CHAPTER 13

Probing the Atom

13.1 THE PERIODIC TABLE

It has been hypothesized for millennia that all matter is made of tiny, indivisible, smallest bits of matter called atoms. Great progress had been made during the nineteenth century in attributing the thermodynamic laws and some of the properties of matter, especially gases, to the kinetic-molecular theory (Chapter 7). In addition, it was known for centuries that there are different types of so-called fundamental "elements" in nature—gold, silver, copper, sodium, etc. There are the smallest units into which substances can be divided by chemical means. Eventually it was found useful to give the elements special symbols, for example, "C" for carbon, "O" for oxygen, "H" for hydrogen, and so on.

For many people, such as the English chemist John Dalton (1766–1844), these different elements indicated that nature is also made up of different types of atoms, one type of atom for each element. Each element was considered a collection of identical, indestructible atoms, and this idea was confirmed in chemical studies during the nineteenth century. When two or more atoms link together, they form a molecule. The molecule may be an element itself if both atoms are the same, such as O_2, or it may be a com-

FIGURE 13.1 John Dalton (1766–1844). Born in Eaglesfield, England, the son of a weaver, Dalton attended a Quaker school until the age of 12. He taught at New College, Manchester. He is best known for his theory that matter is composed of atoms of differing weights that combine in simple, whole-number ratios by weight. He listed the atomic weights of known elements relative to the weight of hydrogen. Dalton's work formed the basis of the periodic table of the elements.

pound if the atoms are different, such as H_2O. Since atoms are not divisible, the idea of joining two atoms of hydrogen to, say, 1½ atoms of oxygen instead of exactly one atom of oxygen is meaningless. Dalton's law of fixed proportions follows quite naturally from the hypothesis that elements are made up of identical, indestructible atoms. When elements combine to form compounds, precisely the same ratio of masses of the constituents is required each time to make a particular compound. For example, 23.0 g of sodium always combine with 35.5 g of chlorine to produce 58.5 g of salt. If you start with 25.0 g of sodium and 35.5 g of chlorine, you still get only 58.5 g of salt, but there will be 2.0 g of sodium left over. This rule, the *law of fixed proportions*, applies to all compounds.

The law of fixed proportions indicated that the atoms of the elements that combine to form compounds such as NaCl do not have the same mass. Since only 23.0 g of sodium combined with 35.5 g of chlorine to form 58.5 g of NaCl, the chlorine atoms must be more massive than the sodium atoms. In fact, if one sodium atom has 23.0 "units" of mass, then one chlorine atom would have 35.5 units of mass. The units of mass could be grams, kilograms, pounds, or whatever. For convenience, they are called *atomic mass units*, for which the standard symbol is u.

We now know that there are just 92 different elements that occur naturally throughout nature, and in fact throughout the entire Universe. These have a variety of atomic mass units (amu, abbreviated to u). Today, the atomic mass units are measured relative to one type, or isotope, of carbon known as carbon-12 (^{12}C), which is *defined* as having atomic mass of 12.0000 u. The atomic masses of all other elements can then be given relative to carbon-12. For instance, relative to carbon-12, natural sodium has an atomic mass of 22.99 u, hydrogen has an atomic mass of 1.008 u, oxygen has an atomic mass of 15.999 u, uranium has a mass of 238.03 u, and so on. Some of these elements have similar properties: some are gases at room temperature; some are solid metals. There are differences in densities, melting and boiling points, electrical conductivity, and so on.

In the 1869, the Russian chemist Dimitri Mendeleev (1834–1907) arranged the 63 elements then known in a table according to their physical properties from the lightest (hydrogen) to the heaviest (uranium). He found that when arranged by chemical properties, the elements tended to line up vertically and to vary horizontally in a periodic fashion, forming what is now called the *periodic table*. Although some elements were missing at the time and were discovered only later, Mendeleev courageously assigned a position on the table to each element known to him, leaving blanks

FIGURE 13.2 Dimitri Mendeleev (1834–1907), founder of the modern periodic table.

where he hypothesized that an element had not yet been discovered. These elements were discovered later, confirming Mendeleev's hypothesis!

Only one element can occupy a particular position on the periodic table. Mendeleev assigned a number to each element at each position. This number, called the *atomic number*, now goes from 1 for hydrogen to 92 for uranium. It is given the symbol Z. The atomic number immediately identifies the element, since it is unique for each element. It is always an integer and is given at the top of the space in most periodic tables. Thus, when someone refers to element $Z = 3$, we know immediately that they are referring to the element lithium; element 53 is iodine, and so on.

When the elements are arranged on the modern periodic table as shown in the color insert, Plate 6, the elements below one another in each column or group share physical properties to a remarkable degree, as Mendeleev had discovered. Therefore, these elements can be considered to belong to the same "family" of elements. For instance, Group I on the left contains the family of *alkali* metals: lithium, sodium, potassium, rubidium, and cesium. This is a group of soft metals with very low densities, low melting points, and similar chemical behavior. Another family of elements, called the *halogens*, is found in Group VII: fluorine, chlorine, bromine, and iodine. These elements combine violently with many metals and form

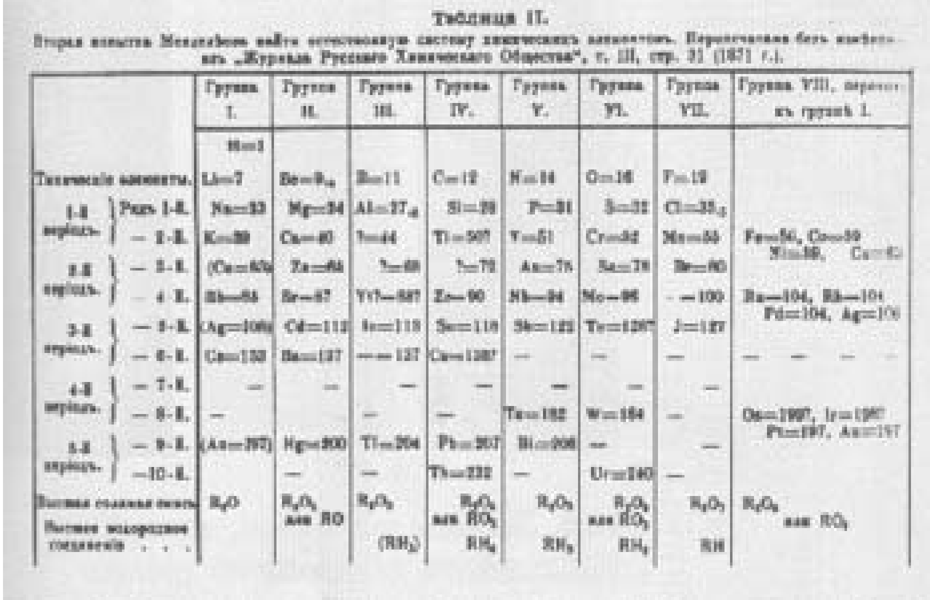

FIGURE 13.3 Mendeleev's original periodic table of elements as it appeared in 1872.

white, crystalline salts (*halogen* means "salt former"). These salts have similar formulas, such as NaF, NaCl, NaBr, $MgCl_2$, $MgBr_2$, etc.

Occasionally, for reasons then not known, it was necessary to depart from the overall scheme of ordering the elements. For example, the chemical properties of argon (Ar) and potassium (K) demand that they be placed in the eighteenth and nineteenth positions in order to fall into groups characteristic of their families. On the basis of their atomic masses alone (39.948 u for argon; 39.102 u for potassium) their positions would have been reversed.

The beautiful regularity and symmetry of the periodic table indicate that a model of the atoms that make up the different elements will probably also display a similar beauty and symmetry. Such a model did arise during the early twentieth century, and it did not disappoint our expectations. But we shall see that an enormous amount of ingenuity and detective work was required to comprehend the structure of the atom.

13.2 THE IDEA OF ATOMIC STRUCTURE

Chemistry in the nineteenth century had succeeded remarkably in accounting for combining proportions and in predicting chemical reactions. This success had convinced most scientists that matter is indeed composed of atoms. But there remained a related question: Are atoms really indivisible, as had been assumed, or do they consist of still smaller particles?

You can see how this question arose by thinking a little more about the periodic table. Mendeleev had arranged the elements in the order of increasing atomic mass. But the atomic masses of the elements cannot explain the *periodic* features of Mendeleev's table.

Why, for example, do the 3rd, 11th, 19th, 37th, 55th, and 87th elements, with quite different atomic masses, have similar chemical properties (e.g., they burn when exposed to air)?

