## CHAPTER 17

# Probing the Nucleus

## 17.1 QUESTIONS ABOUT THE NUCLEUS

We saw in Chapter 14 that studies of the atom indicated that the atom consists of a very small, positively charged nucleus surrounded by negatively charged electrons. Experiments on the scattering of $\alpha$ particles revealed that the nucleus has dimensions of the order of $10^{-14}$ m. Since the diameter of an atom is of the order of $10^{-10}$ m, the nucleus takes up only a minute fraction of the volume of an atom. The nucleus, however, contains nearly all of the mass of the atom, as was also shown by the scattering experiments.

The existence of the atomic nucleus (named by analogy with the nucleus of a living cell) and its properties raised many questions similar to those raised about the atom. Is the nucleus itself made up of still smaller units? If so, what are these units, and how are they arranged in the nucleus? What methods can be used to get answers to these questions? What experimental evidence can be used as a guide?

The study of the properties and structure of atoms needed new physical methods. The methods that could be used to study the properties of bodies of ordinary size, that is, those with dimensions of the order of centimeters or meters, could not yield information about the structure of atoms. It is reasonable to expect that it is still more difficult to get information about what goes on inside the nucleus, which is such a small part of the atom. New kinds of experimental data must be obtained. New theories must be devised to help correlate and explain the data. In these respects, the study of the nucleus is still another step on the long road from the very large to the very small, along which we have been traveling in this course, following the historical development of physical science.

## 17.2 BECQUEREL'S DISCOVERY

One of the first and most important clues to an understanding of the nucleus occurred with the discovery of the phenomenon later known as *radioactivity* in early 1896 by the French physicist Henri Becquerel. It was an-

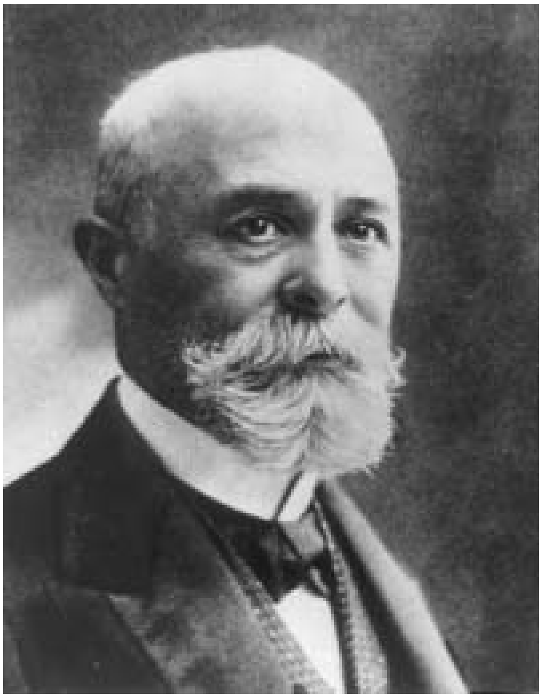

**FIGURE 17.1** Henri Becquerel (1852–1908) received the 1903 Nobel Prize in physics together with Pierre and Marie Curie (for the discovery of natural radioactivity).

other of those "accidents" that illustrate how the trained and prepared mind is able to respond to an unexpected observation.

Only 2 months before, in November 1895, Röntgen had discovered X rays. In doing so, he had unwittingly set the stage for the discovery of radioactivity. Röntgen had found that X rays came from the glowing spot on a glass tube where a beam of cathode rays (high-speed electrons) was hitting (see Section 13.7). When the cathode-ray beam was turned off, the spot of light on the face of the glass tube disappeared; the X rays coming from that spot also stopped.

The emission of light by the glass tube when it is excited by the cathode-ray beam is an example of the phenomenon called *fluorescence*, which was well known before Röntgen's work. A considerable amount of research had been done on fluorescence during the latter part of the nineteenth century. A substance is said to be fluorescent if it immediately emits visible light when struck by:

(1) visible light of shorter wavelength;
(2) invisible radiations, such as ultraviolet light; or
(3) the beam of electrons that make up cathode rays.

Fluorescence stops when the exciting light is turned off. (The term *phosphorescence* is generally applied to a related phenomenon, the emission of visible light that continues *after* the exciting light is turned off.)

Röntgen's observation that the X rays came from the spot that also showed fluorescence raised his suspicion that there was a close connection between X rays and fluorescence. Becquerel was fortunate in having the necessary materials and training to study this problem. In addition, he was the son and grandson of physicists who had made important contributions to the field of fluorescence and phosphorescence. In his Paris laboratory, Becquerel had devised an instrument for examining materials in complete darkness a small fraction of a second after they had been exposed to a brilliant light. The question occurred to Becquerel: When bodies are made to fluoresce (or phosphoresce) in the visible region with sufficient intensity, do they also emit X rays in addition to the light rays? He tested a number of substances by exposing them to sunlight; his method of checking whether they also emitted invisible X rays followed Röntgen's idea: Is a well-wrapped photographic plate exposed when near the source of such invisible rays? One of the samples Becquerel used happened to be a salt of the metal uranium, a sample of potassium-uranyl sulfate. In his words:

> I wrapped a . . . photographic plate . . . with two sheets of thick black paper, so thick that the plate did not become clouded by ex-

> posure to the sun for a whole day. I placed on the paper a crust of the phosphorescent substance, and exposed the whole thing to the sun for several hours. When I developed the photographic plate I saw the silhouette of the phosphorescent substance in black on the negative. If I placed between the phosphorescent substance and the paper a coin or a metallic screen pierced with an open-work design, the image of these objects appeared on the negative. The same experiment can be tried with a thin sheet of glass placed between the phosphorescent substance and the paper, which excludes the possibility of a chemical action resulting from vapors which might emanate from the substance when heated by the sun's rays.
>
> We may therefore conclude from these experiments that the phosphorescent substance in question emits radiations which penetrate paper that is opaque to light. . . .

In his published paper, Becquerel was careful to conclude from his experiment only that "penetrating radiations" were emitted from the phosphorescent substance. He did not write that the substance emitted X rays while it phosphoresced, because he had not fully verified that the radiations were X rays—though the radiations were transmitted through the black paper, just as X rays are—or that they were actually related to the phosphorescence (though he strongly suspected that they were). Before he could investigate these possibilities, he made this discovery:

> . . . among the preceding experiments some had been made ready on Wednesday the 26th and Thursday the 27th of February [1896]; and as on those days the sun only showed itself intermittently, I kept my arrangements all prepared and put back the holders in the dark in the drawer of the case, and left in place the crusts of uranium salt. Since the sun did not show itself again for several days, I developed the photographic plates on the 1st of March, expecting to find the images very feeble. On the contrary, the silhouettes appeared with great intensity. I at once thought that the action might be able to go on in the dark. . . .

Further experiments verified this surprising thought. Even when the uranium compound was not being excited by sunlight to phosphoresce, it continually emitted something that could penetrate black paper and other substances opaque to light, such as thin plates of aluminum or copper. Becquerel found that all the compounds of uranium, many of which were not phosphorescent at all, and metallic uranium itself had the same property. The amount of action on the photographic plate did not depend on what the particular compound of uranium was, but only on the amount of uranium present in it!

Becquerel also found that the persistent radiation from a sample of uranium did not appear to change, either in intensity or character, with the passing of time during days, weeks, and months. Nor was a change in the activity observed when the sample of uranium or of one of its compounds was exposed to ultraviolet light, infrared light, or X rays. Moreover, the intensity of the uranium radiation (or "Becquerel rays," as they came to be known) was the same at room temperature (20°C), at 200°C, and at the temperature at which oxygen and nitrogen (air) liquefy, about −190°C. Thus, *these rays seemed unaffected by physical (and chemical) changes of the source*.

Becquerel also showed that the radiations from uranium produced ionization in the surrounding air. They could discharge a positively or negatively charged body such as an electroscope. So the uranium rays resemble X rays in two important respects: their penetrating power and their ionization power. Both kinds of rays were invisible to the unaided eye, but both affected photographic plates. Still, X rays and Becquerel rays differed

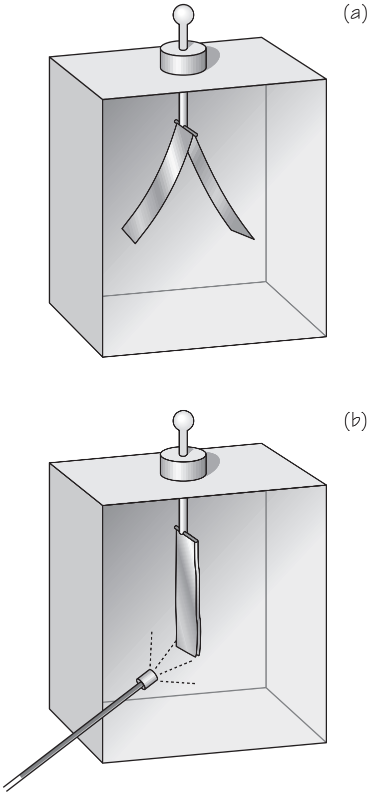

**FIGURE 17.2** The ionizing effect of the Becquerel rays could be demonstrated with a charged electroscope. When a sample of uranium is held near the electroscope leaves, the rays cause gas molecules in the air to ionize—that is, to become electrically charged. Ions, with a charge opposite to that on the leaves, drift to the leaves and neutralize their charge. The time taken for the leaves to fall is a measure of the rate of ionization of the gas and, therefore, of the activity of the uranium source.

in at least two important ways: Compared to X rays, these newly discovered rays from uranium needed no cathode-ray tube or even light to start them, and they could not be turned off. Becquerel showed that even after a period of 3 years a given piece of uranium and its compounds continued to emit radiations spontaneously.

The years 1896 and 1897 were years of great excitement in physics, to a large extent because of the interest in the recently discovered X rays and in cathode rays (electrons). It quickly became evident that X rays could be used in medicine, and they were the subject of much research. In comparison, the properties of the Becquerel rays were less spectacular, and little work was done on them in the period from the end of May 1896 until the end of 1897. In any case, it seemed that somehow Becquerel rays were special cases of X-ray emission. Even Becquerel himself turned to other work. But attention began to be attracted by the fact that the invisible rays from the uranium and its compounds appeared spontaneously.

Two questions were asked. First, what was the source of the energy creating the uranium rays and making it possible for them to penetrate opaque substances? Second, did any other of the 70 or more elements known then have properties similar to those of uranium? The first question was not answered for some time, although it was considered seriously. The second question was answered early in 1898 by the Curies, who, by doing so, opened a whole new field of research in physical science.

