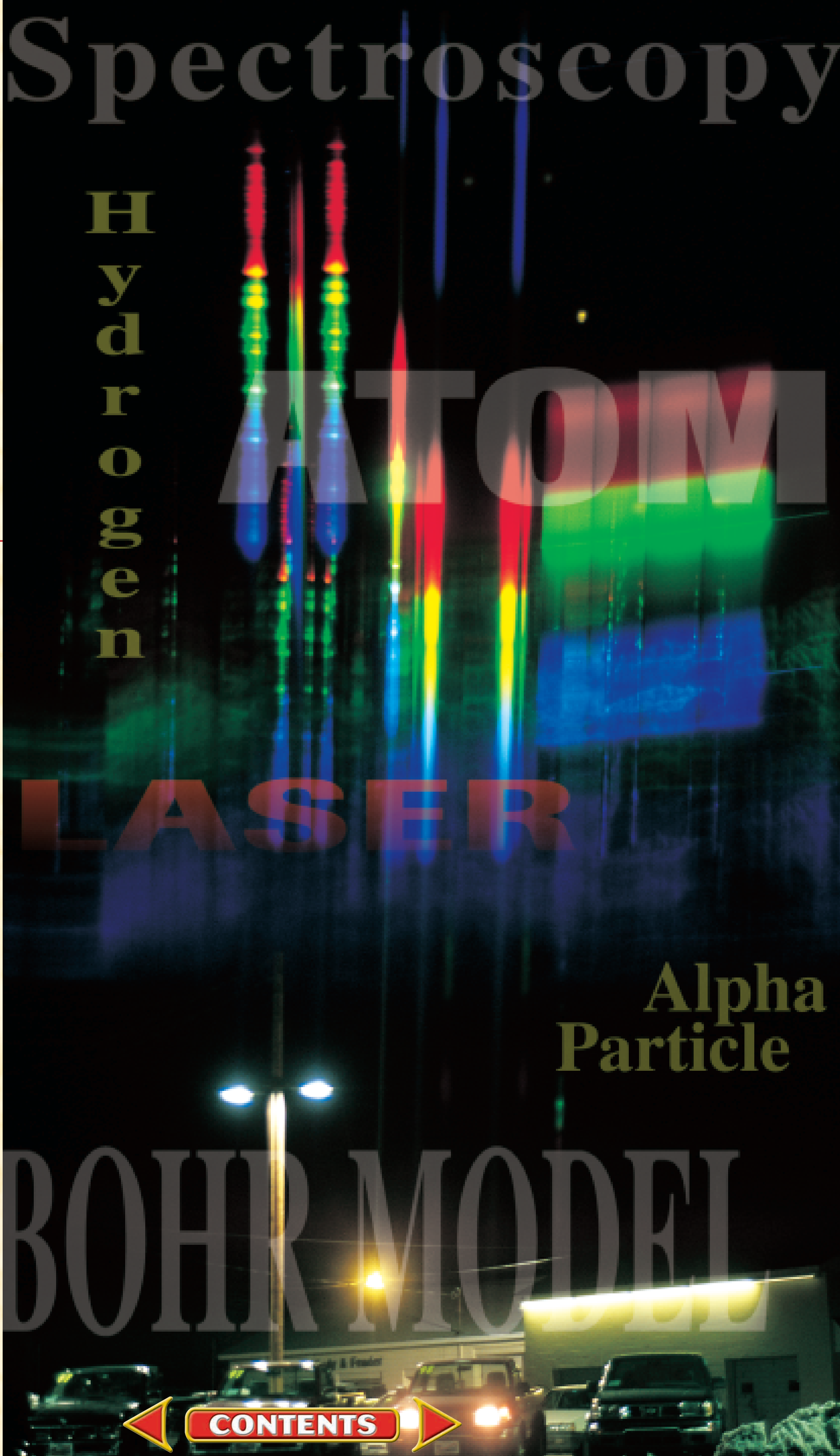
Atomic Secrets

The street lamps, the building lights, and the car headlights were photographed with a diffraction grating over the camera lens. Why are the images of the lights different, and how could these differences be used to identify the types of lights?

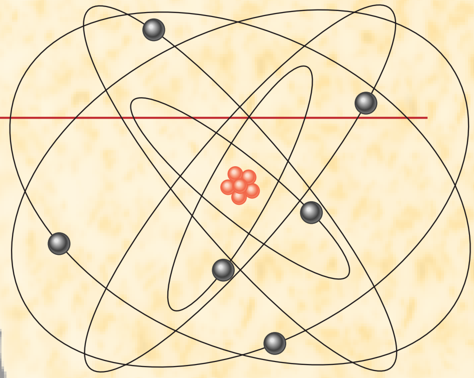
Look at the text on page 649 for the answer.



CONTENTS

CHAPTER

28 The Atom



You have seen that both prisms and diffraction gratings separate light into its component colors. But do all sources of light produce the same spectrum, that is, the same mix of colors? Look at the light sources in the photograph at the left—headlights, street lamps, and building lights. The spectra of these light sources were produced by placing a diffraction grating over the camera lens. Note the differences among the spectra. Some spectra include every color; in others, there are brighter bands at specific colors.

For more than 100 years, researchers have studied spectra such as these to identify sources of light. Because the light sources are made up of different materials, scientists concluded that different atoms produce different spectra. This relationship between chemical composition and spectra has been helpful in astronomy. Because analyzing the spectrum of the light produced by an unknown material can reveal its chemical makeup, this process has been used to determine the array of atoms present in the sun and in distant stars. In fact, astronomers can even use the spectra of stars to determine the surface temperature of a star.

Today scientists use spectral analysis to identify and characterize matter. Spectral analysis also helps delineate impurities in manufactured items such as molten iron or glass.

When the study of spectra began, the structure of atoms was unknown. However, the study of the spectrum produced by the simplest atom—hydrogen—led to the discovery of the structure of not only hydrogen, but also of all other atoms.

WHAT YOU'LL LEARN

- You will examine the composition of the atom.
- You will determine specific energies of electrons.
- You will explore the probability of finding the electrons at specific locations or with specific momenta.
- You will study lasers.

WHY IT'S IMPORTANT

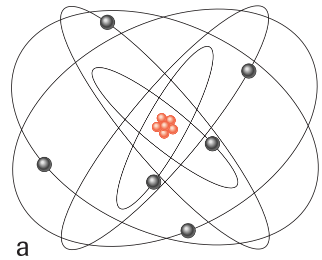
- The quantum model of the atom and the transition of electrons between orbitals are responsible for the ability to produce artificial lighting, determine the composition of materials, operate electronic equipment, and produce special effects for movies.

PHYSICS Online

To find out more about the atom, visit the Glencoe Science Web site at **science.glencoe.com**

CLICK HERE

28.1 The Bohr Model of the Atom



OBJECTIVES

- **Explain** the structure of the atom.
- **Distinguish** continuous spectra from line spectra.
- **Contrast** emission and absorption spectra.
- **Solve** problems using the orbital radius and energy-level equations.

By the end of the 19th century, most scientists agreed that atoms exist. Furthermore, as a result of the discovery of the electron by J. J. Thomson, they agreed that the atom could not be an indivisible particle. All of the atoms that Thomson tested contained electrons that had a negative charge. Yet atoms were known to be electrically neutral and much more massive than electrons. Therefore it followed that atoms must contain not only electrons, but massive, positively charged parts as well.

The Nuclear Model

Discovering the nature of the massive part of the atom and the arrangement of the electrons was a major challenge. Physicists and chemists from many countries both cooperated and competed in searching for the solution to this puzzle. The result provided not only knowledge of the structure of the atom, but also a totally new approach to both physics and chemistry. The history of this work 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 electrons arranged within this substance like raisins in a muffin. However, Ernest Rutherford, who was working in England at the same time, performed a series of brilliant experiments showing that the atom had a very different structure.

Compounds containing uranium had been found to emit penetrating rays. Some of these emissions were found to be massive, positively charged particles moving at high speed. These were later named **alpha (α) particles.** The α particles could be detected by a screen coated with zinc sulfide that emitted a small flash of light, or **scintillation,** each time an α particle hit it, as shown in **Figure 28–1.**

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J. J. Thomson, the discoverer of the electron, first called these subatomic particles "corpuscles."

