

# Light, Quanta, and Atoms

People create games and puzzles because it is fun to try to figure out the unknown. In Mastermind® (Pressman Toys) you guess patterns of colored pegs; in Battleship® (Milton Bradley) you figure out the location of enemy battleships. Some video games include hidden evils for you to discover before they discover you. These games differ in how clues are gathered, but all depend on our natural human enjoyment of figuring out something we cannot see.

Physicists do physics because they experience this same sense of enjoyment. Much of modern science has dealt with figuring out the nature of things we cannot see—stars and galaxies too far away to see in any detail, atoms and molecules too small to see even when viewed up close. Following up clues provided by nature, physicists build models in much the same fashion that we solve puzzles. While the models themselves often seem abstract, the processes used to develop them are familiar.

Within the context of solving puzzles, we can see how the current model of the atom evolved. *Emission* and *absorption spectra* provide information about how atoms absorb and release energy. Niels Bohr combined these observations, the quantum nature of light and Newtonian physics to build a model of the atom, called the *Bohr model*. By applying de Broglie's concept of matter waves to Bohr's model of the atom, electron orbits can be explained

in terms of *circular standing waves*. *Quantum mechanics* now interprets these standing waves as *probability clouds*, which describe the probability of finding the electron at various locations in the atom. This quantum mechanical model, the result of many people working on different pieces of the puzzle, is the currently accepted picture of the atom.

## LIGHT SPECTRA AND ATOMIC STRUCTURE

Some street lamps produce a yellowish light, while others have a bluish-white color. A campfire looks yellow until you throw in the color comics from last Sunday's paper, producing a multicolored display. From common experiences we know that the colors of burning gases depend on the material being burned. So consistent are these colors that nineteenth-century chemists were able to devise a series of flame tests to identify the elements in unknown materials by the colors of light they emit when they burn.

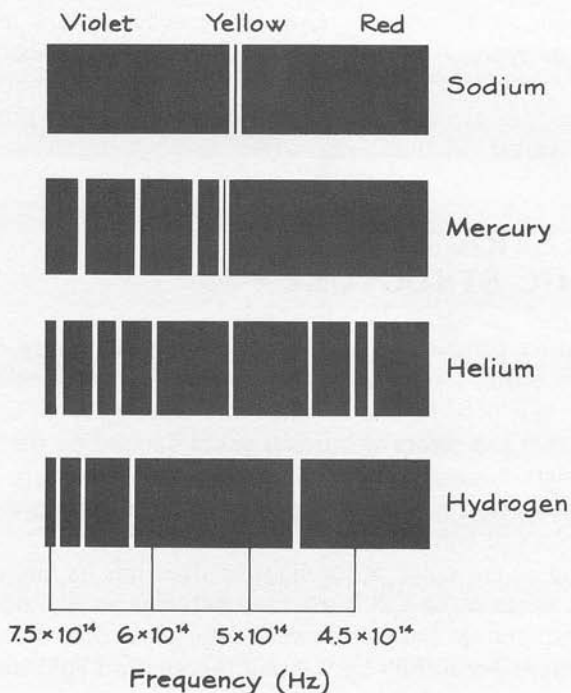
Just looking at the color of the light emitted has its limitations, though. Two colors that seem to be the same may actually be slightly different mixtures of light frequencies. Scientists eventually found that they could analyze substances far more accurately by passing the emitted light through a prism. The resulting spectra provided unmistakable signatures for each of the elements. And, though no one suspected it at the time, these spectra were also a major clue in the puzzle of atomic structure.

### Emission and Absorption Spectra

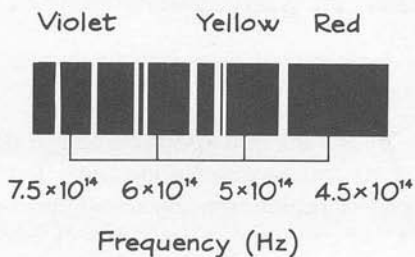
Physicists study light spectra by means of a spectroscope, a device that separates or spreads light out into its component frequencies. The pattern can be captured on a screen or on photographic film. By using film sensitive to infrared or ultraviolet frequencies, we can extend the range of frequencies we can detect into regions below and above the visible range.

The spectrum of light emitted from a light source is called an **emission spectrum**. When we look at emission spectra from different sources, we see two different types. A **continuous spectrum** is a smear of colors, each merging with the next. These colors usually range from red ( $4.3 \times 10^{14}$  Hz) on one end to violet ( $7.5 \times 10^{14}$  Hz) on the other. The spectrum of light emitted by the sun—a rainbow—is an example of a continuous spectrum. Campfires and incandescent lamps also produce continuous spectra. In contrast to the continuous spectrum, a **discrete spectrum** reveals isolated colors with regions of black between them. Because the regions of color look like lines against a black background, a discrete spectrum is often called a *line spectrum*. Emission spectra produced by pure samples of chemical elements are always discrete. A few of these are shown in Figure 18-1.

A discrete spectrum is always unique to the chemical element that produced it. Hydrogen always produces exactly the same spectrum, and it produces a spectrum that is distinctly different from that of neon. The spectral lines combine to produce the colors characteristic of chemical elements used



**Figure 18-1**  
Emission spectra for selected chemical elements.



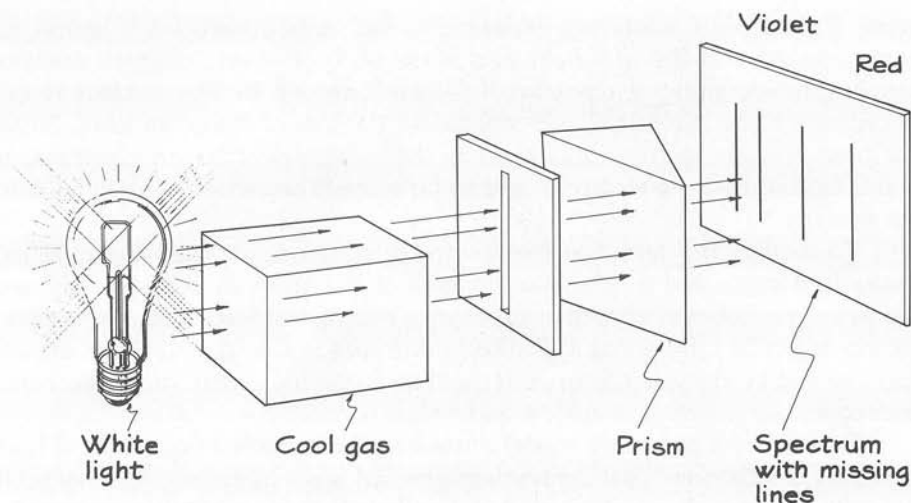
**Figure 18-2**

in the flame tests. Emission spectra can be used to identify all the elements in any substance.

### SELF-CHECK 18A

The emission spectrum shown in Figure 18-2 was obtained from a mixture of two gases. Use the spectra in Figure 18-1 to identify the gases.

Absorption, the reverse of emission, occurs when light travels through a gas. Figure 18-3 shows an experiment in which white light (which contains all visible frequencies) passes through a sample of hydrogen gas and then through a prism. The resulting pattern looks like a continuous spectrum, ex-

**Figure 18-3**

An absorption spectrum is produced when light emitted by an incandescent bulb passes through a sample of cool gas before striking the screen. The pattern of dark lines against a continuous spectrum shows which frequencies the gas has absorbed.

**Figure 18-4**

The emission and absorption spectrum of hydrogen.

cept that it has several dark lines in it. Some of the frequencies of light have been removed, absorbed as the light traveled through the gas. A continuous spectrum in which a few frequencies have been removed by absorption is called an **absorption spectrum**.

When we compare the emission and absorption spectra of hydrogen (Figure 18-4), we see a striking similarity between the two. The emission and absorption lines occur at precisely the *same* frequencies. The same is true for the emission and absorption spectra produced by other chemical elements. The frequencies of light absorbed by any chemical element match exactly the frequencies of light that it emits. Similar processes must be involved in light emission and light absorption. The changes in the atom that produce emission need only be reversed to produce absorption.

Discrete spectra baffled nineteenth-century scientists. It was not hard to understand how burning gases give off light; that is simply a matter of converting thermal energy into electromagnetic energy. But no one had any idea about why each chemical element produced or absorbed only a few of the frequencies of all those possible within the enormous range of the electromagnetic spectrum.

## Discrete Spectra and Atomic Structure

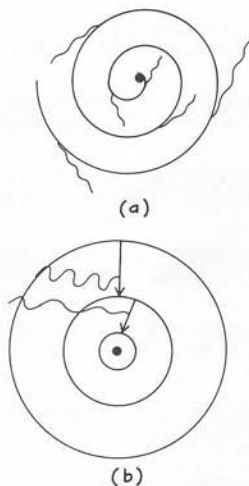
As some physicists were puzzling over discrete emission and absorption spectra, others were working on a related problem—the internal structure of the



atom. Rutherford's scattering experiments with alpha particles (Chapter 8) demonstrated that the atom contains a massive, positively charged nucleus with negatively charged electrons distributed around it. The nucleus is extraordinarily tiny, about one ten-thousandth the size of the atom itself. Since the positively charged nucleus attracts the negatively charged electrons, it seems curious that the electrons stay so far away. They should simply fall into the nucleus.

To explain the fact that the electrons remain apart from the nucleus, Rutherford suggested a "planetary" model of the atom. In our solar system, the sun's gravitational attraction provides a centripetal force that keeps planets moving in its orbit about it. Rutherford proposed that the electrical attraction exerted by the nucleus provides a centripetal force that keeps electrons circling it.