## 17.3 THE CURIES DISCOVER OTHER RADIOACTIVE ELEMENTS

One of Becquerel's colleagues in Paris was the physicist Pierre Curie, who had recently married a Polish-born physicist, Marja Sklodowska, who became Marie Curie. Marie Curie undertook a systematic study of the Becquerel rays and looked for other elements and minerals that might emit them. Using a sensitive type of electrometer that her husband had recently invented, she measured the small electric current produced when the rays ionized the air. She assumed this current is proportional to the intensity of the rays (and it actually is). With this new technique, Curie could give a numerical value to the ionizing effect produced by the rays. These values were reproducible within a few percent from one experiment to the next with the same sample.

One of Marie Curie's first results was the discovery that the element thorium (Th) and its compounds emitted radiations with properties similar to those of the uranium rays. (The same finding was made independently in Germany by Gerhardt C. Schmidt, at about the same time.) The fact that thorium emits rays like those of uranium was of great importance; it showed that the mysterious rays were not a property peculiar just to one element. The discovery spurred the search for still other elements that might emit similar rays. The fact that uranium and thorium were the elements with the greatest known atomic masses indicated that the very heavy elements might have special properties different from those of the lighter elements.

The evident importance of the problems raised by the discovery of the uranium and thorium rays led Pierre Curie to lay aside his researches in other fields of physics and to join his wife to work on these new problems. They began a herculean task. First, they found that the intensity of the emission from any thorium compound was directly proportional to the fraction by weight of the metallic element thorium present. (Recall that Becquerel had found a similar result for uranium compounds.) Moreover, the amount of radiation was independent of the physical conditions or the chemical combination of the active elements. These results led the Curies to the conclusion that the emission of the rays depended only on the presence of *atoms* of either of the two elements—uranium or thorium. Atoms of other elements that were present were simply inactive or absorbed some of the radiation.

These ideas were especially important because they helped the Curies interpret their later experiments. For example, in their studies of the radiation from minerals they examined *pitchblende*, an ore containing about 80% uranium oxide ($U_3O_8$). They found that the emission from pitchblende, as measured by its effect in ionizing air, was about four or five times as great as that to be expected from the amount of uranium in the ore. The other elements known at the time to be associated with uranium in pitchblende, such as bismuth and barium, had been shown to be not active, or as Marie Curie now called it, not "radioactive." If emission of rays is an atomic phenomenon, the unexpected activity of pitchblende could be explained only by the presence of another, hitherto undiscovered, element in pitchblende, an element more active than uranium itself.

To explore this hypothesis, the Curies applied chemical separation processes to a large quantity of pitchblende to try to isolate this hypothetical active substance. After each separation process, the products were tested, the inactive part discarded, and the active part analyzed further. Finally, the Curies obtained a highly active product that presumably consisted mainly of the unknown element. In a note titled "On a New Radioactive Substance Contained in Pitchblende," which they submitted to the French Academy of Sciences in July of 1898, they reported:

> By carrying on these different operations . . . finally we obtained a substance whose activity is about 400 times greater than that of uranium. . . .

## MARIE AND PIERRE CURIE

Pierre Curie (1859–1906) studied at the Sorbonne in Paris. In 1878, he became an assistant teacher in the physical laboratory there, and some years later, professor of physics. He was well known for his research on crystals and magnetism. Pierre and Marie Sklodowska were married in 1895 (she was 28 years old). After their marriage, Marie chose radioactivity for her doctoral research. In 1898, Pierre joined his wife in this work. Their collaboration was so successful that in 1903 they were awarded the Nobel Prize in physics, which they shared with Becquerel. Pierre Curie was run over and killed by a horse-drawn vehicle in 1906. Marie Curie was appointed to a professorship at the Sorbonne, the first woman there to have this post.

In 1911, Marie Curie was awarded the Nobel Prize in chemistry for the discovery of the two new elements, radium and polonium. She was the first person to win two Nobel Prizes in science. The rest of her career was spent in the supervision of the Paris Institute of Radium, a center for research on radioactivity and the use of radium in the treatment of cancer. During her visit to the U.S., a group of women

(*a*) (*b*)

**FIGURE 17.3** (a) Marie Curie. (b) Marie and Pierre Curie on a bicycling holiday. (*Continued*) presented her with the precious gifts of 1 gram of radium.

Marie Curie died in 1934 of leukemia, a form of cancer of the leukocyte-forming cells of the body, probably caused by overexposure to the radiations from radioactive substances during her research over more than three decades.

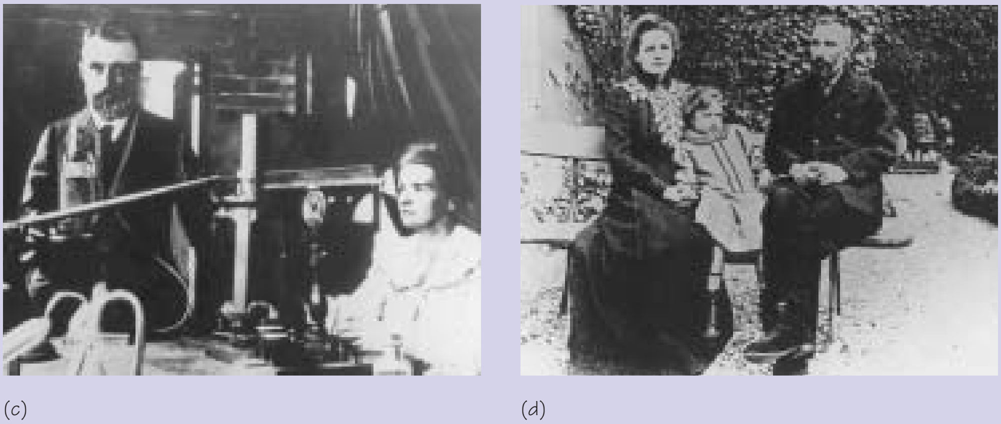
(*c*) (*d*)

**FIGURE 17.3** (*Continued*) (c) Marie and Pierre Curie in their laboratory; (d) Marie and Pierre Curie with their daughter, Irène, who later also received a Nobel Prize.

> We believe, therefore, that the substance which we removed from pitchblende contains a metal which has not yet been known, similar to bismuth in its chemical properties. If the existence of this new metal is confirmed, we propose to call it *polonium*, after the name of the native country of one of use.

Six months after the discovery of polonium, the Curies chemically separated another substance from pitchblende. They found the emission from it so intense as to indicate they had found yet another new element, even more radioactive than polonium! This substance had an activity per unit mass nine hundred times that of uranium and was chemically entirely different from uranium, thorium, or polonium. Spectroscopic analysis of this substance revealed spectral lines characteristic of the inactive element barium, but also a line in the ultraviolet region that did not seem to belong to any known element. The Curies reported their belief that the substance, "although for the most part consisting of barium, contains in addition a new element which produced radioactivity and, furthermore, is very near barium in its chemical properties." For this new element, so extraordinarily radioactive, they proposed the name *radium*.

The next step in making the evidence for the newly discovered elements more convincing was to determine their properties, especially their atomic masses. The Curies had made it clear that they had not yet isolated either polonium or radium in pure metallic form, or even yet obtained a pure sample of a compound of either element. From the substance containing the strongly radioactive substance that they called radium, they had separated a part consisting of barium chloride mixed with a presumably very small quantity of radium chloride. Additional separations by chemical means yielded an increasing proportion of radium chloride. The difficulty of this task is indicated by the Curies' remark that radium "is very near barium in its chemical properties," for it is very difficult to separate elements whose chemical properties are similar. Moreover, to obtain their highly radioactive substances in usable amounts, they had to start with a very large amount of pitchblende.

With an initial 100-kg shipment of pitchblende (from which the uranium salt had been removed to be used in the manufacture of glass) the Curies went to work in an abandoned woodshed at the School of Physics where Pierre Curie taught. Having failed to obtain financial support, the Curies made their preparations without technical help in this "laboratory." Marie Curie wrote later:

> I came to treat as many as twenty kilograms of matter at a time, which had the effect of filling the shed with great jars full of precipitates and liquids. It was killing work to carry the receivers, to pour off the liquids and to stir, for hours at a stretch, the boiling material in a smelting basin.

From the mixture of radium chloride and barium chloride they produced, only the average atomic mass of the barium and radium could be computed. At first an average value of 146 u was obtained, as compared to 137 u for the atomic mass of barium. After many additional purifications that increased the proportion of radium chloride, the average value for atomic mass rose to 174 u. (The unit u is the standard unit used today for the atomic mass, the hydrogen atom having an atomic mass of about 1 u.) Continuing the tedious purification process for 4 years, during which she treated several tons of pitchblende residue, Marie Curie was able to report, in July 1902, that she had isolated 0.1 g of radium chloride, so pure that spectroscopic examination showed no evidence of any remaining barium. She calculated the atomic mass of radium

The present yield of radium from 1 ton of high-grade uranium ore is about 0.2 g.

to be 225 u (the present-day value is 226.03 u). The activity of radium is more than a million times that of the same mass of uranium.

## 17.4 IDENTIFYING THE RAYS

Once the extraordinary properties of radium became known, they excited interest both inside and outside the scientific world, and the number of people studying radioactivity increased rapidly. The main question that attracted attention was: What are the mysterious radiations emitted by radioactive bodies?

In 1899, Ernest Rutherford, whose later development of the theory of the nuclear atom was discussed in Chapter 14, started to seek answers to this question. Rutherford found that a sample of uranium emits at least two distinct kinds of rays: one that is very readily absorbed, which he called for convenience $\alpha$ rays (alpha rays), and the other more penetrating, which he called $\beta$ rays (beta rays). In 1900, the French physicist P. Villard observed that the emission from radium contained rays much more penetrating than even the $\beta$ rays; this type of emission was given the name $\gamma$ (gamma) rays. The penetrating power of the three types of rays, as known at the time, is compared in the table below, first published by Rutherford in 1903.

**Thickness of Aluminum Required to Reduce the Radiation Intensity to One-half Its Initial Value**

| *Radiation type* | *Thickness of aluminum* |
|---|---|
| $\alpha$ | 0.0005 cm |
| $\beta$ | 0.05 cm |
| $\gamma$ | 8 cm |

Thus, the “Becquerel rays” were more complex than had been thought even before the nature of $\alpha$, $\beta$, and $\gamma$ rays was ascertained. Of the three kinds of rays, the $\alpha$ rays are the most strongly ionizing and the $\gamma$ rays the least. The power of penetration is inversely proportional to the power of ionization. This is to be expected; the penetrating power of the $\alpha$ rays from uranium is low because they expend their energy very rapidly in causing intense ionization. Alpha ($\alpha$) rays can be stopped—that is, almost all are absorbed—by about 0.0005 cm of aluminum, by a sheet of ordinary writing paper, or by a few centimeters of air. Beta ($\beta$) rays are completely stopped only after traveling many meters in air, or 0.05 cm in aluminum. Gamma ($\gamma$) rays can pass through many centimeters of aluminum or lead, or through a meter of concrete, before being almost completely absorbed. One consequence of these properties of the rays is that heavy and expensive shielding is sometimes needed in the study or use of radiations in accelerator and nuclear reactors, to protect people from the harmful effects of the rays. The rays ionize and, consequently, break down molecules in living cells, causing radiation "burns," fatal injuries to cells, damage that can lead to the growth of cancer cells, and dangerous mutations in the structure of the DNA molecules. In some cases, these "radiation shields" are as much as 3 m thick, but they are effective, protecting workers at reactor stations and accelerator research facilities.