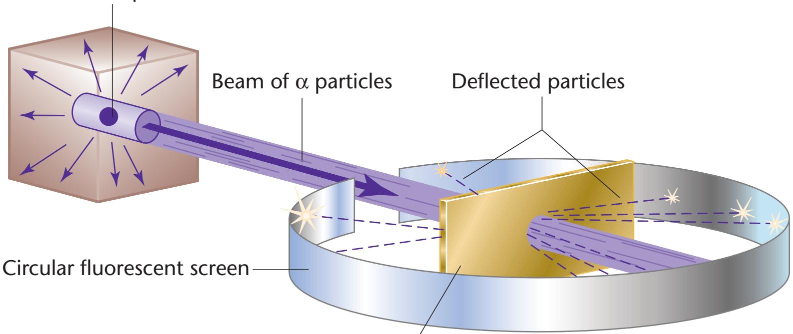


FIGURE 28–1 After bombarding metal foil with alpha particles, Rutherford's team concluded that most of the mass of the atom was concentrated in the nucleus.

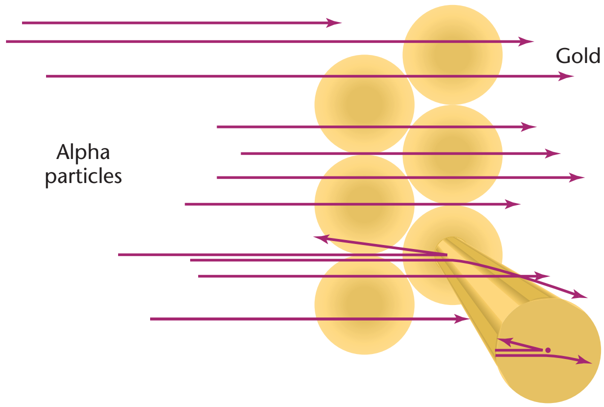


FIGURE 28–2 Most of the alpha particles that Rutherford directed at a thin sheet of gold foil went through it without deflection. A few, however, were deflected at large angles.

Rutherford directed a beam of α particles at a thin sheet of gold foil only a few atoms thick. He observed that while most of the α particles passed through the sheet, the beam was spread slightly by the foil. Detailed studies of the deflection of α particles showed that a few of the particles were deflected at large angles, even larger than 90°. A diagram of this is shown in **Figure 28–2.** Rutherford was amazed. He had assumed that the mass was spread more or less evenly throughout the atom. He 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.

Using Coulomb's force law and Newton's laws of motion, Rutherford concluded that the results could be explained only if all the positive charge of the atom were concentrated in a tiny, massive central core, now called the nucleus. Rutherford's model is therefore called the **nuclear model** of the atom. All the positive charge and more than 99.9 percent of the mass of the atom are in its nucleus. Electrons are outside and far away from the nucleus, and they do not contribute a significant amount of mass. The diameter of the atom is 10 000 times larger than the diameter of the nucleus; thus, the atom is mostly empty space. This discovery accounted for the previous observations that almost all of the α particles had passed through the gold foil, and that some of the particles had been deflected at large angles.

Atomic spectra How are the electrons arranged around the nucleus of the atom? One of the clues scientists used to answer this question came from studying the light emitted by atoms. The set of light wavelengths emitted by an atom is called the atom's **emission spectrum.**

Atoms of gases can be made to emit their characteristic colors by using a gas discharge tube apparatus. A glass tube containing a gas at low pressure has metal electrodes at each end. When a high voltage of electricity is applied across the tube, electrons pass through the gas. The electrons collide with the gas atoms, transferring energy to them. The atoms then give up this extra energy, emitting it in the form of light. The light emitted by mercury is shown in **Figure 28–3.**

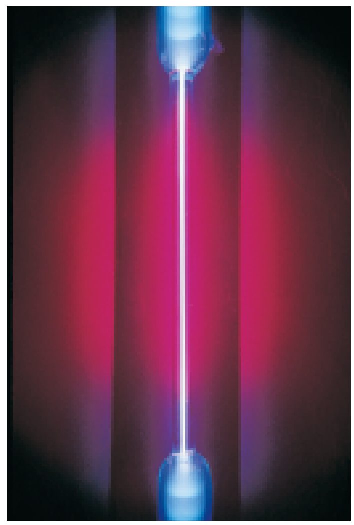


FIGURE 28–3 In a gas discharge tube apparatus, mercury gas glows when high voltage is applied.

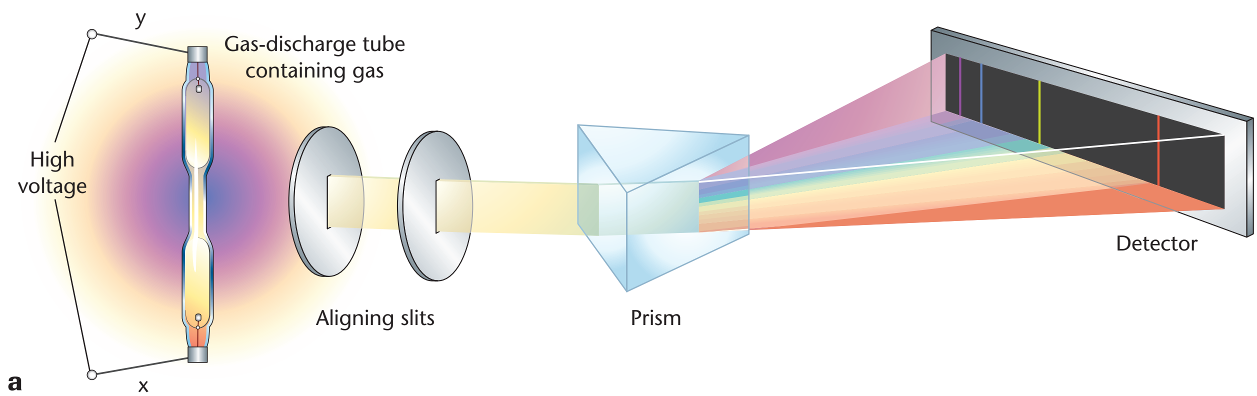
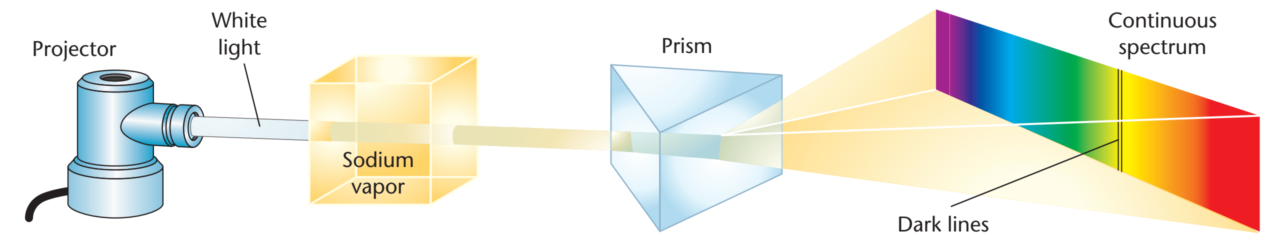


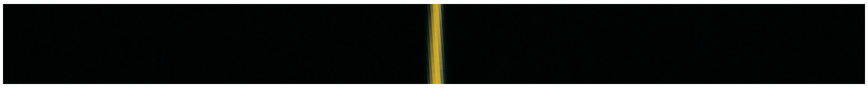
FIGURE 28–4 A prism spectroscope can be used to observe emission spectra **(a, b).** The emission spectra of neon **(c)** and molecular hydrogen **(d)** show characteristic lines.

