as you just learned, the light emitted by atoms is radiated only at specific wavelengths.

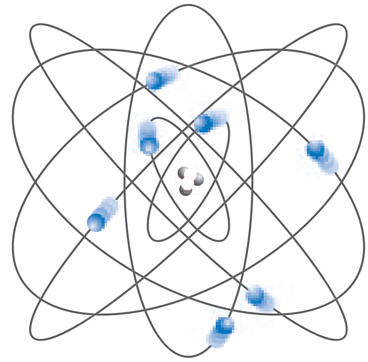
Danish physicist Niels Bohr went to England in 1911 and joined Rutherford's group to work on determining the structure of the atom. He tried to unite the nuclear model with Planck's quantized energy levels and Einstein's theory of light. This was a courageous idea because as of 1911, neither of these revolutionary ideas was widely understood or accepted.

Figure 28-7 The emission spectrum of hydrogen in the visible range has four lines.

Quantized Energy

Bohr began with the planetary arrangement of electrons, as diagrammed in **Figure 28-8,** but then made the bold hypothesis that the laws of electromagnetism do not apply inside the atom. He postulated that an electron in a stable orbit does not radiate energy, even though it is accelerating. Bohr referred to this stable condition as a stationary state. He went on to assume that only stationary states with specific amounts of energy are allowed. In other words, Bohr considered the energy levels in an atom to be quantized.

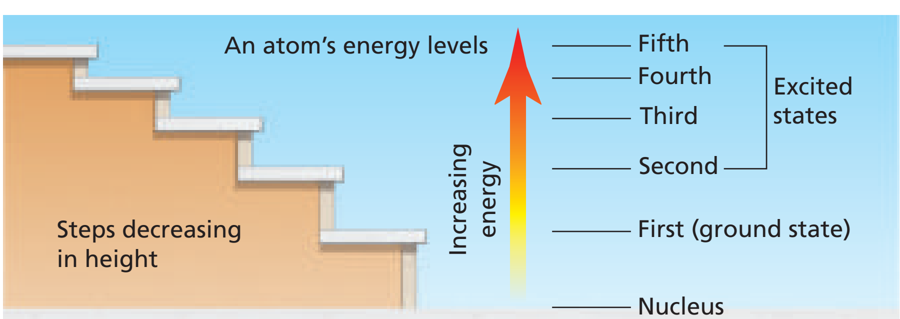
As shown in **Figure 28-9,** the quantization of energy in atoms can be likened to a flight of stairs with decreasing-height steps. To go up the stairs you must move from one step to the next—it is impossible to stop at a midpoint between steps. Instead of steps, atoms have quantized amounts of energy, each of which is called an **energy level.** Just as you cannot occupy a position between steps, an atom's energy cannot have a value between allowed energy levels. An atom with the smallest allowable amount of energy is said to be in the **ground state.** When an atom absorbs energy, it moves, or makes a transition to, a higher energy level. Any energy level above the ground state is called an **excited state.**



■ **Figure 28-8** Bohr's planetary model of the atom was based on the postulation that electrons move in fixed orbits around the nucleus.

Energy of an Atom What determines the amount of energy an atom has? An atom's energy equals the sum of the kinetic energy of the electrons and the potential energy from the attractive force between the electrons and the nucleus. The energy of an atom with electrons in a nearby orbit is less than that of an atom with electrons in a faraway orbit because work must be done to move the electrons away from the nucleus. Thus, atoms in excited, higher-energy states have electrons in larger, or more distant, orbits. Because energy is quantized and energy is related to the size of the orbit, the size of the orbit also is quantized. The model of an atom just described, that of a central nucleus with orbiting electrons having specific quantized energy levels, is known as the Bohr model of the atom.

If Bohr was correct in hypothesizing that stable atoms do not radiate energy, then what is responsible for an atom's characteristic emission spectrum? To answer this question, Bohr suggested that electromagnetic energy is emitted when the atom changes from one stationary state to another. Incorporating Einstein's photoelectric theory, Bohr knew that the energy of every photon is given by the equation, $E_{\text{photon}} = hf$. He then postulated that when an atom absorbs a photon, the atom's energy increases by an amount equal to that of the photon. This excited atom then makes a transition to a lower energy level by emitting a photon.



■ **Figure 28-9** These decreasing-height steps are analogous to the allowed energy levels in an atom. Note how the difference in energy between adjacent energy levels decreases as the energy level increases.

$E_{photon\ 1} = E_3 - E_1$
$E_{photon\ 2} = E_2 - E_1$
$E_{photon\ 3} = E_3 - E_2$

$E_{photon\ 1} > E_{photon\ 2} > E_{photon\ 3}$
$\lambda_1 < \lambda_2 < \lambda_3$

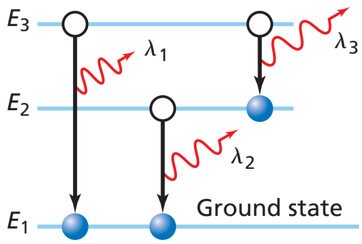


Figure 28-10 The energy of the emitted photon is equal to the difference in energy between the initial and final energy levels of the atom.

When the atom makes the transition from its initial energy level, E_i, to its final energy level, E_f, the change in energy, ΔE_{atom}, is given by the following equation.

$$\Delta E_{atom} = E_f - E_i$$

As shown in **Figure 28-10,** the change in energy of the atom equals the energy of the emitted photon.

$$E_{photon} = \Delta E_{atom}$$
or
$$E_{photon} = E_f - E_i$$

The following equations summarize the relationships between the change in energy states of an atom and the energy of the photon emitted.

Energy of an Emitted Photon $E_{photon} = hf$, or $E_{photon} = \Delta E_{atom}$

The energy of an emitted photon is equal to the product of Planck's constant and the emitted photon's frequency. The energy of an emitted photon also is equal to the loss in the atom's energy.

Predictions of the Bohr Model

A scientific theory must do more than present postulates; it must allow predictions to be made that can be checked against experimental data. A good theory also can be applied to many different problems, and it ultimately provides a simple, unified explanation of some part of the physical world.

Bohr used his theory to calculate the wavelengths of light emitted by a hydrogen atom. The calculations were in excellent agreement with the values measured by other scientists. As a result, Bohr's model was widely accepted. Unfortunately, the model only worked for the element hydrogen; it could not predict the spectrum of helium, the next-simplest element. In addition, there was not a good explanation as to why the laws of electromagnetism should work everywhere but inside the atom. Not even Bohr believed that his model was a complete theory of the structure of the atom. Despite its shortcomings, however, the Bohr model describes the energy levels and wavelengths of light emitted and absorbed by hydrogen atoms remarkably well.

Development of Bohr's model Bohr developed his model by applying Newton's second law of motion, $F_{net} = ma$, to the electron. The net force is described by Coulomb's law for the interaction between an electron of charge $-q$ that is a distance r from a proton of charge $+q$. That force is given by $F = -Kq^2/r^2$. The acceleration of the electron in a circular orbit about a much more massive proton is given by $a = -v^2/r$, where the negative sign shows that the direction is inward. Thus, Bohr obtained the following relationship:

$$\frac{Kq^2}{r^2} = \frac{mv^2}{r}$$

In the equation, K is the constant from Coulomb's law and has a value of $9.0\times10^9\ N\cdot m^2/C^2$.

Next, Bohr considered the angular momentum of the orbiting electron, which is equal to the product of an electron's momentum and the radius of its circular orbit. The angular momentum of the electron is thus given by mvr. Bohr postulated that angular momentum also is quantized; that is, the angular momentum of an electron can have only certain values. He claimed that the allowed values were multiples of $h/2\pi$, where h is Planck's constant. Using n to represent an integer, Bohr proposed that $mvr = nh/2\pi$. Using $Kq^2/r^2 = mv^2/r$ and rearranging the angular momentum equation, $v = nh/2\pi mr$, Bohr found that the orbital radii of the electrons in a hydrogen atom are given by the following equation.

Electron Orbital Radius in Hydrogen $r_n = \dfrac{h^2n^2}{4\pi^2Kmq^2}$

The radius of an electron in orbit n is equal to the product of the square of Planck's constant and the square of the integer n divided by the quantity four times the square of π, times the constant K, times the mass of an electron, times the square of the charge of an electron.

You can calculate the radius of the innermost orbit of a hydrogen atom, also known as the Bohr radius, by substituting known values and $n = 1$ into the above equation.

$$r_1 = \frac{(6.626\times10^{-34}\ \mathrm{J{\cdot}s})^2(1)^2}{4\pi^2(9.0\times10^9\ \mathrm{N{\cdot}m^2/C^2})(9.11\times10^{-31}\ \mathrm{kg})(1.60\times10^{-19}\ \mathrm{C})^2}$$

$$= 5.3\times10^{-11}\ \mathrm{J^2{\cdot}s^2/N{\cdot}m^2{\cdot}kg}$$

$$= 5.3\times10^{-11}\ \mathrm{m,\ or\ 0.053\ nm}$$

By performing a little more algebra you can show that the total energy of the atom, which is the sum of the kinetic energy of the electron and the potential energy, and is given by $-Kq^2/2r$, is represented by the following equation:

$$E_n = \frac{-2\pi K^2mq^4}{h^2} \times \frac{1}{n^2}$$

By substituting numerical values for the constants, you can calculate the total energy of the atom in joules, which yields the following equation:

$$E_n = -2.17\times10^{-18}\ \mathrm{J} \times \frac{1}{n^2}$$

Converting the relationship to units of electron volts yields the following equation.

Energy of a Hydrogen Atom $E_n = -13.6\ \mathrm{eV} \times \dfrac{1}{n^2}$

The total energy of an atom with principal quantum number n is equal to the product of -13.6 eV and the inverse of n^2.

