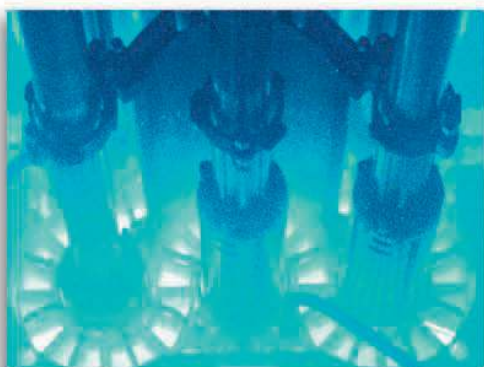


23

Nuclear Reactions and Their Applications



Producing with Nuclear Power In addition to supplying energy, many nuclear reactors, like this one in Idaho, produce medical and industrial isotopes.

Outline

23.1 Radioactive Decay and Nuclear Stability

Components of the Nucleus
Types of Radioactive Decay; Nuclear Equations
Mode of Decay

23.2 The Kinetics of Radioactive Decay

Rate of Radioactive Decay
Radioisotopic Dating

23.3 Nuclear Transmutation: Induced Changes in Nuclei

23.4 The Effects of Nuclear Radiation on Matter

Ionizing Radiation and Living Matter
Sources of Ionizing Radiation

23.5 Applications of Radioisotopes

Radioactive Tracers
Applications of Ionizing Radiation

23.6 The Interconversion of Mass and Energy

Mass Difference Between Nucleus and Nucleons
Nuclear Binding Energy

23.7 Applications of Fission and Fusion

Nuclear Fission
Nuclear Fusion

Key Principles

to focus on while studying this chapter

- **Nuclear reactions** differ markedly from chemical reactions in several ways: (1) Element identity typically *does* change in a nuclear reaction. (2) Nuclear particles and, much less often, electrons participate. (3) Nuclear reactions release so much energy that the mass *does* change. (4) Rates of nuclear reactions are *not* affected by temperature, catalysts, or, except rarely, the compound in which the element occurs (*Introduction*).
- In a balanced nuclear reaction, the total mass number (A) and total charge (Z) of the reactants must equal those of the products. Protons and neutrons are called **nucleons**; a plot of number of neutrons (N) versus number of protons (Z) for all nuclei shows a narrow **band of stability**. Unstable nuclei undergo various types of radioactive decay. The type can often be predicted by taking into account a nuclide's mass relative to the atomic mass and N/Z ratio. Nuclear stability is associated with filled **nucleon levels**. Certain heavy nuclei undergo a **decay series** to reach stability (*Section 23.1*).
- Radioactive decay is a **first-order process**, so the decay rate (**activity**) depends *only* on the number of nuclei. The **half-life**, or time required for half the nuclei present to decay, does *not* depend on the number of nuclei. In **radiocarbon dating**, the age of an object is determined by comparing its ^{14}C activity with that of living things (*Section 23.2*).
- **Particle accelerators** change one element into another (**nuclear transmutation**) by bombarding nuclei with high-energy particles (*Section 23.3*).
- **Ionizing radiation** causes chemical changes in matter. The harm caused in living matter depends on the ionizing ability and penetrating power of the radiation. Cosmic rays and decay of radioactive minerals give rise to a natural **background radiation** (*Section 23.4*).
- **Isotopes** of an element have nearly identical chemical properties. A small amount of a radioactive isotope of an element mixed with a large amount of the stable isotope can act as a **tracer** for studying reaction pathways, physical movements of substances, and medical problems (*Section 23.5*).
- The mass of a nucleus is *less* than the sum of its nucleon masses, and Einstein's equation gives the energy equivalent to this **mass difference**, which is the **nuclear binding energy**. The **binding energy per nucleon** is a measure of nuclide stability. Heavy nuclides split (**fission**) and light nuclides join (**fusion**) to release energy, thus increasing the binding energy per nucleon (*Section 23.6*).
- Nuclear power plants employ a fission **chain reaction** to create steam that generates electricity. Safety concerns center on leaks and long-term disposal of waste. Commercial energy from fusion is still in early development (*Section 23.7*).

Far below the outer fringes of its cloud of electrons lies the atom's tiny, dense core. For nearly the entire text, we have focused on an atom's electrons, treating the nucleus as their electrostatic anchor, examining the effect of its positive charge on atomic properties and, ultimately, chemical behavior. But, for the scientists probing the structure and behavior of the nucleus itself, *there* is the scene of real action, one that holds enormous potential benefit and great mystery and wonder.

Society is ambivalent about the applications of nuclear research, however. The promise of abundant energy and treatments for disease comes hand-in-hand with the threat of nuclear waste contamination, reactor accidents, and unimaginable destruction from nuclear war or terrorism. Can the power of the nucleus be harnessed for our benefit, or are the risks too great? In this chapter, we discuss the principles that can help you consider this vital question knowledgeably.

The changes that occur in atomic nuclei are strikingly different from chemical changes. In chemical reactions, electrons are shared or transferred to form *compounds*, while nuclei sit by passively, never changing their identities. In nuclear reactions, the roles are reversed, as electrons in their orbitals take part much less often, while the nuclei undergo changes that, in nearly every case, form different *elements*. Nuclear reactions are often accompanied by energy changes a million times greater than those in chemical reactions, energy changes so great that changes in mass *are* detectable. Moreover, nuclear reaction yields and rates are *not* subject to the effects of pressure, temperature, and catalysis. Table 23.1 summarizes these general differences.

Concepts & Skills to Review before studying this chapter

- discovery of the atomic nucleus (Section 2.4)
- protons, neutrons, mass number, and the ${}^A_Z\text{X}$ notation (Section 2.5)
- half-life and first-order reaction rate (Section 16.4)

Table 23.1 Comparison of Chemical and Nuclear Reactions

Chemical Reactions	Nuclear Reactions
1. One substance is converted into another, but atoms never change identity.	1. Atoms of one element typically are converted into atoms of another element.
2. Orbital electrons are involved as bonds break and form; nuclear particles do not take part.	2. Protons, neutrons, and other particles are involved; orbital electrons take part much less often.
3. Reactions are accompanied by relatively small changes in energy and no measurable changes in mass.	3. Reactions are accompanied by relatively large changes in energy and measurable changes in mass.
4. Reaction rates are influenced by temperature, concentration, catalysts, and the compound in which an element occurs.	4. Reaction rates depend on number of nuclei, but are not affected by temperature, catalysts, or, except on rare occasions, the compound in which an element occurs.

23.1 RADIOACTIVE DECAY AND NUCLEAR STABILITY

A stable nucleus remains intact indefinitely, but *the great majority of nuclei are unstable*. An unstable nucleus exhibits **radioactivity**: it spontaneously disintegrates, or *decays*, by emitting radiation. In Section 23.2, you'll see that each type of unstable nucleus has its own characteristic *rate* of radioactive decay, which can range from less than a billionth of a second to billions of years. In this section, we consider important terms and notation for nuclei, discuss some of the key events in the discovery of radioactivity, and describe the various types of radioactive decay and how to predict which type occurs for a given nucleus.

The Components of the Nucleus: Terms and Notation

Recall from Chapter 2 that the nucleus contains essentially all the atom's mass but is only about 10^{-5} times its radius (or 10^{-15} times its volume). Obviously, the nucleus is incredibly dense: about 10^{14} g/mL. *Protons* and *neutrons*, the

elementary particles that make up the nucleus, are called **nucleons**. The term **nuclide** refers to a nucleus with a particular composition, that is, with specific numbers of the two types of nucleons. Most elements occur in nature as a mixture of **isotopes**, atoms with the characteristic number of protons of the element but different numbers of neutrons. Therefore, each isotope of an element is a particular nuclide that we identify by its numbers of protons and neutrons. For example, oxygen has three naturally occurring isotopes—the most abundant contains eight protons and eight neutrons, whereas the least abundant contains eight protons and nine neutrons.

The relative mass and charge of a particle—a nucleon, another elementary particle, or a nucleus—is described by the notation ${}_Z^AX$, where X is the *symbol* for the particle, A is the *mass number*, or the total number of nucleons, and Z is the *charge* of the particle; for nuclei, A is the *sum of protons and neutrons* and Z is the *number of protons* (atomic number). In this notation, the three subatomic elementary particles are

$${}_{-1}^0\text{e} \text{ (electron), } {}_1^1\text{p} \text{ (proton), and } {}_0^1\text{n} \text{ (neutron)}$$

(A proton is also sometimes represented as ${}_1^1\text{H}^+$.) The number of neutrons (N) in a nucleus is the mass number (A) minus the atomic number (Z): $N = A - Z$. The two naturally occurring stable isotopes of chlorine, for example, have 17 protons ($Z = 17$), but one has 18 neutrons (${}_{17}^{35}\text{Cl}$, also written ${}^{35}\text{Cl}$) and the other has 20 (${}_{17}^{37}\text{Cl}$, or ${}^{37}\text{Cl}$). Nuclides can also be designated with the element name followed by the mass number, for example, chlorine-35 and chlorine-37. In naturally occurring samples of an element or its compounds, *the isotopes of the element are present in specific proportions* that can vary only very slightly. Thus, in a sample of sodium chloride (or any Cl-containing substance), 75.77% of the Cl atoms are chlorine-35 and the remaining 24.23% are chlorine-37.

To understand this chapter, it's very important that you are comfortable with nuclear notations, so please take a moment to review Sample Problem 2.4 (p. 43) and Problems 2.24 to 2.31 (pp. 65 and 66).