The emission spectrum of an atom can be seen by looking at the light through a prism or a diffraction grating or by putting such a grating in front of a camera lens, as was done in the opening photo. The spectrum can be studied in greater detail using the instrument diagrammed in **Figure 28–4a** and pictured in **Figure 28–4b.** In this **spectroscope,** the light passes through a slit and is then dispersed as it travels through a prism or diffraction grating. A lens system collects the dispersed light for viewing through a telescope or for recording on a photographic plate. Each wavelength of light forms an image of the slit at different positions. The spectrum of an incandescent solid is a continuous band of colors from red through violet. The spectrum of a gas, however, is a series of lines of different colors. Each line corresponds to a particular wavelength of light emitted by the atoms of the gas. Suppose an unidentified gas, perhaps neon or hydrogen, is contained in a tube. When it is excited, the gas will emit light at wavelengths characteristic of the atoms of that gas. Thus, the gas can be identified by comparing its wavelengths with the lines present in the spectrum of a known sample. The emission spectra for neon and hydrogen are shown in **Figure 28–4c,** and **d,** respectively.

How can the differences in these emissions be used to identify the atoms involved? When the emission spectrum of a combination of elements is photographed, 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. Through comparison of the intensities of the lines, the percentage composition of the material can be determined. Thus, an emission spectrum is a useful analytic tool.

FIGURE 28–5 This apparatus is used to produce the absorption spectrum of sodium.




a

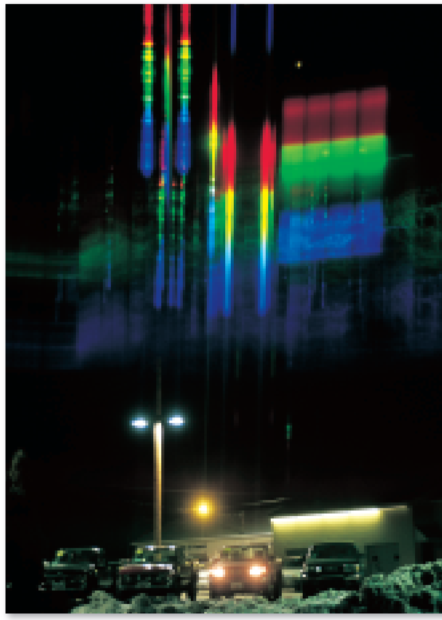

b

FIGURE 28–6 The emission spectrum **(a)** and the absorption spectrum **(b)** of sodium are pictured.

A gas that is cool and does not emit light will absorb light at characteristic wavelengths. This set of wavelengths is called an **absorption spectrum.** To obtain an absorption spectrum, white light is sent through a sample of gas and then through a spectroscope, as shown in **Figure 28–5.** The normally continuous spectrum of the white light then has dark lines in it. These lines show that light of some wavelengths has been absorbed. Often, the bright lines of the emission spectrum and the dark lines of the absorption spectrum of a gas occur at the same wavelengths. Thus, cool gaseous elements absorb the same wavelengths that they emit when excited, as shown in **Figure 28–6** for the emission spectrum and the absorption spectrum of sodium. Analysis of the wavelengths of the dark lines in an absorption spectrum also can be used to indicate the composition of the gas.

Atomic Secrets

Answers question from page 644.



In 1814, while examining the spectrum of sunlight, Josef von Fraunhofer noticed some dark lines. These dark lines, now called Fraunhofer lines, are shown in **Figure 28–7.** To account for these lines, Fraunhofer assumed that the sun has a relatively cool atmosphere of gaseous elements. He reasoned that as light leaves the sun, it passes through these gases, which absorb light at their characteristic wavelengths. As a result, these wavelengths are missing from the sun's absorption spectrum. Through a comparison of the missing lines with the known lines of the various elements, the composition of the atmosphere of the sun was determined. In this manner, the element helium was discovered in the sun before it was found on Earth. Spectrographic analysis also has made it possible to determine the composition of stars.

ASTRONOMY CONNECTION

Both emission and absorption spectra are valuable scientific tools. As a result of the elements' characteristic spectra, chemists are able to analyze , identify, and quantify unknown materials by observing the spectra 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.

FIGURE 28–7 Fraunhofer lines appear in the absorption spectrum of the sun.

FIGURE 28–8 The emission spectrum of hydrogen in the visible range has four lines.

The study of spectra is a branch of science known as **spectroscopy.** Spectroscopists are employed throughout research and industrial communities. Spectroscopy has proven an effective tool to analyze materials on Earth, and it is the only currently available tool to study 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 28–8.**

Any theory that explained the structure of the atom would have to account for these wavelengths, and it would also have to fit Rutherford's nuclear model. Rutherford had suggested that electrons orbited the nucleus much as the planets orbit the sun. There was, however, a serious problem with this planetary model.

An electron in an orbit is constantly accelerated toward the nucleus. As you learned in Chapter 26, electrons that have been accelerated will radiate energy by emitting electromagnetic waves. At the rate that an orbiting electron would lose energy, it would spiral into the nucleus in only 10^{-9} second. But, atoms are known to be stable. Thus, the planetary model was not consistent with the laws of electromagnetism. In addition, if the planetary theory were true, the accelerated electron should radiate energy at all wavelengths. But, as you have seen, the light emitted by atoms is radiated only at specific wavelengths.

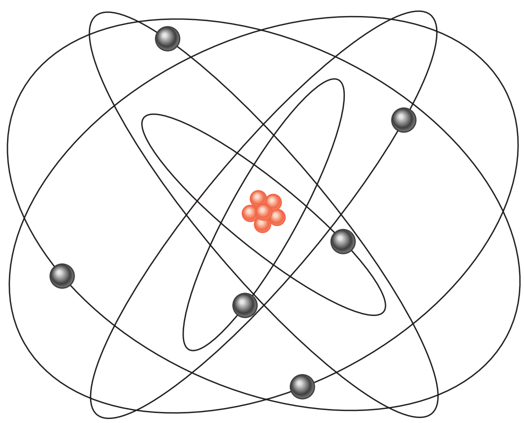


FIGURE 28–9 Bohr's planetary model of the atom postulated that electrons moved in fixed orbits around the nucleus.

Danish physicist Niels Bohr went to England in 1911 and soon joined Rutherford's group to work on determining the structure of the atom. He tried to unite the nuclear model with Einstein's theory of light. This was a courageous idea, because in 1911 Einstein's revolutionary theory of the photoelectric effect had not yet been confirmed by experiments and was not widely accepted. Recall from Chapter 27 that according to Einstein, light and other forms of electromagnetic radiation consist of discrete bundles of energy called photons. That is, light seems to act like a stream of tiny particles.

F.Y.I.

The proton was observed in 1919 as a particle that is emitted when the nucleus of an atom is bombarded by alpha particles.

Bohr energy is quantized Bohr started with the planetary arrangement of electrons, diagrammed in **Figure 28–9,** but he made the bold hypothesis that the laws of electromagnetism do not operate inside atoms. He hypothesized that an electron in a stable orbit does not radiate energy, even though it is accelerating.

If energy was not radiated when electrons were in stable orbits but elements could emit a characteristic energy spectrum, when was the energy radiated? Bohr suggested that light is emitted when the electron's energy changes, as shown in **Figure 28–10.** According to Einstein, the energy of a photon of light is represented by the equation $E = hf = hc/\lambda$. Bohr reasoned that if the emission spectrum contains only certain wavelengths, then an electron can emit or absorb only specific amounts of energy. Therefore, the atomic electrons can have only certain amounts of energy. Recall from Chapter 27 that when energy is found only in certain amounts, it is said to be quantized.

The quantization of energy in atoms is unlike everyday experience. For example, if the energy of a pendulum were quantized, it could only oscillate with certain amplitudes, such as 10 cm or 20 cm, but not for example, 11.3 cm. Electrons in an atom have different quantized amounts of energy that are called **energy levels.** When an electron has the smallest allowable amount of energy, it is in the lowest energy level, called the **ground state.** If an electron absorbs energy, it can make a transition to a higher energy level, called an **excited state.** Atomic electrons usually remain in excited states for only a few billionths of a second before returning to the ground state and emitting energy.

The energy of an orbiting electron in an atom is the sum of the kinetic energy of the electron and the potential energy resulting from the attractive force between the electron and the nucleus. The energy of an electron in an orbit near the nucleus is less than that of an electron in an orbit farther away because work must be done to move an electron to orbits farther away from the nucleus. The electrons in excited states have larger orbits and correspondingly higher energies. The model of an atom having a central nucleus with its electrons occupying specific quantized energy levels is known as the **Bohr model** of the atom.

