32

CHAPTER 32

The Atom and the Quantum

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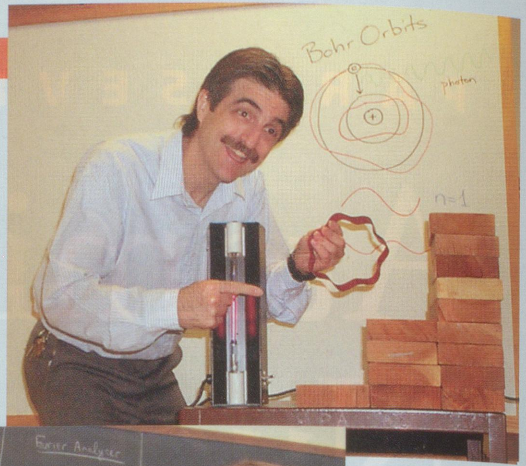

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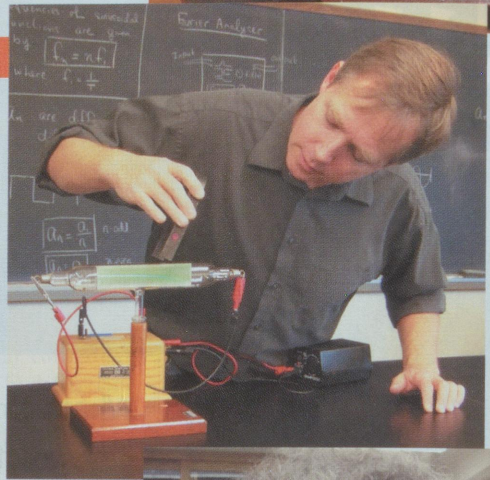

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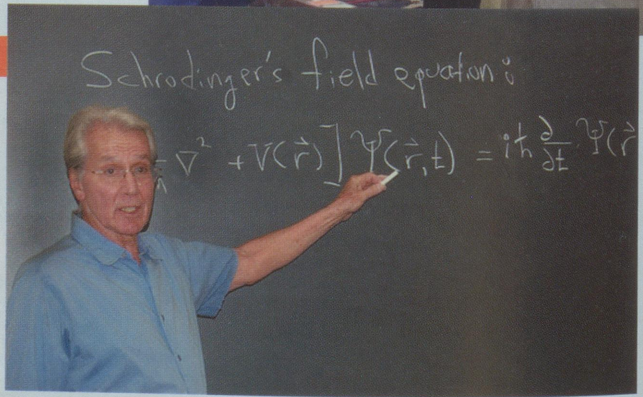


These four physicists, who excel in their teaching, bring quantum physics down to Earth for their students. 1 David Kagan uses a strip of corrugated plastic in class to model an orbiting electron. The stacked wood blocks model electron energy levels. 2 Roger King uses a magnet to bend an electron beam in a Crookes tube. 3 Dean Zollman investigates nuclear properties with a modern version of Rutherford's scattering experiment. 4 Professor and author Art Hobson teaches one of his specialties, quantum physics.

Niels Bohr was born in Copenhagen, Denmark, in 1885. His father, a Lutheran, was a professor of physiology at the University of Copenhagen. His mother came from a wealthy Jewish family prominent in Danish banking and parliamentary circles. His brother, Harald Bohr, was a mathematician and Olympic soccer player on the Danish national team. Niels was passionate about soccer as well, and the two brothers played in a number of national matches in Copenhagen.

Niels Bohr

Bohr earned his physics doctorate in Denmark in 1911. He then worked for a time in the laboratory of J. J. Thomson, the discoverer of the electron, at Trinity College in Cambridge, England, before continuing his research under Ernest Rutherford at the University of Manchester, also in England. Rutherford had just discovered that a tiny, positively charged nucleus sits at the center of every atom, surrounded, presumably, by Thomson's electrons. Bohr pondered this new picture of the atom and added quantum principles to it. He published his model of atomic structure in 1913, in which electrons travel only in certain orbits around the atomic nucleus, and the atom emits light when electrons make "quantum jumps" from one orbit to another. His theory brilliantly accounted for the observed spectral lines of hydrogen and the so-called Balmer series as well as other series.

Bohr won the Nobel Prize in Physics in 1922 for his work on the quantum theory of atoms, a year after Albert Einstein won the Nobel Prize for his work on the photoelectric effect. After quantum theory evolved and matured in the mid-1920s, Einstein had great reservations about its probabilistic nature, much preferring the determinism of classical physics. He and Bohr debated these two views of physics throughout their lives, always maintaining the greatest respect for each other.

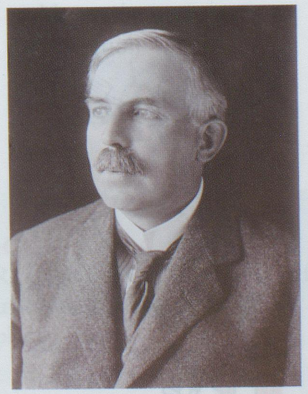
Because Bohr's mother was Jewish, he was in danger in Nazi-occupied Denmark during World War II. In 1943, shortly before an impending arrest, he escaped to Sweden with his family. The Allies, recognizing Bohr's importance, flew him from Sweden to London tucked in the bomb bay of an unarmed Mosquito bomber. Because Bohr forgot to put on his oxygen mask, he passed out. Fortunately, the pilot, sensing that something was wrong when Bohr didn't respond to intercom messages, descended to a lower altitude and delivered a still-living passenger to London. Bohr reportedly said that he had slept like a baby during the flight. He then went to the United States to work on the U.S. Manhattan Project at the top-secret Los Alamos laboratory in New Mexico. For security reasons he was assigned the name of Nicholas Baker during the project.

After the war, Bohr returned to Copenhagen, advocating the peaceful use of nuclear energy and the sharing of nuclear information. When awarded the Order of the Elephant by the Danish government, he designed his own coat of arms, which featured a symbol of yin and yang with the Latin motto *contraria sunt complementa*: "opposites are complementary."