## 17.5 THE CHARGE AND MASS OF THE RAYS

Another method used to study the rays emitted in radioactivity was to direct them through a magnetic field to see if they were deflected or deviated from their initial directions by the action of the field. This method came to provide one of the most widely used tools for the study of atomic and nuclear events. It is based on the now familiar fact that a force acts on a charged particle when it moves across a magnetic field. As was discussed in Section 10.11, this force always acts at right angles to the direction of motion of the charged particle. The particle experiences a continual deflection and, if sent into a uniform field at right angles, it moves along the arc of a circle.

This property had been used in the 1890s by J.J. Thomson in his studies of cathode rays. He showed that these rays consist of very small negatively charged particles, or electrons (Chapter 13). Becquerel, the Curies, and others found that the $\alpha$, $\beta$, and $\gamma$ rays behaved differently from one another in a magnetic field. The behavior of the rays is illustrated in Figure 17.4.

Suppose that some radioactive material, such as a sample of uranium, is placed at the end of a narrow hole in a lead block and that a narrow beam consisting of $\alpha$, $\beta$, and $\gamma$ rays escapes from the opening. If the beam enters a strong, uniform magnetic field (as in the last two drawings), the three types of rays will go along paths separated from one another. The $\gamma$ rays continue in a straight line without any deviation. The $\beta$ rays will be deflected to one side, moving in circular arcs of differing radii. The $\alpha$ rays will be deflected slightly to the other side, moving in a circular arc of large radius; they are rapidly absorbed in the air.

The direction of the deflection of the $\beta$ rays in such a magnetic field is the same as that observed earlier in Thomson's studies of the properties of cathode rays. It was concluded, therefore, that the $\beta$ rays, like cathode rays,

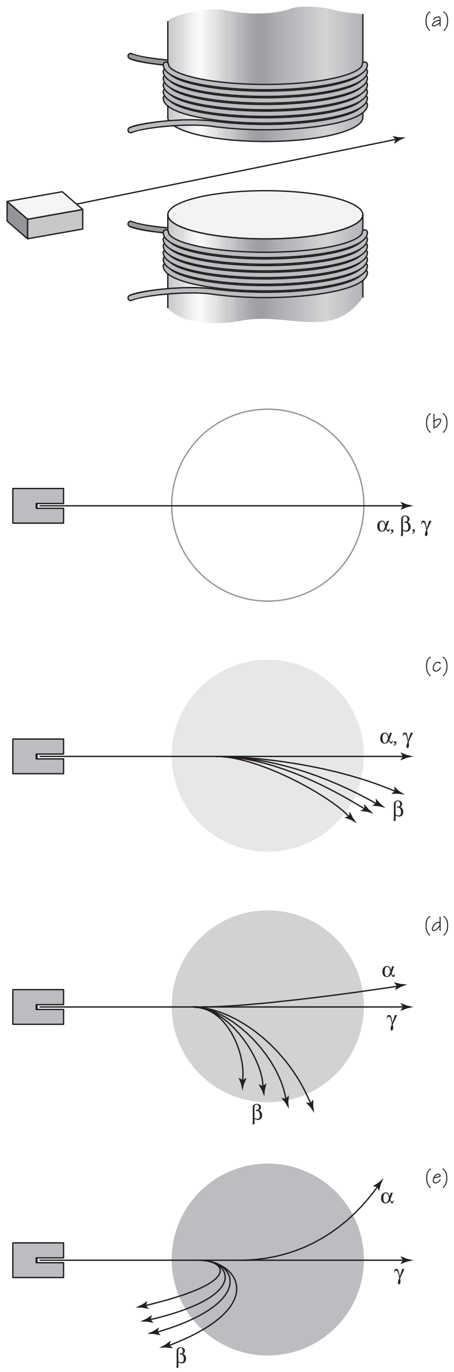

**FIGURE 17.4** (a) Alpha, beta, and gamma rays are separated from a sample of radioactive material by their passage through a magnetic field; (b) no magnetic field present; (c) weak magnetic field; (d) stronger magnetic field; (e) very strong magnetic field.

consist of *negatively charged particles*. (The Curies confirmed the negative charge on the $\beta$ particles in 1900; they caused the beam of the particles to enter an electroscope that became negatively charged.) Since the direction of the deflection of the $\alpha$ rays was opposite that of the $\beta$ rays, it was concluded that the $\alpha$ rays consist of *positively charged particles*. Since the $\gamma$ rays were not deflected at all, it was concluded that they are neutral, that is, they have no electric charge. Electromagnetic radiation is neutral, as are particles that carry equal amounts of positive and negative charge. No conclusion could be drawn from this type of experiment as to whether the $\gamma$ rays are, or are not, particles.

The deflection of a charged particle in electric and magnetic fields depends on both its charge and its mass. Therefore, the ratio of charge ($q$) to mass ($m_q$) for $\beta$ particles can be calculated from measured deflections in fields of known intensity. Becquerel, investigating $\beta$ particles in 1900, used a procedure that was essentially the same as that used by J.J. Thomson in 1897 to obtain a reliable value for the ratio of charge to mass for the particles in cathode rays. At that time, the fact that a consistent single value of $e/m_e$ had been found established quantitatively the *existence* of the electron (see Section 13.3). (Here $e$ is the electron's charge, and $m_e$ is its mass.) By sending $\beta$ rays through electric and magnetic fields, Becquerel was able to calculate the speed of the $\beta$ particles. He obtained a value of $q/m_q$ for $\beta$ particles which was close to that found by Thomson for the electron. This permitted the deduction that the *$\beta$ particles are electrons*. (However, they should not be thought of as existing in the nucleus until emitted. Rather, they are generated during the emission.)

The nature of the $\alpha$ radiation was more difficult to establish. It was necessary to use a very strong magnetic field to produce measurable deflections of $\alpha$ rays. The value of $q/m_q$ found for $\alpha$ particles ($4.8 \times 10^7$ C/kg) was about 4000 times smaller than $q/m_q$ for $\beta$ particles. The reason for the small $q/m_q$ value could be a small value of $q$ or a large value of $m$. Other evidence available at the time indicated that magnitude of the charge $q$ for an $\alpha$ particle was not likely to be smaller than that for a $\beta$ particle. It was therefore concluded that mass $m$ would have to be much larger for the $\alpha$ particle than for the $\beta$ particle.

The value of $q/m_q$ given above for $\alpha$ particles is just one-half that of $q/m_q$ found earlier for a hydrogen ion. The value would be explained in a reasonable way if the $\alpha$ particle were like a hydrogen molecule minus one electron ($H_2{}^+$), or if it were a helium atom (whose mass was known to be about four times that of a hydrogen atom) without its two electrons ($He^{++}$). Other possibilities might have been entertained. In fact, however, the right identification turned out to be that of $\alpha$ particles with $He^{++}$. *The $\alpha$ particle was* *found to be the same as a helium nucleus* and, therefore, has a mass of about four atomic mass units. The clever experiment described in the following section provided the final proof.

## 17.6 RUTHERFORD'S "MOUSETRAP"

The gas helium was first discovered to exist on the Sun through spectroscopic analysis of the Sun's radiation (Section 14.1). Later, it was discovered that helium gas can be found on the Earth, imprisoned in radioactive minerals. In addition, William Ramsey and Frederick Soddy, working in the Cavendish Laboratory in Cambridge, had discovered, in 1903, that helium is given off from a radioactive compound, radium bromide. This led Rutherford, also in Cambridge, to advance the hypothesis that the $\alpha$ particle is a double ionized helium atom ($He^{++}$), that is, a helium atom minus its two electrons, or, as we would now say, the nucleus of a helium atom. In a series of experiments conducted from 1906 to 1909, Rutherford succeeded in proving the correctness of his hypothesis in several different ways. The last and most convincing of these experiments was made in 1909, with T.D. Royds, by constructing what James Jeans later called "a sort of mousetrap for $\alpha$ particles."

The experiment used the radioactive element radon (Rn). Pierre Curie and André Debierne had discovered radon in 1901. They had found that a radioactive gas is given off from radium. A small amount of the gas collected in this way was found to be a strong $\alpha$ emitter. They showed the gas to be a new element, which they called "radium emanation," later called "radon." Ramsey and Soddy then found that when radon is stored in a closed vessel, helium gas always appears in the vessel also. Thus, helium is given off not only by radium but also by radon.

Rutherford and Royds put a small amount of radon in a fine glass tube with a wall only 0.01 mm thick. This wall was thin enough so that $\alpha$ particles could pass through it, but radon itself could not. The tube was sealed into a thick-walled, outer glass tube that had an electric discharge section at the top (Figure 17.5). They pumped the air out of the outer tube and allowed the apparatus to stand for about a week. During this time, while $\alpha$ particles from the radon passed through the thin walls of the inner tube, a gas was found gradually to collect in the previously evacuated space. Then they pumped mercury in at the bottom of the apparatus to compress the very small quantity of gas and confine it in the discharge tube. When a potential difference was applied to the electrodes of the discharge tube, an

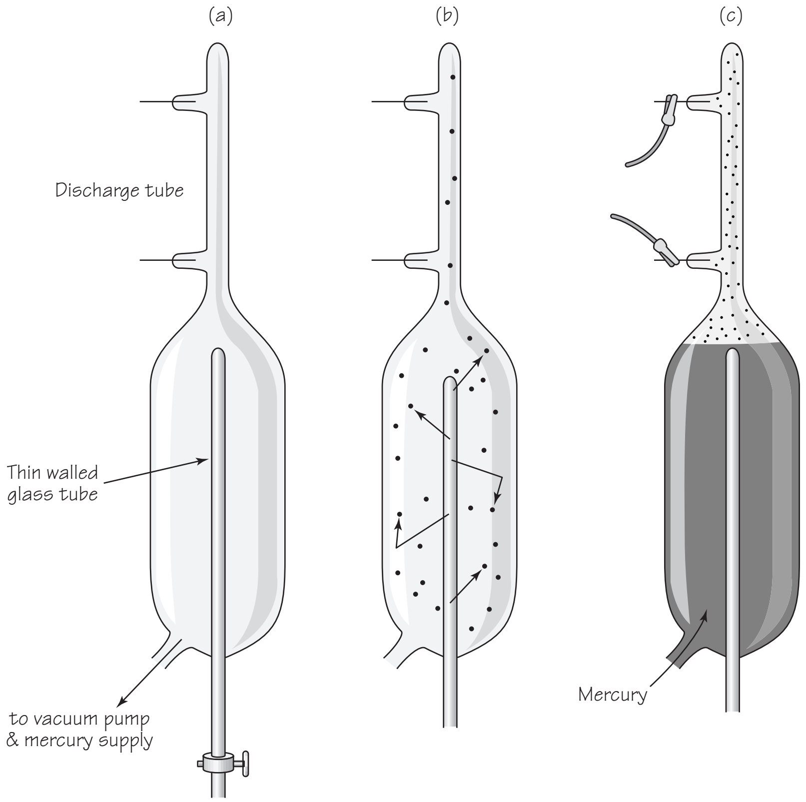

**FIGURE 17.5** Rutherford's "mousetrap" for identifying particles.

electric discharge was produced in the gas. They examined the resulting light with a spectroscope, and they found that the spectral lines were characteristic of the element helium. (In a separate control experiment, helium gas itself was put in the inner, thin-walled tube and did not leak through the wall of the inner tube.)