Einstein's theory holds that the light photon has an energy *hf*. Bohr postulated that the change in the energy of an atomic electron when a photon is absorbed is equal to the energy of that photon. When the electron

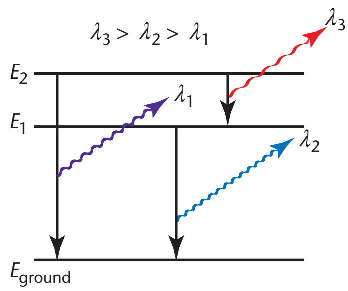


FIGURE 28–10 The energy of the emitted photon is equal to the difference in energy between two energy levels.

makes the return transition to the ground state, a photon is emitted. The energy of the photon is equal to the energy difference between the excited and ground states as defined by the following relationship.

Energy of an Emitted Photon $hf = E_{excited} - E_{ground} = \Delta E$

Molecules have additional discrete energy levels. For example, they can rotate and vibrate, which individual atoms cannot do. As a result, molecules can emit a much wider variety of light frequencies than can individual atoms.

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 was able to calculate the wavelengths of light emitted by hydrogen. 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 could not predict the spectrum of the next simplest element, helium. In addition, there was no reason to suggest that 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, the Bohr model describes the energy levels and wavelengths of light emitted and absorbed by hydrogen atoms remarkably well. You can use this model to calculate the wavelengths of light emitted by an atom. Bohr's calculations start with Newton's law, $F = ma$, applied to an electron of mass m and charge $-q$ in a circular orbit of radius r about a massive particle, a proton, of charge q. Remember that $F = Kq^2/r^2$ and that ma can be represented as mv^2/r. Thus, the following is true.

$$F = ma$$

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

Here, K is the constant, 9.0×10^9 N·m^2/C^2, from Coulomb's law.

Pocket Lab

Nuclear Bouncing

Place a 9" aluminum pie pan on a table. Gently press four glass marbles (protons) into the pie pan so that they sit in small indentations near the center of the pan. Roll a 12-mm steel ball (alpha particle) down a grooved ruler to see if you can hit the marbles. Then remove the marbles and put the steel balls into the indentations (each steel ball represents a nucleus) in the pie pan and roll a marble (alpha particle) down the grooved ruler.

Analyze and Conclude When you roll the steel ball at the marbles, does it change its path? Does the steel ball ever bounce back? When you switch the balls and marbles, how are the results different? Why are the results different? Hypothesize what will happen when an alpha particle hits a proton.

FIGURE 28–11 Ernest Rutherford and Niels Bohr devised the planetary model of the atom.

Bohr proposed that the angular momentum, which is the product of the momentum of the electron and the radius of its circular orbit, mvr, can have only certain values. These values are represented by the equation $mvr = nh/2\pi$. Here, h is Planck's constant and n is an integer. Because the quantity mvr can have only certain values, the angular momentum is said to be quantized.

Because $Kq^2/r^2 = mv^2/r$, and $v = nh/2m\pi r$, Bohr substituted for v and predicted the radii of the orbits of the electrons in the hydrogen atom by using the following equation.

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

By substituting SI values for the quantities into the equation, you can calculate the radius of the innermost orbit of the hydrogen atom, where $n = 1$, by using the following relationship.

$$r_n = \frac{(6.626 \times 10^{-34}\ \mathrm{J{\cdot}s})^2\,(1)^2}{4\pi^2(9.00 \times 10^9\ \mathrm{N{\cdot}m^2/C^2})(9.11 \times 10^{-31}\ \mathrm{kg})(1.60 \times 10^{-19}\ \mathrm{C})^2}$$

$$= 5.30 \times 10^{-11}\ \mathrm{J^2{\cdot}s^2/N{\cdot}m^2{\cdot}kg}$$

$$= 5.30 \times 10^{-11}\ \mathrm{m,\ or\ 0.053\ nm}$$

A little more algebra shows that the total energy of the electron in its orbit, which is the sum of the potential and kinetic energies of the electron and is defined by $-1/2\ Kq^2/r$, is represented by the following equation.

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

By substituting numerical values for the constants, you can calculate the energy of the electron in joules, which yields the following equation.

$$E_n = -2.17 \times 10^{-18}\ \mathrm{J} \times \frac{1}{n^2}$$

This can be written in electron volts by the following equation.

Energy of an Electron in Its Orbit $E_n = -13.6\ \mathrm{eV} \times \dfrac{1}{n^2}$

Both the radius of an orbit and the energy of the electron can have only certain values. That is, both are quantized. The integer n is called the principal **quantum number.** The number n determines the values of r and E. **Figure 28–12** shows that the radius increases as the square of n. The energy depends on $1/n^2$. The energy is negative to indicate the amount of energy that must to be added to the electron in the energy level to free it from the attractive force of the nucleus, that is, to ionize the atom. When an electron moves from a lower energy level to a higher energy level, it gains energy. The value of the energy difference is positive. The energy levels of hydrogen are shown in **Figure 28–13.**

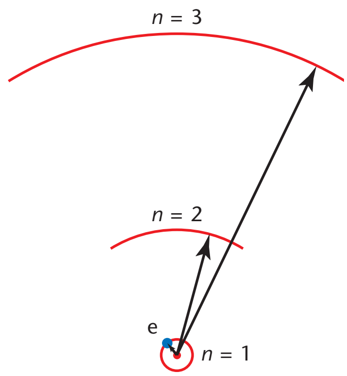


FIGURE 28–12 Radii of electron orbits for the first three energy levels of hydrogen according to the Bohr model.

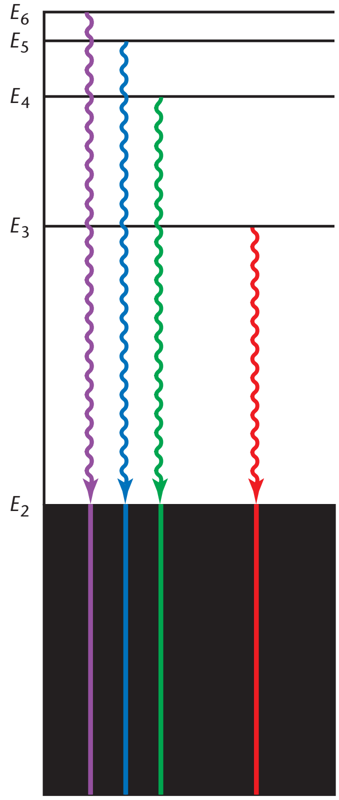


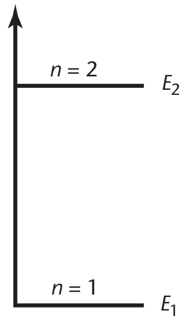
FIGURE 28–13 Bohr's model of the hydrogen atom showed that a definite amount of energy is released when an electron moves from a higher to a lower energy level. The energy released in each transition corresponds to a definite line in the hydrogen spectrum.

The light emitted by hydrogen when the atom drops into its ground state from any excited state is in the ultraviolet range. The four visible lines in the hydrogen spectrum are all produced when the atom drops from the $n = 3$ or higher state into the $n = 2$ state.

Example Problem

Orbital Energy of Electrons in the Hydrogen Atom

For the hydrogen atom, determine the energy of the innermost energy level ($n = 1$), the energy of the second energy level ($n = 2$), and the energy difference between the first and second energy levels.



Sketch the Problem

- Diagram two energy levels 1 and 2. Indicate increasing energy.

Calculate Your Answer

Known:

innermost energy level, $n = 1$

next energy level, $n = 2$

Unknown:

energy of level 1, $E_1 = ?$

energy of level 2, $E_2 = ?$

difference in energy, $\Delta E = ?$

Strategy:

Use $E_n = -13.6\text{ eV} \times 1/n^2$ for each energy level and calculate the difference between the energy levels.

Calculations:

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

$$n = 1,\ E_1 = -13.6\text{ eV} \times \frac{1}{1^2} = -13.6\text{ eV}$$

$$n = 2,\ E_2 = -13.6\text{ eV} \times \frac{1}{2^2} = -3.40\text{ eV}$$

$$\Delta E = E_2 - E_1$$

$$= -3.40\text{ eV} - (-13.6\text{ eV})$$

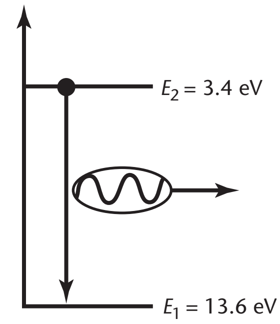
$$= 10.2\text{ eV} = 10.2\text{ eV of energy is absorbed.}$$

Check Your Answer

- Are the units correct? Energy values are 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.