Bohr's son Aage went on to become a very successful physicist and, like his father, won a Nobel Prize in Physics, his in 1975. Niels Bohr died in Copenhagen in 1962. Much of this chapter involves his view of the underlying physics of nature.

32.1 Discovery of the Atomic Nucleus

Half a dozen years after Einstein announced the photoelectric effect, the New Zealand–born British physicist Ernest Rutherford oversaw his now famous gold-foil experiment.[1] This significant experiment showed that the atom is mostly empty space, with most of its mass concentrated in the central region—the **atomic nucleus**.



Ernest Rutherford (1871–1937)

In Rutherford's experiment, a beam of positively charged particles (alpha particles) from a radioactive source was directed through a sheet of extremely thin gold foil. Because alpha particles are thousands of times more massive than electrons, it was expected that the stream of alpha particles would not be impeded as it passed

[1]Why "oversaw"? To indicate that more investigators than Rutherford were involved in this experiment. The widespread practice of elevating a single scientist to the position of sole investigator, which seldom is the case, too often denies the involvement of other investigators. There's substance to the saying, "There are two things more important to people than sex and money—*recognition and appreciation*."

Rutherford later related that the discovery of alpha particles rebounding backward was the most incredible event of his life—as incredible as if a 15-inch cannon shell rebounded from a piece of tissue paper.

through the sea of electrons. This was indeed observed—for the most part. Nearly all alpha particles passed through the gold foil with little or no deflection and produced a spot of light when they hit a fluorescent screen beyond the foil. But some particles were deflected from their straight-line paths as they emerged. A few alpha particles were widely deflected, and a small number were even scattered backward! These alpha particles must have hit something relatively massive—but what? Rutherford reasoned that the undeflected particles traveled through empty space in regions of the gold foil, while the small number of deflected particles were repelled from extremely dense, positively charged central cores. Each atom, he concluded, must contain one of these cores, which he named the *atomic nucleus* .

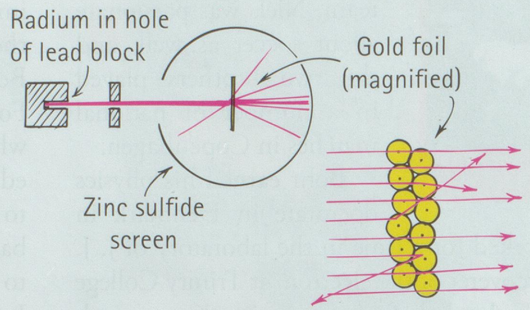


FIGURE 32.1
The occasional large-angle scattering of alpha particles from gold atoms led Rutherford to the discovery of the very massive small nuclei at their centers.

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CHECK POINT

1. What convinced Rutherford that the gold foil was mostly empty space?
2. What convinced him that particles in the gold foil were massive?

CHECK YOUR ANSWERS

1. Finding that most alpha particles were undeflected indicated much empty space.
2. Finding that some alpha particles bounced backward indicated the presence of particles more massive than the alpha particles.

You'll note that much of this chapter is a synthesis of chapters already covered—hopefully a "tying-it-together" chapter.

32.2 Discovery of the Electron

Surrounding the atomic nucleus are electrons. The name *electron* comes from the Greek word for amber, a brownish-yellow fossil resin studied by the early Greeks. They found that when amber was rubbed by a piece of cloth, it attracted such things as bits of straw. This phenomenon, known as the *amber effect*, remained a mystery for almost 2000 years. In the late 1500s, William Gilbert, Queen Elizabeth's physician, found other materials that behaved like amber, and he called them "electrics." The concept of electric charge awaited experiments by the American scientist-statesman Benjamin Franklin nearly two centuries later. Recall from Chapter 22 that Franklin experimented with electricity and postulated the existence of an electric fluid that could flow from place to place. An object with an excess of this fluid he called *electrically positive*, and one with a deficiency of the fluid he called *electrically negative*. The fluid was thought to attract ordinary matter but to repel itself. Although we no longer talk about electric fluid, we still follow Franklin's lead in how we define positive and negative electricity. Franklin's 1752 experiment with the kite in the lightning storm showed that lightning is an electrical discharge between clouds and the ground. This discovery told him that electricity is not restricted to solid or liquid objects and that it can travel through a gas.

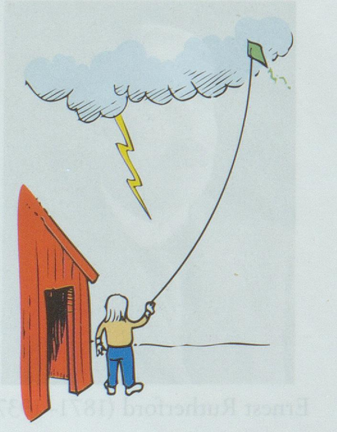

FIGURE 32.2
Franklin's kite-flying experiment.

Franklin's experiments later inspired other scientists to produce electric currents through various dilute gases in sealed glass tubes. Among these, in the 1870s, was Sir William Crookes, an unorthodox English scientist who believed he could communicate with the dead. He is better remembered for his "Crookes tube," a sealed glass tube that contains gas under very low pressure and has electrodes inside the tube near each end (the forerunner to today's neon signs). The gas glowed when the electrodes were connected to a voltage source (such as a battery). Different gases glowed with different colors. Experiments conducted with tubes containing metal slits and plates showed that the gas was made to glow by some sort of a "ray" emerging from the negative terminal (the *cathode*). Slits could make the ray narrow and plates could prevent the ray from reaching the positive terminal (the *anode*). The apparatus was named the *cathode-ray tube*, or CRT (Figure 32.3). When electric charges were brought near the tube, the ray was deflected. It bent toward positive charges and away from negative charges. The ray was also deflected by the presence of a magnet. These findings indicated that the ray consisted of negatively charged particles.

