20

NUCLEAR REACTIONS AND THEIR ROLE IN CHEMISTRY

The nebula RCW49, shown in infrared light in this image from NASA's Spitzer Space Telescope, is a nursery for newborn stars. Nuclear fusion reactions involving hydrogen are responsible for the tremendous amounts of energy given off in stars. Reactions of this type are discussed in this chapter. *(Courtesy NASA/JPL-Caltech/University of Wisconsin-madison)*

CHAPTER OUTLINE

THIS CHAPTER IN CONTEXT From the standpoint of chemistry, our interest in the atomic nucleus stems primarily from its role in determining the number and energies of an atom's electrons. This is because it is the electron distribution in an atom that controls chemical properties. Although the nuclei of most isotopes are exceptionally stable, many of the elements of interest to us also have one or more isotopes with unstable nuclei that have unique properties which make them particularly useful in chemistry. These unstable nuclei tend to emit radiation consisting of particles and/or energy. In the pages ahead you will learn about the different kinds of nuclear radiation, how this radiation is detected and measured, and how the properties of unstable nuclei can be applied to practical problems.

One of the benefits that comes from studying nuclear transformations is an understanding of the enormous amounts of energy associated with certain nuclear changes. As you study this chapter you will come to appreciate the origin of the energy given off by stars, including our own sun. You will also learn how nuclear reactors work and how nuclear reactions have been applied to produce nuclear weapons.

20.1 MASS AND ENERGY ARE CONSERVED IN ALL OF THEIR FORMS

Changes involving unstable atomic nuclei generally involve large amounts of energy, amounts that are considerably greater than in chemical reactions. To understand how these energy changes arise, we begin our study by re-examining two physical laws that, until this chapter, have been assumed to be separate and independent, namely, the laws of conservation of energy and conservation of mass. They may be safely treated as distinct for chemical reactions but not for nuclear reactions. These two laws, however, are only different aspects of a deeper, more general law.

As atomic and nuclear physics developed in the early 1900s, physicists realized that the mass of a particle cannot be treated as a constant in all circumstances. The mass, m, of a particle depends on the particle's velocity, v, meaning the velocity relative to the observer. A particle's mass is related to this velocity, and to the velocity of light, c, by the following equation.

□ $c = 3.00 \times 10^8$ m s^{-1}, the speed of light.

$$m = \frac{m_0}{\sqrt{1 - (v/c)^2}} \tag{20.1}$$

Notice what happens when v is zero and the particle has no velocity (relative to the observer). The ratio v/c is then zero, the whole denominator reduces to a value of 1, and Equation 20.1 becomes

$$m = m_0$$

This is why the symbol m_0 stands for the particle's *rest mass*.

Rest mass is what we measure in all lab operations, because any object, like a chemical sample, is either at rest (from our viewpoint) or is not moving extraordinarily rapidly. Only as the particle's velocity approaches the speed of light, c, does the v/c term in Equation 20.1 become important. As v approaches c, the ratio v/c approaches 1, and so $[1 - (v/c)^2]$ gets closer and closer to 0. The whole denominator, in other words, approaches a value of 0. If it actually reached 0, then m, which would be ($m_0 \div 0$), would become infinity. In other words, the mass, m, of the particle moving at the velocity of light would be infinitely great, a physical impossibility. This is why the speed of light is seen as an absolute upper limit on the speed that any particle can approach.

□ Even at $v = 1000$ m s^{-1} (about 2250 mph), the denominator (not rounded) is 0.99999333, or within 7×10^{-4}% of 1.

At the velocities of everyday experience, the mass of anything calculated by Equation 20.1 equals the rest mass to four of five significant figures. The difference cannot be detected by weighing devices. Thus, in all of our normal work, mass appears to be conserved, and the law of conservation of mass functions this way in chemistry.

As a particle's velocity increases, energy is changed to mass

We know that matter cannot appear from nothing, so the extra mass an object acquires as it goes faster must come from the energy supplied to increase the object's velocity. Physicists therefore realized that mass and energy are interconvertible and that in the world of high energy physics, the laws of conservation of mass and conservation of energy are not separate and independent. What emerged was a single, deeper law now called the **law of conservation of mass–energy**.

Law of Conservation of Mass–Energy The sum of all the energy and of all the mass (expressed as an equivalent in energy) in the universe is a constant.

The Einstein equation quantitatively relates rest mass and energy

Albert Einstein, perhaps the most famous physicist of the twentieth century, was able to show that when mass converts to energy, the change in energy, ΔE, is related to the change in rest mass, Δm_0, by the following equation, now called the **Einstein equation.**

TOOLS
Einstein equation

$$\Delta E = \Delta m_0 c^2 \qquad (20.2)$$

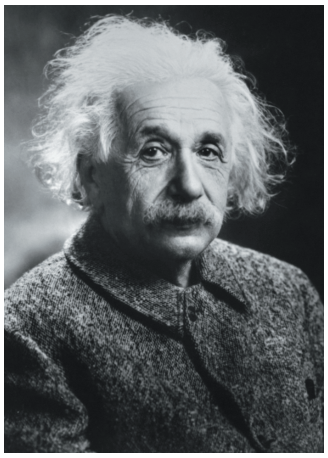
Again, c is the velocity of light, 3.00×10^8 m s^{-1}.

Because the velocity of light is very large, even if an energy change is enormous, the change in mass, Δm_0, is extremely small. For example, the combustion of methane releases considerable heat per mole.

$$\mathrm{CH_4}(g) + 2\mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) + 2\mathrm{H_2O}(l) \qquad \Delta H^\circ = -890\ \mathrm{kJ}$$

The release of 890 kJ of heat energy corresponds to a loss of mass, which by the Einstein equation equals a loss of 9.89 ng. This is about 1×10^{-7}% of the mass of 1 mol of CH_4 and 2 mol of O_2. Such a tiny change in mass is not detectable by laboratory balances, so for all practical purposes, mass is conserved. Although the Einstein equation has no direct applications in chemistry, its importance certainly became clear when atomic fission (the breaking apart of heavy atoms to form lighter fragments) was first observed in 1939.

- The Einstein equation is given as $E = mc^2$ in the popular press.



Albert Einstein (1879–1955); Nobel prize, 1921 (physics). *(Photo Researchers.)*

20.2 THE ENERGY REQUIRED TO BREAK A NUCLEUS INTO SEPARATE NUCLEONS IS CALLED THE NUCLEAR BINDING ENERGY

As we will discuss further in Section 20.3, an atomic nucleus is held together by extremely powerful forces of attraction that are able to overcome the repulsions between protons. To break a nucleus into its individual nucleons therefore requires an enormous input of energy. This energy is called the **nuclear binding energy.**

Absorption of the binding energy would produce the individual nucleons—protons and neutrons—that had made up the nucleus. These nucleons would now carry extra mass corresponding to the mass-equivalent of the energy they had absorbed. If we add up their masses, the sum should be larger than the mass of the nucleus from which they had come. And this is exactly what is observed. *For a given atomic nucleus, the sum of the rest masses of all of its nucleons is always a little larger than the actual mass of the nucleus.* The mass difference is called the **mass defect,** and its energy equivalent is the nuclear binding energy.

Keep in mind that nuclear binding energy is not energy actually possessed by the nucleus but is, instead, the energy the nucleus would have to absorb to break apart. Thus, the *higher* the binding energy, the *more stable* is the nucleus.

The nuclear binding energy can be calculated

We can calculate nuclear binding energy using the Einstein equation. Helium-4, for example, has atomic number 2, so its nucleus consists of 4 nucleons (2 protons and 2 neutrons). The rest mass of one helium-4 nucleus is known to be 4.0015061792 u. However, the sum of the rest masses of its four separated nucleons is slightly more, 4.0318827650 u, which we can show as follows. The rest mass of an isolated proton is 1.0072764669 u and that of a neutron is 1.0086649156 u.

$$\begin{aligned} \text{For 2 protons: } 2 \times 1.0072764669\ \text{u} &= 2.0145529338\ \text{u} \\ \text{For 2 neutrons: } 2 \times 1.0086649156\ \text{u} &= \underline{2.0173298312\ \text{u}} \\ \text{Total rest mass of nucleons in } {}^{4}\text{He} &= 4.0318827650\ \text{u} \end{aligned}$$

The mass defect, the difference between the calculated and measured rest masses for the helium-4 nucleus, is 0.030375858 u, found by

$$\underset{\text{mass of the 4 nucleons}}{4.0318827650\ \text{u}} - \underset{\text{mass of nucleus}}{4.0015061792\ \text{u}} = \underset{\text{mass defect}}{0.0303765858\ \text{u}}$$

Using Einstein's equation, let's calculate to three significant figures the nuclear binding energy that is equivalent to the mass defect for the ${}^{4}\text{He}$ nucleus. We will round the mass defect to 0.0304 u. To obtain the energy in joules, we have to remember that $1\ \text{J} = 1\ \text{kg m}^2\ \text{s}^{-2}$, so the mass in atomic mass units (u) must be converted to kilograms. The table of constants inside the rear cover of the book gives $1\ \text{u} = 1.6605389 \times 10^{-24}$ g, which equals $1.6605389 \times 10^{-27}$ kg. Substituting into the Einstein equation gives

$$\begin{aligned} \Delta E = \Delta mc^2 &= \underbrace{(0.0304\ \cancel{\text{u}}) \times \frac{1.66065389 \times 10^{-27}\ \text{kg}}{1\ \cancel{\text{u}}}}_{\Delta m\ (\text{in kg})} \times \underbrace{(3.00 \times 10^{8}\ \text{m s}^{-1})^2}_{c^2} \\ &= 4.54 \times 10^{-12}\ \text{kg m}^2\ \text{s}^{-2} \\ &= 4.54 \times 10^{-12}\ \text{J} \end{aligned}$$

There are four nucleons in the helium-4 nucleus, so the binding energy *per* nucleon is (4.54×10^{-12} J/4 nucleons) or 1.14×10^{-12} J/nucleon.

The formation of just one nucleus of ${}^{4}\text{He}$ releases 4.54×10^{-12} J. If we could make Avogadro's number or 1 mol of ${}^{4}\text{He}$ nuclei (with a total mass of only 4 g) the net release of energy would be

$$(6.02 \times 10^{23}\ \text{nuclei}) \times (4.54 \times 10^{-12}\ \text{J/nucleus}) = 2.73 \times 10^{12}\ \text{J}$$

This is a huge amount of energy from forming only 4 g of helium. It could keep a 100 watt lightbulb lit for nearly 900 years!

Binding energy per nucleon varies from one element to another

Figure 20.1 shows a plot of binding energies per nucleon versus mass numbers for most of the elements. The curve passes through a maximum at iron-56, which means that the nuclei of iron-56 atoms are the most stable of all. The plot in Figure 20.1, however, does not have a sharp maximum. Thus, a large number of elements with intermediate mass numbers in the broad center of the periodic table include the most stable isotopes in nature.

Nuclei of low mass number have small binding energies per nucleon. Joining two such nuclei to form a heavier nucleus, a process called **nuclear fusion,** leads to a more stable nucleus and a large increase in binding energy per nucleon. This extra energy is released when the two lighter nuclei fuse (join) and is the origin of the energy released in the cores of stars and the detonation of a hydrogen bomb. Nuclear fusion is discussed further in Section 20.8.

As we follow the plot of Figure 20.1 to the highest mass numbers, the nuclei decrease in stability as the binding energies decrease. Among the heaviest atoms, therefore, we might expect to find isotopes that could change to more stable forms by breaking up into lighter nuclei, by undergoing nuclear fission. **Nuclear fission,** also discussed in Section 20.8, is the spontaneous breaking apart of a nucleus to form isotopes of intermediate mass number.

☐ Quite often you'll see the terms *atomic fusion* and *atomic fission* used for nuclear fusion and nuclear fission.

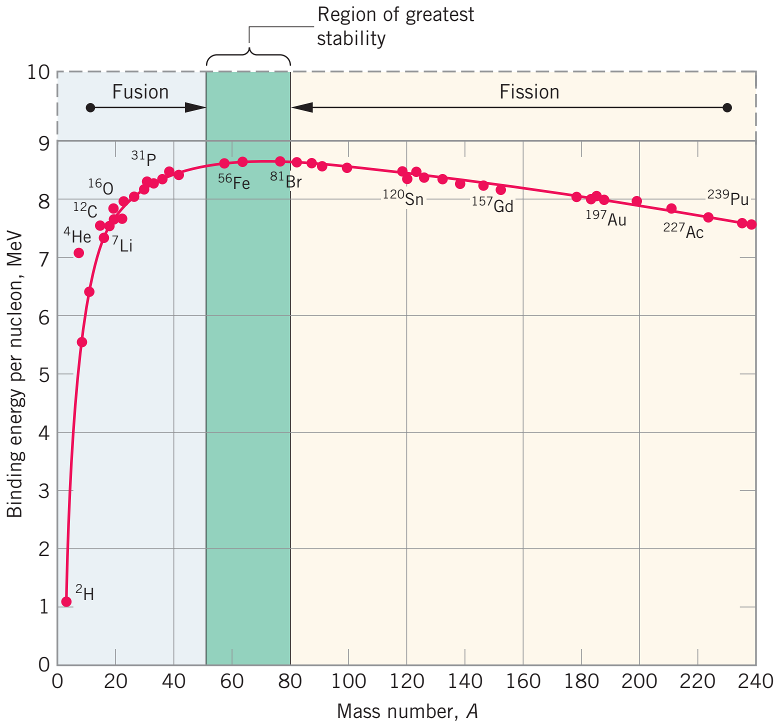


FIG. 20.1 **Binding energies per nucleon.** The energy unit here is the megaelectron volt or MeV: 1 MeV = 10^6 eV = 1.602×10^{-13} J. *(From D. Halliday and R. Resnick, Fundamentals of Physics, 2nd ed., Revised, 1986. John Wiley & Sons, Inc. Used by permission.)*

20.3 RADIOACTIVITY IS AN EMISSION OF PARTICLES AND/OR ELECTROMAGNETIC RADIATION BY UNSTABLE ATOMIC NUCLEI

Except for hydrogen, all atomic nuclei have more than one proton, each of which carries a positive charge. Because like charges repel, we might wonder how any nucleus could be stable. Electrostatic forces of attraction and repulsion, such as the kinds present among the ions in a crystal of sodium chloride, are not the only forces at work in the nucleus, however. Protons do, indeed, repel each other electrostatically, but another force, a force of attraction called the *nuclear strong force,* also acts in the nucleus. The nuclear strong force, effective only at very short distances, overcomes the electrostatic force of repulsion between protons, and it binds both protons and neutrons into a nuclear package. Moreover, the neutrons, by helping to keep the protons farther apart, also lessen repulsions between protons.

☐ Adjacent neutrons experience no electrostatic repulsion between each other, only the strong force (of attraction).

One consequence of the difference between the nuclear strong force and the electrostatic force occurs among nuclei that have large numbers of protons but too few intermingled neutrons to dilute the electrostatic repulsions between protons. Such nuclei are often unstable; their nuclei carry more energy than do other arrangements of the same nucleons. To achieve a lower energy and thus more stability, unstable nuclei have a tendency to eject small nuclear fragments, and many simultaneously release high-energy electromagnetic radiation. The stream of particles (or photons) coming from the sample is called **nuclear radiation** or **atomic radiation** and the phenomenon is called **radioactivity.** Isotopes that exhibit this property are said to be **radioactive** and are called **radionuclides.** About 60 of the approximately 350 naturally occurring isotopes are radioactive.

In a sample of a given radionuclide, not all the atoms undergo change at once. The rate at which radiation is emitted (which translates into the intensity of the radiation) depends on the concentration of the isotope in the sample. Over time, as radioactive nuclei change into stable ones, the number of atoms of the radionuclide remaining in the sample decreases, causing the intensity of the radiation to drop, or *decay*. The radionuclide is said to undergo **radioactive decay.**

Naturally occurring atomic radiation consists principally of three kinds: alpha, beta, and gamma radiation, as discussed below.

Alpha radiation is a stream of helium nuclei

Alpha radiation consists of a stream of helium nuclei called **alpha particles** (α particles), symbolized as $^{4}_{2}\text{He}$, where 4 is the mass number and 2 is the atomic number. The alpha particle bears a charge of 2+, but the charge is omitted from the symbol.

Alpha particles are the most massive of those commonly emitted by radionuclides. When ejected (Figure 20.2), alpha particles move through the atom's electron orbitals, emerging from the atom at speeds of up to one-tenth the speed of light. Their size, however, prevents them from going far. After traveling at most only a few centimeters in air, alpha particles collide with air molecules, lose kinetic energy, pick up electrons, and become neutral helium atoms. Alpha particles cannot penetrate the skin, although enough exposure causes a severe skin burn. If carried in air or on food into the soft tissues of the lungs or the intestinal tract, emitters of alpha particles can cause serious harm, including cancer.