Example Problem

Frequency and Wavelength of Emitted Photons



An electron in an excited hydrogen atom drops from the second energy level to the first energy level. Calculate the energy, the frequency, and the wavelength of the photon emitted.

Sketch the Problem

- Diagram two energy levels showing the electron moving from level 2 to level 1.

Calculate Your Answer

Known:

second energy level, $n = 2$

innermost energy level, $n = 1$

Planck's constant, $h = 6.63 \times 10^{-34}$ J·s

speed of light, $c = 3.00 \times 10^{8}$ m/s

energy of 1 eV $= 1.60 \times 10^{-19}$ J

Unknown:

energy of level 2, $E_2 = ?$

energy of level 1, $E_1 = ?$

difference in energy, $\Delta E = hf = ?$

frequency, $f = ?$

wavelength, $\lambda = ?$

Strategy:

Use $E_n = -13.6 \text{ eV} \times 1/n^2$
to determine the energy of each level.

To determine the energy difference, use the following equations.
$\Delta E = hf$, so $f = \dfrac{\Delta E}{h}$
and $c = \lambda f$, so $\lambda = \dfrac{c}{f}$

Calculations:

for $n = 2$, $E_2 = -13.6 \text{ eV} \times \dfrac{1}{2^2} = -3.40 \text{ eV}$

for $n = 1$, $E_1 = -13.6 \text{ eV} \times \dfrac{1}{1^2} = -13.6 \text{ eV}$

$\Delta E = E_1 - E_2 = -13.6 \text{ eV} - (-3.40 \text{ eV})$

$= -10.2$ eV Energy is emitted.

$$f = \frac{\Delta E}{h} = \frac{(10.2 \text{ eV})(1.60 \times 10^{-19} \text{ J/eV})}{6.63 \times 10^{-34} \text{ J·s}}$$

$$= 2.46 \times 10^{15} \text{ s}^{-1}$$

$$\lambda = \frac{c}{f} = \frac{3.00 \times 10^{8} \text{ m/s}}{2.46 \times 10^{15} \text{ s}^{-1}} = 1.22 \times 10^{-7} \text{ m}$$

$$= 122 \text{ nm}$$

Check Your Answer

- Are the units correct? Wavelengths are measured in multiples of meters. Energy is measured in electron volts.
- Are the signs correct? The energy is released as the electron moves from the second energy level to the first. The energy difference is thus negative.
- Is the magnitude realistic? Energy released in this transition produces light in the ultraviolet region below 400 nm.

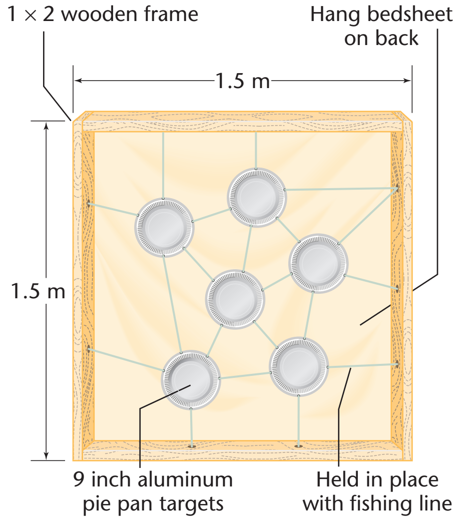
Shots in the Dark

Problem

Given that the atom is mostly empty space, how easy is it to hit a nucleus and cause atomic scattering?

Materials

3 dozen rubber stoppers
bedsheet (or blanket)
blindfold (or darkened goggles)
6 9-inch aluminum pie pans
4 1.5-meter 1 × 2 wood pieces
fishing line



Procedure

1. Construct the model according to the diagram.
2. Each student will be blindfolded, led to a position 3 m directly in front of the target area, and allowed to toss ten rubber stoppers (one at a time) into the target area. If a rubber stopper does not strike within the target area, the shooter should be told "too high," "too low," and so on and be given an extra rubber stopper.
3. Students will be able to hear the nuclear "hit" when the rubber stopper hits the target area. Only one hit will be counted on a single target.
4. When you have completed the lab, dispose of or recycle appropriate materials. Put away materials that can be reused.

Data and Observations

Student's Name	Number of Shots	Number of Hits

Analyze and Conclude

There are six circular targets within the target area. The ratio of hits to shots is represented by the following:

$$\frac{\text{hits}}{\text{shots}} = \frac{\text{total target area}}{\text{total model area}} = \frac{6\pi r^2}{\text{width} \times \text{height}}$$

1. **Analyzing Data** Use the class totals for shots and hits to calculate the total area for the six targets. Estimate the area for each target. Then calculate the radius for each target.
2. **Relating Concepts** The uncertainty for this experiment decreases with more shots. The percentage uncertainty is represented by the following:

$$\%\ \text{uncertainty} = \frac{(\text{shots})^{1/2}}{\text{shots}} \times 100$$

Find the uncertainty for your class.

Apply

1. A recent phone poll sampled 800 people. Estimate the uncertainty in the poll.

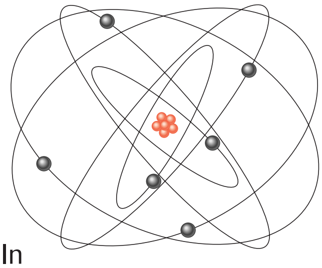
Practice Problems

1. According to the Bohr model, how many times larger is the orbit of a hydrogen electron in the second level than in the first?
2. You learned how to calculate the radius of the innermost orbit of the hydrogen atom. Note that all factors in the equation are constants with the exception of n^2. Use the solution to the Example Problem "Orbital Energy of Electrons in the Hydrogen Atom" to find the radius of the orbit of the second, third, and fourth allowable energy levels in the hydrogen atom.
3. Calculate the energies of the second, third, and fourth energy levels in the hydrogen atom.
4. Calculate the energy difference between E_3 and E_2 in the hydrogen atom. Find the wavelength of the light emitted. Which line in **Figure 28–8** is the result of this transmission?
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^{-9} m. If you use a baseball with diameter of 7.5 cm to represent the nucleus, how far away would the electron be?
6. 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 frequency of the photon emitted by the mercury atom?
   c. What is the wavelength of the photon emitted by the mercury atom?

28.1 Section Review

1. Which of these quantities is quantized: your height, the number of your siblings, or the mass of a sample of gas?
2. Why don't the electrons in Rutherford's nuclear model fly away from the nucleus?
3. Explain how energy is conserved when an atom absorbs a photon of light.
4. Review and analyze the Bohr model and the Rutherford nuclear model. How does the Bohr model differ from the Rutherford nuclear model? Critique each model, and explain their strengths and weaknesses.
5. **Critical Thinking** An emission spectrum can contain wavelengths produced when an electron moves from the third to the second level. Could you see this line in the absorption spectrum with your eyes or will you need a special detector? Explain.

28.2 The Quantum Model of the Atom



OBJECTIVES

- **Describe** the shortcomings of the Bohr model of the atom.
- **Describe** the quantum model of the atom.
- **Explain** how a laser works and **describe** the properties of laser light.

The Bohr model was a major contribution to the 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 free an electron completely from an atom. The value calculated by Bohr's team was in good agreement with experimental data. Using spectrographic data for many elements, Bohr and his team were able to determine the energy levels of the elements. 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.

The postulates that Bohr made could not be explained on the basis of known physics. For example, the theories of electromagnetism required that the accelerated particles radiate energy, which would cause the rapid collapse of the atom. In addition, Bohr could not explain the reason for the quantization of angular momentum. 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 he proposed that particles have wave properties just as light has particle properties. The wavelength of a particle with momentum mv was defined earlier to be $\lambda = h/mv$, and the angular momentum was defined to be $mvr = hr/\lambda$. The Bohr quantization condition, $mvr = nh/2\pi$, can be written in the following way.