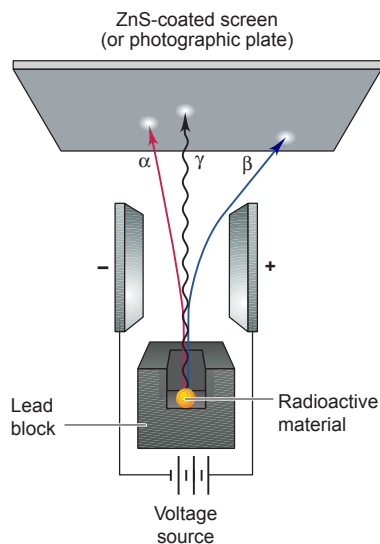


FIGURE 23.1 Three types of radioactive emissions in an electric field. Positively charged α particles curve toward the negative plate; negatively charged β particles curve toward the positive plate. (Later, we'll give these the symbol β^- .) The curvature is greater for β particles because they have much lower mass. The γ rays, uncharged high-energy photons, are unaffected by the field.

Types of Radioactive Decay; Balancing Nuclear Equations

When a nuclide of an element decays, it emits radiation and, under most circumstances, changes into a nuclide of a different element. The three natural types of radioactive emission are

- **Alpha particles** (symbolized α , ${}_2^4\alpha$, or ${}_2^4\text{He}^{2+}$) are identical to helium-4 nuclei.
- **Beta particles** (symbolized β , β^- , or sometimes ${}_{-1}^0\beta$) are high-speed electrons. (The emission of electrons from the nucleus may seem strange, but as you'll see shortly, they result from a nuclear reaction.)
- **Gamma rays** (symbolized γ , or sometimes ${}_0^0\gamma$) are very high-energy photons.

Figure 23.1 illustrates the behavior of these emissions in an electric field: α particles curve to a small extent toward the negative plate, β particles curve to a greater extent toward the positive plate, and γ rays are not affected by the electric field.

When a nuclide decays, it forms a nuclide of lower energy, and the excess energy is carried off by the emitted radiation and the recoiling nucleus. The decaying, or reactant, nuclide is called the *parent*; the product nuclide is called the *daughter*. Nuclides can decay in several ways. As each of the major types of decay is introduced (summarized in Table 23.2), we'll show examples of that type and apply the key principle used to balance nuclear reactions: *the total Z (charge, number of protons) and the total A (sum of protons and neutrons) of the reactants equal those of the products*:

$$\frac{\text{Total } A}{\text{Total } Z} \text{ Reactants} = \frac{\text{Total } A}{\text{Total } Z} \text{ Products} \quad (23.1)$$

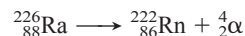
Table 23.2 Modes of Radioactive Decay*

Mode	Emission	Decay Process	Change in		
			A	Z	N
α Decay	α (${}^4_2\text{He}^{2+}$)	 α expelled	-4	-2	-2
β^- Decay†	β^- (${}^0_{-1}\beta$)	 nucleus with xp^+ and yn^0 → nucleus with $(x+1)p^+$ and $(y-1)n^0$ + ${}^0_{-1}\beta$ expelled Net: 1_0n in nucleus → 1_1p in nucleus + ${}^0_{-1}\beta$ expelled	0	+1	-1
Positron (β^+) emission†	β^+ (${}^0_1\beta$)	 nucleus with xp^+ and yn^0 → nucleus with $(x-1)p^+$ and $(y+1)n^0$ + ${}^0_1\beta$ expelled Net: 1_1p in nucleus → 1_0n in nucleus + ${}^0_1\beta$ expelled	0	-1	+1
Electron (e^-) capture (EC)†	x-ray	 nucleus with xp^+ and yn^0 → nucleus with $(x-1)p^+$ and $(y+1)n^0$ Net: ${}^0_{-1}e$ absorbed from low-energy orbital + 1_1p in nucleus → 1_0n in nucleus	0	-1	+1
Gamma (γ) emission	γ	 excited nucleus → stable nucleus + γ photon radiated	0	0	0

*Nuclear chemists consider β^- decay, positron emission, and electron capture to be three decay modes of the more general process known as beta decay (see text).

†Neutrinos (ν) or antineutrinos ($\bar{\nu}$) are also formed during the three types of beta decay. Although we will not include them in other equations in the chapter, keep in mind that antineutrinos are always expelled during β^- decay, and neutrinos are expelled during β^+ emission and e^- capture.

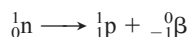
1. **Alpha (α) decay** involves the loss of an α particle from a nucleus. For each α particle emitted by the parent, A decreases by 4 and Z decreases by 2 in the daughter. Every element beyond bismuth (Bi; $Z = 83$) is radioactive and exhibits α decay. Thus, α decay is the most common means for a heavy, unstable nucleus to become more stable. For example, radium undergoes α decay to yield radon (Rn; $Z = 86$):



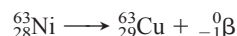
Note that the A value for Ra equals the sum of the A values for Rn and α ($226 = 222 + 4$), and that the Z value for Ra equals the sum of the Z values for Rn and α ($88 = 86 + 2$).

2. **Beta (β) decay** is a more general class of radioactive decay that includes three types: β^- decay, β^+ emission, and electron capture.

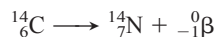
- **β^- decay** (or *negatron emission*) involves the ejection of a β^- particle from the nucleus. This change does not involve expulsion of a β^- particle that was in the nucleus; rather, a neutron is converted into a proton, which remains in the nucleus, and a β^- particle, which is expelled immediately:



As always, the totals of the A and the Z values for reactant and products are equal. Radioactive nickel-63 becomes stable copper-63 through β^- decay:

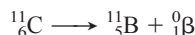


Another example is the β^- decay of carbon-14, used in radiocarbon dating:

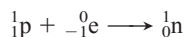


Note that β^- decay results in a product nuclide with A the same but with Z one higher (one more proton) than in the reactant nuclide. In other words, an atom of the element with the next higher atomic number is formed.

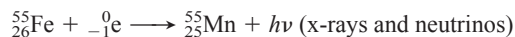
- **Positron (β^+) emission** involves the emission of a β^+ particle from the nucleus. A key idea of modern physics is that most fundamental particles have corresponding *antiparticles* with the same mass but opposite charge. (The neutrino and antineutrino are an example.) The **positron** is the antiparticle of the electron. Positron emission occurs through a process in which a proton in the nucleus is converted into a neutron, and a positron is expelled. In terms of the effect on A and Z , positron emission has the opposite effect of β^- decay: the daughter has the same A but Z is one lower (one fewer proton) than the parent. Thus, an atom of the element with the next lower atomic number forms. Carbon-11, a synthetic radioisotope, decays to a stable boron isotope through β^+ emission:



- **Electron (e^-) capture (EC)** occurs when the nucleus interacts with an electron in an orbital from a low atomic energy level. The net effect is that a proton is transformed into a neutron:

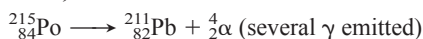


(We use the symbol “e” to distinguish an orbital electron from a beta particle, β .) The orbital vacancy is quickly filled by an electron that moves down from a higher energy level, and that process continues through still higher energy levels, with x-ray photons and neutrinos carrying off the energy difference in each step. Radioactive iron forms stable manganese through electron capture:

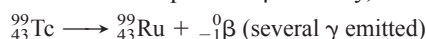


Even though the processes are different, *electron capture has the same net effect as positron emission: Z lower by 1, A unchanged.*

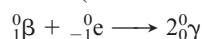
3. **Gamma (γ) emission** involves the radiation of high-energy γ photons from an excited nucleus. Just as an atom in an excited *electronic* state reduces its energy by emitting photons, usually in the UV and visible ranges (see Section 7.2), a nucleus in an excited state lowers its energy by emitting γ photons, which are of much higher energy (much shorter wavelength) than UV photons. Many nuclear processes leave the nucleus in an excited state, so γ emission accompanies many other (but mostly β) types of decay. Several γ photons (also called γ rays) of different energies can be emitted from an excited nucleus as it returns to the ground state, as in this case:



Gamma emission is common subsequent to β^- decay, as in the following:



Because γ rays have no mass or charge, γ emission does not change A or Z . Two gamma rays are emitted when a particle and an antiparticle annihilate each other. In the medical technique *positron-emission tomography* (Section 23.5), a positron and an electron annihilate each other (with all A and Z values shown):



SAMPLE PROBLEM 23.1 Writing Equations for Nuclear Reactions

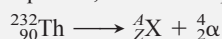
Problem Write balanced equations for the following nuclear reactions:

(a) Naturally occurring thorium-232 undergoes α decay.

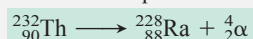
(b) Zirconium-86 undergoes electron capture.

Plan We first write a skeleton equation that includes the mass numbers, atomic numbers, and symbols of all the particles on the correct sides of the equation, showing the unknown product particle as ${}^A_Z\text{X}$. Then, because the total of mass numbers and the total of charges on the left side and the right side must be equal, we solve for A and Z , and use Z to determine X from the periodic table.

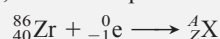
Solution (a) Writing the skeleton equation, with the α particle as a product:



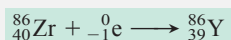
Solving for A and Z and balancing the equation: For A , $232 = A + 4$, so $A = 228$. For Z , $90 = Z + 2$, so $Z = 88$. From the periodic table, we see that the element with $Z = 88$ is radium (Ra). Thus, the balanced equation is



(b) Writing the skeleton equation, with the captured electron as a reactant:



Solving for A and Z and balancing the equation: For A , $86 + 0 = A$, so $A = 86$. For Z , $40 + (-1) = Z$, so $Z = 39$. The element with $Z = 39$ is yttrium (Y), so we have



Check Always read across superscripts and then across subscripts, with the yield arrow as an equal sign, to check your arithmetic. In part (a), for example, $232 = 228 + 4$, and $90 = 88 + 2$.

FOLLOW-UP PROBLEM 23.1 Write a balanced equation for the reaction in which a nuclide undergoes β^- decay and produces cesium-133.

