26 The Nucleus of the Atom

KEY IDEA

The mass of an atom is concentrated in its nucleus, which is held together by powerful forces.

PHYSICS AROUND US . . . Enjoying Empty Spaces

It's great to be lying on the beach, lulled by the sound of the surf, soaking up the Sun. There's a great feeling of peacefulness in the wide-open spaces, with no crowding of people or buildings, just wind and air and water. Away from the pressures of school and work, time seems to stand still.

In such a relaxing setting, it's hard to imagine that your own body is also mostly empty space. Every atom has electrons whizzing around with various amounts of energy, but inside the cloud of electrons is almost nothing but empty space. Yet at the very center of each atom, in a tiny volume only about one-hundred-thousandth of the size of the electron cloud, is the nucleus that holds almost all the mass of the atom.

Most of the time, the nucleus stays in the middle of the atom and nothing happens. Electrons collide and change energy, atoms bond with other atoms, and the nuclei just go along for the ride, never changing. But some nuclei do change, and they emit highly energetic particles. In fact, thousands of energetic particles are passing unnoticed through your body every second. Some of those speeding particles are damaging your cells, breaking apart bonds in the molecules that control critical functions of metabolism and cell division.

But don't lose a moment worrying about this ubiquitous background radioactivity. Since the dawn of life, low levels of radioactivity in rocks, soils, oceans, and air have bathed every living thing. This radioactivity is a natural part of the everyday environment on Earth, like the ocean and beach, and our bodies have evolved mechanisms to repair the damage it causes. But radioactivity also reveals much about the inner structure of the atom.

EMPTY SPACE AND ENORMOUS ENERGY

Imagine that you are holding a basketball, while 25 kilometers (about 15 miles) away, a few grains of sand whiz around. And imagine that all of the vast intervening space—enough to contain a fair-sized city—is absolutely empty. In some respects, that's what the inside of an atom is like, though on a much smaller scale. The basketball is the nucleus and the grains of sand represent the electrons. (Remember, however, that both nuclei and electrons display characteristics of both particles and waves.) The atom, with a diameter 100,000 times that of its nucleus, is almost entirely space.

In previous chapters we explored the properties of atoms in terms of their electrons. Chemical reactions, the way a material responds to electricity, and even the very shape and strength of objects depend on the way that electrons in different atoms interact with one another. In terms of our analogy, all of the properties of the atoms that we have studied so far result from interactions that take place 25 kilometers from where the basketball-sized nucleus is sitting. The incredible emptiness of the atom is a key to understanding two important facts about the relation of the atom to its nucleus:

1. **What goes on in the nucleus of an atom has almost nothing to do with the atom's chemistry, and vice versa.** The chemical bonding of an atom's electrons has virtually no effect on what happens to the nucleus. In most situations you can regard the orbiting electrons and the central nucleus as two separate and independent systems.
2. **The energies available in the nucleus are much greater than those available among electrons.** The particles inside the nucleus are tightly locked in. It takes a great deal more energy to pull them out than it does to remove an electron from an atom. When a particle is released from the nucleus, it carries a lot of energy.

The enormous energy we can get from the nucleus follows from the equivalence of mass and energy (which we described in Chapter 12 and discuss in more detail in Chapter 28). This relationship is defined in Einstein's most famous equation, $E = mc^2$. Remember that the constant c, the speed of light, is a very large number (3×10^8 meters per second) and that this large number is squared in Einstein's equation to give an even larger number. Thus even a very small mass is equivalent to a very large energy.

Einstein's equation tells us that a given amount of mass can be converted into a specific amount of energy in any form, and vice versa. This statement is true for any process involving energy—whenever additional energy E is stored in an object, the mass of the object increases. For example, when hydrogen and oxygen combine to form water, the mass of the water molecule is a tiny bit less than the sum of the masses of the original atoms. This missing mass has been converted to energy holding the atoms together in the molecule. Similarly, when an archer draws a bow, the mass of the bow increases by a tiny amount because of the increased elastic potential energy in the bent material.

The change in mass of objects in everyday events such as these is so small that it is customarily ignored and we speak of the various forms of energy without thinking about their mass equivalents. In nuclear reactions, however, we cannot ignore the mass changes. For example, a nuclear reactor can transform fully

20% of the mass of a proton into energy in each reaction by a process we soon will discuss. Thus nuclear reactions can convert significant amounts of mass into energy, while chemical reactions, which involve only relatively small changes in electric potential energy, involve only infinitesimal changes in mass. This difference explains why an atomic bomb, which derives its destructive force from nuclear reactions, is so much more powerful than conventional explosives, which depend on chemical reactions in materials such as TNT.

THE ORGANIZATION OF THE NUCLEUS

As we have seen in Chapter 21, Ernest Rutherford discovered the atomic nucleus by observing how fast-moving particles scatter off gold atoms. In later experiments with even faster atomic "bullets," physicists found that atomic nuclei sometimes break into smaller fragments. Thus, like the atom itself, the nucleus is made up of smaller pieces. The most important particles in the nucleus are the proton and the neutron. Approximately equal in mass, the proton and neutron can be thought of as the primary building blocks of the nucleus (Figure 26-1).

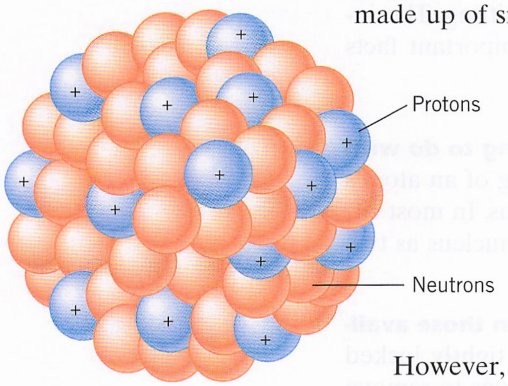


FIGURE 26-1. The massive atomic nucleus incorporates positively charged protons and electrically neutral neutrons.

The **proton** (from Latin for "the first one") has a positive electric charge of +1 and was the first nuclear constituent to be discovered and identified. The number of protons determines the electric charge of the nucleus. An atom in its electrically neutral state has as many negative electrons in orbit as protons in the nucleus. Thus the number of protons in the nucleus determines the chemical identity of an atom.

However, when physicists began studying nuclei, they quickly found that the mass of a nucleus is significantly greater than the sum of the masses of its protons. In fact, for most atoms, the nucleus is more than twice as massive as its protons. What does this observation imply? Physicists concluded that the additional mass had to be in a then undiscovered particle in the nucleus. We now realize that this extra mass is supplied by a particle with no electric charge called the **neutron** (for "the neutral one"). The neutron has approximately the same mass as the proton. Thus a nucleus with equal numbers of protons and neutrons has twice the mass of the protons alone.

The mass of a proton or a neutron is about 2000 times the mass of an electron. Therefore, almost all the mass of the atom is contained within the protons and neutrons in its nucleus. You can think of things this way: electrons give an atom its size, but the nucleus gives an atom its mass.

Element Names and Atomic Numbers

The most important fact in describing any atom is the number of protons in the nucleus—the **atomic number.** This number defines which element you are dealing with. For example, all atoms of gold (atomic number 79) have exactly 79 protons. In fact, for scientists, the name "gold" is simply a convenient shorthand for "atoms with 79 protons." Every element has its own atomic number: all hydrogen atoms have just 1 proton, carbon atoms must have 6 protons, and so on. The periodic table of the elements that we discussed in Chapter 21 can be thought of as a chart in which the number of protons in the atomic nucleus increases as we read from left to right and from top to bottom.

The fixed number of positively charged protons in an atom dictates the number and arrangement of the atom's electrons and thus its chemical properties. In this way, protons define the atom's chemical behavior.

Isotopes and the Mass Number

Each element has a fixed number of protons, but the number of neutrons may vary from atom to atom. In other words, two atoms with the same number of protons may have different numbers of neutrons. Such atoms are said to be **isotopes** of the element, and they have different masses. The total number of protons and neutrons in an atom is called the **mass number.**

The nucleus of every element exists in several different isotopes, each with a different number of neutrons. For example, the most common isotope of carbon has six neutrons, so it has a mass number of 12 (6 protons + 6 neutrons), usually written ^{12}C or carbon-12 and called "carbon twelve." Other isotopes of carbon, such as carbon-13, with seven neutrons, and carbon-14, with eight neutrons, are heavier than carbon-12. However, they have the same electron arrangements and, therefore, the same chemical behavior. A neutral carbon atom, whether carbon-12, carbon-13, or carbon-14, must have six electrons in orbit to balance its six protons.

The complete set of all the isotopes—every known combination of protons and neutrons—is often illustrated on a graph that plots number of protons versus number of neutrons (Figure 26-2). Several features are evident from this graph. First, every chemical element has many known isotopes—in some cases

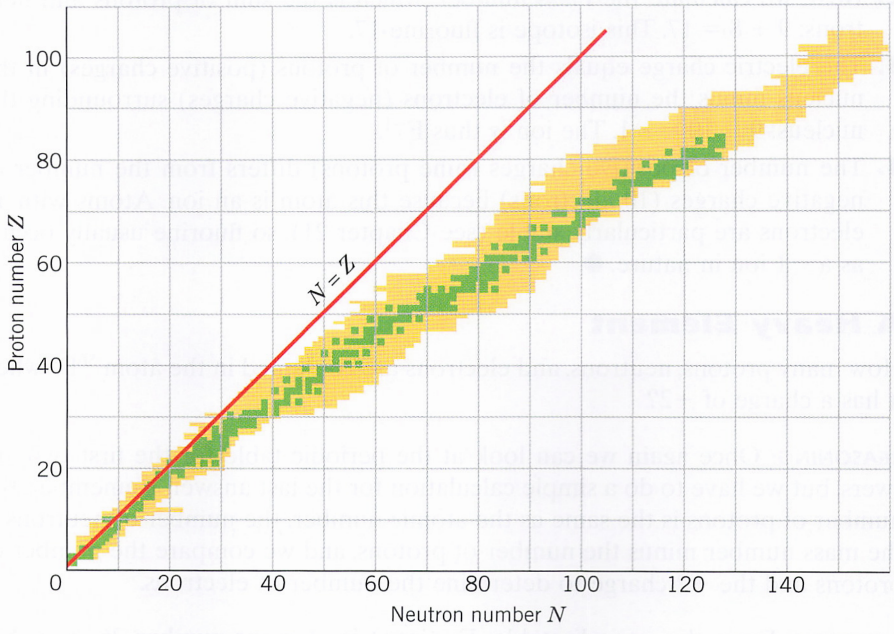


FIGURE 26-2. A chart of the isotopes. Stable isotopes appear in green, and radioactive isotopes are in gold. Each of the approximately 2000 known isotopes has a different combination of protons (*Z*, on the vertical scale) and neutrons (*N*, on the horizontal scale). Isotopes of the light elements (toward the bottom left of the chart) have similar numbers of protons and neutrons and thus lie close to the diagonal $N = Z$ line at 45 degrees. Heavier isotopes (on the upper right part of the chart) tend to have more neutrons than protons and thus lie well below this line.

dozens of them. Close to 2000 isotopes have been documented, compared to the hundred or so different chemical elements. This graph also reveals that for any particular isotope, the number of protons is not generally the same as the number of neutrons. While many light elements, up to about calcium (with 20 protons), often have nearly equal numbers of protons and neutrons, heavier elements tend to have more neutrons than protons. This fact plays a key role in the phenomenon of radioactivity, as we shall see.