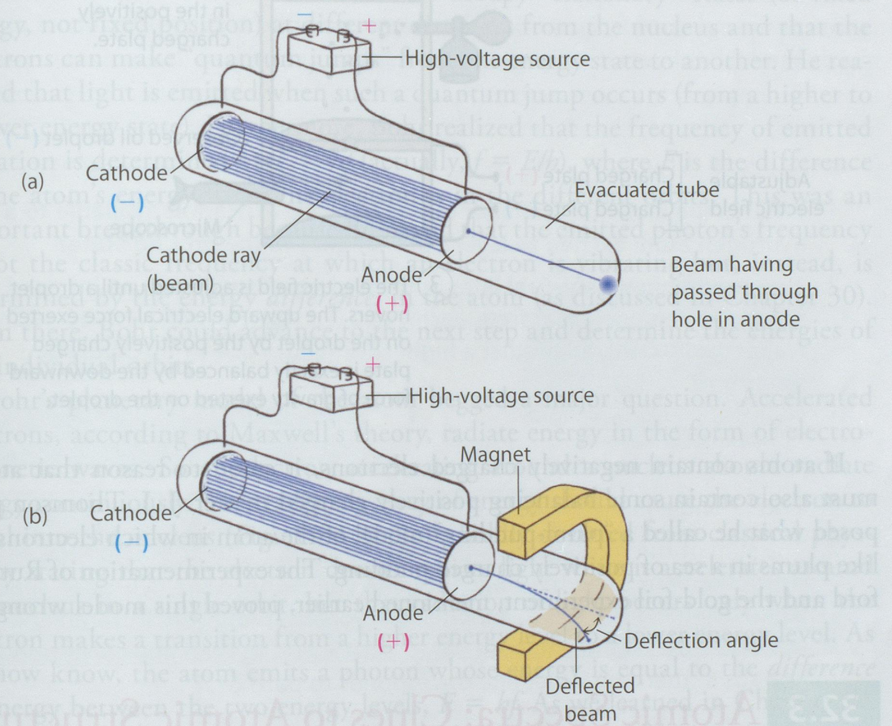

FIGURE 32.3
(a) A simple cathode-ray tube. The small hole in the anode permits the passage of a narrow beam that strikes the end of the tube, producing a glowing dot on the glass.
(b) The cathode ray is deflected by a magnetic field.

In 1897, the English physicist Joseph John Thomson ("J. J.," as his friends called him) showed that the cathode rays were particles smaller and lighter than atoms. He created narrow beams of cathode rays and measured their deflection in electric and magnetic fields. Thomson reasoned that the amount of the beams' deflection depended on the mass of the particles and their electric charge. How? The greater each particle's mass, the greater the inertia and the less the deflection. The greater each particle's charge, the greater the force and the greater the deflection. The greater the speed, the less the deflection.

From careful measurements of the deflection of the beam, Thomson succeeded in calculating the mass-to-charge ratio of the cathode-ray particle, which was named the **electron**. All electrons are identical; they are copies of one another. For establishing the existence of the electron, J. J. Thomson was awarded the Nobel Prize in Physics in 1906.

In 1909, a dozen years after Thomson's mass-to-charge measurements, American physicist Robert Millikan carried out an experiment that enabled him to calculate the numerical value of a single unit of electric charge. In his experiment,

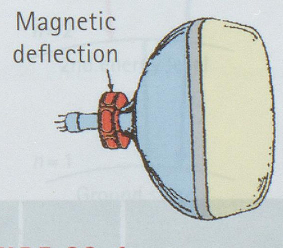

FIGURE 32.4
CRTs like this were common before flat-screen TV displays largely replaced them.

Millikan sprayed tiny oil droplets into a chamber between electrically charged plates—into an *electric field*. When the field was strong, some of the droplets moved upward, indicating that they carried a very slight negative charge. Millikan adjusted the field so that the droplets hovered motionless. He knew that the downward force of gravity on the droplets was exactly balanced by the upward electrical force. Investigation showed that the charge on each drop was always some multiple of a single very small value, which he proposed to be the fundamental unit of charge carried by each electron. Using this value and the mass-to-charge ratio discovered by Thomson, Millikan calculated the mass of an electron to be about 1/2000 the mass of the lightest known atom, hydrogen. This confirmed Thomson's supposition that the electron is a lightweight and it established the quantum unit of charge. For his work in physics, Millikan received the 1923 Nobel Prize.

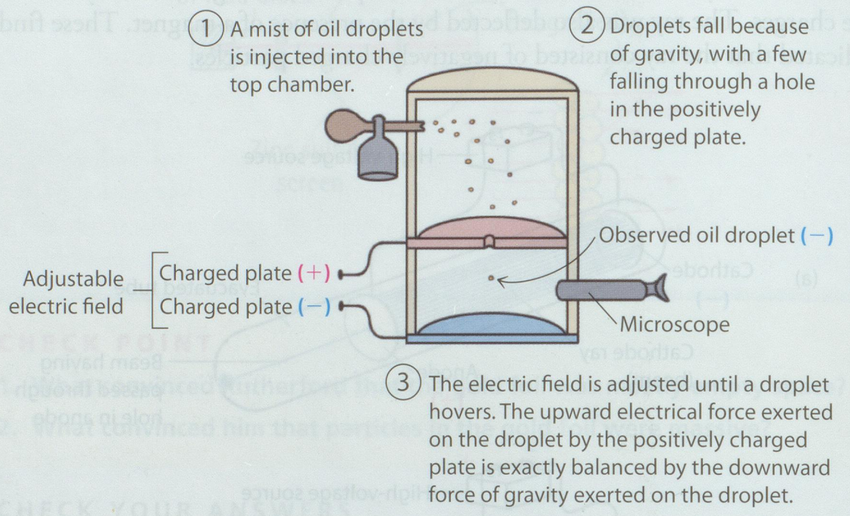

FIGURE 32.5
Millikan's oil-drop experiment for determining the charge on the electron.