Now it was clear to Rutherford how to interpret his results. He could safely conclude that the helium gas that collected in the outer glass tube was formed from $\alpha$ particles that had passed into the outer tube and picked up some electrons to form helium atoms. Rutherford's result implied conclusions more important than just the identity of $\alpha$ particles. Apparently, an atom of an element (radon) can spontaneously emit a fragment (an $\alpha$ particle) that is the nucleus of *another* element (helium)! This was a startling idea, but only the beginning of more startling things to come.

## 17.7 RADIOACTIVE TRANSFORMATIONS

The emission of $\alpha$ and $\beta$ particles raised difficult questions with respect to existing ideas of matter and its structure. The rapid development of chemistry in the nineteenth century had made the atomic–molecular theory of matter highly convincing. According to this theory, a pure element consists of identical atoms, which are indestructible and unchangeable. But if a radioactive atom emits as substantial a fragment as an $\alpha$ particle, which was shown to be an ionized helium atom, can the radioactive atom remain unchanged? That did not seem plausible. Rather, it seemed that there must be a transformation in which the radioactive atom is changed to an atom of a different chemical element.

If an atom emits an $\alpha$ particle, a significant part of its mass—four atomic mass units—will be carried away by the $\alpha$ particle. What about the atoms that emit $\beta$ particles? The $\beta$ particle, which was shown to be an electron, is far less massive than the $\alpha$ particle. However, its mass is not zero; so a radioactive atom must also undergo some change when it emits a $\beta$ particle. It was again difficult to escape the conclusion that radioactive atoms are, in fact, subject to division into two parts of markedly unequal mass, a conclusion contrary to the old, basic concept that the atom is indivisible.

Another fundamental question arose in connection with the energy carried by the rays emitted by radioactive substances. As early as 1903, Rutherford and Soddy, and Pierre Curie and a young coworker, A. Laborde, noted that a sample of radium kept itself at a higher temperature than its surroundings merely by absorbing some of the energy of the $\alpha$ particles emitted by atoms inside the sample. Curie and Laborde found that 1 g of radium can produce about 0.1 kcal of heat per hour. A sample of radium thus can continue to release energy year after year, and evidently for a very long time.

The continuing release of such a quantity of heat could not be explained by treating radioactivity as an ordinary chemical reaction. It was clear that radioactivity did not involve chemical changes in the usual sense. Energy was emitted by samples of pure elements; energy emission by compounds containing radioactive elements did not depend on the type of molecule in which the radioactive element was present. The origin of the production of heat had to be sought in some deep changes *within* the atoms of radioactive elements, rather than in chemical reactions among atoms.

Rutherford and Soddy proposed a bold theory of *radioactive transformation* to explain the nature of these changes. They suggested that when a radioactive atom emits an $\alpha$ or a $\beta$ particle, it really breaks into two parts: the $\alpha$ or $\beta$ particle that is emitted, and a heavy, leftover part that is physically and chemically different from the "parent" atom. There was a good deal of evidence for the last part of the assumption. For example, the formation of radon gas from radium was known, as mentioned earlier. When the atomic mass of radon was determined, it turned out to be smaller than that of radium by just four atomic mass units, the mass of an $\alpha$ particle.

The idea of radioactive transformation can be represented by an "equation" similar to the kind used to represent chemical reactions. For example, using the symbols Ra and Rn to represent atoms of radium and radon, and He to represent an $\alpha$ particle after it has picked up two electrons to form a helium atom, the transformation of radium into radon can be expressed as

$$\mathrm{Ra} \rightarrow \mathrm{Rn} + \mathrm{He}.$$

The process can be described as the transformation—also called a "disintegration," "decay," or "transmutation"—of radium into radon, with the emission of an $\alpha$ particle.

Many decay processes, in addition to the example just cited, had been found and studied by the Curies, by Rutherford and his coworkers, and by others, and these processes fitted easily into the kind of scheme proposed by Rutherford and Soddy (who received Nobel Prizes in chemistry for their work). For example, radon is radioactive also, emitting another $\alpha$ particle and thereby decaying into an atom of an element that was called "radium A" at the time. Radium A was later shown to be polonium (Po):

$$\mathrm{Rn} \rightarrow \mathrm{Po} + \mathrm{He}.$$

Polonium is also a radioactive solid. In fact, the original "parent" radium atoms undergo a series or chain of transformations into generation after generation of new, radioactive, so-called *daughter elements*, ending finally with a daughter element that is nonradioactive or, in other words, stable.

## 17.8 RADIOACTIVE DECAY SERIES

The decay of radium and its daughters was found eventually to lead to a stable end product that was identified by its chemical behavior as *lead.* The chain beginning with radium has 10 members, some emitting $\alpha$ particles and others emitting $\beta$ particles. Some $\gamma$ rays are emitted during the decay series, but $\gamma$ rays do not appear alone; the are emitted only together with an $\alpha$ particle or a $\beta$ particle.

Rutherford and Soddy suggested that, since radium is always found in uranium ores, such as pitchblende, which the Curies had analyzed, radium itself may be a member of a series starting with uranium as the ancestor of all the members. Research showed that this is indeed the case. Each uranium atom may in time give rise to successive daughter atoms, radium being the sixth generation and stable lead being the fifteenth.

The table shows all the members of the so-called *uranium–radium series.* The meaning of some of the symbols will be discussed in later sections. The number following the name of an element, as in uranium-238, indicates the atomic mass in atomic mass units. Notice that there are heavier and lighter varieties of the element, such as uranium-238 and uranium-235, polonium-218, 214, and 210. (More will be said about these varieties in Section 17.10.) Two other naturally occurring radioactive series have been found; one starts with thorium-232 and the other with uranium-235.

**Uranium–Radium Decay Series**

| *Substance* | *Decay mode* | *Half-life* |
|---|---|---|
| Uranium-238 | $\alpha$, $\gamma$ | $4.51 \times 10^9$ yr |
| Thorium-234 | $\beta$, $\gamma$ | 24.1 days |
| Protactinium-234 | $\beta$, $\gamma$ | 1.18 min |
| Uranium-234 | $\alpha$, $\gamma$ | $2.48 \times 10^5$ yr |
| Thorium-230 | $\alpha$, $\gamma$ | $8.0 \times 10^4$ yr |
| Radium-226 | $\alpha$, $\gamma$ | 1620 yr |
| Radon-222 | $\alpha$, $\gamma$ | 3.82 days |
| Polonium-218 | $\alpha$ | 3.05 min |
| Lead-214 | $\beta$, $\gamma$ | 26.8 min |
| Bismuth-214 | $\alpha$, $\beta$, $\gamma$ | 19.7 min |
| Polonium-214 | $\alpha$ | $1.64 \times 10^{-4}$ s |
| Lead-210 | $\beta$, $\gamma$ | 21.4 yr |
| Bismuth-210 | $\beta$, $\gamma$ | 5.0 days |
| Polonium-210 | $\alpha$, $\gamma$ | 138.4 days |
| Lead-206 | stable | stable |

Each member of the series differs physically and chemically from its immediate parent above or daughter below; it should therefore be possible to separate the different members in any radioactive sample. This is by no means impossible to do, but the separation problem is made difficult by the fact that the different radioactive species decay at different rates, some very slowly, some rapidly, others at intermediate rates. These rates and their meaning will be discussed in the next section, but the fact that the rates differ gives rise to important effects that can now be discussed.

Two other naturally occurring radioactive series have been found; one starts with thorium-232 and the other with uranium-235.

## Growth of Daughter Elements

An interesting effect is provided by that portion of the uranium–radium series that starts with the substance called polonium-218. A pure sample of polonium-218 may be collected by exposing a piece of ordinary material, such as a thin foil of aluminum, to the gas radon. Some of the radon atoms decay into polonium-218 atoms, which then stick to the surface of the foil. The graph in Figure 17.6 shows what becomes of the polonium-218. Polonium-218 ($^{218}$Po) decays into lead-214 ($^{214}$Pb), which decays into bismuth-214 ($^{214}$Bi), which decays into polonium-214 ($^{214}$Po), then lead-210 ($^{210}$Pb), then bismuth-210 ($^{210}$Bi), etc. If the original sample contains 1,000,000 atoms of $^{218}$Po when it is formed, after 20 min it will contain about 10,000 $^{218}$Po atoms, about 660,0000 $^{214}$Pb atoms, about 240,000 $^{214}$Bi atoms, and about 90,000 $^{210}$Pb atoms. The number of $^{214}$Po atoms is negligibly small because most of the $^{214}$Po changes into $^{210}$Pb in a small fraction of a second.

A sample of pure, freshly separated radium ($^{226}$Ra) would also change in composition in a complicated way, but much more slowly. Eventually it would consist of a mixture of some remaining radium-226, plus radon-222, polonium-218, lead-214, and all the rest of the members of the chain down to, and including, stable "radium G" (lead-206).

