38 The Atom and the Quantum

The final unit of this book is about the realm of the unimaginably tiny atom. This chapter investigates atomic structure, which is revealed by analyzing light. Light has a dual nature, which in turn radically alters our understanding of the atomic world. The next chapter covers the structure of the atomic nucleus and radioactivity, and the concluding chapter is about the nuclear processes of fission and fusion.

38.1 Models

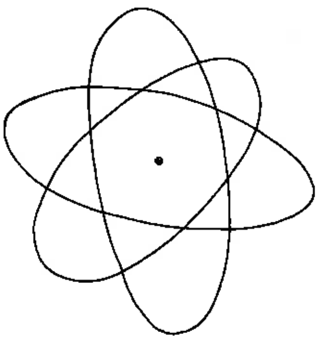

Fig. 38-1 The old planetary model of an atom with electrons orbiting like little planets around a tiny sun.

Nobody knows what an atom looks like, or even if it makes sense to suppose it has an appearance. To visualize the processes that occur at the subatomic realm, we construct models. In the planetary model—the one that most people think of when they picture an atom—the electrons orbit the nucleus like planets going around the sun. This was an early model of the atom suggested by the Danish physicist Niels Bohr in 1913. We still tend to think in terms of this simple picture, even though it has been replaced several times over by models that give progressively better results but become progressively more abstract. Old models sometimes steer us off track in our investigations of nature, and sometimes they provide a scaffolding that allows us to progress to more complicated models. Today's models of the atom, for example, have given way to mathematical theories that cannot be represented by pictures.

A useful model of the atom must be consistent with a model for light. All light has its source in the motion of electrons within the atom. Down through the centuries there have been two primary models of light: the particle model and the wave model. Isaac Newton believed in a particle model of light. He thought that light was composed of a hail of tiny particles. Christian Huygens stated that light was a wave phenomenon. About a century later this was reinforced when Thomas Young demonstrated interference. Later, Maxwell proposed and Hertz proved that light is a part of the electromagnetic spectrum. This seemed to verify the wave nature of light once and for all. But Maxwell's electromagnetic wave model was not the last word on the nature of light. In 1905 Albert Einstein resurrected the particle theory of light.

38.2 Light Quanta

Einstein viewed light as being made up of a hail of tiny particles—concentrated bundles of electromagnetic energy. Einstein extended the idea of a German physicist, Max Planck, who a few years earlier had proposed that energy in an atom is not continuous, but occurs in little chunks he called **quanta** (plural of *quantum*). Planck suggested that when an atom emits light, the energy of the atom changes by quantized amounts. Einstein carried this idea further and proposed that light is composed of quanta also. He called these quanta **photons**.

A quantum is an elemental unit—a smallest amount of something. The idea that certain quantities are quantized—that they come in discrete (separate) units—was known in Einstein's time. Matter is quantized. The mass of a gold ring, for example, is equal to some whole-number multiple of the mass of a single gold atom. Electricity is quantized, as all electric charge is some whole-number multiple of the charge of a single electron.

The newer physics tells us that other quantities are also quantized—quantities such as energy and angular momentum. The energy in a light beam is quantized and comes in packets, or quanta; only a whole number of quanta can exist. The quanta of light, or of electromagnetic radiation in general, are the photons.

The energy of a photon is directly proportional to its frequency. When the energy E of a photon is divided by its frequency f, the quantity that results is always the same, no matter what the frequency. This quantity is a constant known as **Planck's constant**, h. The energy of every photon is therefore

$$E = hf$$

This equation gives the smallest amount of energy that can be converted to light of frequency f. Light is not emitted continuously, but is emitted as a stream of photons, each with an energy hf.

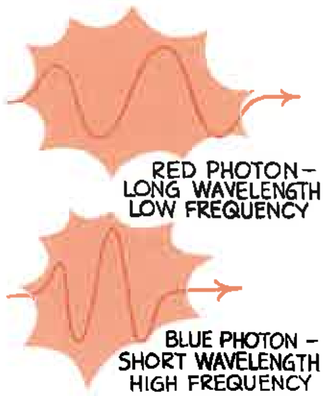

Fig. 38-2 The energy of a photon of light is proportional to its vibrational frequency.