The electron was the first of many fundamental particles that were discovered.

If atoms contain negatively charged electrons, it stood to reason that atoms must also contain some balancing positively charged matter. J. J. Thomson proposed what he called a "plum-pudding" model of the atom in which electrons are like plums in a sea of positively charged pudding. The experimentation of Rutherford and the gold-foil experiment, mentioned earlier, proved this model wrong.

32.3 Atomic Spectra: Clues to Atomic Structure

During the period of Rutherford's experiments, chemists were using the spectroscope (discussed in Chapter 30) for chemical analysis, while physicists were busily occupied trying to find order in the confusing arrays of spectral lines. It had long been known that the lightest element, hydrogen, has a far more orderly spectrum than the other elements (Figure 32.6). An important sequence of lines in the hydrogen spectrum starts with a line in the red region, followed by one in the blue, then by several lines in the violet, and many in the ultraviolet. The spacing between successive lines becomes smaller and smaller from the first in the red to the last in the ultraviolet, until the lines become so close that they seem to merge. A Swiss schoolteacher, Johann Jakob Balmer, first expressed the wavelengths of these lines in a single mathematical formula in 1884. Balmer, however, was unable to provide a reason why his formula worked so successfully. His guess that his formula could be extended to predict other lines of hydrogen proved to be correct, leading to the prediction of lines that had not yet been measured.

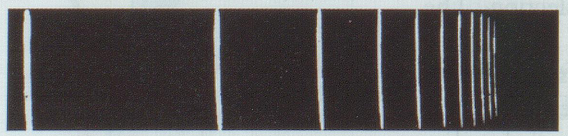

FIGURE 32.6
A portion of the hydrogen spectrum. Each line, an image of the slit in the spectroscope, represents light of a specific frequency emitted by hydrogen gas when excited (higher frequency is to the right).

Another regularity in atomic spectra was found by the Swedish physicist and mathematician Johannes Rydberg. He noticed that the frequencies of lines in certain series in other elements followed a formula similar to Balmer's and that the sum of the frequencies of two lines in such series often equals the frequency of a third line. This relationship was later advanced as a general principle by the Swiss physicist Walter Ritz and is called the **Ritz combination principle**. It states that the spectral lines of any element include frequencies that are either the sum or the difference of the frequencies of two other lines. Like Balmer, Ritz was unable to offer an explanation for this regularity. However, these regularities were the clues that the Danish physicist Niels Bohr used to understand the structure of the atom itself.

32.4 Bohr Model of the Atom

In 1913, Bohr applied the quantum theory of Planck and Einstein to the nuclear atom of Rutherford and formulated the well-known planetary model of the atom.[2] Bohr reasoned that electrons occupy "stationary" states (of fixed energy, not fixed position) at different distances from the nucleus and that the electrons can make "quantum jumps" from one energy state to another. He reasoned that light is emitted when such a quantum jump occurs (from a higher to a lower energy state). Furthermore, Bohr realized that the frequency of emitted radiation is determined by $E = hf$ (actually, $f = E/h$), where E is the difference in the atom's energy when the electron is in the different orbits. This was an important breakthrough because Bohr said that the emitted photon's frequency is not the classic frequency at which an electron is vibrating but, instead, is determined by the energy *differences* in the atom (as discussed in Chapter 30). From there, Bohr could advance to the next step and determine the energies of the individual orbits.

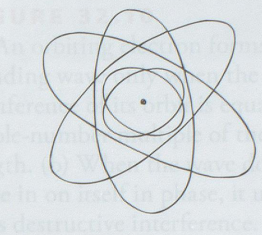

FIGURE 32.7
The Bohr model of the atom. Although this model is very oversimplified, it is still useful in understanding light emission.

Bohr's planetary model of the atom begged a major question. Accelerated electrons, according to Maxwell's theory, radiate energy in the form of electromagnetic waves. So an electron accelerating around a nucleus should radiate energy continuously. This radiating away of energy should cause the electron to spiral into the nucleus (Figure 32.8). Bohr boldly deviated from classical physics by stating that the electron doesn't radiate light while it accelerates around the nucleus in a single orbit, but that radiation of light occurs only when the electron makes a transition from a higher energy level to a lower energy level. As we now know, the atom emits a photon whose energy is equal to the *difference* in energy between the two energy levels, $E = hf$. As we learned in Chapter 30, the frequency of the emitted photon, its color, depends on the size of the jump. So the quantization of light energy neatly corresponds to the quantization of electron energy.

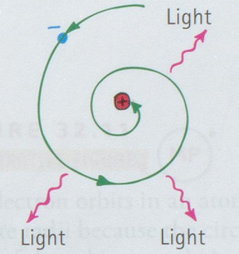

FIGURE 32.8
According to classical theory, an electron accelerating around its orbit should continuously emit radiation. This loss of energy should cause it to spiral rapidly into the nucleus. But this does not happen.

Bohr's views, as outlandish as they seemed at the time, explained the regularities found in atomic spectra. Bohr's explanation of the Ritz combination principle is shown in Figure 32.9. If an electron is raised to the third energy level, it can return to its initial level either by a single jump from the third to the first level or by a double jump, first to the second level and then to the first level. These two return paths will produce three spectral lines. Note that the sum of the two energy jumps along paths A and B is equal to the single energy jump along path C. Since frequency is proportional to energy, the frequencies of light emitted along paths A and B, when added, equal the frequency of light emitted when the transition is along path C. Now we can see why the sum of two frequencies in the spectrum is equal to a third frequency in the spectrum.

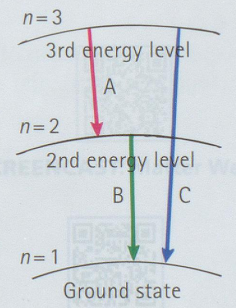

FIGURE 32.9
Three of many energy levels in an atom. An electron jumping from the third level to the second level (red A), and one jumping from the second level to the ground state (green B). The sum of the energies (and the frequencies) for these two jumps equals the energy (and the frequency) of the single jump from the third level to the ground state (blue C).