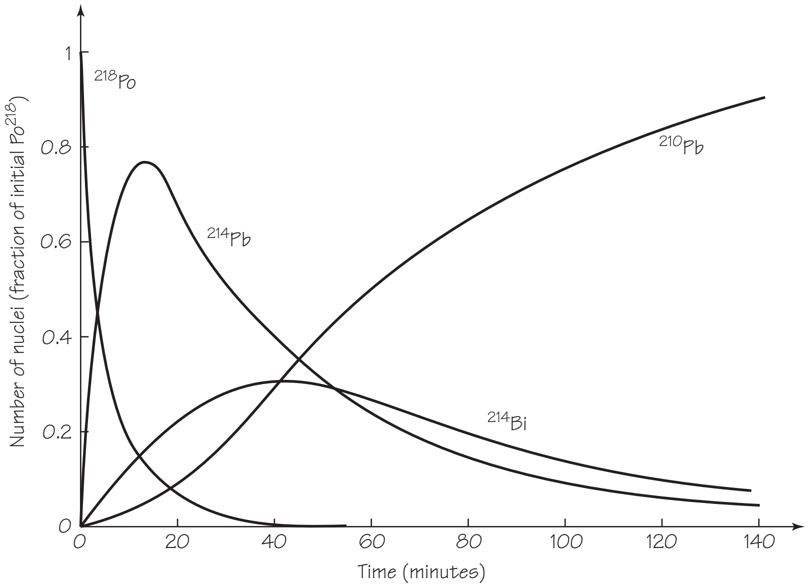

**FIGURE 17.6** Graph of decay of polonium-218 with build-up of its daughter elements over time.

Similarly, a sample of pure uranium may contain, after a time, 14 other elements of which 13 (all but the last, stable portion) contribute to the radioactive emission, each in its own way. In all such cases, a complicated mixture of elements results. After starting as a pure $\alpha$ emitter, a sample eventually emits many $\alpha$ particles, $\beta$ particles, and $\gamma$ rays, apparently continuously and simultaneously.

It is evident that the separation of the different members of a radioactive chain from one another would be difficult, especially if some members of the chain decay rapidly. The determination of the chemical nature and the radioactive properties of each member required great experimental ingenuity. One successful method depended on the skillful chemical purification of a particular radioactive substance, as the Curies had done with radium and polonium. For example, suppose that a sample has been obtained from which all the radioactive atoms except those of radium-226 have been removed. The sample immediately starts to give off radon gas. The latter can be drawn off and its properties examined before it becomes seriously contaminated by the distintegration of many of its atoms into polonium-218. If this is done, it is found that radon decays (through several transformations) into lead much more quickly than radium decays into radon.

## 17.9 DECAY RATE AND HALF-LIFE

In the last section, we noted that of 1,000,000 $^{218}$Po atoms present in a freshly prepared sample of that radioactive substance, only about 10,000 would remain after 20 min, the rest having decayed into atoms of $^{214}$Pb and its daughter products. It would take only 3 min following the preparation of the pure sample of $^{218}$Po for 50% of the atoms originally present in the sample to have decayed. In the case of radium ($^{226}$Ra), it would take 1620 years for half of the radium atoms in a freshly prepared sample of radium to be transformed into radon atoms.

These two examples illustrate the experimental fact that samples of radioactive elements show great difference in their rates of decay. These different rates are the result of *averages* of many individual, different decay events going on at random in a sample. Looking at *one* atom of any radioactive element, one never can tell when it will decay; some may decay as soon as they are produced, while others may never decay. Still, it has been found experimentally that for a large group of atoms of one kind, *the fraction of these atoms that decay per second is unchangeable* and always the same for any large group of atoms of that kind. This fraction is almost completely independent of all physical and chemical conditions, such as temperature, pressure, and form of chemical combination. These remarkable properties of radioactivity deserve special attention, and the meaning of the italicized statement above now will be discussed in detail because it is basic to an understanding of radioactivity.

Say, for example, that 1/1000 of the atoms in a freshly prepared pure sample decay during the first second. Then you would expect that 1/1000 of the remaining atoms will decay during the next second. But also, 1/1000 of the atoms remaining after 10 s will decay during the eleventh second, and so on. In fact, during any subsequent second of time, 1/1000 of the atoms remaining at the beginning of that second will decay, at least until the number of remaining atoms becomes so small that predictions become very uncertain.

Since the *fraction* of the atoms that decay per unit time is a constant for each element, the *number* of atoms that decay per unit time will decrease in proportion to the diminishing number of atoms that have not yet changed. Consequently, if the percentage of surviving, unchanged atoms is plotted as a function of time, a curve like the one in the sketch is obtained. The number of atoms in a sample that decays per unit time is termed the *activity* of the sample. Thus, the graph also represents the way in which the measured activity of a sample would decrease with time.

The curve that shows the number of atoms that have not decayed as a function of time approaches the time axis asymptotically; that is, the number of survivors becomes small, but only approaches zero. This is another way of saying that a definite "lifetime" in which all of the original atoms for a sample will have decayed cannot be assigned.

However, it is possible to specify the time required for any particular *fraction* of a sample to decay, one-half, one-third, or 37%, for instance. For convenience in making comparisons, the fraction one-half has been chosen. Rutherford called the time required for the decay of one-half of the original atoms of a pure sample the *half-life* (symbol $T_{1/2}$). Each kind of radioactive atom has a unique half-life, and thus the half-life of an element can be used to identify a radioactive element. As the table on p. 741 shows, a wide variety of half-lives have been found.

For $^{238}$U, the parent of the uranium series, the half-life is 4.5 billion years. This means that after $4.5 \times 10^9$ yr, half of the $^{238}$U atoms will have decayed. For $^{214}$Po, the half-life is of the order of $10^{-4}$ s; that is, in only 1/10,000 of a second, half of an original sample of $^{214}$Po atoms will have decayed. If pure samples of each, containing the same number of atoms, were available, the initial activity (atoms decaying per second) of $^{214}$Po would be very strong and that of $^{238}$U very feeble. If left for even 1 min, though, the polonium would have decayed so thoroughly and, therefore, the number of its surviving atoms would be so small, that at this point the

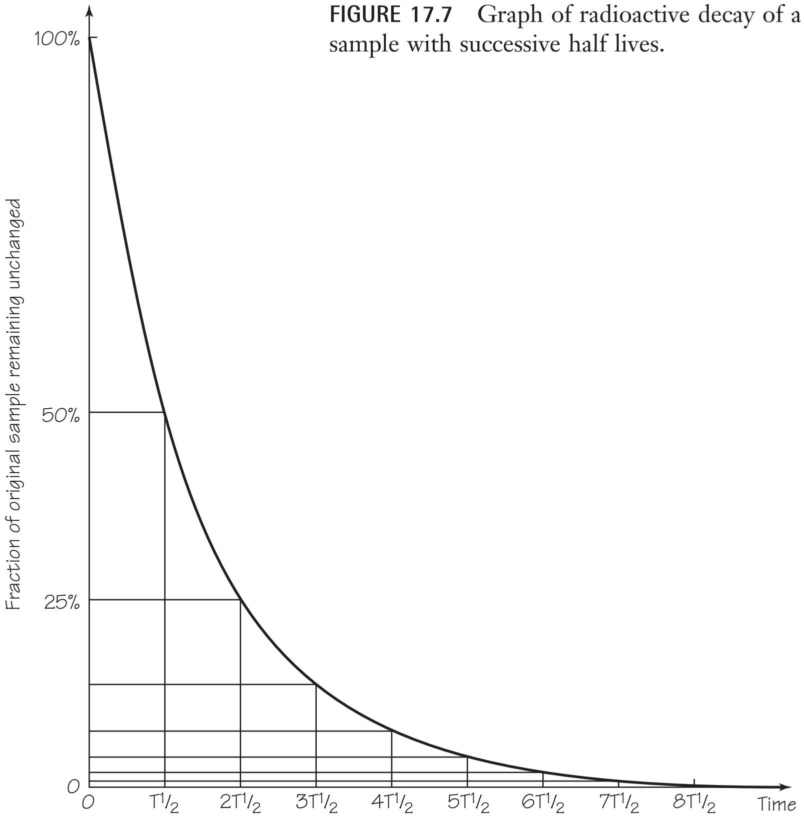

**FIGURE 17.7** Graph of radioactive decay of a sample with successive half lives.

activity due to polonium would now be less than the activity of the uranium atoms. Perhaps some radioactive elements, present in great quantities long ago, decayed so rapidly that no measurable traces are now left. On the other hand, many radioactive elements decay so slowly that during any ordinary experimentation time the counting rates that indicate decay seem to remain constant. (That is why Becquerel did not notice any change in the activity of his uranium salt samples.)

The principal advantage of the concept of a half-life lies in the experimental result implied in the graph: For any element of half-life $T_{1/2}$, no matter how old a sample is, on average half of the atoms will still have survived after an additional time interval $T_{1/2}$. Thus, the half-life is not to be thought of as an abbreviation for "half a life." If one-half the original atoms remain unchanged after a time $T_{1/2}$, one-fourth, $\frac{1}{2} \times \frac{1}{2}$, will remain after two consecutive half-life intervals $2T_{1/2}$, one-eighth after $3T_{1/2}$, and so on. Note how different the situation is for a population of, say, human beings instead of radioactive atoms. In a group of $N_0$ babies, half the number may survive to their 70th birthday; of these $N_0/2$ senior citizens, none is likely to celebrate a 140th birthday. But of $N_0$ radioactive atoms with a half-life of 70 years, on average $N_0/4$ will have remained intact after 140 yr, $N_0/8$ after 210 yr, etc. To put it differently, *the statistical probability of survival for atoms is unchanged by the age they have already reached*. In humans, of course, the probability of survival (say, for another year) depends strongly on age, and so the concept "half-life" is not usable in this case.

We have been considering the behavior not of individual atoms, but of a very large number of them. As discussed in connection with the behavior of gases in Chapter 7, this method allows us to use laws of statistics to describe the average behavior of the group. If a hundred thousand people were to flip coins simultaneously just once, you could predict with good accuracy that about one-half of them would get heads. But you could not accurately predict that one particular person in this crowd would obtain heads on a single flip. If the total number of coins tossed is small (10), the observed count is likely to differ considerably from the prediction of 50% heads. From experiments in radioactivity, you can predict that a certain fraction of a relatively large number of atoms in a sample will survive in any given time interval (e.g., one-half will survive to reach the age $T_{1/2}$), but not whether a particular atom will be among the survivors. As the sample of survivors decreases in size owing to disintegrations, predictions become less precise. Eventually, when only a few unchanged atoms are left, you could no longer make useful predictions at all. In short, the disintegration law is a *statistical* law and is thus applicable only to large populations of the radioactive atoms. Moreover, it makes no assumptions as to *why* the atoms disintegrate.

In the discussion of the kinetic theory of matter, you saw that it is a hopeless and meaningless task to try to describe the motions of each individual molecule, but you could calculate the *average* pressure of a gas containing a very large number of molecules. Similarly, in dealing with radioactivity, the inability to specify when each of the tremendous number of atoms in a normal sample will disintegrate makes a statistical treatment necessary and useful.

## 17.10 THE CONCEPT OF ISOTOPES

The discovery that there are three radioactive series, each containing apparently new substances, created a serious problem. In 1910, there were still some empty spaces in the periodic table of the elements, but not enough spaces for the many new substances. The periodic table represents an arrangement of the elements according to their chemical properties, and, if it could not include the radioactive elements, it would have to be revised, perhaps in some drastic and fundamental way.