38.3 The Photoelectric Effect

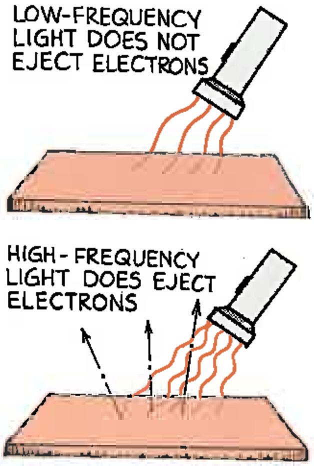

Fig. 38-3 The photoelectric effect depends on the frequency of light.

Einstein was led to the quantum theory of light by his study of the photoelectric effect. The **photoelectric effect** is the ejection of electrons from certain metals when light falls upon them. These metals are said to be *photosensitive* (that is, sensitive to light). This effect is used in electric eyes, in the photographer's light meter, and in the sound tracks of motion pictures.

Early investigators knew that high-frequency light, even from a dim source, was capable of ejecting electrons from a photosensitive metal surface; yet low-frequency light, even from a very bright source, could not dislodge electrons. They knew that bright light carried more energy than dim light. It was thought that very bright red light, for example, should dislodge electrons more easily than dim blue or violet light could. But this was not the case. Only light of high frequencies was capable of supplying sufficient energy to eject electrons from the metals.

Einstein explained the photoelectric effect by thinking of light in terms of photons. The absorption of a photon by an atom in the metal surface is an all-or-nothing process. One and only one photon is completely absorbed by each electron ejected from the metal. This means that the *number* of photons that hit the metal has nothing to do with whether a given electron will be ejected. If the energy per photon is too small, then the brightness or intensity of light does not matter. The critical factor is the frequency or color of the light. A few photons of blue or violet light can eject a few electrons, but hordes of red or orange photons cannot eject a single electron. Only high-freqency photons have the concentrated energy needed for ejection.

A wave has a broad front, and its energy is spread out along this front. For the energy of a light wave to be concentrated enough to eject a single electron from a metal surface is as unlikely as for an ocean wave to hit a beach and knock a single seashell far inland with an energy equal to the energy of the whole wave. The photoelectric effect suggests that we think of light as a succession of particle-like photons, rather than as a continous train of waves. The number of photons in a light beam controls the brightness of the beam, while the energy of each photon is proportional to the frequency of the light.

Experimental verification of Einstein's explanation of the photoelectric effect was made 11 years later by the American physicist Robert Millikan. Every aspect of Einstein's interpretation was confirmed. It was for this (and not for his theory of relativity) that Einstein received the Nobel prize.

▸ **Questions**

1. Will brighter light eject more electrons from a photosensitive surface than dimmer light of the same frequency?
2. Will high-frequency light eject a greater number of electrons than low-frequency light?

38.4 Waves as Particles

Figure 38-4 is a striking example of the particle nature of light. The photograph was taken with exceedingly feeble light. Each frame shows the image progressing photon by photon. Note also that the photons seem to strike the film in an independent and random manner.