Rutherford's planetary model answered the immediate question of how atoms were structured, yet inadvertently raised other questions. As they orbit the nucleus, electrons accelerate. Accelerating electrons emit electromagnetic waves—that is how we get X rays, for example. In so doing, however, the electrons would constantly lose energy. Each emission of electromagnetic energy would decrease the electron's kinetic energy. Like the satellite that gradually loses energy to friction and falls back to earth, the electron *should* gradually spiral back into the nucleus. Rutherford still could not explain why atoms continue to exist—why they do not simply collapse.



**Figure 18-5**

**(a)** Electrons that release energy as they spiral into the nucleus should produce a continuous spectrum.  
**(b)** Bohr proposed that the electrons make jumps from one allowed orbit to another, producing a discrete emission spectrum.

### The Bohr Model of the Atom

Early in the twentieth century, Niels Bohr examined the problems posed by Rutherford's planetary model of the atom in terms of the discrete emission and absorption spectra of chemical elements. Collections of atoms, like those found in a sample of hydrogen gas for example, do release energy in the form of discrete lines in the electromagnetic spectrum. They also absorb energy—precisely the same lines in the electromagnetic spectrum. If we assume that the electrons are responsible for these lines, then the details of emission and absorption spectra are clues to the internal structure of the atom.

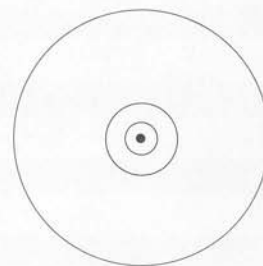
Bohr linked the frequency of emission and absorption lines to a specific quantity of energy, using the relationship that energy is equal to Planck's constant times the frequency of the light. If electrons emitted electromagnetic energy continuously as they spiraled inward toward the nucleus, they would emit all possible frequencies of light (Figure 18-5(a)). We would expect to see a continuous spectrum of light, a rainbow of colors. The observation of discrete rather than continuous spectra suggested to Bohr that electrons release energy while making jumps from one orbit to another, as shown in Figure 18-5(b). However, even jumps from one orbit to another could produce continuous spectra. If all orbits were equally possible, then all possible jumps could occur. The electrons would emit all energies and we would again see a continuous spectrum.

Bohr needed a radical modification to Rutherford's planetary model of the atom. He suggested that electrons move only in certain allowed orbits at specific distances from the nucleus, like those shown for the hydrogen atom in Figure 18-6. While moving in these allowed orbits, electrons do not emit any

electromagnetic energy. When moving from one orbit to another, however, the electrons absorb or release energy. If the electron moves to an orbit farther from the nucleus, it absorbs an amount of energy equal to the energy of a photon in one of the absorption lines. When the electron jumps to an orbit closer to the nucleus, it releases an amount of energy equal to the energy of a photon in one of the emission lines. Allowed orbits restrict the possible jumps to just a few. Consequently, the light spectrum emitted by atoms would be discrete. Since Bohr's model uses the same orbits to explain both emission and absorption of energy, it predicts emission and absorption lines that have identical frequencies.

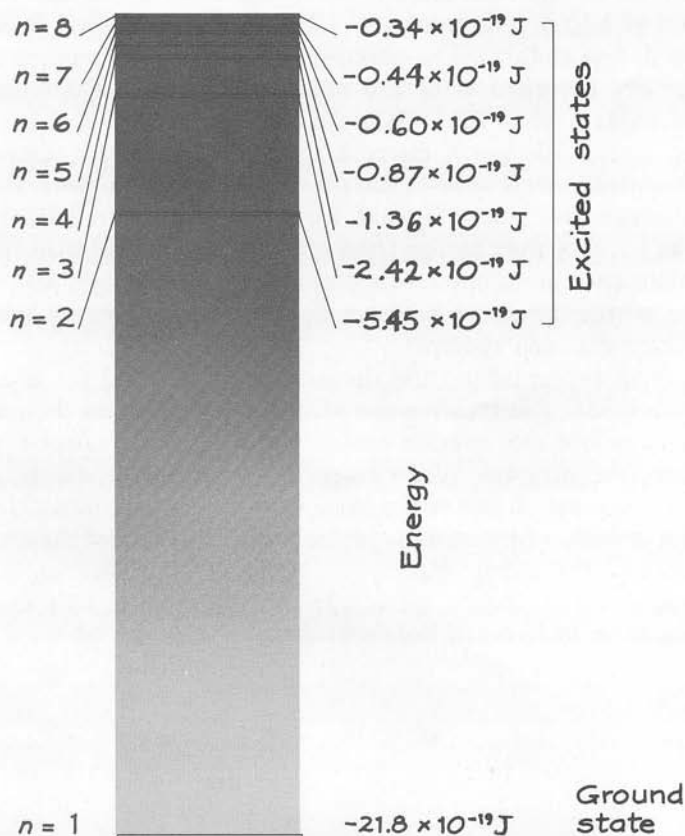
In proposing that electrons move only in allowed orbits, Bohr was suggesting that the energy of the electron is quantized. In any given orbit, an electron has a certain amount of kinetic energy associated with its motion and electrical potential energy associated with the electrical force of attraction exerted by the nucleus. If electrons move only in allowed orbits, their energy takes on selected or discrete values. Bohr saw the same quantization of energy in the atom that Planck saw in the radiation emitted by solids and Einstein saw in the photoelectric effect. At the atomic level, energy appears in discrete amounts.

The energies associated with each allowed orbit provide a useful way of describing what happens as atoms absorb or emit light. Figure 18-7 shows an



**Figure 18-6**

The allowed orbits for the hydrogen atom.



**Figure 18-7**

The energy level diagram for hydrogen.

**energy-level diagram** that illustrates the allowed energies for electrons in the hydrogen atom. The  $n = 1$  line corresponds to the energy of an electron when it is moving in the smallest orbit, the orbit closest to the nucleus. This is often referred to as the **ground state** of the electron because it represents the smallest energy the electron can have. The  $n = 2$  line corresponds to the energy the electron has while in the next orbit;  $n = 3$  while in the third orbit; and so on. Collectively, the energy levels above ground state are called **excited states**. The energy of the electron increases as we move to higher excited states.

The energy-level diagram provides a convenient way to describe the energy emitted or absorbed by the atom. The distance between states is proportional to the energy difference between states. So, for example, the energy difference between  $n = 3$  and  $n = 2$  is smaller than that between  $n = 2$  and  $n = 1$ . Electrons absorb and emit light as they move from one orbit to another. When they move from a lower state to a higher state, electrons absorb energy. When they drop from a higher state back down into a lower state, they release energy. The energies absorbed or released with each jump can be found from the energy level diagram.

In a sample of hydrogen gas at room temperature, each atom is most likely to have its electron in its ground state. An external energy source—heat, light, or electricity—can provide energy to these electrons. When we pass white light through a sample of hydrogen gas for example, electrons “select” photons of energy needed to move into excited states. Electrons might move from  $n = 1$  to  $n = 2$ ; from  $n = 1$  to  $n = 3$ ; from  $n = 1$  to  $n = 2$  and then to  $n = 3$ ; and so forth. The photons with energies equivalent to each of these jumps are absorbed while the others are ignored. Consequently, the light that emerges is white light minus the frequencies used to move electrons into excited states. We see a discrete absorption spectrum. Once excited, electrons eventually drop back to the ground state. Again, several paths are possible. Electrons in the  $n = 3$  state, for example, can drop from  $n = 3$  to  $n = 1$  in one jump or they can go from  $n = 3$  to  $n = 2$  and then from  $n = 2$  to  $n = 1$ . With each jump, the electron emits a photon of light with an energy equal to the energy difference between the two states. Discrete energy levels lead to discrete emission spectra.

In addition to explaining the discrete spectra found for chemical elements, Bohr’s model also resolves the problem of why atoms do not collapse. The electron can lose only specific amounts of energy equivalent to the jumps in the energy-level diagram. When it reaches some lowest energy, called the *ground state*, the electron can neither lose more energy nor move closer to the nucleus. The stability of the atom is assured. Bohr’s model of the atom moved several pieces of the atomic puzzle into place. It did, however, leave some questions unanswered. Before turning to those questions, let’s examine the design of the laser in terms of Bohr’s model of the atom.

## A STEP FURTHER—MATH

MATCHING THE  
EMISSION SPECTRUM

Bohr showed that the energy ( $E$ ) of a given state ( $n$ ) can be expressed by a simple relationship:

$$E_n = -\frac{E_{\text{ground state}}}{n^2} \quad \text{where } n = 1, 2, 3 \dots$$

The  $n = 1$  state is called the ground state of the electron. At  $n = 2$ , the first excited state, the electron's energy is one-fourth that at the ground state. At  $n = 3$ , the energy is one-ninth that of the ground state and so forth. (By convention, the total energy of the atom is expressed as a negative number and approaches zero as the electron moves farther from the nucleus.) For hydrogen, the measured ground state energy of the electron is  $21.8 \times 10^{-19}$  J. Using Bohr's relationship:

$$E_2 = -\frac{21.8 \times 10^{-19} \text{ J}}{(2)^2} = -5.45 \times 10^{-19} \text{ J}$$

$$E_3 = -\frac{21.8 \times 10^{-19} \text{ J}}{(3)^2} = -2.42 \times 10^{-19} \text{ J}$$

We could continue this process until the energy of the electron becomes zero. The energies associated with the first few states of the hydrogen atom are shown in its energy-level diagram (Figure 18-7).