Why are these properties somewhat different from those of the 4th, 12th, 20th, 38th, 56th, and 88th elements in the list (which react slowly with air or water), but greatly different from the properties of the 2nd, 10th, 18th, 36th, 54th, and 86th elements (which rarely combine with any other element)?

The periodicity in the properties of the elements led to speculation that atoms might have structure, that they might be made up of smaller pieces. The properties changed gradually from group to group. This fact suggested that some unit of atomic structure might be added from one element to the next, until a certain portion of the structure is completed. The completed condition would occur in the atom of a noble gas (Group VIII in Plate 6). In an atom of the next heavier element, a new portion of the structure would be started, and so on. The methods and techniques of classical chemistry could not supply experimental evidence for such structure. In the nineteenth century, however, discoveries and new techniques in physics opened the way to prove that atoms do actually consist of smaller pieces. Evidence piled up to support the conclusion that the atoms of different elements differ in the number and arrangement of these pieces.

13.3 CATHODE RAYS

In 1855, the German physicist Heinrich Geissler invented a powerful vacuum pump. This pump could remove enough gas from a strong glass tube to reduce the pressure to 0.01% of normal air pressure. It was the first major improvement in vacuum pumps after Guericke's invention of the air pump, two centuries earlier. So we noted, using Geissler's new pump made possible the electric light bulb, the electron tube, and other technologically valuable inventions over the next 50 years. It also opened new fields to pure scientific research.

Geissler's friend Julius Plücker connected one of Geissler's evacuated tubes to a battery. He was surprised to find that, at the very low pressure obtained with Geissler's pump, electricity flowed through the tube. Plücker used apparatus similar to sketch (a) on page 591. He sealed a wire into each end of a strong glass tube. Inside the tube, each wire ended in a metal plate, called an electrode. Outside the tube, each wire ran to a source of high voltage. The negative plate inside the tube is called the *cathode*, and the positive plate is called the *anode*. A meter indicated the current going through the tube.

Plücker and his student Johann Hittorf noticed that when an electric current passed through the low-pressure gas in a tube, the tube itself glowed with a pale green color. Several other scientists observed this effect, but two decades passed before anyone undertook a thorough study of the glowing tubes. By 1875, William Crookes had designed new tubes for studying the glow. When he used a bent tube the most intense green glow appeared on the part of the tube that was directly opposite the cathode (at g in sketch (b) on page 591). This suggested that the green glow is produced by something that comes out of the cathode and travels down the tube until it hits the glass. Another physicist, Eugen

Substances that glow when exposed to light, particularly ultraviolet, are called fluorescent. "Fluorescent lights" are essentially Geissler tubes with an inner coating of fluorescent powder.

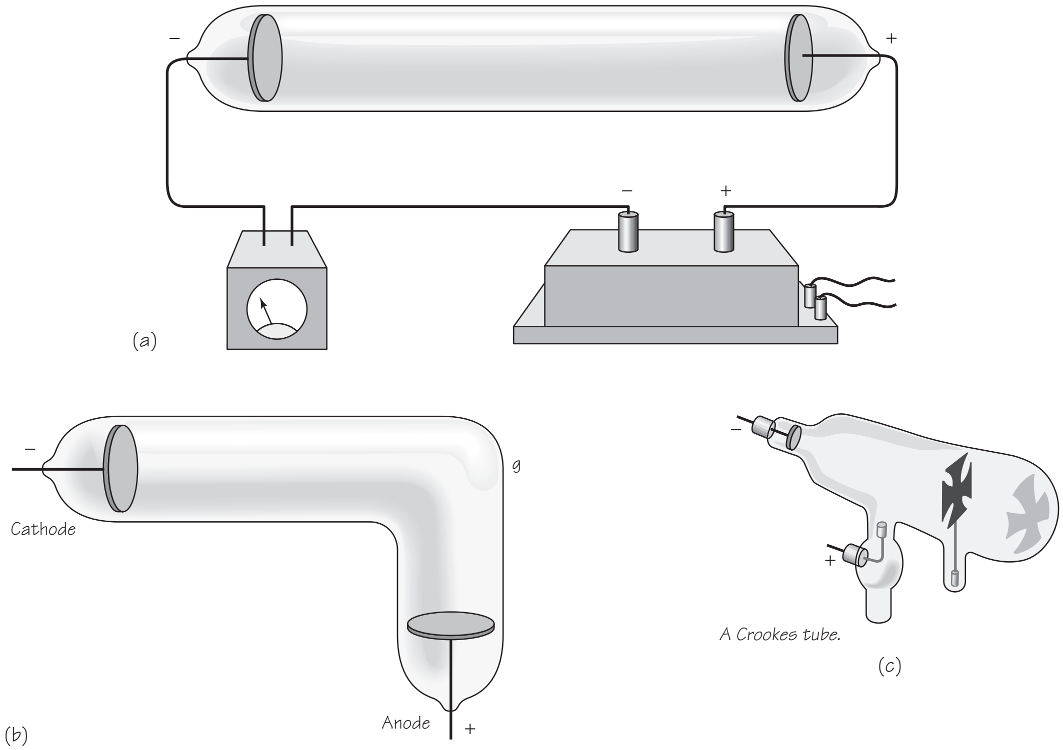

FIGURE 13.4 (a) Geissler-Plücker tube; (b) bent Geissler-Plücker tube; (c) Crookes tube with Maltese cross as barrier.

Goldstein, was also studying the effects of passing an electric current through a gas at low pressure. Goldstein coined a term for whatever it was that appeared to be coming from the cathode, hence the name *cathode rays*. But what could they be?

To study the nature of the rays, Crookes did some clever experiments. He reasoned that if cathode rays could be stopped before they reached the end of the tube, the intense green glow would disappear. He therefore introduced barriers like the Maltese cross (made of metal). A shadow of the barrier appeared in the midst of the green glow at the end of the tube. The cathode seemed to act like a source that radiates a kind of light; the cross acted like a barrier blocking the light. The shadow, cross, and cathode appeared along one straight line. Therefore, Crookes concluded, cathode rays, like light rays, travel in straight lines. Next, Crookes moved a magnet near the tube, and the shadow moved. Thus, he found that magnetic fields deflect cathode rays (which does not happen with light).

In the course of many experiments, Crookes found the following properties of cathode rays:

- No matter what material the cathode is made of, it produces rays with the same properties.
- In the absence of a magnetic field, the rays travel in straight lines perpendicular to the surface that emits them.
- A magnetic field deflects the path of the cathode rays.
- The rays can produce some chemical reactions similar to the reactions produced by light. For example, certain silver salts change color when hit by the rays.
- In addition, Crookes suspected (but did not succeed in showing) that electrically charged objects deflect the path of cathode rays.

The cathode rays fascinated physicists at the time. Some thought that the rays must be a form of light. After all, they have many of the properties of light: they travel in straight lines and produce chemical changes and fluorescent glows just as light does. According to Maxwell's theory, light consists of electromagnetic waves. So the cathode rays might, for example, be electromagnetic waves of frequency much higher than that of visible light.

However, while magnetic fields do not bend light, they do bend the path of cathode rays. Chapter 10 described how magnetic fields exert forces on currents, that is, on moving electric charges. A magnetic field deflects cathode rays in the same way that it deflects negative charges. Therefore, some physicists believed that cathode rays consisted of negatively charged particles.

The debate over whether cathode rays are a form of electromagnetic waves or a stream of charged particles continued for 25 years. Finally, in

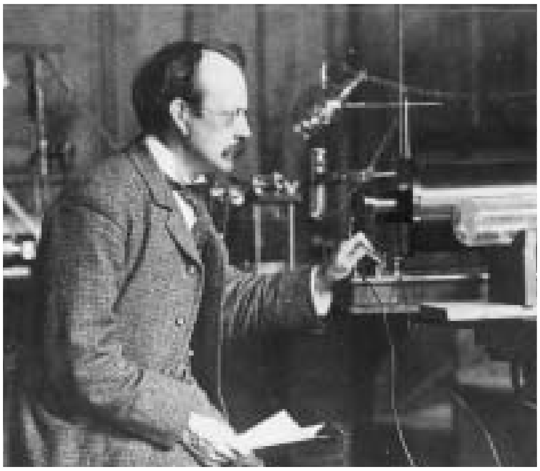

FIGURE 13.5 Joseph John Thomson (1856–1940), one of the greatest British physicists, attended Owens College in Manchester, England, and then Cambridge University. He worked on the conduction of electricity through gases, on the relation between electricity and matter, and on atomic models. His greatest single contribution was the discovery of the electron. Thomson was the head of the famous Cavendish Laboratory at Cambridge University, where one of his students was Ernest Rutherford.