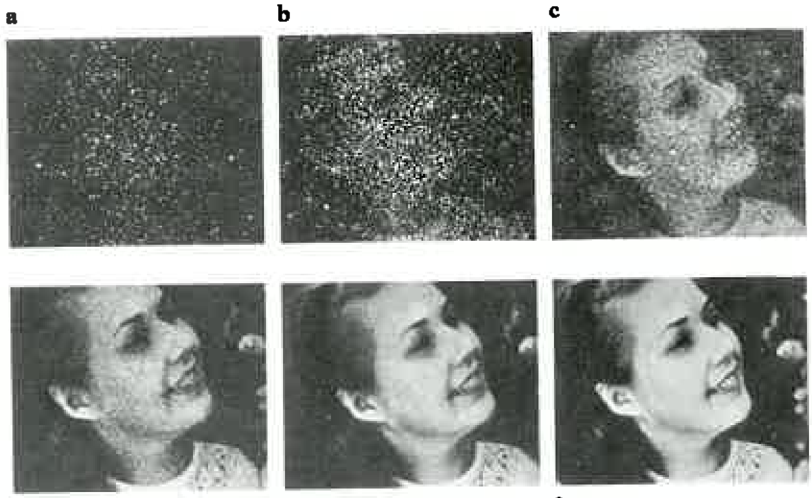

Fig. 38-4 Stages of exposure reveal the photon-by-photon production of a photograph. The approximate numbers of photons at each stage were (a) 3×10^3, (b) 1.2×10^4, (c) 9.3×10^4, (d) 7.6×10^5, (e) 3.6×10^6, (f) 2.8×10^7.

▸ **Answers**

1. Yes (if the frequency is great enough for any electrons to be ejected). The number of ejected electrons depends on the number of incident photons.
2. Not necessarily. The answer is yes if electrons are ejected by the high-frequency light but not by the low-frequency light, because its photons do not have enough energy. If the light of both frequencies can eject electrons, then the number of electrons ejected depends on the brightness of the light, not on its frequency.

38.5 Particles as Waves

If light can have particle properties, cannot particles have wave properties? This question was posed by the French physicist Louis de Broglie in 1924, while he was still a student. His answer gave him a Ph.D. in physics, and later won him the Nobel prize in physics.

De Broglie suggested that all matter could be viewed as having wave properties. All particles—electrons, protons, atoms, bullets, and even humans—have a wavelength that is related to the momentum of the particles by

$$\text{wavelength} = \frac{h}{\text{momentum}}$$

where h is, lo and behold, Planck's constant again. The wavelength of a particle is called the *de Broglie wavelength*. A particle of large mass and ordinary speed has too small a wavelength to be detected by conventional means. However, a tiny particle—such as an electron—moving at high speed has a detectable wavelength.* It is smaller than the wavelength of visible light but large enough for noticeable diffraction. A beam of electrons, interestingly enough, behaves like a beam of light. It can be diffracted and undergoes wave interference under the same conditions that light does (Figure 38-5).

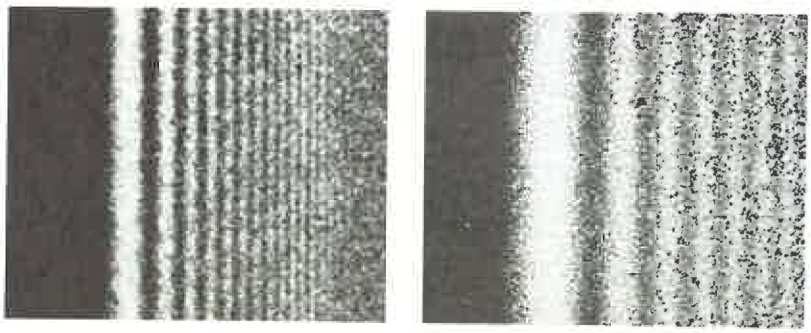

Fig. 38-5 Fringes produced by the diffraction (left) of an electron beam and (right) of a light.

* A bullet of mass 0.10 kg traveling at 330 m/s, for example, has a de Broglie wavelength of $(6.6 \times 10^{-34}\ \text{J·s})/[(0.02\ \text{kg}) \times (330\ \text{m/s})] = 10^{-34}$ m, an incredibly small size that is a million million million millionth the diameter of a hydrogen atom. An electron traveling at 2 percent of the speed of light, on the other hand, has a wavelength of 10^{-10} m, which is equal to the diameter of the hydrogen atom. Diffraction effects for electrons are measurable, whereas diffraction effects for bullets are not.