EXAMPLE 26-1

Inside the Atom

An atom has nine protons and eight neutrons in its nucleus and 10 electrons in orbit.

1. What element is it?
2. What is its mass number?
3. What is its electric charge?
4. How is it possible that the number of protons and electrons is different?

REASONING: We can find the first three answers by looking at the periodic table (Appendix C). For the last answer, we refer back to Chapter 21 and the discussion of stable electron states.

SOLUTION:

1. The element name depends on the number of protons, which is nine. A glance at the periodic table reveals that element number 9 is fluorine.
2. Next, we calculate the mass number, which is the sum of protons and neutrons: $9 + 8 = 17$. This isotope is fluorine-17.
3. The electric charge equals the number of protons (positive charges) in the nucleus minus the number of electrons (negative charges) surrounding the nucleus: $9 - 10 = -1$. The ion is thus F^{-1}.
4. The number of positive charges (nine protons) differs from the number of negative charges (10 electrons) because this atom is an ion. Atoms with 10 electrons are particularly stable (see Chapter 21), so fluorine usually occurs as a -1 ion in nature. ●

EXAMPLE 26-2

A Heavy Element

How many protons, neutrons, and electrons are contained in the atom ^{56}Fe when it has a charge of $+2$?

REASONING: Once again we can look at the periodic table for the first two answers, but we have to do a simple calculation for the last answer. Remember, the number of protons is the same as the atomic number, the number of neutrons is the mass number minus the number of protons, and we compare the number of protons and the $+2$ charge to determine the number of electrons.

SOLUTION: From the periodic table, Fe (iron) is element number 26, so it has 26 protons.

The number of neutrons is the mass number minus the number of protons: $56 - 26 = 30$ neutrons.

The number of electrons surrounding the nucleus is equal to the number of protons minus the charge on the ion, which in this case is $+2$. Thus there are $26 - 2 = 24$ electrons in orbit. ●

The Strong Force

In Chapter 16 we learned that one of the fundamental laws of electricity is that like charges repel one another. If you think about the structure of the nucleus for a moment, you will realize that the nucleus is made up of a large number of positively charged objects (the protons) in close proximity to one another. Why doesn't the electrical repulsion between the protons push them apart and disrupt the nucleus completely?

The nucleus can be stable only if there is an attractive force capable of balancing or overcoming the electrical repulsion at the incredibly small scale of the nucleus. Whatever the force is, it must be vastly stronger than gravity or electromagnetism, the only two forces we've encountered up to this point. For this reason it is called the **strong force.** The strong force must operate over only very short distances (characteristic of the size of the nucleus) because our everyday experience tells us that the strong force doesn't act on large objects. Both with respect to its magnitude and its range, the strong force is somehow confined to the nucleus. In this respect, the strong force is unlike electricity or magnetism.

The strong force has another distinctive feature. If you weigh a dozen apples and a dozen oranges, their total weight is simply the sum of the individual pieces of fruit. But this is not true of protons and neutrons in the nucleus. The mass of the nucleus is always slightly less than the sum of the masses of the protons and neutrons. When protons and neutrons come together, some of their mass is converted into the energy that binds them together. We know this must be true, because it requires energy to pull most nuclei apart. This *binding energy* varies from one nucleus to another. The iron nucleus is the most tightly bound of all nuclei. This fact is important in the life cycle of stars, which can obtain energy by combining lighter nuclei into heavier ones (a process called *fusion,* discussed later in this chapter) until they form iron. Because iron is the most stable of nuclei, stars cannot obtain energy by combining iron nuclei into something else and so fusion in the star stops.

RADIOACTIVITY

The vast majority of atomic nuclei in objects around you—more than 99.999% of the atoms in our everyday surroundings—are stable. In all probability, the nuclei in those atoms will not change over the lifetime of the universe. But some kinds of atomic nuclei are not stable. Uranium-238, for example, which is the most common isotope of the rather common element uranium, has 92 protons and 146 neutrons in its nucleus. If you put a block of uranium-238 on a table in front of you and watch it for a while, you will find that a few of the uranium nuclei in that block will disintegrate spontaneously. One moment there is a normal uranium atom in the block and the next moment there are fragments of smaller atoms and no uranium atom. At the same time, fast-moving particles speed away from the uranium block into the surrounding environment. This spontaneous release of energetic particles is called **radioactivity** or **radioactive decay.** The emitted particles themselves are referred to as *radiation*. The term *radiation* used in this sense is somewhat different from the electromagnetic radiation that we introduced in Chapter 11. In this case, *radiation* refers to whatever is produced from the spontaneous decay of nuclei, be it electromagnetic waves or actual particles with mass.

What Is Radioactive?

Almost all the atoms around you are stable, but most everyday elements have at least a few isotopes that are radioactive. Carbon, for example, is stable in its most common isotopes carbon-12 and carbon-13; however, carbon-14, which constitutes about one in every 10^{12} carbon atoms in living things, is radioactive. A few elements, such as uranium, radium, and thorium, have no stable isotopes at all. Even though most of our surroundings are composed of stable isotopes, a quick glance at the chart of isotopes (Figure 26-2) reveals that most of the 2000 or so known natural and laboratory-produced isotopes are unstable and undergo radioactive decay of one kind or another.

Physics in the Making

Becquerel and Curie

The nature of radioactivity was discovered in 1896 by Antoine Henri Becquerel (1852–1908), who studied minerals that incorporate uranium and other radioactive elements. He placed some of these samples in a drawer of his desk along with an unexposed photographic plate and a metal coin. When he developed the photographic plate some time later, the silhouette of the coin was clearly visible. From this photograph he concluded that some unknown form of radiation had traveled from the sample to the plate. The coin seemed to have absorbed the radiation and blocked it off, but the radiation that got around the coin and reached the plate delivered enough energy to cause the chemical reactions that normally go into photographic development. Becquerel knew that whatever had exposed the plate must have originated in the minerals and traveled at least as far as the plate.

The Curie family, with Marie Sklodowska, Pierre, and their child, Irene. Both parents received the Nobel Prize in physics in 1903, and Marie was the sole recipient of the 1911 prize in chemistry. Their daughter received the 1935 Nobel Prize with her husband, Frederic Joliot-Curie.

Becquerel's discovery was followed by an extraordinarily exciting time for chemists, who began an intensive effort to isolate and study the elements from which the radiation originated. The leader in the field we now call radiochemistry was also one of the best-known scientists of the modern era, Marie Sklodowska Curie (1867–1934). Born in Poland and married to Pierre Curie, a distinguished French physicist (see Chapter 24), she conducted her pioneering research in France, often under extremely difficult conditions because of the unwillingness of her colleagues to accept her. She worked with tons of exotic uranium-bearing minerals from mines in Bohemia, and she isolated minute quantities of previously unknown elements such as radium and polonium. One of her crowning achievements was the isolation of 22 milligrams of pure radium chloride, which became an international standard for measuring radiation levels. She also pioneered the use of X rays for medical diagnosis during World War I.

For her work, Madame Curie became the first scientist to be awarded two Nobel prizes, one in physics and one in chemistry. One of the most common units used to measure radioactivity is named the curie, in her honor. She also was one of the first scientists to die from prolonged exposure to radiation, whose harmful effects were not known at that time. Her fate, unfortunately, was shared by many other pioneers in nuclear physics. ●

The Kinds of Radioactive Decay

Physicists who studied radioactive rocks and minerals soon discovered three distinct kinds of radioactive decay, each of which changes the nucleus in its own characteristic way and each of which plays an important role in modern science and technology. These three kinds of radioactivity were dubbed *alpha, beta,* and *gamma radiation* (after the first three letters of the Greek alphabet) to indicate that they were unknown and mysterious when first discovered (Figure 26-3).

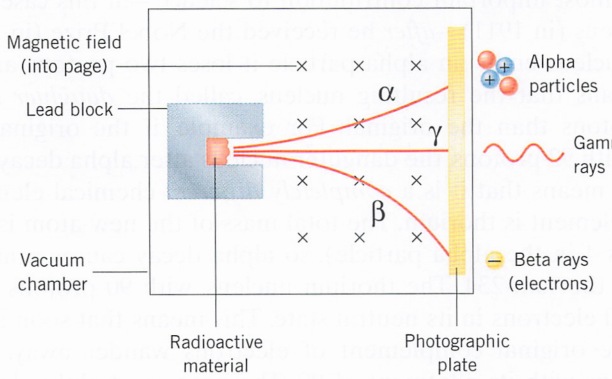


FIGURE 26-3. The three common types of radioactivity are designated alpha, beta, and gamma radiation. Alpha particles have a positive charge and beta rays have a negative charge, so they are deflected in opposite directions by a magnetic field. Gamma rays have no charge and pass through a magnetic field undeflected.

Alpha Decay Some radioactive decays involve the emission of a relatively large and massive particle composed of two protons and two neutrons. Such a particle is exactly the same as the nucleus of a helium-4 atom. It is called an "alpha particle," and the process by which it is emitted is called **alpha decay.** (An alpha particle is usually represented in equations and diagrams by the Greek letter alpha, α.)

Ernest Rutherford, whom we've met as the discoverer of the nucleus, discovered the nature of alpha decay in the first decade of the twentieth century. His simple and clever experiment, sketched in Figure 26-4, began with placing a small amount of radioactive material known to emit alpha particles in a sealed

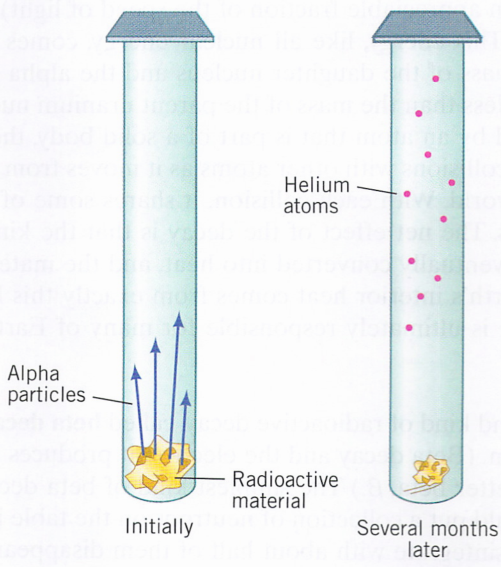


FIGURE 26-4. The Rutherford experiment led to the identification of the alpha particle, which is the same as a helium nucleus.

tube. After several months, careful chemical analysis revealed the presence of a small amount of helium in the tube—helium that hadn't been present when the tube was sealed. From this observation, Rutherford concluded that alpha particles must be associated with the helium atom. Today we say that Rutherford observed the emission of the helium nucleus in radioactive decay, followed by the acquisition of two electrons to form an atom of helium gas.

Rutherford received the Nobel Prize in chemistry for his chemical studies and his work in sorting out radioactivity. He is one of the few people in the world who made his most important contribution to science—in this case, the discovery of the nucleus (in 1911)—*after* he received the Nobel Prize (in 1908).