1897, J.J. Thomson, head of the famous Cavendish Laboratory at Cambridge University, made a series of experiments that convinced physicists that cathode rays are negatively charged particles.

By then, it was well known that the paths of charged particles are affected by both magnetic and electric fields. By assuming that cathode rays were negatively charged particles, Thomson could predict what should happen when they passed through such fields. For example, an electric field of just the right magnitude and direction should exactly balance the deflection of a beam of cathode rays by a magnetic field. As Thomson discovered, the predictions were verified. *Thomson could therefore conclude that cathode rays were indeed made up of negatively charged particles.* He was then able to calculate, from the experimental data, the ratio of the charge of a particle to its mass. This ratio is represented by q/m, where q is the charge and m is the mass of the particle. Thomson found that the rays coming from cathodes made of different materials all had the same value of q/m: 1.76×10^{11} C/kg.

Thomson found that

$$\frac{q}{m} = 1.76 \times 10^{11}\ \text{C/kg}.$$

According to Millikan's experiment, the magnitude of q is 1.6×10^{-19} C. Therefore, the mass of an electron is

$$m = \frac{1.6 \times 10^{-19}\ \text{C}}{1.76 \times 10^{11}\ \text{C/kg}} = 0.91 \times 10^{-30}\ \text{kg}.$$

(The mass of a hydrogen ion is 1.66×10^{-27} kg. This is approximately the value of one "atomic mass unit".)

Thus, it was clear that cathode rays must be made of something all materials have in common. The value of q/m for the cathode-ray particles was about 1800 times larger than the value of q/m for charged hydrogen atoms (ions), which had been known to be 9.6×10^{7} C/kg, as measured in chemical electrolysis experiments. It was therefore obvious that either the charge of the cathode-ray particles is much greater than that of the hydrogen ion, or the mass of the cathode-ray particles is much less than the mass of the hydrogen ion.

To decide between the two possibilities, Thomson also measured the charge q on the negatively charged particles in cathode rays with methods other than deflection by electric and magnetic fields. His experiments were not very accurate, but they were good enough to indicate that the magnitude of the negative charge of a cathode-ray particle was the same or not much different from that of the magnitude of the charge of the positive hydrogen ion in electrolysis. In view of the large value of q/m for the cathode-ray particles, Thomson concluded that the mass m of cathode-ray particles is much less than the mass of hydrogen ions.

In sum, the cathode-ray particles were found to have three important properties:

1. The same types of particles were emitted by a wide variety of different cathode materials.

2. They were much smaller in mass than the hydrogen atom, which has the smallest known mass.
3. They had the same magnitude of charge as the charge on the hydrogen ion.

Thomson therefore concluded that the cathode-ray particles must be a hitherto unobserved constituent of all matter. Since it carries electric charge, the cathode-ray particle was later named the *electron*. It has a negative charge of -1.6×10^{-19} C. The *magnitude* of the electron charge is given the symbol e, where $e = 1.6 \times 10^{-19}$ C.

Thomson's finding meant that the atom was not the ultimate limit to the subdivision of matter. Rather, the electron is part of an atom, and could perhaps even be a basic building block of all atoms. The electron, whose existence Thomson had first proved by quantitative experiment, is one of the fundamental or "elementary" particles of which all matter is made. How do we know this?

13.4 THE SMALLEST CHARGE

In Section 10.2, you read that an electrified comb can pick up a small piece of paper. Obviously, the electric force on the paper must exceed the gravitational force exerted on the paper by the Earth. This observation indicates that electric forces generally are stronger than gravitational forces. Using the same principle, the gravitational force on a microscopically small object (which still contains several billion atoms) can be balanced against the electrical force on the same object when the object has a net electric charge of only a single excess charge. This single charge is the electron. It is one of the basic constituents of all atoms.

The fact that the gravitational force on a small object can be balanced by the electric force is the basis of a method for actually measuring the electron's charge. The method was first employed by the American physicist Robert A. Millikan in 1909.

Suppose a small body of mass m, for example, a tiny drop of oil, has a net negative electric charge of magnitude q. (Millikan used fine droplets of oil from an atomizer. The droplets became charged as they formed a spray. The oil was convenient because of the low rate of evaporation of the droplet.) The negatively charged oil drop is placed in an electric field $\mathbf{E}$ directed downward. A force $\mathbf{F}_{\text{el}}$ of magnitude $q\mathbf{E}$ is now exerted on the drop in the *upward* direction. Of course, there is also a downward gravitational force $\mathbf{F}_{\text{grav}}$ of magnitude mg on the drop. The oil drop will accelerate

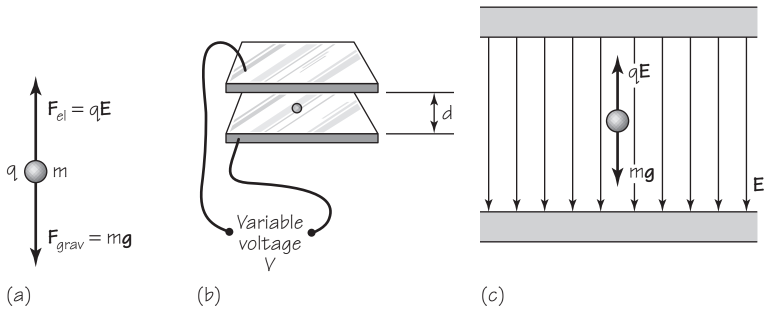

FIGURE 13.6 Schematic diagram of Millikan's oil-drop experiment.

upward or downward, depending on whether the electric force or the gravitational force is greater. By adjusting the magnitude of the electric field strength **E** (i.e., by changing the source that sets up **E**), the two forces, one electric and one gravitational, can be balanced.

What happens when the two forces acting on the drop are balanced? Remember that if a zero *net* force acts on a body, the body can have no acceleration; that is, it would be at rest or continue to move at some constant velocity. In this case, air resistance is also acting as long as the drop moves at all and will soon bring the drop or sphere to rest. The drop will then be in equilibrium. In fact, it will be suspended in mid-air. When this happens, the magnitude of the electric field strength **E** which was applied to produce this condition can be recorded.

Since now the electric force balances the gravitational force, the following must hold:

$$qE = mg.$$

You can calculate the charge q from this equation if you know the quantities E, m, and g, since

$$q = \frac{mg}{E}.$$

Thus, you can find, in the laboratory, what different values of charge q a very small test object can carry.

When you do this, you will discover a remarkable fact: *All possible measurable charges in nature are made up of whole multiples of a smallest charge*. This smallest possible charge is called the *magnitude of the charge on one electron*. By repeating the experiment many times with a variety of small charges,

FIGURE 13.7 The original apparatus used by Robert Millikan to conduct his oil-drop experiment.

you can find the value of the charge on one electron. By convention, the charge on an electron is represented by the symbol e.

In effect, this is what Millikan did. He obtained the value of $e = 1.6024 \times 10^{-19}$ C for the electron charge, and he found that the sign of the charge on the electron is negative. Any charge q is therefore given by $q = ne$ where n is the whole number of individual charges, each of magnitude e. Therefore, for example, 1 C is the magnitude of the charge on $1/(1.6 \times 10^{-19})$ electrons. For most purposes you can use the value $e = 1.6 \times 10^{-19}$ C. This value agrees with the results of many other experiments done since then.

No experiment has yet revealed the existence of a smaller unit of charge on ordinary matter, such as an oil drop. However, scientists have found that most elementary particles are in fact composed of smaller entities, known as *quarks*, each of which can have a charge of either $+2/3e$ or $-1/3e$. But quarks cannot exist in isolation, so the smallest charge that can be observed on ordinary mater is still the charge of the electron, e.

13.5 THOMSON'S MODEL OF THE ATOM

What is the atom like? By the beginning of the twentieth century, enough chemical and physical information was available to allow many physicists to devise models of atoms. It was known that negative particles with identical properties (electrons) could be obtained from many different substances and in different ways. This suggested that electrons are parts of all atoms. Electrons are negatively charged. But samples of an element are ordinarily electrically neutral. Therefore, the atoms making up such samples are also presumably neutral. If so, the presence of negative electrons in an atom would seem to require the presence of an equal amount of positive charge.

As mentioned earlier, hydrogen atoms are about 1800 times more massive than electrons, which have a mass of 9.1×10^{-28} g. Thus electrons make up only a very small part of the atomic mass in any atom. Consequently, any model of an atom must take into account the following information:

- an electrically neutral atom contains equal amounts of positive and negative charge;
- the negative charge is associated with only a small part of the mass of the atom, that carried by the electrons.