□ In Chapter 2 you learned that an isotope is identified by writing its mass number, *A*, as a superscript and its atomic number, *Z*, as a subscript in front of the chemical symbol, *X*, as in $^{A}_{Z}X$. We use this same notation in representing particles involved in nuclear reactions. For particles that are not atomic nuclei, *Z* stands for the charge on the particle.

Nuclear equations describe the decay of radioactive nuclei

To symbolize the decay of a nucleus, we construct a **nuclear equation**, which we can illustrate by the alpha decay of uranium-238 to thorium-234.

$$^{238}_{92}\text{U} \longrightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He}$$

Unlike chemical reactions, nuclear reactions produce new isotopes, so we need separate rules for balancing nuclear equations.

□ A nuclear transformation such as the alpha decay of ^{238}U does not depend on the chemical environment. The same nuclear equation applies whether the uranium is in the form of the free element or in a compound.

Rules for balancing a nuclear equation

1. The sum of the mass numbers on each side of the arrow must be the same.
2. The sum of the atomic numbers (nuclear charge) on each side of the arrow must be the same.

TOOLS *Nuclear equations*

In the nuclear equation for the decay of uranium-238, the atomic numbers balance (90 + 2 = 92), and the mass numbers balance (234 + 4 = 238). Notice that electrical charges are not shown, even though they are there (initially). The alpha particle, for example, has a charge of 2+. If it was emitted by a neutral uranium atom, then the thorium particle initially has a charge of 2−. These charged particles, however, eventually pick up or lose electrons either from each other or from molecules in the matter through which they travel.

Beta radiation is a stream of electrons

Naturally occurring **beta radiation** consists of a stream of electrons, which in this context are called **beta particles**. In a nuclear equation, the beta particle has the symbol $^{0}_{-1}e$, because the electron's mass number is 0 and its charge is 1−. Hydrogen-3 (tritium) is a beta emitter that decays by the following equation.

$$\underset{\text{tritium}}{^{3}_{1}\text{H}} \longrightarrow \underset{\text{helium-3}}{^{3}_{1}\text{He}} + \underset{\substack{\text{beta particle}\\ \text{(electron)}}}{^{0}_{-1}e} + \underset{\text{antineutrino}}{\bar{\nu}}$$

□ Sometimes the beta particle is given the symbol $^{0}_{-1}\beta$, or simply $^{-}\beta$.

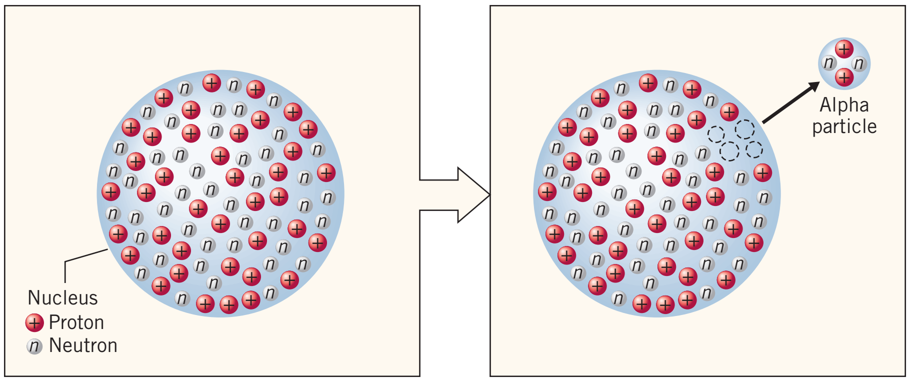


FIG. 20.2 Emission of an alpha particle from an atomic nucleus. Removal of $^{4}_{2}\text{He}$ from a nucleus decreases the atomic number by 2 and the mass number by 4.

Both the antineutrino (to be described further shortly) and the beta particle come from the atom's nucleus, not its electron shells. We do not think of them as having a prior existence in the nucleus, any more than a photon exists before its emission from an excited atom (see Figure 20.3). Both the beta particle and the antineutrino are created during the decay process in which a neutron is transformed into a proton.

$$\underset{\substack{\text{neutron}\\\text{(in the}\\\text{nucleus)}}}{{}^{1}_{0}n} \longrightarrow \underset{\substack{\text{beta particle}\\\text{(emitted)}}}{{}^{0}_{-1}e} + \underset{\substack{\text{proton}\\\text{(remains in}\\\text{the nucleus)}}}{{}^{1}_{1}p} + \underset{\substack{\text{antineutrino}\\\text{(emitted)}}}{\overline{\nu}}$$

Unlike alpha particles, which are all emitted with the same discrete energy from a given radionuclide, beta particles emerge from a given beta emitter with a continuous spectrum of energies. Their energies vary from zero to some characteristic fixed upper limit for each radionuclide. This fact once gave nuclear physicists considerable trouble, partly because it was an apparent violation of energy conservation. To solve this problem, Wolfgang Pauli proposed in 1927 that beta emission is accompanied by emission of yet another decay particle, this one electrically neutral and almost massless. Enrico Fermi suggested the name *neutrino* ("little neutral one"), but eventually it was named the *antineutrino*, symbolized by $\overline{\nu}$.

An electron is extremely small, so a beta particle is less likely to collide with the molecules of anything through which it travels. Depending on its initial kinetic energy, a beta particle can travel up to 300 cm (about 10 ft) in dry air, much farther than alpha particles. Only the highest energy beta particles can penetrate the skin, however.

Gamma radiation is very high energy electromagnetic radiation

□ Gamma photons have the symbol ${}^{0}_{0}\gamma$ because they have no charge or mass.

Gamma radiation, which accompanies most nuclear decays, consists of high-energy photons given the symbol ${}^{0}_{0}\gamma$ or, often, simply γ in equations. Gamma radiation is extremely penetrating and is effectively blocked only by very dense materials, like lead.

The emission of gamma radiation involves transitions between energy levels *within* the nucleus. Nuclei have energy levels of their own, much as atoms have orbital energy levels. When a nucleus emits an alpha or beta particle, it sometimes is left in an excited energy state. By the emission of a gamma-ray photon, the nucleus relaxes into a more stable state.

The electron volt is an energy unit

The energy carried by a given radiation is usually described by an energy unit new to our study, the **electron volt,** abbreviated **eV;** 1 eV is the energy an electron receives when accelerated under the influence of 1 volt. It is related to the joule as follows.

$$1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}$$

□ In the interconversion of mass and energy,

$1\text{eV} \Leftrightarrow 1.783 \times 10^{-36}$ kg

$1\text{MeV} \Leftrightarrow 1.78 \times 10^{-27}$ g

$1\text{ GeV} \Leftrightarrow 1.783 \times 10^{-24}$ g

As you can see, the electron volt is an extremely small amount of energy, so multiples are commonly used, like the kilo-, mega-, and gigaelectron volt.

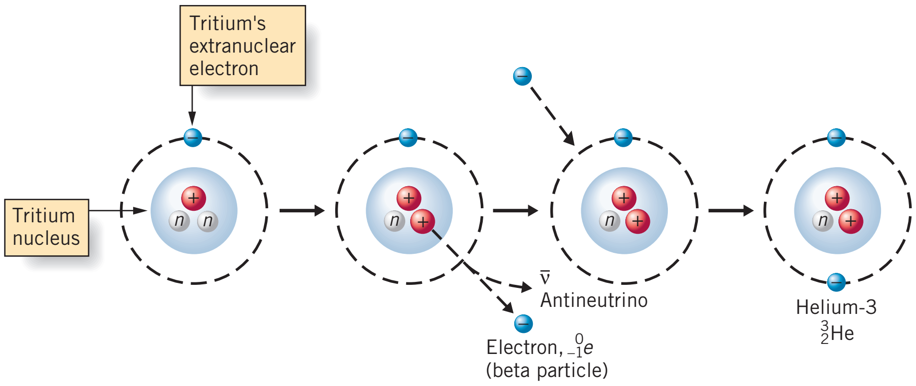


FIG. 20.3 Emission of a beta particle from a tritium nucleus. Emission of a beta particle changes a neutron into a proton. This results in a positively charged ion, which picks up an electron to become a neutral atom.

$$1 \text{ keV} = 10^3 \text{ eV}$$
$$1 \text{ MeV} = 10^6 \text{ eV}$$
$$1 \text{ GeV} = 10^9 \text{ eV}$$

An alpha particle emitted by radium-224 has an energy of 5 MeV. Hydrogen-3 (tritium) emits beta radiation at an energy of 0.05 to 1 MeV. The gamma radiation from cobalt-60, the radiation currently used to kill bacteria and other pests in certain foods, consists of photons with energies of 1.173 MeV and 1.332 MeV.

X rays are high energy electromagnetic radiation

X rays, like gamma rays, consist of high-energy electromagnetic radiation, but their energies are usually less than those of gamma radiation. Although X rays are emitted by some synthetic radionuclides, when needed for medical diagnostic work they are generated by focusing a high-energy electron beam onto a metal target.

☐ X rays used in diagnosis typically have energies of 100 keV or less.

Some radioisotopes emit positrons or neutrons; others capture electrons

Many *synthetic isotopes* (isotopes not found in nature, but synthesized by methods discussed in Section 20.5) emit *positrons*, which are particles with the mass of an electron but a positive instead of a negative charge. A **positron** is a positive beta particle, a positive electron, and its symbol is $^{0}_{1}e$. It forms in the nucleus by the conversion of a proton to a neutron (Figure 20.4). Positron emission, like beta emission, is accompanied by a chargeless and virtually massless particle, a *neutrino* (ν), the counterpart of the antineutrino ($\overline{\nu}$) in the realm of antimatter (defined below). Cobalt-54, for example, is a positron emitter and changes to a stable isotope of iron.

☐ The symbol $^{+}\beta$ is sometimes used for the positron.

$$^{54}_{27}\text{Co} \longrightarrow {}^{54}_{26}\text{Fe} + \underset{\text{positron}}{^{0}_{1}e} + \underset{\text{neutrino}}{\nu}$$

A positron, when emitted, eventually collides with an electron, and the two annihilate each other (Figure 20.5). Their masses change entirely into the energy of two gamma-ray photons called *annihilation radiation photons*, each with an energy of 511 keV.

$$^{0}_{-1}e + {}^{0}_{1}e \longrightarrow 2\,^{0}_{0}\gamma$$

Because a positron destroys a particle of ordinary matter (an electron), it is called a particle of antimatter. To be classified as **antimatter,** a particle must have a counterpart among one of ordinary matter, and the two must annihilate each other when they collide. For example, a neutron and an antineutron represent such a pair and annihilate each other when they come in contact.

Neutron emission, another kind of nuclear reaction, does not lead to an isotope of a different element. Krypton-87, for example, decays as follows to krypton-86.

$$^{87}_{36}\text{Kr} \longrightarrow {}^{86}_{36}\text{Kr} + \underset{\text{neutron}}{^{1}_{0}n}$$

Electron capture, yet another kind of nuclear reaction, is very rare among natural isotopes but common among synthetic radionuclides. For example, an electron can be captured from the orbital electron shell having $n = 1$ or $n = 2$ by a vanadium-50 nucleus, causing it to change to a stable ^{50}Ti nucleus. The transformation is accompanied by the emmision of X rays and a neutrino.

$$^{50}_{23}\text{V} + {}^{0}_{-1}e \xrightarrow{\text{Electron capture}} {}^{50}_{22}\text{Ti} + \text{X rays} + \nu$$

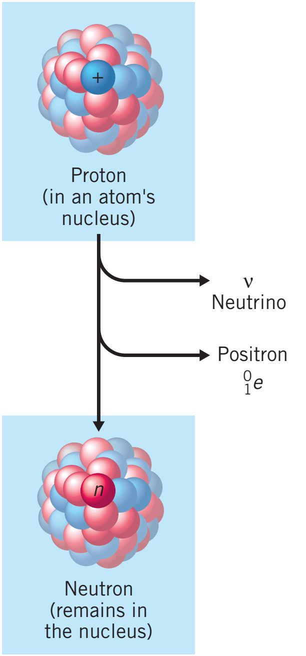


FIG. 20.4 The emission of a positron replaces a proton by a neutron.

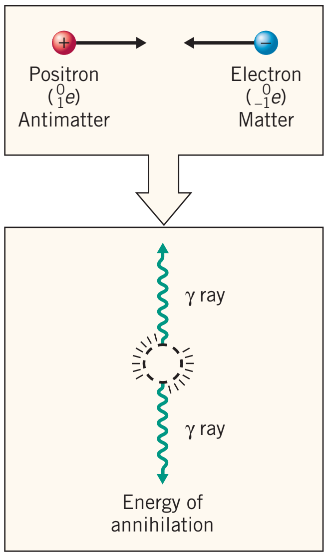


FIG. 20.5 Gamma radiation is produced when a positron and an electron collide. Annihilation of a positron and an electron leads to two gamma-ray photons.

The net effect in the nucleus of electron capture is the conversion of a proton into a neutron (Figure 20.6).

$$^1_1p + {}^0_{-1}e \longrightarrow {}^1_0n$$

proton	electron	neutron
(in the nucleus)	(captured from the 1*s* orbital)	(in the nucleus)

Electron capture does not change an atom's mass number, only its atomic number. It also leaves a hole in the first or second electron shell, and the atom emits photons of X rays as other orbital electrons drop down to fill the hole. Moreover, the nucleus that has just captured an orbital electron may be in an excited energy state and so can emit a gamma ray photon.

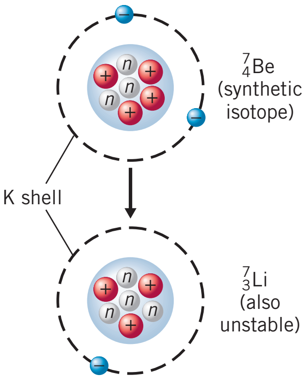


FIG. 20.6 Electron capture. Electron capture is the collapse of an orbital electron into the nucleus, and this changes a proton into a neutron. A gap is left in a low-energy electron orbital. When an electron drops from a higher orbital to fill the gap, an X-ray photon is emitted.

A radioactive disintegration series is a sequence of successive nuclear reactions

Sometimes a radionuclide does not decay directly to a stable isotope, but decays instead to another unstable radionuclide. The decay of one radionuclide after another will continue until a stable isotope forms. The sequence of such successive nuclear reactions is called a **radioactive disintegration series.** Four series occur naturally. Uranium-238 is at the head of one of them (Figure 20.7).

The rates of decay of radionuclides vary and are usually described by specifying their half lives, $t_{1/2}$, a topic we studied in Section 13.4. One *half-life* period in nuclear science is

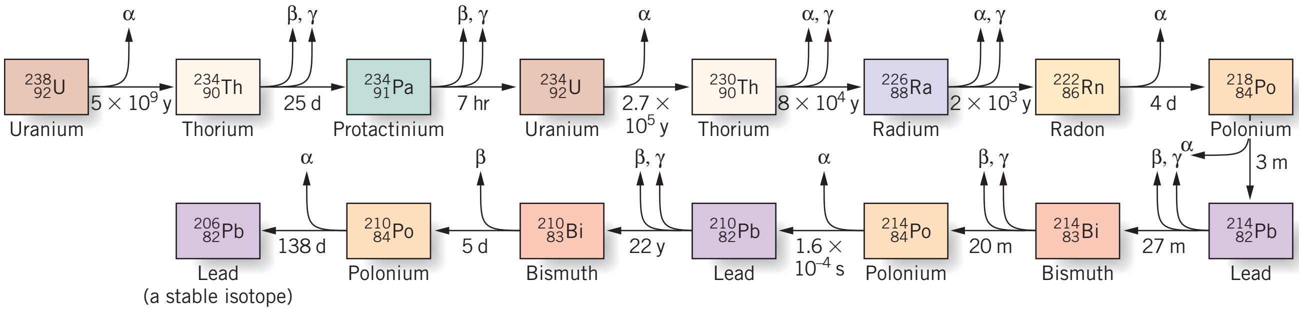


FIG. 20.7 The uranium-238 radioactive disintegration series. The time given beneath each arrow is the half-life period of the preceding isotope (y = year, m = month, d = day, hr = hour, and s = second).