Nuclear Stability and the Mode of Decay

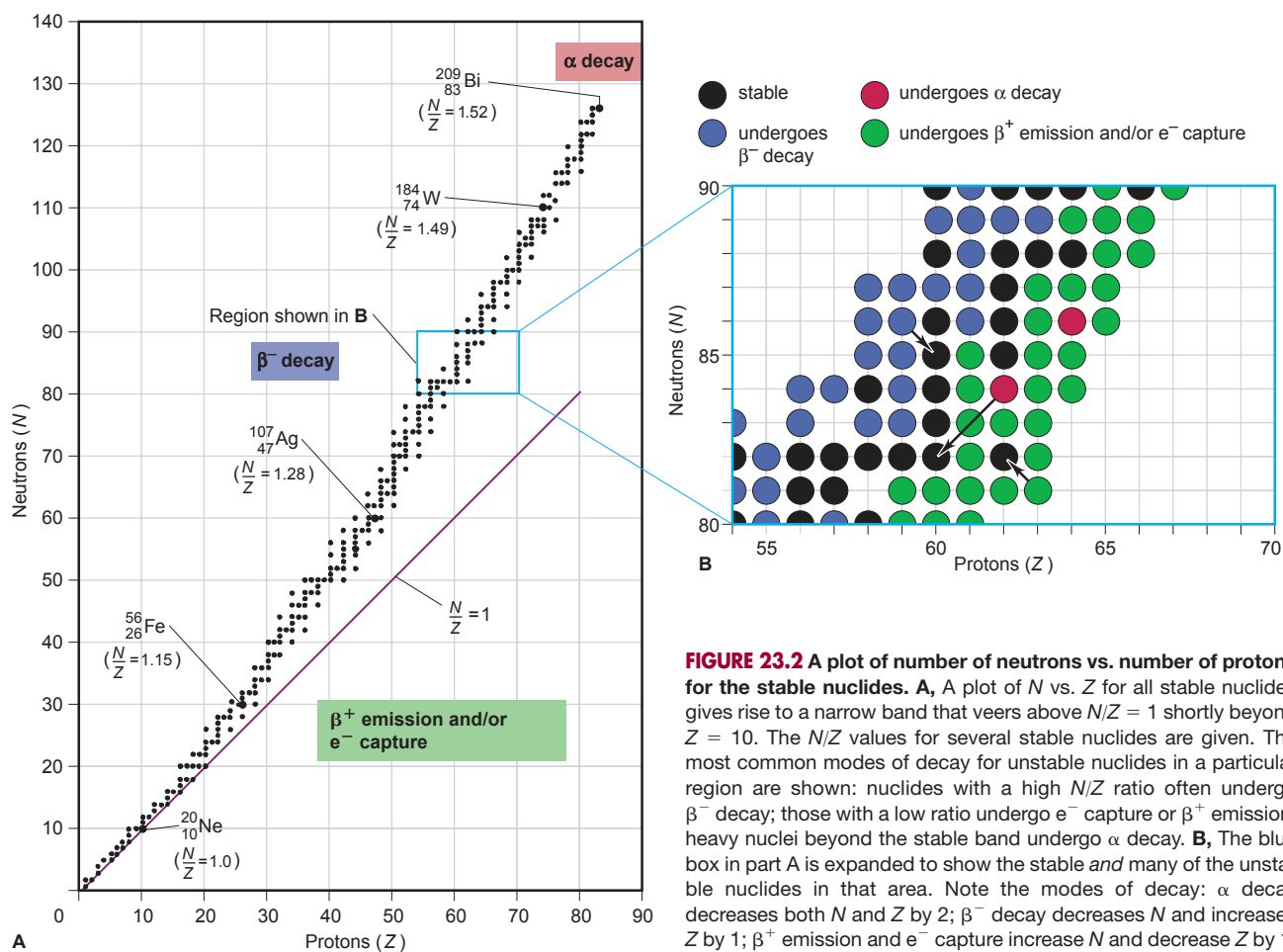
There are several ways that an unstable nuclide might decay, but can we predict how it will decay? Indeed, can we predict whether a given nuclide will decay at all? Our knowledge of the nucleus is much less complete than that of the whole atom, but some patterns emerge by observing the naturally occurring nuclides.

The Band of Stability Two key factors determine the stability of a nuclide. The first is the number of neutrons (N), the number of protons (Z), and their ratio (N/Z), which we calculate from $(A - Z)/Z$. This factor relates primarily to nuclides

that undergo one of the three modes of β decay. The second factor affecting stability is the total mass of the nuclide, which mostly relates to nuclides that undergo α decay.

Figure 23.2A is a plot of number of neutrons vs. number of protons for all stable nuclides. Note the following:

- The points form a narrow **band of stability** that gradually curves above the line for $N = Z$ ($N/Z = 1$).
- Very few stable nuclides exist with $N/Z < 1$; the only two are ${}^1_1\text{H}$ and ${}^3_2\text{He}$.
- Many lighter nuclides with $N = Z$ are stable, such as ${}^4_2\text{He}$, ${}^{12}_6\text{C}$, ${}^{16}_8\text{O}$, and ${}^{20}_{10}\text{Ne}$; the heaviest of these is ${}^{40}_{20}\text{Ca}$. Thus, for lighter nuclides, one neutron for each proton ($N = Z$) is enough to provide stability.
- The N/Z ratio of stable nuclides gradually increases as Z increases. A few examples are noted on the figure: for ${}^{56}_{26}\text{Fe}$, $N/Z = 1.15$; for ${}^{107}_{47}\text{Ag}$, $N/Z = 1.28$; for ${}^{184}_{74}\text{W}$, $N/Z = 1.49$, and, finally, for ${}^{209}_{83}\text{Bi}$, $N/Z = 1.52$. Thus, for heavier stable nuclides, $N > Z$ ($N/Z > 1$), and N increases faster than Z . As we discuss below, if N/Z of a nuclide is either too high (above the band) or not high enough (below the band), the nuclide is unstable and undergoes one of the three modes of beta decay.
- All nuclides with $Z > 83$ are unstable. Thus, the largest members of main groups 1A(1) through 8A(18), actinium and the actinides ($Z = 89$ –103), and the other elements of the fourth ($6d$) transition series ($Z = 104$ –112), are radioactive and (as discussed below) undergo α decay.



Stability and Nuclear Structure The oddness or evenness of N and Z values is related to some important patterns of nuclear stability. Two interesting points become apparent when we classify the known stable nuclides:

- Elements with an even Z (number of protons) usually have a larger number of stable nuclides than elements with an odd Z . Table 23.3 demonstrates this point for cadmium ($Z = 48$) through xenon ($Z = 54$).
- Well over half the stable nuclides have *both* even N and even Z . Only four nuclides with odd N and odd Z are stable: ${}^2_1\text{H}$, ${}^6_3\text{Li}$, ${}^{10}_5\text{B}$, and ${}^{14}_7\text{N}$.

To explain the stability from even values of N and Z , one model of nuclear structure postulates that protons and neutrons lie in *nucleon energy levels*, and that greater stability results from the *pairing of spins* of like nucleons. (Note the analogy to electron energy levels and the stability from pairing of electron spins.)

Just as noble gases—with 2, 10, 18, 36, 54, and 86 electrons—are exceptionally stable because they have filled *electron* levels, nuclides with N or Z values of 2, 8, 20, 28, 50, 82 (and $N = 126$) are exceptionally stable. These so-called *magic numbers* are thought to correspond to the numbers of protons or neutrons in filled *nucleon* levels. A few examples are ${}^{50}_{22}\text{Ti}$ ($N = 28$), ${}^{88}_{38}\text{Sr}$ ($N = 50$), and the ten stable nuclides of tin ($Z = 50$). Some extremely stable nuclides have doubly magic numbers: ${}^4_2\text{He}$, ${}^{16}_8\text{O}$, ${}^{40}_{20}\text{Ca}$, and ${}^{208}_{82}\text{Pb}$ ($N = 126$).

Table 23.3 Number of Stable Nuclides for Elements 48 to 54*

Element	Atomic No. (Z)	No. of Nuclides
Cd	48	8
In	49	2
Sn	50	10
Sb	51	2
Te	52	8
I	53	1
Xe	54	9

*Even Z shown in boldface.

SAMPLE PROBLEM 23.2 Predicting Nuclear Stability

Problem Which of the following nuclides would you predict to be stable and which radioactive: (a) ${}^{18}_{10}\text{Ne}$; (b) ${}^{32}_{16}\text{S}$; (c) ${}^{236}_{90}\text{Th}$; (d) ${}^{123}_{56}\text{Ba}$? Explain.

Plan In order to evaluate the stability of each nuclide, we determine the N and Z values, the N/Z ratio from $(A - Z)/Z$, the value of Z , stable N/Z ratios (from Figure 23.2), and whether Z and N are even or odd.

Solution (a) **Radioactive.** This nuclide has $N = 8$ ($18 - 10$) and $Z = 10$, so $N/Z = \frac{18 - 10}{10} = 0.8$. Except for hydrogen-1 and helium-3, no nuclides with $N < Z$ are stable; despite even N and Z , this nuclide has too few neutrons to be stable.

(b) **Stable.** This nuclide has $N = Z = 16$, so $N/Z = 1.0$. With $Z < 20$ and even N and Z , this nuclide is most likely stable.

(c) **Radioactive.** This nuclide has $Z = 90$, and every nuclide with $Z > 83$ is radioactive.

(d) **Radioactive.** This nuclide has $N = 67$ and $Z = 56$, so $N/Z = 1.20$. For Z values of 55 to 60, Figure 23.2A shows $N/Z \geq 1.3$, so this nuclide has too few neutrons to be stable.