In addition, any atomic model should answer at least two questions:

1. How many electrons are there in an atom?
2. How are the electrons and the positive charge arranged in an atom?

During the first 10 years of the twentieth century, several atomic models were proposed, but none was satisfactory. Until 1911, the most popular model for the atom was one proposed by J.J. Thomson in 1904. Thomson suggested that an atom consisted of a sphere of positive electricity in which an equal amount of negative charge was distributed in the form of small electrons. From studies of the ionization of atoms in chemistry and through X radiation (discussed in Section 13.8). The amount of positive charge in each atom was surmised to be equal to the atomic number of the atom Z (an integer) times the magnitude of the electron charge e, that is, an atom's positive charge is equal to Ze. In order to render the atom electrically neutral, there were an equal number of negatively charged electrons.

Under these assumptions, Thomson's atom was like a "plum pudding" of positive electricity, with the negative electrons scattered in it like plums or raisins. Thus, hydrogen ($Z = 1$) consisted of one electron, charge $-1e$, embedded in a sphere of positive charge, $1e$. Helium ($Z = 2$) consisted of two electrons embedded in a sphere of positive charge of $2e$, and so on.

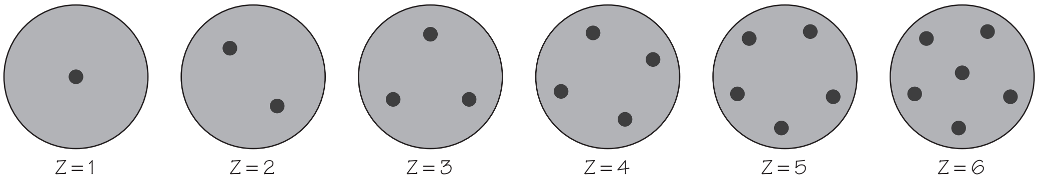

FIGURE 13.8 Thomson's "plum pudding" model of the atom for the first six elements.

The positive "pudding" was assumed to act on the negative electrons, holding them in the atom by electric forces only. Thomson did not know how the positive "pudding" itself was held together, though it should fall apart by electrical repulsion. He took the radius of the atom to be of the order of 10^{-8} cm, based on information from the kinetic theory of gases and other considerations. But his incomplete model was unable to account for either the stability of the atom or for its chemical properties. Much more experimental information and a radical new concept—the concept of the energy quantum—was required to construct a much more satisfactory model of the atom. This concept entered through the use of results from work on an entirely different set of problems—thermal radiation and the puzzle of the photoelectric effect—pursued at that time by Max Planck and Albert Einstein.

13.6 THE PHOTOELECTRIC EFFECT

In 1887, the German physicist Heinrich Hertz was testing Maxwell's theory of electromagnetic waves (see Section 12.4). He noticed that a metallic surface can emit electric charges when light of very short wavelength falls on it. Because light and electricity are both involved, this phenomenon is called the *photoelectric effect*.

When the electric charges so produced passed through electric and magnetic fields, their paths changed in the same ways as the paths of cathode rays. It was therefore deduced that the electric charges emitted by the photoelectric effect consist of negatively charged particles. In 1898, J.J. Thomson measured the value of the ratio q/m for these particles. Using the same method that he had used for cathode-ray particles, Thomson got the same value for the photoelectric particles as he had for the cathode-ray particles. These experiments (and others) demonstrated that photoelectric particles had the same properties as electrons. In fact, physicists now consider that these particles *are* electrons, although they are often referred to as *photoelectrons*, to indicate their origin. Later work showed that all substances (solids, liquids, and gases) exhibit the photoelectric effect under appropriate conditions. However, it is more convenient to study the effect with metallic surfaces.

The photoelectric effect has had an important place in the development of atomic physics because the effect could not be explained in terms of the ideas of classical physics. New ideas had to be introduced to account for the experimental results. In particular, the introduction of the revolutionary concept of *quanta* initiated a new branch of physics called *quantum* *theory*. This theory and the notion of quanta arose, at least in part, because of the explanation provided for the photoelectric effect.

The basic information for studying the photoelectric effect comes from two kinds of measurements:

1. measurements of the photoelectric current (the number of photoelectrons emitted per unit time); and
2. measurements of the kinetic energies of the photoelectrons after they are emitted.

The photoelectric current can be studied with an apparatus like that sketched in Figure 13.9(a). Two metal plates, C and A, are sealed inside a well-evacuated quartz tube. (Quartz glass is transparent to ultraviolet light as well as to visible light.) The two plates are connected to a source of potential difference (e.g., a high-voltage battery). In the circuit is also an ammeter, which measures the electric current. As long as light strikes plate C, as in sketch (b), electrons with a range of kinetic energies are emitted from

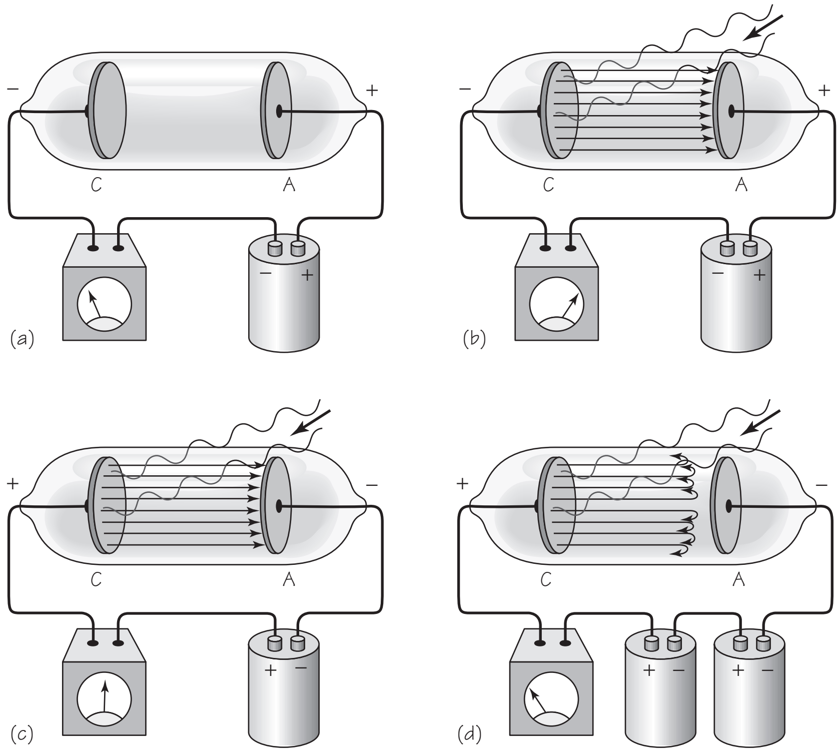

FIGURE 13.9 Schematic diagram of apparatus for photoelectric experiments.

it. If the potential of plate A is positive relative to plate C, these emitted photoelectrons will accelerate to plate A. (Some emitted electrons will reach plate A even if it is not positive relative to C.) The resulting electric current is indicated by the ammeter.

The results of the experiment include the following: the stronger (or more intense) the beam of light of a given color (frequency), the greater the photoelectric current. Any metal used as plate C shows a photoelectric effect, but only if the light has a frequency greater than a certain value. This value of the frequency is called the threshold frequency for that metal. Different metals have different threshold frequencies. But if the incident light has a frequency lower than the threshold frequency, *no* photoelectrons are emitted, no matter how great the intensity of the light or how long the light is left on! This is the first set of surprising discoveries.

The kinetic energies of the electrons can be measured in a slightly modified version of the apparatus, sketched in (c). The battery is reversed so that plate A now tends to repel the photoelectrons. The voltage can be changed from zero to a value just large enough to keep any electrons from reaching plate A, as indicated in (d).

When the voltage across the plates is zero, the meter will indicate a current. This reading shows that the photoelectrons, emerging with kinetic energy from the metallic surface, can reach plate A. As the repelling voltage is increased, the photoelectric current decreases. Eventually a certain voltage is reached at which even the swiftest electrons are repelled and thus the current becomes zero, as indicated in (d). This voltage, which is called the *stopping voltage*, is a measure of the maximum kinetic energy of the emitted photoelectrons (KE_{max}). Using V_{stop} to indicate the stopping voltage, then maximum kinetic energy is given by the relation

$$KE_{\text{max}} = eV_{\text{stop}},$$

where e is the magnitude of the electron's charge. The measured maximum kinetic energy of the emitted electrons was found to be proportional to the frequency of the incident light.

The above experimental results can be summarized in the following statements. Only the important experimental results are listed here. Their theoretical interpretation will be discussed later.

1. A metal shows a photoelectric effect only if the incident light has a frequency above a certain threshold frequency (symbol f_0).
2. If light of a given frequency does produce a photoelectric effect, the photoelectric current from the surface is proportional to the intensity of the light falling on it.