We can use the energies associated with each state to predict the energy released when the electron jumps from a higher to a lower state. Consequently, we could check to see whether Bohr's model predicts the emission lines observed for hydrogen. When an electron moves from a higher state to a lower state, it emits energy equivalent to the energy difference between the two states. For example, an electron that moves from the  $n = 3$  to the  $n = 2$  state loses:

$$\begin{aligned} E_3 - E_2 &= (-2.42 \times 10^{-19} \text{ J}) - (-5.45 \times 10^{-19} \text{ J}) \\ &= 3.03 \times 10^{-19} \text{ J} \end{aligned}$$

We can determine the frequency of light to which this energy corresponds by using the relationship:

$$E = hf \quad f = \frac{E}{h}$$

Since the energy released by the electron is  $3.03 \times 10^{-19}$  J, we have:

$$f = \frac{E}{h} = \frac{3.03 \times 10^{-19} \text{ J}}{6.62 \times 10^{-34} \text{ J} \cdot \text{s}} = 4.58 \times 10^{14} \text{ Hz}$$

This corresponds to the frequency of the red line in the hydrogen spectrum!

Bohr's model of the atom predicted the frequencies of light found in the emission spectrum of hydrogen. Three other lines are visible. They have frequencies of  $6.2 \times 10^{14}$  Hz,  $6.9 \times 10^{14}$  Hz, and  $7.3 \times 10^{14}$  Hz. All occur when the electron jumps from a higher state down to the  $n = 2$  state. Can you figure out to which jump each line corresponds?



**SELF-CHECK 18B**

Use the energy-level diagram for hydrogen shown in Figure 18-7 to describe:

- the possible jumps an electron could make in moving from the  $n = 4$  level to the  $n = 1$  level
- the number of different lines in the emission spectrum that could be produced by these jumps
- the order of jumps, from least to most energy released

**LASERS**

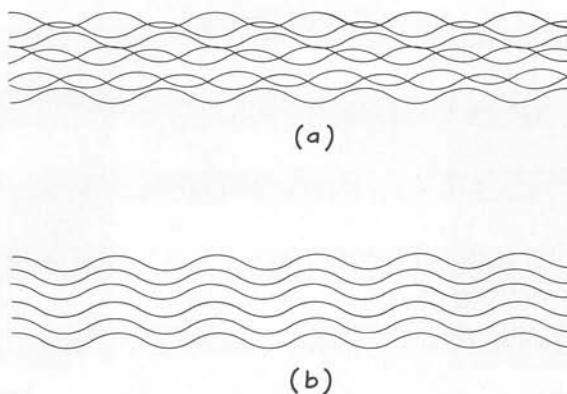
The laser is one of the most interesting practical results of our increased understanding of the internal structure of the atom. Invented in 1960, the **laser** (Light Amplification through Stimulated Emission of Radiation) has found many applications. The fundamental process of light emission is the same in a laser as in an ordinary light bulb. Electrons move from a higher energy level to a lower one, releasing energy in the form of light. However, our understanding of energy levels in the atom has enabled us to manipulate this process to produce coherent light.

**Lasers and Coherent Light**

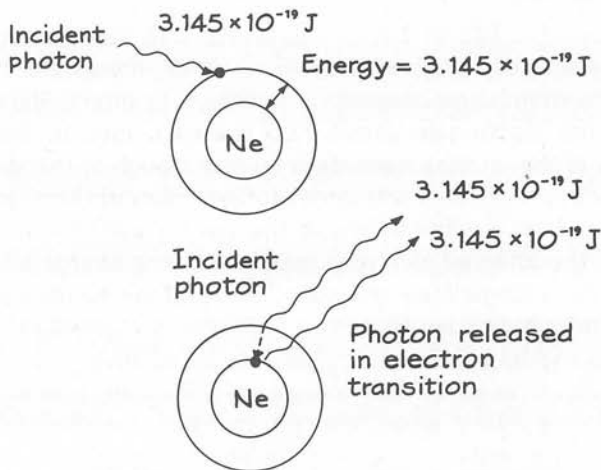
Light emitted by a laser differs from light emitted by a light bulb because of a single property, called *coherence*. Laser light is coherent; light emitted by a light bulb is incoherent. **Coherent light** has two characteristics: (1) it has a single wavelength, and (2) the waves produced by individual atoms are in phase, or in step, with one another. We examine each of these characteristics in more detail.

Light emitted by most light sources does not consist of a single wavelength. While the light has a characteristic color, this color is usually a composite of at least several discrete lines in the emission spectrum. The whitish blue color of a mercury-vapor lamp, for example, comes from the violet lines in the mercury spectrum; but other lines (green and yellow) are also present. We can convert ordinary light into single-wavelength light by using a filter to block all wavelengths but one. The light is then a single wavelength, but it is not yet coherent.

The second characteristic of coherent light is that the waves produced by individual atoms are in step with one another. Ordinarily, excited atoms emit light spontaneously. The light waves produced by individual atoms move independently of one another (Figure 18-8(a)). The waves often interfere destructively, reducing the intensity of the light finally emitted. In a laser, atoms are stimulated to emit each photon so that its wave has a definite relation to all other waves. The crests of all waves line up (Figure 18-8(b)) and interfere

**Figure 18-8**

(a) Ordinary light sources produce waves that are independent of one another. (b) Coherent light consists of waves that are in step with one another.

**Figure 18-9**

Neon in its excited state can be stimulated to drop to a lower energy state by a photon equal in energy to the energy lost by the neon atom.

constructively with one another. Using an analogy drawn from rock concerts, coherent light is similar to applause that occurs when people keep time with the music. It is stimulated by the music, so everyone claps at the same time. Incoherent light is like the general applause that concludes a performance.

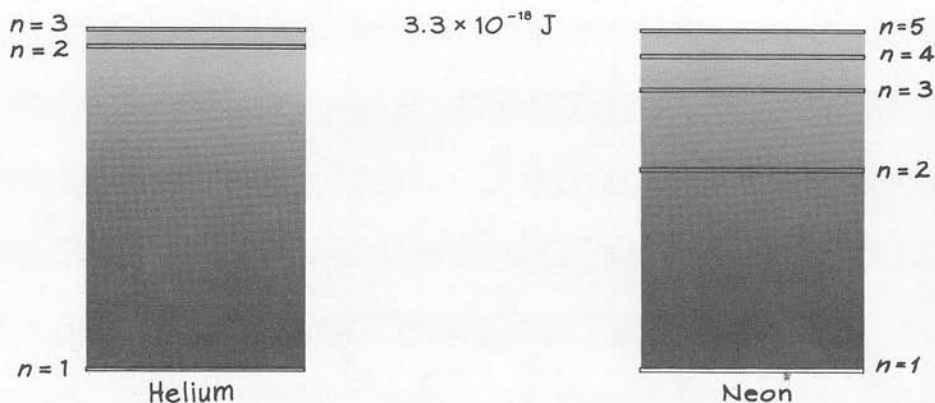
The process by which coherent light is produced needs to do two things. First of all, it needs to emphasize one specific electron transition in order to produce light of a single wavelength. This can be accomplished by moving a majority of the atoms into a single excited state. Secondly, each atom must be induced to emit its photon so that the crests of each light wave are lined up with the others. This occurs as a result of stimulated emission.

### Stimulated Emission of Radiation

**Stimulated emission** of radiation occurs when an atom in an excited state interacts with a photon of precisely the energy that would be emitted by the atom as it drops to a lower energy state. This interaction stimulates the atom to make the transition and to release a photon that is coherent with the incident photon. For example, suppose one of the electrons of a neon atom is in an excited state, as illustrated in Figure 18-9. The energy emitted by the electron if it moves down to the next lowest orbit is  $3.145 \times 10^{-19}$  joules (J).

**Figure 18-10**

The  $n = 3$  state of helium and the  $n = 5$  state of neon are both  $3.3 \times 10^{-18}$  J above the ground state. Helium atoms in the  $n = 3$  state can excite the neon atoms to move to the  $n = 5$  state.



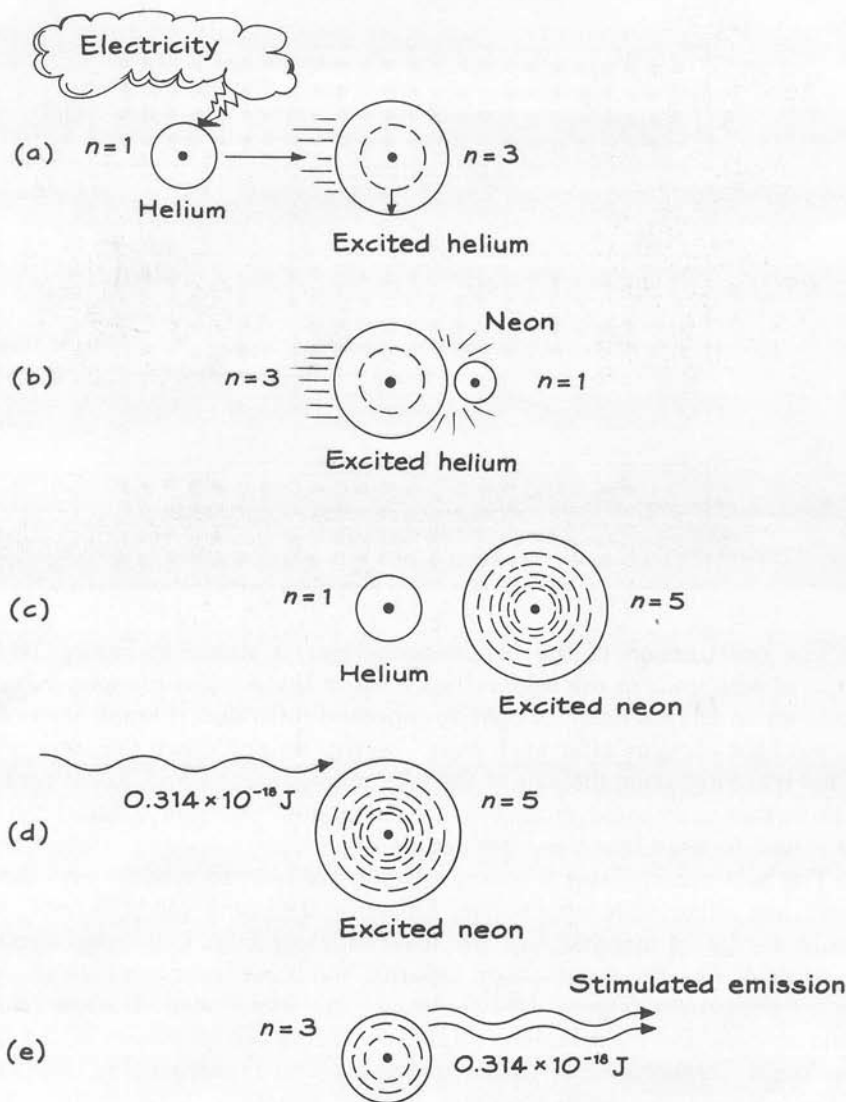
When a photon of  $3.145 \times 10^{-19}$  J interacts with this excited atom, it will stimulate the electron to drop to the lower orbit. A photon of  $3.145 \times 10^{-19}$  J is emitted. The stimulation does more, however. In effect, the incoming photon “shakes” the electron to cause it to change orbits. In the process, the electron emits a photon that matches crest and trough of the incident photon. The two photons are in step with one another. Each of these photons can, in turn, stimulate other excited atoms in the same state to emit identical photons. Since all the affected electrons make the same energy jump, the resulting radiation is of a single frequency. And since all the photons are emitted by stimulation, the radiation is coherent.