Both the electron's orbital radius and the energy of the atom are quantized. The integer, n, that appears in these equations is called the **principal quantum number.** It is the principal quantum number that determines the quantized values of r and E. In summary, the radius, r, increases as the square of n, whereas the energy, E, depends on $1/n^2$.

MINI LAB

Bright-Line Spectra

Turn on a gas-discharge tube power supply attached to a gas tube so that the tube glows. ***CAUTION: Handle gas tube carefully to avoid breaking. Do not touch any exposed metal when the power supply is turned on. Dangerous voltages are present. Always turn off the power supply before changing gas tubes.***

Turn off the room lights.

1. **Describe** the color that you observe.
2. **Observe** the gas-discharge tube through a diffraction grating.
3. **Sketch** the results of viewing the gas-discharge tube through the diffraction grating.
4. **Predict** whether the observed spectrum will change when a different gas-discharge tube is viewed through the diffraction grating.
5. **Test** your prediction.

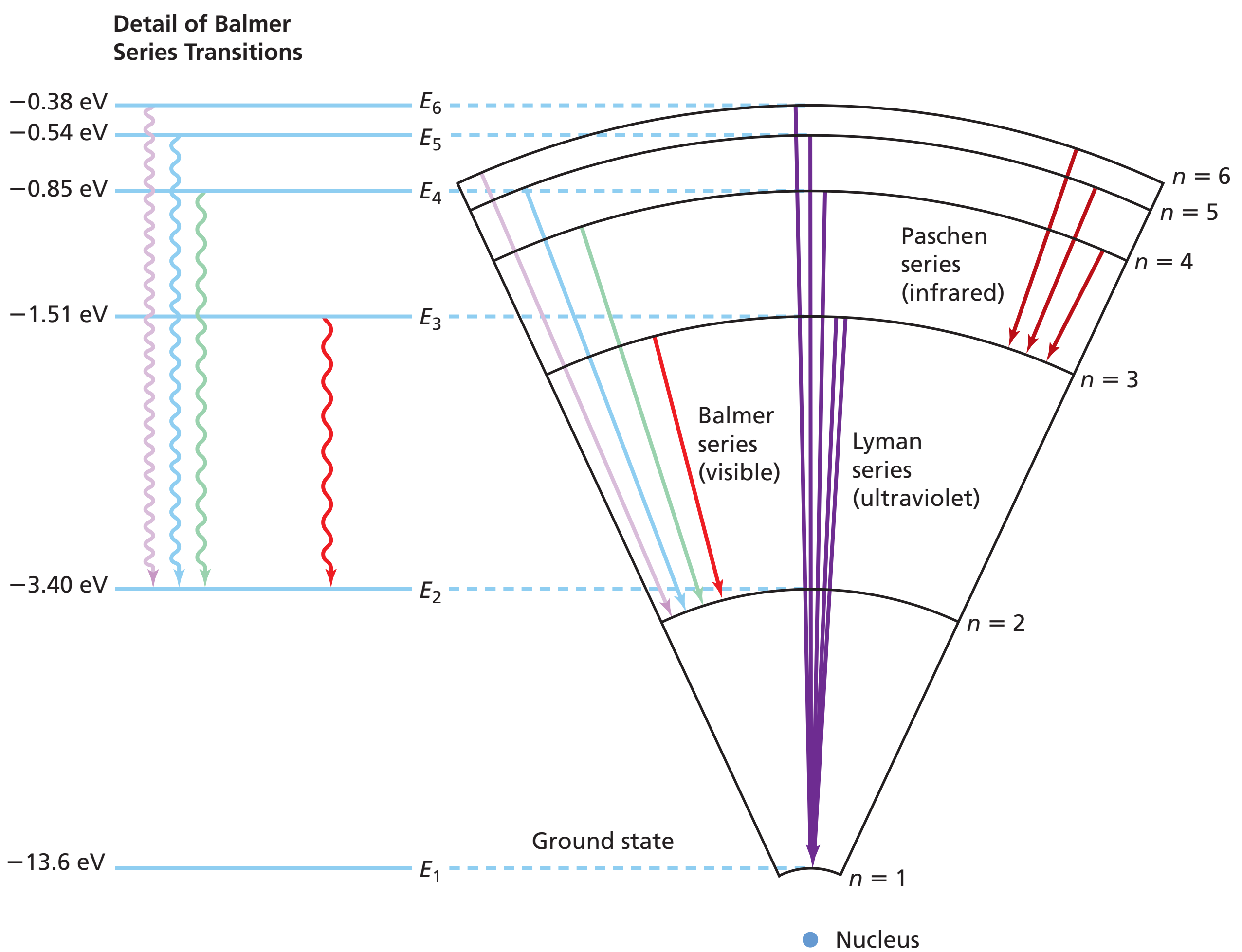
Analyze and Conclude

6. **Sketch** the results of viewing the new gas-discharge tube through the diffraction grating.
7. **Explain** why there are differences in the two spectra.

Energy and electron transitions You may be wondering why the energy of an atom in the Bohr model has a negative value. Recall from Chapter 11 that only energy differences have meaning. The zero energy level can be chosen at will. In this case, zero energy is defined as the energy of the atom when the electron is infinitely far from the nucleus and has no kinetic energy. This condition exists when the atom has been ionized; that is, when the electron has been removed from the atom. Because work has to be done to ionize an atom, the energy of an atom with an orbiting electron is less than zero. The energy of an atom has a negative value. When an atom makes a transition from a lower to a higher energy level, the total energy becomes less negative, but the overall energy change is positive.

Some of hydrogen's energy levels and the possible energy level transitions that it can undergo are shown in **Figure 28-11.** Note that an excited hydrogen atom can emit electromagnetic energy in the infrared, visible, or ultraviolet range depending on the transition that occurs. Ultraviolet light is emitted when the atom drops into its ground state from any excited state. The four visible lines in the hydrogen spectrum are produced when the atom drops from the $n = 3$ or higher energy state into the $n = 2$ energy state.

Figure 28-11 The distinct set of color lines that make up a hydrogen atom's visible spectrum are known as the Balmer series. This visible light is the result of the photons emitted when electrons make transitions to the second energy level, $n = 2$. Other electron transitions in a hydrogen atom result in the emission of ultraviolet (Lyman series) and infrared (Paschen series) electromagnetic energy.



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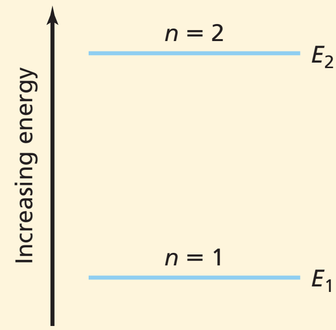
EXAMPLE Problem 1

Energy Levels **A hydrogen atom absorbs energy, causing its electron to move from the innermost energy level ($n = 1$), to the second energy level ($n = 2$). Determine the energy of the first and second energy levels, and the energy absorbed by the atom.**

1 Analyze and Sketch the Problem

- Diagram the energy levels E_1 and E_2.
- Indicate the direction of increasing energy on the diagram.



Known:	**Unknown:**
Quantum number of innermost energy level, $n = 1$	Energy of level $E_1 = ?$
Quantum number of second energy level, $n = 2$	Energy of level $E_2 = ?$
	Energy difference, $\Delta E = ?$

2 Solve for the Unknown

Use the equation for the energy of an electron in its orbit to calculate the energy of each level.

$$E_n = -13.6 \text{ eV} \times \frac{1}{n^2}$$

$$E_1 = -13.6 \text{ eV} \times \frac{1}{(1)^2} \qquad \textbf{Substitute } n = 1$$

$$= -13.6 \text{ eV}$$

$$E_2 = -13.6 \text{ eV} \times \frac{1}{(2)^2} \qquad \textbf{Substitute } n = 2$$

$$= -3.40 \text{ eV}$$

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The energy absorbed by the atom, ΔE, is equal to the energy difference between the final energy level of the atom, E_f, and the initial energy level of the atom, E_i.

$$\Delta E = E_f - E_i$$

$$= E_2 - E_1 \qquad \textbf{Substitute } E_f = E_2, E_i = E_1$$

$$= -3.40 \text{ eV} - (-13.6 \text{ eV}) \qquad \textbf{Substitute } E_2 = -3.40 \text{ eV}, E_1 = -13.6 \text{ eV}$$

$$= 10.2 \text{ eV} \quad \text{Energy is absorbed.}$$

3 Evaluate the Answer

- **Are the units correct?** Orbital energy values should be measured in electron volts.
- **Is the sign correct?** The energy difference is positive when electrons move from lower energy levels to higher energy levels.
- **Is the magnitude realistic?** The energy needed to move an electron from the first energy level to the second energy level should be approximately 10 eV, which it is.

PRACTICE Problems

Additional Problems, Appendix B

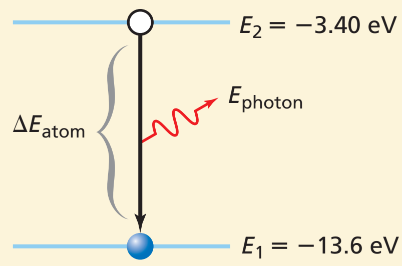
1. Calculate the energies of the second, third, and fourth energy levels in the hydrogen atom.
2. Calculate the energy difference between E_3 and E_2 in the hydrogen atom.
3. Calculate the energy difference between E_4 and E_2 in the hydrogen atom.
4. The text shows the solution of the equation $r_n = \frac{h^2n^2}{4\pi^2Kmq^2}$ for $n = 1$, the innermost orbital radius of the hydrogen atom. Note that with the exception of n^2, all factors in the equation are constants. The value of r_1 is 5.3×10^{-11} m, or 0.053 nm. Use this information to calculate the radii of the second, third, and fourth allowable energy levels in the hydrogen atom.
5. The diameter of the hydrogen nucleus is 2.5×10^{-15} m, and the distance between the nucleus and the first electron is about 5×10^{-11} m. If you use a ball with a diameter of 7.5 cm to represent the nucleus, how far away will the electron be?