Bohr Quantization Condition $\dfrac{hr}{\lambda} = \dfrac{nh}{2\pi}$ or $n\lambda = 2\pi r$

Note that the circumference of the Bohr orbit, $2\pi r$, is equal to a whole number multiple, n, of the wavelength of the electron, λ.

In 1926, Austrian physicist Erwin Schroedinger 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 picture 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 wave-particle nature of matter means 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 at a specific location. The most probable

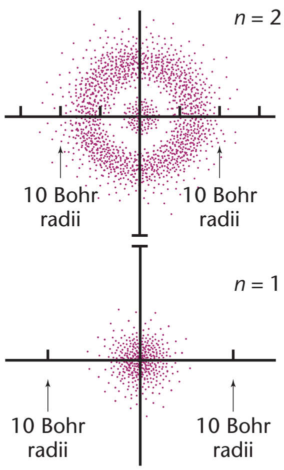


FIGURE 28–14 These plots show the probability of finding the electron in a hydrogen atom at any given location. The denser the points, the higher the probability of finding the electron.

distance of the electron from the nucleus in hydrogen is found to be the same as the radius of the Bohr orbit. The probability that the electron is at any radius can be calculated, and a three-dimensional plot can be constructed that shows regions of equal probability. The region in which there is a high probability of finding the electron is called the **electron cloud. Figure 28–14** shows a slice through the electron cloud for the two lowest states of hydrogen.

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, uses this model and has been extremely successful in predicting many details of the structure of the atom. 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 structure 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 are not otherwise available.

Quantum mechanics is also used to analyze the details of the emission and absorption of light by atoms. As a result of this theory, a new source of light was developed.

Lasers

Light emitted by an incandescent source has many wavelengths and travels in all directions. Light produced by an atomic gas consists of only a few different wavelengths, but it also is emitted in all directions. The light waves emitted by atoms at one end of a discharge tube are not necessarily in step with the waves from the other end. That is, the waves are not necessarily all at the same point in their cycle. Some will be in step, with the minima and maxima of the waves coinciding. Such light is called **coherent light.** Others will be out of step. Such light is called **incoherent light.** This is 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.

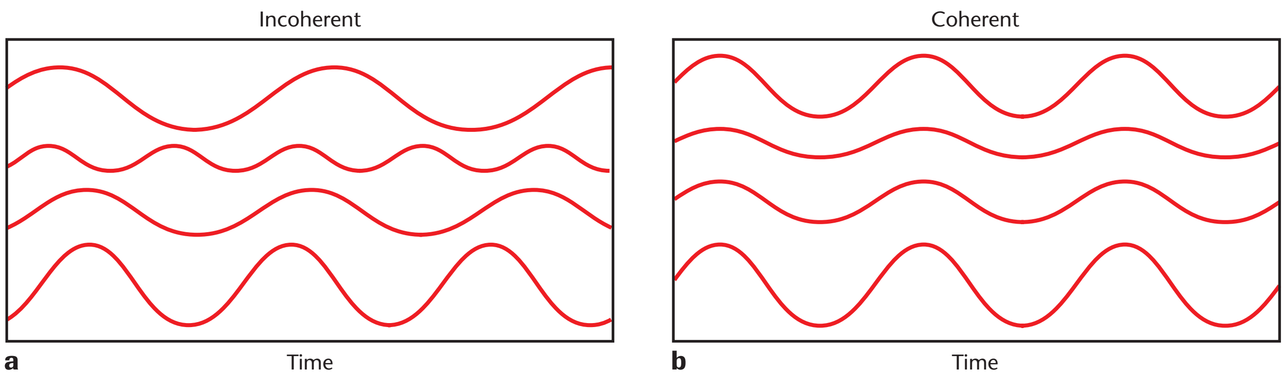
Pocket Lab

Bright Lines

Turn on a gas-discharge tube power supply attached to a gas tube so that the tube glows. **CAUTION:** *Do not touch any exposed metal when the power supply is turned on. Dangerous high voltages are present. Always turn off the power supply before changing gas tubes.* Turn off the room lights. Describe the color. Now look through a diffraction grating at the tube.

Analyze and Conclude Make a sketch of the results. Repeat this activity with a different gas tube. Explain the differences.

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



F.Y.I.

If enough atoms could be squeezed together to fill one cubic centimeter with each nucleus touching another nucleus, the total weight would be over 250 million metric tons.

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 is called spontaneous emission.

In 1917, Einstein considered what would happen to an atom already in an excited state that is struck by another photon of the same energy as the original photon. He showed that the atom will emit a photon of the same energy and move to a lower state. The photon that caused, or stimulated, the emission will not be affected. This process is called **stimulated emission.** The two photons leaving the atom not only will have the same wavelength, but also they will be in step, as shown in **Figure 28–16b.**

Either of the two photons can now strike other excited atoms, producing additional photons that are in step with the original photons. This process can continue, producing an avalanche of photons, all of the same wavelength and all having their maxima and minima at the same times.

To make this process happen, certain conditions must be met. First, of course, there must be other atoms in the excited state. Second, the photons must be collected so that they strike the excited atoms. A device that fulfills both these conditions was invented in 1959 and is called a **laser.** The word *laser* is an acronym. It 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.

The atoms in a laser can be put into the excited state, or pumped, in different ways, as outlined in **Figure 28–17.** An intense flash of light with a wavelength shorter than that of the laser can pump the atoms. The more energetic photons produced by the flash collide with and excite the lasing atoms. One atom decays to a lower energy state, starting the avalanche. As a result, a brief flash or pulse of laser light is emitted. Alternatively, a continuous electric discharge such as that in a neon sign can be used to put atoms in the excited state. The laser light resulting from this process is continuous rather than pulsed. The helium-neon

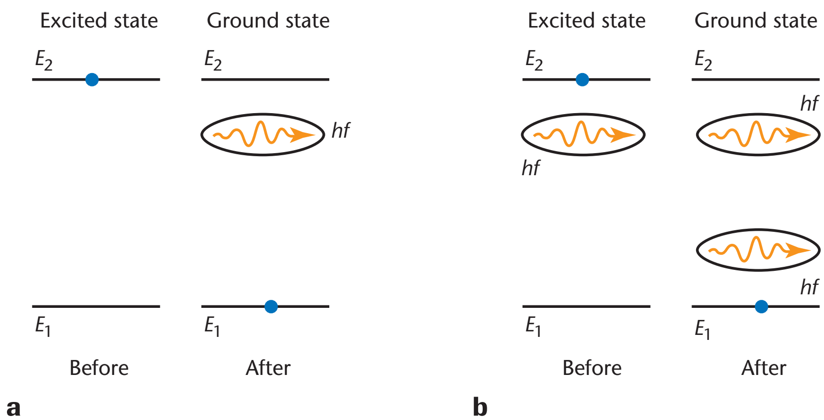


FIGURE 28–16 The spontaneous emission of a photon with energy hf when an electron in an atom drops from an excited state E_2 to the ground state E_1 is shown **(a).** The stimulated emission of a photon when an excited atom is struck by a photon with energy hf is shown **(b).** For both, $hf = E_2 - E_1$.

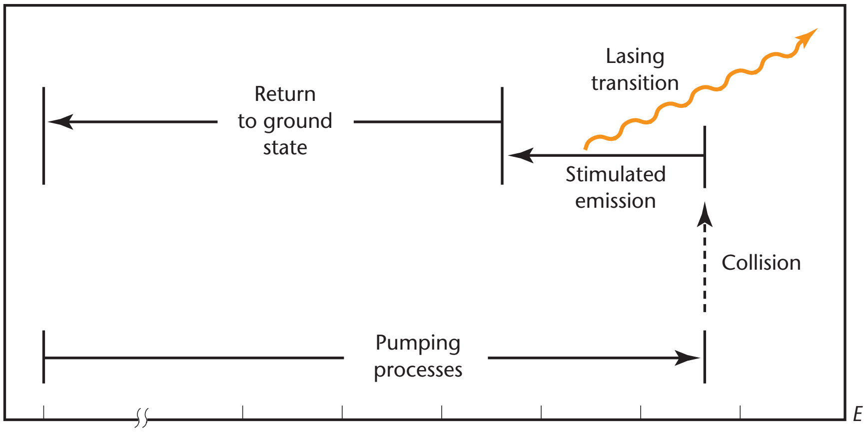


FIGURE 28–17 A photon striking an atom in the excited state stimulates it to make a transition to a lower state and emit a second coherent photon.

lasers often seen in science classrooms are continuous lasers. An electric discharge excites the helium atoms. They collide with the neon atoms, pumping them to an excited state and causing them to lase.