Check By consulting a table of isotopes, such as the one in the *CRC Handbook of Chemistry and Physics*, we find that our predictions are correct.

FOLLOW-UP PROBLEM 23.2 Why is ${}^{31}_{15}\text{P}$ stable but ${}^{30}_{15}\text{P}$ unstable?

Predicting the Mode of Decay An unstable nuclide generally decays in a mode that shifts its N/Z ratio toward the band of stability. This fact is illustrated in Figure 23.2B, which expands a small region of Figure 23.2A to show the stable *and* many of the unstable nuclides in that region, as well as their modes of decay.

Note the following points:

1. **Neutron-rich nuclides.** Nuclides with too many neutrons for stability (a high N/Z) lie above the band of stability. They undergo β^- decay, which converts a neutron into a proton, thus reducing the value of N/Z .
2. **Proton-rich nuclides.** Nuclides with too many protons for stability (a low N/Z) lie below the band. They undergo β^+ emission and/or e^- capture, both of which convert a proton into a neutron, thus increasing the value of N/Z . (The rate of e^- capture increases with Z , so β^+ emission is more common among lighter elements and e^- capture more common among heavier elements.)
3. **Heavy nuclides.** Nuclides with $Z > 83$ are too heavy to be stable and undergo α decay, which reduces their Z and N values by two units per emission.

If we have the information in Figure 23.2, predicting the mode of decay of an unstable nuclide is just a matter of comparing its N/Z ratio with those in the nearby region of the band of stability. But, even when Figure 23.2 is not available, we can often make an educated guess at the mode of decay. The atomic mass of an element represents the weighted average of its naturally occurring isotopes, so we would expect the mass number A of a *stable* nuclide to be relatively close to the atomic mass. If an *unstable* nuclide of the element (given Z) has an A value higher than the atomic mass, it is neutron rich and will probably decay by β^- emission. If, on the other hand, the unstable nuclide has an A value lower than the atomic mass, it is proton rich and will probably decay by β^+ emission and/or e^- capture. In the next sample problem, we compare the mass number with the atomic mass to help us predict the mode of decay of an unstable nuclide and then check the prediction with the N/Z values in Figure 23.2.

SAMPLE PROBLEM 23.3 Predicting the Mode of Nuclear Decay

Problem Use the atomic mass of the element to predict the mode(s) of decay of the following radioactive nuclides: (a) $^{12}_5\text{B}$; (b) $^{234}_{92}\text{U}$; (c) $^{81}_{33}\text{As}$; (d) $^{127}_{57}\text{La}$.

Plan If the nuclide is too heavy to be stable ($Z > 83$), it undergoes α decay. For other cases, we use the Z value to locate the element in the periodic table and obtain its atomic mass. If the mass number of the nuclide is higher than the atomic mass, the nuclide has too many neutrons: N too high $\Rightarrow \beta^-$ decay. If the mass number is lower than the atomic mass, the nuclide has too many protons: Z too high $\Rightarrow \beta^+$ emission and/or e^- capture.

Solution (a) This nuclide has $Z = 5$, which is boron (B), and the atomic mass is 10.81. The nuclide's A value of 12 is higher than its atomic mass, so this nuclide is neutron rich. It will probably undergo β^- decay lowering N to 6 and raising Z to 6 to form the stable $^{12}_6\text{C}$.

(b) This nuclide has $Z = 92$, so it will undergo α decay and decrease its total mass.

(c) This nuclide has $Z = 33$, which is arsenic (As), and the atomic mass is 74.92. The A value of 81 is higher, so this nuclide is neutron rich and will probably undergo β^- decay.

(d) This nuclide has $Z = 57$, which is lanthanum (La), and the atomic mass is 138.9. The A value of 127 is lower, so this nuclide is proton rich and will probably undergo β^+ emission and/or e^- capture.

Check To confirm our predictions in (a), (c), and (d), let's compare each nuclide's N/Z ratio to those in the band of stability. (a) This nuclide has $N = 7$ and $Z = 5$, so $N/Z = 1.40$, which is too high for this region of the band; it will undergo β^- decay. (c) This nuclide has $N = 48$ and $Z = 33$, so $N/Z = 1.45$, which is too high for this region of the band; it undergoes β^- decay. (d) This nuclide has $N = 70$ and $Z = 57$, so $N/Z = 1.23$, which is too low for this region of the band; it undergoes β^+ emission and/or e^- capture. Our predictions based on N/Z values were the same as those based on atomic mass.

Comment Both possible modes of decay are observed for the nuclide in part (d).

FOLLOW-UP PROBLEM 23.3 Use the A value for the nuclide and the atomic mass in the periodic table to predict the mode of decay of (a) $^{61}_{26}\text{Fe}$; (b) $^{241}_{95}\text{Am}$.

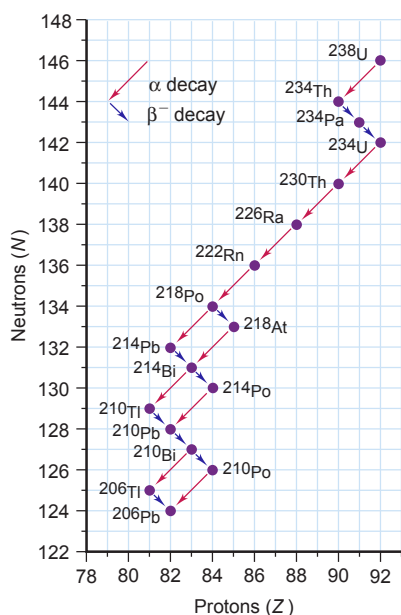


FIGURE 23.3 The ^{238}U decay series. Uranium-238 (top right) undergoes a series of 14 steps involving α or β^- decay until lead-206 forms (bottom left).

Decay Series A parent nuclide may undergo a series of decay steps before a stable daughter nuclide forms. The succession of steps is called a **decay series**, or **disintegration series**, and is typically depicted on a gridlike display. Figure 23.3 shows the decay series from uranium-238 to lead-206. Numbers of neutrons (N) are plotted against numbers of protons (Z) to form the grid, which displays a series of α and β^- decays. The typical zigzag pattern arises because $N > Z$, which means that α decay, which reduces both N and Z by two units, decreases Z by a slightly greater percentage than it does N . Therefore, α decays result in neutron-rich daughters, which undergo β^- decay to gain more stability. Note that a given nuclide can undergo both types of decay. (Gamma emission accompanies many of these steps but does not affect the type of nuclide.) This series is one of three that occur in nature. All end with isotopes of lead whose nuclides all have one

($Z = 82$) or two ($N = 126$, $Z = 82$) magic numbers. A second series begins with uranium-235 and ends with lead-207, and a third begins with thorium-232 and ends with lead-208. (Neptunium-237 began a fourth series, but its half-life is so much less than the age of Earth that only traces of it remain today.)

SECTION 23.1 SUMMARY

Nuclear reactions are normally not affected by reaction conditions or chemical composition and release much more energy than chemical reactions. • To become more stable, a radioactive nuclide may emit α particles (${}^4_2\text{He}$ nuclei), β particles (β^- or ${}^0_{-1}\beta$; high-speed electrons), positrons (β^+ or ${}^0_{+1}\beta$), or γ rays (high-energy photons) or may capture an orbital electron (${}^0_0\text{e}$). • A narrow band of neutron-to-proton ratios (N/Z) includes those of all the stable nuclides. • Certain “magic numbers” of neutrons and protons are associated with very stable nuclides. • By comparing a nuclide’s mass number with the atomic mass and its N/Z ratio with those in the band of stability, we can predict that, in general, neutron-rich nuclides undergo β^- decay and proton-rich nuclides undergo β^+ emission and/or e^- capture. Heavy nuclides ($Z > 83$) undergo α decay. • Three naturally occurring decay series all end in isotopes of lead.

23.2 THE KINETICS OF RADIOACTIVE DECAY

Chemical and nuclear systems both tend toward maximum stability. Just as the concentrations in a chemical system change in a predictable direction to give a stable equilibrium ratio, the type and number of nucleons in an unstable nucleus change in a predictable direction to give a stable N/Z ratio. As you know, however, the tendency of a chemical system to become more stable tells nothing about how long that process will take, and the same holds true for nuclear systems. In this section, we examine the kinetics of nuclear change; later, we’ll examine the energetics of nuclear change.

The Rate of Radioactive Decay

Radioactive nuclei decay at a characteristic rate, regardless of the chemical substance in which they occur. The *decay rate*, or **activity** ($\mathcal{A}$), of a radioactive sample is the change in number of nuclei ($\mathcal{N}$) divided by the change in time (t). As we saw with chemical reaction rates, because the number of nuclei is *decreasing*, a minus sign precedes the expression for the decay rate:

$$\text{Decay rate } (\mathcal{A}) = -\frac{\Delta \mathcal{N}}{\Delta t}$$

The SI unit of radioactivity is the **becquerel (Bq)**; it is defined as one disintegration per second (d/s): $1 \text{ Bq} = 1 \text{ d/s}$. A much larger and more common unit of radioactivity is the **curie (Ci)**. Originally, the curie was defined as the number of disintegrations per second in 1 g of radium-226, but it is now a fixed quantity:

$$1 \text{ Ci} = 3.70 \times 10^{10} \text{ d/s} \quad (23.2)$$

Because the curie is so large, the millicurie (mCi) and microcurie (μCi) are commonly used. We often express the radioactivity of a sample in terms of *specific activity*, the decay rate per gram.