When a nucleus emits an alpha particle it loses two protons and two neutrons. This means that the resulting nucleus, called the *daughter nucleus,* has two fewer protons than the original. For example, if the original nucleus is uranium-238 with 92 protons, the daughter nucleus after alpha decay has only 90 protons, which means that it is a *completely different* chemical element. In this case, the new element is thorium. The total mass of the new atom is 234 (238 in uranium minus 4 in the alpha particle), so alpha decay causes uranium-238 to transform into thorium-234. The thorium nucleus, with 90 protons, can accommodate only 90 electrons in its neutral state. This means that soon after the decay, two of the original complement of electrons wander away, leaving the daughter nucleus with its allotment of 90. The process of alpha decay reduces the mass and changes the chemical identity of the decaying nucleus.

Radioactivity is nature's philosopher's stone. According to medieval alchemists, the philosopher's stone was supposed to turn lead into gold (among other wonderful properties). The alchemists never found their stone because almost all their work involved what we today call *chemical reactions;* that is, they were trying to change one element into another by manipulating electrons. Given what we now know about the structure of atoms, we realize that they were approaching the problem from the wrong end. If you really want to change one chemical element into another, you have to manipulate the nucleus, precisely what happens in the process of radioactivity.

When the alpha particle leaves the parent nucleus, it typically travels at very high speed (often at an appreciable fraction of the speed of light), so it carries a lot of kinetic energy. This energy, like all nuclear energy, comes from the conversion of mass: the mass of the daughter nucleus and the alpha particle added together is somewhat less than the mass of the parent uranium nucleus. If the alpha particle is emitted by an atom that is part of a solid body, then the particle undergoes a series of collisions with other atoms as it moves from the parent nucleus into the wider world. With each collision, it shares some of its kinetic energy with other atoms. The net effect of the decay is that the kinetic energy of the alpha particle is eventually converted into heat, and the material warms up. About one-half of Earth's interior heat comes from exactly this kind of energy transfer, and this heat is ultimately responsible for many of Earth's major surface features.

Beta Decay The second kind of radioactive decay, called **beta decay,** involves the emission of an electron. (Beta decay and the electron it produces are usually denoted by the Greek letter beta, β.) The simplest kind of beta decay is for a single neutron. If you could put a collection of neutrons on the table in front of you, they would start to disintegrate with about half of them disappearing in the first

10 minutes or so. The most obvious products of this decay are a proton and an electron. Both particles carry an electric charge and are therefore very easy to detect. This production of one positive and one negative particle from a neutral one does not change the total electric charge of the entire system.

In the 1930s, when beta decay of the neutron was first seen in a laboratory, the experimental equipment available at the time easily detected and measured the energies of the electron and proton. Physicists looking at beta decay were troubled to find that the process appeared to violate the law of conservation of energy, as well as conservation of momentum. When they added up the momenta of the proton and electron, the result was less than the momentum of the original neutron. If only the electron and proton were given off, the conservation of momentum would be violated.

Rather than face this possibility, physicists at the time followed the lead of Wolfgang Pauli (see Chapter 6) and postulated that another particle had to be emitted in the decay, a particle that they could not detect at the time, but that carried away the missing momentum. It wasn't until 1956 that physicists were able to detect this missing particle—the *neutrino* ("little neutral one")—in the laboratory. This particle has no electric charge, travels almost as fast as the speed of light, and, if stopped, would have almost no mass. Today, at giant particle accelerators (see Chapter 27), neutrinos are routinely produced in other experiments.

When beta decay takes place inside a nucleus, one of the neutrons in the nucleus is converted into a proton, an electron, and a neutrino. (Note that a neutron does not consist of a proton, electron, and neutrino. These particles are formed only at the moment the decay reaction takes place.) The lightweight electron and the neutrino speed out of the nucleus while the proton remains. The electron that comes off in beta decay is not one of the electrons that originally circled the nucleus in an allowed orbit: the electron emitted from the nucleus comes out so fast that it is long gone from the atom before any of the electrons in orbit have time to react. The new atom has a net positive charge, however, and eventually acquires a stray electron from its environment.

The net effect of beta decay is that the daughter nucleus has approximately the same mass as the parent (it has the same total number of protons plus neutrons), but has one more proton and one less neutron. It is therefore a different element than it was before. For example, carbon-14 (six protons) undergoes beta decay to become an atom of nitrogen-14 (seven protons). If you were to place a small pile of carbon-14 powder—it looks like black soot—in a sealed jar and come back in 20,000 years, most of the powder would have disappeared and the jar would be filled with colorless, odorless nitrogen gas. Beta decay, therefore, is a transformation in which the chemical identity of the atom is changed, but its mass is virtually the same before and after. (Remember, the electron and neutrino that are emitted are extremely lightweight and make almost no difference in the atom's total mass.)

What force in nature could cause an uncharged particle such as the neutron to fly apart? The force is certainly not gravitational attraction between masses. It is not the electromagnetic force that causes oppositely charged particles to fly away from one another. And beta decay seems to be quite different from the strong nuclear force that holds protons together in the nucleus. In fact, beta decay is an example of the operation of the fourth fundamental force in nature, called the *weak force.* (We discuss the weak force again in Chapter 27.)

A safety officer in protective clothing uses a Geiger counter to examine waste for radioactivity.

Gamma Radiation The third kind of radioactivity, called **gamma radiation,** is different in character from alpha and beta decay. (Gamma decay and gamma radiation are usually denoted by the Greek letter gamma, γ.) *Gamma ray* is simply a generic term for a very energetic photon—electromagnetic radiation. In Chapter 19 we have seen that all electromagnetic radiation comes from the acceleration of charged particles, and that is what happens in gamma radioactivity.

When an electron in an atom shifts from a higher energy level to a lower one, we know that a photon is emitted, typically in the range of visible or ultraviolet light. In just the same way, the particles in a nucleus can shift between different energy levels. These shifts, or nuclear quantum leaps, involve energy differences thousands or millions of times greater than those of orbiting electrons. When particles in a nucleus undergo shifts from higher to lower energy levels, some of the emitted gamma radiation is in the range of X rays, while others are even more energetic.

A nucleus emits gamma rays any time its protons and neutrons reshuffle. Neither the protons nor the neutrons change their identity, so the daughter atom has the same mass, the same isotope number, and the same chemical identity as the parent. Nevertheless, this process produces highly energetic radiation.

Moving Down the Chart of the Isotopes

The three kinds of radioactivity, summarized in Table 26-1, affect the nuclei of isotopes. Gamma radiation changes the energy of the nucleus without altering the number of protons and neutrons—the isotope doesn't change into a different element. Alpha and beta radiation result in a new element because they change the numbers of protons and neutrons. Think about how alpha and beta radiation might be represented on the chart of isotopes (Figure 26-2). Alpha decay removes two protons and two neutrons, so we shift diagonally down and to the left by two units on the chart. Alpha decay thus provides a way to move from heavier to lighter elements.

Beta decay, on the other hand, results in one less neutron and one more proton. On the chart of the isotopes, we move diagonally up and to the left by one unit. Beta decay provides a way to reduce the ratio of neutrons to protons, which is generally greater for heavier elements. Combinations of beta and alpha decays, therefore, provide a means to move from heavier radioactive isotopes to lighter stable isotopes.

TABLE 26-1 Types of Radioactive Decay

Type of Decay	Particle Emitted	Net Change
Alpha	Alpha particle	New element with two fewer protons and two fewer neutrons
Beta	Beta particle	New element with one more proton and one fewer neutron
Gamma	Photon (gamma ray)	Same element, but with less energy

Radiation and Health

Why is nuclear radiation so dangerous? Alpha, beta, and gamma radiation all carry a great deal of energy—enough energy to damage the molecules that are essential to the workings of the cells that make up your body. The most common type of damage is *ionization,* the stripping away of one or more of an atom's electrons. An ionized atom cannot bond in its normal way, and any structures that depend on that atom—a cell wall or a piece of genetic material, for example—will be damaged. Prolonged exposure to ionizing radiation can so disrupt an organism's cells that it dies. You can think of the damage caused by the different kinds of radiation as being analogous to the damage different sorts of vehicles would inflict in moving through a crowded alley (Figure 26-5). The alpha particle is like a truck that demolishes everything in its path but doesn't get very far. Beta radiation is like a car that does less damage but goes farther, while gamma radiation penetrates the farthest.

It takes a great deal of radiation to cause sickness or death. Only in unusual circumstances, such as the aftermath of nuclear weapons used on the Japanese cities of Hiroshima and Nagasaki at the end of World War II or the nuclear reactor accident at Chernobyl in Ukraine in 1986, do people die shortly after exposure. However, there are possible long-term effects of exposure to lower levels of radiation. While such exposure does not cause immediate death, it may affect

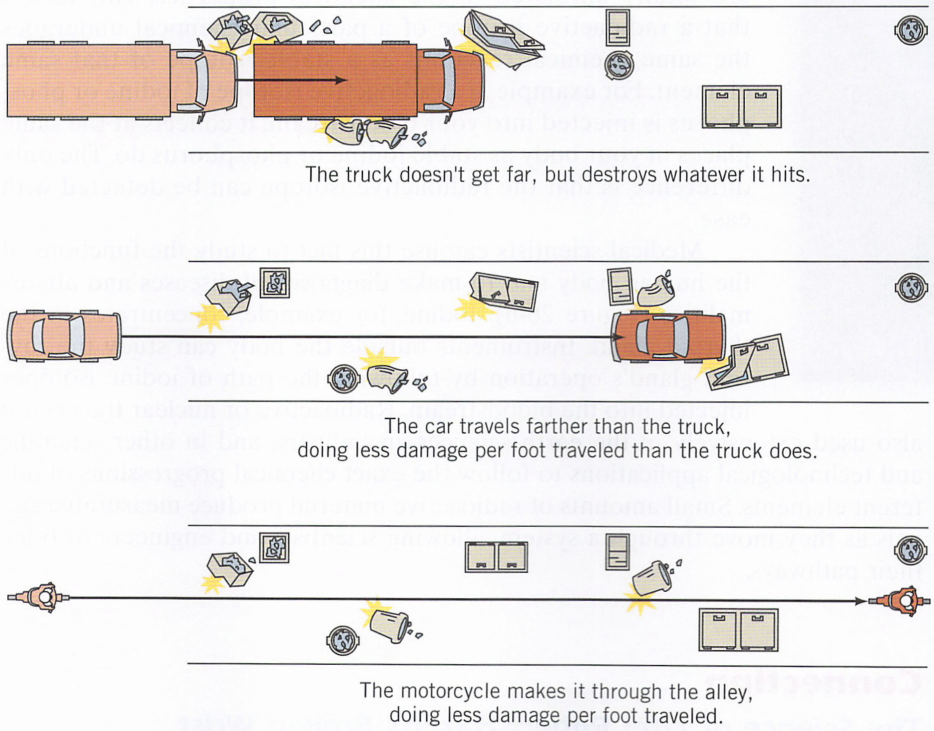


FIGURE 26-5. The damage to atoms and molecules from different kinds of radiation can be compared to the damage to objects in an alleyway caused by different types of vehicles. The truck is analogous to an alpha particle, the car is analogous to a beta particle, and the motorcycle is analogous to a gamma ray. Although you might conclude that the gamma ray does the least amount of damage, its high energy and ability to penetrate deeply makes it especially dangerous in large quantities.

a body's ability to replace old cells and can increase the risk of cancer. Radiation may also damage the genetic material that carries the coded information required to reproduce. Birth defects may not appear until decades after parental exposure to radiation.