The clue to the solution of the puzzle lay in the observation that some of the newly found materials that were members of a radioactive series have *chemical* properties identical to those of well-known elements, although some of their *physical* properties are different. For example, what was then called Uranium II, the "great-granddaughter" of Uranium I, was found to have the same chemical properties as Uranium I itself. When both were mixed together, the two could not be separated by chemical means. No chemist has detected, by chemical analysis, any difference between these two substances. But the two substances, now known as uranium-238 and uranium-234, do differ from each other in certain physical properties. As the table on p. 741 shows, uranium-238 and -234 have quite different radioactive half-lives: $4.5 \times 10^9$ yr and $2.5 \times 10^5$ yr, respectively. The mass of a uranium-234 atom must be smaller than that of a uranium-238 atom by the mass of one $\alpha$ particle and two $\beta$ particles.

Another pair of radioactive substances, called then radium B and radium G, were found to have the same chemical properties as lead; when mixed with lead they could not be separated from it by chemical means. These substances are now known as lead-214 and lead-206, respectively. Lead-214 is radioactive, and lead-206 is stable. The decay series on p. 741 indicates that the atoms must differ from each other in mass by the mass of two $\alpha$ particles and four $\beta$ particles. There are many other examples of such physical differences among two or more radioactive substances with the same chemical behavior.

Soddy suggested a solution that threw a flood of light on the nature of matter and on the relationship of the elements in the periodic table. He proposed that a chemical element could be regarded as a pure substance only in the sense that all of its atoms have the same chemical properties; that is, a chemical element may in fact be a *mixture of atoms* having different radioactive behavior and different atomic masses, but all having the same chemical properties. This idea meant that one of the basic postulates of Dalton's atomic theory would have to be changed, namely, the postulate that the atoms of a pure element are alike in *all* respects. According to Soddy, it is only in chemical properties that the atoms of a given element are identical. The several physically different species of atoms making up a particular element occupy the same place in the periodic table, that is, have the same atomic number $Z$. Because of this, Soddy called them *isotopes* of the element, from the Greek words *isos* and *topos*, meaning *same* and *place*—the same place in the periodic table. Thus, uranium-238 ($^{238}U$) and uranium-234 ($^{234}U$) are isotopes of uranium ($_{92}U$); lead-214 ($^{214}Pb$) and

**FIGURE 17.8** Frederick Soddy (1877–1956), an English chemist, studied at Oxford and went to Canada in 1899 to work under Rutherford at McGill University in Montreal. There he worked out his explanation of radioactive decay. Soddy returned to England in 1902 to work with William Ramsay, the discoverer of the rare gases argon, neon, krypton, and xenon. Ramsay and Soddy showed, in 1903, that helium was continuously produced by naturally radioactive substances. In 1921, Soddy was awarded the Nobel Prize in chemistry for his discovery of isotopes. He was a professor of chemistry at Oxford from 1919 to 1936.

lead-206 ($^{206}$Pb) are isotopes of lead ($_{82}$Pb). They are *chemically the same*; they occupy the same place on the periodic table and have the same atomic number $Z$, given by the subscript. But they are *physically different*, because they have different atomic masses A, given by the superscript in atomic mass units.

With this idea in mind, the many species of radioactive atoms in the three radioactive series were soon shown by chemical analysis to be isotopes of one or another of the last 11 naturally occurring elements in the periodic table, from lead ($Z = 82$) to uranium ($Z = 92$). For example, the 2nd and 5th members of the uranium series (see the table on p. 741) were shown to be isotopes of thorium, with $Z = 90$; the 8th, 11th, and 14th members turned out to be isotopes of polonium ($Z = 84$). The old names and symbols given to the members of radioactive series upon their discovery were therefore rewritten to represent both the chemical similarity and physical difference among isotopes. The present names for uranium $X_1$ and ionium, for example, are thorium-234 and thorium-230 (as shown in the table, p. 741).

A modern “shorthand” form for symbolizing any species of atom, called a *nuclide*, is also given in the same table—for example, $^{234}_{90}\mathrm{Th}$ and $^{230}_{90}\mathrm{Th}$ for two of the isotopes of thorium. The subscript (90 in both cases for thorium) is the atomic number $Z$, the place number in the periodic table; the superscript (234 or 230) is the mass number $A$, the approximate atomic mass in atomic mass units.

Note that, when writing the symbol for a nucleus, the atomic mass is always given as a whole number (such as U-238); but the atomic mass of an element as given in the periodic table (see color plate 6) is in most cases a decimal fraction (such as Uranium, 238.03). This is because the atomic mass given in the periodic table refers to the relative mass of the element in its natural state, which is a mixture of the various naturally occurring isotopes of the element. The atomic mass of the natural element is then an average of the atomic masses of the individual isotopes, weighted according to their abundance in relation to the other isotopes. For example, natural hydrogen has 1.0080 atomic mass units. It is a mixture of the abundant isotope $^{1}_{1}\mathrm{H}$ and the much less abundant isotopes $^{2}_{1}\mathrm{H}$ and $^{3}_{1}\mathrm{H}$. (However, the masses of these atoms are *not exactly* 2.0000 u and 3.0000 u, respectively, because of the mass of the electron orbiting the nucleus in the neutral atom.)

## 17.11 TRANSFORMATION RULES

Two questions then arose: How do changes in chemical nature come about as an atom undergoes radioactive decay? More specifically, what determines whether the atomic number $Z$ increases or decreases in a given radioactive transformation?

In 1913, Soddy in England and A. Fajans in Germany answered these questions independently. They each proposed two rules that systematized all the relevant observations for natural radioactivity. They are called the *transformation rules of radioactivity*. By 1913 Rutherford’s nuclear model of the atom was generally accepted. Using this model, one could consider a radioactive atom to have an unstable nucleus that emits an $\alpha$ or a $\beta$ particle (sometimes with emission of a $\gamma$ ray). Every nucleus has a positive charge given by $Ze$, where $Z$ is the atomic number and $e$ is the magnitude of the charge of an electron. The nucleus is surrounded by $Z$ electrons that make the atom as a whole electrically neutral and determine the chemical behavior of the atom. As discussed above, an $\alpha$ particle has an atomic mass of about four units and a positive charge of two units, $+2e$. A $\beta$ particle has a negative charge of one unit, $e$, and very little mass compared to an $\alpha$ particle.

With this information in mind, the transformation rules may now be stated as follows:

1. When a nucleus emits an $\alpha$ particle, the mass of the atom decreases by four atomic mass units, and the atomic number $Z$ of the nucleus decreases by two units; the resulting atom belongs to an element two spaces back in the periodic table (see color plates).
2. When a nucleus emits a $\beta$ particle, the mass of the atom is changed very little, but the atomic number $Z$ increases by one unit; the resulting atom belongs to an element one place forward in the periodic table. (We will discuss, in Section 18.10, how this behavior can be used to produce nonnaturally occurring elements beyond uranium, that is, so-called transuranium elements.)
3. When only a $\gamma$ ray is emitted, there is no change in the number corresponding to the atomic mass, and none in the atomic number.

The table of the uranium–radium series (p. 741) shows how these rules apply to that series, at least as far as the atomic number is concerned.

These rules, applied to the Rutherford–Bohr model of the atom, help to explain why a change in chemical nature occurs as a result of $\alpha$ or $\beta$ emission. Emission of an $\alpha$ particle takes two positive charges from the nucleus and four atomic mass units from the atom. An example is the following:

$$^{218}_{84}\mathrm{Po} \rightarrow {}^{214}_{82}\mathrm{Pb} + \alpha.$$

The resulting new atom ($^{82}$Pb) with its less positive nucleus can hold in its outer shells two fewer electrons than before, so two excess electrons drift away. The chemical behavior of atoms is controlled by the number of electrons; therefore, the new atom acts chemically like an atom of an element with an atomic number two units *less* than that of the parent atom.

On the other hand, in the case of $\beta$ emission, the nucleus, and with it the whole atom, becomes *more* positively charged, by one unit. An example is the following:

$$^{234}_{90}\mathrm{Th} \rightarrow {}^{234}_{91}\mathrm{Pa} + \beta.$$

The number of electrons that the atom can hold around the nucleus has increased by one. After it has picked up an extra electron to become neutral again, the atom acts chemically as an atom with an atomic number one unit greater than that of the atom before the radioactive change occurred.

By using the transformation rules, Soddy and Fajans were able to determine the place in the periodic table for every one of the substances (or nuclides) in the radioactive series; no revision of the existing periodic table was needed. Many of the nuclides between $Z = 82$ (lead) and $Z = 92$ (uranium) are now known to contain several isotopes each. These results were expected from the hypothesis of the existence of isotopes, but direct, independent evidence was also sought and obtained in 1914.

## 17.12 SUMMARY OF NOTATION FOR NUCLIDES AND NUCLEAR REACTIONS

We summarize and recapitulate some of the above ideas and notations. (See the next section for some applications.) Because of the existence of isotopes, it is no longer possible to designate an atomic species only by means of the *atomic number* $Z$. To distinguish among the isotopes of an element some new symbols were introduced. One is the *mass number*, $A$, defined as the whole number closest to the measured atomic mass. For example, the lighter and heavier isotopes of neon are characterized by the pairs of values: $Z = 10$, $A = 20$, and $Z = 10$, $A = 22$. (An element that consists of a single isotope can, of course, also be characterized by its $Z$ and $A$ values.)

These values of $Z$ and $A$ are determined by the properties of the atomic nucleus. According to the Rutherford–Bohr model of the atom, the atomic number $Z$ is the magnitude of the positive charge of the nucleus in elementary charge units. The mass number $A$ is very nearly equal to (but a bit less than) the atomic mass of the atom, because the total mass of the electrons around the nucleus is very small compared to the mass of the nucleus.

The term *nuclide* is used to denote an atomic species characterized by particular values of $Z$ and $A$. An *isotope* is then *one* of a group of two or more nuclides, each having the same atomic number $Z$ but different mass numbers $A$. A radioactive atomic species is called a *radioactive nuclide*, or *radionuclide* for short. A nuclide is usually denoted by the chemical symbol with a subscript at the lower left giving the atomic number and a superscript at the upper left giving the mass number. Thus, in the symbol ${}^{A}_{Z}\mathrm{X}$ for a certain nuclide, $Z$ stands for the atomic number, $A$ stands for the mass number, and X stands for the chemical symbol. For example, ${}^{9}_{4}\mathrm{Be}$ is the nuclide beryllium with atomic number 4 and mass number 9; the symbols ${}^{20}_{10}\mathrm{Ne}$ and ${}^{22}_{10}\mathrm{Ne}$ represent the neon isotopes discussed above. Since the $Z$ value is the same for all the isotopes of a given element, it is often omitted, except when needed for balancing equations (as you will shortly see). Thus, you can write ${}^{20}\mathrm{Ne}$ for ${}^{20}_{10}\mathrm{Ne}$, or ${}^{238}\mathrm{U}$ for ${}^{238}_{92}\mathrm{U}$.