An electron microscope makes practical use of the wave nature of electrons. The wavelength of electron beams is typically thousands of times shorter than the wavelength of visible light, so the electron microscope is able to distinguish detail not possible with optical microscopes (Figure 38-6).

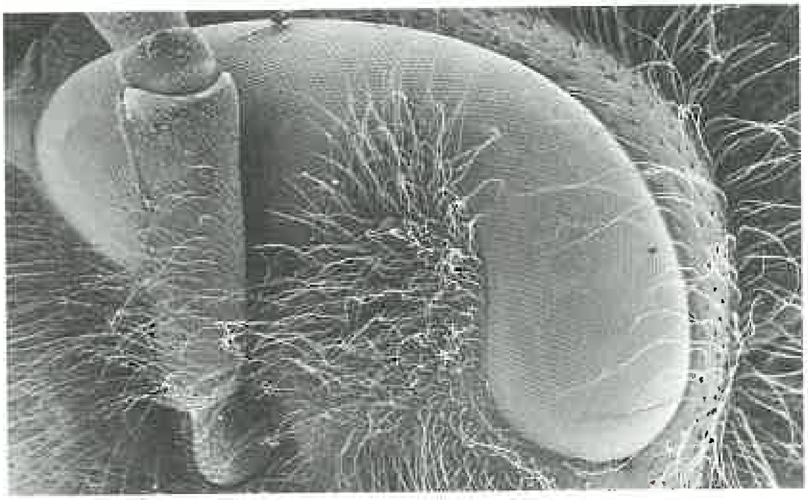

Fig. 38-6 Detail of a wasp eye as seen with a scanning electron microscope at a "low" magnification of 200 times.

38.6 Electron Waves

More far-reaching than the diffraction of electrons is de Broglie's model of matter waves in the atom. The planetary model of the atom developed by Niels Bohr was useful in explaining the atomic spectra of the elements. It explained why elements emitted only certain frequencies of light. An electron has different amounts of energy when it is in different orbits around the nucleus. From an energy point of view, an electron is said to be in different *energy levels* when it is in different orbits. The electrons in an atom normally occupy the lowest energy levels available.

An electron is boosted by various means to higher energy levels. As it returns to its stable level, it emits a photon. The energy of the photon is exactly equal to the difference in the energy levels in the atom. The characteristic pattern of lines in the spectrum of an element corresponds to electron transitions between the energy levels characteristic of the atoms of that element. By examining spectra, physicists were able to determine the various energy levels in the atom. This was a tremendous triumph for atomic physics.

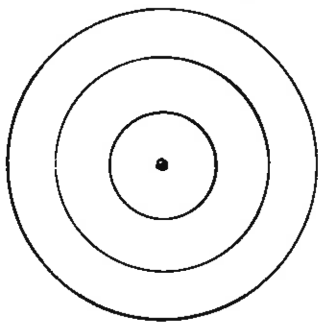

Fig. 38-7 The Bohr model of the atom.

One of the difficulties of this model of the atom, however, was reconciling why electrons occupied certain energy levels in the atom. Evidence was that they occupied one level or the other, but never a level in between. The orbital paths of electrons in the atom seemed to be discrete. That is, they were at definite distances from the atomic nucleus.

This seemed strange. There is nothing to suggest, as a comparative example, that a body couldn't orbit at any distance from the sun. There are no places in the solar system where orbiting is forbidden. Why, then, were planetary orbits in the atom so different? The planetary model of the atom was on shaky ground. This was because the electron was considered to be a particle, a tiny BB whirling around the nucleus like a planet whirling around the sun.

Why the electron occupies only discrete levels is understood by considering the electron to be not a particle but a *wave*. According to de Broglie's theory of matter waves, an orbit exists where an electron wave closes in on itself in phase. In this way it reinforces itself constructively in each cycle, just as the standing wave on a music string is constructively reinforced by its successive reflections. In this view, the electron is visualized not as a particle located at some point in the atom, but as though its mass and charge were spread out into a standing wave surrounding the nucleus. The wavelength of the electron wave must fit evenly into the circumferences of the orbits (Figure 38-8).