No one can escape high-energy radiation entirely. Radioactive decay of uranium and other elements in rocks and soils, as well as cosmic radiation from space, bombard us all the time. Radioactive elements even occur in our bones and tissues. Certain kinds of radiation—medical X rays, for example, and radioactive isotopic tracers used for diagnosis—have saved countless lives. Human beings have always lived in a radioactive environment. Nevertheless, no matter how small the added risk, it is wise to minimize unnecessary exposure to sources of radioactivity.

APPLICATIONS OF RADIOACTIVITY

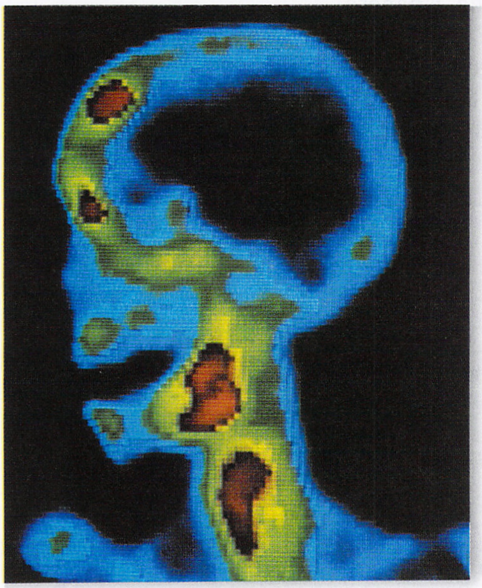


FIGURE 26-6. Radioactive tracers at work. The patient has been given a radioactive tracer that concentrates in the bone and emits radiation that can be measured on a film. The dark spot in the front part of the skull indicates the presence of a bone cancer.

The Use of Radioactive Tracers

Today, numerous radioactive materials find special uses in medicine and industry because all radioactive isotopes are also chemical elements. The chemistry of atoms is governed by their electrons, while the radioactive properties of a material are totally unrelated to the chemical properties. This means that a radioactive isotope of a particular chemical undergoes the same chemical reactions as a stable isotope of that same element. For example, if a radioactive isotope of iodine or phosphorus is injected into your bloodstream, it collects at the same places in your body as stable iodine or phosphorus do. The only difference is that the radioactive isotope can be detected with ease.

Medical scientists can use this fact to study the functions of the human body and to make diagnoses of diseases and abnormalities (Figure 26-6). Iodine, for example, concentrates in the thyroid gland. Instruments outside the body can study the thyroid gland's operation by following the path of iodine isotopes injected into the bloodstream. Radioactive or nuclear tracers are also used extensively in the earth sciences, in industry, and in other scientific and technological applications to follow the exact chemical progressions of different elements. Small amounts of radioactive material produce measurable signals as they move through a system, allowing scientists and engineers to trace their pathways.

Connection

The Science of Life: Robert Hazen's Broken Wrist

I once had an experience that gave me a whole new perspective on radioactivity. A decade ago, while playing beach volleyball, I dove for a ball and bent back my wrist. It hurt a lot, but it was early in the season, so I taped up the wrist and kept on playing. After a couple of weeks it didn't hurt too much, so I forgot about the injury.

Years later, when the wrist started hurting again, I went to a doctor, who said, "Your wrist has been broken for a long time. When did it happen?" Because the break was so old, my doctor had to find out whether the broken bone surfaces were still able to mend. They sent me to a specialized hospital facility where I was given a shot of a fluid containing a radioactive phosphorus compound—a compound that accumulates on active growth surfaces of bone. After a few minutes, this material circulated through my body and some of the phosphorus compound concentrated on unset regions of my wrist bones. Radioactive molecules constantly emitted particles that moved through my skin to the outside; as I lay on a table, an image of my broken wrist glowed on the overhead monitor. The process produced a clear picture of the fracture, so my doctor was able to reset the bones. My wrist has healed, and I'm back to playing volleyball. ●

Half-Life

Left to itself, a single nucleus of an unstable isotope eventually decays in a spontaneous event. That is, the original nucleus persists up until a specific time, then radioactive decay occurs and from that point on you see only the fragments of the decay.

Watching a single nucleus undergo decay is like watching one kernel in a batch of popcorn. Each kernel pops at a specific time, but not all the kernels pop at the same time. Even though you can't predict when any one kernel will pop, you can predict the average time it takes for a kernel to pop. A collection of radioactive nuclei behaves in a similar way. Some nuclei decay almost as soon as you start watching; others persist for much longer times. The percentage of nuclei that decay in each second after you start watching remains more or less the same.

Physicists use the term **half-life** to describe the average time it takes for one-half of a batch of radioactive isotopes to undergo decay. For example, if you have 100 nuclei at the beginning of your observation and it takes 20 minutes for 50 of them to undergo radioactive decay, then the half-life of that nucleus is 20 minutes. If you were to watch that sample for another 20 minutes, however, not all the nuclei would have decayed. You would find that you had about 25 nuclei at the end. After another 20 minutes you would most likely have 12 or 13, and so on (Figure 26-7).

FIGURE 26-7. The graph shows the percentage of radioactive nuclei left in a sample as the number of half-lives increases.

Saying that a nucleus has a half-life of 1 hour does *not* mean that all the nuclei sit there for 1 hour, at which point they all decay. The nuclei, like the popcorn kernels in our example, decay at different times. The half-life is simply an indication of how long on average it will be before an individual nucleus decays.

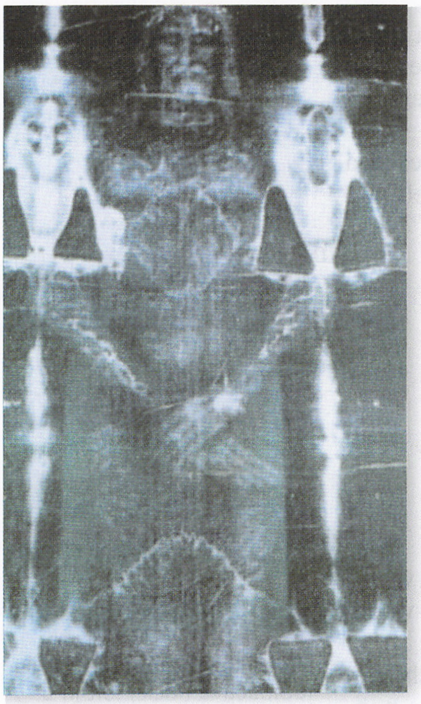
The situation is something like giving identical coins to, say, 10,000 people and seeing who can flip heads the most consecutive times. Everyone left in the game flips a coin every minute. After the first minute, about one-half the people have flipped tails and are out of the game, while the other half have flipped heads and continue in the game. (We assume fair coins for which the probability of landing heads or tails is 50%.) After 2 minutes, one-half of the people left after the first minute are still playing, or one-quarter of the original 10,000 people. Radioactive decay is similar to this game because after every half-life, half of the nuclei that have not decayed are still available to decay in the next half-life.

Radioactive nuclei display a wide range of half-lives. Some nuclei, such as uranium-222, are so unstable that they persist only a tiny fraction of a second. Others, such as uranium-238, have half-lives that range into the billions of years, comparable to the age of Earth. Between these two extremes you can find a radioactive isotope that has almost any half-life you wish.

We do not yet understand enough about the nucleus to be able to predict half-lives. On the other hand, the half-life is a fairly easy number to measure and therefore can be determined experimentally for any nucleus. The fine print on most charts of isotopes (expanded versions of Figure 26-2) usually includes the half-life for each radioactive isotope.

Radiometric Dating

Radioactive decay has provided scientists who study Earth and human history with one of their most important methods of determining the age of materials. This remarkable technique, which depends on our knowledge of the half-life of radioactive materials, is called **radiometric dating.**



The Shroud of Turin, with its ghostly image of a man, was dated by carbon-14 techniques to centuries after the death of Christ.

The best-known radiometric dating scheme involves the isotope carbon-14. Every living organism takes in carbon during its lifetime. At this moment your body is taking the carbon in your food and converting it to tissue, and the same is true of all other animals. Plants are taking in carbon dioxide from the air and doing the same thing. Most of this carbon, about 99%, is in the form of carbon-12, while perhaps 1% is carbon-13. But a certain small percentage, no more than one carbon atom in every trillion, is in the form of carbon-14, a radioactive isotope of carbon with a half-life of about 5700 years.

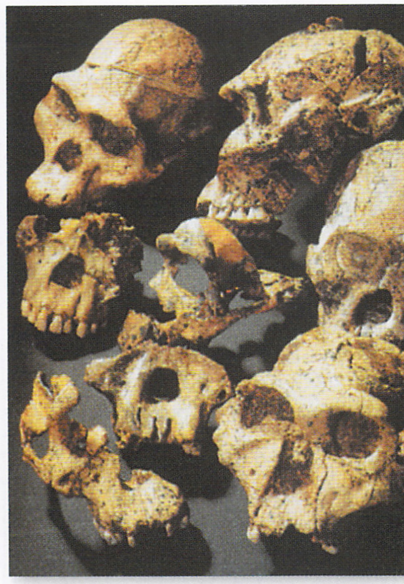
As long as an organism is alive, the carbon-14 in its tissues is constantly renewed in the same small proportion that is found in the general environment. All the isotopes of carbon behave the same way chemically, so the proportions of carbon isotopes in the living tissue are the same everywhere, for all living things. When an organism dies, however, it stops taking in carbon of any form. From the time of death, therefore, the carbon-14 in the tissues is no longer replenished. Like a ticking clock, carbon-14 disappears atom by atom to form an ever-smaller percentage of the total carbon. We can determine the approximate age of a bone, piece of wood, cloth, or other carbon-containing object by carefully measuring the fraction of carbon-14 that remains and comparing it to the amount of carbon-14 that we know must have been in that material when it was alive. If the material happens to be a piece of wood taken from an Egyptian tomb, for example, we have a pretty good estimate of how old the artifact is and, probably, when the tomb was built.

Carbon-14 dating often appears in the news when a reputedly ancient artifact is shown to be from more recent times. In one highly publicized experiment, the Shroud of Turin, a fascinating cloth artifact reputed to be involved in the burial of Jesus, was shown by carbon-14 techniques to date from the fourteenth century A.D.

Develop Your Intuition: Dating Stonehenge

In Chapter 3, we described the monument known as Stonehenge and gave an age for it. This age came from the carbon dating technique we've just described. The monument is made of stone; how do you suppose this dating was done?