We might expect stimulated emission to be relatively simple to produce. In fact, however, it is somewhat complex. First of all, you have to move a great many atoms into a single excited energy state. Secondly, the atoms must remain in that state long enough to enable the photon to interact with the electron. Most atoms emit light approximately one-millionth of a second after having been excited, too short a time for the photon-electron interaction. However, a few excited states last a few thousandths to a few hundredths of a second. These states can be used for stimulated emission.

### Helium-Neon Lasers

Lasers use the process of stimulated emission to produce the intense, coherent light for which they are famous. While several materials can be used to establish this process, helium and neon—used in the helium-neon laser—are the most common. Let’s examine the role of each element in establishing stimulated emission of light.

Helium atoms are used to raise the neon atoms to the particular excited state required for stimulated emission. The neon atoms then emit the photons that we collectively call the laser light. Figure 18-10 shows the energy level diagrams of helium and neon. For convenience, we have called the ground-state energy zero and expressed the energy of higher levels relative to it. Helium and neon both have excited states at  $3.3 \times 10^{-18}$  J above the ground state—the  $n = 3$  state for helium and the  $n = 5$  state for neon. When a neon electron reaches the  $n = 5$  state, it remains there for a relatively long time.

**Figure 18-11**

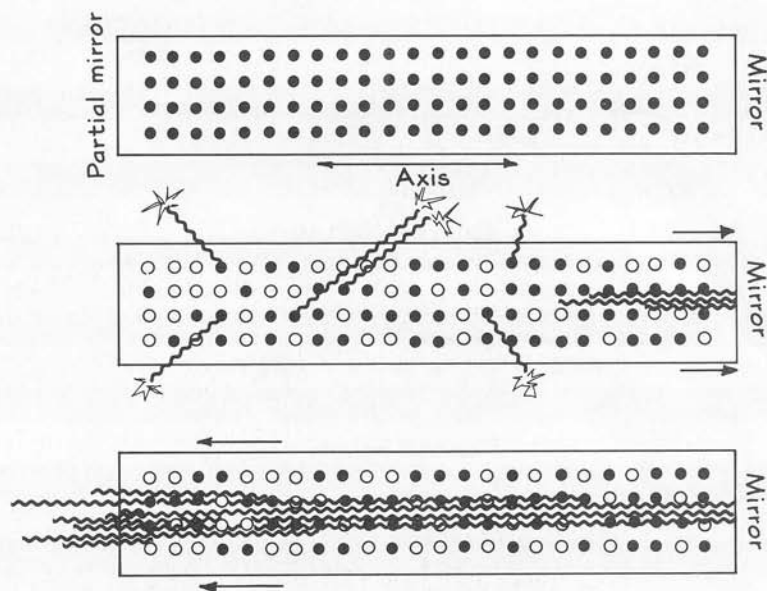
A high voltage is applied to a mixture of helium and neon gas, exciting some helium atoms to the  $n = 3$  state. Collisions between excited helium atoms and neon atoms excite neon atoms to the  $n = 5$  state. A single photon of the proper energy can then stimulate neon atoms to make a transition from the  $n = 5$  state to the  $n = 3$  state, emitting red light.

Thus, it is an excellent candidate for stimulated emission. When it does change states, the electron moves down to the  $n = 3$  state and releases a photon of red light.

Figure 18-11 shows the complete process by which a helium-neon laser emits coherent, red light. A relatively high voltage (1000 volts) is applied to a tube containing a mixture of helium and neon gas. (Usually the mixture consists of approximately 85% helium and 15% neon.) The high voltage excites some of the helium atoms to the  $n = 3$  state (Figure 18-11(a)). These excited helium atoms collide with neon atoms (Figure 18-11(b)), exciting the neon atoms to their  $n = 5$  state. As the excited neon atoms are stimulated to drop down to the  $n = 3$  state (Figure 18-11(d)), each of them emits a single photon of  $0.314 \times 10^{-18} \text{ J}$  (Figure 18-11(e)). Each of these photons will stimulate the emission of other  $0.314 \times 10^{-18} \text{ J}$  photons by other excited neon atoms.

**Figure 18-12**

Mirrors reflect photons traveling along the axis of the tube. Those not traveling along the axis merely escape and do not contribute to further stimulated emissions.



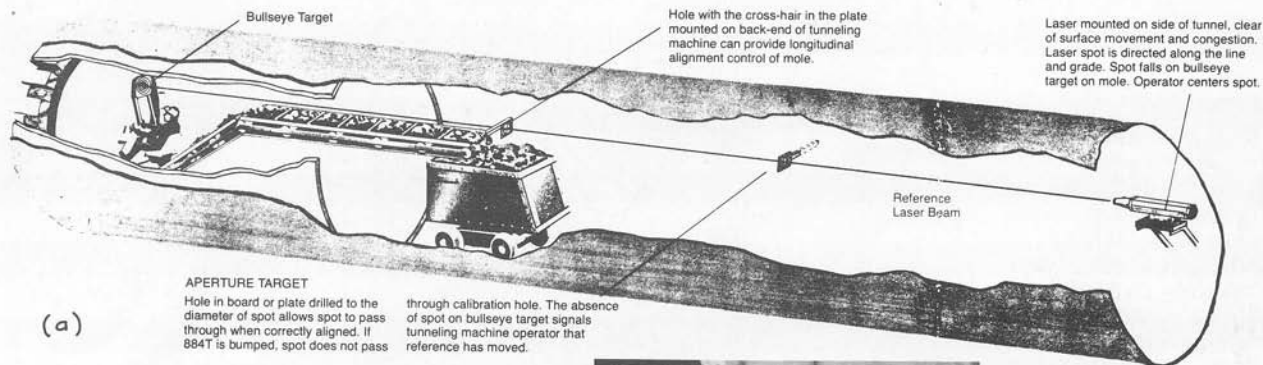
The construction of the helium-neon laser is shown in Figure 18-12. Mirrors at both ends of the tube reflect most of the emitted photons in order to build up an intense beam of light by repeated reflection. A small amount of coherent light escapes after each complete trip up and down the laser. Photons not traveling along the axis of the tube merely escape and do not contribute to further stimulated emissions. Consequently, the light emitted by the laser is well focused into a narrow beam.

The helium-neon laser is a very inexpensive laser to produce and consequently has found wide application. Other lasers, using solids, liquids, and gases as the lasing material, are also used widely. Lasers find broad applications in fields like the construction industry, medicine, energy research, and even entertainment (Figure 18-13). Lasers are often used to align underground tunnels, the most famous application being the alignment of the Bay Area Rapid Transit (BART) tunnel under the San Francisco Bay. Surgeons use the intensity and small diameter of laser beams to assist them in performing microsurgery. Grocery store cash registers read product codes with laser light. Lasers are now used as a type of "record needle" for videodiscs. The puzzle of atomic spectra has led to an enormously useful array of laser devices.

## THE QUANTUM-MECHANICAL ATOM

Rarely do we solve a puzzle on our first attempt. Mastermind requires at least two tries and sometimes many, many more. We do crossword puzzles with pencils so we can erase. In a sense, each try is a temporary model. We take the observations we have, build a temporary model, and then test it with additional observations. Eventually we reach a final model, a solution to the puzzle. Building models in science is in some ways the same. We take the available observations, build a model, and then look for additional observa-





(a)

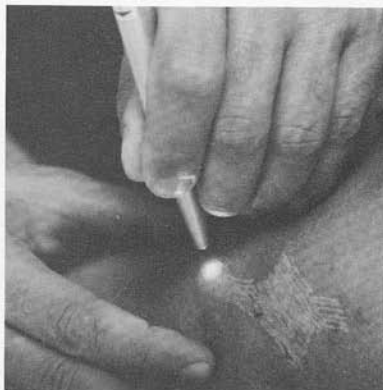
## APERTURE TARGET

Hole in board or plate drilled to the diameter of spot allows spot to pass through when correctly aligned. If 864T is bumped, spot does not pass

through calibration hole. The absence of spot on bullseye target signals tunneling machine operator that reference has moved.