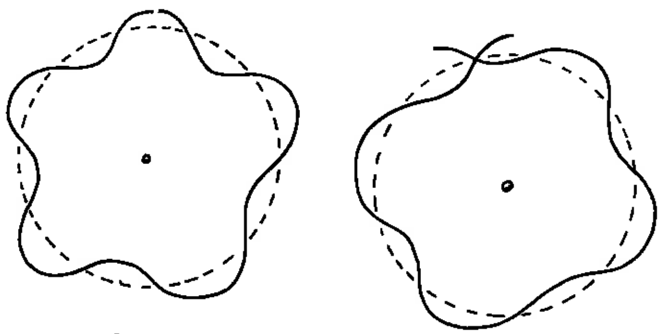

Fig. 38-8 (a) Orbital electrons form standing waves only when the circumference of the orbit is equal to a whole-number multiple of wavelengths. (b) When the wave does not close in on itself in phase, it undergoes destructive interference.

The circumference of the innermost orbit is equal to one wavelength of the electron wave. The second orbit has a circumference of two electron wavelengths, the third three, and so on (Figure 38-9). 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.*

* Electron wavelengths are successively longer for orbits of increasing radii; so for a more accurate analogy, the construction of longer necklaces requires using not only *more* paper clips, but *larger* paper clips as well.

Since the circumferences of electron orbits are discrete, it follows that the radii of these orbits, and hence the energy levels, are also discrete.

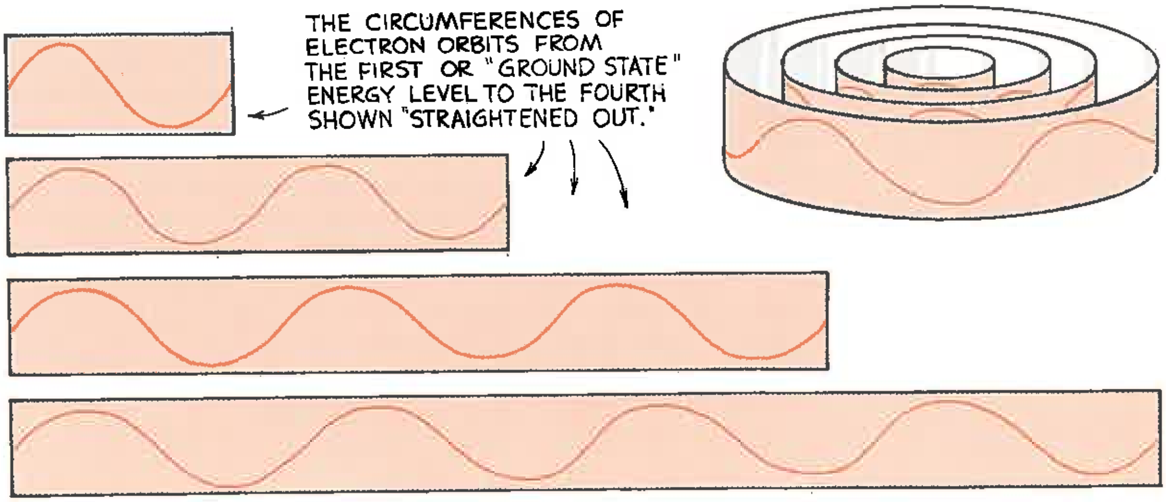

Fig. 38-9 The electron orbits in an atom have discrete radii because the circumferences of the orbits are whole-number multiples of the electron wavelengths, which differ for the various elements (and also for different orbits within the elements). This results in discrete energy levels, which characterize each element. The figure is greatly oversimplified, as the standing waves make up spherical and ellipsoidal shells rather than flat, circular ones.

This view explains why electrons do not spiral closer and closer to the nucleus when photons are emitted. If each electron orbit is described by a standing wave, 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.