TABLE 20.1 **Typical Half-Life Periods**

Element	Isotope	Half-Life	Radiations or Mode of Decay
Naturally occurring radionuclides			
Potassium	$^{40}_{19}\mathrm{K}$	1.3×10^{9} yr	beta, gamma
Tellurium	$^{123}_{52}\mathrm{Te}$	1.2×10^{13} yr	electron capture
Neodymium	$^{144}_{60}\mathrm{Nd}$	5×10^{15} yr	alpha
Samarium	$^{149}_{62}\mathrm{Sm}$	4×10^{14} yr	alpha
Rhenium	$^{187}_{75}\mathrm{Re}$	7×10^{10} yr	beta
Radon	$^{222}_{86}\mathrm{Rn}$	3.82 day	alpha
Radium	$^{226}_{88}\mathrm{Ra}$	1590 yr	alpha, gamma
Thorium	$^{230}_{90}\mathrm{Th}$	8×10^{4} yr	alpha, gamma
Uranium	$^{238}_{92}\mathrm{U}$	4.51×10^{9} yr	alpha
Synthetic radionuclides			
Tritium	$^{3}_{1}\mathrm{T}$	12.26 yr	beta
Oxygen	$^{15}_{8}\mathrm{O}$	124 s	positron
Phosphorus	$^{32}_{15}\mathrm{P}$	14.3 day	beta
Technetium[a]	$^{99m}_{43}\mathrm{Tc}$	6.02 hr	gamma
Iodine	$^{131}_{53}\mathrm{I}$	8.07 day	beta
Cesium	$^{137}_{55}\mathrm{Cs}$	30 yr	beta, gamma
Strontium	$^{90}_{38}\mathrm{Sr}$	28.1 yr	beta
Plutonium	$^{238}_{94}\mathrm{Pu}$	87.8 yr	alpha
Americium	$^{243}_{95}\mathrm{Am}$	7.37×10^{3} yr	alpha

[a] The superscript 99m refers to a metastable isotope of technetium, which has a higher energy than technetium-99.

the time it takes for a given sample of a radionuclide to decay to one-half of its initial amount. Radioactive decay is a first-order process, so the period of time taken by one half-life is independent of the initial number of nuclei. The huge variations in the half-lives of several radionuclides are shown in Table 20.1.

EXAMPLE 20.1
Writing a Balanced Nuclear Equation

Cesium-137, $^{137}_{55}\mathrm{Cs}$, one of the radioactive wastes from a nuclear power plant or an atomic bomb explosion, emits beta and gamma radiation. Write the nuclear equation for the decay of cesium-137.

Ions of cesium, which is in the same family as sodium, travel in the body to many of the same sites where sodium ions go.

ANALYSIS: We will start with an incomplete equation using the given information. Then we will use the requirements for a balanced nuclear equation as a tool to figure out any other data needed to complete the equation.

SOLUTION: The incomplete nuclear equation is

$$^{137}_{55}\mathrm{Cs} \longrightarrow {}^{\;\;0}_{-1}e + {}^{0}_{0}\gamma + {}^{\overset{\text{Mass number goes here.}}{\_\_}}_{\underset{\text{Atomic number goes here.}}{\_\_}}\overset{\text{Atomic symbol goes here.}}{\_\_}$$

The atomic symbol can be obtained from the table inside the front cover after we have determined the atomic number, *Z*. *Z* is found using the fact that the atomic number (55) on the left side of the equation must equal the sum of the atomic numbers on the right side.

$$55 = -1 + 0 + Z$$
$$Z = 56$$

The periodic table tells us that element 56 is Ba (barium). To determine which isotope of barium forms, we recall that the sums of the mass numbers on either side of the equation must also be equal. Letting *A* be the mass number of the barium isotope,

$$137 = 0 + 0 + A$$
$$A = 137$$

The balanced nuclear equation, therefore, is

$$^{137}_{55}\text{Cs} \longrightarrow {}^{0}_{-1}e + {}^{0}_{0}\gamma + {}^{137}_{56}\text{Ba}$$

IS THE ANSWER REASONABLE? Besides double-checking that element 56 is barium, the answer satisfies the requirements of a nuclear equation, namely, the sums of the mass numbers, 137, are the same on both sides as are the sums of the atomic numbers.

Practice Exercise 1: Marie Curie earned one of her two Nobel prizes for isolating the element radium, which soon became widely used to treat cancer. Radium-226, ${}^{226}_{88}\text{Ra}$, emits a gamma photon plus a particle to give radon-222. Write a balanced nuclear equation for its decay and identify the particle that's emitted. (Hint: Be sure to balance atomic number and mass.)

Practice Exercise 2: Write the balanced nuclear equation for the decay of strontium-90, a beta emitter. (Strontium-90 is one of the many radionuclides present in the wastes of operating nuclear power plants.)

20.4 STABLE ISOTOPES FALL WITHIN A "BAND OF STABILITY" ON A PLOT BASED ON NUMBERS OF PROTONS AND NEUTRONS

When all known isotopes of each element, both stable and unstable, are arrayed on a plot according to numbers of protons and neutrons, an interesting zone can be defined (Figure 20.8). The two curved lines in the array of Figure 20.8 enclose this zone, called the **band of stability,** within which lie all stable nuclei. (No isotope above element 83, bismuth, is included in Figure 20.8 because none has a *stable* isotope.) Within the band of stability are also some unstable isotopes, because smooth lines cannot be drawn to exclude them.

Any isotope not represented anywhere on the array, inside or outside the band of stability, probably has a half-life too short to permit its detection. For example, an isotope with 50 neutrons and 60 protons would lie well below the band of stability and would likely be extremely unstable. Any attempt to make it would likely be a waste of time and money.

Notice that the band curves slightly upward as the number of protons increases. The curvature means that the *ratio* of neutrons to protons gradually increases from 1:1, a ratio indicated by the straight line in Figure 20.8. The reason is easy to understand. More protons require more neutrons to provide a compensating nuclear strong force and to dilute electrostatic proton–proton repulsions.

Isotopes occurring above and to the left of the band of stability tend to be beta emitters. Isotopes lying below and to the right of the band are positron emitters. The isotopes with atomic numbers above 83 tend to be alpha emitters. Are there any reasons for these tendencies?

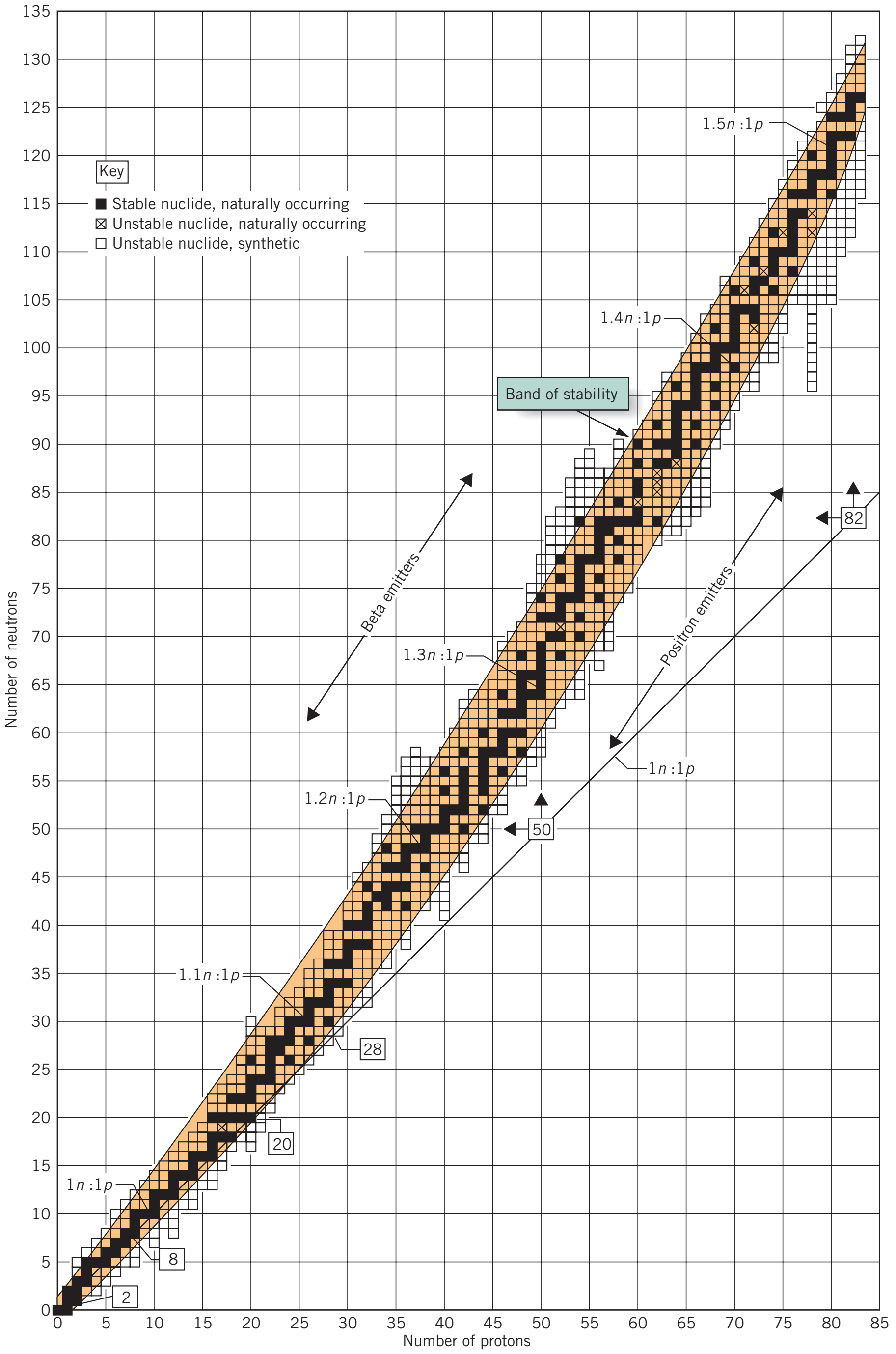


FIG. 20.8 The band of stability. Stable nuclei fall within a narrow band in a plot of number of neutrons versus number of protons. Nuclei far outside the band are too unstable to exist.

Alpha Emitters

The alpha emitters, as we said, occur mostly among the radionuclides above atomic number 83. Their nuclei have too many protons, and the most efficient way to lose protons is by the loss of an alpha particle.

Beta Emitters

Beta emitters are generally above the band of stability and so have neutron-to-proton ratios that evidently are too high. By beta decay a nucleus loses a neutron and gains a proton, thus decreasing the ratio.

The proton can also be given the symbol ${}^1_1$H in nuclear equations.

$${}^1_0n \longrightarrow {}^1_1p + {}^0_{-1}e + \overline{\nu}$$

For example, by beta decay fluorine-20 decreases its neutron-to-proton ratio from 11/9 to 10/10.

$$\begin{array}{rccc} & {}^{20}_{9}\text{F} & \longrightarrow & {}^{20}_{10}\text{Ne} + {}^{0}_{-1}e + \overline{\nu} \\ \dfrac{\text{neutron}}{\text{proton}} = & \dfrac{11}{9} & & \dfrac{10}{10} \end{array}$$

The surviving nucleus, that of neon-20, is closer to the center of the band of stability. Figure 20.9, an enlargement of the fluorine part of Figure 20.8, explains this change further. Figure 20.9 also shows how the beta decay of magnesium-27 to aluminum-27 also lowers the neutron-to-proton ratio.

Positron Emitters

In nuclei with too few neutrons to be stable, positron emission increases the neutron-to-proton ratio by changing a proton into a neutron. A fluorine-17 nucleus, for example, increases its neutron-to-proton ratio, improves its stability, and moves into the band of stability by emitting a positron and a neutrino and changing to oxygen-17 (see Figure 20.9).

$$\begin{array}{rccc} & {}^{17}_{9}\text{F} & \longrightarrow & {}^{17}_{8}\text{O} + {}^{0}_{1}e + \nu \\ \dfrac{\text{neutron}}{\text{proton}} = & \dfrac{8}{9} & & \dfrac{9}{8} \end{array}$$

Positron decay by magnesium-23 to sodium-23 produces a similarly favorable shift (also shown in Figure 20.9).

Nuclei with even numbers of protons and neutrons are likely to be stable

Study of the compositions of stable nuclei reveals an interesting relationship: Nature favors even numbers for protons and neutrons. This is summarized by the **odd–even rule**.

TOOLS
Odd–even rule

Odd–Even Rule When the numbers of neutrons and protons in a nucleus are both even, the isotope is far more likely to be stable than when both numbers are odd.

Of the 264 stable isotopes, only five have odd numbers of both protons and neutrons, whereas 157 have even numbers of both. The rest have an odd number of one nucleon and an even number of the other. To see this, notice in Figure 20.8 how the horizontal lines

▫ These five stable isotopes, ${}^{2}_{1}$H, ${}^{6}_{3}$Li, ${}^{10}_{5}$B, ${}^{14}_{7}$N, and ${}^{138}_{57}$La, all have odd numbers of both protons and neutrons.

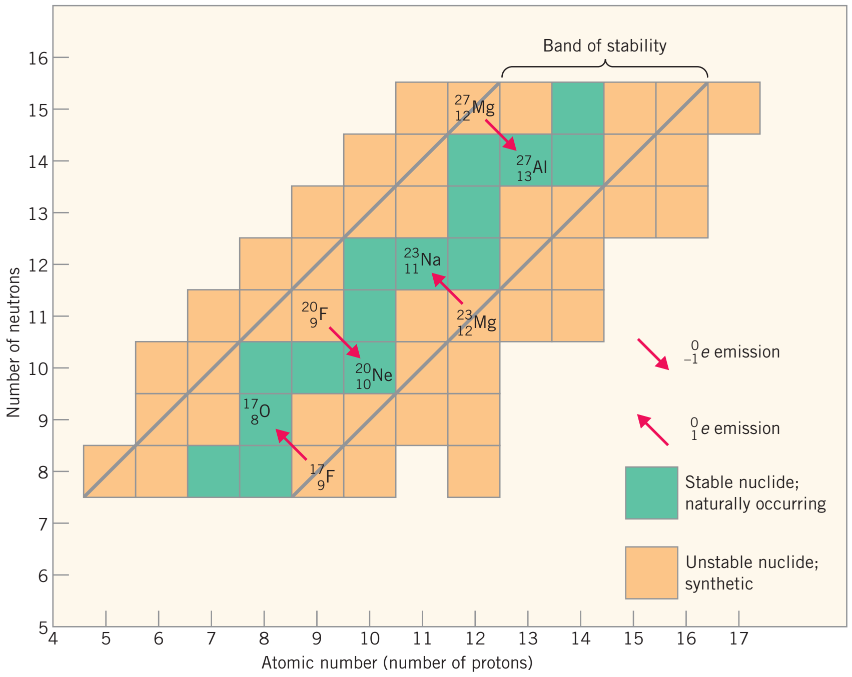


FIG. 20.9 **An enlarged section of the band of stability.** Beta decay from magnesium-27 and fluorine-20 reduces their neutron-to-proton ratio and moves them closer to the band of stability. Positron decay from magnesium-23 and fluorine-17 increases the ratio and moves those nuclides closer to the band of stability, too.

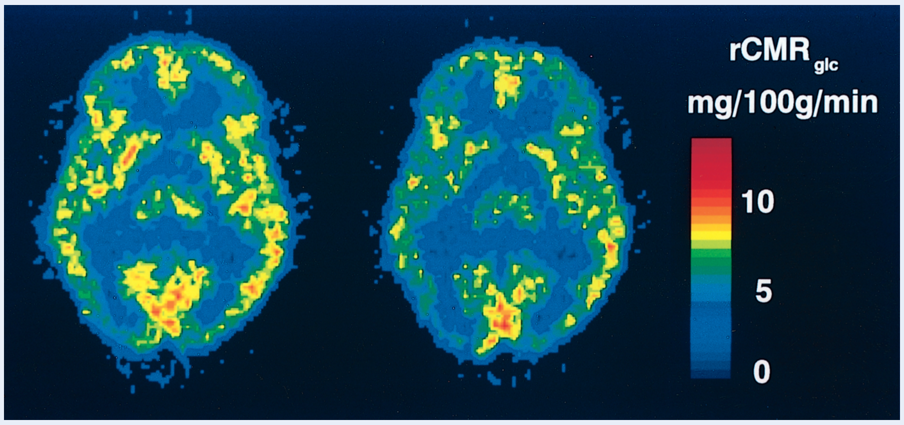
FACETS OF CHEMISTRY

20.1

Positron Emission Tomography (PET)

Positron emitters are used in an important method for studying brain function called the PET scan, for *positron emission tomography*. The technique begins by chemically incorporating positron-emitting radionuclides into molecules, like glucose, that can be absorbed by the brain directly from the blood. It's like inserting radiation generators that act from within the brain rather than focusing X rays or gamma rays from the outside. Carbon-11, for example, is a positron emitter whose atoms can be used in place of carbon-12 atoms in glucose molecules, $C_6H_{12}O_6$. (One way to prepare such glucose is to let a leafy vegetable, Swiss chard, use ${}^{11}CO_2$ to make the glucose by photosynthesis.)

A PET scan using tiny amounts of carbon-11 glucose detects situations where glucose is not taken up normally, for example, in manic depression, schizophrenia, and Alzheimer's disease. After the carbon-11 glucose is ingested by the patient, radiation detectors outside the body pick up the annihilation radiation produced when electrons react with positrons emitted at specifically those brain sites that use glucose. By mapping the locations of the brain sites using the tagged glucose, a picture showing brain function can be formed. PET scan technology, for example, demonstrated that the uptake of glucose by the brains of smokers is less than that of nonsmokers, as shown in Figure 1.