(b)



(c)

Hank Morgan, Rainbow

Figure 18-13

Lasers have found wide application in the construction industry, medicine, energy research, and entertainment: (a) an alignment laser for underground tunneling; (b) a Laserdisc™ Videodisc player; (c) a laser used in surgery.

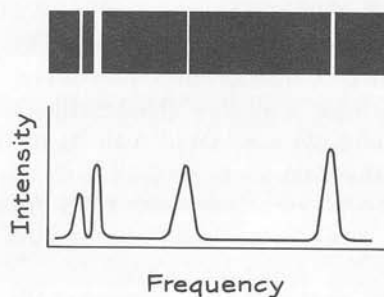


Figure 18-14

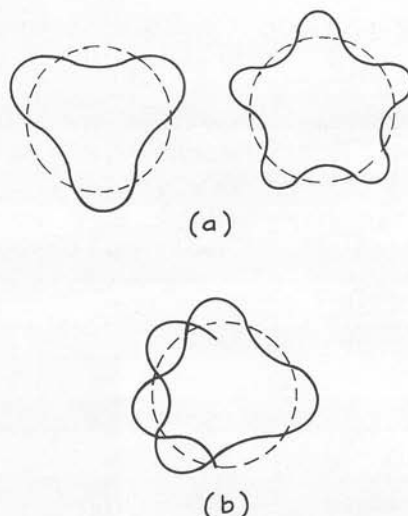
Experimentalists found a small range of frequencies associated with each line in emission and absorption spectra.

tions that can test the model's validity. Sometimes the model survives; sometimes it does not.

Bohr's model of the atom was a temporary one. While the concept of allowed orbits explains both the stability of the atom and discrete emission and absorption spectra, it is not based on any known principles of physics. There seems to be no reason for electrons to jump instantaneously from one energy state to another, shunning all energies in between. Furthermore, additional observations revealed limitations to the model. While it was enormously accurate in describing the hydrogen spectrum, Bohr's model was far less successful in describing the spectra produced by more complex atoms. In addition, more detailed measurements showed that the so-called lines in atomic spectra are not confined to the frequencies predicted by Bohr's allowed orbits. Each line is really a narrow band centered at the predicted frequency (Figure 18-14). Moreover, the intensity of individual spectral lines varied.

**Figure 18-15**

Because the electron moves in a circular orbit, the de Broglie wavelength must close upon itself. **(a)** Orbits exist only where a wave can close on itself in phase, so that crest meets crest or trough meets trough. **(b)** Other matches produce destructive interference and the wave dies out.



Some lines were brighter than others—an observation for which Bohr's model offered no explanation. Bohr's picture began to grow fuzzy. The resolution to these problems lay in yet another model of the atom—the quantum mechanical model.

### De Broglie Wavelengths and Circular Standing Waves

As we saw in Chapter 17, de Broglie wavelengths explain the wavelike behavior of electrons as they move through a thin metal foil. Like light waves and X rays, electrons produce diffraction and interference patterns. If electrons orbiting a nucleus also have wavelike characteristics, each of them should have a de Broglie wavelength associated with its momentum. An electron's momentum varies with the distance of its orbit from the nucleus; consequently a unique wavelength can be associated with each possible orbit. Electrons in orbits near the nucleus have longer wavelengths than electrons in orbits far from the nucleus.

If we forget about electrons for a moment and think just in terms of waves, circular paths demand that a wave close in upon itself. Closing waves into circles presents circumstances similar to those that result in resonance and standing waves. If the wave closes so that a crest meets a crest or a trough meets a trough (Figure 18-15(a)), the wave reinforces itself. We would have a **circular standing wave**. If a wave closes so that a crest meets a trough, it cancels itself. Other matches (Figure 18-15(b)) create a jumble that gradually dies out.

In applying his concept of matter waves to the Bohr model of the atom, de Broglie showed that Bohr's allowed orbits were those orbits for which de Broglie wavelengths could establish circular standing waves. For a given orbit, an electron has a de Broglie wavelength that depends on its momentum in that orbit. When a whole number of these wavelengths fit into the circumference of the orbit ( $n\lambda = 2\pi r$ , where  $n = 1, 2, 3, \dots$ ), the orbit is stable. When a whole number of de Broglie wavelengths do not fit into the circumference of

the orbit, the orbit is unstable and the electron moves on. Electrons select those orbits in which their de Broglie wavelengths can establish circular standing waves.

De Broglie's matter waves provided a much-needed mechanism to explain the discrete orbits found in atoms. Once a wavelike character can be associated with the electron, the identification of acceptable orbits becomes a mathematical puzzle in which the orbit's circumference (determined from the radius of the orbit) must match a whole number of de Broglie wavelengths (determined by the electron's velocity at that radius). Classical concepts of force, momentum, and circular waves are integrated with the modern concept of matter waves.

### SELF-CHECK 18C

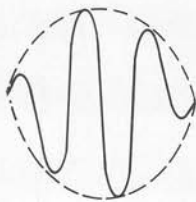
De Broglie wavelengths and circumferences have been calculated for a series of possible orbits in the Bohr atom. Select the orbits that satisfy de Broglie's criteria and indicate, in each case, how many de Broglie wavelengths fit into the orbit.

Orbit	de Broglie Wavelengths	Circumference
A	$3.33 \times 10^{-10} \text{ m}$	$3.33 \times 10^{-10} \text{ m}$
B	$4.71 \times 10^{-10} \text{ m}$	$6.66 \times 10^{-10} \text{ m}$
C	$6.67 \times 10^{-10} \text{ m}$	$13.34 \times 10^{-10} \text{ m}$
D	$5.72 \times 10^{-10} \text{ m}$	$9.80 \times 10^{-10} \text{ m}$
E	$9.99 \times 10^{-10} \text{ m}$	$29.97 \times 10^{-10} \text{ m}$

### Waves and Probabilities

While de Broglie's circular standing waves give us a way of understanding why some orbits are allowed and others are not, they thrust us back into the ambiguous world of wave-particle duality. Added to the diffraction and interference phenomena discussed in Chapter 17, the wave nature of orbiting electrons provides impressive evidence that particles do behave like waves. But most of us end up wondering just what all this means. When we say that a moving ball or electron has a wavelength, what does this wavelength represent? A few years after de Broglie published his work, his entire concept of matter waves was incorporated into a much broader model called quantum mechanics. **Quantum mechanics** presents a mathematical description of phenomena in terms of wave equations—a complex description of a wave view of reality. It is extremely abstract and by itself offers us little additional insight. But the probabilistic interpretation given to its results can be very helpful and, in a sense, offers the only real bridge between wave and particle views of matter.

The **probabilistic interpretation** given to the wave view of matter proposes that the waves associated with electrons and other particles describe

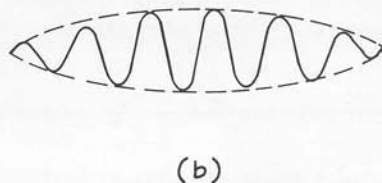
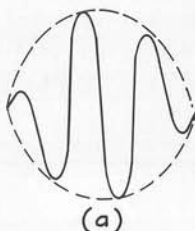
**Figure 18-16**

A wave packet describes the wave associated with a particle. The larger the amplitude of the wave within the packet, the greater the probability of locating the particle at that point.

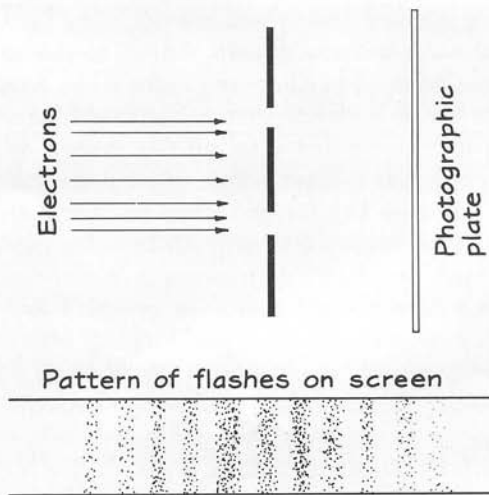
the probability of locating that particle at each point in space. For an isolated particle, quantum mechanics predicts a wave like that shown in Figure 18-16, called a **wave packet**. The wavelength of the wave is the de Broglie wavelength, determined from the momentum of the particle. The amplitude of the wave varies from zero at the ends of the packet to a maximum in the middle. According to the probabilistic interpretation, the amplitude squared at each point in space describes the probability of finding the particle at that point. For the wave packet shown in Figure 18-16, the probability of finding the electron is the largest at the center of the packet and essentially zero near the ends and outside the packet.

### SELF-CHECK 18D

The sketches in Figure 18-17 describe the wave packets associated with two electrons. Which has the greater momentum? Use the probabilistic interpretation of the wave amplitude to describe the probable locations of each electron. Which can be located within a narrower region of space?

**Figure 18-17****Figure 18-18**

We can think of an interference pattern as describing the most probable paths of a collection of electrons. Most go to regions of constructive interference; few go to regions of destructive interference.