38.7 Relative Sizes of Atoms

The radii of the electron orbits in the Bohr model of the atom are determined by the amount of electric charge in the nucleus. For example, the single positively charged proton in the hydrogen atom holds one negatively charged electron in an orbit at a particular radius. If we double the positive charge in the nucleus, the orbiting electron will be pulled into a tighter orbit with half its former radius since the electrical attraction is doubled. This doesn't quite happen, however, because the double charge in the nucleus normally holds a second orbital electron, which diminishes the effect of the positive nucleus. This added electron makes the atom electrically neutral. The atom is no longer hydrogen, but is helium. The two orbital electrons assume an orbit characteristic of helium. An additional proton in the nucleus pulls the electrons into an even closer orbit and, furthermore, holds a third electron in a second orbit. This is the lithium atom, atomic number 3. We can continue with this process, increasing the positive charge of the nucleus and adding successively more electrons and more orbits all the way up to atomic numbers above 100, to the synthetic radioactive elements.*

As the nuclear charge increases and additional electrons are added in outer orbits, the inner orbits shrink in size because of the stronger electrical attraction to the nucleus. This means that the heavier elements are not much larger in diameter than the lighter elements. The diameter of the uranium atom, for example, is only about three hydrogen diameters even though it is 238 times more massive. The schematic diagrams in Figure 38-10 are drawn approximately to the same scale.

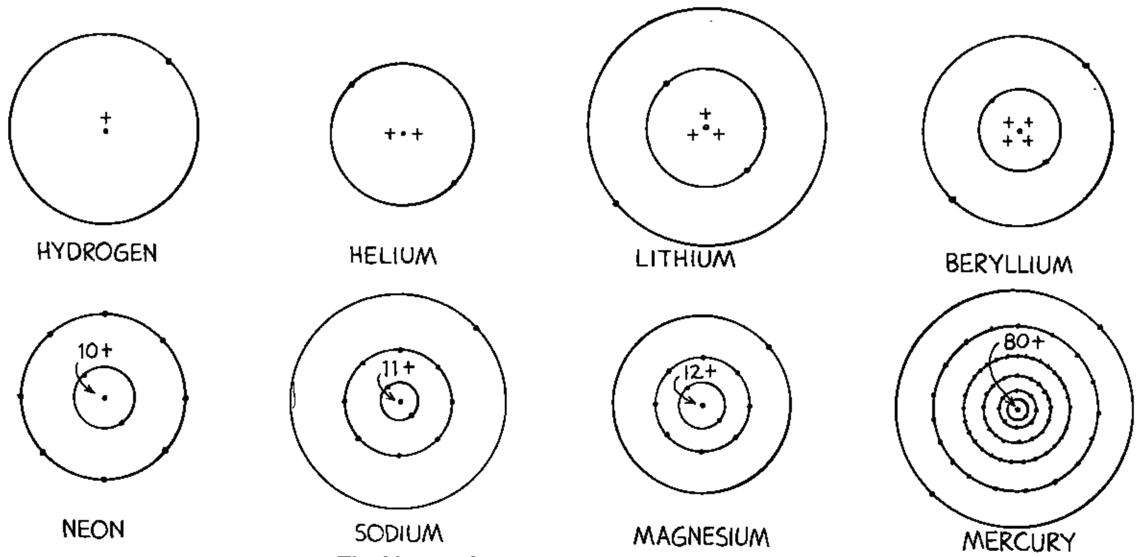

Fig. 38-10 The orbital model illustrated for some light and heavy atoms drawn to approximate scale. Note that the heavier atoms are not appreciably larger than the lighter atoms.

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* Each orbit will hold only so many electrons. A rule of quantum mechanics states that an orbit is filled when it contains a number of electrons given by $2n^2$, where n is 1 for the first orbit, 2 for the second orbit, 3 for the third orbit, and so on. For $n = 1$, there are 2 electrons; for $n = 2$, there are $2(2^2)$, or 8, electrons; for $n = 3$, there are a maximum of $2(3^2)$, or 18, electrons, etc. The number n is called the *principal quantum number*.