An activity is meaningful only when we consider the large number of nuclei in a macroscopic sample. Suppose there are 1×10^{15} radioactive nuclei of a particular type in a sample and they decay at a rate of 10% per hour. Although any particular nucleus in the sample might decay in a microsecond or in a million hours, the *average* of all decays results in 10% of the entire collection of nuclei disintegrating each hour. During the first hour, 10% of the *original* number, or 1×10^{14} nuclei, will decay. During the next hour, 10% of the remaining 9×10^{14} nuclei, or 9×10^{13} nuclei, will decay. During the next hour, 10% of those remaining



Animation: Radioactive Decay

will decay, and so forth. Thus, for a large collection of radioactive nuclei, *the number decaying per unit time is proportional to the number present*:

$$\text{Decay rate } (\mathcal{A}) \propto \mathcal{N} \quad \text{or} \quad \mathcal{A} = k\mathcal{N}$$

where k is called the **decay constant** and is characteristic of each type of nuclide. The larger the value of k , the higher is the decay rate: **larger $k \Rightarrow$ higher $\mathcal{A}$** .

Combining the two rate expressions just given, we obtain

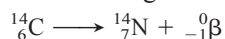
$$\mathcal{A} = -\frac{\Delta\mathcal{N}}{\Delta t} = k\mathcal{N} \quad (23.3)$$

Note that the activity depends only on $\mathcal{N}$ raised to the first power (and on the constant value of k). Therefore, *radioactive decay is a first-order process* (see Section 16.3). The only difference in the case of nuclear decay is that we consider the *number* of nuclei rather than their concentration.



Animation: Half-Life

Half-Life of Radioactive Decay Decay rates are also commonly expressed in terms of the fraction of nuclei that decay over a given time interval. The **half-life** ($t_{1/2}$) of a nuclide is the time it takes for half the nuclei present in a sample to decay. *The number of nuclei remaining is halved after each half-life.* Thus, half-life has the same meaning for a nuclear change as for a chemical change (Section 16.4). Figure 23.4 shows the decay of carbon-14, which has a half-life of 5730 years, in terms of number of ^{14}C nuclei remaining:



We can also consider the half-life in terms of mass of substance. As ^{14}C decays, the mass of ^{14}C decreases while the mass of ^{14}N increases. If we start with 1.0 g of ^{14}C , half that mass of ^{14}C (0.50 g) will be left after 5730 years, half of that mass (0.25 g) after another 5730 years, and so on. The activity depends on the number of nuclei, so the activity is halved after each succeeding half-life as well.

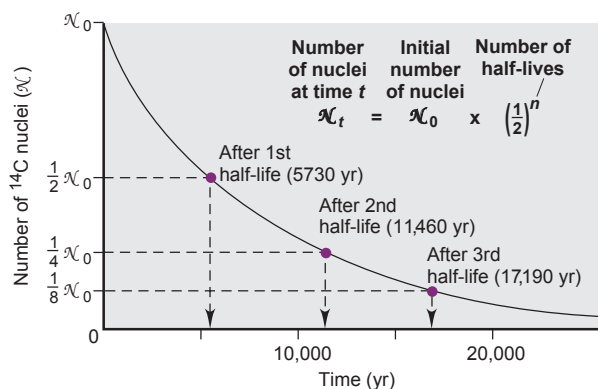
We determine the half-life of a nuclear reaction from its rate constant. Rearranging Equation 23.3 and integrating over time gives an expression for finding the number of nuclei remaining after a given time t , $\mathcal{N}_t$:

$$\ln \frac{\mathcal{N}_t}{\mathcal{N}_0} = -kt \quad \text{or} \quad \mathcal{N}_t = \mathcal{N}_0 e^{-kt} \quad \text{and} \quad \ln \frac{\mathcal{N}_0}{\mathcal{N}_t} = kt \quad (23.4)$$

where $\mathcal{N}_0$ is the number of nuclei at $t = 0$. (Note the similarity to Equation 16.4, p. 521.) To calculate the half-life ($t_{1/2}$), we set $\mathcal{N}_t$ equal to $\frac{1}{2}\mathcal{N}_0$ and solve for $t_{1/2}$:

$$\ln \frac{\mathcal{N}_0}{\frac{1}{2}\mathcal{N}_0} = kt_{1/2} \quad \text{so} \quad t_{1/2} = \frac{\ln 2}{k} \quad (23.5)$$

FIGURE 23.4 Decrease in number of ^{14}C nuclei over time. A plot of number of ^{14}C nuclei vs. time gives a downward-sloping curve. In each half-life (5730 years), half the ^{14}C nuclei present undergo decay. A plot of mass of ^{14}C (or any other variable proportional to number of nuclei) vs. time has the same shape.



Exactly analogous to the half-life of a first-order chemical change, *this half-life is not dependent on the number of nuclei and is inversely related to the decay constant*:

$$\text{large } k \Rightarrow \text{short } t_{1/2} \quad \text{and} \quad \text{small } k \Rightarrow \text{long } t_{1/2}$$

The decay constants and half-lives of radioactive nuclides vary over a very wide range, even for the nuclides of a given element (Table 23.4).

Table 23.4 Decay Constants (k) and Half-Lives ($t_{1/2}$) of Beryllium Isotopes

Nuclide	k	$t_{1/2}$
${}^7_4\text{Be}$	$1.30 \times 10^{-2}/\text{day}$	53.3 days
${}^8_4\text{Be}$	$1.0 \times 10^{16}/\text{s}$	6.7×10^{-17} s
${}^9_4\text{Be}$	Stable	
${}^{10}_4\text{Be}$	$4.3 \times 10^{-7}/\text{yr}$	1.6×10^6 yr
${}^{11}_4\text{Be}$	$5.02 \times 10^{-2}/\text{s}$	13.8 s

SAMPLE PROBLEM 23.4 Finding the Number of Radioactive Nuclei

Problem Strontium-90 is a radioactive byproduct of nuclear reactors that behaves biologically like calcium, the element above it in Group 2A(2). When ${}^{90}\text{Sr}$ is ingested by mammals, it is found in their milk and eventually in the bones of those drinking the milk. If a sample of ${}^{90}\text{Sr}$ has an activity of 1.2×10^{12} d/s, what are the activity and the fraction of nuclei that have decayed after 59 yr ($t_{1/2}$ of ${}^{90}\text{Sr} = 29$ yr)?

Plan The fraction of nuclei that have decayed is the change in number of nuclei, expressed as a fraction of the starting number. The activity of the sample ($\mathcal{A}$) is proportional to the number of nuclei ($\mathcal{N}$), so we know that

$$\text{Fraction decayed} = \frac{\mathcal{N}_0 - \mathcal{N}_t}{\mathcal{N}_0} = \frac{\mathcal{A}_0 - \mathcal{A}_t}{\mathcal{A}_0}$$

We are given $\mathcal{A}_0$ (1.2×10^{12} d/s), so we find $\mathcal{A}_t$ from the integrated form of the first-order rate equation (Equation 23.4), in which t is 59 yr. To solve that equation, we first need k , which we can calculate from the given $t_{1/2}$ (29 yr).

Solution Calculating the decay constant k :

$$t_{1/2} = \frac{\ln 2}{k} \quad \text{so} \quad k = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{29 \text{ yr}} = 0.024 \text{ yr}^{-1}$$

Applying Equation 23.4 to calculate $\mathcal{A}_t$, the activity remaining at time t :

$$\ln \frac{\mathcal{N}_0}{\mathcal{N}_t} = \ln \frac{\mathcal{A}_0}{\mathcal{A}_t} = kt \quad \text{or} \quad \ln \mathcal{A}_0 - \ln \mathcal{A}_t = kt$$

$$\begin{aligned} \text{So,} \quad \ln \mathcal{A}_t &= -kt + \ln \mathcal{A}_0 = -(0.024 \text{ yr}^{-1} \times 59 \text{ yr}) + \ln (1.2 \times 10^{12} \text{ d/s}) \\ \ln \mathcal{A}_t &= -1.4 + 27.81 = 26.4 \end{aligned}$$

$$\mathcal{A}_t = 2.9 \times 10^{11} \text{ d/s}$$

(All the data contain two significant figures, so we retained two in the answer.) Calculating the fraction decayed:

$$\text{Fraction decayed} = \frac{\mathcal{A}_0 - \mathcal{A}_t}{\mathcal{A}_0} = \frac{1.2 \times 10^{12} \text{ d/s} - 2.9 \times 10^{11} \text{ d/s}}{1.2 \times 10^{12} \text{ d/s}} = 0.76$$

Check The answer is reasonable: t is about 2 half-lives, so $\mathcal{A}_t$ should be about $\frac{1}{4}\mathcal{A}_0$, or about 0.3×10^{12} ; therefore, the activity should have decreased by about $\frac{3}{4}$.

Comment 1. A useful substitution of Equation 23.4 for finding $\mathcal{A}_t$, the activity at time t , is $\mathcal{A}_t = \mathcal{A}_0 e^{-kt}$.

2. Another way to find the fraction of activity (or nuclei) remaining incorporates the number of half-lives ($t/t_{1/2}$). By combining Equations 23.4 and 23.5 and substituting $(\ln 2)/t_{1/2}$ for k , we obtain

$$\ln \frac{\mathcal{N}_0}{\mathcal{N}_t} = \left(\frac{\ln 2}{t_{1/2}} \right) t = \frac{t}{t_{1/2}} \ln 2 = \ln 2^{t/t_{1/2}}$$

$$\text{Inverting the ratio gives} \quad \ln \frac{\mathcal{N}_t}{\mathcal{N}_0} = \ln \left(\frac{1}{2} \right)^{t/t_{1/2}}$$

Taking the antilog gives

$$\text{Fraction remaining} = \frac{\mathcal{N}_t}{\mathcal{N}_0} = \left(\frac{1}{2} \right)^{t/t_{1/2}} = \left(\frac{1}{2} \right)^{59/29} = 0.24$$

$$\text{So,} \quad \text{Fraction decayed} = 1.00 - 0.24 = 0.76$$

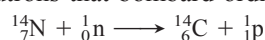
FOLLOW-UP PROBLEM 23.4 Sodium-24 ($t_{1/2} = 15$ h) is used to study blood circulation. If a patient is injected with an aqueous solution of ${}^{24}\text{NaCl}$ whose activity is 2.5×10^9 d/s, how much activity is in the patient's body and excreted fluids after 4.0 days?