3. If light of a given frequency releases photoelectrons, the emission of these electrons is *immediate*.
4. The kinetic energies of the emitted electrons display a maximum value, which is proportional to the frequency of the incident light (above the threshold frequency).

The measured time interval between the instant the light strikes the metallic surface and the appearance of electrons is at most 3×10^{-9} s and probably much less. In some experiments, the light intensity used was extremely low. According to the classical wave theory of light, it should take several hundred seconds for an electron to accumulate enough energy from such light to be emitted. But even in these cases, electrons are emitted practically right after the light strikes the surface.

5. The maximum kinetic energy of the photoelectrons increases in direct proportion to the frequency of the light that causes their emission. The maximum *KE* is *not* dependent on the *intensity* of the incident light, as the classical wave theory of light would require.

The way in which the maximum kinetic energy of the electrons varies with the frequency of the incident light is shown in Figure 13.10. The symbols $(f_0)_1$, $(f_0)_2$, $(f_0)_3$ stand for the different threshold frequencies of three different substances. For each substance, the experimental data points fall on a straight line. All the lines have the same slope.

What is most astonishing about the experimental results is that there is a threshold frequency, and that the photoelectrons are emitted if the light frequencies are just above the threshold frequency, no matter how weak the beam of light. But if the light frequencies are just below the threshold

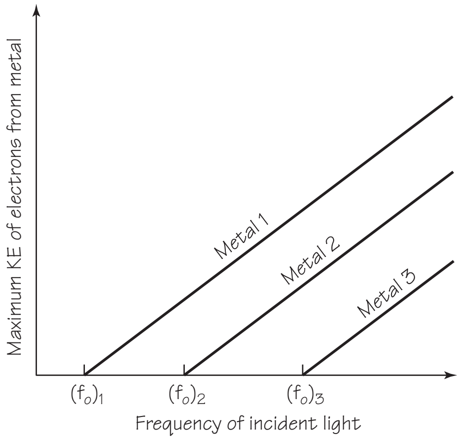

FIGURE 13.10 Photoelectric effect: Maximum kinetic energy of the electrons as a function of the frequency of the incident light. Different metals yield lines that are parallel but have different threshold frequencies.

frequency, no electrons are emitted no matter how great the intensity of the light beam is.

Statements 1, 3, and 4 above could not be explained on the basis of the classical electromagnetic theory of light. How could a low-intensity train of light waves, spread out over a large number of atoms, concentrate, in a very short time interval, enough energy on one electron to knock the electron out of the metal?

Furthermore, the classical wave theory could not account for the existence of a threshold frequency. There seemed to be no reason why a high-intensity beam of low-frequency radiation should not produce photoelectricity if low-intensity radiation of higher frequency could produce it. Neither could classical theory explain why the maximum kinetic energy of the photoelectrons increases directly with the frequency of the light, but is independent of the intensity. Thus, the photoelectric effect posed an important challenge that the classical wave theory of light was not able to meet.

13.7 EINSTEIN'S THEORY OF THE PHOTOELECTRIC EFFECT

The explanation of the photoelectric effect was the major work cited in the award to Albert Einstein of the Nobel Prize in Physics in 1921. Einstein's explanation, proposed in 1905, played a major role in the development of atomic physics. He based his theory on a daring hypothesis, for few of the experimental details were known in 1905. Moreover, the key point of Einstein's explanation contradicted the classical ideas of the time.

Einstein assumed that the energy of light is not distributed evenly over the whole expanding wave front (as the classical theory assumed). Instead, the light energy is concentrated in separate "lumps." In addition, the amount of energy in each of these regions is not just any amount, but a definite amount of energy that is proportional to the frequency f of the light wave. The proportionality factor is a constant (symbol h); it is called *Planck's constant* for reasons which will be discussed later. Thus, in this model, the light energy in a beam of frequency comes in pieces, each of amount

$$E = hf$$

where

$$h = 6.6 \times 10^{-34}\ \text{J/s}.$$

The amount of radiant energy in each piece is called a *light quantum* (*quantum* is Latin for *quantity*), or a quantum of light energy. The quantum of light energy was later called a *photon*.

There is no explanation clearer or more direct than Einstein's. A quote from his first paper (1905) on this subject is given here. Only the notation is changed, in order to agree with modern practice (including the notation used in this text):

> . . . According to the idea that the incident light consists of quanta with energy hf, the ejection of cathode rays [photoelectrons] by light can be understood in the following way. Energy quanta penetrate the surface layer of the body, and their energy is converted, at least in part, into kinetic energy of electrons. The simplest picture is that a light quantum gives up all its energy to a single electron; we shall assume that this happens. . . . An electron provided with kinetic energy inside the body may have lost part of its kinetic energy by the time it reaches the surface. In addition, it is to be assumed that each electron, in leaving the body, has to do an amount of work W (which is characteristic of the body). The electrons ejected directly from the surface and at right angles to it will have the greatest velocities perpendicular to the surface. The maximum kinetic energy of such an electron is
>
> $$KE_{\text{max}} = hf - W.$$
>
> If the plate C is charged to a positive potential, V_{stop}, just large enough to keep the body from losing electric charge, we must have
>
> $$KE_{\text{max}} = hf - W = eV_{\text{stop}},$$
>
> where e is the magnitude of the electronic charge. . . .
>
> If the derived formula is correct, then V_{stop}, when plotted as a function of the frequency of the incident light, should yield a straight line whose slope should be independent of the nature of the substance illuminated.

The first equation in the above quotation is usually called *Einstein's photoelectric equation*. Let us compare Einstein's photoelectric equation with the experimental results, in the order given in the previous section, to test whether or not his theory accounts for those results:

1. According to the photoelectric equation, the kinetic energy of the photoelectrons is greater than zero only when the photon energy hf is greater than the work W, which is the work the electron must do against the forces of attraction of the material in the cathode C as it leaves the metal. The energy required to escape the metal is known as the *work*

function. Therefore, an electron can be emitted only when the frequency of the incident light is greater than a certain lowest value corresponding to the work required to exit from the metal. In symbols, the minimum frequency f_0 is defined by the relation $hf_0 = W$.

2. Next, according to Einstein's photon model of light, it is an individual photon that ejects an electron (if above the threshold frequency of light). Now, the intensity of the light is proportional to the number of the photons in the light beam. In addition, the number of photoelectrons ejected is proportional to the number of photons incident on the surface. Therefore, the number of electrons ejected (and with it the photoelectric current) is proportional to the intensity of the incident light. (However, not every photon in a light beam strikes an electron, emitting it from the metal; only about 1 in 50 photons does so.)
3. In Einstein's model, the light energy is concentrated in a stream of light quanta (photons). So no time is needed for the electron to collect light energy. Instead, the quanta transfer their energy immediately to the photoelectrons, which emerge after the very short time required for them to escape from the surface.
4. Finally, the photoelectric equation predicts that the greater the frequency of the incident light, the greater the maximum kinetic energy of the ejected electrons. According to the photon model, the photon's energy is directly proportional to the light frequency. The minimum energy needed to eject an electron is the energy required for the electron to escape from the metal surface. This explains why light of frequency less than some frequency f_0 cannot eject any electrons. The kinetic energy of the escaping electron is the difference between the energy of the absorbed photon and the energy lost by the electron in escaping the surface.

Testing Einstein's Theory by Experiment

As you can see, Einstein's photoelectric equation agreed qualitatively with the experimental results. But could it hold up under quantitative experimental testing? In particular:

(1) Does the maximum kinetic energy of the electrons vary in direct proportion to the light frequency?
(2) Is the proportionality factor h really the same for all substances?

For 10 years, experimental physicists attempted these quantitative tests. One experimental difficulty was that the value of W for a metal changes greatly if there are impurities (e.g., a layer of oxidized metal) on the sur-

ALBERT EINSTEIN

Albert Einstein (1879–1955) was born in the city of Ulm, in Germany. Like Newton, he showed no particular intellectual promise as a youngster. He received his early education in Germany, but at the age of 15, dissatisfied with the discipline in school and militarism in the nation, he left and went to Switzerland. After graduation from the Polytechnic Institute in Zurich, Einstein (in 1902) found work in the Swiss Patent Office in Berne. This job gave Einstein a salary to live on, to get married, and an opportunity to use his spare time for working in physics on his own. In 1905, he published three papers of immense importance. One dealt with quantum theory and included his theory of the photoelectric effect. Another treated the problem of molecular motions and sizes, and worked out a mathematical analysis of the phenomenon of "Brownian motion." Einstein's analysis and the subsequent experimental work by Jean Perrin, a French physicist, provided a strong argument for the molecular motions assumed in the kinetic theory. Einstein's third 1905 paper provided the theory of special relativity, which revolutionized modern thought about the nature of space and time, and of physical theory itself.