You might consider measuring radioactive elements within the stones themselves, perhaps some radioactive isotopes of silicon or oxygen. But that would tell you the age of the stones, not the time at which they were used to form Stonehenge. The actual dating of Stonehenge involved measuring amounts of carbon-14 in some of the artifacts found at the site, including tools made out of animal bone and charcoal from buried campfires. Archeologists have found digging tools made from deer antlers and cattle shoulder bones buried at the bottom of ditches used to help raise the stones. Even a human skeleton and fragments of pottery have been found in holes clearly dug to help maneuver the great stones into place. Radiometric dating of all these objects has given us the date of about 2800 B.C. for the building of the monument.



The oldest human fossils are too ancient to be dated by carbon-14 methods. An alternative technique, potassium-argon dating, is employed for dating the rocks in which these skulls, which are up to 3.7 million years old, were found.

Carbon-14 dating has been instrumental in mapping human history over the last several thousand years. When an object is more than about 50,000 years old, however, the amount of carbon-14 left in it is so small that this dating scheme cannot be used. This means that carbon-14 can only be used to find the age of objects that are relatively young.

In order to date rocks and minerals that are millions of years old, scientists must rely on similar techniques that use radioactive isotopes of much greater half-life. Among the most widely used radiometric clocks in geology are those based on the decay of potassium-40 (half-life of 1.25 billion years), uranium-238 (half-life of 4.5 billion years), and rubidium-87 (half-life of 49 billion years). In these cases, we measure the total number of atoms of a given element, together with the relative percentage of a given isotope, to determine how many radioactive nuclei were present at the beginning.

Connection

Dating a Frozen Mammoth

Russian paleontologists occasionally discover beautifully preserved mammoths frozen in Siberian ice. Carbon isotope analyses from these mammoths often show that only about one-quarter of the original carbon-14 is still present in the mammoth tissues and hair. If the half-life of carbon-14 is 5700 years, how old is the mammoth?

To solve this problem it is necessary to determine how many half-lives have passed, with the predictable decay rate of carbon-14 serving as a clock. In this case, only one-quarter of the original carbon-14 isotopes remain ($\frac{1}{4} = \frac{1}{2} \times \frac{1}{2}$), so the carbon-14 isotopes have passed through two half-lives. After 5700 years, about one-half of the original carbon-14 isotopes remain. Similarly, after another 5700 years only one-half of these remaining carbon-14 isotopes (or one-quarter of the original amount) remain. The age of the preserved mammoth is thus two half-lives, or about 11,400 years. ●

DECAY CHAINS

When a parent nucleus decays, the daughter nucleus is not necessarily stable (it is radioactive). In fact, in the great majority of cases the daughter nucleus is as unstable as the parent. The original parent decays into the daughter, the daughter decays into a second daughter, on and on, perhaps for more than a dozen different radioactive events. Even if you start with a pure collection of atoms of the same isotope of the same chemical element, nuclear decay guarantees that eventually you'll have many different chemical species in the sample. A series of decays of this sort is called a "decay chain." The sequence of decays continues until a stable isotope appears. Given enough time, all the atoms of the original element eventually decay into that stable isotope.

To get a sense of a decay chain, consider the case of uranium-238 with a half-life of approximately 4.5 billion years. Uranium-238 decays by alpha emission into thorium-234, another radioactive isotope. In the process, uranium-238 loses two protons and two neutrons. Thorium-234 undergoes beta decay (half-life of 24.1 days) into protactinium-234 (half-life of about 7 hours), which in turn undergoes beta decay to uranium-234. Each of these beta decays results in the conversion of a neutron into a proton and an electron. After three radioactive decays, we are back to uranium, albeit a lighter isotope with a 247,000-year half-life.

The rest of the uranium decay chain is shown in Figure 26-8. It follows a long path through eight different elements before it winds up as stable lead-206. Given

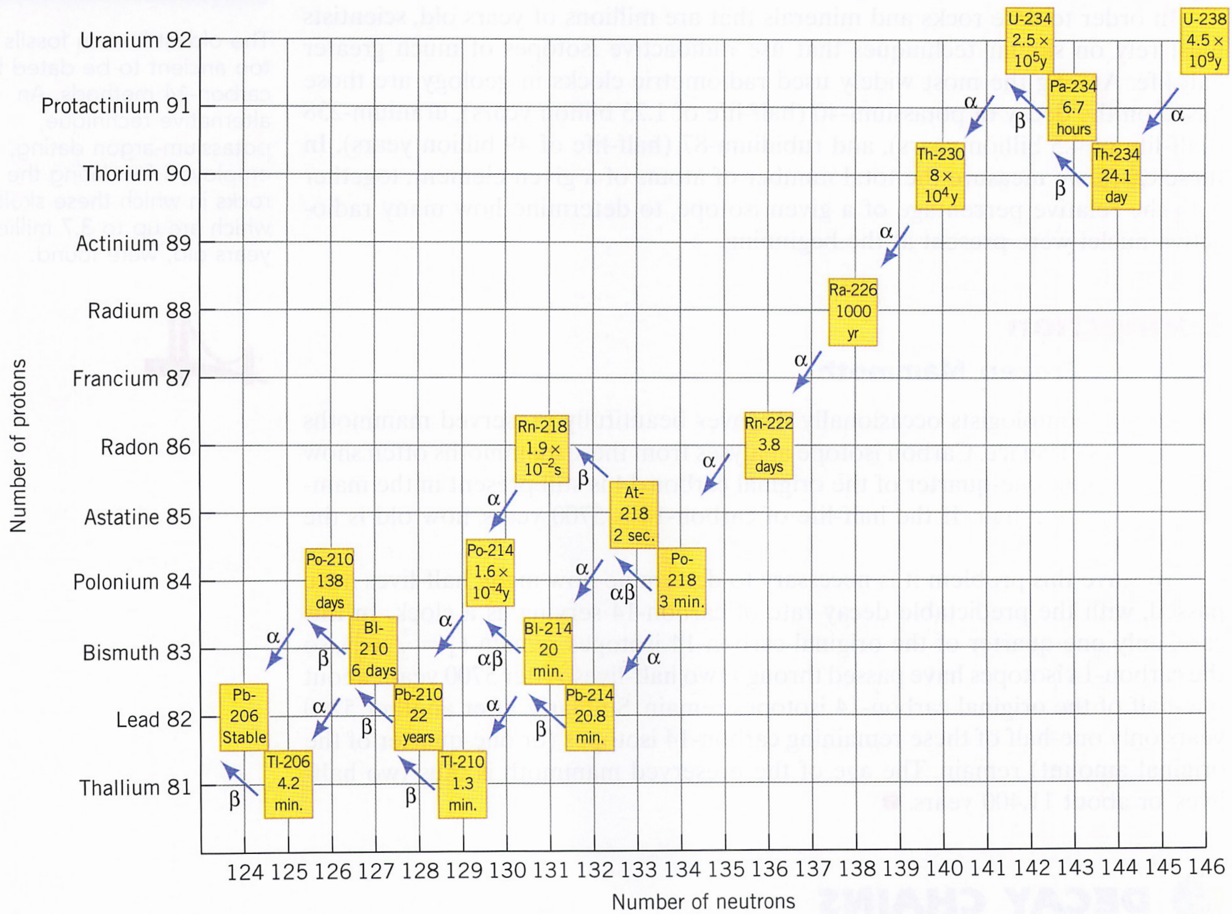


FIGURE 26-8. The uranium-238 decay chain. The nuclei in the chain decay by both alpha and beta emission until they reach lead-206, a stable isotope. Some isotopes may undergo either alpha or beta decay, as indicated by splits in the chain. Nevertheless, all paths eventually arrive at lead-206 after 14 decay events.

enough time, all the uranium-238 now on Earth will eventually decay into lead-206. Earth is only about 4.5 billion years old, however, so there's only been time for about one-half of the original uranium to decay. For the next several billion years we can expect to have all the members of the uranium decay chain in existence on Earth.

Indoor Radon

The uranium-238 decay chain is not an abstract concept, of interest only to theoretical physicists. In fact, the widely publicized health concerns over indoor radon pollution is a direct consequence of the uranium decay chain. Uranium is a fairly common element—about 2 grams out of every ton of rocks at Earth's surface are uranium. The first steps in the uranium-238 decay chain produce thorium, radium, and other elements that remain sealed in ordinary rocks and soils. The principal health concern arises from the production of radon-222, about halfway along the path to stable lead. Radon is a colorless, odorless, inert gas that does not chemically bond to its host rock.

As radon is formed, it seeps out of its mineral host and moves into the atmosphere, where it undergoes alpha decay (half-life of about 4 days) into polonium-218 and a dangerous sequence of short-lived, highly radioactive isotopes. Historically, winds and weather quickly dispersed radon atoms and they posed no serious threat to human health. However, in our modern age of well-insulated, tightly sealed buildings, radon gas can seep in and build up, occasionally to hundreds of times normal levels, in poorly ventilated basements. Exposure to such high radon levels is dangerous because each radon atom undergoes at least five more radioactive decay events in just a few days.

The solution to the radon problem is relatively simple. First, any basement or other sealed-off room should be tested for radon. Simple test kits are available at your local hardware store. If high levels of radon are detected, then the area's ventilation should be improved.

ENERGY FROM THE NUCLEUS

Most scientists who worked on understanding the nucleus and its decays were involved in basic research (see Chapter 1). They were interested in acquiring knowledge for its own sake. But, as frequently happens, knowledge pursued for its own sake is quickly turned to practical use. This certainly happened with the physics of the nucleus.

The atomic nucleus holds vast amounts of energy. One of the defining achievements of the twentieth century was the understanding of and ability to harness that energy. Two very different nuclear processes can be exploited in our search for energy—processes called nuclear fission and nuclear fusion.

Nuclear Fission

Fission means "splitting" and nuclear fission means the "splitting of a nucleus." In most cases, energy is required to tear apart a nucleus. However, some heavy isotopes have nuclei that can be split apart into products that have less mass than the original nucleus. From such nuclei, energy can be obtained from the mass difference.

The most common nucleus from which energy is obtained by fission is uranium-235, an isotope of uranium that constitutes about 7 of every 1000 uranium atoms in the world. If a neutron hits uranium-235, the nucleus splits into two roughly equal-sized large pieces and several smaller fragments. Among these fragments are two or three more neutrons. If these neutrons go on to hit other uranium-235 nuclei, the process is repeated and a *chain reaction* can begin, with each split nucleus producing the neutrons that cause more splittings. By this basic process, uranium can produce large amounts of energy.

The device that allows us to extract energy from controlled nuclear fission is called a **nuclear reactor** (Figure 26-9). The uranium in most reactors contains primarily uranium-238, but it has been processed so it contains much more uranium-235 than it would if it were found in nature. This uranium has been processed into long fuel rods, about the thickness of a lead pencil, surrounded by a metallic protector. Typical reactors incorporate many thousands of fuel rods. Between the fuel rods is a substance called a "moderator," usually water, whose function is to slow down the neutrons that leave the rods.

The nuclear reactor works like this: A neutron strikes a uranium-235 nucleus in one fuel rod, causing that nucleus to split apart. These fragments include several fast-moving neutrons. Fast neutrons are very inefficient at producing fission, but as the neutrons move through the moderator they slow down. In this way, they can initiate other fission events in other uranium-235 nuclei. A chain reaction in a reactor proceeds as neutrons cascade from one fuel rod to another. In the process, the energy released by the conversion of matter goes into heating the fuel rods and the water. The hot water is pumped to another location in the nuclear plant, where it is used to produce steam.