▶ EXAMPLE Problem 2

Frequency and Wavelength of Emitted Photons An excited hydrogen atom drops from the second energy level ($n = 2$) to the first energy level ($n = 1$). Calculate the energy and the wavelength of the emitted photon. Use E_1 and E_2 values from Example Problem 1.

1 Analyze and Sketch the Problem

- Diagram the energy levels E_1 and E_2.
- Indicate the direction of increasing energy and show the emission of a photon on the diagram.



Known:	Unknown:
Energy of level $E_1 = -13.6$ eV	Frequency, $f = ?$
Energy of level $E_2 = -3.40$ eV	Wavelength, $\lambda = ?$
	Energy difference, $\Delta E = ?$

2 Solve for the Unknown

The energy of the emitted photon is equal to ΔE, the energy difference between the final energy level of the atom, E_f, and the initial energy level of the atom, E_i.

$$\Delta E = E_f - E_i$$
$$= E_1 - E_2 \quad \text{Substitute } E_f = E_1, E_i = E_2$$
$$= -13.6 \text{ eV} - (-3.40 \text{ eV}) \quad \text{Substitute } E_1 = -13.6 \text{ eV}, E_2 = -3.40 \text{ eV}$$
$$= -10.2 \text{ eV} \quad \text{Energy is emitted.}$$

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Isolating a Variable
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To determine the wavelength of the photon, use the following equations.

$$\Delta E = hf, \text{ so } f = \frac{\Delta E}{h} \quad \text{Solve the photon energy equation for frequency.}$$
$$c = \lambda f, \text{ so } \lambda = \frac{c}{f} \quad \text{Solve the wavelength-frequency equation for wavelength.}$$
$$\lambda = \frac{c}{(\Delta E/h)} \quad \text{Substitute } f = \frac{\Delta E}{h}$$
$$= \frac{hc}{\Delta E}$$
$$= \frac{1240 \text{ eV}\cdot\text{nm}}{10.2 \text{ eV}} \quad \text{Substitute } hc = 1240 \text{ eV}\cdot\text{nm}, \Delta E = 10.2 \text{ eV}$$
$$= 122 \text{ nm}$$

3 Evaluate the Answer

- **Are the units correct?** Energy is measured in electron volts. The prefix *nano*- modifies the base SI unit, the meter, which is the correct unit for the wavelength.
- **Are the signs correct?** Energy is released when the atom emits a photon during the transition from the second energy level to the first; thus, the energy difference is negative.
- **Is the magnitude realistic?** Energy released in this transition produces light in the ultraviolet region below 400 nm.

▶ PRACTICE Problems

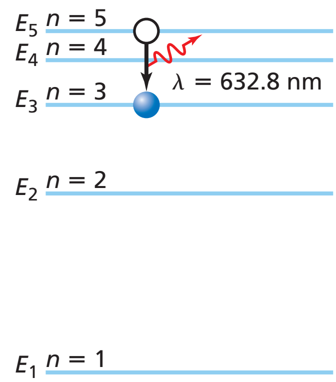
Additional Problems, Appendix B

6. Find the wavelength of the light emitted in Practice Problems 2 and 3. Which lines in Figure 28-7 correspond to each transition?
7. For a particular transition, the energy of a mercury atom drops from 8.82 eV to 6.67 eV.
   - **a.** What is the energy of the photon emitted by the mercury atom?
   - **b.** What is the wavelength of the photon emitted by the mercury atom?
8. The ground state of a helium ion is −54.4 eV. A transition to the ground state emits a 304-nm photon. What was the energy of the excited state?

CHALLENGE PROBLEM

Although the Bohr atomic model accurately explained the behavior of a hydrogen atom, it was unable to explain the behavior of any other atom. Verify the limitations of the Bohr model by analyzing an electron transition in a neon atom. Unlike a hydrogen atom, a neon atom has ten electrons. One of these electrons makes a transition between the $n = 5$ and the $n = 3$ energy states, emitting a photon in the process.

1. Assuming that the neon atom's electron can be treated as an electron in a hydrogen atom, what photon energy does the Bohr model predict?
2. Assuming that the neon atom's electron can be treated as an electron in a hydrogen atom, what photon wavelength does the Bohr model predict?
3. The actual wavelength of the photon emitted during the transition is 632.8 nm. What is the percent error of the Bohr model's prediction of photon wavelength?



The Bohr model was a major contribution to scientists' understanding of the structure of the atom. In addition to calculating the emission spectrum, Bohr and his students were able to calculate the ionization energy of a hydrogen atom. The ionization energy of an atom is the energy needed to completely free an electron from an atom. The calculated ionization value closely agreed with experimental data. The Bohr model further provided an explanation of some of the chemical properties of the elements. The idea that atoms have electron arrangements unique to each element is the foundation of much of our knowledge of chemical reactions and bonding. Niels Bohr, whose accomplishments are commemorated in the stamps shown in **Figure 28-12,** was awarded a Nobel prize in 1922.

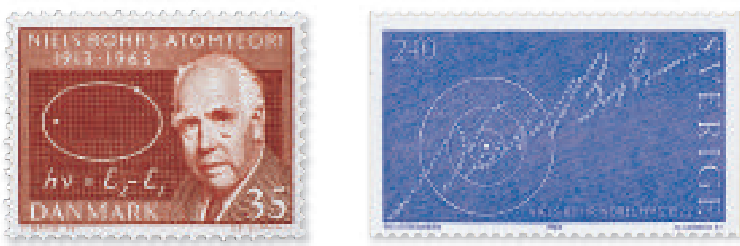


Figure 28-12 Honored in these postage stamps from Denmark and Sweden, Niels Bohr's great contribution to our understanding of the atom earned him worldwide recognition and a Nobel prize.

28.1 Section Review

9. **Rutherford's Nuclear Model** Summarize the structure of the atom according to Rutherford's nuclear model.

10. **Spectra** How do the emission spectra of incandescent solids and atomic gases differ? In what ways are they similar?

11. **Bohr Model** Explain how energy is conserved when an atom absorbs a photon of light.

12. **Orbit Radius** A helium ion behaves like a hydrogen atom. The radius of the ion's lowest energy level is 0.0265 nm. According to Bohr's model, what is the radius of the second energy level?

13. **Absorption Spectrum** Explain how the absorption spectrum of a gas can be determined. Describe the reasons for the spectrum's appearance.

14. **Bohr Model** Hydrogen has been detected transitioning from the 101st to the 100th energy levels. What is the radiation's wavelength? Where in the electromagnetic spectrum is this emission?

15. **Critical Thinking** The nucleus of a hydrogen atom has a radius of about 1.5×10^{-15} m. If you were to build a model of the hydrogen atom using a softball ($r = 5$ cm) to represent the nucleus, where would you locate an electron in the $n = 1$ Bohr orbit? Would it be in your classroom? On school property?

28.2 The Quantum Model of the Atom

▶ **Objectives**

- **Describe** the shortcomings of Bohr's atomic model.
- **Describe** the quantum model of the atom.
- **Explain** how a laser works.
- **Describe** the properties of laser light.

▶ **Vocabulary**

quantum model
electron cloud
quantum mechanics
coherent light
incoherent light
stimulated emission
laser

The postulates that Bohr made could not be explained on the basis of accepted physics principles of the time. For example, electromagnetic theory required that the accelerated particles radiate energy, causing the rapid collapse of the atom. In addition, the idea that electron orbits have well-defined radii was in conflict with the Heisenberg uncertainty principle. How could Bohr's work be put on a firm foundation?

From Orbits to an Electron Cloud

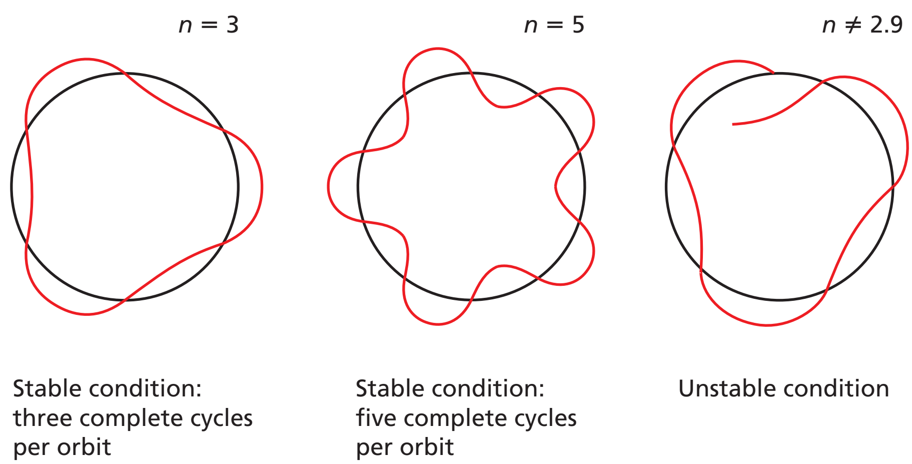
The first hint to the solution of these problems was provided by Louis de Broglie. Recall from Chapter 27 that de Broglie proposed that particles have wave properties, just as light has particle properties. The de Broglie wavelength of a particle with momentum mv is defined as $\lambda = h/mv$. The angular momentum of a particle can be defined as $mvr = hr/\lambda$. Thus, Bohr's required quantized angular-momentum condition, $mvr = nh/2\pi$, can be written in the following way:

$$\frac{hr}{\lambda} = \frac{nh}{2\pi} \quad \text{or} \quad n\lambda = 2\pi r$$

That is, the circumference of the Bohr orbit, $2\pi r$, is equal to a whole number multiple, n, of de Broglie wavelengths, λ. **Figure 28-13** illustrates this relationship.