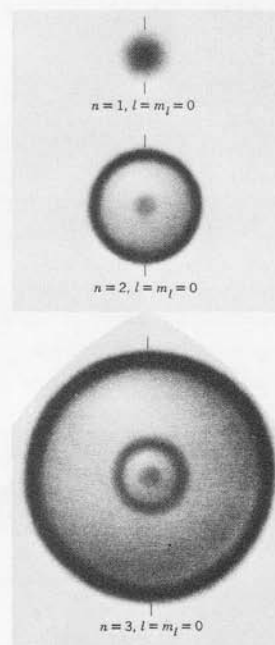
We can extend this probabilistic interpretation to the behavior of a group of electrons, like those that create diffraction and interference phenomena. Let's imagine that we fire a continuous stream of electrons at a barrier with two slits in it. The electrons that make it through the slits strike a photographic plate and leave a dot exposed on the plate. At first the dots seem to be spaced across the plate somewhat randomly. But gradually a pattern builds up (Figure 18-18). The electrons are not spread randomly across the plate—they are clustered in bands that are identical to interference bands produced by waves. Regions where a lot of electrons go alternate with regions where few electrons go. We can think of the interference pattern produced by waves as a description of the probability that an electron will reach each point of the screen.

### Probabilities and the Atom

When applied to the atom, quantum mechanics presents a model of the atom that is more sophisticated than the Bohr model. Quantum mechanics replaces the discrete orbits in the Bohr atom with probabilities. Imagine for a moment that we measure the position of the electron in the hydrogen atom and record that position with a small dot. If we repeat this experiment hundreds of times, a pattern like that in Figure 18-19(a) builds up. Dots are spread about throughout the space outside the nucleus, but they are more dense in some regions than others. The region of space in which the density of the dots is the largest corresponds to the first Bohr orbit. In this sense quantum mechanics presents a model of the atom that corresponds to Bohr's model. However, Bohr's restriction that electrons exist only in discrete orbits is replaced with a description that says that electrons are most probably found in regions that correspond to Bohr's orbits. Theoretically, the electron can exist anywhere in the atom. Quantum mechanics replaces Bohr's discrete orbits with fuzzy **electron probability clouds**.

Changes in state can be described by sketches like those in Figure 18-19. Bohr's model of the atom leads us to imagine an electron spontaneously jumping from one orbit to another, from one energy level to another. Quantum mechanics replaces this image with one in which one picture dissolves into another. As an electron moves from the  $n = 2$  state to the  $n = 1$  state, the distribution in (b) fades into and becomes the distribution in (a). The transition occurs over an extremely short time interval,  $10^{-8}$  seconds, during which time light is emitted. The reverse happens when light is absorbed. As the electron gains energy, the probability distribution in (a) fades into and becomes the distribution shown in (b). Our knowledge of the electron remains incomplete. The different probability distributions reflect the energy of the electron but provide only a guess at where the electron might actually be found.

Ultimately, the success of the quantum mechanical model of the atom rests with its ability to take up where Bohr's model left off. Bohr's model was unable to describe the spectra of the more complex atoms, the different intensities of spectral lines, and the broadening of individual spectral lines. The explanation of the broadened spectral lines lies in the probabilistic nature of



**Figure 18-19**

The electron probability clouds for  $n = 1$ ,  $n = 2$ , and  $n = 3$  states of the hydrogen atom.



electron orbits. Since the quantum mechanical atom predicts only probabilities, the electron can be in a variety of locations about the nucleus. When electrons change state, the most probable transition they will make is from one Bohr orbit to another. In reality, of course, they can make any number of jumps—some larger and some smaller than the discrete jumps predicted by Bohr's model of the atom. The spectral lines will be broadened slightly because electrons are making transitions slightly larger and slightly smaller than the transition predicted by Bohr orbits. The different intensities of the spectral lines reflect the different probabilities associated with each transition. Some transitions are very improbable and, thus, occur very seldom. Others occur more frequently, to the point that we see a dim spectral line. Finally, highly probable transitions occur frequently. These yield the most intense spectral lines. The different intensities of spectral lines arise from the different probabilities associated with each transition. While the mathematics becomes increasingly complex, quantum mechanics has been applied to complex atoms as well as to the collisions among atoms and detailed structures of materials. Quantum mechanics provides us with a significantly more powerful model of atomic structure.

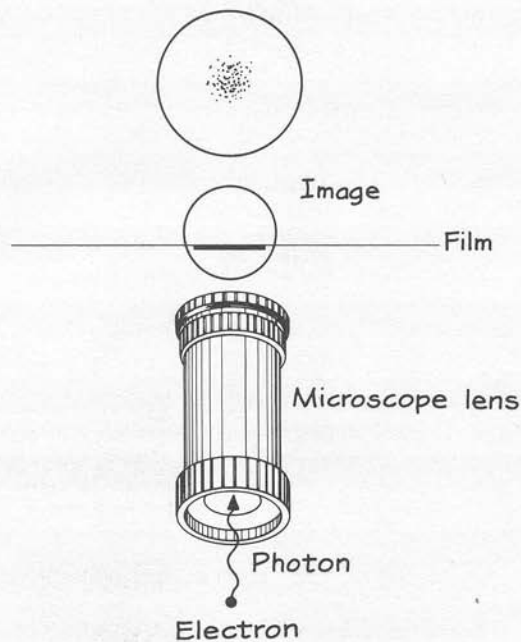
Today the quantum mechanical model of the atom is regarded as the most complete picture of atoms and their interactions. It fits observations and experimental results extremely well. But, is it the correct solution to the puzzle of the atom? We do not know. Here our analogy between games and puzzles and model-building in science fails. Ultimately, there are no "correct solutions" in science. All scientific models and theories are temporary—reflections of the observations available at the time they were built. As we learn more, our theories change. Perhaps that is what keeps the puzzle so interesting—what keeps physicists doing physics.

## PROBABILITY AND UNCERTAINTY

The concept that matter waves describe probabilities is at once comforting and disquieting. When applied to diffraction and interference effects where we are dealing with many electrons at once, a probabilistic interpretation seems reasonable. After all, a lot of electrons can travel a lot of different paths, and we can use probability to show the collective result, just as we can use it to show the collective results of tossing a coin a thousand times. When applied to a single electron, however, the concept seems disquieting. Quantum mechanics predicts the probability distribution for a single electron, not a collection of one thousand one-electron atoms. We have replaced the concept of knowing where an electron is with knowing the probability of its being at each point in space. This is rather like saying we can never know whether an individual coin has landed heads or tails up, only how likely it is to end up one way or the other. Such an idea is at odds with our everyday experience!

### Limits to Measuring

Traditionally, physicists believed that they measured, and consequently described, phenomena from an objective viewpoint. According to the traditional

**Figure 18-20**

We "see" an electron by causing light to interact with it. Light strikes the electron, is reflected into a microscope lens and transmitted to a piece of photographic film.

view, it should be possible to measure the position and velocity of an object as precisely as the measuring instruments allow. The fact that these measurements do not tell us the exact position or velocity of a particle seemed to be a reflection of the measuring instruments, not of our ability to know these quantities.

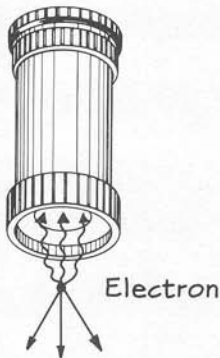
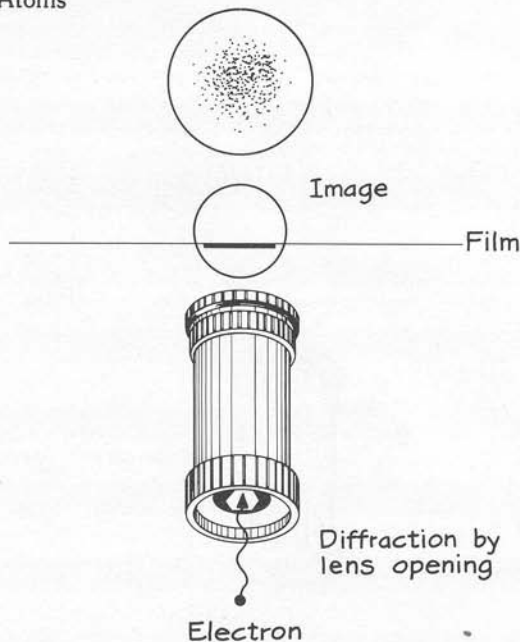
Quantum mechanics and its probabilistic interpretation force us to conclude that we cannot know things as exactly as we once thought. To understand how this inherent uncertainty arises, we conduct a thought experiment. Suppose we want to measure the position of a stationary electron as precisely as possible. Measuring the position of something usually involves seeing it in some way, so we use light to help us detect the electron. Light strikes the electron and is reflected into a microscope lens and transmitted to a piece of photographic film (Figure 18-20). There, the magnified image of the electron's location is recorded. We generally trace backward from the film image to infer the original location of the electron.

Difficulties arise when we think about the interaction between the light and the electron. Scientists have always conceded that the process of measurement alters the thing being measured, but traditionally they have contended that this alteration is quite small. When light strikes an electron, the energy and momentum carried by its quanta can be transferred to the electron, thus altering the electron's position and momentum. We can minimize this effect by allowing only one quantum of light to strike the electron, but that single quantum will still affect the electron.

We can examine what happens to our measurements of position and momentum when we try to minimize the effect the experimenter has by looking at the electron with a single quantum of light. In order to restrict ourselves to this single quantum, we look through narrower and narrower microscope lenses. Because of diffraction, however, smaller and smaller lenses produce

**Figure 18-21**

Because of diffraction by the microscope lens, a well-defined image of the electron is spread into a diffuse circle. The diffuse circle describes an uncertainty associated with locating the electron.

**Figure 18-22**

Once it has collided with the photon, the electron can move away in one of many possible directions.

images of the electron that are more diffuse (Figure 18-21). As we try to minimize the momentum transferred to the electron during the process of measuring, our knowledge of where the electron is becomes more and more vague. The uncertainty in locating the electron increases. Suppose we relax our one quantum restriction and allow more quanta to interact with the electron. Now our knowledge of the position of the electron improves, but our knowledge of the direction it will move once struck by the photons (Figure 18-22) deteriorates. What we have gained in being more able to locate the electron precisely, we have lost in knowing where it will be after we have found it. Large uncertainties in momentum make it more difficult for us to predict where the electron goes after the collisions with the quanta.