The control, a PET scan of a normal brain.

The PET scan of the brain of a volunteer injected with nicotine.

The color code indicates the rates of glucose metabolism.

FIG. 1 **Positron emission tomography (PET) in the study of brain activity.** (*Left*) Normal brain. (*Right*) Brain affected by nicotine. The PET scan reveals widespread reduction in the rate of glucose metabolism when nicotine is present. *(Courtesy of E.D. London, National Institute of Drug Abuse.)*

with the largest numbers of dark squares (stable isotopes) most commonly correspond to even numbers of neutrons. Similarly, the vertical lines with the most dark squares most often correspond to even numbers of protons.

The odd–even rule is related to the spins of nucleons. Both protons and neutrons behave as though they spin, like orbital electrons. When two protons or two neutrons have paired spins, meaning the spins are opposite, their combined energy is less than when the spins are unpaired. Only when there are even numbers of protons and neutrons can all the spins be paired and so give the nucleus a lower energy and greater stability. The least stable nuclei tend to be those with both an odd number of protons and an odd number of neutrons.

Isotopes with "magic numbers" are especially stable

Another rule of thumb for nuclear stability is based on *magic numbers* of nucleons. Isotopes with specific numbers of protons or neutrons, the **magic numbers,** are more stable than the rest. The magic numbers of nucleons are 2, 8, 20, 28, 50, 82, and 126, and where they fall is shown in Figure 20.8 (except for magic number 126).

When the numbers of both protons and neutrons are the same magic number, as they are in ${}^{4}_{2}He$, ${}^{16}_{8}O$, and ${}^{40}_{20}Ca$ the isotope is very stable. ${}^{100}_{50}Sn$ also has two identical magic numbers. Although this isotope of tin is unstable, having a half-life of only several seconds, it is much more stable than nearby radionuclides, whose half-lives are in milliseconds. Thus, although tin-100 lies well outside the band of stability, it is stable enough to be observed. One stable isotope of lead, ${}^{208}_{82}Pb$, involves two different magic numbers, 82 protons and 126 neutrons.

□ Magic numbers do not cancel the need for a favorable neutron-to-proton ratio. An atom with 82 protons and 82 neutrons lies far outside the band of stability, and yet 82 is a magic number.

The existence of magic numbers supports the hypothesis that a nucleus has a shell structure with energy levels analogous to electron energy levels. Electron levels, as you already know, are associated with their own special numbers, those that equal the maximum number of electrons allowed in a principal energy level: 2, 8, 18, 32, 50, 72, and 98 (for principal levels with n equal to 1, 2, 3, 4, 5, 6, and 7, respectively). The total numbers of electrons in the atoms of the most chemically stable elements—the noble gases—also make up a special set: 2, 10, 18, 36, 54, and 86 electrons. Thus, special sets of numbers are not unique to nuclei.

20.5 TRANSMUTATION IS THE CHANGE OF ONE ISOTOPE INTO ANOTHER

The change of one isotope into another is called **transmutation,** and radioactive decay is only one cause. Transmutation can also be forced by the bombardment of nuclei with high-energy particles, such as alpha particles from natural emitters, neutrons from atomic reactors, or protons made by stripping electrons from hydrogen. To make them better bombarding missiles, protons and alpha particles can be accelerated in an electrical field (Figure 20.10). This gives them greater energy and enables them to sweep through the target atom's orbital electrons and become buried in its nucleus. Although beta particles can be accelerated, their disadvantage is that they are repelled by a target atom's own electrons.

Transmutation occurs when compound nuclei decay

▫ *Compound* here refers only to the idea of *combination*, not to a chemical.

Both the energy and the mass of a bombarding particle enter the target nucleus at the moment of capture. The energy of the new nucleus, called a **compound nucleus,** quickly becomes distributed among all of the nucleons, but the nucleus is nevertheless rendered somewhat unstable. To get rid of the excess energy, a compound nucleus generally ejects something (a neutron, proton, or electron) and often emits gamma radiation as well. This leaves a new nucleus of an isotope different than the original target, so a transmutation has occurred overall.

Ernest Rutherford observed the first example of artificial transmutation. When he let alpha particles pass through a chamber containing nitrogen atoms, an entirely new radiation was generated, one much more penetrating than alpha radiation. It proved to be a stream of protons (Figure 20.11). Rutherford was able to show that the protons came from the decay of the compound nuclei of fluorine-18, produced when nitrogen-14 nuclei captured bombarding alpha particles.

▫ The asterisk, *, symbolizes a high-energy nucleus, a *compound nucleus.*

$$\underset{\substack{\text{alpha}\\\text{particle}}}{{}^{4}_{2}\text{He}} + \underset{\substack{\text{nitrogen}\\\text{nucleus}}}{{}^{14}_{7}\text{N}} \longrightarrow \underset{\substack{\text{fluorine}\\\text{(compound}\\\text{nucleus)}}}{{}^{18}_{9}\text{F}^{*}} \longrightarrow \underset{\substack{\text{oxygen}\\\text{(a rare but}\\\text{stable isotope)}}}{{}^{17}_{8}\text{O}} + \underset{\substack{\text{proton}\\\text{(high}\\\text{energy)}}}{{}^{1}_{1}p}$$

FIG. 20.10 Linear accelerator. In this linear accelerator at Brookhaven National Laboratory on Long Island, New York, protons can be accelerated to just under the speed of light and given up to 33 GeV of energy before they strike their target. (1 GeV = 10^9 eV.) *(Courtesy Brookhaven National Laboratory.)*

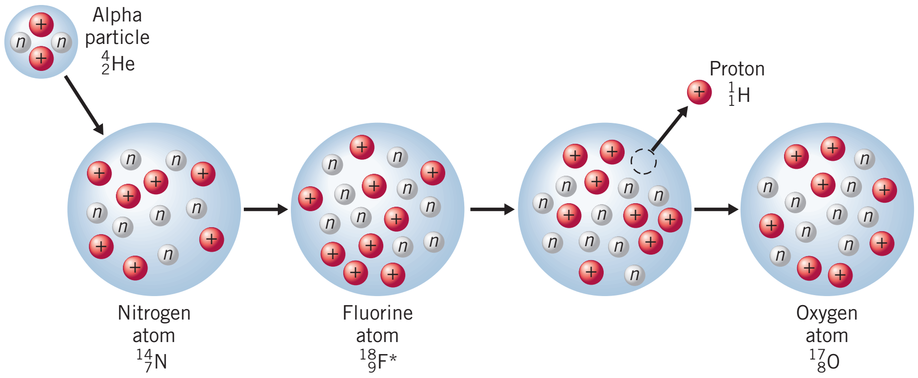


FIG. 20.11 Transmutation of nitrogen into oxygen. When the nucleus of nitrogen-14 captures an alpha particle it becomes a compound nucleus of fluorine-18. This then expels a proton and becomes the nucleus of oxygen-17.

In the synthesis of alpha particles from lithium-7, protons are used as bombarding particles. The resulting compound nucleus, that of beryllium-8, splits in two.

$$\underset{\text{proton}}{^1_1p} + \underset{\text{lithium}}{^7_3\text{Li}} \longrightarrow \underset{\text{beryllium}}{^8_4\text{Be}^*} \longrightarrow \underset{\text{alpha particles}}{2\,^4_2\text{He}}$$

Compound nuclei can follow various decay modes

A given compound nucleus can be made in a variety of ways. Aluminum-27, for example, forms by any of the following routes.

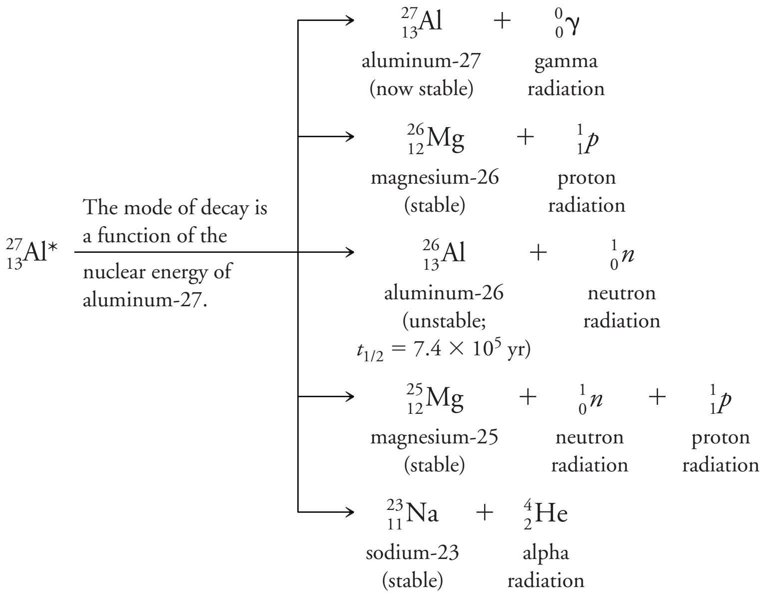
$$^4_2\text{He} + ^{23}_{11}\text{Na} \longrightarrow ^{27}_{13}\text{Al}^*$$

$$^1_1p + ^{26}_{12}\text{Mg} \longrightarrow ^{27}_{13}\text{Al}^*$$

$$\underset{\text{deuteron}}{^2_1\text{H}} + ^{25}_{12}\text{Mg} \longrightarrow ^{27}_{13}\text{Al}^*$$

A deuteron is the nucleus of a deuterium atom, just as a proton is the nucleus of a protium (hydrogen) atom.

Each path gives the compound nucleus $^{27}_{13}\text{Al}^*$ a different amount of energy. Depending on this energy, different paths of decay are available, and all of the following routes have been observed. They illustrate how the synthesis of such a large number of synthetic isotopes, some stable and others unstable, has been possible.



Transmutation is used to produce synthetic elements

Over a thousand isotopes have been made by transmutations. Most do not occur naturally; they appear in the band of stability (Figure 20.8) as open squares, nearly 900 in number. The naturally occurring radioactive isotopes above atomic number 83 all have very long half-lives. Others might have existed, but their half-lives probably were too short to permit them to last into our era. All of the elements beyond neptunium (atomic number 93 and higher, known as the **transuranium elements**) are synthetic. Elements from atomic numbers 93 to 103 complete the actinide series of the periodic table, which starts with element 90, thorium. Beyond this series, elements 104–116 and 118 have also been made.

To make the heaviest elements, bombarding particles larger than neutrons are used, such as alpha particles or the nuclei of heavier atoms. For example, element 110—darmstadtium, Ds—was made when a neutron was ejected from the compound nucleus formed by the fusion of nickel-62 and lead-208.

Element 110 was named after the place of its discovery, Darmstadt, Germany. Element 111 was named in honor of Wilhelm Roentgen, who discovered X rays in 1895.

$$^{62}_{28}\text{Ni} + ^{208}_{82}\text{Pb} \longrightarrow ^{270}_{110}\text{Ds} \longrightarrow ^{269}_{110}\text{Ds} + ^{1}_{0}n$$

Similarly, some atoms of element 111 (roentgenium, Rg) formed when a neutron was lost from the compound nucleus made by bombarding bismuth-209 with nickel-64. Most of these heavy atoms are extremely unstable, with half-lives measured in fractions of milliseconds. An exception is element 114; two isotopes have been detected with half-lives reported to be in seconds.

20.6 How is radiation measured?

Atomic radiation is often described as **ionizing radiation** because it creates ions by knocking electrons from molecules in the matter through which the radiation travels. The generation of ions is behind some of the devices for detecting radiation.

The Geiger–Müller tube, one part of a **Geiger counter,** detects beta and gamma radiation having energy high enough to penetrate the tube's window. Inside the tube is a gas under low pressure in which ions form when radiation enters. The ions permit a pulse of electricity to flow, which activates a current amplifier, a pulse counter, and an audible click.

A **scintillation counter** (Figure 20.12) contains a sensor composed of a substance called a *phosphor* that emits a tiny flash of light when struck by a particle of ionizing radiation. These flashes can be magnified electronically and automatically counted.

The darkening of a photographic film exposed to radiation over a period of time is proportional to the total quantity of radiation received. **Film dosimeters** work on this principle (Figure 20.13), and people who work near radiation sources use them. Each person keeps a log of total exposure, and if a predetermined limit is exceeded, the worker must be reassigned to an unexposed workplace.

The rate of radioactive decay is measured in disintegrations per second

The **activity** of a radioactive material is the number of disintegrations per second. The SI unit of activity is the **becquerel (Bq),** and it equals one disintegration per second. A liter

The becquerel is named after Henri Becquerel (1852–1908), the discoverer of radioactivity, who won a Nobel prize in 1903.

FIG. 20.12 Scintillation probe. Energy received from radiations striking the phosphor at the end of the probe is amplified by a photomultiplier unit and sent to the instrument where the intensity of the radiation is displayed on a meter. *(Research Products International Corp.)*

FIG. 20.13 **Badge dosimeter.** *(Cliff Moore/Photo Researchers, Inc.)*

Marie Curie (1867–1934); Nobel prizes 1903 (physics) and 1911 (chemistry). *(Courtesy College of Physicians of Philadelphia.)*

of air has an activity of about 0.04 Bq, due to carbon-14 in its carbon dioxide. A gram of natural uranium has an activity of about 2.6×10^4 Bq.

The **curie (Ci),** named after Marie Curie, the discoverer of radium, is an older unit, being equal to the activity of a 1.0 g sample of radium-226.

$$1 \text{ Ci} = 3.7 \times 10^{10} \text{ disintegrations s}^{-1} = 3.7 \times 10^{10} \text{ Bq} \qquad (20.3)$$

For a sufficiently large sample of a radioactive material, the activity is experimentally found to be proportional to the number of radioactive nuclei, N:

$$\text{Activity} = kN$$

The constant of proportionality, k, is called the **decay constant.** The decay constant is characteristic of the particular radioactive nuclide, and it gives the activity per nuclide in the sample. Since activity is the number of disintegrations per second, and therefore the change in the number of nuclei per second, we can write

TOOLS *Law of radioactive decay*

$$\text{Activity} = -\frac{\Delta N}{\Delta t} = kN \qquad (20.4)$$

which is called the **law of radioactive decay.**[1] The law shows that radioactive decay is a first-order kinetic process, and the decay constant is really just a first-order rate constant in terms of number of nuclei, rather than concentrations.

Recall from Chapter 13 that the half-life of a first-order reaction is given by Equation 13.5:

TOOLS *Half-life of a radionuclide*

$$t_{1/2} = \frac{\ln 2}{k}$$

If we know the half-life of a radioisotope, we can use this relationship to compute its decay constant and also the activity of a known mass of the radioisotope, as Example 20.2 demonstrates.

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[1] Note that the minus sign is introduced to make the activity a positive number, since the change in the number of radioactive nuclei ΔN is negative.

EXAMPLE 20.2

Using the Law of Radioactive Decay

Deep space probes such as NASA's Cassini spacecraft need to keep their instruments warm enough to operate effectively. Since solar power is not available in the darkness of deep space, these craft generate heat from the radioactive decay of small pellets of plutonium dioxide. Each pellet is about the size of a pencil eraser and weighs about 2.7 g. If the pellets are pure PuO_2, with the plutonium being ^{238}Pu, what is the activity of one of the fuel pellets, in bequerels?

ANALYSIS: The tool for finding the activity is Equation 20.4. But to use the equation, we'll need the decay constant, k, and the number of plutonium-238 atoms, N, in the fuel pellet.

From Table 20.1, the half-life of ^{238}Pu is 87.8 years. We can rearrange Equation 13.5 to compute the decay constant k from the half-life:

$$k = \frac{\ln 2}{t_{1/2}}$$

We'll want the decay constant in terms of seconds because the becquerel is defined as disintegrations per second. We should therefore convert the half-life into seconds before substituting it into Equation 13.5.

We know that the pellet contains 2.7 g of PuO_2. To calculate N, we'll have to perform the following stoichiometric conversion:

$$2.7 \text{ g PuO}_2 \Leftrightarrow \text{? atoms Pu}$$

The tools here come from Chapter 3. We'll use the molar mass to find moles of PuO_2, the chemical formula to relate this to moles of Pu, and then Avogadro's number to find the number of atoms of Pu. Our strategy is to perform the following conversions:

$$2.7 \text{ g PuO}_2 \longrightarrow \text{mol PuO}_2 \longrightarrow \text{mol Pu} \longrightarrow \text{atoms Pu}$$

We'll need the formula mass of PuO_2 (270 g mol^{-1}) to perform the first conversion. For the second conversion, the formula of PuO_2 tells us that there is 1 mol Pu in 1 mol PuO_2. For the final conversion, recall that 1 mole of atoms is 6.02×10^{23} atoms.