In 1926, Austrian physicist Erwin Schröedinger used de Broglie's wave model to create a quantum theory of the atom based on waves. This theory did not propose a simple planetary model of an atom, as Bohr's model had. In particular, the radius of the electron orbit was not likened to the radius of the orbit of a planet about the Sun.

The Heisenberg uncertainty principle states that it is impossible to know both the position and momentum of an electron at the same time. Thus, the modern **quantum model** of the atom predicts only the probability that an electron is in a specific region. Interestingly, the quantum model predicts that the most probable distance between the electron and the nucleus for a hydrogen atom is the same as the radius predicted by the Bohr model.



■ **Figure 28-13** For an electron to have a stable orbit around the nucleus, the circumference of the orbit must be a whole-number multiple, n, of the de Broglie wavelength. Note that the whole-number multiples $n = 3$ and $n = 5$ are stable, whereas $n = 2.9$ is not.

The probability of the electron being at any specific radius can be calculated, and a three-dimensional plot can be constructed showing regions of equal probability. The region in which there is a high probability of finding the electron is called the **electron cloud.** A slice through the electron cloud for the two lowest energy states of a hydrogen atom is shown in **Figure 28-14.**

Even though the quantum model of the atom is difficult to visualize, **quantum mechanics,** the study of the properties of matter using its wave properties, makes use of this model and has been extremely successful in predicting many details of atomic structure. These details are very difficult to calculate precisely for all but the simplest atoms. Only very sophisticated computers can make highly accurate approximations for the heavier atoms. Quantum mechanics also enables the structures of many molecules to be calculated, allowing chemists to determine the arrangement of atoms in the molecules. Guided by quantum mechanics, chemists have been able to create new and useful molecules that do not occur in nature. Quantum mechanics also is used to analyze the details of the emission and absorption of light by atoms. As a result of quantum mechanical theory, a new source of light was developed.

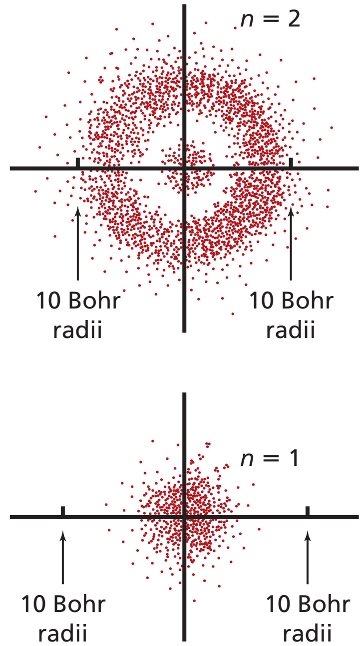


Figure 28-14 These plots show the probability of finding the electron in a hydrogen atom for distances up to approximately 10 Bohr radii from the nucleus for the first and second energy levels. The density of the distribution of dots corresponds to the probability of finding the electron. Note that the Bohr radius = 0.053 nm.

Lasers

As you know, light emitted by an incandescent source has a continuum of wavelengths, whereas light produced by an atomic gas consists of only a few distinct wavelengths. Light from both sources travels in all directions. Furthermore, light waves emitted by atoms at one end of a gas-discharge tube are not necessarily in step, or synchronized, with waves from the other end of the tube. That is, the waves are not necessarily all at the same point in their cycle at the same time. Recall from Chapter 19 that waves that are in step, with their minima and maxima coinciding, are said to be coherent. Light waves that are coherent are referred to as **coherent light.** Out-of-step light waves produce **incoherent light.** Both types of light waves are shown in **Figure 28-15.**

Light is emitted by atoms that have been excited. So far, you have learned about two ways in which atoms can be excited: thermal excitation and electron collision. Atoms also can be excited by collisions with photons of exactly the right energy.

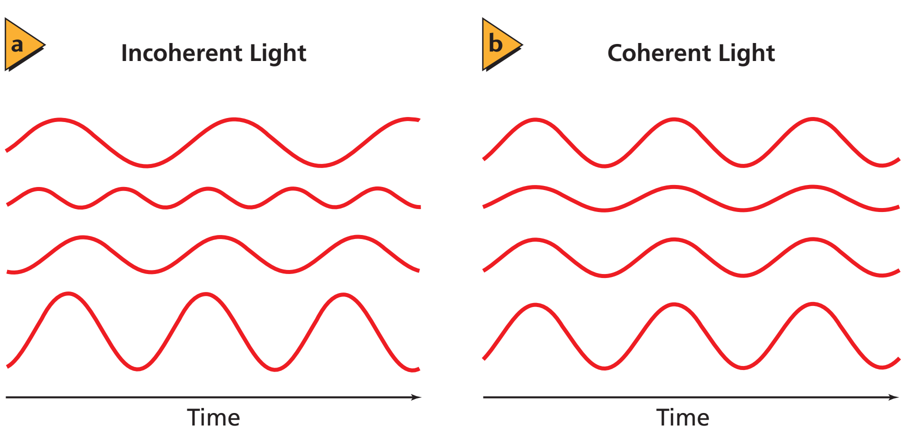


Figure 28-15 Waves of incoherent light **(a)** and coherent light **(b)** are shown.

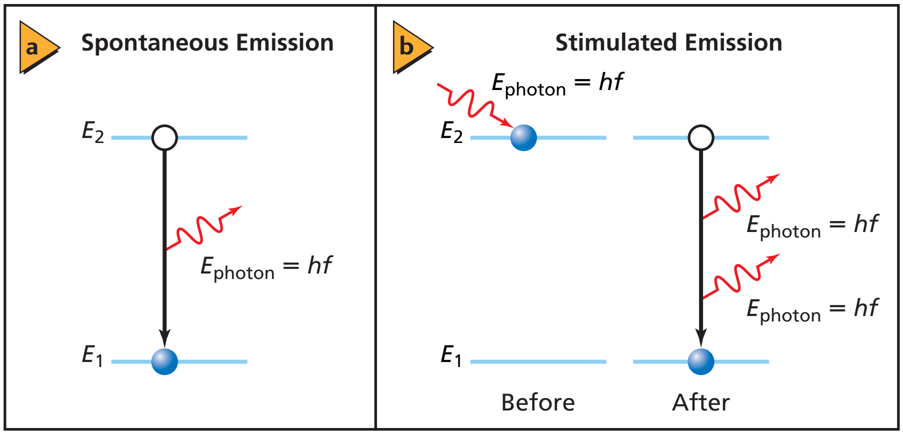


Figure 28-16 During spontaneous emission, an electron in an atom drops from the excited state, E_2, to the ground state, E_1, by spontaneously emitting a photon with energy, hf **(a).** During stimulated emission, an excited atom is struck by a photon with energy $E_2 - E_1$. The atom drops to the ground state and emits a photon. Both the incident and emitted photon have the same energy: $E_{photon} = E_2 - E_1$ **(b).**

Spontaneous and stimulated emission What happens when an atom is in an excited state? After a very short time, it normally returns to the ground state, giving off a photon of the same energy that it absorbed, as shown in **Figure 28-16a.** This process is called spontaneous emission.

In 1917, Einstein considered what happens to an atom already in an excited state that is struck by a photon with an energy equal to the energy difference between the excited state and the ground state. He showed that the atom, by a process known as **stimulated emission,** returns to the ground state and emits a photon with an energy equal to the energy difference between the two states. The photon that caused, or stimulated, the emission is not affected. The two photons leaving the atom not only will have the same frequency, but they will be in step, or coherent, as shown in **Figure 28-16b.** Either of the two photons can now strike other excited atoms, thereby producing additional photons that are in step with the original photons. This process can continue and produce an avalanche of photons, all of the same wavelength and all having their maxima and minima at the same times.

For this process to occur, certain conditions must be met. First, there must be other atoms in the excited state. Second, the atoms must remain in the excited state long enough to be struck by a photon. Third, the photons must be contained so that they are able to strike other excited atoms. In 1959, a device called a **laser** was invented that fulfilled all the conditions needed to produce coherent light. The word *laser* is an acronym that stands for *l*ight *a*mplification by *s*timulated *e*mission of *r*adiation. An atom that emits light when it is stimulated in a laser is said to lase.

Atom excitation The atoms in a laser can be excited, or pumped, as outlined in **Figure 28-17.** An intense flash of light with a wavelength shorter than that of the laser can be used to pump the atoms. The shorter wavelength, higher energy photons produced by the flash collide with and excite the lasing atoms. When one of the excited atoms decays to a lower energy state by emitting a photon, the avalanche of photons begins. This results in the emission of a brief flash, or pulse, of laser light. Alternatively, the lasing atoms can be excited by collisions with other atoms. In the helium-neon lasers often seen in science classrooms, an electric discharge excites the helium atoms. These excited helium atoms collide with the neon atoms, pumping them to an excited state and causing them to lase. The laser light resulting from this process is continuous rather than pulsed.

Lasing The photons emitted by the lasing atoms are contained by confining the lasing atoms within a glass tube that has parallel mirrors at each end. One of the mirrors is more than 99.9% reflective and reflects nearly all of the light hitting it, whereas the other mirror is partially reflective and allows only about 1 percent of the light hitting it to pass through. Photons that are emitted in the direction of the ends of the tube will be reflected back into the gas by the mirrors. The reflected photons strike more atoms, releasing more photons with each pass between the mirrors. As the process continues, a high intensity of photons builds. The photons that exit the tube through the partially reflecting mirror produce the laser beam. **Figure 28-18** shows a laser being used in a laboratory.

Because all the stimulated photons are emitted in step with the photons that struck the atoms, laser light is coherent light. The light is also all of one wavelength, or monochromatic, because the transition of electrons between only one pair of energy levels in one type of atom is involved. The parallel mirrors used in the laser result in the emitted laser light being highly directional. In other words, laser light does not diverge much as it travels. Because a typical laser beam is very small, often only about 2 mm in diameter, the light is very intense. Many solid, liquid, and gas substances can be made to lase. Most substances produce laser light at only one wavelength. The light from some lasers, however, can be tuned, or adjusted, over a range of wavelengths.