The photons emitted by atoms are collected by placing a glass tube containing the atoms between two parallel mirrors. One mirror is 100 percent reflective and will reflect all the light hitting it while the other is only partially reflective and will allow only about one percent of the light to pass through. When a photon strikes an atom in the excited state, it stimulates the atom to make a transition to the lower state. Thus, two photons leave the atom. These photons can strike other atoms and produce more photons, thereby starting the avalanche. Photons that are directed toward the ends of the tube will be reflected back into the gas by the mirrors. The reflected photons reinforce one another with each pass between the mirror and build to a high intensity. The photons that exit the tube through the partially reflecting mirror produce the laser beam. This is shown in **Figure 28–18.**

Laser light is highly directional because of the parallel mirrors. The light beam is very small, typically only about 1/2 mm in diameter, so the light is very intense. The light is 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. Because all the stimulated photons are emitted in step with the photons that struck the atoms, laser light is coherent light.

Pocket Lab

Laser Diffraction

Caution: *Avoid staring directly into the laser beam or at bright reflections.* A diffraction grating separates light from discharge tubes into individual wavelengths. Predict what will happen when you shine a laser light through a diffraction grating. Shine the laser at a sheet of white paper about 1 foot away. Then place the diffraction grating next to the laser to see what happens.

Analyze and Conclude
Describe and explain your results. Predict how your results would be similar and different with a green laser light.

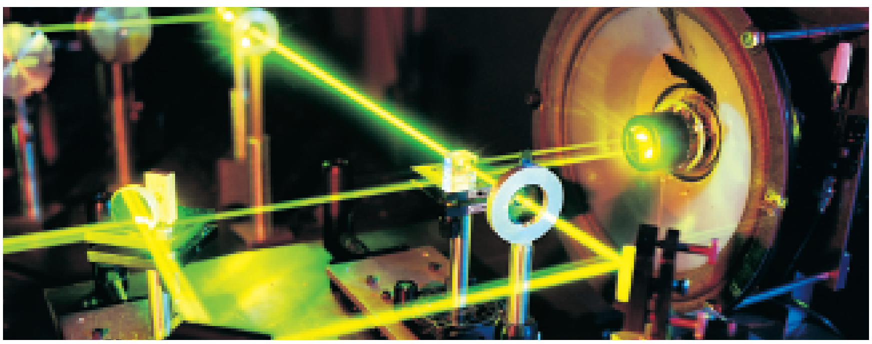


FIGURE 28–18 A laser produces a beam of coherent light.

HELP WANTED

LASER TECHNICIAN

Will produce, test, operate, and/or repair lasers. May work in a hospital, the fiber optics industry, a research lab, a manufacturing plant, the military, or on space projects or construction sites. Some laser technicians create holograms for charge cards, diagnosis of medical problems, or identification of flaws in machine parts. Position requires a two-year degree in laser technology. Four-year degree preferred. For more information, contact:

Laser Institute of America
13501 Ingenuity Drive
Orlando, FL 32826

Many substances—solids, liquids, and gases—can be made to lase in this way. Most produce laser light at only one wavelength. For example, red is produced by a neon laser, blue by a helium-cadmium laser, and green by an argon-ion laser. The light from some lasers, however, can be tuned, or adjusted, over a range of wavelengths. **Table 28–1** shows the types of wavelengths produced by some common lasers.

All lasers are very inefficient. No more than one percent of the electrical energy delivered to a laser is converted to light energy. Despite this 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. Surveyors use laser beams for this reason. Laser beams are also used to check the straightness of long tunnels and pipes. When astronauts visited the moon, they left a mirror which was used by scientists to reflect a laser beam from Earth. The distance between Earth and the moon was thus accurately determined.

Laser light is used in fiber optics as well. A fiber uses total internal reflection to transmit light over many kilometers with little loss. The laser is switched on and off rapidly, transmitting information through the fiber. In many cities, optical fibers have replaced copper wires for the transmission of telephone calls, computer data, and even television pictures.

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

The concentrated power of laser light is used in a variety of ways. In medicine, lasers can be used to repair the retina in an eye. Lasers also can be used in surgery in place of a knife to cut flesh with little loss of

TABLE 28–1

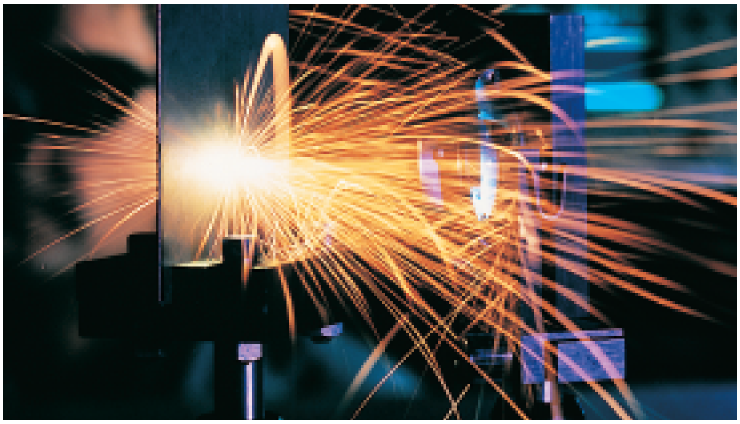
Common Lasers

Medium	Wavelength (nm)	Type
Nitrogen	337 (UV*)	Pulsed
Helium-cadmium	441.6	Continuous
Argon ion	476.5, 488.0, 514.5	Continuous
Krypton ion	524.5	Continuous
Neon	632.8	Continuous
Gallium aluminum arsenide	680	Continuous
Ruby	694.3	Pulsed
Gallium arsenide+	840–1350 (IR*)	Continuous
Neodymium	1040 (IR*)	Pulsed
Carbon dioxide	10 600 (IR*)	Continuous

* UV means ultraviolet, IR means infrared.

+ The wavelength of gallium arsenide depends on temperature.

FIGURE 28–19 Because laser light is so concentrated and directed, lasers can be used in industry to cut and weld steel.



Physics & Technology

All Aglow!

Fluorescence is a phenomenon that occurs when certain substances are exposed to radiation. As a result of their being stimulated by the radiation, fluorescent substances reemit the radiation to produce that glow-in-the dark quality found on various decals and stickers; T-shirts, tennis shoes and other pieces of clothing; light-switch plates; and children's books, puzzles, and toys; just to name a few examples.

In addition to those just named above, fluorescent substances have many practical applications. In medicine, for example, the fluorescent protein found in jellyfish can be used to mark and subsequently change, or mutate, the amino acids that make up a protein. Staining chromosomes with fluorescent dyes allows physicians to identify DNA abnormalities that cause diseases such as Down's syndrome and leukemia. Another medical application of fluorescence involves using the glowing substances to count proteins and other chemicals to provide early warning signs of certain types of cancers.

Fluorescence also has its importance in the art world. Jewelers and gemologists have used fluorescent techniques to study the Hope diamond, the largest known gem of its kind at 45.5 carats. Determining the conditions under which this diamond formed could lead to the discovery of other such diamonds. Diamonds, the hardest known natural substances, have uses in dentistry as polishers, and in jewelry and industry as industrial drills and abrasives.

Fluorescence can sometimes solve or prevent certain crimes. Forensic labs, for example, use fluorescent powders to lift fingerprints. Glow-in-the-dark decals are used by some cities to deter auto theft, which generally occurs between midnight and 6:00 A.M. Users of the decals sign a statement saying that they normally don't drive between those hours. Thus, if a car with these special fluorescent stickers is seen on the road, police are alerted to a possible theft.

Thinking Critically Think about what you've just read. Compile a list of at least five other practical applications of fluorescence. Explain the usefulness of each of your ideas.