[2]This model, like most models, has major defects because the electrons do not revolve in planes as planets do. The model was revised; "orbits" became "shells" and "clouds." We use the term *orbit* because it was, and still is, commonly used. Electrons are not just bodies, like planets, but rather behave like waves concentrated in certain parts of the atom.

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Bohr was able to account for X-rays in heavier elements, showing that they are emitted when electrons jump from outer to innermost orbits. He predicted X-ray frequencies that were later experimentally confirmed. Bohr was also able to calculate the "ionization energy" of a hydrogen atom—the energy needed to knock the electron out of the atom completely. This also was verified by experiment.

Using measured frequencies of X-rays as well as visible, infrared, and ultraviolet light, scientists could map the energy levels of all the atomic elements. Bohr's model had electrons orbiting in neat circles (or ellipses) arranged in groups or shells. This model of the atom accounted for the general chemical properties of the elements. It also led to the discovery of hafnium.

Bohr solved the mystery of atomic spectra while providing an extremely useful model of the atom. He was quick to stress that his model was to be interpreted as a crude beginning, and the picture of electrons whirling about the nucleus like planets about the Sun was not to be taken literally (to which popularizers of science paid no heed). His sharply defined orbits were conceptual representations of an atom whose later description involved waves—quantum mechanics. His ideas of quantum jumps and frequencies being proportional to energy differences remain part of today's modern theory.

CHECK POINT

1. What is the maximum number of paths for de-excitation available to a hydrogen atom excited to level 3 in changing to the ground state?
2. Two predominant spectral lines in the hydrogen spectrum, an infrared one and a red one, have frequencies of 2.7×10^{14} Hz and 4.6×10^{14} Hz, respectively. Can you predict a higher-frequency line in the hydrogen spectrum?

CHECK YOUR ANSWERS

1. Two (a single jump and a double jump), as shown in Figure 32.9.
2. The sum of the frequencies is $2.7 \times 10^{14} + 4.6 \times 10^{14} = 7.3 \times 10^{14}$ Hz, which happens to be the frequency of a violet line in the hydrogen spectrum. Using Figure 32.9 as a model, can you see that if the infrared line is produced by a transition similar to path A and the red line corresponds to path B, then the violet line corresponds to path C?

32.5 Explanation of Quantized Energy Levels: Electron Waves

VIDEO: Electron Waves

Back in Chapter 11 we discussed the different sizes of atoms; this was shown in Figure 11.10. In Chapter 30 we discussed atomic excitation and how atoms emit photons when their electrons make energy-level transitions. The idea that electrons may occupy only certain energy levels was very perplexing to early investigators and to Bohr himself. It was perplexing because the electron was at first thought to be analogous to a particle, a tiny BB, whirling around the nucleus like a planet whirling around the Sun. Just as a satellite can orbit at any distance from the Sun, it would seem that an electron should be able to orbit around the nucleus at any radial distance—depending, of course, like the satellite, on its speed. Moving among all orbits would enable the electrons to emit all energies of light. But this doesn't happen (recall Figure 32.8). Why the electron occupies only discrete levels is understood by considering the electron to be not a particle but a *wave*.

Louis de Broglie introduced the concept of matter waves in 1924. He hypothesized that a wave is associated with every particle and that the wavelength of a matter wave is inversely related to a particle's momentum. These *matter waves* behave just like other waves; they can be reflected, refracted, diffracted, and caused to interfere. Using the idea of interference, de Broglie showed that the discrete values of radii of Bohr's orbits are a natural consequence of standing electron waves. A Bohr orbit exists where an electron wave closes on itself constructively. The electron wave becomes a standing wave, like a wave on a musical string. In this view, the electron is thought of not as a particle located at some point in the atom but as if its mass and charge were spread out into a standing wave surrounding the atomic nucleus—with an integral number of wavelengths fitting evenly into the circumferences of the orbits (Figure 32.10). The circumference of the innermost orbit, according to this picture, is equal to one wavelength. The second orbit has a circumference of two electron wavelengths, the third has three, and so forth (Figure 32.11). This is similar to a chain necklace made of paper clips. No matter what size necklace is made, its circumference is equal to some multiple of the length of a single paper clip.[3] Since the circumferences of electron orbits are discrete, it follows that the radii of these orbits, and hence the energy levels, are also discrete.

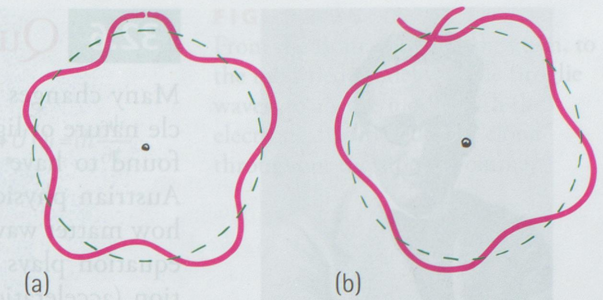

FIGURE 32.10
(a) An orbiting electron forms a standing wave only when the circumference of its orbit is equal to a whole-number multiple of the wavelength. (b) When the wave does not close in on itself in phase, it undergoes destructive interference. Hence, orbits exist only where waves close in on themselves in phase.