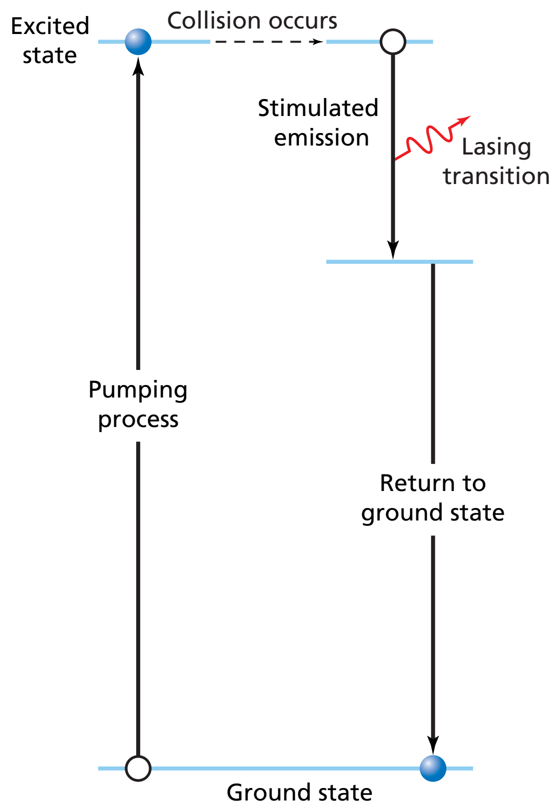


Figure 28-17 When a photon strikes an excited atom, it stimulates the atom to emit a second coherent photon and to make a transition to a lower state.

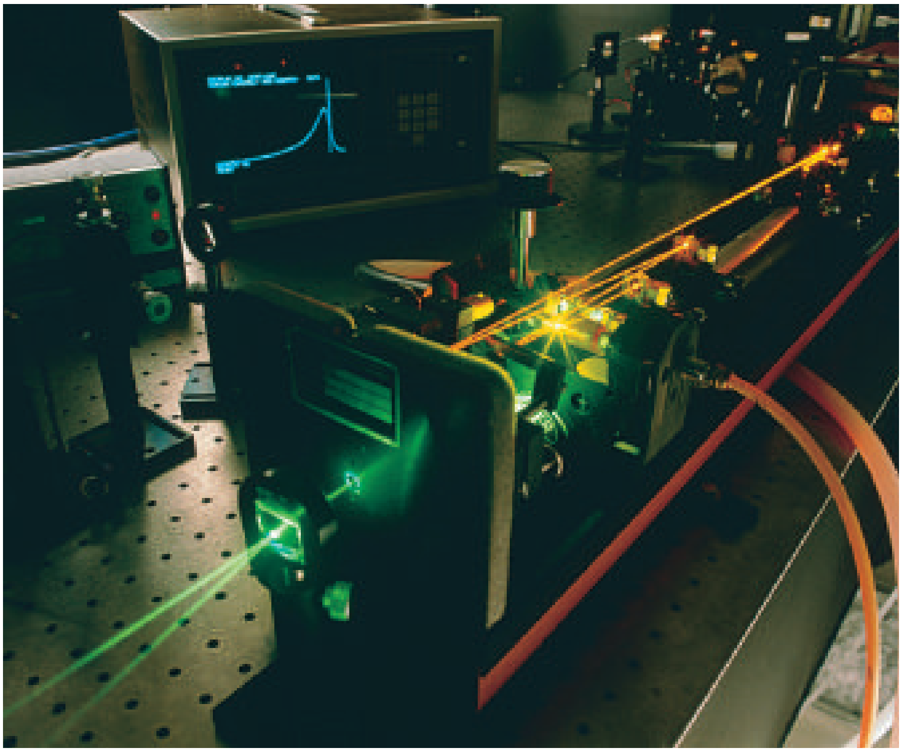


Figure 28-18 This argon laser produces a beam of coherent light.

Table 28-1		
Common Lasers		
Medium	**Wavelength (nm)**	**Type**
Krypton-fluoride excimer (KrF gas)	248 (ultraviolet)	Pulsed
Nitrogen (N_2 gas)	337 (ultraviolet)	Pulsed
Indium gallium nitride (InGaN crystal)	420	Continuous
Argon ion (Ar^+ gas)	476.5, 488.0, 514.5	Continuous
Neon (Ne gas)	632.8	Continuous
Gallium aluminum arsenide (GaAlAs crystal)	635, 680	Continuous
Gallium arsenide (GaAs crystal)	840–1350 (infrared)	Continuous
Neodymium (Nd: YAG crystal)	1064 (infrared)	Pulsed
Carbon dioxide (CO_2 gas)	10,600 (infrared)	Continuous

Laser Applications

If you have used a CD or a DVD player, then you have used a laser. These lasers, as well as the ones used in laser pointers, are made of semiconducting solids. The laser in a CD player is made of layers of gallium arsenide (GaAs) and gallium aluminum and arsenide (GaAlAs). The lasing layer is only 200-nm thick, and each side of the crystal is only 1–2 mm long. The atoms in the semiconducting solid are pumped by an electric current, and the resulting photons are amplified as they bounce between the polished ends of the crystal. **Table 28-1** shows the wavelength and type, pulsed or continuous, produced by some common lasers.

Most lasers are very inefficient. For example, no more than 1 percent of the electrical energy delivered to a gas laser is converted to light energy. Although crystal lasers have efficiencies near 20 percent, they often have much less power than gas lasers do. Despite their inefficiency, the unique properties of laser light have led to many applications. Laser beams are narrow and highly directional—they do not spread out over long distances. For this reason, surveyors use laser beams in applications such as checking the straightness of long tunnels and pipes. When astronauts visited the Moon, they left behind mirrors on the Moon's surface. Scientists on Earth have used the mirrors to reflect a laser beam transmitted from Earth. The distance between Earth and the Moon was thus accurately determined. By tracking the Moon's location from different parts of Earth, the movement of Earth's tectonic plates has been measured.

Laser light also is commonly used in fiber-optics communications. A fiber optic cable makes use of total internal reflection to transmit light over distances of many kilometers with little loss of signal energy. The laser, typically with a wavelength of 1300–1500 nm, is rapidly switched on and off, transmitting information as a series of pulses through the fiber. All over the world, optical fibers have replaced copper wires for the transmission of telephone calls, computer data, and even television pictures.

APPLYING PHYSICS

► **Laser Eye Surgery** Excimer lasers are used in laser eye surgery because the energy of the photons that they emit are able to destroy abnormal tissue without causing harm to surrounding healthy tissue. Thus, a skilled surgeon can use the laser to remove extremely thin layers of tissue in order to reshape a cornea. ◄

■ **Figure 28-19** The ultraviolet photons emitted by this laser are able to strip electrons from atoms of the targeted tissue. The photons break chemical bonds and vaporize the tissue.

The single wavelength of light emitted by lasers makes them valuable in spectroscopy. Laser light is used to excite other atoms. The atoms then return to the ground state and emit characteristic spectra. Samples with extremely small numbers of atoms can be analyzed in this way. In fact, single atoms have been detected and held almost motionless by means of laser excitation.

The concentrated power of laser light is used in a variety of ways. In medicine, for example, lasers are used to reshape the corneas of eyes. Lasers also can be used in surgery, as shown in **Figure 28-19,** in place of a knife to cut flesh with little loss of blood. In industry, lasers are used to cut materials such as steel and to weld materials together. In the future, lasers may be able to produce nuclear fusion to create an almost inexhaustible energy source.

A hologram, shown in **Figure 28-20,** is a photographic recording of both the phase and the intensity of light. Holograms are made possible by the coherent nature of laser light. Holograms form realistic three-dimensional images and can be used, among other applications, in industry to study the vibrations of sensitive equipment and their components.

■ **Figure 28-20** A hologram is formed when interference of two laser beams records both the intensity and phase of light from the object on film.

28.2 Section Review

16. Lasers Which of the lasers in Table 28-1 emits the reddest light (visible light with the longest wavelength)? Which of the lasers emit blue light? Which of the lasers emit beams that are not visible to the human eye?

17. Pumping Atoms Explain whether green light could be used to pump a red laser. Why could red light not be used to pump a green laser?

18. Bohr Model Limitations Although it was able to accurately predict the behavior of hydrogen, in what ways did Bohr's atomic model have serious shortcomings?

19. Quantum Model Explain why the Bohr model of the atom conflicts with the Heisenberg uncertainty principle, whereas the quantum model does not.

20. Lasers Explain how a laser makes use of stimulated emission to produce coherent light.

21. Laser Light What are the four characteristics of laser light that make it useful?

22. Critical Thinking Suppose that an electron cloud were to get so small that the atom was almost the size of the nucleus. Use the Heisenberg uncertainty principle to explain why this would take a tremendous amount of energy.

Physics Online physicspp.com/self_check_quiz

CONTENTS

PHYSICS LAB

Finding the Size of an Atom

Ernest Rutherford used statistical analysis and probability to help analyze the results of his gold foil experiment. In this experiment, you will model the gold foil experiment using BBs and cups. You will then analyze your results in terms of probability to estimate the size of an object that cannot be seen.

QUESTION

How can probability be used to determine the size of an object that cannot be seen?

Objectives

- **Interpret data** to determine the probability of a BB striking an unseen object.
- **Calculate** the size of an unseen object based on probability.