FIGURE 28–20 When a hologram is made on film, a laser beam is split into two parts by a half-silvered mirror. Interference occurs on the film as the direct laser light meets laser light reflected off an object. The interference of both beams of light allows the film to record both the intensity and phase of light from the object.

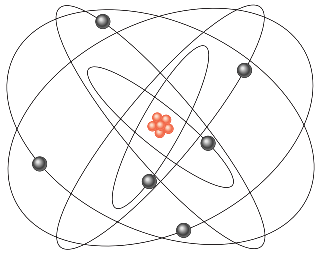
blood. In industry, lasers are used to cut materials such as steel, as shown in **Figure 28–19,** and to weld materials together. In the future, lasers may be able to produce nuclear fusion to create an almost inexhaustible energy source.

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

28.2 Section Review

1. Which of the lasers in **Table 28–1** emits the most red light (visible light with the longest wavelength)? Which emits in the blue region of the spectrum?
2. Could green light be used to pump a red laser? Why could red light not be used to pump a green laser?
3. Why does the Bohr model of the atom conflict with the uncertainty principle while the quantum model does not?
4. Research and evaluate how lasers have impacted society. Describe the connection between lasers and future careers in physics.
5. **Critical Thinking** Suppose that an electron cloud were to get so small that the atom was almost the size of the nucleus. Use the uncertainty principle to explain why this would take a tremendous amount of energy.

CHAPTER 28 REVIEW



Summary

Key Terms

28.1

- alpha (α) particle
- scintillation
- nuclear model
- emission spectrum
- spectroscope
- absorption spectrum
- spectroscopy
- energy level
- ground state
- excited state
- Bohr model
- quantum number

28.2

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

28.1 The Bohr Model of the Atom

- Ernest Rutherford directed positively charged, high-speed alpha particles at thin metal foils. By studying the paths of the reflected particles, he showed that atoms are mostly empty space with a tiny, massive, positively-charged nucleus at the center.
- The spectra produced by atoms of an element can be used to identify that element.
- If white light passes through a gas, the gas absorbs the same wavelengths that it would emit if it were excited. If light leaving the gas goes through a prism, an absorption spectrum is visible.
- In the model of the atom developed by Niels Bohr, the electrons can have only certain energy levels.
- In the Bohr model of the atom, electrons can make transitions between energy levels. As they do, they emit or absorb electromagnetic radiation.
- The frequency and wavelength of absorbed and emitted radiation can be calculated for the hydrogen atom using the Bohr model.

28.2 The Quantum Model of the Atom

- The quantum mechanical model of the atom cannot be visualized easily. Only the probability that an electron is at a specific location can be calculated.
- 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.

Key Equations

28.1

$$\Delta E = hf \qquad r_n = \frac{h^2n^2}{4\pi^2Kmq^2} \qquad E_n = -13.6\ \text{eV} \times \frac{1}{n^2}$$

28.2

$$n\lambda = 2\pi r$$

Reviewing Concepts

Section 28.1

1. Describe how Rutherford determined that the positive charge in an atom is concentrated in a tiny region rather than spread throughout the atom.
2. How does the Bohr model explain why the absorption spectrum of hydrogen contains exactly the same frequencies as its emission spectrum?
3. Review and critique the planetary model atom. What are some of the problems with a planetary model of the atom?
4. Analyze and critique the Bohr model of the atom. What three assumptions did Bohr make in developing his model?
5. Explain how line spectra from gas-discharge tubes are produced.
6. How does the Bohr model account for the spectra emitted by atoms?
7. Explain why line spectra produced by hydrogen gas-discharge tubes are different from those produced by helium gas-discharge tubes.

Section 28.2

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

9. A device similar to a laser that emits microwave radiation is called a maser. What words are likely to make up this acronym?

10. What properties of laser light led to its use in light shows?

Applying Concepts

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

12. The northern lights are the result of high-energy particles coming from the sun and striking atoms high in Earth's atmosphere. If you looked at these lights through a spectrometer, would you expect to see a continuous or line spectrum? Explain.

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

14. Suppose you wanted to explain quantization to a younger brother or sister. Would you use money or water as an example? Explain.

15. A photon with energy of 6.2 eV enters a mercury atom in the ground state. Will it be absorbed by the atom? See **Figure 28–21.** Explain.

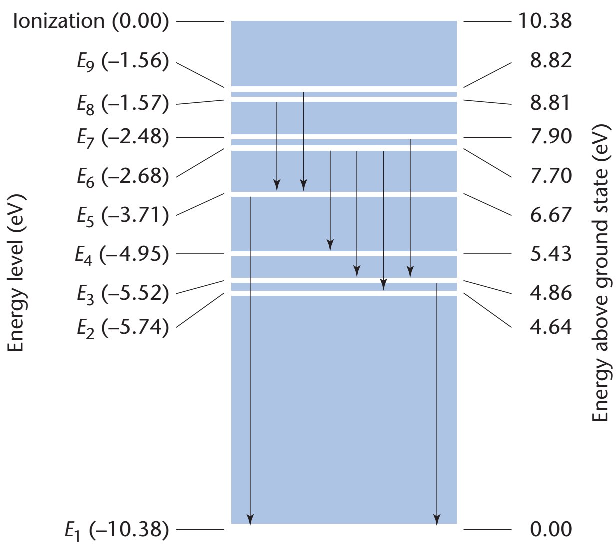


FIGURE 28–21

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

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

18. When electrons fall from higher energy levels to the third energy level within hydrogen atoms, are the photons that are emitted infrared, visible, or ultraviolet light? Explain.

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

20. You have a laser that emits a red light, a laser that emits a green light, and a laser that emits a blue light. Which laser, produces photons with the highest energy?

Problems

Section 28.1

*See **Figure 28–21** for Problems 21, 22, and 23.*

21. A mercury atom is in an excited state when its energy level is 6.67 eV above the ground state. A photon of energy 2.15 eV strikes the mercury atom and is absorbed by it. To what energy level is the mercury atom raised?

22. 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 electron dropped down to the E_2 energy level instead?

23. 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.

a. What is the energy of the photon?

b. What is the photon's frequency?

24. 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?

25. Calculate the radius of the orbital associated with the energy levels E_5 and E_6 of the hydrogen atom.
26. What energies are associated with a hydrogen atom's energy levels E_2, E_3, E_4, E_5, and E_6?
27. Using the values that are calculated in problem 26, calculate the following energy differences for a hydrogen atom.
    - **a.** $E_6 - E_5$
    - **b.** $E_6 - E_3$
    - **c.** $E_4 - E_2$
    - **d.** $E_5 - E_2$
    - **e.** $E_5 - E_3$
28. Use the values from problem 27 to determine the frequencies of the photons emitted when an electron in a hydrogen atom makes the level changes listed.
29. Determine the wavelengths of the photons of the frequencies that you calculated in problem 28.
30. Determine the frequency and wavelength of the photon emitted when an electron drops
    - **a.** from E_3 to E_2 in an excited hydrogen atom.
    - **b.** from E_4 to E_3 in an excited hydrogen atom.
31. What is the difference between the energies of the E_4 and E_1 energy levels of the hydrogen atom?
32. From what energy level did an electron fall if it emits a photon of 94.3 nm wavelength when it reaches ground state within a hydrogen atom?
33. For a hydrogen atom in the $n = 3$ Bohr orbital, find
    - **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.
34. A hydrogen atom has its electron 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?

Section 28.2

35. Gallium arsenide lasers are used in CD players. If such a laser emits at 840 nm, what is the difference in eV between the two lasing energy levels?
36. A carbon dioxide laser emits very high-power infrared radiation. What is the energy difference in eV between the two lasing energy levels?

Extra Practice For more practice solving problems, go to Extra Practice Problems, Appendix B.

Critical Thinking Problems

37. The four brightest lines in the mercury spectrum have wavelengths of 405 nm, 436 nm, 546 nm, and 578 nm. What are the differences in energy levels for each of these lines?
38. After the emission of these visible photons, the mercury atom continues to emit photons until it reaches the ground state. From inspection of **Figure 28–21,** determine whether or not any of these photons would be visible. Explain.

Going Further

Measuring Energy Levels 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 1/7 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?

Essay Research and describe the history of the model of the atom. Be sure to include the strengths and weaknesses of each model.

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