SOLUTION: First, let's compute the decay constant from the half-life. We must convert the half-life into seconds:

$$87.8 \cancel{\text{years}} \times \frac{365 \cancel{\text{days}}}{1 \cancel{\text{year}}} \times \frac{24 \cancel{\text{hours}}}{1 \cancel{\text{day}}} \times \frac{60 \cancel{\text{min}}}{1 \cancel{\text{hour}}} \times \frac{60 \text{ s}}{1 \cancel{\text{min}}} = 2.77 \times 10^9 \text{ s}$$

Now we can solve Equation 13.5 for the decay constant:

$$\begin{aligned} k &= \frac{\ln 2}{t_{1/2}} \\ &= \frac{0.693}{2.77 \times 10^9 \text{ s}} \\ &= 2.50 \times 10^{-10} \text{ s}^{-1} \end{aligned}$$

Next, we need the number of ^{238}Pu atoms in the fuel pellet:

$$2.7 \cancel{\text{g PuO}_2} \times \frac{1 \cancel{\text{mol PuO}_2}}{270 \cancel{\text{g PuO}_2}} \times \frac{1 \cancel{\text{mol Pu}}}{1 \cancel{\text{mol PuO}_2}} \times \frac{6.02 \times 10^{23} \text{ atoms Pu}}{1 \cancel{\text{mol Pu}}} = 6.0 \times 10^{21} \text{ atoms Pu}$$

From the law of radioactive decay, Equation 20.4, the activity of the fuel pellet is

$$\text{Activity} = kN = 2.50 \times 10^{-10} \text{ s}^{-1} \times 6.0 \times 10^{21} \text{ atoms Pu} = 1.5 \times 10^{12} \text{ atoms Pu/s}$$

Since the becquerel is defined as the number of disintegrations per second, and each plutonium atom corresponds to one disintegration, we can report the activity as 1.5×10^{12} Bq.

IS THE ANSWER REASONABLE? There is no easy way to check the size of the decay constant beyond checking the arithmetic. The number of Pu atoms in the pellet makes sense, because if we have 2.7 g PuO_2 and the formula mass is 270 g mol^{-1}, we have 1/100th of a mole of PuO_2 and so 1/100 a mole of Pu. We should have (1/100) of 6.02×10^{23} atoms, or 6.02×10^{21} atoms of Pu.

The fuel pellet has about 40 times the activity of a gram of radium-226, as indicated by Equation 20.3, so the activity of ^{238}Pu per gram is about 15 times the activity of ^{226}Ra per gram.

Practice Exercise 3: A 2.00 g sample of a mixture of plutonium with a nonradioactive metal has an activity of 6.22×10^{11} Bq. What is the percentage by mass of plutonium the sample? (Hint: How many atoms of Pu are in the sample?)

Practice Exercise 4: The EPA limit for radon-222 is 6 pCi (picocuries) per liter of air. How many atoms of ^{222}Rn per liter of air will produce that activity?

Other units are used to express amounts of radiation and its effects on tissue

Nuclear radiation can have varying effects depending on the energy of the radiation and its ability to be absorbed. The **gray (Gy)** is the SI unit of *absorbed dose*, and 1 gray corresponds to 1 joule of energy absorbed per kilogram of absorbing material. The **rad** is an older unit of absorbed dose, 1 rad being the absorption of 10^{-2} joule per kilogram of tissue. Thus, 1 Gy equals 100 rad. In terms of danger, if every individual in a large population received 450 rad (4.5 Gy), roughly half of the population would die in 60 days.

□ The *gray* is named after Harold Gray, a British radiologist. *Rad* stands for **r**adiation **a**bsorbed **d**ose.

The gray is not a good basis for comparing the biological effects of radiation in tissue, because these effects depend not just on the energy absorbed but also on the kind of radiation and the tissue itself. The **sievert (Sv),** the SI unit of *dose equivalent,* was invented to meet this problem. The **rem** is an older unit of dose equivalent, one still used in medicine. Its value is generally taken to equal 10^{-2} Sv. The U.S. government has set guidelines of 0.3 rem per week as the maximum exposure workers may receive. (For comparison, a chest X ray typically involves about 0.007 rem or 7 mrem.)

□ *Rem* stands for **r**oentgen **e**quivalent for **m**an, where the *roentgen* is a unit related to X ray and gamma radiation.

Radiation damages living tissue

A whole body dose of 25 rem (0.25 Sv) induces noticeable changes in human blood. A set of symptoms called *radiation sickness* develops at about 100 rem, becoming severe at 200 rem. Among the symptoms are nausea, vomiting, a drop in the white cell count, diarrhea, dehydration, prostration, hemorrhaging, and loss of hair. If each person in a large population absorbed 400 rem, half would die in 60 days. A 600 rem dose would kill everyone in the group in a week. Many workers received at least 400 rem in the moments following the steam explosion that tore apart one of the nuclear reactors at the Ukraine energy park near Chernobyl in 1986.

Radiation produces free radicals

Even small absorbed doses can be biologically harmful. The danger does not lie in the heat energy associated with the dose, which is usually very small. Rather, the harm is in the ability of ionizing radiation to create unstable ions or neutral species with odd (unpaired) electrons, species that can set off other reactions. Water, for example, can interact as follows with ionizing radiation.

$$\mathrm{H{-}\underset{\cdot\cdot}{\ddot{O}}{-}H} \xrightarrow{\text{radiation}} \left[\mathrm{H{-}\underset{\cdot\cdot}{\dot{O}}{-}H}\right]^{+} + {}^{\;\,0}_{-1}e$$

The new cation, $\left[\mathrm{H{-}\underset{\cdot\cdot}{\dot{O}}{-}H}\right]^{+}$, is unstable and one breakup path is

$$\left[\mathrm{H{-}\underset{\cdot\cdot}{\dot{O}}{-}H}\right]^{+} \longrightarrow \underset{\text{proton}}{\mathrm{H}^{+}} + \underset{\text{hydroxyl radical}}{\mathrm{:\underset{\cdot\cdot}{\dot{O}}{-}H}}$$

The proton might pick up a stray electron to become a hydrogen atom, H•.

Free radicals are discussed in more detail in Facets of Chemistry 13.1 on page 553.

Both the hydrogen atom and the hydroxyl radical are examples of **free radicals** (often simply called *radicals*), which are neutral or charged particles having one or more unpaired electrons. They are chemically very reactive. What they do once formed depends on the other chemical species nearby, but radicals can set off a series of totally undesirable chemical reactions inside a living cell. This is what makes the injury from absorbed radiation of far greater magnitude than the energy alone could inflict. A dose of 600 rem is a lethal dose in a human, but the same dose absorbed by pure water causes the ionization of only one water molecule in every 36 million.

Background radiation comes from a variety of sources

The presence of radionuclides in nature makes it impossible for us to be free from all exposure to ionizing radiation. Cosmic rays composed of high-energy photons shower on us from the sun and interstellar space. They interact with the air's nitrogen molecules to produce carbon-14, a beta emitter, which enters the food chain via photosynthesis, which converts CO_2 to sugars and starch. From soil and from building stone comes the radiation of radionuclides native to the earth's crust. The top 40 cm of soil holds an average of 1 gram of radium, an alpha emitter, per square kilometer. Naturally occurring potassium-40, a beta emitter, adds its radiation wherever potassium ions are found in the body. The presence of carbon-14 and potassium-40 together produce about 5×10^5 nuclear disintegrations per minute inside an adult human. Radon gas seeps into basements from underground formations. In fact, a little over half of the radiation we experience, on average, is from radon-222 and its decay products.

Diagnostic X rays, both medical and dental, also expose us to ionizing radiation. All these sources produce a combined **background radiation** that averages 360 mrem per person annually in the United States. The averages are roughly 82% from natural radiation and 18% from medical sources.

Radiation can be reduced by shielding and by distance

Gamma radiation and X rays are so powerful that they are effectively shielded only by very dense materials, like lead. Otherwise, one should stay as far from a source as possible, because the intensity of radiation diminishes with the *square* of the distance. This relationship is the **inverse square law,** which can be written mathematically as follows, where d is the distance from the source.

$$\text{Radiation intensity} \propto \frac{1}{d^2}$$

When the intensity, I_1, is known at distance d_1, then the intensity I_2 at distance d_2 can be calculated by the following equation.

TOOLS
Inverse square law

$$\frac{I_1}{I_2} = \frac{d_2^{\,2}}{d_1^{\,2}} \qquad (20.5)$$

This law applies only to a small source that radiates equally in all directions, with no intervening shields.

When the ratio is taken, the proportionality constant cancels, so we don't have to know what its value is.

Practice Exercise 5: If the intensity of radiation from a radioactive source is 4.8 units at a distance of 5.0 m, how far from the source would you have to move to reduce the intensity to 0.30 units? (Hint: How does intensity vary with distance?)

Practice Exercise 6: If an operator 10 m from a small source is exposed to 1.4 units of radiation, what will be the intensity of the radiation if he moves to 1.2 m from the source?

20.7 RADIONUCLIDES HAVE MEDICAL AND ANALYTICAL APPLICATIONS

Because chemical properties depend on the number and arrangement of orbital electrons and not on the specific makeup of nuclei, both radioactive and stable isotopes of an element behave the same chemically. This fact forms the basis for some uses of radionuclides. The chemical and physical properties enable scientists to get radionuclides into place in systems of interest. Then the radiation is exploited for medical or analytical purposes. Tracer analysis is an example.

Tracer analysis is used in medicine to study specific tissues in the body

In **tracer analysis,** the chemical form and properties of a radionuclide enable the system to distribute it to a particular location. The intensity of the radiation then tells something about how that site is working. In the form of the iodide ion, for example, iodine-131 is carried by the body to the thyroid gland, the only user of iodide ion in the body. The gland takes up the iodide ion to synthesize the hormone thyroxine. An underactive thyroid gland is unable to concentrate iodide ion normally and will emit less intense radiation under standard test conditions.

Tracer analyses are also used to pinpoint the locations of brain tumors, which are uniquely able to concentrate the pertechnetate ion, TcO_4^-, made from technetium-99*m*.[2] This strong gamma emitter, which resembles the chloride ion in some respects, is one of the most widely used radionuclides in medicine.

□ Technetium-99*m* is also used in bone scans to detect and locate bone cancer. Active cancer sites concentrate the Tc, which can be detected using external scanning devices.

Neutron activation analysis is used to detect trace elements

A number of stable nuclei can be changed into emitters of gamma radiation by capturing neutrons, and this makes possible a procedure called **neutron activation analysis**. Neutron capture followed by gamma emission can be represented by the following equation (where A is a mass number and X is a hypothetical atomic symbol).

$$\underset{\substack{\text{isotope}\\ \text{of element}\\ X\text{ being}\\ \text{analyzed}}}{{}^{A}X} + \underset{\text{neutron}}{{}^{1}_{0}n} \longrightarrow \underset{\substack{\text{compound}\\ \text{nucleus}\\ \text{(unstable)}}}{{}^{(A+1)}X^{*}} \longrightarrow \underset{\substack{\text{more stable}\\ \text{form of a new}\\ \text{isotope of }X}}{{}^{(A+1)}X} + \underset{\substack{\text{gamma}\\ \text{radiation}}}{{}^{0}_{0}\gamma}$$

A neutron-enriched compound nucleus emits gamma radiation at its own set of unique frequencies, and these sets of frequencies are now known for each isotope. (Not all isotopes, however, become gamma emitters by neutron capture.) The element can be identified by measuring the specific *frequencies* of gamma radiation emitted. The *concentration* of the element can be determined by measuring the *intensity* of the gamma radiation.

Neutron activation analysis is so sensitive that concentrations as low as $10^{-9}\%$ can be determined. A museum might have a lock of hair of some famous but long dead person suspected of having been slowly murdered by arsenic poisoning. If so, some arsenic would be in the hair, and neutron activation analysis could find it without destroying the specimen of hair.

Radiological dating determines the age of a sample using kinetics

The determination of the age of a geological deposit or an archaeological find by the use of the radionuclides that are naturally present is called **radiological dating.** It is based partly on the premise that the half-lives of radionuclides have been constant throughout the entire geological period. This premise is supported by the finding that half-lives are insensitive to all environmental forces such as heat, pressure, magnetism, or electrical stresses. Radiological dating of archaeological objects by carbon-14 analyses was described in Chapter 13 (page 539).

In geological dating, a pair of isotopes is sought that are related as a "parent" to a "daughter" in a radioactive disintegration series, like the uranium-238 series (Figure 20.7). Uranium-238 (as "parent") and lead-206 (as "daughter") have thus been used as a radiological dating pair of isotopes. The half-life of uranium-238 is very long, a necessary criterion for geological dating. Put simply, after the concentrations of uranium-238

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[2] The *m* in technetium-99*m* means that the isotope is in a metastable form. Its nucleus is at a higher energy level than the nucleus in technetium-99, to which technetium-99*m* decays.

and lead-206 are determined in a rock specimen, the *ratio* of the concentrations together with the half-life of uranium-238 is used to calculate the age of the rock.

Probably the most widely used isotopes for dating rock are the potassium-40/argon-40 pair. Potassium-40 is a naturally occurring radionuclide with a half-life nearly as long as that of uranium-238. One of its modes of decay is electron capture, and argon-40 forms.

$${}^{40}_{19}\text{K} + {}^{0}_{-1}e \longrightarrow {}^{40}_{18}\text{Ar}$$

The argon produced by the reaction remains trapped within the crystal lattices of the rock and is freed only when the rock sample is melted. How much has accumulated is measured with a mass spectrometer (page 47), and the observed ratio of argon-40 to potassium-40, together with the half-life of the parent, permits the age of the specimen to be estimated. Because the half-lives of uranium-238 and potassium-40 are so long, samples have to be at least 300,000 years old for either of the two parent–daughter isotope pairs to provide reliable results.

For ${}^{238}\text{U}$, $t_{1/2} = 4.51 \times 10^9$ yr.
For ${}^{40}\text{K}$, $t_{1/2} = 1.3 \times 10^9$ yr.

Carbon-14 dating determines the age of organic objects

As discussed in detail in Chapter 13, measurements of the ${}^{14}\text{C}$ to ${}^{12}\text{C}$ ratio in an ancient organic sample, such as an object made of wood or bone, permits calculation of the sample's age.

There are two approaches to carbon-14 dating. The older method, introduced by its discoverer, Willard F. Libby (Nobel Prize in Chemistry, 1960), measures the *radioactivity* of a sample taken from the specimen. The radioactivity is proportional to the concentration of carbon-14.

The newer and current method of carbon-14 dating relies on a device resembling a mass spectrometer that is able to separate the atoms of carbon-14 from the other isotopes of carbon (as well as from nitrogen-14) *and count all of them*, not just the carbon-14 atoms that decay. This approach permits the use of smaller samples (0.5–5 mg versus 1–20 g for the Libby method); it works at much higher efficiencies; and it gives more precise dates. Objects of up to 70,000 years old can be dated, but the highest accuracy involves systems no older than 7000 years.

For carbon dating to be accurate, extraordinary precautions must be taken to ensure that specimens are not contaminated by more recent sources of carbon or carbon compounds.

20.8 NUCLEAR FISSION AND NUCLEAR FUSION RELEASE LARGE AMOUNTS OF ENERGY

Nuclear fission is a process whereby a heavy atomic nucleus splits into two lighter fragments. **Nuclear fusion,** on the other hand, is a process whereby very light nuclei join to form a heavier nucleus. Both processes release large amounts of energy, as we will discuss shortly.

Nuclear fission is initiated by absorption of a neutron by an unstable nucleus

Because of their electrical neutrality, neutrons penetrate an atom's electron cloud relatively easily and so are able to enter the nucleus. Enrico Fermi discovered in the early 1930s that even slow-moving, *thermal neutrons* can be captured. (Thermal neutrons are those whose average kinetic energy puts them in thermal equilibrium with their surroundings at room temperature.) When he directed thermal neutrons at a uranium target, Fermi discovered that several different species of nuclei, all much lighter than uranium, were produced.

Without realizing it, what Fermi had observed was the nuclear fission of one particular isotope, uranium-235, present in small concentrations in naturally occurring uranium. The general reaction can be represented as follows.

$${}^{235}_{92}\text{U} + {}^{1}_{0}n \longrightarrow X + Y + b\,{}^{1}_{0}n$$

X and Y can be a large variety of nuclei with intermediate atomic numbers. Over 30 have been identified. The coefficient b has an average value of 2.47, the average number of neutrons produced by fission events. A typical specific fission is

$${}^{235}_{92}\text{U} + {}^{1}_{0}n \longrightarrow {}^{236}_{92}\text{U}^* \longrightarrow {}^{94}_{36}\text{Kr} + {}^{139}_{56}\text{Ba} + 3\,{}^{1}_{0}n$$

What actually undergoes fission is the compound nucleus of uranium-236. It has 144 neutrons and 92 protons, giving it a neutron-to-proton ratio of roughly 1.6. Initially, the emerging krypton and barium isotopes have the same ratio, and this is much too high for them. The neutron-to-proton ratio for stable isotopes with 36 to 56 protons is nearer 1.2 to 1.3 (Figure 20.8). Therefore, the initially formed, neutron-rich krypton and barium nuclides promptly eject neutrons, called *secondary neutrons,* that generally have much higher energies than thermal neutrons.