Radioisotopic Dating

The historical record fades rapidly with time and virtually disappears for events of more than a few thousand years ago. Much of our understanding of prehistory comes from a technique called **radioisotopic dating**, which uses **radioisotopes** to determine the age of an object. The method supplies data in fields as diverse as art history, archeology, geology, and paleontology.

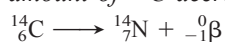
The technique of *radiocarbon dating*, for which the American chemist Willard F. Libby won the Nobel Prize in chemistry in 1960, is based on measuring the amounts of ^{14}C and ^{12}C in materials of biological origin. The accuracy of the method falls off after about six half-lives of ^{14}C ($t_{1/2} = 5730$ yr), so it is used to date objects up to about 36,000 years old.

Here is how the method works. High-energy cosmic rays, consisting mainly of protons, enter the atmosphere from outer space and initiate a cascade of nuclear reactions; some produce neutrons that bombard ordinary ^{14}N atoms to form ^{14}C :



Through the competing processes of formation and radioactive decay, the amount of ^{14}C in the atmosphere has remained nearly constant.

The ^{14}C atoms combine with O_2 , diffuse throughout the lower atmosphere, and enter the total carbon pool as gaseous $^{14}\text{CO}_2$ and aqueous $\text{H}^{14}\text{CO}_3^-$. They mix with ordinary $^{12}\text{CO}_2$ and $\text{H}^{12}\text{CO}_3^-$, reaching a constant $^{12}\text{C}/^{14}\text{C}$ ratio of about $10^{12}/1$. Carbon atoms in the CO_2 are taken up by plants during photosynthesis, and then taken up and excreted by animals that eat the plants. Thus, the $^{12}\text{C}/^{14}\text{C}$ ratio of a living organism is the same as the ratio in the environment. When an organism dies, however, it no longer absorbs or releases CO_2 , so the $^{12}\text{C}/^{14}\text{C}$ ratio steadily increases because *the amount of ^{14}C decreases as it decays*:



The difference between the $^{12}\text{C}/^{14}\text{C}$ ratio in a dead organism and the ratio in living organisms reflects the time elapsed since the organism died.

As you saw in Sample Problem 23.4, the first-order rate equation can be expressed in terms of a ratio of activities:

$$\ln \frac{\mathcal{N}_0}{\mathcal{N}_t} = \ln \frac{\mathcal{A}_0}{\mathcal{A}_t} = kt$$

We use this expression in radiocarbon dating, where $\mathcal{A}_0$ is the activity in a living organism and $\mathcal{A}_t$ is the activity in the object whose age is unknown. Solving for t gives the age of the object:

$$t = \frac{1}{k} \ln \frac{\mathcal{A}_0}{\mathcal{A}_t} \quad (23.6)$$

To determine the ages of more ancient objects or of objects that do not contain carbon, different radioisotopes are measured. For example, by comparing the ratio of ^{238}U ($t_{1/2} = 4.5 \times 10^9$ yr) to its final decay product, ^{206}Pb , geochemists found that the oldest known surface rocks on Earth—granite in western Greenland—are about 3.7 billion years old. The ratio $^{238}\text{U}/^{206}\text{Pb}$ in samples from meteorites gives 4.65 billion years for the age of the Solar System, and thus Earth.

SAMPLE PROBLEM 23.5 Applying Radiocarbon Dating

Problem The charred bones of a sloth in a cave in Chile represent the earliest evidence of human presence at the southern tip of South America. A sample of the bone has a specific activity of 5.22 disintegrations per minute per gram of carbon (d/min·g). If the $^{12}\text{C}/^{14}\text{C}$ ratio for living organisms results in a specific activity of 15.3 d/min·g, how old are the bones ($t_{1/2}$ of $^{14}\text{C} = 5730$ yr)?

Plan We first calculate k from the given $t_{1/2}$ (5730 yr). Then we apply Equation 23.6 to find the age (t) of the bones, using the given activities of the bones ($\mathcal{A}_t = 5.22$ d/min·g) and of a living organism ($\mathcal{A}_0 = 15.3$ d/min·g).

Solution Calculating k for ^{14}C decay:

$$k = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{5730 \text{ yr}} = 1.21 \times 10^{-4} \text{ yr}^{-1}$$

Calculating the age (t) of the bones:

$$t = \frac{1}{k} \ln \frac{\mathcal{A}_0}{\mathcal{A}_t} = \frac{1}{1.21 \times 10^{-4} \text{ yr}^{-1}} \ln \left(\frac{15.3 \text{ d/min}\cdot\text{g}}{5.22 \text{ d/min}\cdot\text{g}} \right) = 8.89 \times 10^3 \text{ yr}$$

The bones are about 8900 years old.

Check The activity of the bones is between $\frac{1}{2}$ and $\frac{1}{4}$ the activity of a living organism, so the age should be between one and two half-lives (5730 to 11,460 yr).

FOLLOW-UP PROBLEM 23.5 A sample of wood from an Egyptian mummy case has a specific activity of 9.41 d/min·g. How old is the case?

SECTION 23.2 SUMMARY

The decay rate (activity) of a sample is proportional to the number of radioactive nuclei. Nuclear decay is a first-order process, so the half-life is a constant for each nuclide. • Radioisotopic methods, such as ^{14}C dating, determine the ages of objects by measuring the ratio of specific isotopes in the sample.

23.3 NUCLEAR TRANSMUTATION: INDUCED CHANGES IN NUCLEI

The alchemists' dream of changing base metals into gold was never realized, but in the early 20th century, nuclear physicists found that they *could* change one element into another. Research into **nuclear transmutation**, the *induced* conversion of one nucleus into another, was closely linked with research into atomic structure and led to the discovery of the neutron and to the production of artificial radioisotopes. Later, high-energy bombardment of nuclei in particle accelerators began a scientific endeavor, which continues to this day, of creating many new nuclides and a growing number of new elements.

During the 1930s and 1940s, researchers probing the nucleus bombarded elements with neutrons, α particles, protons, and **deuterons** (nuclei of the stable hydrogen isotope deuterium, ^2H). Neutrons are especially useful as projectiles because they have no charge and thus are not repelled as they approach a target nucleus. The other particles are all positive, so early researchers found it difficult to give them enough energy to overcome their repulsion by the target nuclei. Beginning in the 1930s, however, **particle accelerators** were invented to impart high kinetic energies to particles by placing them in an electric field, usually in combination with a magnetic field.

A major advance was the *linear accelerator*, a series of separated tubes of increasing length that, through a source of alternating voltage, change their charge from positive to negative in synchrony with the movement of the particle through them (Figure 23.5). A proton, for example, exits the first tube just when that tube becomes positive and the next tube negative. Repelled by the first tube and

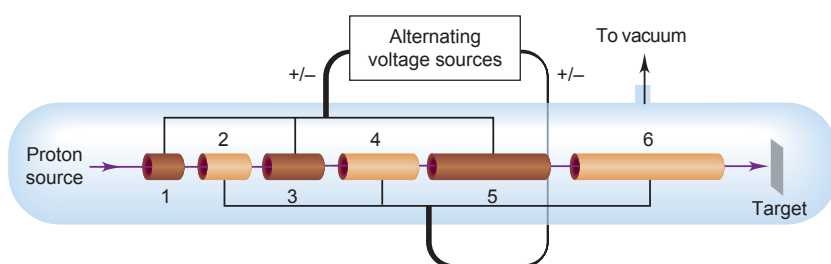
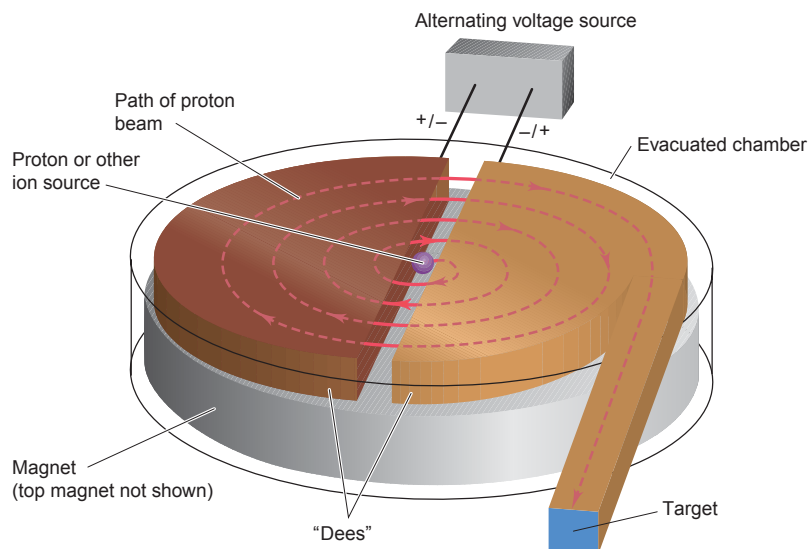


FIGURE 23.5 Schematic of a linear accelerator. The voltage of each tubular section is alternated, so that the positively charged particle (a proton here) is repelled from the section it is leaving and attracted to the section it is entering. As a result, the particle's speed is continually increased.

FIGURE 23.6 Schematic of a cyclotron accelerator. The magnets lie within an evacuated chamber above and below two “dees,” open, D-shaped electrodes that act like the tubes in the linear design. The particle is accelerated as it passes from one dee, which is momentarily positive, to the other, which is momentarily negative. Its speed and radius increase until it is deflected toward the target nucleus.



attracted by the second, the proton accelerates across the gap between them. The process is performed in stages to achieve high particle energies without having to apply a single high voltage. A 40-ft linear accelerator with 46 tubes, built in California after World War II, accelerated protons to speeds several million times faster than earlier accelerators. The *cyclotron* (Figure 23.6), invented by E. O. Lawrence in 1930, applies the principle of the linear accelerator but uses electromagnets to give the particle a spiral path, thus saving space.