In 1915, Einstein published a paper on the theory of general relativity. In it he provided a new theory of gravitation that included Newton's theory as a special case.

When Hitler and the Nazis came to power in Germany in 1933, Einstein went to the United States and became a member of the Institute for Advanced Study at Princeton. He spent the rest of his working life seeking a unified theory which would include gravitation and electromagnetism. Shortly before the Germans launched World War II, Einstein signed a letter to President Roosevelt, warning of the war potential of an "atomic bomb," for which the Germans had all necessary knowledge and motivation to build. After World War II, Einstein devoted much of his time to organizations advocating world agreements to end the threat of atomic warfare.

face. Finally, in 1916, Robert A. Millikan established that there is indeed a straight-line relationship between the frequency of the absorbed light and the maximum kinetic energy of the photoelectrons, as required by Einstein's equation. If KE_{max} is plotted along the y-axis and f along the x-axis, then Einstein's equation exhibits the familiar form of the equation for a straight line

$$y = mx + b.$$

In a graph of Einstein's equation, the slope should be equal to h, and the KE-intercept should be equal to $-W$. This is exactly what Millikan found.

In order to obtain his data, Millikan designed an apparatus in which the

FIGURE 13.11 Robert Andrews Millikan (1868–1953), an American physicist, attended Oberlin College, where his interest in physics was only mild. After his graduation, he became more interested in physics, taught at Oberlin while taking his master's degree, and then obtained his doctorate degree from Columbia University in 1895. After more study in Germany, Millikan went to the University of Chicago, where he became a professor of physics in 1910. His work on the determination of the electron's charge took place from 1906 to 1913. He was awarded the Nobel Prize in physics in 1923 for this research and for the very careful experiments which resulted in the verification of the Einstein photoelectric equation. In 1921, Millikan moved to what became known as the California Institute of Technology, eventually to become its president.

metal photoelectric surface was cut clean while in a vacuum. A knife inside the evacuated volume was manipulated by an electromagnet outside the vacuum to make the cuts. This rather intricate arrangement was required to achieve a pure metal surface.

The straight-line graphs Millikan obtained for different metals all had the same slope, h, even though the threshold frequencies (related to W) were different. The value of h could be obtained from Millikan's measurements, and it was the same for each metal surface. Also, it agreed very well with a value obtained by means of other, independent methods. Much to Millikan's own surprise (who started out believing in the classical theory of light), his experiment provided a quantitative verification of Einstein's theory of the photoelectric effect. Thus, it can be said that Einstein's equation led to two Nobel Prizes: one to Einstein, who derived it theoretically, and one to Millikan, who verified it experimentally.

The Idea of Energy Quanta

Historically, the first suggestion that the energy in electromagnetic radiation is "quantized" (comes in definite quanta) did not come from the photoelectric effect. Rather, it came from studies of the heat and light radiated by hot solids. Max Planck, a German physicist, introduced the concept of quanta of energy (though in a different context) in 1900, 5 years before Einstein's theory. Thus, the constant h is known as Planck's constant.

Planck was trying to explain how the heat (and light) energy radiated by a hot body is related to the frequency of the radiation. Classical physics (nineteenth-century thermodynamics and electromagnetism) could not account for the experimental facts. Planck found that the facts could be interpreted only by assuming that atoms, on radiating, change their energy not in varying, but in separate, quantized amounts, E. Einstein's theory of the photoelectric effect was actually an extension and application of Planck's quantum theory of thermal radiation. The essential point was that Einstein postulated the change in the atom's energy E being carried off was located in a distinct photon of energy $E = hf$, where f is the frequency of the light emitted by the atom, rather than being spread continuously over the light wave.

The experiments and the theory of radiation are much more difficult to describe than the experiments and theory of the photoelectric effect. By now, many tests have been made of both Planck's and Einstein's conceptions. In all cases, Planck's constant h is regarded as one of the universal constants of nature. It is found to have the same basic position in quantum physics that Newton's universal constant G has in the physics of gravitation. As Planck said, it is an absolute constant in the sense that even if extraterrestrials

FIGURE 13.12 Max Planck (1858–1947), a German physicist, was the originator of the quantum theory, one of the two great revolutionary physical theories of the twentieth century (along with Einstein's relativity theory). Planck won the Nobel Prize in 1918 for his work on quantum theory. He tried for many years to show that this theory could be understood in terms of the classical physics of Newton and Maxwell, but this attempt did not succeed. Quantum physics is fundamentally different because of its postulate that energy in light and matter exists in finite quanta and is not continuously divisible.

exist on other planets, they would find the same value for h (as indeed they would find for G, e, c, and all other physical constants of nature).

The Physicists' Dilemma

The photoelectric effect presented physicists with a real dilemma. According to the classical wave theory, light consists of electromagnetic waves extending continuously throughout space. This theory was highly successful in explaining optical phenomena (reflection, refraction, polarization, interference). Light behaves like a wave experimentally, and Maxwell's theory gives a good account of this wave behavior. But Maxwell's theory could not account for the photoelectric effect. Einstein's theory, which postulated the existence of separate quanta of light energy, accounted for the photoelectric effect. But it could not account for other properties of light, such as interference. The result was that there were two models for light whose basic concepts seemed to contradict each other. According to one, light is a wave phenomenon; according to the other, light has particle-like properties. Each model had its successes and limits. What, if anything, could be done about the contradictions between the two models? You will see later that this problem and its treatment occupy a central position in modern physics.

13.8 X RAYS

In 1895, the German physicist Wilhelm Konrad Röntgen made a discovery that surprised him and all the world. Like the photoelectric effect, it did not fit with accepted ideas about electromagnetic waves, and eventually it too required the introduction of quanta for a complete explanation. The discovery was that of X rays, which were often called "Röntgen" rays, after their discoverer. The consequences for atomic physics, medicine, and technology were dramatic and important.

On November 8, 1895, Röntgen was experimenting with the newly found cathode rays, as were many physicists all over the world. According to a biographer:

> . . . he had covered the all-glass pear-shaped tube [Crookes tube] with pieces of black cardboard, and had darkened the room in order to test the opacity of the black paper cover. Suddenly, about a yard from the tube, he saw a weak light that shimmered on a little bench he knew was nearby. Highly excited, Röntgen lit a match and, to his great surprise, discovered that the source of the mysterious light was a little barium platinocyanide screen lying on the bench.

FIGURE 13.13 Wilhelm Röntgen (1845–1923).

Barium platinocyanide, a mineral, is one of the many chemicals known to *fluoresce* (emit visible light when illuminated with ultraviolet light). But no source of ultraviolet light was present in Röntgen's experiment. Cathode rays had not been observed to travel more than a few centimeters in air. So, neither ultraviolet light nor the cathode rays themselves could have caused the fluorescence. Röntgen therefore deduced that the fluorescence involved rays of a new kind. He named them *X rays*, since the rays were an unknown nature.

In an intensive and thorough series of experiments over the next 7 weeks, he determined the properties of this new radiation. Röntgen reported his results on December 28, 1895, in a paper whose title (translated) is "On a New Kind of Rays."

Röntgen's paper described nearly all of the properties of X rays that are known even now. It described the method of producing the rays and proved that they originated in the glass wall of the tube, where the cathode rays struck it. Röntgen showed that X rays travel in straight lines from their place of origin, and that they darken a photographic plate. He reported in detail the ability of X rays to penetrate various substances, such as paper, wood, aluminum, platinum, and lead. Their penetrating power was greater through light materials (paper, wood, flesh) than through dense materials (platinum, lead, bone). He described and exhib-

FIGURE 13.14 X-ray image of Frau Röntgen's hand with rings. © Bettmann/CORBIS.

ited photographs showing "the shadows of bones of the hand, of a set of weights inside a small box, and of a piece of metal whose inhomogeneity becomes apparent with X rays." He gave a clear description of the shadows cast by the bones of the hand on the fluorescent screen. Röntgen also reported that the X rays were not deflected by a magnetic field. He also could not show their reflection, refraction, or interference effects in ordinary optical apparatus.

J.J. Thomson discovered one of the most important properties of X rays a month or two after the rays themselves had become known. He found that when the rays pass through a gas, they make it a conductor of electricity. Thomson attributed this effect to "a kind of electrolysis, the molecule being split up, or nearly split up by the Röntgen rays." The X rays, in passing through the gas, knock electrons loose from some of the atoms or molecules of the gas. The atoms or molecules that lose these electrons become positively charged. They are called *ions* because they resemble the positive ions in electrolysis, and the gas is said to be *ionized*. Also, the freed electrons may attach themselves to previously neutral atoms or molecules, giving them negative charges.