An isotope that can undergo fission after neutron capture is called a **fissile isotope.** The naturally occurring fissile isotope of uranium used in reactors is uranium-235, whose abundance among the uranium isotopes today is only 0.72%. Two other fissile isotopes, uranium-233 and plutonium-239, can be made in nuclear reactors.

Nuclear chain reactions require a critical mass of fissile material

The secondary neutrons released by fission become thermal neutrons as they are slowed by collisions with surrounding materials. They can now be captured by unchanged uranium-235 atoms. Because each fission event produces, on the average, more than two new neutrons, the potential exists for a **nuclear chain reaction** (Figure 20.14). A *chain reaction* is a self-sustaining process whereby products from one event cause one or more repetitions of the process.

If the sample of uranium-235 is small enough, the loss of neutrons to the surroundings is sufficiently rapid to prevent a chain reaction. However, at a certain *critical mass* of uranium-235, about 50 kilograms, this loss of neutrons is insufficient to prevent a sustained reaction. A virtually instantaneous fission of the sample ensues, in other words, an atomic bomb explosion. To trigger an atomic bomb, therefore, two or more subcritical masses of uranium-235 (or plutonium-239) are forced together to form a critical mass.

The energy yield from fission is very large

The binding energy per nucleon (Figure 20.1) in uranium-235 (about 7.6 MeV) is less than the binding energies of the new nuclides (about 8.5 MeV). The net change for a single

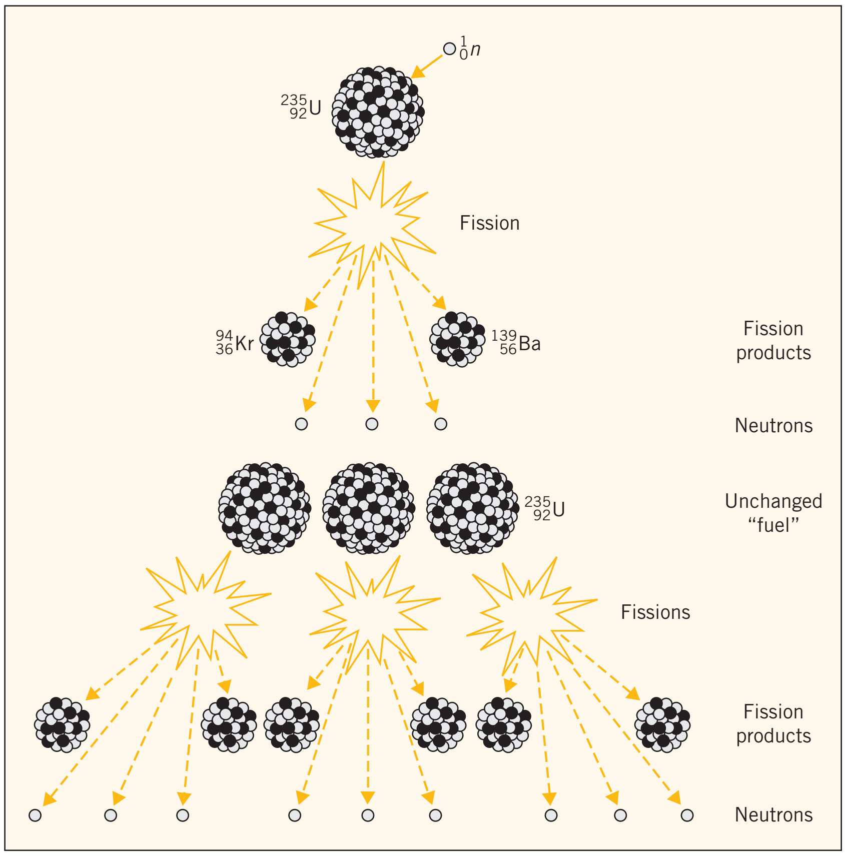


FIG. 20.14 Nuclear chain reaction. Whenever the concentration of a fissile isotope is high enough (at or above the critical mass), the neutrons released by one fission can be captured by enough unchanged nuclei to cause more than one additional fission event. In civilian nuclear reactors, the fissile isotope is too dilute for this to get out of control. In addition, control rods of nonfissile materials that are able to capture excess neutrons can be inserted or withdrawn from the reactor core to make sure that the heat generated can be removed as fast as it forms.

fission event can be calculated as follows (where we intend only *two* significant figures in each result).

Binding energy in krypton-94:

■ 1 MeV = 1.602×10^{-13} J

$$(8.5\ \text{MeV/nucleon}) \times 94\ \text{nucleons} = 800\ \text{MeV}$$

Binding energy in barium-139:

$$(8.5\ \text{MeV/nucleon}) \times 139\ \text{nucleons} = \underline{1200\ \text{MeV}}$$

$$\text{Total binding energy of products:} \quad 2000\ \text{MeV}$$

Binding energy in uranium-235:

$$(7.6\ \text{MeV/nucleon}) \times 235\ \text{nucleons} = 1800\ \text{MeV}$$

The difference in total binding energy is (2000 MeV − 1800 MeV), or 200 MeV (3.2×10^{-11} J), which has to be taken as just a rough calculation. This is the energy released by each fission event going by the equation given. The energy produced by the fission of 1 kg (4.25 mole) of uranium-235 is calculated to be roughly 8×10^{13} J, enough to keep a 100 watt lightbulb in energy for 3000 years.

■ The energy available from 1 kg of uranium-235 is equivalent to the energy of combustion of 3000 tons of soft coal or 13,200 barrels of oil.

Heat from fission reactions can be used to drive electrical generators

Virtually all civilian nuclear power plants throughout the world operate on the same general principles. The energy of fission is used, as heat, either directly or indirectly to increase the pressure of some gas that then drives an electrical generator.

The heart of a nuclear power plant is the *reactor,* where fission takes place in the *fuel core*. The nuclear fuel is generally uranium oxide, enriched to 2–4% in uranium-235 and formed into glasslike pellets. These are housed in long, sealed metal tubes called *cladding*. Bunches of tubes are assembled in spacers that permit a coolant to circulate around the tubes. A reactor has several such assemblies in its fuel core. The coolant carries away the heat of fission.

There is no danger of a nuclear power plant undergoing an atomic bomb explosion. An atomic bomb requires uranium-235 at a concentration of 85% or greater or plutonium-239 at a concentration of at least 93%. The concentration of fissile isotopes in a reactor is in the range of 2 to 4%, and much of the remainder is the common, nonfissile uranium-238. However, if the coolant fails to carry away the heat of fission, the reactor core can melt, and the molten mass might even go through the thick-walled containment vessel in which the reactor is kept. Or the high heat of the fission might split molecules of coolant water into hydrogen and oxygen, which, on recombining, would produce an immense explosion.

To convert secondary neutrons to thermal neutrons, the fuel core has a *moderator*, which is the coolant water itself in nearly all civilian reactors. Collisions between secondary neutrons and moderator molecules heat up the moderator. This heat energy eventually generates steam that enables an electric turbine to run. Ordinary water is a good moderator, but so are heavy water (D_2O) and graphite.

■ D_2O is deuterium oxide. Deuterium is an isotope of hydrogen, ${}^{2}_{1}H$.

Two main types of reactors dominate civilian nuclear power, the *boiling water reactor* and the *pressurized water reactor*. Both use ordinary water as the moderator and so are sometimes called *light water reactors.* Roughly two-thirds of the reactors in the United States are the pressurized-water type (Figure 20.15). Such a reactor has two loops, and water circulates in both. The primary loop moves water through the reactor core, where it picks up thermal energy from fission. The water is kept in the *liquid* state by being under high pressure (hence the name *pressurized* water reactor).

The hot water in the primary loop transfers thermal energy to the secondary loop at the steam generator (Figure 20.15). This makes steam at high temperature and pressure, which is piped to the turbine. As the steam drives the turbine, the steam pressure drops. The condenser at the end of the turbine, cooled by water circulating from a river or lake or from huge cooling towers, forces a maximum pressure drop within the turbine by condensing the steam to liquid water. The returned water is then recycled to high-pressure steam. (In the boiling water reactor, there is only one coolant loop. The water heated in the reactor itself is changed to the steam that drives the turbine.)

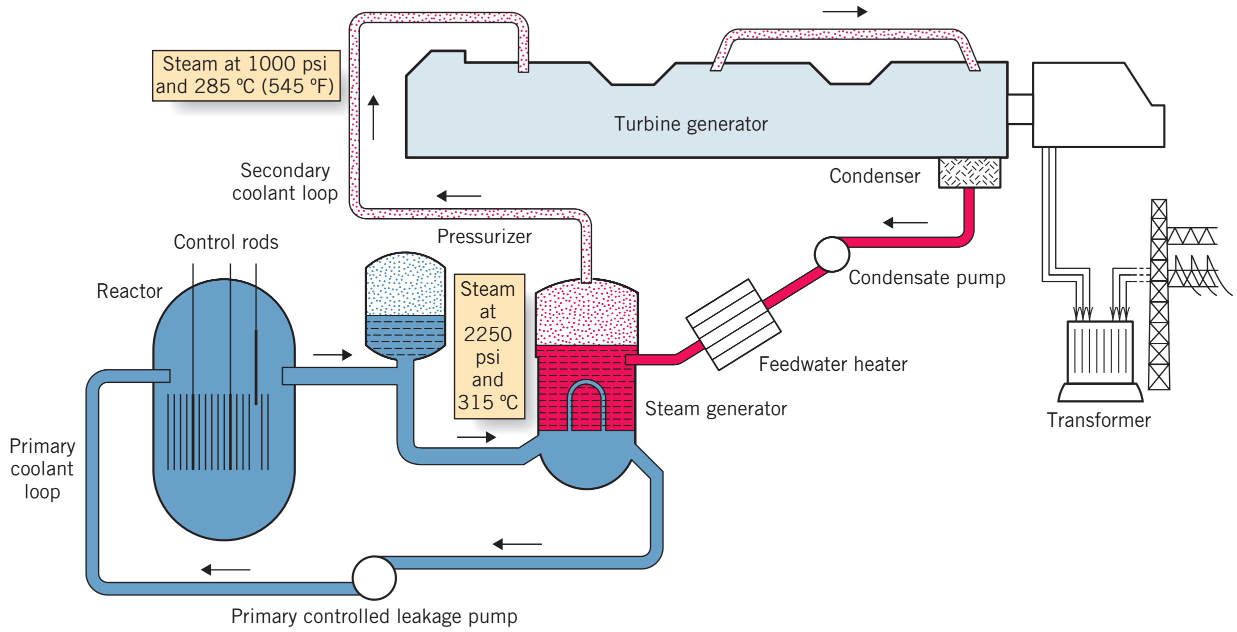


FIG. 20.15 **Pressurized water reactor, the type used in most of the nuclear power plants in the United States.** Water in the primary coolant loop is pumped around and through the fuel elements in the core, and it carries away the heat of the nuclear chain reactions. The hot water delivers its heat to the cooler water in the secondary coolant loop, where steam is generated to drive the turbines. *(Drawing from WASH-1261, U.S. Atomic Energy Commission, 1973.)*

Nuclear power plants produce several types of radioactive waste

Radioactive wastes from nuclear power plants occur as gases, liquids, and solids. The gases are mostly radionuclides of krypton and xenon but, with the exception of xenon-85 ($t_{1/2}$ = 10.4 years), the gases have short half-lives and decay quickly. During decay, they must be contained, and this is one function of the cladding. Other dangerous radionuclides produced by fission include iodine-131, strontium-90, and cesium-137.

□ Cesium-137 has a half-life of 30 years; that of strontium-90 is 28.1 years. Both are beta emitters.

Iodine-131 must be contained because the human thyroid gland concentrates iodide ion to make the hormone thyroxine. Once the isotope is in the gland, beta radiation from iodine-131 could cause harm, possibly thyroid cancer and possibly impaired thyroid function. An effective countermeasure to iodine-131 poisoning is to take doses of ordinary iodine (as sodium iodide). Statistically, this increases the likelihood that the thyroid will take up stable iodide ion rather than the unstable anion of iodine-131.

Cesium-137 and strontium-90 also pose problems to humans. Cesium is in Group IA together with sodium, so radiating cesium-137 cations travel to some of the same places in the body where sodium ions go. Strontium is in Group IIA with calcium, so strontium-90 cations can replace calcium ions in bone tissue, sending radiation into bone marrow and possibly causing leukemia. The half-lives of cesium-137 and strontium-90, however, are relatively short.

Some radionuclides in wastes are so long-lived that solid reactor wastes must be kept away from all human contact for dozens of centuries, longer than any nation has ever yet endured. Probably the most intensively studied procedure for making solid radioactive wastes secure is to convert them to glasslike or rocklike solids, and bury them deeply within a rock stratum or in a mountain believed to be geologically stable with respect to earthquakes or volcanoes. Finding a location for such a site has been the subject of much scientific and political debate.

Nuclear fusion occurs when light nuclei join to form a heavier nucleus

In Section 20.2 we mentioned that joining, or *fusing,* two light nuclei that lie to the left of the peak in the nuclear binding energy curve in Figure 20.1 leads to a net increase in nuclear binding energy and a corresponding release of energy. The process is called *nuclear* *fusion,* and the amount of energy released is considerably greater than in fission. Harnessing this energy for peaceful purposes is still a long way off, however, because many immensely difficult scientific and engineering problems remain to be solved.

Deuterium, ${}^{2}_{1}\mathrm{H}$, an isotope of hydrogen, is a key fuel in all approaches to fusion. It is naturally present as 0.015% of all hydrogen atoms, including those in water. Despite this low percentage, the earth has so much water that the supply of deuterium is virtually limitless.

The fusion reaction most likely to be used in a successful fusion reactor involves fusion of deuterium and another isotope of hydrogen, tritium, ${}^{3}_{1}\mathrm{H}$.

$$\underset{\text{deuterium}}{{}^{2}_{1}\mathrm{H}} + \underset{\text{tritium}}{{}^{3}_{1}\mathrm{H}} \longrightarrow \underset{\text{helium}}{{}^{4}_{2}\mathrm{He}} + \underset{\text{neutron}}{{}^{1}_{0}n} + 17.6\ \mathrm{MeV}$$

This corresponds to an energy yield of 2.82×10^{-12} J for each atom of helium formed, or 1.70×10^{9} kJ per mole of He formed. One problem with this reaction is that tritium is radioactive with a relatively short half-life, so it doesn't occur naturally. It can be made in several ways, however, from lithium or even deuterium via other nuclear reactions.

Comparing fission and fusion on a mass basis, fission of one kilogram of ${}^{235}\mathrm{U}$ yields approximately 8×10^{13} J, whereas forming one kilogram of ${}^{4}\mathrm{He}$ by the fusion reaction above yields 4.2×10^{14} J. Therefore, on a mass basis, fusion yields more than five times as much energy as fission. The potential energy yield from fusion is so great that the deuterium in just 0.005 cubic kilometers of ocean would supply the energy needs of the United States for one year.

Thermonuclear fusion uses high temperatures to overcome electrostatic repulsions between nuclei

The central scientific problem with fusion is to get the fusing nuclei close enough for a long enough time that the nuclear strong force (of attraction) can overcome the electrostatic force (of repulsion). As we learned in Section 20.3, the strong force acts over a much shorter range than the electrostatic force. Two nuclei on a collision course, therefore, repel each other virtually until they touch and get into the range of the strong force. The kinetic energies of two approaching nuclei must therefore be very substantial if they are to overcome this electrostatic barrier. Achieving such energies, moreover, must be accomplished by large numbers of nuclei all at once in batch after batch if there is to be any practical generation of electrical power by nuclear fusion. Relatively isolated fusion events achieved in huge accelerators will not do. The only practical way to give batch quantities of nuclei enough energy is to transfer *thermal* energy to them, and so the overall process is called *thermonuclear fusion.* Temperatures required to provide such thermal energy are very high—more than 100 million degrees Celsius!

The atoms whose nuclei we want to fuse must first be stripped of their electrons. Thus, a high energy cost is exacted from the start but, overall, the energy yield will more than pay for it. The product is an electrically neutral, gaseous mixture of nuclei and unattached electrons called a **plasma.** The plasma must then be made so dense that like-charged nuclei are within 2 fm (2×10^{-15} m) of each other, meaning a plasma density of roughly 200 g cm^{-3} as compared with 200 mg cm^{-3} under ordinary conditions. To achieve this, the plasma must be confined at a pressure of several billion atmospheres long enough for the separate nuclei to fuse. The temperature needed is several times the temperature at the center of our sun.

▫ The interior of the sun is at a temperature of approximately 15 million kelvins (15 MK).

Although practical peaceful uses for thermonuclear fusion are still in the distant future, military applications have been around for over 60 years. Thermonuclear fusion is the source of the energy released in the explosion of a hydrogen bomb. The energy needed to trigger the fusion is provided by the explosion of a fission bomb based on either uranium or plutonium.