The steam is used to run a generator to produce electricity, as described in Chapter 17 (see Figure 17-6). In fact, the only significant difference between a nuclear reactor and a coal-fired generating plant is the way in which steam is made. In a nuclear reactor, the energy to produce steam comes from the con-

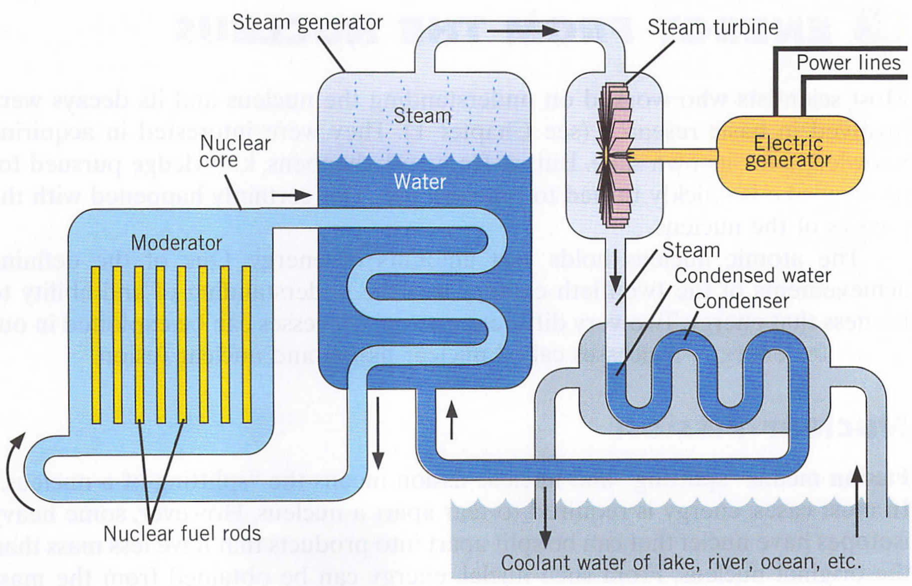


FIGURE 26-9. A nuclear reactor, shown here schematically, produces heat that converts water to steam. The steam powers a turbine, just as in a conventional coal-burning plant.

The nuclear power plant at Three Mile Island, near Harrisburg, Pennsylvania, had to shut down after suffering a partial meltdown. Safety measures ensured that only a very small amount of radioactive material was released into the environment.

version of mass in uranium nuclei; in a conventional generating plant it comes from the burning of coal.

Nuclear reactors must keep a tremendous amount of nuclear potential energy under control while confining dangerously radioactive material. Modern reactors are thus designed with numerous safety features. For example, the water that is in contact with the uranium is sealed in a self-contained system and does not touch the rest of the reactor. Another built-in safety feature is that nuclear reactors cannot function without the presence of the moderator. If there should be an accident in which the water evaporated from the reactor vessel, the chain reaction would shut off. Thus a reactor cannot explode and is *not* analogous to the explosion of an atomic bomb (see Connection: Nuclear Weapons, page 574).

The most serious accident that can occur at a nuclear reactor involves processes in which the flow of water to the fuel rods is interrupted. When this happens, the enormous heat stored in the central part of the reactor can cause the fuel rods to melt. Such an event is called a *meltdown.* In 1979, a nuclear reactor at Three Mile Island in Pennsylvania suffered a partial meltdown, but released only a very small amount of radioactive material to the environment. In 1986, a less carefully designed reactor at Chernobyl, Ukraine, underwent a meltdown accompanied by large releases of radioactivity.

Fusion

Fusion refers to a process in which two nuclei come together (fuse) to form a third, larger nucleus. Under special circumstances it is possible to push two nuclei together and make them fuse. When elements with low atomic numbers fuse, the mass of the final nucleus is less than the mass of its constituent parts. In these cases, it's possible to extract energy from the fusion reaction by conversion of that "missing" mass.

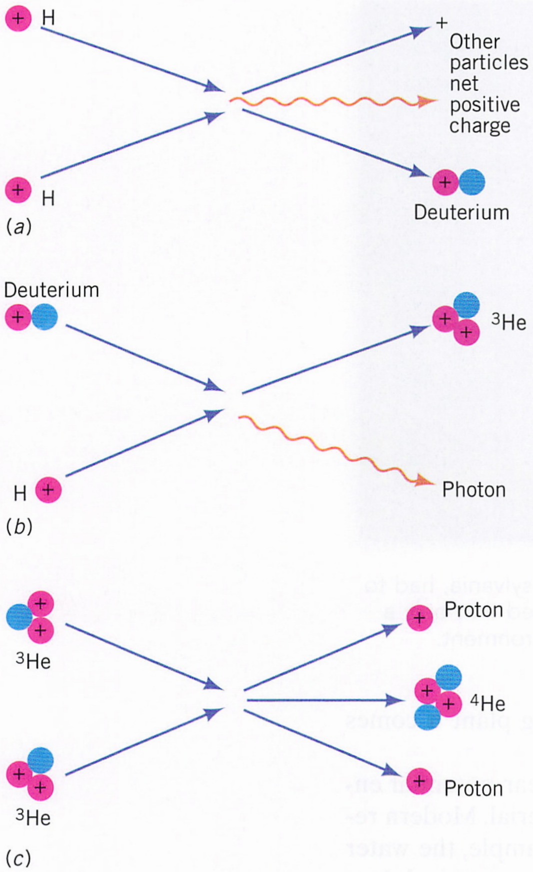


FIGURE 26-10. A fusion reaction releases energy as nuclei combine. Hydrogen nuclei enter into a multistep process. (*a*) First two protons combine to form a deuterium nucleus. (*b*) Then two deuterium nuclei combine to form helium-3. (*c*) Finally, two helium-3 nuclei combine to form helium-4. The red balls are protons; the blue ones are neutrons.

The most common fusion reaction combines four hydrogen nuclei to form a helium nucleus (Figure 26-10). (Remember that an ordinary hydrogen nucleus is a single proton with no neutron. Thus we use the terms "hydrogen nucleus" and "proton" interchangeably.) Note that the reaction requires several steps, starting with two protons fusing to form deuterium, which is the isotope of hydrogen that contains one neutron. This is followed by a second collision in which two deuterium nuclei form helium-3, and a final collision produces helium-4. This nuclear reaction powers the Sun and other stars and thus is ultimately responsible for all life on Earth.

You cannot just put hydrogen in a container and expect it to form helium, however. Two positively charged protons must collide with tremendous speed in order to overcome their electrostatic repulsion and allow the strong force to kick in. (Remember, the strong force operates only over extremely short distances.) In the Sun, high pressures and temperatures in the star's interior trigger the fusion reaction. The sunlight falling outside your window is generated by the conversion of 600 million tons of hydrogen into helium each second. The helium nucleus has a mass about one-half percent less than the original hydrogen nuclei had. The "missing" mass is converted into the energy that eventually radiates out into space.

Connection
Fusion in Stars

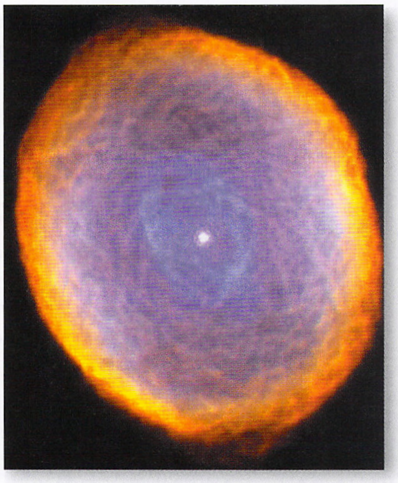
Stars are incredibly massive compared to Earth and the other planets that form our solar system. In fact, over 99.9% of all the mass in the solar system is located in the Sun. So even this high rate of converting hydrogen to helium (600 million tons per second) will continue for literally billions of years. However, all the hydrogen in the Sun's core eventually will get used up, leaving only helium in its place. What will happen then?

The high pressure in the Sun's interior is due to the gravitational attraction of its own mass, as with any star. The inward pressure is kept in check by the heat generated by fusion, which acts to expand the Sun's gas outward. When all the hydrogen is used up, the gravitational force will squeeze the Sun's interior further, generating greater pressure and greater temperature to the point that helium nuclei begin to fuse and form carbon atoms.

In a star such as the Sun, fusion never goes beyond the stage of helium fusing to carbon. The carbon core becomes extremely dense, but even the high pressure and temperature (about 300 million kelvins) are not enough for carbon nuclei to fuse. Eventually, the nuclear fusion reactions cease and the star collapses down to a small hot object called a "white dwarf star." Astronomers expect this is how the Sun will eventually end.

However, in more massive stars than the Sun, the greater mass does generate the higher pressure and temperature needed for carbon fusion. In fact, a series of fusion reactions take place in concentric shells within the star, with fusion of carbon, oxygen, neon, and so on until an inner core of iron forms. As we mentioned earlier in this chapter, iron has the most stable nucleus of all the elements. The fusion of iron does not produce energy; instead, less energy is produced than you have to put in. So fusion stops even in the largest stars, even at temperatures of several billion kelvins.

When fusion stops in a star, gravitational collapse takes over, the core becomes unstable, and the star's life ends in an epic explosion called a "supernova." During the actual explosion, the incredible energy released (about 10^{43} joules!) does briefly allow fusion of heavy nuclei to take place. Astronomers think that all the elements beyond lithium are ejected into the galaxy in these explosions, and that all the elements beyond iron in the periodic table are formed by fusion reactions during the supernova explosion itself. Astronomers observe the gaseous debris of such explosions in the form of a planetary nebula. ●

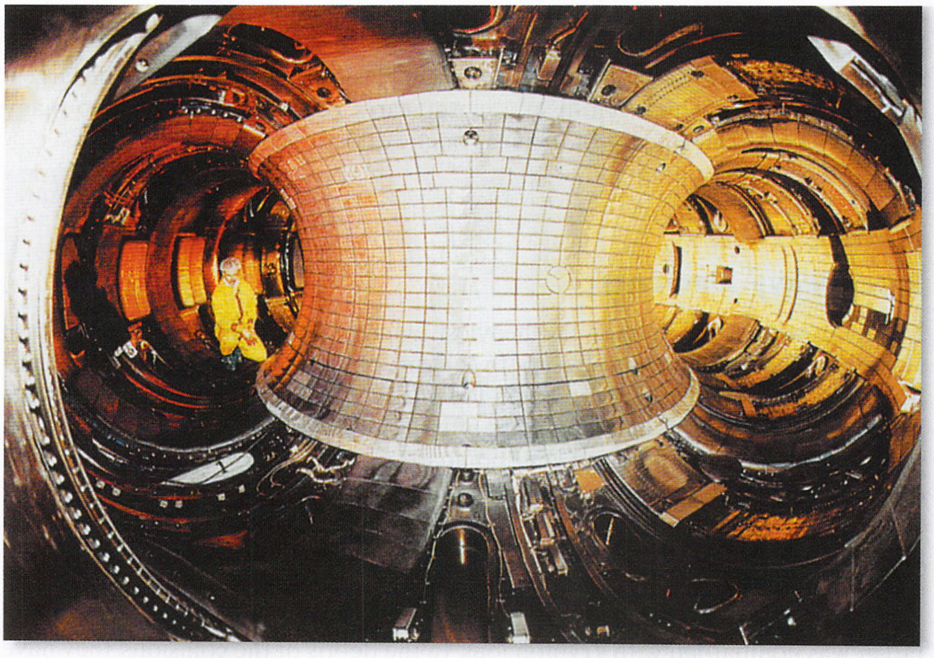


Gaseous debris from exploding stars produce a planetary nebula such as the spectacular Spirograph Nebula.