Scientists use accelerators for many applications, from producing radioisotopes used in medical applications to studying the fundamental nature of matter. Perhaps the most specific application for chemists is the synthesis of **transuranium elements**, those with atomic numbers higher than uranium, the heaviest naturally occurring element. Some reactions that were used to form several of these elements appear in Table 23.5.

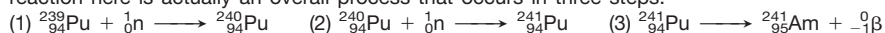
SECTION 23.3 SUMMARY

One nucleus can be transmuted to another through bombardment with high-energy particles. • Accelerators increase the kinetic energy of particles. They are used to produce transuranium elements and radioisotopes for medical use.

Table 23.5 Formation of Some Transuranium Nuclides*

Reaction							Half-life of Product
$^{239}_{94}\text{Pu}$	+	^1_0n	$\longrightarrow$	$^{241}_{95}\text{Am}$	+	$^0_{-1}\beta$	432 yr
$^{239}_{94}\text{Pu}$	+	$^4_2\alpha$	$\longrightarrow$	$^{242}_{96}\text{Cm}$	+	^1_0n	163 days
$^{241}_{95}\text{Am}$	+	$^4_2\alpha$	$\longrightarrow$	$^{243}_{97}\text{Bk}$	+	^1_0n	4.5 h
$^{242}_{96}\text{Cm}$	+	$^4_2\alpha$	$\longrightarrow$	$^{245}_{98}\text{Cf}$	+	^1_0n	45 min
$^{253}_{99}\text{Es}$	+	$^4_2\alpha$	$\longrightarrow$	$^{256}_{101}\text{Md}$	+	^1_0n	76 min
$^{243}_{95}\text{Am}$	+	$^{18}_8\text{O}$	$\longrightarrow$	$^{256}_{103}\text{Lr}$	+	5^1_0n	28 s

*Like chemical reactions, nuclear reactions may occur in several steps. For example, the first reaction here is actually an overall process that occurs in three steps:

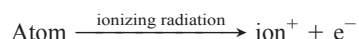


23.4 THE EFFECTS OF NUCLEAR RADIATION ON MATTER

In 1986, an accident at the Chernobyl nuclear facility in the former Soviet Union released radioactivity that, according to the World Health Organization, will cause thousands of cancer deaths. In the same year, isotopes used in medical treatment emitted radioactivity that prevented thousands of cancer deaths. In this section and Section 23.5, we examine radioactivity's harmful and beneficial effects.

The key to both of these effects is that *nuclear changes cause chemical changes in surrounding matter*. In other words, even though the nucleus of an atom may undergo a reaction with little or no involvement of the atom's electrons, the emissions from that reaction *do* affect the electrons of nearby atoms.

Virtually all radioactivity causes **ionization** in surrounding matter, as the emissions collide with atoms and dislodge electrons:



From each ionization event, a cation and a free electron result, and the number of such *cation-electron pairs* produced is directly related to the energy of the incoming **ionizing radiation**.

Effects of Ionizing Radiation on Living Matter

Ionizing radiation has a destructive effect on living tissue, and if the ionized atom is part of a key biological macromolecule or cell membrane, the results can be devastating to the cell and perhaps the organism.

Units of Radiation Dose and Its Effects To measure the effects of ionizing radiation, we need a unit for radiation dose. Units of radioactive decay, such as the becquerel and curie, measure the number of decay events in a given time but not their energy or absorption by matter. The number of cation-electron pairs produced in a given amount of living tissue is a measure of the energy absorbed by the tissue. The SI unit is the **gray (Gy)**, equal to 1 joule of energy absorbed per kilogram of body tissue: $1 \text{ Gy} = 1 \text{ J/kg}$. A more widely used unit is the **rad (radiation-absorbed dose)**, which is equal to 0.01 Gy:

$$1 \text{ rad} = 0.01 \text{ J/kg} = 0.01 \text{ Gy}$$

To measure actual tissue damage, we must account for differences in the strength of the radiation, the exposure time, and the type of tissue. To do this, we multiply the number of rads by a *relative biological effectiveness* (RBE) factor, which depends on the effect of a given type of radiation on a given tissue or body part. The product is the **rem (roentgen equivalent for man)**, the unit of radiation dosage equivalent to a given amount of tissue damage in a human:

$$\text{no. of rems} = \text{no. of rads} \times \text{RBE}$$

Doses are often expressed in millirems (10^{-3} rem). The SI unit for dosage equivalent is the **sievert (Sv)**. It is defined in the same way as the rem but with absorbed dose in grays; thus, $1 \text{ rem} = 0.01 \text{ Sv}$.

Penetrating Power of Emissions The effect on living tissue of a radiation dose depends on the penetrating power *and* ionizing ability of the radiation. Figure 23.7 depicts the differences in penetrating power of the three common emissions. Note, in general, that *penetrating power is inversely related to the mass, charge, and energy of the emission*. In other words, if a particle interacts strongly with matter, it penetrates only slightly, and vice versa:

- **α Particles.** Alpha particles are massive and highly charged, which means that they interact with matter most strongly of the three common types of emissions. As a result, they penetrate so little that a piece of paper, light clothing, or the outer layer of skin can stop α radiation from an external source. Internally, however, such as from ingestion, an α emitter can cause grave localized damage through extensive ionization.



FIGURE 23.7 Penetrating power of radioactive emissions. Penetrating power is often measured in terms of the depth of water that stops 50% of the incoming radiation. (Water is the main component of living tissue.) Alpha particles, with the highest mass and charge, have the lowest penetrating power, and γ rays have the highest. (Average values of actual penetrating distances are shown.)

- *Beta Particles and Positrons.* Beta particles (β^-) and positrons (β^+) have less charge and much less mass than α particles, so they interact less strongly with matter. Even though a given particle has less chance of causing ionization, a β^- (or β^+) emitter is a more destructive external source because the particles penetrate deeper. Specialized heavy clothing or a thick (0.5 cm) piece of metal is required to stop these particles.
- *γ Rays.* Neutral, massless γ rays interact least with matter and, thus, penetrate most. A block of lead several inches thick is needed to stop them. Therefore, an external γ ray source is the most dangerous because the energy can ionize many layers of living tissue.

Sources of Ionizing Radiation

We are continuously exposed to ionizing radiation from natural and artificial sources (Table 23.6). Indeed, life evolved in the presence of natural ionizing radiation, called **background radiation**. The same radiation that alters bonds in DNA and causes harmful mutations also causes beneficial mutations that, over evolutionary time, allow species to change.

Background radiation has several sources. One source is *cosmic radiation*, which increases with altitude because of decreased absorption by the atmosphere. Thus, people in Denver absorb twice as much cosmic radiation as people in Los Angeles; even a jet flight involves measurable absorption. The sources of most background radiation are thorium and uranium minerals present in rocks and soil.

Table 23.6 Typical Radiation Doses from Natural and Artificial Sources

Source of Radiation	Average Adult Exposure
Natural	
Cosmic radiation	30–50 mrem/yr
Radiation from the ground	
From clay soil and rocks	~25–170 mrem/yr
In wooden houses	10–20 mrem/yr
In brick houses	60–70 mrem/yr
In concrete (cinder block) houses	60–160 mrem/yr
Radiation from the air (mainly radon)	
Outdoors, average value	20 mrem/yr
In wooden houses	70 mrem/yr
In brick houses	130 mrem/yr
In concrete (cinder block) houses	260 mrem/yr
Internal radiation from minerals in tap water and daily intake of food (^{40}K , ^{14}C , Ra)	~40 mrem/yr
Artificial	
Diagnostic x-ray methods	
Lung (local)	0.04–0.2 rad/film
Kidney (local)	1.5–3 rad/film
Dental (dose to the skin)	≤1 rad/film
Therapeutic radiation treatment	Locally ≤ 10,000 rad
Other sources	
Jet flight (4 h)	~1 mrem
Nuclear testing	<4 mrem/yr
Nuclear power industry	<1 mrem/yr
Total average value	100–200 mrem/yr

Radon, the heaviest noble gas [Group 8A(18)], is a radioactive product of uranium and thorium decay. Its concentration in the air we breathe is associated with certain common building materials and types of local soil and rocks. Its decay products cause most of the damage. About 150 g of K^+ ions is dissolved in the water in the tissues of an average adult, and 0.0118% of these ions are radioactive ^{40}K . The presence of these substances and of atmospheric $^{14}CO_2$ makes all food, water, clothing, and building materials slightly radioactive.

The largest artificial source of radiation, and the easiest to control, is medical diagnostic procedures, especially x-rays. The radiation dosage from nuclear testing and radioactive waste disposal is miniscule for most people, but exposures for those living near test sites or disposal areas may be much higher.

SECTION 23.4 SUMMARY

All radioactive emissions cause ionization. • The effect of ionizing radiation on living matter depends on the quantity of energy absorbed and its penetrating power, and the extent of ionization in a given type of tissue. Radiation dose for the human body is measured in rem. • All organisms are exposed to varying quantities of natural ionizing radiation.

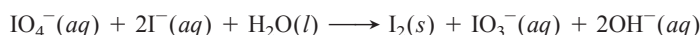
23.5 APPLICATIONS OF RADIOISOTOPES

Our ability to detect minute amounts of radioisotopes makes them powerful tools for studying processes in biochemistry, medicine, materials science, environmental studies, and many other scientific and industrial fields. Such uses depend on the fact that *isotopes of an element exhibit very similar chemical and physical behavior*. In other words, except for having a less stable nucleus, a radioisotope has nearly the same chemical properties as a nonradioactive isotope of that element. For example, the fact that $^{14}CO_2$ is utilized by a plant in the same way as $^{12}CO_2$ forms the basis of radiocarbon dating.