Fusion reactions are the source of energy in stars

Nature has used thermonuclear fusion since the origin of the universe as the source of energy in stars, where high temperatures (over 15 megakelvins) and huge gravity provide the kinetic energy and high density needed to initiate fusion reactions. The chief process in solar-mass stars like our sun is called the proton–proton cycle:

The Proton–Proton Cycle

$$2\left[{}^{1}_{1}\text{H} + {}^{1}_{1}\text{H}\right] \longrightarrow 2\,{}^{2}_{1}\text{H} + 2\,{}^{\;0}_{1}e + 2\nu$$

$$2\left[{}^{1}_{1}\text{H} + {}^{2}_{1}\text{H}\right] \longrightarrow 2\,{}^{3}_{2}\text{He} + 2\gamma$$

$${}^{3}_{2}\text{He} + {}^{3}_{2}\text{He} \longrightarrow {}^{4}_{2}\text{He} + {}^{1}_{1}\text{H} + {}^{1}_{1}\text{H}$$

$$\text{Net:} \quad 4\,{}^{1}_{1}\text{H} \longrightarrow {}^{4}_{2}\text{He} + 2\,{}^{\;0}_{1}e + 2\nu + 2\gamma$$

The positrons produced combine with electrons in the plasma, annihilate each other, and generate additional energy and gamma radiation. Virtually all the neutrinos escape the sun and move into the solar system (and beyond), carrying with them a little less than 2% of the energy generated by the cycle. Not counting the energy of the neutrinos, each operation of one cycle generates 26.2 MeV or 4.20×10^{-12} J, which is equivalent to 2.53×10^{12} J per *mole* of alpha particles produced. This is the source of the solar energy radiated throughout our system, which can continue in this way for another 5 billion years.

☐ Even at a temperature of 15 megakelvins, the rate of energy production per cubic centimeter by fusion in the sun is quite small, only about 10^{-4} J s^{-1} cm^{-3}. (That's thousands of times less than the rate at which a human body generates heat!) But because the sun has such a large volume, the *total* rate of energy production is enormous.

SUMMARY

The Einstein Equation. Mass and energy are interconvertible. The **Einstein equation,** $\Delta E = \Delta mc^2$ (where *c* is the speed of light), lets us calculate one from the other. The total of the energy in the universe and all the mass calculated as an equivalent of energy is a constant, which is the **law of conservation of mass–energy.**

Nuclear Binding Energies. When a nucleus forms from its nucleons, some mass changes into energy. This amount of energy, the **nuclear binding energy,** would be required to break up the nucleus again. The higher the binding energy per nucleon, the more stable is the nucleus.

Radioactivity. The *electrostatic force* by which protons repel each other is overcome in the nucleus by the *nuclear strong force.* The ratio of neutrons to protons is a factor in nuclear stability. By radioactivity, the naturally occurring **radionuclides** adjust their neutron-to-proton ratios, lower their energies, and so become more stable by emitting **alpha** or **beta radiation,** sometimes gamma radiation as well.

The loss of an **alpha particle** leaves a nucleus with four fewer units of mass number and two fewer of atomic number. Loss of a **beta particle** leaves a nucleus with the same mass number and an atomic number one unit higher. **Gamma radiation** lets a nucleus lose some energy without a change in mass or atomic number. Depending on the specific isotope, synthetic radionuclides emit alpha, beta, and gamma radiation. Some emit **positrons** (positive electrons, a form of **antimatter**) that produce gamma radiation by annihilation collisions with electrons. Other synthetic radionuclides decay by **electron capture** and emit **X rays.** Some radionuclides emit neutrons.

Nuclear equations are balanced when the mass numbers and atomic numbers on either side of the arrow respectively balance. The energies of emission are usually described in **electron volts (eV)** or multiples thereof (1 eV = 1.602×10^{-19} J).

A few very long-lived radionuclides in nature, like ^{238}U, are at the heads of **radioactive disintegration series,** which represent the successive decays of "daughter" radionuclides until a stable isotope forms.

Nuclear Stability. Stable nuclides generally fall within a curving band, called the **band of stability,** when all known nuclides are plotted according to their numbers of neutrons and protons. Radionuclides that have too high neutron-to-proton ratios eject beta particles to adjust their ratios downward. Those with neutron-to-proton ratios too low generally emit positrons to change their ratios upward.

Isotopes whose nuclei consist of even numbers of both neutrons and protons are generally much more stable than those with odd numbers of both; this is the **odd–even rule.** Having all neutrons paired and all protons paired is energetically better (more stable) than having any nucleon unpaired. Isotopes with specific numbers of protons or neutrons, the **magic numbers** of 2, 8, 20, 28, 50, 82, and 126, are generally more stable than others.

Transmutation. When a bombardment particle—a proton, deuteron, alpha particle, or neutron—is captured, the resulting **compound nucleus** contains the energy of both the captured particle and its nucleons. The mode of decay of the compound nucleus is a function of its extra energy, not its extra mass. Many radionuclides have been made by these nuclear reactions, including all of the elements from atomic number 93 and higher.

Detecting and Measuring Radiations. Instruments to detect and measure **ionizing radiation—Geiger counters** or **scintillation counters,** for example—take advantage of the ability of radiation to generate ions in air or other matter. Such radiation can harm living tissue by producing **free radicals.**

The **curie (Ci)** and the **becquerel (Bq),** the SI unit, describe how active a source is; 1 Ci = 3.7×10^{10} Bq where 1 Bq = 1 disintegration s^{-1}.

The SI unit of absorbed dose, the **gray (Gy),** is used to describe how much energy is absorbed by a unit mass of absorber; 1 Gy = 1 J kg^{-1}. An older unit, the **rad,** is equal to 0.01 Gy.

The **sievert,** an SI unit, and the **rem,** an older unit, are used to compare doses absorbed by different tissues and caused by different kinds of radiation. A 600 rem whole body dose is lethal.

The normal **background radiation** causes millirem exposures per year. Naturally occurring radon, cosmic rays, radionuclides in soil and stone building materials, medical X rays, and releases from nuclear tests or from nuclear power plants all contribute to this background.

Protection against radiation can be achieved by using dense shields (e.g., lead or thick concrete), by avoiding overuse of radionuclides or X rays in medicine, and by taking advantage of the **inverse square law.** This law tells us that the intensity of radiation decreases with the square of the distance from its source.

Applications. **Tracer analysis** uses small amounts of a radionuclide, which can be detected using devices like the scintillation counter, to follow the path of chemical and biological processes. In **neutron activation analysis,** neutron bombardment causes some elements to become γ emitters. The radiation can be detected and measured, giving the identity and concentration of the activated elements. **Radiological dating** uses the known half-lives of naturally occurring radionuclides to date geological and archeological objects.

Fission and Fusion. Uranium-235, which occurs naturally, and plutonium-239, which can be made from uranium-238, are **fissile isotopes** that serve as the fuel in present-day reactors. When either isotope captures a thermal neutron, the isotope splits in one of several ways to give two smaller isotopes plus energy and more neutrons. The neutrons can generate additional fission events, enabling a nuclear chain reaction. If a critical mass of a fissile isotope is allowed to form, the **nuclear chain reaction** proceeds out of control, and the material detonates as an atomic bomb explosion. *Pressurized water reactors* are the most commonly used fission reactors for power generation, and have two loops of circulating fluids. In the primary loop, water circulates around the reactor core and absorbs the heat of fission. In the secondary loop, water accepts the heat and changes to high-pressure steam, which drives the electrical generator. One major problem with nuclear energy is the storage of radioactive wastes.

Thermonuclear fusion joins two light nuclei to form a heavier nucleus with the release of more energy than nuclear fission. A typical reaction combines ${}^{2}_{1}H$ and ${}^{3}_{1}H$ to give ${}^{4}_{2}He$ and a neutron. High temperatures and pressures are necessary to initiate the fusion reaction. In stars, gravity is able to contain the high temperature **plasma** and allow fusion to occur. In a hydrogen bomb, the reaction is initiated by a fission bomb. Scientific and engineering hurdles must still be overcome before fusion can be a viable peaceful energy source.

TOOLS FOR PROBLEM SOLVING

In this chapter you learned to apply the following concepts as tools in solving problems related to nuclear changes and their applications. Study each tool carefully so that you know what each is used for. When faced with solving a problem, recall what each tool does and consider whether it will be helpful in finding a solution. This will aid you in selecting the tools you need.

The Einstein equation *(page 822)* This equation, $\Delta E = \Delta m_0 c^2$ (or often just $E = mc^2$), is used when you have to relate an amount of mass to its equivalent in energy. Be careful of units in using the equation. To obtain joules, mass must be in units of kilograms and c must have units of m/s (because $1 \text{ J} = 1 \text{ kg m}^2 \text{ s}^{-2}$).

Nuclear equations *(page 825)* When you have to write and balance a nuclear equation, remember to apply the following two criteria:

1. The sums of the mass numbers on each side of the arrow must be equal.
2. The sums of the atomic numbers on each side must be the same.

The odd–even rule *(page 832)* This rule allows you to judge and compare the likely stability of nuclei according to their numbers of protons and neutrons.

When the numbers of neutrons and protons in a nucleus are both even, the isotope is far more likely to be stable than when both numbers are odd.

Law of radioactive decay *(page 837)* Use this law to relate the *activity* (in units of disintegrations per second, or Bq) to the *decay constant, k* (a first-order rate constant), and the number of atoms of the radionuclide in the sample, N. The activity is a quantity that can be measured using a Geiger or scintillation counter.

$$\text{Activity} = -\frac{\Delta N}{\Delta t} = kN$$

Half-life of a radionuclide *(page 837)* When you know the half-life of an isotope (which is available in tables), you can calculate the decay constant, k. This is useful when you need to apply the law of radioactive decay (see above).

$$t_{1/2} = \frac{\ln 2}{k}$$

Inverse square law *(page 840)* This simple law lets you compute radiation intensity at various distances from a radioactive source. If the intensity, I_1, is known at distance d_1, then the intensity I_2 at distance d_2 can be calculated by

$$\frac{I_1}{I_2} = \frac{d_2^{\,2}}{d_1^{\,2}}$$

QUESTIONS, PROBLEMS, AND EXERCISES

Answers to problems whose numbers are printed in color are given in Appendix B. More challenging problems are marked with asterisks. ILW = Interactive Learningware solution is available at www.wiley.com/college/brady. OH = an Office Hours video is available for this problem.

REVIEW QUESTIONS

Conservation of Mass–Energy

20.1 In chemical calculations involving chemical reactions we can regard the law of conservation of mass as a law independent of the law of conservation of energy despite Einstein's union of the two. What fact(s) makes this possible?

20.2 How can we know that the speed of light is the absolute upper limit on the speed of any object?

20.3 State the following.
(a) law of conservation of mass–energy
(b) Einstein equation

20.4 Why isn't the sum of the masses of all nucleons in one nucleus equal to the mass of the actual nucleus?

Radioactivity

20.5 When a substance is described as *radioactive*, what does that mean? Why is the term *radioactive decay* used to describe the phenomenon?

20.6 Three kinds of radiation make up nearly all of the radiation observed from naturally occurring radionuclides. What are they?

20.7 Give the composition of each of the following.
(a) alpha particle
(b) beta particle
(c) positron
(d) deuteron

20.8 Why is the penetrating ability of alpha radiation less than that of beta or gamma radiation?

20.9 With respect to their formation, how do gamma rays and X rays differ?

20.10 How does electron capture generate X rays?

Nuclear Stability

20.11 What data are plotted and what criterion is used to identify the actual band in the band of stability?

20.12 Both barium-123 and barium-140 are radioactive, but which is more likely to have the *longer* half-life? Explain your answer.

20.13 Tin-112 is a stable nuclide but indium-112 is radioactive and has a very short half-life ($t_{1/2} = 14$ min). What does tin-112 have that indium-112 does not to account for this difference in stability?

20.14 Lanthanum-139 is a stable nuclide but lanthanum-140 is unstable ($t_{1/2} = 40$ hr). What rule of thumb concerning nuclear stability is involved?

20.15 As the atomic number increases, the neutron-to-proton ratio increases. What does this suggest is a factor in nuclear stability?

20.16 Radionuclides of high atomic number are far more likely to be alpha emitters than those of low atomic number. Offer an explanation for this phenomenon.

20.17 Although lead-164 has two magic numbers, 82 protons and 82 neutrons, this isotope is unknown. Lead-208, however, is known and stable. What problem accounts for the nonexistence of lead-164?

20.18 What decay particle is emitted from a nucleus of low to intermediate atomic number but a relatively high neutron-to-proton ratio? How does the emission of this particle benefit the nucleus?

20.19 What decay particle is emitted from a nucleus of low to intermediate atomic number but a relatively low neutron-to-proton ratio? How does the emission of this particle benefit the nucleus?

20.20 What does electron capture do to the neutron-to-proton ratio in a nucleus, increase it, decrease it, or leave it alone? Which kinds of radionuclides are more likely to undergo this change, those above or those below the band of stability?

Transmutations

20.21 Compound nuclei form and then decay almost at once. What accounts for the instability of a compound nucleus?

20.22 What explains the existence of several decay modes for the compound nucleus aluminum-27?

20.23 Rutherford theorized that a compound nucleus forms when helium nuclei hit nitrogen-14 nuclei. If this compound nucleus decayed by the loss of a neutron instead of a proton, what would be the other product?

Detecting and Measuring Radiations

20.24 What specific property of nuclear radiation is used by the Geiger counter?

20.25 Dangerous doses of radiation can actually involve very small quantities of energy. Explain.

20.26 What units, SI and common, are used to describe each of the following?
(a) the *activity* of a radioactive sample
(b) the *energy* of a particle or of a photon of radiation given off by a nucleus
(c) the amount of energy absorbed by a given mass from a dose of radiation
(d) dose equivalents for comparing biological effects

20.27 A sample giving 3.7×10^{10} disintegrations s^{-1} has what activity in Ci and in Bq?

20.28 Explain the necessity in health sciences for the *sievert.*

Applications of Radionuclides

20.29 Why should a radionuclide used in diagnostic work have a short half-life? If the half-life is too short, what problem arises?

20.30 An alpha emitter is not used in diagnostic work. Why?

20.31 In general terms, explain how neutron activation analysis is used and how it works.

20.32 What is one assumption in the use of the uranium/lead ratio for dating ancient geologic formations?

*__20.33__ If a sample used for carbon-14 dating is contaminated by air, there is a potentially serious problem with the method. What is it?

20.34 List some of the kinds of radiation that make up our background radiation.

Nuclear Fission and Fusion

20.35 Why is it easier for a nucleus to capture a neutron than a proton?

20.36 What do each of the following terms mean?

(a) thermal neutron
(c) fissile isotope
(b) nuclear fission
(d) nuclear fusion

20.37 Which fissile isotope occurs in nature?

20.38 What fact about the fission of uranium-235 makes it possible for a *chain reaction* to occur?

20.39 Explain in general terms why fission generates more neutrons than needed to initiate it.

20.40 Why would there be a *subcritical mass* of a fissile isotope? (Why isn't *any* mass of uranium-235 critical?)

20.41 What purpose is served by a *moderator* in a nuclear reactor?

20.42 Why is there no possibility of an atomic bomb explosion from a nuclear power plant?

20.43 Write the nuclear equation for the fusion reaction between deuterium and tritium. Why must tritium be synthesized for this reaction?

20.44 What obstacles make constructing a reactor for controlled nuclear fusion especially difficult?

REVIEW PROBLEMS

Conservation of Mass–Energy

20.45 Calculate the mass equivalent in grams of 1.00 kJ.

***20.46** Calculate the mass in kilograms of a 1.00 kg object when its velocity, relative to us, is (a) 3.00×10^7 m s^{-1}, (b) 2.90×10^8 m s^{-1}, and (c) 2.99×10^8 m s^{-1}. (Notice the progression of these numbers toward the velocity of light, 3.00×10^8 m s^{-1}.)

OH **20.47** Calculate the amount of mass in nanograms that is changed into energy when one mole of liquid water forms by the combustion of hydrogen, all measurements being made at 1 atm and 25 °C. What percentage is this of the total mass of the reactants?

20.48 Show that the mass equivalent to the energy released by the complete combustion of 1 mol of methane (890 kJ) is 9.89 ng.

Nuclear Binding Energies

ILW OH **20.49** Calculate the binding energy in joules per nucleon of the deuterium nucleus, whose mass is 2.0135 u.