Safety Precautions

- **Be sure to immediately pick up BBs that have fallen onto the floor.**

Materials

shoe box
three identical small paper cups
200 BBs
centimeter ruler
large towel or cloth

Procedure

1. Use the centimeter ruler to measure the length and width of the inside of the shoe box. Record the measurements in the data table.
2. Use the centimeter ruler to measure the diameter of the top of one of the cups. Record the measurement in the data table.
3. Place the shoe box in the center of a folded towel, such that the towel extends at least 30 cm beyond each side of the shoe box.
4. Randomly place the three paper cups in the bottom of the shoe box.
5. Have your lab partner randomly drop 200 BBs into the shoe box. Make sure he or she distributes the BBs evenly over the area of the shoe box. Note that some of the BBs may miss the shoe box and land on the towel.
6. Count the number of BBs in the cups and record the value in the data table.

Analyze

1. Calculate the area of the shoe box. The area of a rectangular shape is given by the equation Area $=$ Length $\times$ Width.
2. Calculate the area of a cup using the diameter you measured. The area of a circle is given by the equation Area $= \frac{\pi(\text{Diameter})^2}{4}$.
3. Calculate the total area of the cups by multiplying the area per cup by the total number of cups.
4. Calculate the percentage of shoe box that is occupied by the three cups by dividing the total area of the cups by the area of the shoe box and then multiplying by 100.
5. Calculate the percentage of BBs that landed in the cups by dividing the number of BBs in the cups by the number of BBs dropped, and then multiplying by 100.

Horizons Companies

Data Table

	Your Data	Data from Group 2	Data from Group 3	Data from Group 4	Data from Group 5	Class Average
Shoe box length (cm)						
Shoe box width (cm)						
Shoe box area (cm^2)						
Measured diameter of cup (cm)						
Calculated area of a cup (cm^2)						
Total number of cups	3	3	3	3	3	3
Total calculated area of cups (cm^2)						
Percentage of shoe box occupied by cups (%)						
Number of BBs dropped	200	200	200	200	200	200
Number of BBs in cups						
Percent of BBs in cups						
Percent of shoe box occupied by cups based on probability						
Total area of cups based on probability (cm^2)						
Number of cups	3	3	3	3	3	3
Area of one cup based on probability (cm^2)						

6. Determine the percentage of the shoe box occupied by the cups based on probability. Note that this percentage is (ideally) equal to the percentage of BBs that landed in cups.
7. Calculate the total area of the cups based on probability. To calculate this value, multiply the percentage of the shoe box occupied by the cups (based on probability) by the area of the shoe box.
8. Calculate the area of each cup based on probability by dividing the total area of the cups based on probability by three.
9. Record the experimental data from the other groups in the data table and then calculate classroom averages for all of the data.
10. **Error Analysis** Compare your calculated value for the area of the cup based on probability (experimental value) with the area of the cup calculated from the measured diameter (accepted value). What is the percent error in your value based upon probability? Calculate the percent error using the following equation:

$$\text{Percent error} = \frac{\left|\text{Accepted value} - \text{Experimental value}\right|}{\text{Accepted value}} \times 100$$

Conclude and Apply

1. Were you able to accurately determine the area of a cup based on probability? Explain in terms of the percent error.
2. List the error sources in this experiment and describe their effects on the results.

Going Further

If larger cups were used in the experiment, do you think you would need fewer, the same, or more BBs to achieve accurate results? Explain.

Real-World Physics

Your teacher polls the class about postponing a test. Does the accuracy of the poll depend on how many students are surveyed? Explain.

Physics Online

To find out more about the atom, visit the Web site: physicspp.com

Future Technology

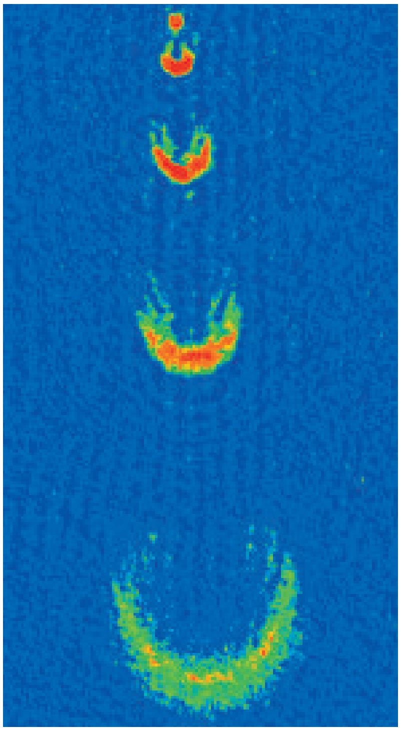
Atom Laser

The recently developed atom laser is a technology with a promising future. Unlike traditional lasers that emit beams or pulses of coherent photons, atom lasers emit beams or pulses of coherent atoms. As explained below, coherent atoms are different from the incoherent atoms that make up ordinary matter.

History In 1923, Louis Victor de Broglie predicted that all particles, including atoms, have wave properties. Wavelength is inversely related to the particle's mass and velocity, and it is too small to be observed at room temperature. As the atom is cooled, its speed is reduced and its wavelength increases.

In the 1920s, Albert Einstein and Satyendra Nath Bose researched particles known as bosons. They predicted that if bosons could be cooled to their lowest possible energy state, then all of the particles would have the same phase and the same wavelength. In other words, the particles would have coherent properties. This unusual phase of matter is called a Bose-Einstein condensate.

The first Bose-Einstein condensates were formed in 1995 by Eric Cornell and Carl Wieman, and independently, by Wolfgang Ketterle. In further research, Ketterle brought two separate Bose-Einstein condensate samples near each other. Ketterle observed wavelike interference patterns from the atoms in the condensates. He went on to prove that all of the atoms in the condensate had the same wavelength and were in phase. The atoms in the condensate were coherent, just as Bose and Einstein predicted.



An atom laser emits pulses of coherent sodium atoms. Each pulse contains 10^5 to 10^6 atoms, and the pulses are accelerated downward due to gravity. The spreading of the pulses is a result of repulsive forces.

The First Atom Laser In 1997, Ketterle and his collaborators announced the first steps in the development of an atom laser. They developed a way to eject small pulses (between 100,000 and 1,000,000 atoms) of coherent atoms from a Bose-Einstein condensate into a beam.

In this first atom laser, pulses of coherent atoms could only travel in a single direction. The emitted atoms behaved like projectiles, following downward arcing paths due to the influence of gravity. As shown in the photo, the coherent atoms in each pulse tend to spread apart as the beam propagates. In 1999, William Phillips found a way to send pulses of coherent atoms in any direction and how to prevent the atoms from spreading out as the beam propagated. By making a series of very short pulses, Phillips was able to form a continuous beam of coherent atoms.

The Future Bose-Einstein condensates and atom lasers will be used to study the basic properties of quantum mechanics and matter waves. Scientists anticipate that atom lasers will be useful in making more-precise atomic clocks and in building small electronic circuits. The atom laser also may be used in atomic interferometry to precisely measure gravitational forces and to test relativity.

Going Further

1. **Research** Investigate what fermions are and if they can form Bose-Einstein condensates. (*Hint: See how the Pauli exclusion principle applies to fermions.*)
2. **Critical Thinking** Atom lasers operate in an ultrahigh vacuum environment. Why do you think this is so?

Chapter 28 Study Guide

28.1 The Bohr Model of the Atom

Vocabulary

- alpha particles (p. 748)
- nucleus (p. 749)
- absorption spectrum (p. 751)
- energy level (p. 753)
- ground state (p. 753)
- excited state (p. 753)
- principal quantum number (p. 755)

Key Concepts

- Ernest Rutherford directed positively charged, high-speed alpha particles at thin metal foils. By studying the paths of the reflected particles, he showed that an atom is mostly empty space with a tiny, massive, positively charged nucleus at its center.
- The spectrum produced by atoms of an element can be used to identify an unknown sample of that element.
- If white light passes through a gas, the gas absorbs the same wavelengths that it would emit if it were excited. If the light is then passed through a prism, the absorption spectrum of the gas is visible.
- Niels Bohr's model of the atom correctly showed that the energy of an atom can have only certain values; thus, it is quantized. He showed that the energy of a hydrogen atom in energy level n is equal to the product of -13.6 eV and the inverse of n^2.

$$E_n = -13.6\text{ eV} \times \frac{1}{n^2}$$

- According to the Bohr model, atoms make transitions between allowable energy levels, absorbing or emitting energy in the form of photons (electromagnetic waves). The energy of the photon is equal to the difference between the initial and final states of the atom.

$$E_{\text{photon}} = E_{\text{f}} - E_{\text{i}}$$

- According to the Bohr model, the radius of an electron's orbit can have only certain (quantized) values. The radius of the electron in energy level n of a hydrogen atom is given by the following equation.

$$r_n = \frac{h^2n^2}{4\pi^2Kmq^2}$$

28.2 The Quantum Model of the Atom

Vocabulary

- quantum model (p. 760)
- electron cloud (p. 761)
- quantum mechanics (p. 761)
- coherent light (p. 761)
- incoherent light (p. 761)
- stimulated emission (p. 762)
- laser (p. 762)

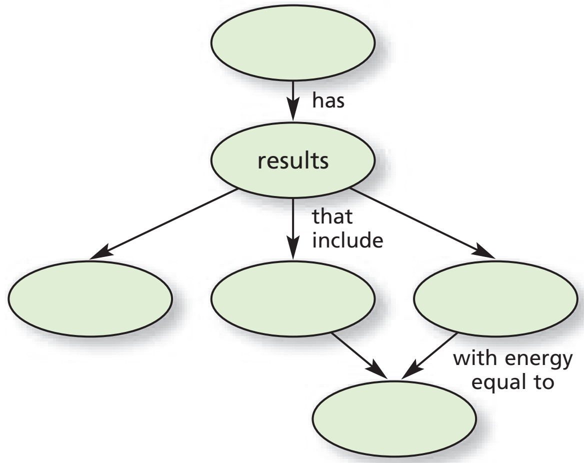
Key Concepts

- In the quantum-mechanical model of the atom, the atom's energy has only specific, quantized values.
- In the quantum-mechanical model of the atom, only the probability of finding the electron in a specific region can be determined. In the hydrogen atom, the most probable distance of the electron from the nucleus is the same as the electron's orbital radius in the Bohr model.
- Quantum mechanics is extremely successful in calculating the properties of atoms, molecules, and solids.
- Lasers produce light that is directional, powerful, monochromatic, and coherent. Each property gives the laser useful applications.