The introduction of the mass number and the symbol for a nuclide makes it possible to represent radioactive nuclides in an easy and consistent way.

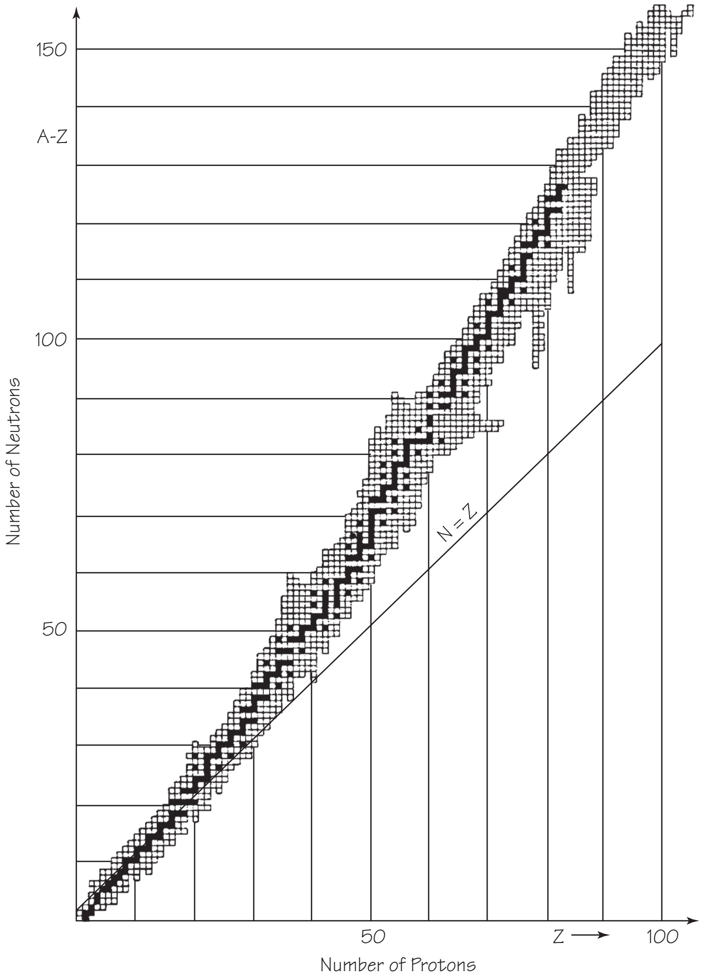

**FIGURE 17.10** A chart of the known nuclides. Each black square represents a stable natural nuclide. Each open square represents a known, unstable nuclide, with only a small number of these found naturally (the rest being artificial). Note that all isotopes of a given element are found in a vertical column centered on the element's atomic number, *Z*. As we will see in the next chapter, the *Z* number is the number of protons in the nucleus, and the difference between the atomic mass and the atomic number is the number of neutrons.

In addition, radioactive decay can be expressed by a simple "equation" representing the changes that occur in the decay process. For example, the first step in the uranium–radium series, namely, the decay of uranium-238 into thorium-234, may be written

$$^{238}_{92}\mathrm{U} \rightarrow\, ^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}.$$

The symbol $^{4}_{2}\mathrm{He}$ stands for the helium nucleus ($\alpha$ particle); the other two symbols represent the initial and final atomic nuclei, each with the appropriate charge and mass number. The arrow stands for "decays into." The "equation" represents a nuclear *reaction* and is analogous to an equation for a chemical reaction. The atomic numbers on the two sides of the equation must balance because the electric charge of the nucleus must be conserved: In the example above, 92 = 90 + 2. Also, the mass numbers must balance because of conservation of mass: 238 = 234 + 4.

For another example, in the table of the uranium–radium series, ${}^{234}_{90}\mathrm{Th}$ (thorium-234) decays by $\beta$ emission, becoming ${}^{234}_{91}\mathrm{Pa}$ (protactinium-234). Since a $\beta$ particle (electron) has charge $-e$ and has an extremely small mass, the symbol ${}^{\;0}_{-1}e$ is used for it. This $\beta$-decay process may then be represented by the equation

$${}^{234}_{90}\mathrm{Th} \rightarrow {}^{234}_{91}\mathrm{Pa} + {}^{\;0}_{-1}e + {}^{-0}_{\;0}\bar{\nu}.$$

(The ${}^{-0}_{\;0}\bar{\nu}$ 0 is another particle, a so-called antineutrino, given off in $\beta$ decay. It will be discussed in Chapter 18.)

## 17.13 SOME USEFUL APPLICATIONS OF RADIOACTIVITY

The dangers of exposure to radioactivity were noted briefly in Section 17.4. The “fall-out” of radioactive dust and elements from testing of nuclear weapons on or above ground during the 1950s was so harmful to the public that an international treaty to stop such testing was made. For example, strontium-90 is a radioactive isotope that is produced in fission reactions that can be transported into the upper atmosphere by above-ground nuclear explosions. The element strontium is just below calcium on the periodic table. When strontium-90 falls to the ground as fall-out from nuclear testing, it is ingested by cows eating grass and can replace calcium in the formation of milk, thus getting into the food chain where it can damage the internal organs of children and adults who drink the milk.

These processes of radiation damage to biological organisms are the subject of considerable research today, and they have important applications in agriculture, medicine, and other areas. One important area of research involves the study to find out, for example, how radiation produces genetic changes. Since it has been discovered that many of the key chemical processes in cells are organized by single chains of molecules, including DNA, it is clear that a single particle of radiation can, by breaking a chemical bond in such a chain, cause a permanent and perhaps disastrous change in the cell.

The metabolism of plants and animals is being studied with the aid of extremely small amounts of radioactive nuclides called *isotopic tracers*. A radioactive isotope, for example, ${}^{14}\mathrm{C}$, acts chemically (and therefore physiologically) like a stable isotope (${}^{12}\mathrm{C}$). Thus, by following a radioactive tracer with detectors, the behavior of a chemical material can be followed as it goes through various metabolic processes. The role of micronutrients (elements that are essential, in extremely small amounts, for the well-being of plants and animals) can be studied in this way. Agricultural experiments with fertilizers containing radioactive isotopes have shown at what point in the growth of a plant the fertilizer is essential. In chemistry, radioactive isotopes help in the determination of the details of chemical reactions and of the structure of complex molecules, such as proteins, vitamins, and enzymes.

Perhaps the most rewarding uses of radioisotopes have been in medical research, diagnosis, and therapy. For example, tracers can help to determine the rate of flow of blood through the heart and to the limbs, thus aiding in the diagnosis of abnormal conditions. Intense doses of radiation can do serious damage to all living cells, but diseased cells are often more easily damaged than normal cells. Radiation can therefore be used to treat some diseases, for example, destroying cancerous tumors. Some parts of the body take up particular elements preferentially. For example, the thyroid gland absorbs iodine easily. Specially prepared radioisotopes of such elements can be administered to the victims of certain diseases, thus supplying desired radiation right at the site of the disease. This method has been used in the treatment of cancer of the thyroid gland, blood diseases, and brain tumors and in the diagnosis of thyroid, liver, and kidney ailments. To destroy a malignancy in a prostate gland, "seeds" containing radioactive materials may be inserted into it.

Some Typical Isotope Applications

| *Isotope* | *Half-life* | *Important uses* |
|---|---|---|
| ${}^{3}_{1}\mathrm{H}$ | 11 yr | Used as a tag in organic substances. |
| ${}^{14}_{6}\mathrm{C}$ | 5730 yr | Used as a tag in studying the synthesis of many organic substances. When ${}^{14}_{6}\mathrm{C}$ is incorporated in food material, its presence can be traced in the metabolic products. |
| ${}^{24}_{11}\mathrm{Na}$ | 15 hr | Useful in a wide variety of biochemical investigations because of its solubility and chemical properties. |
| ${}^{32}_{15}\mathrm{P}$ | 14 days | For the study of bone metabolism, the treatment of blood diseases and the diagnosis of tumors. |
| ${}^{35}_{16}\mathrm{S}$ | 87 days | Has numerous chemical and industrial applications. |
| ${}^{60}_{27}\mathrm{Co}$ | 5.3 yr | Because of its intense $\gamma$ emission, may be used as a low-cost substitute for radium in radiography and therapy. |
| ${}^{131}_{53}\mathrm{I}$ | 8 days | For the study of thyroid metabolism and the treatment of thyroid diseases. |

One of the most well-known applications of radioactivity is in the determination of the age of organic and other materials. The rare isotope of carbon known as carbon-14 is often used for this purpose. This isotope has a half-life of 5730 years. The more abundant isotope of carbon is carbon-12, which is stable. All living things contain the element carbon, and they all absorb carbon from their environment while they are alive. Most of the carbon is in stable form, but a known fraction of it is in the form of carbon-14. When the living thing, such as a muscle or tree, dies, it ceases to absorb any new carbon, while the carbon-14 in its cells undergoes radioactive decay. As time advances, the amount of carbon-14 in the dead tree, or a piece of wood from the tree, decreases owing to its radioactive decay. By comparing the amount of carbon-14 remaining to the amount that is normally found in living trees of that species, scientists can determine its approximate age by referring to the half-life. For instance, if only one-half the original amount of carbon-14 remains, then the tree is about 5730 yr old.

This was the method used to determine the age of the so-called "Ice Man" in Europe, a prehistoric man who was found after having been frozen for centuries in the ice of a glacier in the Alps. The glacier eventually melted, revealing the dead man's body. Scientists determined that the amount of carbon-14 remaining in his body and in the objects found with him was a little more than half of what it would be if he were alive. Thus they placed the time of his death at about 5000 yr ago; it is the oldest preserved human body ever found.

Radioactive carbon-14 has been used with great success to date once-living materials, but it is limited to time spans in the thousands of years. This is because the more half-lives that pass by, the fewer the number of carbon-14 atoms that remain. So the statistical error increases greatly. For much longer time scales, on the order of millions or even billions of years, scientists have turned to the uranium–radium decay series itself. You will notice that the ultimate "ancestor" of this series, uranium-238, has a half-life of about 4.5 billion years. This is also the approximate age of the Earth, so there is still about one-half of the uranium left on Earth compared to the amount there was at Earth's formation some 4.5 billion years ago. (Much of the original uranium may have contributed to the warming of the Earth in the early years, and in fact the total radioactive amount in the Earth has kept it from being much less cold than it otherwise would have been after its formation.) Enough time has passed for the entire uranium–radium series to have become active and so build up a substantial amount of the end product, the stable lead isotope, lead-206. However, lead-208 is the more common isotope of lead, while lead-206 arises only from the uranium–radium decay series. So the amounts of uranium-238 and lead-206 present in ancient rocks in which prehistoric creatures, such as dinosaurs, are found can be used to determine the approximate age of the rocks, hence the approximate age of the dinosaur fossils found in them.