20.50 Calculate the binding energy in joules per nucleon of the tritium nucleus, whose mass is 3.01550 u.

Radioactivity

20.51 Complete the following nuclear equations by writing the symbols of the missing particles

(a) ${}^{211}_{82}\mathrm{Pb} \longrightarrow {}^{0}_{-1}e + \_\_\_\_$

(b) ${}^{177}_{73}\mathrm{Ta} \xrightarrow{\text{electron capture}} \_\_\_\_$

(c) ${}^{220}_{86}\mathrm{Rn} \longrightarrow {}^{4}_{2}\mathrm{He} + \_\_\_\_$

(d) ${}^{19}_{10}\mathrm{Ne} \longrightarrow {}^{0}_{1}e + \_\_\_\_$

20.52 Write the symbols of the missing particles to complete the following nuclear equations.

(a) ${}^{245}_{96}\mathrm{Cm} \longrightarrow {}^{4}_{2}\mathrm{He} + \_\_\_\_$

(b) ${}^{146}_{56}\mathrm{Ba} \longrightarrow {}^{0}_{-1}e + \_\_\_\_$

(c) ${}^{58}_{29}\mathrm{Cu} \longrightarrow {}^{0}_{1}e + \_\_\_\_$

(d) ${}^{68}_{32}\mathrm{Ge} \xrightarrow{\text{electron capture}} \_\_\_\_$

20.53 Write a balanced nuclear equation for each of the following changes.

(a) alpha emission from plutonium-242
(b) beta emission from magnesium-28
(c) positron emission from silicon-26
(d) electron capture by argon-37

OH **20.54** Write the balanced nuclear equation for each of the following nuclear reactions.

(a) electron capture by iron-55
(b) beta emission by potassium-42
(c) positron emission by ruthenium-93
(d) alpha emission by californium-251

20.55 Write the symbols, including the atomic and mass numbers, for the radionuclides that would give each of the following products.

(a) fermium-257 by alpha emission
(b) bismuth-211 by beta emission
(c) neodymium-141 by positron emission
(d) tantalum-179 by electron capture

20.56 Each of the following nuclides forms by the decay mode described. Write the symbols of the parents, giving both atomic and mass numbers.

(a) rubidium-80 formed by electron capture
(b) antimony-121 formed by beta emission
(c) chromium-50 formed by positron emission
(d) californium-253 formed by alpha emission

20.57 Krypton-87 decays to krypton-86. What other particle forms?

20.58 Write the symbol of the nuclide that forms from cobalt-58 when it decays by electron capture.

Nuclear Stability

OH **20.59** If an atom of potassium-38 had the option of decaying by positron emission or beta emission, which route would it likely take, and why? Write the nuclear equation.

20.60 Suppose that an atom of argon-37 could decay by either beta emission or electron capture. Which route would it likely take, and why? Write the nuclear equation.

20.61 If we begin with 3.00 mg of iodine-131 ($t_{1/2} = 8.07$ days), how much remains after 6 half-life periods?

20.62 A sample of technetium-99*m* with a mass of 9.00 ng will have decayed to how much of this radionuclide after 4 half-life periods (about 1 day)?

Transmutations

20.63 When vanadium-51 captures a deuteron (${}^{2}_{1}\mathrm{H}$), what compound nucleus forms? (Write its symbol.) This particle expels a proton (${}^{1}_{1}p$). Write the nuclear equation for the overall change from vanadium-51.

20.64 The alpha-particle bombardment of fluorine-19 generates sodium-22 and neutrons. Write the nuclear equation, including the intermediate compound nucleus.

OH **20.65** Gamma-ray bombardment of bromine-81 causes a transmutation in which a neutron is one product. Write the symbol of the other product.

20.66 Neutron bombardment of cadmium-115 results in neutron capture and the release of gamma radiation. Write the nuclear equation.

20.67 When manganese-55 is bombarded by protons, each ^{55}Mn nucleus releases a neutron. What else forms? Write the nuclear equation.

20.68 Which nuclide forms when sodium-23 is bombarded by alpha particles and the compound nucleus emits a gamma-ray photon?

20.69 The nuclei of which isotope of zinc-70 would be needed as bombardment particles to make nuclei of element 112 from lead-208 if the intermediate compound nucleus loses a neutron?

20.70 Write the symbol of the nuclide whose nuclei would be the target for bombardment by nickel-64 nuclei to produce nuclei of $^{272}_{111}$Rg after the intermediate compound nucleus loses a neutron.

Detecting and Measuring Radiations

20.71 Suppose that a radiologist who is 2.0 m from a small, unshielded source of radiation receives 2.8 units of radiation. To reduce the exposure to 0.28 units of radiation, to what distance from the source should the radiologist move?

20.72 By what percentage should a radiation specialist increase the distance from a small unshielded source to reduce the radiation intensity by 10.0%?

OH **20.73** If exposure from a distance of 1.60 m gave a worker a dose of 8.4 rem, how far should the worker move away from the source to reduce the dose to 0.50 rem for the same period?

20.74 During work with a radioactive source, a worker was told that he would receive 50 mrem at a distance of 4.0 m during 30 min of work. What would be the received dose if the worker moved closer, to 0.50 m, for the same period?

Law of Radioactive Decay

20.75 Smoke detectors contain a small amount of americium-241, which has a half-life of 1.70 × 10^5 days. If the detector contains 0.20 mg of ^{241}Am, what is the activity, in becquerels? In microcuries?

20.76 Strontium-90 is a dangerous radioisotope present in fallout produced by nuclear weapons. ^{90}Sr has a half-life of 1.00 × 10^4 days. What is the activity of 1.00 g of ^{90}Sr, in becquerels? In microcuries?

20.77 Iodine-131 is a radioisotope present in radioactive fallout that targets the thyroid gland. If 1.00 mg of ^{131}I has an activity of 4.6 × 10^{12} Bq, what is the decay constant for ^{131}I? What is the half-life, in seconds?

20.78 A 10.0 mg sample of thallium-201 has an activity of 7.9 × 10^{13} Bq. What is the decay constant for ^{201}Tl? What is the half-life of ^{201}Tl, in seconds?

Applications of Radionuclides

OH **20.79** What percentage of cesium chloride made from cesium-137 ($t_{1/2}$ = 30 y; beta emitter) remains after 150 y? What *chemical* product forms?

20.80 A sample of waste has a radioactivity, caused solely by strontium-90 (beta emitter, $t_{1/2}$ = 28.1 yr), of 0.245 Ci g^{-1}. How many years will it take for its activity to decrease to 1.00 × 10^{-6} Ci g^{-1}?

20.81 A worker in a laboratory unknowingly became exposed to a sample of radiolabeled sodium iodide made from iodine-131 (beta emitter, $t_{1/2}$ = 8.07 days). The mistake was realized 28.0 days after the accidental exposure, at which time the activity of the sample was 25.6 × 10^{-5} Ci g^{-1}. The safety officer needed to know how active the sample was at the time of the exposure. Calculate that value in curies per gram.

20.82 Technetium-99*m* (gamma emitter, $t_{1/2}$ = 6.02 hr) is widely used for diagnosis in medicine. A sample prepared in the early morning for use that day had an activity of 4.52 × 10^{-6} Ci. What will its activity be at the end of the day, that is, after 8.00 hr?

20.83 A 0.500 g sample of rock was found to have 2.45 × 10^{-6} mol of potassium-40 ($t_{1/2}$ = 1.3 × 10^9 yr) and 2.45 × 10^{-6} mol of argon-40. How old was the rock? (What assumption is made about the origin of the argon-40?)

20.84 If a rock sample was found to contain 1.16 × 10^{-7} mol of argon-40, how much potassium-40 would also have to be present for the rock to be 1.3 × 10^9 years old?

20.85 A tree killed by being buried under volcanic ash was found to have a ratio of carbon-14 atoms to carbon-12 atoms of 4.8 × 10^{-14}. How long ago did the eruption occur?

20.86 A wooden door lintel from an excavated site in Mexico would be expected to have what ratio of carbon-14 to carbon-12 atoms if the lintel is 9.0 × 10^3 y old?

Nuclear Energy

OH **20.87** Complete the following nuclear equation by supplying the symbol for the other product of the fission.

$$^{235}_{92}\text{U} + {}^{1}_{0}n \longrightarrow {}^{94}_{38}\text{Sr} + \underline{\qquad} + 2\,{}^{1}_{0}n$$

20.88 Both products of the fission in the previous problem are unstable. According to Figures 20.8 and 20.9, what is the most likely way for each of them to decay, by alpha emission, beta emission, or by positron emission? Explain. What are some of the possible fates of the extra neutrons produced by the fission shown in the previous problem?

ADDITIONAL EXERCISES

20.89 What is the nuclear equation for each of the following changes?
(a) beta emission from aluminum-30
(b) alpha emission from einsteinium-252
(c) electron capture by molybdenum-93
(d) positron emission by phosphorus-28

***20.90** Calculate to three significant figures the binding energy in joules per nucleon of the nucleus of an atom of iron-56. The observed mass of one *atom* is 55.9349 u. What information lets us know that no isotope has a *larger* binding energy per nucleon?

***20.91** Calculate to five significant figures the binding energy in joules per nucleon of uranium-235. The observed mass of one *atom* is 235.0439 u.

20.92 Give the nuclear equation for each of these changes.
(a) positron emission by carbon-10
(b) alpha emission by curium-243
(c) electron capture by vanadium-49
(d) beta emission by oxygen-20

***20.93** If a positron is to be emitted spontaneously, how much more *mass* (as a minimum) must an *atom* of the parent have than an *atom* of the daughter nuclide? Explain.

20.94 If a proton and an antiproton were to collide and produce two annihilation photons, what would the wavelength of the photons be, in meters? Which of the decimal multipliers in Table 1.4 (page 12) would be most appropriate for expressing this wavelength?

20.95 There is a gain in binding energy per nucleon when light nuclei fuse to form heavier nuclei. Yet, a tritium atom and a deuterium atom, in a mixture of these isotopes, does not spontaneously fuse to give helium (and energy). Explain why not.

20.96 ^{214}Bi decays to isotope *A* by alpha emission; *A* then decays to *B* by beta emission, which decays to *C* by another beta emission. Element *C* decays to *D* by still another beta emission, and *D* decays by alpha emission to a stable isotope, *E*. What is the proper symbol of element *E*? (Contributed by Prof. W. J. Wysochansky, Georgia Southern University.)

20.97 ^{15}O decays by positron emission with a half-life of 124 s. (a) Give the proper symbol of the product of the decay. (b) How much of a 750 mg sample of ^{15}O remains after 5.0 min of decay? (Contributed by Prof. W. J. Wysochansky, Georgia Southern University.)

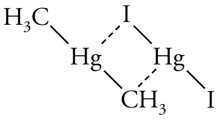
20.98 Alpha decay of ^{238}U forms ^{234}Th. What kind of decay from ^{234}Th produces ^{234}Ac? (Contributed by Prof. Mark Benvenuto, University of Detroit—Mercy.)

20.99 A sample of rock was found to contain 2.07×10^{-5} mol of ^{40}K and 1.15×10^{-5} mol of ^{40}Ar. If we assume that all of the ^{40}Ar came from the decay of ^{40}K, what is the age of the rock in years ($t_{1/2} = 1.3 \times 10^9$ years for ^{40}K.)

20.100 The ^{14}C content of an ancient piece of wood was found to be one-eighth of that in living trees. How many years old is this piece of wood ($t_{1/2} = 5730$ years for ^{14}C)?

20.101 Dinitrogen trioxide, N_2O_3, is largely dissociated into NO and NO_2 in the gas phase where there exists the equilibrium, $N_2O_3 \rightleftharpoons NO + NO_2$. In an effort to determine the structure of N_2O_3, a mixture of NO and *NO_2 was prepared containing isotopically labeled N in the NO_2. After a period of time the mixture was analyzed and found to contain substantial amounts of both *NO and *NO_2. Explain how this is consistent with the structure for N_2O_3 being ONONO.

20.102 The reaction $(CH_3)_2Hg + HgI_2 \longrightarrow 2CH_3HgI$ is believed to occur through a transition state with the structure



If this is so, what should be observed if $(CH_3)_2Hg$ and *HgI_2 are mixed, where the asterisk denotes a radioactive isotope of Hg? Explain your answer.

***20.103** A large, complex piece of apparatus has built into it a cooling system containing an unknown volume of cooling liquid. It is desired to measure the volume of the coolant without draining the lines. To the coolant was added 10.0 mL of methanol whose molecules included atoms of ^{14}C and that had a specific activity of 580 counts per minute per gram (cpm/g), determined using a Geiger counter. The coolant was permitted to circulate to assure complete mixing before a sample was withdrawn that was found to have a specific activity of 29 cpm/g. Calculate the volume of coolant in the system in milliliters. The density of methanol is 0.792 g/mL, and the density of the coolant is 0.884 g/mL.

***20.104** A complex ion of chromium(III) with oxalate ion was prepared from ^{51}Cr-labeled $K_2Cr_2O_7$, having a specific activity of 843 cpm/g (counts per minute per gram), and ^{14}C-labeled oxalic acid, $H_2C_2O_4$, having an specific activity of 345 cpm/g. Chromium-51 decays by electron capture with the emission of gamma radiation, whereas ^{14}C is a pure beta emitter. Because of the characteristics of the beta and gamma detectors, each of these isotopes may be counted independently. A sample of the complex ion was observed to give a gamma count of 165 cpm and a beta count of 83 cpm. From these data, determine the number of oxalate ions bound to each Cr(III) in the complex ion. (Hint: For the starting materials calculate the cpm per mole of Cr and oxalate, respectively.)

20.105 Iodine-131 is used to treat Graves disease, a disease of the thyroid gland. The amount of ^{131}I used depends on the size of the gland. If the dose is 86 microcuries per gram of thyroid gland, how many grams of ^{131}I should be administered to a patient with a thyroid gland weighing 20 g? Assume all the iodine administered accumulates in the thyroid gland.

20.106 The fuel for a thermonuclear bomb (hydrogen bomb) is lithium deuteride, a salt composed of the ions ${}^{6}_{3}\text{Li}^{+}$ and ${}^{2}_{1}\text{H}^{-}$. Considering the nuclear reaction

$${}^{6}_{3}\text{Li} + {}^{1}_{1}n \longrightarrow {}^{4}_{2}\text{He} + {}^{3}_{1}\text{H}$$

explain how a ^{235}U fission bomb could serve as a trigger for a fusion bomb. Write appropriate nuclear equations.

20.107 In 2006, the confirmed synthesis of ${}^{294}_{118}\text{Uuo}$ (an isotope of element 118) was reported to involve the bombardment of ^{245}Cf with ^{48}Ca. Write an equation for the nuclear reaction, being sure to include any other products of the reaction.

***20.108** Radon, a radioactive noble gas, is an environmental problem in some areas, where it can seep out of the ground and into homes. Exposure to radon-222, an alpha emitter with a half-life of 3.823 days, can increase the risk of lung cancer. At an exposure level of 4 pCi per liter (the level at which the EPA recommends action), the lifetime risk of death from lung cancer due to radon exposure is estimated to be 62 out of 1,000 for current smokers, compared with 73 out of 10,000 for nonsmokers. If the air in a home was analyzed and found to have an activity of 4.1 pCi L^{-1}, how many atoms of ^{222}Rn are there per liter of air?

***20.109** The isotope ^{145}Pr decays by emission of beta particles with an energy of 1.80 MeV each. Suppose a person swallowed, by accident, 1.0 mg of Pr having a specific activity (activity per gram) of 140 Bq g^{-1}. What would be the absorbed dose from ^{145}Pr in units of Gy and rad over a period of 10 minutes? Assume all the beta particles are absorbed by the person's body.

EXERCISES IN CRITICAL THINKING

20.110 A silver wire coated with nonradioactive silver chloride is placed into a solution of sodium chloride that is labeled with radioactive ^{36}Cl. After a while, the AgCl was analyzed and found to contain some ^{36}Cl. How do you interpret the results of this experiment?

20.111 Suppose you were given a piece of cotton cloth and told that it was believed to be 2000 years old. You performed a carbon dating test on a tiny piece of the cloth and your data indicated that it was only 800 years old. If the cloth really was 2000 years old, what factors might account for your results?

20.112 In 2006, the former Soviet spy Alexander Litvinenko was poisoned by the polonium isotope ^{210}Po. He died 23 days after ingesting the isotope, which is an alpha emitter. Find data on the Internet to answer the following: Assuming Litvinenko was fed 1 μg of ^{210}Po and that it became uniformly distributed through the cells in his body, how many atoms of ^{210}Po made their way into each cell in his body? (Assume his body contained the average number of cells found in an adult human.) Also calculate the number of cells affected by the radiation each second, being sure to take into account an estimate of the average number of cells affected by an alpha particle emitted by a ^{210}Po nucleus.

20.113 What would be the formula of the simplest hydrogen compound of element 116? Would a solution of that compound in water be acidic, basic, or neutral? Explain your reasoning.

20.114 Astatine is a halogen. Its most stable isotope, ^{210}At, has a half-life of only 8.3 hours, and only very minute amounts of the element are ever available for study ($<$0.001 μg). This amount is so small as to be virtually invisible. Describe experiments you might perfom that would tell you whether the silver salt of astatine, AgAt, is insoluble in water.