You might imagine that we could design an experimental apparatus that would reduce the uncertainties in position and in momentum. For example, we can reduce the uncertainty in momentum by using light quanta of lower and lower energies. Consequently, less energy is transferred to the electron. However, quanta of lower energy correspond to waves of longer wavelength. Since the spreading of light due to diffraction depends directly on the wavelength, the diffraction pattern spreads as the wavelength increases. The location of the electron would now become less certain. If we decrease the wavelength to locate the electron more precisely, we simultaneously increase the frequency and transmit more energy to it, making the momentum less certain. We cannot know both position and momentum with precision at the same time.

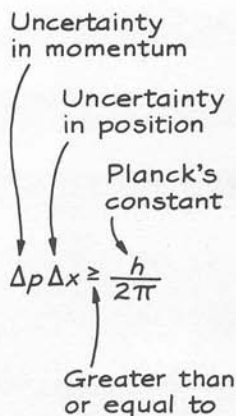
### Heisenberg Uncertainty Principle

The lack of certainty inherent in physical measurement was first described by Werner Heisenberg. The **Heisenberg uncertainty principle** states that in any measurement, the uncertainty in momentum times the uncertainty in po-

sition must be greater than a fixed number, Planck's constant divided by  $2\pi$ , or  $1 \times 10^{-34} \text{ J} \cdot \text{s}$ . In equation form:

$$\left( \text{Uncertainty in momentum} \right) \times \left( \text{uncertainty in position} \right) \geq \frac{\text{Planck's constant}}{2\pi}$$

This principle places a lower limit on the product of the uncertainty in the momentum and the uncertainty in the position. While the uncertainties can change, they must change in response to one another. If the uncertainty in position decreases, the uncertainty in momentum increases, and vice versa. Measuring one quantity precisely limits our knowledge of the other. Expressed in words, Heisenberg's principle says that we can never know both the exact position and the exact momentum of any object at the same time.



## A STEP FURTHER—MATH

### HEISENBERG'S UNCERTAINTY PRINCIPLE

An automobile with a mass of 1000 kg moves at a speed of about 25 m/s. Its momentum is 25,000 kg · m/s. A reasonable uncertainty in measuring this momentum might be 5% — 1250 kg · m/s. The uncertainty in the position of the car would then be:

$$\begin{aligned} \Delta x \Delta p &\geq \frac{h}{2\pi} \\ \Delta x(1250 \text{ kg} \cdot \text{m/s}) &\geq \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{s}}{2\pi} \\ \Delta x &\geq 8.4 \times 10^{-38} \text{ m} \end{aligned}$$

It is not hard to find the automobile within  $10^{-38}$  meters. This uncertainty is too small to be noticeable.

An electron with a mass of  $10^{-30}$  kg moves at a speed of about  $6 \times 10^6$  m/s. Its momentum is  $6 \times 10^{-24}$  kg · m/s. Assume the same uncertainty in momentum as we did for the automobile—5%. The uncertainty in momentum would be  $0.3 \times 10^{-24}$  kg · m/s. The uncertainty in position would then be:

$$\begin{aligned} \Delta x \Delta p &\geq \frac{h}{2\pi} \\ \Delta x(0.3 \times 10^{-24} \text{ kg} \cdot \text{m/s}) &\geq \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{s}}{2\pi} \\ \Delta x &\geq 3.5 \times 10^{-10} \text{ m} \end{aligned}$$

This uncertainty is about the size of the atom. While it seems small to us, it is enormous to the electron. With such a large uncertainty, we would not be able to locate the electron within its atom.

The size of Planck's constant makes the uncertainty in position too small to notice for automobiles but too large to ignore for electrons. At the atomic level we begin to notice an inherent uncertainty or unknowability in nature.



"BUT YOU CAN'T GO THROUGH LIFE APPLYING HEISENBERG'S UNCERTAINTY PRINCIPLE TO EVERYTHING."

The size of Planck's constant limits the situations in which the uncertainty principle is noticeable. If we calculate the uncertainties in position and momentum associated with objects larger than subatomic particles, we find them to be incredibly small. For example, we can locate a grain of wheat within about a millimeter,  $10^{-3}$  meters (m). Heisenberg's uncertainty principle then predicts an uncertainty in momentum of no less than  $10^{-31}$  kilogram · meters/second ( $\text{kg} \cdot \text{m/s}$ ). As it is being blown through the air, the wheat grain might have a momentum of  $10^{-6}$   $\text{kg} \cdot \text{m/s}$ . The uncertainty,  $10^{-31}$   $\text{kg} \cdot \text{m/s}$ , is so small compared to the momentum itself that we cannot possibly notice it. It is, therefore, quite easy to follow the motion of the wheat grain.

On a more philosophical level, Heisenberg's principle can be interpreted as either an uncertainty or an indeterminacy. When we say that position and momentum are uncertain, we mean simply that we cannot measure them as precisely as we might like. The principle describes limitations of the observer, not of nature. When we say that position and momenta are indeterminate, we mean that we will never be able to measure them precisely. The limitation is one of nature. If you think of position as a measure of an object's "present" and momentum as a measure of its "future," then the Heisenberg uncertainty principle describes the inherent unknowability of the future based on an accurate knowledge of the present.

## CHAPTER SUMMARY

Atoms produce discrete spectra characteristic of the chemical element they represent. Light that has been emitted by an atom produces a series of colored lines against a black background, called an *emission spectrum*. Light that has been absorbed by an atom produces a series of black lines against a rainbowlike smear of color, called an *absorption spectrum*. The emission and



absorption lines occur at identical frequencies for all atoms of a given chemical element. Emission and absorption spectra serve as fingerprints for the chemical element that produced them.

Light spectra produced a significant clue as to the internal structure of the atom. Niels Bohr combined Rutherford's nuclear model of the atom with the information provided by emission spectra to develop a model of the atom. The *Bohr model* proposes that electrons orbit atomic nuclei much like planets orbit the sun. Unlike planetary orbits, however, only *selected electron orbits* are permitted. As an atom emits or absorbs energy, the electron jumps into orbits nearer or farther from the nucleus. The energy of the atom associated with these permitted orbits is quantized. The smallest orbit corresponds to the *ground state* of the atom. Higher orbits,  $n = 2$ ,  $n = 3$ , and so on, are called *excited states*. Bohr's model correctly predicts the centers of the emission and absorption lines for hydrogen.

Lasers can be described in terms of Bohr's discrete orbits. Lasers produce coherent light through the stimulated emission of radiation. *Coherent light* is light of a single wavelength that is in step. It is produced when atoms in an excited state are stimulated to emit a photon by a photon identical to it. Helium-neon lasers accomplish this by using excited helium atoms to excite a large number of neon atoms to the  $n = 5$  state. A photon with the same energy as the  $n = 5$  to  $n = 3$  transition will stimulate other neon atoms to emit photons in phase with it.

Matter waves can explain Bohr's discrete orbits in terms of *circular standing waves* produced by de Broglie wavelengths associated with the electron. The orbits selected by Bohr were, in fact, the only orbits for which a whole number of de Broglie wavelengths would fit into the circumference of the orbit. The use of matter waves to justify Bohr's discrete orbits led to the development of the *quantum mechanical model* of the atom. This model replaces discrete orbits with *electron probability clouds* that describe the probability of finding the electron at each location in space. The regions of greatest probability correspond to the radii of Bohr orbits. The quantum mechanical model is regarded as the most complete model of the atom.

Quantum mechanics has led to fundamental questions regarding the extent to which we can know the physical world. The *Heisenberg uncertainty principle* places limits on the preciseness with which we can simultaneously measure the location and momentum of a particle. The uncertainty in momentum times the uncertainty in position must always exceed Planck's constant divided by  $2\pi$ . In a sense, Heisenberg's uncertainty principle limits the extent to which we can know the present and future of an object simultaneously.

## ANSWERS TO SELF-CHECKS

- 18A.** The spectrum in Figure 18-2 shows emission lines characteristic of mercury and hydrogen.
- 18B.** a. There are six possible jumps: 4 to 1; 4 to 2; 4 to 3; 3 to 1; 3 to 2; and 2 to 1.

- b. Each jump produces a unique spectral line. Six spectral lines can be produced by these jumps.
- c. We can estimate the relative energy released in terms of the space between the energy level. In order from least to most energy released, we have 4 to 3, 3 to 1, 4 to 2, 2 to 1, 3 to 1, and 4 to 1.
- 18C.** Orbit A: One wavelength fits into the circumference; orbit C: Two wavelengths fit into the circumference; orbit E: Three wavelengths fit into the circumference.
- 18D.** Since the two waves have the same wavelength, the two electrons have the same momentum. The extent of the wave packet in (a) is narrower than that in (b), telling us that the particle's location is restricted more in (a) than in (b). The variation in amplitude of the wave in (a) tells us that the greatest probability for locating the electron is within the center of the packet. There is little variation in amplitude of the wave in (b). Consequently, the probability of finding the particle is about the same for all locations within the packet.