▶ **Question**
What fundamental force dictates the size of an atom?

Each element has an arrangement of electron orbits unique to that element. For example, the radii of the orbits for the sodium atom are the same for all sodium atoms, but different from the radii of the orbits for other kinds of atoms. When we consider the 92 naturally occurring elements, we find that there are 92 distinct patterns or orbits. There is a different pattern for each element.

The Bohr model of the atom solved the mystery of the atomic spectra of the elements. It accounted for X rays that were emitted when electrons made transitions from outermost to innermost orbits. Bohr was able to predict X-ray frequencies that were later experimentally confirmed. He calculated the *ionization energy* of the hydrogen atom—the energy needed to knock the electron out of the atom completely. This also was verified by experiment. The Bohr model accounted for the general chemical properties of the elements and predicted properties of a missing element (hafnium), which led to its discovery.

The Bohr model was impressive. Nonetheless, Bohr was quick to point out that his model was to be interpreted as a crude beginning, and the picture of electrons whirling like planets about the sun was not to be taken literally (to which popularizers of science paid no heed). His discrete orbits were conceptual representations of an atom whose later description involved a purely mathematical rather than a visual model. Still, his planetary model of the atom with electrons occupying discrete energy levels underlies the more complex models of the atom today, which are built upon a completely different structure from that built by Newton and other physicists before the twentieth century. This is the structure called *quantum mechanics*.

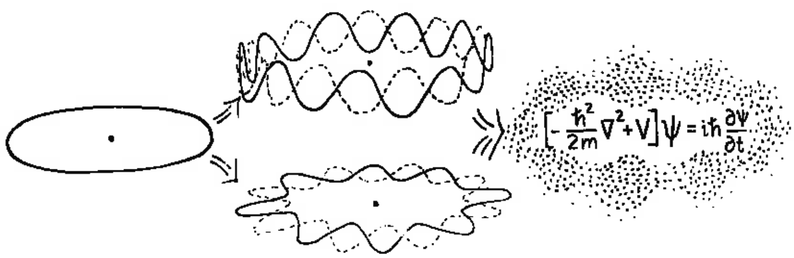

Fig. 38-11 The model of the atom has evolved from the Bohr planetary model (left) to the modified model with de Broglie waves (center) to the mathematical model (the Schroedinger wave equation, right).

▶ **Answer**
The electrical force.

38.8 Quantum Physics

The more that physicists studied the atom, the more convinced they became that the Newtonian laws that work so well for large objects such as baseballs and planets simply do not apply to events in the microworld of the atom. Whereas in the macroworld the study of motion is called *mechanics*, in the microworld the study of the motion of quanta is called **quantum mechanics**. The more general study of quanta in the microworld is simply called **quantum physics**.

While one can be quite certain about careful measurements in the macroscopic world, there are fundamental uncertainties in the measurements of the atomic domain. Macroscopic measurements, such as the temperature of materials, the vibrational frequencies of certain crystals, and the speeds of light and sound, can be made as accurate as the experimenter wishes. But subatomic measurements such as the momentum and position of electrons in an atom and the decay rate of individual radioactive atoms are entirely different. In this domain, the uncertainties in many measurements are comparable to the magnitudes of the quantities themselves. The structure of quantum mechanics is based on probabilities, a notion that is difficult for many people to accept. Even Einstein did not accept this, which prompted his often-quoted statement, "I cannot believe that God plays dice with the universe."

If you continue to expose yourself to physics, you will likely study quantum mechanics in the future. Then you'll find that subatomic interactions are unpredictable, and that the notion of certainty is replaced with probability. It's fascinating material.

38 Chapter Review

Concept Summary

Early models attempted to explain atoms and light.
- In Bohr's planetary model of the atom, electrons orbit the nucleus.
- Newton proposed a particle model of light.
- Huygens' wave model of light was reinforced by Maxwell's electromagnetic wave model.