Chapter 28 Assessment

Concept Mapping

23. Complete the following concept map using these terms: *energy levels, fixed electron radii, Bohr model, photon emission and absorption, energy-level difference.*



Mastering Concepts

24. Describe how Rutherford determined that the positive charge in an atom is concentrated in a tiny region, rather than spread throughout the atom. (28.1)

25. How does the Bohr model explain why the absorption spectrum of hydrogen contains exactly the same frequencies as its emission spectrum? (28.1)

26. Review the planetary model of the atom. What are some of the problems with a planetary model of the atom? (28.1)

27. Analyze and critique the Bohr model of the atom. What three assumptions did Bohr make in developing his model? (28.1)

28. **Gas-Discharge Tubes** Explain how line spectra from gas-discharge tubes are produced. (28.1)

29. How does the Bohr model account for the spectra emitted by atoms? (28.1)

30. Explain why line spectra produced by hydrogen gas-discharge tubes are different from those produced by helium gas-discharge tubes. (28.1)

31. **Lasers** A laboratory laser has a power of only 0.8 mW (8×10^{-4} W). Why does it seem more powerful than the light of a 100-W lamp? (28.2)

32. A device similar to a laser that emits microwave radiation is called a maser. What words likely make up this acronym? (28.2)

33. What properties of laser light led to its use in light shows? (28.2)

Applying Concepts

34. As the complexity of energy levels changes from atom to atom, what do you think happens to the spectra that they produce?

35. **Northern Lights** The northern lights are caused by high-energy particles from the Sun striking atoms high in Earth's atmosphere. If you looked at these lights through a spectrometer, would you see a continuous or line spectrum? Explain.

36. If white light were emitted from Earth's surface and observed by someone in space, would its spectrum appear to be continuous? Explain.

37. Is money a good example of quantization? Is water? Explain.

38. Refer to **Figure 28-21.** A photon with energy of 6.2 eV enters a mercury atom in the ground state. Will it be absorbed by the atom? Explain.

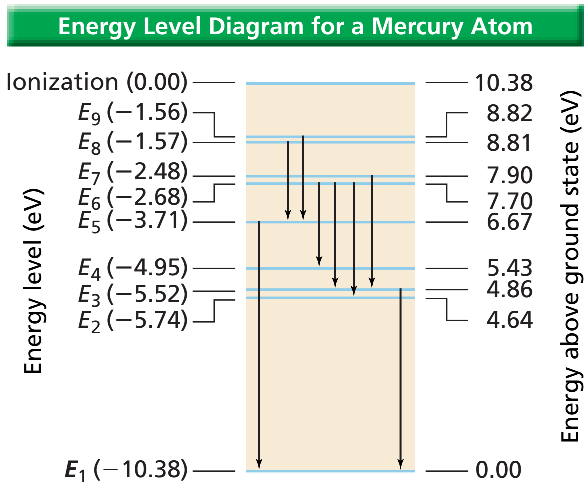


Figure 28-21

39. A certain atom has four energy levels, with E_4 being the highest and E_1 being the lowest. If the atom can make transitions between any two levels, how many spectral lines can the atom emit? Which transition produces the photon with the highest energy?

40. A photon is emitted when an electron in an excited hydrogen atom drops through energy levels. What is the maximum energy that the photon can have? If this same amount of energy were given to the atom in the ground state, what would happen?

41. Compare the quantum mechanical theory of the atom with the Bohr model.

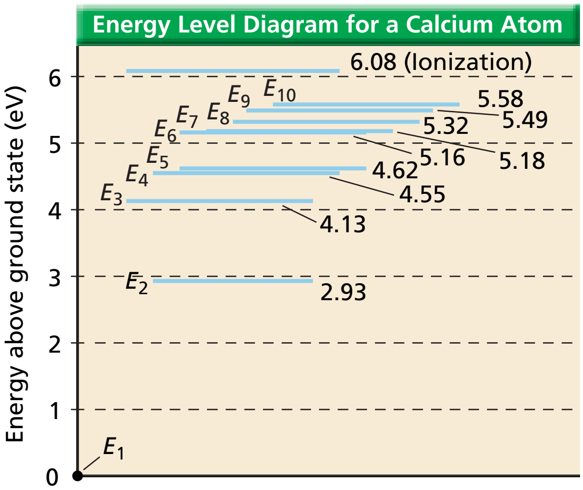
42. Given a red, green, and blue laser, which produces photons with the highest energy?

Mastering Problems

28.1 The Bohr Model of the Atom

43. A calcium atom drops from 5.16 eV above the ground state to 2.93 eV above the ground state. What is the wavelength of the photon emitted?

44. A calcium atom in an excited state, E_2, has an energy level 2.93 eV above the ground state. A photon of energy 1.20 eV strikes the calcium atom and is absorbed by it. To what energy level is the calcium atom raised? Refer to **Figure 28-22.**



■ **Figure 28-22**

45. A calcium atom is in an excited state at the E_6 energy level. How much energy is released when the atom drops down to the E_2 energy level? Refer to Figure 28-22.

46. A photon of orange light with a wavelength of 6.00×10^2 nm enters a calcium atom in the E_6 excited state and ionizes the atom. What kinetic energy will the electron have as it is ejected from the atom?

47. Calculate the energy associated with the E_7 and the E_2 energy levels of the hydrogen atom.

48. Calculate the difference in energy levels in the previous problem.

Refer to Figure 28-21 for Problems 49 and 50.

49. A mercury atom is in an excited state at the E_6 energy level.

a. How much energy would be needed to ionize the atom?

b. How much energy would be released if the atom dropped down to the E_2 energy level instead?

50. A mercury atom in an excited state has an energy of −4.95 eV. It absorbs a photon that raises it to the next-higher energy level. What is the energy and the frequency of the photon?

51. What energies are associated with a hydrogen atom's energy levels of E_2, E_3, E_4, E_5, and E_6?

52. Using the values calculated in problem 51, calculate the following energy differences.

a. $E_6 - E_5$ **c.** $E_4 - E_2$ **e.** $E_5 - E_3$
b. $E_6 - E_3$ **d.** $E_5 - E_2$

53. Use the values from problem 52 to determine the frequencies of the photons emitted when an electron in a hydrogen atom makes the energy level changes listed.

54. Determine the wavelengths of the photons having the frequencies that you calculated in problem 53.

55. A hydrogen atom emits a photon with a wavelength of 94.3 nm when its falls to the ground state. From what energy level did the electron fall?

56. For a hydrogen atom in the $n = 3$ Bohr orbital, find the following.

a. the radius of the orbital

b. the electric force acting between the proton and the electron

c. the centripetal acceleration of the electron

d. the orbital speed of the electron (Compare this speed with the speed of light.)

28.2 The Quantum Model of the Atom

57. CD Players Gallium arsenide lasers are commonly used in CD players. If such a laser emits at 840 nm, what is the difference in eV between the two lasing energy levels?

58. A GaInNi laser lases between energy levels that are separated by 2.90 eV.

a. What wavelength of light does it emit?

b. In what part of the spectrum is this light?

59. A carbon-dioxide laser emits very high-power infrared radiation. What is the energy difference in eV between the two lasing energy levels? Consult Table 28-1.

60. The power in a laser beam is equal to the energy of each photon times the number of photons per second that are emitted.

a. If you want a laser at 840 nm to have the same power as one at 427 nm, how many times more photons per second are needed?

b. Find the number of photons per second in a 5.0-mW 840-nm laser.

61. HeNe Lasers The HeNe lasers used in many classrooms can be made to lase at three wavelengths: 632.8 nm, 543.4 nm, and 1152.3 nm.

a. Find the difference in energy between the two states involved in the generation of each wavelength.

b. Identify the color of each wavelength.

Mixed Review

62. A photon with an energy of 14.0 eV enters a hydrogen atom in the ground state and ionizes it. With what kinetic energy will the electron be ejected from the atom?

63. Calculate the radius of the orbital associated with the energy levels E_5 and E_6 of the hydrogen atom.

64. A hydrogen atom is in the $n = 2$ level.
- **a.** If a photon with a wavelength of 332 nm strikes the atom, show that the atom will be ionized.
- **b.** When the atom is ionized, assume that the electron receives the excess energy from the ionization. What will be the kinetic energy of the electron in joules?

65. A beam of electrons is directed onto a sample of atomic hydrogen gas. What minimum energy of the electrons is needed for the hydrogen atoms to emit the red light produced when the atom goes from the $n = 3$ to the $n = 2$ state?

66. The most precise spectroscopy experiments use "two-photon" techniques. Two photons with identical wavelengths are directed at the target atoms from opposite directions. Each photon has half the energy needed to excite the atoms from the ground state to the desired energy level. What laser wavelength would be needed to make a precise study of the energy difference between $n = 1$ and $n = 2$ in hydrogen?

Thinking Critically

67. **Apply Concepts** The result of projecting the spectrum of a high-pressure mercury vapor lamp onto a wall in a dark room is shown in **Figure 28-23.** What are the differences in energy levels for each of the three visible lines?

Figure 28-23

68. **Interpret Scientific Illustrations** After the emission of the visible photons described in problem 67, the mercury atom continues to emit photons until it reaches the ground state. From an inspection of Figure 28-21, determine whether or not any of these photons would be visible. Explain.