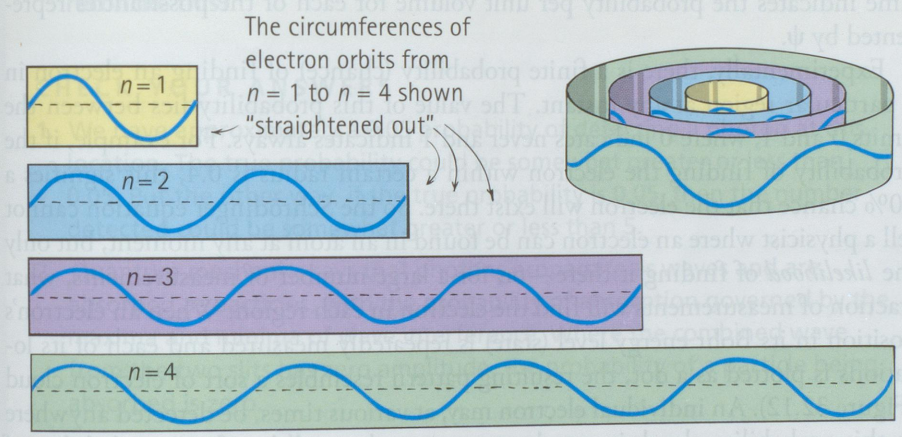

FIGURE 32.11 INTERACTIVE FIGURE MP
The electron orbits in an atom have discrete radii because the circumferences of the orbits are whole-number multiples of the electron wavelength. This results in a discrete energy state for each orbit. (The figure is greatly oversimplified because the standing waves make up spherical and ellipsoidal shells rather than flat, circular ones.)

This model explains why electrons don't spiral closer and closer to the nucleus, causing atoms to shrink to the size of the tiny nucleus. If each electron orbit is described by a standing wave, then the circumference of the smallest orbit can be no smaller than one wavelength—no fraction of a wavelength is possible in a circular (or elliptical) standing wave. As long as an electron carries the momentum necessary for wave behavior, atoms don't shrink in on themselves.

In the still more modern wave model of the atom, electron waves move not only around the nucleus but also in and out, toward and away from the nucleus. The electron wave is spread out in three dimensions, leading to the picture of an electron "cloud." As we shall see, this is a cloud of *probability*, not a cloud made up of a pulverized electrons scattered over space. The electron, when detected, remains a point particle.

SCREENCAST: Matter Waves

[3] For each orbit, the electron has a unique speed, which determines its wavelength. Electron speeds are lower, and wavelengths are longer, for orbits of increasing radii, so for our analogy to be accurate, we'd have to use not only more paper clips to make increasingly longer necklaces but increasingly larger paper clips as well.

Erwin Schrödinger (1887–1961)

32.6 Quantum Mechanics

Many changes in physics occurred in the mid-1920s. Not only was the particle nature of light established experimentally but also particles of matter were found to have wave properties. Starting with de Broglie's matter waves, the Austrian physicist Erwin Schrödinger formulated an equation that describes how matter waves change under the influence of external forces. Schrödinger's equation plays the same role in **quantum mechanics** that Newton's equation (acceleration = force/mass) plays in classical physics.[4] The matter waves in Schrödinger's equation are mathematical entities that are not directly observable, so the equation provides us with a purely mathematical rather than a visual model of the atom—which places it beyond the scope of this book. So our discussion will be brief.[5]

In **Schrödinger's wave equation**, the thing that "waves" is the nonmaterial *matter wave amplitude*—a mathematical entity called a *wave function*, represented by the symbol ψ (the Greek letter psi). The wave function given by Schrödinger's equation represents the possibilities that can occur in a system. For example, the location of the electron in a hydrogen atom may be anywhere from the center of the nucleus to a radial distance far away. An electron's possible position and its probable position at a particular time are not the same. A physicist can calculate its probable position by multiplying the wave function by itself ($|\psi|^2$). This produces a second mathematical entity called a *probability density function*, which at a given time indicates the probability per unit volume for each of the possibilities represented by ψ.

Experimentally, there is a finite probability (chance) of finding an electron in a particular region at any instant. The value of this probability lies between the limits 0 and 1, where 0 indicates never and 1 indicates always. For example, if the probability of finding the electron within a certain radius is 0.4, this signifies a 40% chance that the electron will exist there. So the Schrödinger equation cannot tell a physicist where an electron can be found in an atom at any moment, but only the *likelihood* of finding it there—or, for a large number of measurements, what fraction of measurements will find the electron in each region. When an electron's position in its Bohr energy level (state) is repeatedly measured and each of its locations is plotted as a dot, the resulting pattern resembles a sort of electron cloud (Figure 32.12). An individual electron may, at various times, be detected anywhere in this probability cloud; it even has an extremely small but finite probability of momentarily existing inside the nucleus. It is detected most of the time, however, close to an average distance from the nucleus that fits the orbital radius described by Niels Bohr.

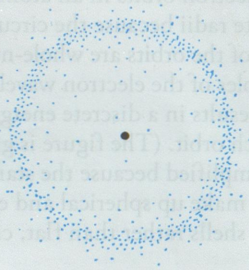

FIGURE 32.12
The probability distribution of an electron cloud for a particular excited state.

Most physicists, but not all, view quantum mechanics as a fundamental theory of nature. Interestingly enough, Albert Einstein, one of the founders of quantum physics, never accepted it as fundamental; he considered the probabilistic nature of quantum phenomena as the outcome of a deeper, as yet undiscovered, physics. He stated, "Quantum mechanics is certainly imposing. But an inner voice tells me it is not yet the real thing. The theory says a lot, but does not really bring us closer to the secret of 'the Old One.'"[6]

[4]Schrödinger's wave equation, strictly for math types, is $\left[-\frac{\hbar^2}{2m}\nabla^2 + U\right]\Psi = i\hbar\frac{\partial\Psi}{\partial t}$.

[5]Our short treatment of this complex subject is hardly conducive to any real understanding of quantum mechanics. At best, it serves as a brief overview and possible introduction to further study. For example, read Ken Ford's *The Quantum World: Quantum Physics for Everyone* (Harvard University Press, paperback edition, 2005).

[6]Although Einstein practiced no religion, he often invoked God as the "Old One" in his statements about the mysteries of nature.