## SOME NEW IDEAS AND CONCEPTS

accelerator
activity
alpha rays
beta rays
carbon dating
daughter element
fluorescence
gamma rays
half-life
isotope
isotopic tracers
nuclide
phosphorescence
radioactive transformation
radioactivity
radionuclide
uranium–radium series

## SOME FURTHER READINGS

E. Curie, *Madame Curie: A Biography*, V. Sheean, transl. (New York: Da Capo Press, 1986).

G. Holton and S.G. Brush, *Physics, The Human Adventure* (Piscataway, NJ: Rutgers University Press, 2001), Chapter 27.

S. Quinn, *Marie Curie: A Life* (New York: Perseus Press, 1996).

T. Trenn, *The Self-Splitting Atom: The History of the Rutherford–Soddy Collaboration* (London: Taylor and Francis, 1977).

### Web Site

Marie Curie: http://www.aip.org/history/curie/contents.htm.

## STUDY GUIDE QUESTIONS

### 17.1 Questions about the Nucleus

1. Why was it difficult to learn about the interior structure of the nucleus?
2. What tools were needed to understand the nucleus?

### 17.2 Becquerel's Discovery

1. Why was Becquerel experimenting with a uranium compound? Describe his experiment.

2. How did uranium compounds have to be treated in order to emit the "Becquerel rays"?
3. What were the properties of the "Becquerel rays"?
4. In what ways were their properties similar to those of X rays?
5. Why couldn't the "Becquerel rays" be X rays?

### 17.3 The Curies Discover Other Radioactive Elements

1. How is the radioactive emission of an element affected by being combined into different chemical compounds?
2. Why did the Curies suspect the existence of another radioactive material in uranium ore, in addition to uranium itself?
3. What was the main difficulty in producing a pure sample of the element radium?
4. Why did the Curies conclude that radioactivity is related to the presence of atoms of uranium or thorium in a sample?
5. Make a list of at least a dozen of the elements mentioned in this section, then find them on the periodic table.

### 17.4 Identifying the Rays

1. List $\alpha$, $\beta$, and $\gamma$ rays in order of penetrating ability.
2. Why is penetrating power inversely related to ionizing power?
3. Why is shielding required when doing research on radioactive materials?
4. What damage can radioactivity cause to living tissue?

### 17.5 The Charge and Mass of the Rays

1. What was the evidence to support the theory that $\beta$ particles are electrons?
2. What observation led to the suggestion that $\alpha$ particles are much more massive than $\beta$ particles?
3. What evidence led to the tentative conclusion that $\alpha$ particles are helium nuclei?
4. What evidence was available to identify the $\gamma$ rays? Why was it inconclusive?

### 17.6 Rutherford's "Mousetrap"

1. What hypothesis were Rutherford and Royds investigating?
2. Briefly describe their experiment.
3. How did Rutherford determine that the gas that appeared in the tube was helium?

### 17.7 Radioactive Transformations

1. Why was radioactive decay believed not to be an ordinary chemical reaction?
2. What was Rutherford's and Soddy's bold theory of radioactive transformation? Why was it considered bold at the time?

3. Give an example of a radioactive transformation. Why is it contrary to the ideas of nineteenth-century chemistry?

### 17.8 Radioactive Decay Series

1. Give at least three reasons for the difficulty of separating decay products.
2. If you start with a sample made entirely of pure uranium-238 atoms, what emission is observed at the start? How will the emission change as time goes on?
3. Look back at the discoveries the Curies made in their study of pitchblende (Section 17.3). How would you account for what they found?
4. What else might the Curies have discovered in pitchblende?
5. Why is it difficult to separate the decay products of uranium?

### 17.9 Decay Rate and Half-Life

1. Why can one not specify the "lifetime" of a sample of radioactive atoms? of a single radioactive atom?
2. A laboratory starts with a certain amount of a radioactive substance. What fraction of the substance will be left unchanged after a period equal to four times its half-life?
3. Another group in the laboratory has a different amount of the same radioactive substance. What fraction of their original amount will be left unchanged after a period equal to four times its half-life?
4. If, after many half-lives, only two atoms of a radioactive substance remain, what will happen during an additional period equal to one half-life of the substance?
5. How does the concept of statistical probability of survival for atoms differ from that for the survival of humans?

### 17.10 The Concept of Isotopes

1. Why was it not necessary, after all, to expand the periodic table to fit in the newly discovered radioactive substances?
2. The symbol for the carbon-12 nuclide is $^{12}_{6}\mathrm{C}$. What is the approximate atomic mass of carbon-12 in atomic mass units? What is its position in the list of elements?
3. On the periodic table the atomic mass of the element carbon is given as 12.011. Why is this mass different from the atomic mass of the carbon-12 nuclide? What does this number say about the other nuclides of carbon?

### 17.11 Transformation Rules

1. By how many units does the mass of an atom change during $\alpha$ decay? during $\beta$ decay?
2. By how many units does the charge of a nucleus change during $\alpha$ decay? during $\beta$ decay?

3. The following $\alpha$, $\beta$, and $\gamma$ decays occur. Using the transformation rules, determine the daughter nuclide in each case:
$^{234}_{92}\mathrm{U} \rightarrow ? + \alpha + \gamma$,
$^{210}_{82}\mathrm{Pb} \rightarrow ? + \beta + \gamma$.

### 17.12 Summary of Notation for Nuclides and Nuclear Reactions

1. Write the complete symbol for the nuclide with atomic mass 194 and atomic number 78. Of which element is it an isotope?
2. Complete the following equation for $\alpha$ decay by replacing the question mark (?). Tell what law or rule you relied on. (X and Y are two different unspecified elements.)
$^{A}_{Z}\mathrm{X} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{?}_{Z-2}\mathrm{Y}$
3. In the same way, complete the following equation for $\beta$ decay by replacing the question mark (?):
$^{A}_{Z}\mathrm{X} \rightarrow {}^{0}_{-1}e + {}^{A}_{?}\mathrm{Y} + {}^{-0}_{0}\overline{\nu}$.

### 17.13 Some Useful Applications of Radioactivity

1. Scientists have discovered the remains of an extinct mammal. They measured the amount of carbon-14 remaining in its bone structure and found that it is only one-fourth the amount in the bones of similar living creatures. How old did they estimate the mammal's remains to be?
2. Why can't the carbon-dating method be used to date the ages of dinosaur fossils?

## DISCOVERY QUESTIONS

1. Radon gas is known to build up in the basements of some homes. Since radon is a strong $\alpha$-emitter, it can cause damage to living tissue. It may be one cause of lung cancer. What types of locations would be likely to have such a problem? From the table of half-lives in the uranium–radium decay series, what would you recommend might be done about this?
2. Suppose that after the Curies had isolated polonium and radium in uranium ore (pitchblende), they had found another radioactive substance. What steps might they take to identify whether it is a new element or a previously known element?
3. In researching the properties of an element, a scientist observes emissions from the element. How could he/she determine if these emissions are from the nucleus or from the electron region of the atom?
4. After performing the research in answer to Question 3 the scientist determines that the emissions are coming from the nucleus. How might he/she identify these emissions?

5. Why does the uranium–radium decay series start with the element with the highest atomic number and end at the element with the lowest atomic number of the series?
6. Soddy's proposal of the existence of isotopes meant that for some elements not all the atoms are identical. Explain why this proposal does not require that the atoms of a given element show differences in chemical behavior.
7. A scientist is researching a series of radioactive decays and obtains an unknown element. Explain how he/she could determine whether the element is a new element and should be given a new place in the periodic table, or is simply an isotope of an already known element.
8. Refer to the table showing the uranium–radium decay series. Look at the parent and daughter nuclei of each of the three types of decays—alpha, beta, and gamma. Write down in your own words the general rules about what happens to the atomic mass and atomic number in each type of decay.
9. A scientist has a lump of pure lead-214 in a container. He determines that the lump contains $6.02 \times 10^{23}$ atoms. The half-life is 26.8 min. After 26.8 min, approximately how many atoms of lead-214 are left? Why is this only approximate?
10. At the National Bureau of Standards, in 1932, 3.8 l of liquid hydrogen was evaporated slowly until only about 1 g remained. This residue allowed the first experimental check on the existence of the "heavy" hydrogen isotope 2H.
    (a) With the help of the kinetic theory of matter, explain why the evaporation should leave in the residue an increased concentration of the isotope of greater atomic mass.
    (b) Why should the evaporation method be especially effective with hydrogen?
11. Supply the missing data indicated by these transformation equations:
    (a) ${}^{212}_{?}\mathrm{Pb} \rightarrow {}^{212}\mathrm{Bi} + ?.$
    (b) ${}^{212}_{?}\mathrm{Bi} \rightarrow ? + {}^{0}_{-1}e.$
    (c) $? \rightarrow {}^{208}\mathrm{Pb} + {}^{4}_{2}\mathrm{He}.$
12. For each part below, select the most appropriate radiation(s): $\alpha$, $\beta$, or $\gamma$:
    (a) most penetrating radiation;
    (b) most easily absorbed by aluminum foil;
    (c) most strongly ionizing radiation;
    (d) may require thick "radiation shields" for protection;
    (e) cannot be deflected by a magnet;
    (f) largest radius of curvature when traveling across a magnetic field;
    (g) highest $q/m$ value;
    (h) important in the Rutherford-Royd "mousetrap" experiment;
    (i) involved in the transmutation of radium to radon;
    (j) involved in the transmutation of bismuth-210 to polonium-210.
13. A Geiger counter shows that the rate of emission of $\beta$ particles from an initially pure sample of a radioactive element decreases to one-half the initial rate in 25 hr.
    (a) What fraction of the original number of radioactive atoms is still unchanged at that time?

(b) What fraction of the original number will have disintegrated in a total of 50 hr?
(c) What assumptions have you made in giving these answers? How might you check them?

### Quantitative

1. A scientist has obtained 100 g of bismuth-210, which has a half-life of 5.0 days. Bismuth-210 decays into polonium-210, which has a half-life of 138.4 days. Bismuth-210 decays into lead-206, which is stable. The scientist leaves the bismuth in its container for 140 days. After that time:
   (a) approximately how much bismuth-210 is left?
   (b) how much polonium-210 is present?
   (c) how much lead-206 is present?