69. **Anaylze and Conclude** A positronium atom consists of an electron and its antimatter relative, the positron, bound together. Although the lifetime of this "atom" is very short—on the average it lives one-seventh of a microsecond—its energy levels can be measured. The Bohr model can be used to calculate energies with the mass of the electron replaced by one-half its mass. Describe how the radii of the orbits and the energy of each level would be affected. What would be the wavelength of the E_2 to E_1 transition?

Writing in Physics

70. Do research on the history of models of the atom. Briefly describe each model and identify its strengths and weaknesses.

71. Green laser pointers emit light with a wavelength of 532 nm. Do research on the type of laser used in this type of pointer and describe its operation. Indicate whether the laser is pulsed or continuous.

Cumulative Review

72. The force on a test charge of $+3.00\times10^{-7}$ C is 0.027 N. What is the electric field strength at the position of the test charge? (Chapter 21)

73. A technician needs a 4-Ω resistor but only has 1-Ω resistors of that value. Is there a way to combine what she has? Explain. (Chapter 23)

74. A 1.0-m-long wire is moved at right angles to Earth's magnetic field where the magnetic induction is 5.0×10^{-5} T at a speed of 4.0 m/s. What is the *EMF* induced in the wire? (Chapter 25)

75. The electrons in a beam move at 2.8×10^{8} m/s in an electric field of 1.4×10^{4} N/C. What value must the magnetic field have if the electrons pass through the crossed fields undeflected? (Chapter 26)

76. Consider the modifications that J. J. Thomson would need to make to his cathode-ray tube so that it could accelerate protons (rather than electrons), then answer the following questions. (Chapter 26)
- **a.** To select particles of the same velocity, would the ratio E/B have to be changed? Explain.
- **b.** For the deflection caused by the magnetic field alone to remain the same, would the B field have to be made smaller or larger? Explain.

77. The stopping potential needed to return all the electrons ejected from a metal is 7.3 V. What is the maximum kinetic energy of the electrons in joules? (Chapter 27)

 For more problems, go to Additional Problems, Appendix B.

Standardized Test Practice

Multiple Choice

1. Which model of the atom was based on the results of Rutherford's gold foil experiment?

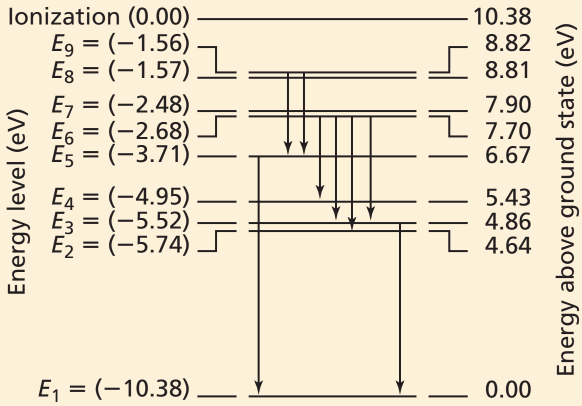
   (A) Bohr model
   (B) nuclear model
   (C) plum pudding model
   (D) quantum mechanical model

2. A mercury atom emits light with a wavelength of 405 nm. What is the energy difference between the two energy levels involved with this emission?

   (A) 0.22 eV
   (B) 2.14 eV
   (C) 3.06 eV
   (D) 4.05 eV

3. The diagram below shows the energy levels for a mercury atom. What wavelength of light is emitted when a mercury atom makes a transition from E_7 to E_4?

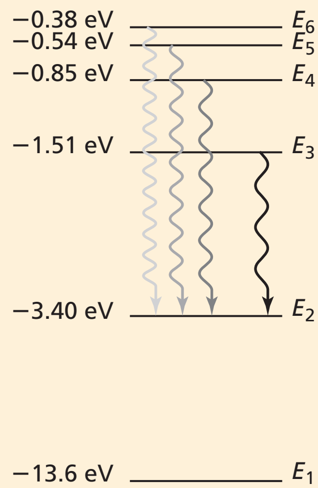
   (A) 167 nm
   (B) 251 nm
   (C) 500 nm
   (D) 502 nm



4. Which statement about the quantum model of the atom is false?

   (A) The possible energy levels of the atom are quantized.
   (B) The locations of the electrons around the nucleus are known precisely.
   (C) The electron cloud defines the area where electrons are likely to be located.
   (D) Stable electron orbits are related to the de Broglie wavelength.

Questions 5 and 6 refer to the diagram showing the Balmer series electron transitions in a hydrogen atom.



5. Which energy level transition is responsible for the emission of light with the greatest frequency?

   (A) E_2 to E_5
   (B) E_3 to E_2
   (C) E_3 to E_6
   (D) E_6 to E_2

6. What is the frequency of the Balmer series line related to the energy level transition from E_4 to E_2? (Note that 1 eV = 1.60×10^{-19} J.)

   (A) 2.55×10^{14} Hz
   (B) 4.32×10^{14} Hz
   (C) 6.15×10^{14} Hz
   (D) 1.08×10^{15} Hz

Extended Answer

7. Determine the wavelength of light emitted when a hydrogen atom makes a transition from the $n = 5$ energy level to the $n = 2$ energy level.

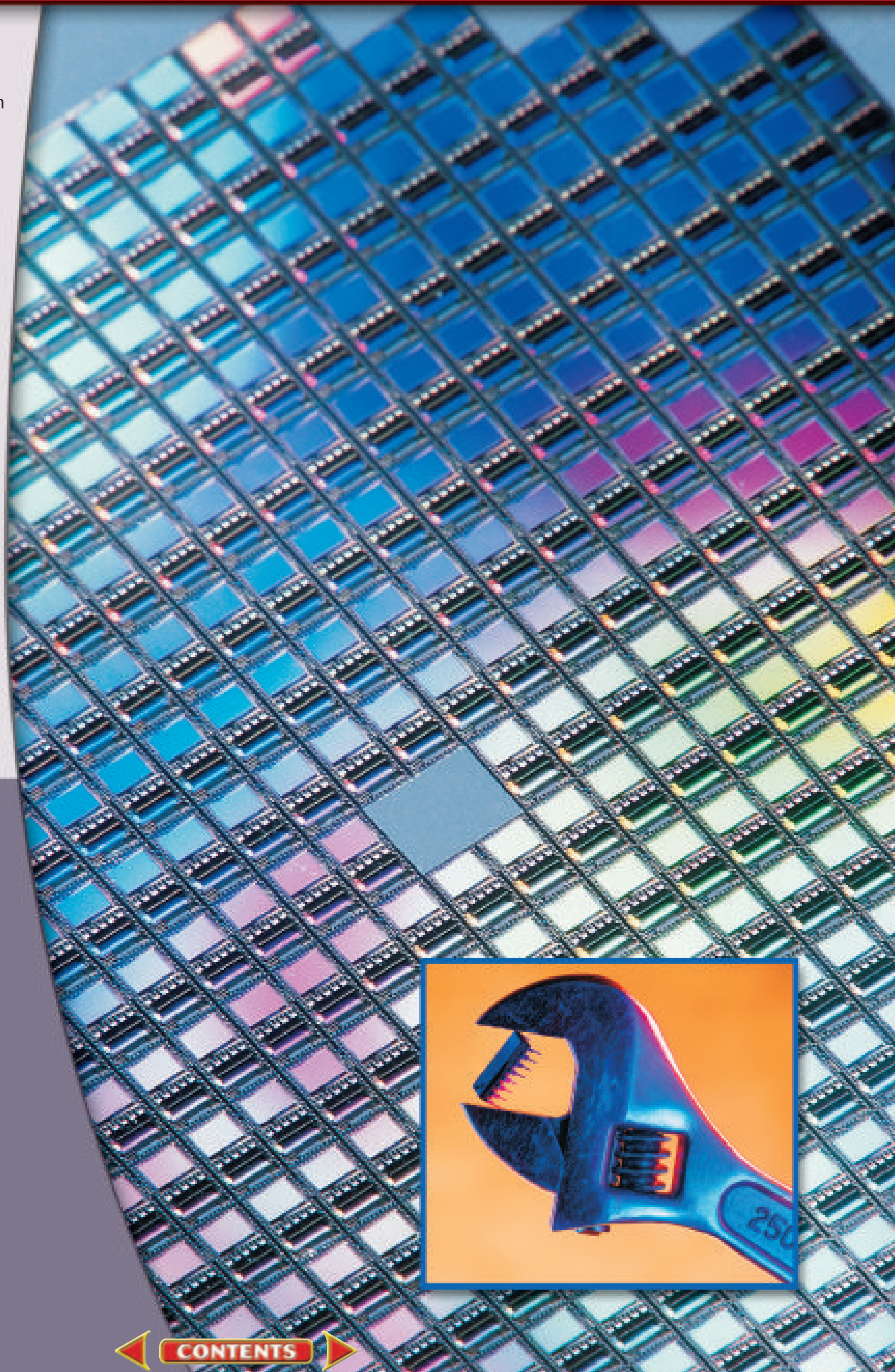
✓ Test-Taking TIP

Stumbling Is Not Failing

Occasionally, you will encounter a question that you have no idea how to answer. Even after reading the question several times, it still may not make sense. If the question is multiple choice, focus on the part of the question that you know something about. Eliminate as many choices as you can. Then take your best guess and move on.

Chapter 29

Solid-State Electronics



What You'll Learn

- You will be able to distinguish among electric conductors, semiconductors, and insulators.
- You will examine how pure semiconductors are modified to produce desired electric properties.
- You will compare diodes and transistors.

Why It's Important

Semiconductors have electric properties that allow them to act as one-way conductors to amplify weak signals in many common electronic devices.

Fast Math Computers and electronic devices use the controlled movement of electrons and holes in semiconductors to do quick calculations and logical operations.

Think About This ▶

A silicon microchip might be small, but it may contain the equivalent of millions of resistors, diodes, and transistors. How can this level of complexity be produced in such a tiny structure?

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