Since the 1950s, there have been many attempts to harness nuclear fusion reactions to produce energy for human use. The problem has always been that it's very difficult to get protons to collide with enough energy to overcome the electrical repulsion between them and initiate the nuclear reaction.

One promising but technically difficult method is to confine protons in a very strong magnetic field while heating them with high-powered radio waves. Alternatively, powerful lasers have been used to heat pellets of liquid hydrogen. This rapid heating produces shock waves in the liquid, and these shock waves raise the temperature and pressure to the point where fusion is initiated.

Several rather expensive programs are now under way in the United States, Europe, and Japan, where researchers are investigating ways to produce



The Princeton Tokamak is a ring-shaped magnetic chamber. Hydrogen plasma is confined in this ring during attempts to achieve nuclear fusion reactions.

commercially feasible nuclear fusion reactors, perhaps sometime in the twenty-first century. The development is slow and costly, but, if researchers are successful, then the energy crisis will be over forever because there is enough hydrogen in the oceans to power fusion reactions virtually forever.

Connection
Nuclear Weapons

Whenever humans have discovered a new source of energy, that source has quickly been turned into weapons. Nuclear energy, the most concentrated energy source known, is no exception. The earliest atomic bombs, developed in the United States as part of World War II's Manhattan Project, relied on fission reactions in concentrated uranium-235 or plutonium-239. The problem faced by the designers of the atomic bomb was that there has to be a certain number of uranium-235 atoms to sustain a chain reaction to the point where large amounts of energy can be released. This quantity of uranium-235, called the *critical mass,* is about 52 kilograms—a chunk of uranium about the size of a cantaloupe.

At the instant an atomic bomb explodes, a critical mass of the radioactive isotope is compressed into a small volume by a conventional chemical explosive. A fission chain reaction then releases a flood of neutrons, which splits much of the nuclear material and releases a destructive wave of heat, light, and other forms of radiation. The first atomic bombs carried the explosive power of many thousands of tons of chemical explosive.

Shortly after World War II, the United States and the Soviet Union developed the hydrogen bomb, which relies on nuclear fusion. In a hydrogen bomb, fission-bomb triggers surround a compact hydrogen-rich compound. When the atomic fission bombs are set off, they trigger much more intense fusion reactions in the hydrogen. Unlike fission bombs, which are limited by the critical mass of nuclear explosive, hydrogen bombs have no intrinsic limit to their size. The largest nuclear warheads, which can fit in a closet, carry payloads equivalent to many millions of tons (megatons) of conventional explosive. ●

Connection
Superheavy Elements

Uranium, with 92 protons, is the heaviest element commonly found in nature. However, ever since the mid-twentieth century, physicists have been able to build heavier ones in the laboratory. If you look at the periodic table of the elements in Appendix C, all the elements past uranium are seen only in specialized experiments. The technique used is to bombard a heavy nucleus such as lead with other nuclei and hope that, occasionally, enough protons and neutrons will stick together to form the nucleus of a still heavier atom. Although these heavier atoms are unstable and decay quickly, they can last long enough to be identified by their spectra.

Groups in California and Germany have run experiments in which a lead target was bombarded with krypton ions. In the debris of the collisions, they have seen evidence for elements up to 116. Some of these superheavy isotopes have half-lives less than 1 thousandth of a second and are the most massive and complex nuclei ever seen. ●

THINKING MORE ABOUT

The Nucleus: Nuclear Waste

When power is generated in a nuclear reactor, many more nuclear changes take place than those associated with the chain reaction itself. Fast-moving debris from the fission of uranium-235 strikes other nuclei in the system—both the ordinary uranium-238 that makes up most of the fuel rods and the nuclei in the concrete and metal that make up the reactor. In these collisions, the original nuclei may undergo fission or absorb neutrons to become isotopes of other elements. Many of these newly produced isotopes are radioactive. The result is that even when all the uranium-235 has been used to generate energy, a lot of radioactive material remains in the reactor. This sort of material is called *high-level nuclear waste.* (The production of nuclear weapons is another source of this kind of waste.) The half-lives of some of the materials in the waste can run to hundreds of thousands of years. How can we dispose of this waste in a way that keeps it away from living things?

The management of nuclear waste begins with storage. Power companies usually store spent fuel rods at a reactor site for tens of years to allow the short-lived isotopes to decay. At the end of this period the long-lived isotopes that are left behind must be isolated from the environment. Scientists have developed techniques for incorporating these nuclei into stable solids, either minerals or glass. The idea is that the electrons in radioactive isotopes form the same kind of bonds as stable isotopes, so that with a judicious choice of materials radioactive nuclei can be locked into a solid mass for long periods of time.

Plans now call for nuclear waste disposal by the incorporation of radioactive atoms into stable glass that is surrounded by successive layers of steel and concrete. These stable containers are to be buried deep under the Earth's surface in stable rock formations. Ultimately, if a long sequence of public hearings, construction permits, and other hurdles are passed, the United States Department of Energy hopes to confine much of the nation's nuclear waste at the Yucca Mountain repository in a remote desert region of Nevada. The hope is that long-lived wastes can be sequestered from the environment until after they are no longer dangerous to human beings.

The Yucca Mountain project continues to be a controversial subject. Supporters of the site argue that a single, remote long-term site is vastly preferable to the present 131 temporary repositories now located in 39 different states. Such scattered sites are difficult to monitor and protect from terrorist threats. Opponents of Yucca Mountain counter that hauling thousands of tons of nuclear waste on interstate highways poses a far greater danger to the public than the present sites. Some geologists, furthermore, fear that Yucca Mountain may be subject to occasional earthquakes and that its location, less than 100 miles from Las Vegas, is not sufficiently remote. In 2002, Congress voted to proceed with the construction of the waste disposal site at this location.

What should we do with our increasing quantities of nuclear waste? What responsibility do we have to future generations to ensure that the waste we bury stays where we put it? Should the existence of nuclear waste restrain us in our development of nuclear energy? Should we, as some scientists argue, keep nuclear waste materials at the surface and use them for applications such as medical tracers and fuel for reactors?

Summary

The nucleus is a tiny collection of massive particles, including positively charged **protons** and electrically neutral **neutrons.** The nucleus plays a role independent of the orbiting electrons that control chemical reactions, and the energies associated with nuclear reactions are much greater. The number of protons—the **atomic number**—determines the nuclear charge and, therefore, the type of element; each element in the periodic table has a different number of protons. The number of neutrons plus protons—the **mass number**—determines the mass of the **isotope.** Nuclear particles are held together by the **strong force,** which operates only over extremely short distances.

While most of the atoms in objects around us have stable, unchanging nuclei, many isotopes are **radioactive**—they spontaneously change through **radioactive decay.** In **alpha decay,** a nucleus loses two protons and two neutrons. In **beta decay,** a neutron spontaneously transforms into a proton, an electron, and a neutrino. A third kind of radioactivity, involving the emission of energetic electromagnetic radiation, is called **gamma radiation.** The rate of radioactive decay is measured by the **half-life,** which is the time it takes for one-half of a collection of isotopes to decay. Radioactive half-lives provide the key for **radiometric dating** techniques based on carbon-14 and other isotopes. Unstable isotopes are also used as radioactive tracers in medicine and other areas of science. Indoor radon pollution and nuclear waste are two problems that arise from the existence of radioactive decay.

We can produce nuclear energy in two ways. **Fission** reactions, as controlled in **nuclear reactors,** produce energy when heavy radioactive nuclei split apart into fragments that together weigh less than the original isotopes. **Fusion** reactions, on the other hand, combine light elements to make heavier ones, as in the conversion of hydrogen into a smaller mass of helium in the Sun. In each case, the lost nuclear mass is converted into energy.

Key Terms

alpha decay The spontaneous release of an alpha particle (two protons and two neutrons) from an atomic nucleus. (p. 559)

atomic number The number of protons in the nucleus of an atom. (p. 554)

beta decay The spontaneous transformation of a neutron into a proton in an atomic nucleus, accompanied by the release of an electron and a neutrino. (p. 560)

fission The splitting of a nucleus into two or more smaller pieces; usually associated with the splitting of uranium to get energy. (p. 569)

fusion The process in which two nuclei join together to form a larger nucleus. (p. 571)

gamma radiation The spontaneous release of a high-energy photon from an atomic nucleus. (p. 562)

half-life The time it takes for one-half of a sample of radioactive material to decay. (p. 565)

isotope Atoms of the same element whose nuclei have the same number of protons but a different number of neutrons. (p. 555)

mass number The number of protons plus the number of neutrons in the nucleus of an atom. (p. 555)

neutron One of the two particles that make up the atomic nucleus; it has no electric charge. (p. 554)

nuclear reactor A device that allows us to extract energy from nuclear fission in a controlled fashion. (p. 570)

proton One of the two particles that make up the atomic nucleus; it has a charge of +1. (p. 554)

radioactivity or **radioactive decay** The spontaneous release of energetic particles from an atomic nucleus. (p. 557)

radiometric dating The determination of the age of materials using the known half-lives of radioactive elements. (p. 566)

strong force One of the four fundamental forces; it is the force that binds the nucleus together. (p. 557)

Review

1. What was the hypothesis behind Rutherford's experiment on alpha decay? What did he prove?
2. What particles make up the nucleus of an atom? Compare the size of a nucleus to the size of an atom.
3. Which are heavier, electrons or protons?
4. Which is easier to remove from an atom, an electron or a proton?
5. Which is most responsible for the chemistry that an atom undergoes, its nucleus or its electrons? Explain.
6. What determines the electric charge of the nucleus? What determines its mass?
7. What fact about atomic nuclei suggests the existence of the neutron?
8. What is an isotope?
9. What does the atomic number of an isotope represent? The mass number?
10. What is the strong force and how do we know it exists?
11. How is the strong force different from gravity and electromagnetism?
12. What happens to atomic nuclei during radioactive decay?
13. What is alpha decay? How does it change the nucleus?
14. What is beta decay?

15. Why does beta decay not change the total electric charge of an atom?
16. What is a neutrino? What led physicists to hypothesize its existence?
17. What is gamma radiation?
18. How does gamma radiation differ from alpha and beta radiation?
19. Why is radiation dangerous? How does it affect biological tissue?
20. Explain the term "half-life."
21. Explain how radioactivity can be used to date (a) a piece of wood and (b) a rock.
22. What is a decay chain? How do decay chains help explain the dangers presented by radon?
23. How is the principle of conservation of energy seen in (a) fission reactions and (b) fusion reactions?
24. What is nuclear fission? How can we obtain energy from it?
25. What is a chain reaction?
26. How does a nuclear reactor work? How do the fuel rods work, and what is the function of the moderator?
27. What are some of the safety features built into nuclear reactors, and how do they function?
28. "Critical mass" is a term that is widely used outside of nuclear science. What is its everyday meaning, and how does that relate to its scientific meaning?
29. How do fusion reactions produce energy? What is the "missing" mass, and what is its significance?
30. How does the Sun generate energy?
31. Do nuclear power plants use fission or fusion?
32. What are superheavy elements and how are they made? Are they stable?
33. What is a radioactive tracer?
34. How are radioactive tracers useful in medicine? Give an example.
35. What is nuclear waste? Why is it a serious problem for society?