Röntgen and Thomson found, independently, that electrified bodies lose their charges when the air around them is ionized by X rays. (It is now easy to see why: the electrified body attracts ions of the opposite charge from the air.) The rate of discharge depends on the intensity of the rays (hence the amount of ionization). This property was therefore used—and still is—as a convenient quantitative means of measuring the intensity of an X ray beam. As a result, careful quantitative measurements of the properties and effects of X rays could be made.

Are X Rays Waves or Particles?

One problem that aroused keen interest following the discovery of X rays concerned the nature of the mysterious rays. Unlike charged particles (electrons, for example) they were not deflected by magnetic or electric fields. Therefore, it seemed that they had to be either neutral particles or electromagnetic waves. It was difficult to choose between these two possibilities. On the other hand, no neutral particles of atomic size (or smaller) that had the penetrating power of X rays were then known. The existence of such particles would be extremely hard to prove, because there was no way of getting at them. On the other hand, if the X rays were electromagnetic waves, they would have to have extremely short wavelengths because only in this case, according to theory, could they have high penetrating power and show no refraction or interference effects with ordinary optical apparatus.

As discussed in Chapter 8, distinctly wave-like properties become apparent only when waves interact with objects, such as slits in a barrier, that are smaller than several wavelengths across. The wavelengths hypothesized for X rays would be on the order of 10^{-10} m (see the diagram of the electromagnetic spectrum in Section 12.5). So a demonstration of their wave behavior would require a diffraction grating with slits spaced about 10^{-10} m apart. Evidence from kinetic theory and from chemistry indicated that atoms were about 10^{-10} m in diameter. It was suggested, therefore, that X rays might be diffracted measurably by crystals in which the atoms form orderly layers about 10^{-10} m apart.

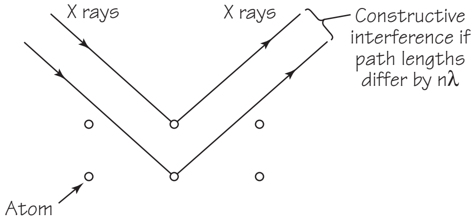

FIGURE 13.15 X-ray diffraction patterns from a metal crystal formed on a photographic film. The black spots are produced by constructive interference of X rays scattered from atoms.

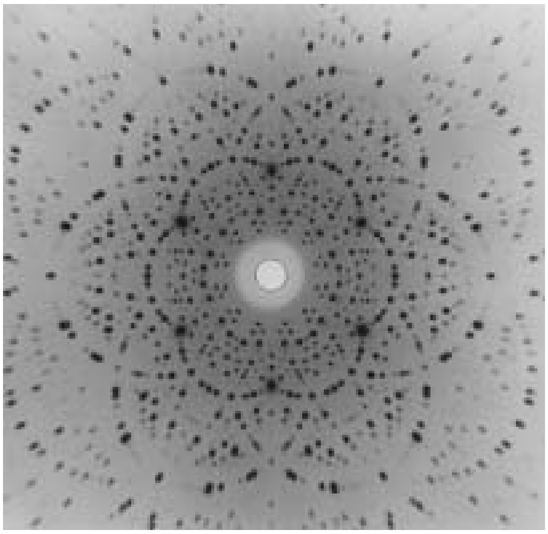

FIGURE 13.16 X-ray diffraction pattern from beryllium crystal.

Such experiments succeeded in 1912. The layers of atoms did act like diffraction gratings, and X rays did, indeed, act like electromagnetic radiations of very short wavelength (like *ultra*-ultraviolet light). These experiments are more complicated to interpret than diffraction of a beam of light by a single, two-dimensional optical grating. The diffraction effect occurs in three dimensions instead of two. Therefore, the diffraction patterns are far more elaborate (see Figure 13.16).

However, in addition to wave properties, X rays were found to have quantum properties, which meant they also exhibited particle-like behavior. For example, they can cause the emission of electrons from metals. These electrons have greater kinetic energies than those produced by ultraviolet light. (The ionization of gases by X rays is also an example of the photoelectric effect. In this case, the electrons are freed from the atoms and molecules of the gas.) Thus, X rays also require quantum theory for the explanation of some of their behavior. So, like light, X rays were shown to have both wave and particle properties.

The Discovery Causes a Sensation

Röntgen's discovery of X rays excited scientists throughout the world. His experiments were immediately repeated and extended in many laboratories in Europe and America. Scientific journals during the year 1896 were filled with letters and articles describing new experiments or confirming the results of earlier experiments. The passage of electricity through gases had been a popular topic for study by physicists; this made widespread experimentation much easier during the years after Röntgen's discovery, because many physics laboratories had cathode-ray tubes and so could produce X rays easily.

The spectacular use of these rays in medicine generated intense interest in X rays. Within 3 months of Röntgen's discovery, X rays were put to practical use in surgical operations in a hospital in Vienna. The use of this new aid to surgery spread rapidly. Since Röntgen's time, X rays have revolutionized some phases of medical practice, especially the diagnosis of some diseases and the treatment of some forms of cancer (because X rays also can destroy malignant tissue). Extremely important uses of X rays occur as well in other fields of applied science, both physical and biological. Among these are the study of the crystal structure of materials; "industrial diagnosis," such as the search for possible defects in materials and engineering structures; the study of what is behind the optically visible surface of old paintings and sculptures; and many others.

The public reaction to the discovery of X rays was sensational also. Many people rushed to have their bodies irradiated by the new rays, thinking they have miraculous properties, while others worried about moral decline if modesty gave way to "X-ray vision." A Berlin newspaper even published a full-body X ray of the Kaiser, giving new meaning in retrospect to the old story of "The Emperor's New Clothes."

From the apparatus used in Röntgen's original discovery, there emerged two paths of development in medical technology. One concentrated on a fluoroscope, the other focused on improving the radiograph; Thomas Edison was crucial in both.

It was Edison's familiarity with the Crookes tube, which was very similar to his 1879 invention of the electric light bulb, that enabled him to make one of the first improvements to X-ray technology. By constructing a tube using thinner glass, Edison found that more X rays could escape. Edison also led the investigation which found that calcium tungstate could produce a clearer image on the fluorescent screen than the previously used barium platinocyanide. Edison took this knowledge and applied it to the manufacture of a "fluoroscope," a device which allowed a person to peer through a box at a screen coated in calcium tungstate, and see a moving image of the inside of their, or another person's, body, as it moved beneath the screen.

A friend of Edison's, Michael Pupin, took this improvement in the fluorescent screen and combined it with a photographic plate, which reduced the patient exposure time for a radiograph from 1 hour to just a few minutes, while also increasing the clarity of the picture. This greatly reduced the danger of damage to tissues.

X-RAY APPLICATIONS

Originally, X rays were produced in Röntgen's laboratory when cathode rays (electrons) struck a target (the glass wall of the tube). Today, X rays usually are produced by directing a beam of high-energy electrons onto a metal target. As the electrons are deflected and stopped, X rays of various energies are produced. The maximum energy a single ray can have is the total kinetic energy the incident electron gives up on being stopped. So the greater the voltage across which the electron beam is accelerated, the more energetic and penetrating are the X rays.

X rays are used in a wide range of fields. Computer-aided tomography (CAT) allows doctors to create three-dimensional images of a patient's body. X rays are also used in art restoration, revealing structural defects in sculptures and enabling the assessment of previous restoration efforts of paintings and frescoes.

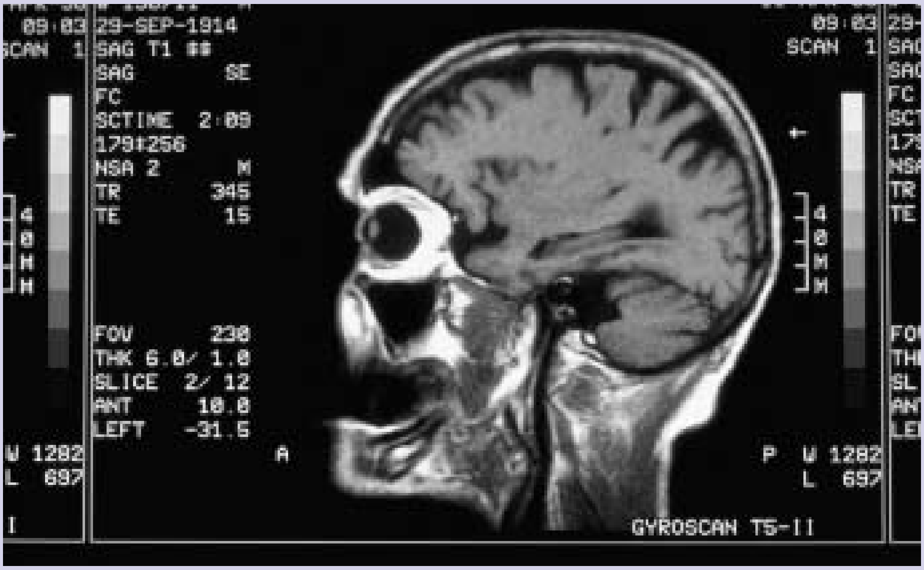

FIGURE 13.17 Computed tomography scan of a person's head in which X rays are used.