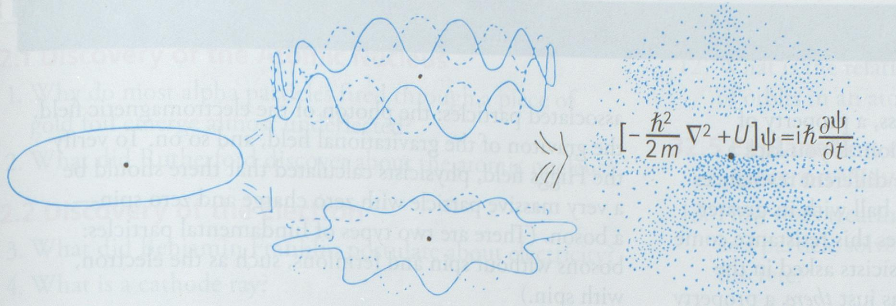

FIGURE 32.13
From the Bohr model of the atom, to the modified model with de Broglie waves, to a wave model with the electrons distributed in a "cloud" throughout the atomic volume.

CHECK POINT

1. **Consider 100 photons diffracting through a thin slit to form a diffraction pattern. If we detect five photons in a certain region in the pattern, what is the probability (between 0 and 1) of detecting a photon in this region?**
2. **Suppose that you open a second identical slit and the diffraction pattern is one of bright and dark bands. Suppose the region where five photons hit before now contains none. A wave theory says that the waves that hit before are now canceled by waves from the other slit—that crests and troughs combine to zero. But our measurement is of photons that either make a hit or don't. How does quantum mechanics reconcile this?**

CHECK YOUR ANSWERS

1. We have approximately a 0.05 probability of detecting a photon at this location. The true probability could be somewhat greater or less than 0.05. Put the other way, if the true probability is 0.05, then the number detected could be somewhat greater or less than 5.
2. Quantum mechanics says that photons propagate as waves and are absorbed as particles, with the probability of absorption governed by the maxima and minima of wave interference. Where the combined wave from the two slits has zero amplitude, the probability of a particle being absorbed is zero.

Considering something to be impossible may reflect a *lack* of understanding, as when scientists thought a single atom could never be seen. Or it may represent a *deep* understanding, as when scientists (and the Patent Office!) reject perpetual motion machines.

32.7 Correspondence Principle

The correspondence principle is a general rule not only for good science but for all good theory—even in areas as far removed from science as government, religion, and ethics. If a new theory is valid, it must account for the verified results of the old theory. This is the **correspondence principle**, first articulated by Bohr. New theory and old theory must correspond; that is, they must overlap and agree in the region where the results of the old theory have been fully verified.

Bohr introduced the correspondence principle in connection with his 1913 theory of the hydrogen atom. He reasoned that when an electron is in a highly excited state, orbiting far from the atomic nucleus, its behavior should resemble (correspond to) classical behavior. And indeed when an electron in such a highly excited state makes a series of quantum jumps, from one state to the next lower one and on downward, it emits photons of gradually increasing frequency that match its own frequency of motion. It seems to spiral inward, as classical physics predicts.

When the techniques of quantum mechanics are applied to still larger systems, the results are essentially identical to those of classical mechanics. The two

HIGGS BOSON

In Chapter 4 we learned about mass, a property of matter that resists changes in motion. If you kick a soccer ball and a bowling ball, the different resistances to motion are easily felt. The bowling ball with its greater mass has greater resistance. Where does this resistance come from? This wasn't a question that physicists asked in the first half of the 20th century. Mass was just *there*, a property of nearly all particles and therefore of larger objects built from those particles. But in the 1960s theorists constructed what seemed like a satisfying theory of all fundamental particles but it had a serious flaw—it predicted that all particles, not just the photon, were massless. What to do? Rejecting the theory was one possibility, but physicists are captivated by mathematical beauty.

In 1964, the Scottish physicist Peter Higgs (soon joined by others with similar ideas) suggested a way out. A new field filling all of space could provide a kind of "viscosity" that would imbue particles with mass. This has come to be called the "Higgs field" and physicists were inclined to believe in it. Verifying its existence, however, was a monumental problem. To find a subatomic field you hunt for the particle that is a manifestation of the field. All fields have associated particles: the photon of the electromagnetic field, the graviton of the gravitational field, and so on. To verify the Higgs field, physicists calculated that there should be a very massive particle with zero charge and zero spin—a boson. (There are two types of fundamental particles: bosons without spin and fermions, such as the electron, with spin.)

The search for the "Higgs boson," as that particle came to be called, stretched over many decades. Finally, on July 4, 2012, scientists at the Large Hadron Collider at the CERN laboratory in Geneva, Switzerland, announced the discovery of a particle that had all the expected properties and a mass 133 times that of the proton (which, in turn, is nearly 2000 times the mass of an electron). Less than a year later, on March 14, 2013, experimental evidence confirmed the Higgs boson.

The scientific community has been ecstatic because the discovery of this elusive particle lends support to the "standard model" of fundamental particles and gives hope that physicists are on track to one day understand gravity more deeply and learn more about dark matter and dark energy (briefly discussed in Chapter 9).

domains blend when the de Broglie wavelength is small compared with the dimensions of the system or of the pieces of matter in the system. It is satisfying to know that quantum theory and classical theory, which make such completely different predictions at the level of a single atom, blend smoothly into a description of nature that extends from the smallest to the largest things in the universe.

For instructor-assigned homework, go to www.masteringphysics.com MP

SUMMARY OF TERMS (KNOWLEDGE)

Atomic nucleus The positively charged center of an atom, containing protons and neutrons and almost the entire mass of the atom, but only a tiny fraction of its volume.

Electron The negative particle in the outer part of an atom.

Ritz combination principle The statement that the frequencies of some spectral lines of the elements are either the sums or the differences of the frequencies of two other lines.

Quantum mechanics The theory of the microworld based on wave functions and probabilities developed especially by Werner Heisenberg (1925) and Erwin Schrödinger (1926).

Schrödinger's wave equation A fundamental equation of quantum mechanics that relates probability wave amplitudes to the forces acting on a system. It is as basic to quantum mechanics as Newton's laws of motion are to classical mechanics.