## PROBLEMS AND QUESTIONS

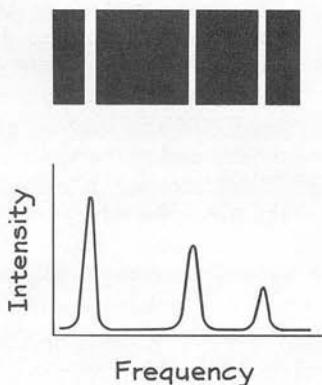
### A. Review of Chapter Material

- A1. Define the following terms:  
 Emission spectrum  
 Continuous spectrum  
 Discrete spectrum  
 Absorption spectrum  
 Bohr model  
 Ground state  
 Excited state  
 Coherent light  
 Stimulated emission  
 Laser  
 Wave packet  
 Quantum mechanical model  
 Probabilistic interpretation  
 Electron probability clouds  
 Heisenberg uncertainty principle  
 Energy level diagram
- A2. In what ways are the processes we use in playing games and solving puzzles similar to the processes scientists use in building models? In what ways are they different?
- A3. How are emission and absorption spectra different?
- A4. How can emission or absorption spectra be used to identify chemical elements?
- A5. How did Bohr's model of the atom draw upon:  
 a. Rutherford's nuclear model of the atom?  
 b. emission and absorption spectra?
- A6. How are jumps from one Bohr orbit to another related to the discrete lines in emission and absorption spectra?
- A7. Describe the role of helium atoms and neon atoms in the helium-neon laser.
- A8. How are de Broglie wavelengths used to explain Bohr's discrete orbits?
- A9. What characteristics of a wave packet describe:  
 a. the space within which the particle will be found?  
 b. the momentum of the particle?  
 c. the probability of finding the particle at each position inside the wave packet?
- A10. In the quantum mechanical model of the atom, what do the Bohr orbits represent?
- A11. In what ways do electron probability clouds explain emission and absorption spectra more completely than Bohr's orbits?
- A12. According to the Heisenberg uncertainty principle, if we measure the position of an object more precisely, what happens to our knowledge of its momentum?

### B. Using the Chapter Material

- B1. A lithium atom has a ground state energy of  $-8.0 \times 10^{-19}$  J. Use Bohr's relation-

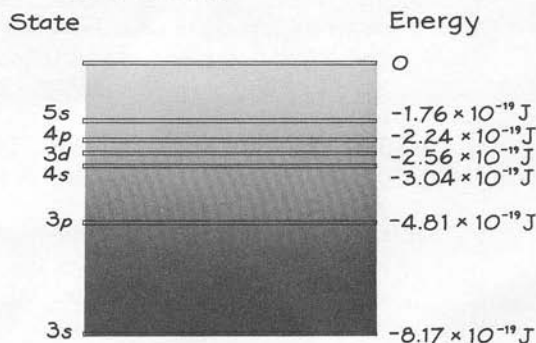
- ship to calculate the energy of the first two excited states.
- B2. An electron moves from one orbit, total energy of  $-0.137 \times 10^{-19}$  J, to a second orbit, total energy of  $-0.550 \times 10^{-19}$  J.
- What energy is associated with this transition?
  - Will the energy be emitted or absorbed by the atom?
- B3. An electron is initially in an orbit for which the atom has a total energy of  $-3.6 \times 10^{-18}$  J. The atom absorbs a photon with an energy of  $0.5 \times 10^{-18}$  J.
- What is the total energy of the atom now?
  - Will the electron be closer to or further from the nucleus?
- B4. An electron is in the  $n = 5$  orbit of a hydrogen atom. List all the possible ways that the electron can return to the  $n = 1$  orbit. What is the total energy associated with each possible route?
- B5. An electron orbit has a circumference of  $3 \times 10^{-10}$  m.
- List the series of de Broglie wavelengths that would fit into the circumference.
  - If the orbit corresponds to the  $n = 3$  state, which of the possible wavelengths in (a) represents the de Broglie wavelength for an electron in that orbit?
- B6. The figure below shows the emission spectrum taken with a photographic plate above a graph showing the intensity associated with each frequency.
- In what ways are the two spectra similar?
  - What additional information is provided by the lower spectrum?



- List the spectral lines in order from most probable to least probable transitions.
- B7. An atom has a relatively long-lived state at  $-3.1 \times 10^{-18}$  J. What energy photon is needed to stimulate emission to the  $-5.6 \times 10^{-18}$  J state?
- B8. Diffraction effects are the major limitations we face in locating a particle precisely. How is diffraction affected if we use light of longer wavelengths? Shorter wavelengths?
- B9. The energy exchanged between the particle and the quantum is the major limitation we face in measuring the momentum of a particle precisely. How are momentum measurements affected if we use light of longer wavelengths (lower frequencies)? Shorter wavelengths (higher frequencies)?

### C. Extensions to New Situations

- C1. Helium was first identified in the absorption spectrum of the sun before its presence was detected on earth. Use the emission spectra in Figure 18-1 to identify the location of the helium absorption lines.
- C2. Art historians are frequently faced with determining whether paintings are originals or forgeries. Suppose you have a painting that was supposedly painted in 1704. However, the colors look remarkably like pigments from cobalt blue, a paint mixture not used until 1804. How might you use the ideas presented in this chapter to determine whether the painting contains cobalt blue?
- C3. Sodium has 10 inner electrons held in orbit close to the nucleus and a single outer electron that behaves much like the single-electron hydrogen atom. The figure below shows the first six energy levels for the outer electron.

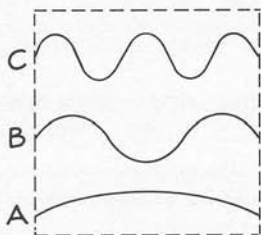


- a. How much energy would be required to remove the outer electron from the atom?
  - b. Sodium produces a characteristic yellow color because of the intensity of a spectral emission line of  $5.09 \times 10^{14}$  Hz. To what energy does this correspond?
  - c. Use the figure to determine which transition corresponds to the yellow emission line.
  - d. List the possible energies of photons that could be absorbed or released as the outer electron moves from one orbit to another.
- C4. Electrons can be released from gases much as they are released from metal plates in the photoelectric effect.
- a. Use Bohr's energy levels for hydrogen to predict the amount of energy needed to release the electron from the atom if the electron were in the  $n = 1$  orbit.
  - b. What would be the threshold frequency needed to release this electron?
  - c. If we illuminate hydrogen with light of  $4 \times 10^{15}$  Hz, how much kinetic energy will the released electrons have?
  - d. How would the threshold frequency change if we could somehow keep the hydrogen atoms in excited states?
- C5. Black-light posters use the process of fluorescence to emit light. When illuminated with black light, the poster gives off light.
- a. A black light emits frequencies in the ultraviolet region of the spectrum. Suppose photons with energy of  $5.0 \times 10^{-19}$  J are incident on the poster. If its atoms need only  $4.0 \times 10^{-19}$  J to move into an excited state, how much energy will be left over from each interaction?
  - b. This leftover energy is released. Based on its frequency, in what form will this energy be released? (An electromagnetic spectrum can be found in Chapter 15.)
  - c. The excited atom now returns to its ground state. What frequency of light is emitted? Is it visible?
  - d. Compare the frequency of the incident light (black-light source) with the frequency of the light emitted by the poster. Explain how the process of fluorescence works.
- C6. Many fluorescent lamps are mercury-vapor lamps with a fluorescent coating along the outside. The mercury-vapor lamp emits ultraviolet light, as well as visible frequencies of light.
- a. Use the results in Problem C5 to explain how the ultraviolet frequencies can be converted into visible frequencies.
  - b. Does the fluorescent coating increase or decrease the amount of visible light available?
  - c. Many large mercury lamps place the fluorescent coating along the inside of the container and a mercury light source inside the container. What happens if the container breaks but the mercury light source continues to operate? In these situations, why do people frequently get sunburns? (Ultraviolet light is the primary cause of sunburns.)
- C7. The first laser used a ruby crystal that contained chromium impurities. This crystal was surrounded by a flash tube, which emitted sudden flashes of bright green light.
- a. The light emitted by the flash was absorbed by the chromium atoms. Based upon the frequency of green light, approximately how much energy did the chromium atoms absorb?
  - b. After excitation, the chromium electrons would spontaneously drop to a lower energy state but not to ground state. This intermediate state was long-lived, allowing it to serve as a state for stimulated emission. How could electrons be stimulated to move from this state?
  - c. The light emitted by the chromium atoms when they moved from the intermediate state to ground state was the light characteristically emitted by the laser. Will this light be in the red end or blue end of the spectrum? Why?
  - d. To get laser action to occur, the ends of the ruby tube reflect light, but the sides do not. Explain why.
- C8. Another light-emission process is phosphorescence. In this case a material that has been exposed to light will continue to emit



light long after the original light source has been removed. Use the model of the atom to explain how this could occur. (The discussion of long-lived states used in lasers will be helpful.)

- C9. Standing waves can exist for particle waves. One example is a particle that is confined to a box. Since the particle cannot leave the box, the amplitude of its matter wave must be zero at the box walls and outside the box. The figure below shows three examples of wave functions for a particle confined to a box.
- Which of these wave functions represent the particle with the greatest momentum? With the least momentum?
  - Why can wavelengths with values between those of A and B or between those of B and C not exist in this box?
  - Why can momenta between those represented by the waves of A, B, and C not exist in the box?



- What conclusions can you draw about the particle momenta that can exist in the box? How does this relate to the electron momenta that can exist in the atom?

#### D. Activities

- A diffraction grating will separate light into its separate wavelengths much like a prism. See if you can borrow an inexpensive diffraction grating. Take it outside and view the various lights you see. Describe the spectra and compare them to spectra in the book or emitted by laboratory sources. Identify the types of gases in the lamps.
- Tape a diffraction grating on the front of a camera. Use high-speed color film to take photographs of the spectra of interesting lamps.
- George Gamow's stories "Quantum Billiards" and "Quantum Jungle" (see Activity D1 in Chapter 17) give you some concept of how Heisenberg's uncertainty principle would affect our experience in the macroscopic world if Planck's constant were a larger number. You might want to read these stories again.