FIGURE 13.18 Another application of X rays. Here, a bronze sculpture by Gian Lorenzo Bernini is being examined by comparison with X-radiograph images (displayed on screen). Such non-invasive procedures have become essential tools in the field of art preservation.

One of the most problematic aspects of X-ray technology for the first two decades was the unreliability of the glass tubes, which often cracked when heated. The high incidence of cracking was eliminated in 1913 when William Coolidge, while working for General Electric, invented the high-vacuum, hot-cathode, tungsten-target X-ray tube. As part of his research into electric-light-bulb filaments, Coolidge found that because tungsten vaporized less than any other metals, it could reduce the buildup of gas residue. Applying this knowledge, Coolidge replaced platinum with tungsten in the cathode-ray tube. When these "Coolidge tubes" came on the market in 1913, they had many benefits over the previous design; they could produce clearer duplications of previous images, be adjusted much more accurately, and because of their increased flexibility, they could arrange to go instantly from high to low penetration.

After World War II the monopoly of X rays as a medical imaging tool began to be challenged by new technologies that combined computers and television to help in obtaining an image of those parts of the human body which X rays had failed to reach. They include CT (computed tomography) scanning, MRI (magnetic resonance imaging), PET (positron emission tomography), and ultrasound.

FIGURE 13.19 Edison examining his hand under a fluoroscope, ca. 1896.

Many of the scientific and technological advances, on which these "daughter technologies" depended, were taking place throughout the first half of this century. However, it took the invention and subsequent improvements in computer technology, before the complex algorithms required for the imaging techniques involved in CT scanning and MRI could be calculated. These advanced medical imaging techniques, whose importance is now established in medical diagnostics, are a world away from the simple technology that allowed Röntgen to see through his skin in 1895, still, despite these advances, X rays remain the most commonly used imaging tools, largely because of the simplicity, and the size, and the financial cost of its "daughter technologies."

SOME NEW IDEAS AND CONCEPTS

atomic mass units
atomic number
cathode
cathode rays
Einstein's photoelectric equation
electron
law of fixed proportions
light quantum
periodic table
photoelectric effect
photoelectron
photon
stopping voltage
Thomson's model of the atom
work function

SOME IMPORTANT EQUATIONS

$$E = hf,$$

$$KE_{\max} = hf - W,$$

$$KE_{\max} = eV_{\text{stop}}.$$

FURTHER READING

G. Holton and S.G. Brush, *Physics, The Human Adventure* (Piscataway, NJ: Rutgers University Press, 2001), Chapter 26.

B.H. Kevles, *Naked to the Bone: Medical Imaging in the Twentieth Century*. Sloan Technology Series (New Brunswick, NJ: Rutgers University Press, 1997).

STUDY GUIDE QUESTIONS

13.1 The Periodic Table

1. How did Mendeleev arrange the elements on the periodic table?
2. Describe some of the common features of different groups of elements.
3. Which element is element 56? What is its atomic number? atomic mass?

13.2 The Idea of Atomic Structure

1. How did the periodic table suggest that atoms might have a structure?
2. What changes with respect to the structure of the atoms would one find as one progressed through the periodic table?

13.3 Cathode Rays

1. What was the most convincing evidence to support the fact that cathode rays were not electromagnetic radiation?
2. What was the reason given for the ratio q/m for electrons being about 1800 times larger than q/m for hydrogen ions?
3. What were two main reasons for Thomson's belief that electrons may be "building blocks" from which all atoms are made?

13.4 The Smallest Charge

1. How can the small oil drops or plastic spheres used in the Millikan experiment experience an electric force upward if the electric field is directed downward?
2. What did the results of the Millikan experiment indicate about the nature of electric charge?
3. In Millikan's experiment, what is necessary to keep a charged droplet suspended in place?

13.5 Thomson's Model of the Atom

1. Describe Thomson's model of the atom. Give the reasoning behind the inclusion of each component and their arrangement.
2. What was one problem with this model?
3. Draw a picture of an atom of oxygen, according to Thomson, and label everything in it.

13.6 The Photoelectric Effect

1. Light falling on a certain metal surface causes electrons to be emitted. What happens to the photoelectric current as the intensity of the light is decreased?
2. What happens as the frequency of the light is decreased?
3. Which observations could not be explained by the classical wave theory of light?
4. Sketch a rough diagram of the equipment and circuit used to demonstrate the main facts of photoelectricity.

13.7 Einstein's Theory of the Photoelectric Effect

1. High-frequency light falls on the metal surface of a cathode in a cathode-ray tube, producing a photocurrent. Use Einstein's theory to describe what happens in each step of this phenomenon.
2. Briefly explain how Einstein's theory accounts for each of the observations listed in Section 13.6.
3. Describe an experiment that tested Einstein's theory. What was the result?
4. Einstein's idea of a quantum of light had a definite relation to the wave model of light. What was it?
5. Why does the photoelectron freed from the metal surface not have as much energy as the quantum of light that causes it to be ejected?
6. What does a stopping voltage of 2.0 V indicate about the photoelectrons emerging from a metal surface?
7. Write down Einstein's photoelectric equation and define the meaning of every symbol in it.
8. Einstein's equation contains the maximum kinetic energy, $KE_{\max}$ of the photoelectrons. Under what conditions would the photoelectrons have less than maximum kinetic energy?
9. The success of Einstein's theory posed a dilemma for physicists. What was it?

13.8 X Rays

1. What processes can produce X rays?
2. X rays were the first "ionizing" radiation discovered. What does "ionizing" mean?
3. What are three properties of X rays that led to the conclusion that X rays were electromagnetic waves?
4. What was the experimental and theoretical evidence to support the notion that X rays had very short wavelengths?

DISCOVERY QUESTIONS

1. If classical physics was so successful, why did new theories arise?
2. Evaluate Thomson's "plum-pudding model" as a model of atoms. Was it a good model in terms of what scientists wanted from a model and what they knew at the time about the atom? What were its strengths and weaknesses?
3. A television tube or a computer-monitor tube is often called a CRT, which stands for "cathode-ray tube." Look up the operation of one of these tubes in an encyclopedia or on-line information source. What do these tubes have to do with cathode rays? How do they work?
4. At light frequencies below the threshold frequency no photoelectrons are emitted. What might happen to the light energy?
5. Planck's constant h has the value 6.6×10^{-34} J/s. What does this tell us about the magnitude of the energy carried by a light quantum?

Quantitative

1. The charge on an electron is 1.6×10^{-19} C. To gain an idea of the small size of an electron charge consider the following. A current of 1 A flows for 1 s down a wire. How many electrons enter the wire?
2. How many joules of energy does one photon of ultraviolet light carry? Assume its frequency to be 1.5×10^{15} Hz.
3. An average photon of visible light has a frequency of about 1×10^{14} Hz. To gain an idea of the small size of a quantum of energy, consider the following. A 100-W light bulb is turned on for 1 s. Only about 5% of the electric energy input to the tungsten filament is given off as visible light. How many photons does the bulb emit?
4. For most metals, the work function W in Einstein's photoelectric equation is about 10^{-18} J. Light of what frequency will cause photoelectrons to leave the metal with virtually no kinetic energy? In what region of the spectrum is this frequency?
5. Monochromatic light of wavelength 5×10^{-7} m falls on a metal cathode to produce photoelectrons. The light intensity at the surface of the metal is 10^2 J/m^2 s.
   (a) What is the frequency of the light?
   (b) What is the energy (in joules) of a single photon of the light?
   (c) How many photons fall on 1 m^2 in 1 s?
   (d) If the diameter of an atom is about 10^{-10} m, how many photons fall on one atom in 1 s, on the average?
   (e) How often would one photon fall on one atom, on the average?
   (f) How many photons fall on one atom in 10^{-10} s, on the average?
   (g) Suppose the cathode is a square 0.05 m on a side. How many electrons are released per second, assuming every photon releases a photoelectron? (In fact, only about 1 in 50 photons does so.) How big a current would this be in amperes?