Einstein proposed that light is composed of discrete quanta of energy called photons.
- The energy of a photon is proportional to its frequency.
- Energy and frequency are related by Planck's constant, h.

The photoelectric effect, the ejection of electrons from certain metals that are struck by light, reinforced the particle theory of light.
- When the frequency of light is so low that the photons have insufficient energy to cause the effect, then increasing the intensity of the light does not matter.

De Broglie suggested that all matter has wave properties.
- Electrons have a detectable wavelength, can be diffracted, and undergo interference.

Bohr's planetary model explained the atomic spectra of the elements.
- Spectral lines are due to transitions of electrons between energy levels that correspond to different electron orbits.
- Bohr's model could not explain why only certain electron energy levels were possible, but de Broglie's concept of unique electron wavelengths for each element could.

The radius of the atoms of each element is unique to that element.
- As the nuclear charge increases, inner electron orbits shrink. As a result, heavier elements are not much larger in diameter than lighter elements.

Quantum mechanics is the study of motion of quanta.
- Newtonian mechanics does not apply to events on the atomic scale.
- Quantum measurements, such as the momentum and position of an electron, involve fundamental uncertainties.

Important Terms

photoelectric effect (38.3)
photon (38.2)
Planck's constant (38.2)
quantum (pl. quanta) (38.2)
quantum mechanics (38.8)
quantum physics (38.8)

Review Questions

1. What is a model? Give two examples for the nature of light. (38.1)
2. What exactly is a quantum? Give two examples. (38.2)
3. What is a quantum of light called? (38.2)
4. What is Planck's constant, and how does it relate to the frequency and energy of a quantum of light? (38.2)
5. Which has more energy per photon—red light or blue light? (38.2)

6. What is the photoelectric effect? (38.3)

7. Why does blue light eject electrons from a photosensitive surface, whereas red light has no effect? (38.3)

8. Will bright blue light eject more electrons than dim light of the same frequency? (38.3)

9. Does the photoelectric effect support the particle model or the wave model of light? (38.3)

10. a. Do particles have wave properties?
    b. Who was the first physicist to give a convincing answer to this question? (38.4)

11. As the speed of a particle increases, does its associated wavelength increase or decrease? (38.5)

12. Does the diffraction of an electron beam support the particle model or the wave model of electrons? (38.5)

13. How does the energy of a photon compare to the difference in energy levels of the atom from which it is emitted? (38.6)

14. What does it mean to say that an electron occupies discrete energy levels in an atom? (38.6)

15. Does the particle view of an electron or the wave view of an electron better explain the discreteness of electron energy levels? Why? (38.6)

16. What does wave interference have to do with the electron energy levels in an atom? (38.6)

17. Why is a helium atom smaller than a hydrogen atom? (38.7)

18. Why are the heaviest elements not appreciably larger than the lightest elements? (38.7)

19. What is quantum mechanics? (38.8)

20. Can the momentum and position of electrons in an atom be measured with certainty? (38.8)

Think and Explain

1. What does it mean to say that a certain quantity is quantized?

2. What evidence can you cite for the wave nature of light? For the particle nature of light?

3. A very bright source of red light has much more energy than a dim source of blue light, but the red light has no effect in ejecting electrons from a photosensitive surface. Why is this so?

4. Which photon has the most energy—one from infrared, visible, or ultraviolet light?

5. If a beam of red light and a beam of green light have exactly the same energy, which beam contains the greater number of photons?

6. Suntanning produces cell damage in the skin. Why is ultraviolet light capable of producing this damage while infrared radiation is not?

7. Electrons in one electron beam have a greater speed than those in another. Which electrons have the longer de Broglie wavelength?

8. We do not notice the wavelength of moving matter in our ordinary experience. Is this because the wavelength is extraordinarily large or extraordinarily small?

9. The equation $E = hf$ describes the energy of each photon in a beam of light. If Planck's constant, h, were larger, would photons of light of the same frequency be more energetic or less energetic?

10. Why will helium rather than hydrogen more readily leak through an inflated rubber balloon?