Correspondence principle The rule that a new theory must produce the same results as the old theory where the old theory is known to be valid.

READING CHECK QUESTIONS (COMPREHENSION)

32.1 Discovery of the Atomic Nucleus

1. Why do most alpha particles fired through a piece of gold foil emerge almost undeflected?
2. What did Rutherford discover about the atomic nucleus?

32.2 Discovery of the Electron

3. What did Benjamin Franklin postulate about electricity?
4. What is a cathode ray?
5. What property of a cathode ray is indicated when a magnet is brought near the tube?
6. What discovery of J. J. Thomson won him the Nobel Prize?
7. What did Robert Millikan discover about the electron?

32.3 Atomic Spectra: Clues to Atomic Structure

8. What did Johann Jakob Balmer discover about the spectrum of hydrogen?
9. What did Johannes Rydberg and Walter Ritz discover about atomic spectra?

32.4 Bohr Model of the Atom

10. What relationship between electron orbits and light emission did Bohr postulate?
11. According to Niels Bohr, can a single electron in one excited state give off more than one photon when it jumps to a lower energy state?
12. What is the relationship between the energy differences of orbits in an atom and the light emitted by the atom?

32.5 Explanation of Quantized Energy Levels: Electron Waves

13. How does treating the electron as a wave rather than as a particle solve the riddle of why electron orbits are discrete?
14. According to the simple de Broglie model, how many wavelengths are there in an electron wave in the first orbit? In the second orbit? In the *n*th orbit?
15. How can we explain why electrons don't spiral into the attracting nucleus?

32.6 Quantum Mechanics

16. What does the wave function ψ represent?
17. How does the probability density function differ from the wave function?
18. How does the probability cloud of the electron in a hydrogen atom relate to the orbit described by Niels Bohr?

32.7 Correspondence Principle

19. Exactly what is it that "corresponds" in the correspondence principle?
20. Would Schrödinger's equation be valid if applied to the solar system? Would it be useful?

THINK AND EXPLAIN (SYNTHESIS)

21. Consider photons emitted from an ultraviolet lamp and a TV transmitter. Which has the greater (a) wavelength, (b) energy, (c) frequency, and (d) momentum?
22. Which color light is the result of a greater energy transition: red or blue?
23. In what way did Rutherford's gold-foil scattering experiment show that the atomic nucleus is both small and very massive?
24. How does Rutherford's model of the atom account for the back-scattering of alpha particles directed at the gold foil?
25. At the time of Rutherford's gold-foil experiment, scientists knew that negatively charged electrons exist within the atom, but they did not know where the positive charge resides. What information about the positive charge was provided by Rutherford's experiment?
26. Why are spectral lines often referred to as "atomic fingerprints"?
27. When an electron makes a transition from its first quantum level to ground level, the energy difference is carried by the emitted photon. In comparison, how much energy is needed to return an electron at ground level to the first quantum level?
28. Figure 32.9 shows three transitions among three energy levels that would produce three spectral lines in a spectroscope. If the energy spacing between the levels were equal, would this affect the number of spectral lines?
29. In terms of wavelength, what is the smallest orbit that an electron can have about the atomic nucleus?
30. Which better explains the photoelectric effect: the particle nature or the wave nature of the electron? Which better explains the discrete levels in the Bohr model of the atom? Defend your answers.
31. How does the wave model of electrons orbiting the nucleus account for discrete energy values rather than a continuous range of energy values?
32. Why do helium and lithium exhibit very different chemical behaviors, even though they differ by only one electron? Why is this question in this chapter rather than in Chapter 11?
33. The Ritz combination principle can be considered to be a statement of energy conservation. Explain.
34. Does the de Broglie model assert that an electron must be moving in order to have wave properties? Defend your answer.
35. Why doesn't a stable electron orbit with a circumference of 2.5 de Broglie wavelengths exist in any atom?
36. An orbit is a distinct path followed by an object in its revolution around another object. An atomic orbital is an electron spread out over a *volume of space* in which the electron is most likely to be found. What do orbits and orbitals have in common?
37. Can a particle be diffracted? Can it exhibit interference?

38. How does the amplitude of a matter wave relate to probability?
39. If Planck's constant, h, were larger, would atoms be larger also? Defend your answer.
40. What is it that waves in the Schrödinger wave equation?
41. When only a few photons are observed, classical physics fails. When many are observed, classical physics is valid. Which of these two facts is consistent with the correspondence principle?
42. When and where do Newton's laws of motion and quantum mechanics overlap?
43. What does Bohr's correspondence principle say about quantum mechanics versus classical mechanics?
44. Does the correspondence principle have application to macroscopic events in the everyday macroworld?
45. What does the wave nature of matter have to do with the fact that we can't walk through solid walls, as Hollywood movies often show using special effects?
46. Largeness or smallness has meaning only relative to something else. Why do we usually call the speed of light "large" and Planck's constant "small"?
47. Make up a multiple-choice question to check a classmate's understanding of the difference between the domains of classical mechanics and quantum mechanics.

THINK AND DISCUSS (EVALUATION)

48. If the electron in a hydrogen atom obeyed classical mechanics instead of quantum mechanics, would it emit a continuous spectrum or a line spectrum? Explain.
49. How can elements with low atomic numbers have so many spectral lines?
50. If the world of the atom is so uncertain and subject to the laws of probabilities, how can we accurately measure such things as light intensity, electric current, and temperature?
51. When we say that electrons have particle properties and then continue to say that electrons have wave properties, aren't we contradicting ourselves? Explain.
52. Why does classical physics predict that atoms should collapse?
53. Did Einstein support quantum mechanics as being fundamental physics, or did he think quantum mechanics was incomplete?
54. Richard Feynman, in his book *The Character of Physical Law*, states: "A philosopher once said, 'It is necessary for the very existence of science that the same conditions always produce the same results.' Well, they don't!" Who was speaking of classical physics, and who was speaking of quantum physics?