Questions

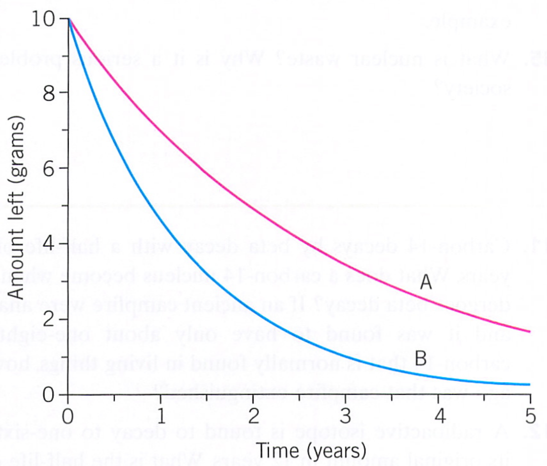
1. Mixing copper and zinc atoms forms the alloy brass. What would form if you fused the nucleus of a copper atom with the nucleus of a zinc atom?
2. Reacting hydrogen and oxygen atoms together produces water, H_2O. What would result from the fusion of two hydrogen nuclei with an oxygen nucleus?
3. An atom has six neutrons, six protons, and six electrons. What element is it? Is it an ion?
4. An atom has 143 neutrons, 92 protons, and 91 electrons. What element is it? Is it an ion?
5. We know that the strong force acts over very short distances. Suppose the range of the strong force were twice what it is now. (In other words, it would attract the same particles with the same force as now, even though they were twice as far away.) Do you think this would have the effect of increasing or decreasing the half-life of uranium-238? Explain.
6. Why does radioactivity seem to be more common with heavier elements?
7. Almost all of the atoms with which we come into daily contact have stable nuclei. Given that most known isotopes are unstable, how could this state of affairs have arisen?
8. Discuss the pros and cons of nuclear power. Refer to specific concepts discussed in this chapter.
9. Can nuclear radiation escape from nuclear power plants? If so, how?
10. Suppose you are a scientist from the future who has discovered the ruins of the Empire State Building. How would you go about estimating the date when it was built?
11. Carbon-14 decays by beta decay with a half-life of 5700 years. What does a carbon-14 nucleus become when it undergoes beta decay? If an ancient campfire were analyzed, and it was found to have only about one-eighth the carbon-14 that is normally found in living things, how long ago was that campfire extinguished?
12. A radioactive isotope is found to decay to one-sixteenth its original amount in 12 years. What is the half-life of this isotope?
13. A radioactive isotope is found to decay to one-eighth its original amount in 30 years. What is the half-life of this isotope?
14. Is it possible for a radioactive nucleus to decay two times and end up as the same element that it started as? Explain. What if the nucleus were restricted to alpha and beta decays only?
15. Hydrogen-3 (also known as tritium) decays by beta decay and has a half-life of about 12 years. What does hydrogen-3 become when it beta decays? If you had a sample of 160 grams of hydrogen-3, how much would still be hydrogen-3 after 36 years?
16. When uranium-238 emits an alpha particle during radioactive decay, what element results?
17. Uranium-235 has a half-life of about 700 million years. What can you say about the likelihood of generating nuclear power from uranium billions of years from now?
18. Suppose you gathered 100,000 of your closest friends to play a game at your local pro-football stadium. Each per-

son has a coin. Every person flips his or her coin once each hour, and everyone that flips tails has to leave. About how many people are left after 1 hour? 2 hours? 3 hours? Can you say exactly how long it will be until everyone is gone?

19. You are given 16 grams of the isotope cobalt-60. Cobalt-60 is radioactive and is used in the process of food irradiation. It has a half-life of 5 years and decays by alpha decay. How long will it take until only 1 gram of cobalt-60 is left? What has the cobalt-60 turned into?

20. If a U-238 nucleus split into two identical pieces and then each piece underwent an alpha decay, what would the two final pieces be?

21. Suppose you are given 10 grams each of two different radioactive isotopes, A and B. You monitor the amount left as a function of time and plot your results (see the figure). What is (approximately) the half-life of each isotope?



22. A 52-kilogram sphere of pure uranium-235 constitutes a critical mass. If this sphere were flattened out into a thin disk, would it still be a critical mass? Explain.

23. Why can't carbon dating be used to estimate the age of dinosaur bones? Why can't it be used to date the rocks that made Stonehenge?

24. What change takes place within the nucleus during beta radiation? If a hydrogen-3 nucleus decays by beta emission, what is the resulting nucleus?

25. Carbon dating is based on the radioactive element carbon-14, which has a half-life of about 5700 years. If the skeletal remains of a human are found to have only one-eighth the amount of carbon-14 that a living human is expected to have, approximately how old is this skeleton?

26. Cesium-137 is radioactive, with a half-life of 30 years, and is used in the food irradiation process. Cesium-137 decays by beta radiation. What does cesium-137 turn into when it decays? How long would it take for 60 grams of cesium-137 to decay to 30 grams?

27. The heaviest naturally occurring element is uranium, atomic number 92. In terms of the forces within the nucleus, why don't heavier elements exist in appreciable amounts?

Problems

1. Use the periodic table to identify the element, atomic number, mass number, and electric charge of each of the following combinations.

a. 6 protons, 7 neutrons, 10 electrons
b. 6 protons, 8 neutrons, 10 electrons
c. 9 protons, 8 neutrons, 10 electrons
d. 9 protons, 8 neutrons, 9 electrons

2. Use the periodic table to determine how many protons and neutrons are in each of the following atoms.

a. C-13 c. Ag-108
b. Ni-56 d. Rn-222

3. The average atomic weight of cobalt atoms (atomic number 27) is actually slightly greater than the average atomic weight of nickel atoms (atomic number 28). How could this situation arise?

4. One atomic mass unit (amu) equals 1.66×10^{-27} kg. Calculate the equivalent of 1 amu in terms of energy in joules.

5. When 1 gram of hydrogen is converted into helium through nuclear fusion, the difference between the masses of the original hydrogen and the final helium is 0.0072 g. How much energy is released when 1 gram of hydrogen undergoes nuclear fusion?

6. Imagine that a collection of 1,000,000 atoms of uranium-238 was sealed in a box at the formation of Earth 4.5 billion years ago. Use the uranium-238 decay chain to predict some of the things you would find if you opened the box today.

7. An ancient archaeological site was being dated using the carbon-14 dating methods on a fire pit. Only 12.5% of the original carbon-14 was detected. How old is this ancient site?

8. Strontium-90 undergoes beta decay with a half-life of 28.8 years. What is the product of this decay? Include in your answer the number of protons (atomic number), the number of neutrons, and the mass number of both strontium-90 and its daughter decay product.

9. Bismuth-214 undergoes beta decay with a half-life of 19.7 minutes. Repeat Problem 8 for the beta decay of this isotope.

10. Thorium-230 undergoes alpha decay with a half-life of about 80,000 years. What is the product of this decay? Include the atomic number and the mass number of both thorium-230 and the product of the alpha decay. Also, if there are 10 grams of thorium-230 to start with, how much of this will remain after 160,000 years has passed? After 240,000 years?

11. Polonium-210 undergoes alpha decay in the uranium decay chain. What is the product of this decay? Include the atomic number and the mass number in your answer.

12. Thorium-232 undergoes three beta decays and one alpha decay in a decay chain. Does the final daughter element depend on the order of the decay modes?

13. If gamma radiation is released from an isotope, how does this change the atomic number of the isotope? The mass number?

14. Isotope X has a half-life of 100 days. A sample is known to have contained about 1,500,000 atoms of isotope X when it was put together, but is now observed to have only about 100,000 atoms of isotope X. Estimate how long ago the sample was assembled. Explain the relevance of this problem to the technique of radiometric dating.

Investigations

1. Read a historical account of the Manhattan Project. What was the principal technical problem in obtaining the nuclear fuel? Why did chemistry play a major role? What techniques are now used to obtain nuclear fuel?

2. What is the current status of U.S. progress toward developing a depository for nuclear waste? How do your representatives in Congress vote on matters relating to this issue?

3. What sorts of isotopes are used for diagnostics in your local hospital? Where are supplies of those radioisotopes purchased? What are the half-lives of the isotopes, and how often are supplies replaced? What is the hospital's policy regarding the disposal of radioactive waste?

4. How much of the electricity in your area comes from nuclear reactors? What fuel do they use? Where are the used fuel rods taken when they are replaced? If the facility offers public tours, visit the reactor and observe the kinds of safety procedures that are used.

5. Obtain a radon test kit from your local hardware store and use it in the basement of two different buildings. How do the values compare? Is either at a dangerous level? If the values differ, what might be the reason?

6. Only about 90 elements occur naturally on Earth, but scientists are able to produce more elements in the laboratory. Investigate the discovery and characteristics of one of these synthetic elements.

7. Soon the United States government will take over responsibility for the nuclear wastes of the 50 states. What options do we have for waste storage? Do you think all the waste should be stored in one place? Should we try to separate and use the radioactive isotopes? What are the factors—social, political, and economic—that will help determine what happens to this nuclear waste? What is the current status of Yucca Mountain as you read this? Do you think the impact of the events of September 11, 2001, changed the debate on this? If so, how?

WWW Resources

See the *Physics Matters* home page at **www.wiley.com/college/trefil** for valuable web links.

1. **http://www.aip.org/history/curie/** An online exhibit, *Marie Curie and the Science of Radioactivity,* by the American Institute of Physics.

2. **http://www.aip.org/history/sakharov/** An online exhibit dedicated to Andrei Sakharov, "father of the Soviet hydrogen bomb," who later became a dissenter and received the Nobel Peace Prize.

3. **http://www.colorado.edu/physics/PhysicsInitiative/Physics2000/isotopes/index.html** A lavishly illustrated and Java-rich (and often slow) site includes a detailed storyline describing isotopes and radioactive decay.

4. **http://www.lbl.gov/abc/index.html** The *ABCs of Nuclear Science,* a lavish site containing tutorials and simulations from Lawrence Berkeley National Laboratory.
5. **http://FusEdWeb.pppl.gov/CPEP/chart.html** *Plasma Physics and Fusion,* another lavish site containing tutorials and simulations from Lawrence Livermore Berkeley National Laboratory.
6. **http://www.xray.hmc.psu.edu/rci/centennial.html** A site indexing the many sites devoted to the celebration of 100 years of medical radiology since the discovery of the X ray in 1895.
7. **http://ippex.pppl.gov/fusion/default.htm** *About Fusion!* A tutorial from the internet plasma physics education experience.