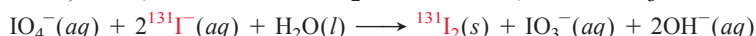
Radioactive Tracers

A tiny amount of a radioisotope mixed with a large amount of the stable isotope can act as a **tracer**, a chemical “beacon” emitting radiation that signals the presence of the substance.

Reaction Pathways Tracers help us choose from among possible reaction pathways. One well-studied example is the reaction between periodate and iodide ions:



Is IO_3^- the result of IO_4^- reduction or I^- oxidation? When we add “cold” (non-radioactive) IO_4^- to a solution of I^- that contains some “hot” (radioactive, indicated in red) $^{131}I^-$, we find that the I_2 is radioactive, not the IO_3^- :

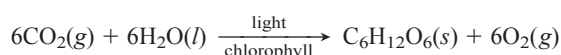


These results show that IO_3^- forms through the reduction of IO_4^- , and that I_2 forms through the oxidation of I^- . To confirm this pathway, we add IO_4^- containing some $^{131}IO_4^-$ to a solution of I^- . As we expected, the IO_3^- is radioactive, not the I_2 :



Thus, tracers act like “handles” we can “hold” to follow the changing reactants.

Far more complex pathways can be followed with tracers as well. The photosynthetic pathway, the most essential and widespread metabolic process on Earth, in which energy from sunlight is used to form the chemical bonds of glucose, has an overall reaction that looks quite simple:



However, the actual process is extremely complex: 13 enzyme-catalyzed steps are required to incorporate each C atom from CO_2 , so the six CO_2 molecules incorporated to form a molecule of $\text{C}_6\text{H}_{12}\text{O}_6$ require six repetitions of the pathway. Melvin Calvin and his coworkers took seven years to determine the pathway. He won the Nobel Prize in chemistry in 1961 for this remarkable achievement.

Material Flow Tracers are used in studies of solid surfaces and the flow of materials. Metal atoms hundreds of layers deep within a solid have been shown to exchange with metal ions from the surrounding solution within a matter of minutes. Chemists and engineers use tracers to study material movement in semiconductor chips, paint, and metal plating, in detergent action, and in the process of corrosion, to mention just a few of many applications.

Hydrologic engineers use tracers to study the volume and flow of large bodies of water. By following radionuclides that formed during atmospheric nuclear bomb tests (^3H in H_2O , $^{90}\text{Sr}^{2+}$, and $^{137}\text{Cs}^+$), scientists have mapped the flow of water from land to lakes and streams to oceans. They also use tracers to study the surface and deep ocean currents that circulate around the globe, as well as the mechanisms of hurricane formation and the mixing of the troposphere and stratosphere. Industries employ tracers to study material flow during the manufacturing process, such as the flow of ore pellets in smelting kilns, the paths of wood chips and bleach in paper mills, the diffusion of fungicide into lumber, and in a particularly important application, the porosity and leakage of oil and gas wells in geological formations.

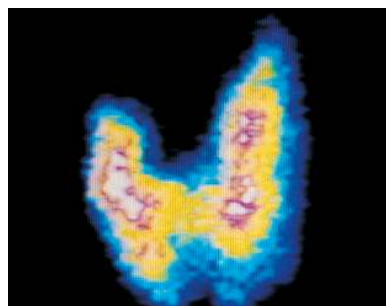


FIGURE 23.8 The use of radioisotopes to image the thyroid gland. Thyroid scanning is used to assess nutritional deficiencies, inflammation, tumor growth, and other thyroid-related ailments. In ^{131}I scanning, the thyroid gland absorbs $^{131}\text{I}^-$ ions, which undergo β^- decay that exposes a photographic film. The asymmetric image indicates disease.

Table 23.7 Some Radioisotopes Used as Medical Tracers

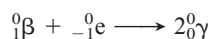
Isotope	Body Part or Process
^{11}C , ^{18}F , ^{13}N , ^{15}O	PET studies of brain, heart
^{60}Co , ^{192}Ir	Cancer therapy
^{64}Cu	Metabolism of copper
^{59}Fe	Blood flow, spleen
^{67}Ga	Tumor imaging
^{123}I , ^{131}I	Thyroid
^{111}In	Brain, colon
^{42}K	Blood flow
$^{81\text{m}}\text{Kr}$	Lung
$^{99\text{m}}\text{Tc}$	Heart, thyroid, liver, lung, bone
^{201}Tl	Heart muscle
^{90}Y	Cancer, arthritis

Activation Analysis Another use of tracers is in *neutron activation analysis* (NAA). In this method, neutrons bombard a nonradioactive sample, converting a small fraction of its atoms to radioisotopes, which exhibit characteristic decay patterns, such as γ -ray spectra, that reveal the elements present. Unlike chemical analysis, NAA leaves the sample virtually intact, so the method can be used to determine the composition of a valuable object or a very small sample. For example, a painting thought to be a 16th-century Dutch masterpiece was shown through NAA to be a 20th-century forgery, because a microgram-sized sample of its pigment contained much less silver and antimony than the pigments used by the Dutch masters. Forensic chemists use NAA to detect traces of ammunition on a suspect's hand or traces of arsenic in the hair of a victim of poisoning.

Medical Diagnosis The largest use of radioisotopes is in medical science. In fact, over 25% of U.S. hospital admissions are for diagnoses based on data from radioisotopes. Tracers with half-lives of a few minutes to a few days are employed to observe specific organs and body parts. For example, a healthy thyroid gland incorporates dietary I^- into iodine-containing hormones at a known rate. To assess thyroid function, the patient drinks a solution containing a trace amount of Na^{131}I , and a scanning monitor follows the uptake of $^{131}\text{I}^-$ into the thyroid (Figure 23.8).

Tracers are also used to measure physiological processes, such as blood flow. The rate at which the heart pumps blood, for example, can be observed by injecting ^{59}Fe , which becomes incorporated into the hemoglobin of blood cells. Several radioisotopes used in medical diagnosis are listed in Table 23.7.

Positron-emission tomography (PET) is a powerful imaging method for observing brain structure and function. A biological substance is synthesized with one of its atoms replaced by an isotope that emits positrons. The substance is injected into a patient's bloodstream, from which it is taken up into the brain. The isotope emits positrons, each of which annihilates a nearby electron. In this process, two γ photons are emitted simultaneously 180° from each other:



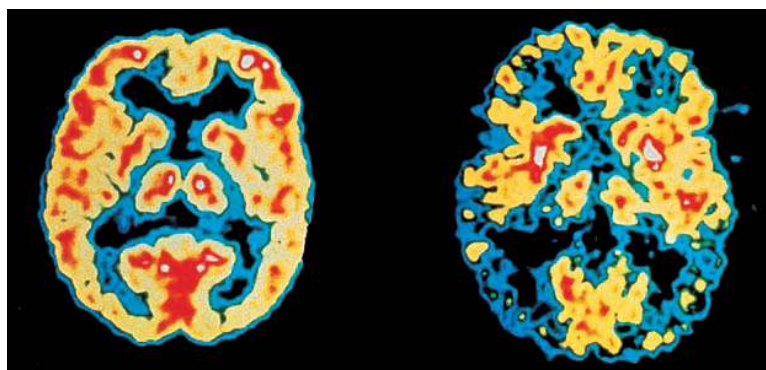


FIGURE 23.9 PET and brain activity. These PET scans show brain activity in a normal person (*left*) and in a patient with Alzheimer's disease (*right*). Red and yellow indicate relatively high activity within a region.

 Animation: Nuclear Medicine

An array of detectors around the patient's head pinpoints the sites of γ emission, and the image is analyzed by computer. Two of the isotopes used are ^{15}O , injected as H_2^{15}O to measure blood flow, and ^{18}F bonded to a glucose analog to measure glucose uptake, which is an indicator of energy metabolism. Among many fascinating PET findings are those that show how changes in blood flow and glucose uptake accompany normal or abnormal brain activity (Figure 23.9). In a recent nonmedical development, substances incorporating ^{11}C and ^{15}O are being investigated using PET to learn how molecules interact with and move along the surface of a catalyst.

Additional Applications of Ionizing Radiation

To be used as a tracer, a radioisotope need only emit radiation that can be detected. Many other uses of radioisotopes depend on ionizing radiation of higher energy.

Cancer cells divide more rapidly than normal cells, so radioisotopes that interfere with cell division kill more cancer cells than normal ones. Implants of ^{198}Au or of ^{90}Sr , which decays to the γ -emitting ^{90}Y , have been used to destroy pituitary and breast tumor cells, and γ rays from ^{60}Co have been used to destroy tumors of the brain and other body parts.

Irradiation of food increases shelf life by killing microorganisms that cause rotting or spoilage (Figure 23.10), but the practice is quite controversial. Advocates point to the benefits of preserving fresh foods, grains, and seeds for long periods, whereas opponents suggest that irradiation might lower the food's nutritional content or produce harmful byproducts. Irradiation also provides a way to destroy newer, more resistant bacterial strains that survive the increasing use of the more common antibiotics in animal feed. The United Nations has approved irradiation for potatoes, wheat, chicken, and strawberries, and, in 2003, the U.S. Food and Drug Administration approved it as well.

Ionizing radiation has been used to control harmful insects. Captured males are sterilized by radiation and released to mate, thereby reducing the number of offspring. This method has been used to control the Mediterranean fruit fly in California and disease-causing insects, such as the tsetse fly and malarial mosquito, in other parts of the world.

SECTION 23.5 SUMMARY

Radioisotopic tracers have been used to study reaction mechanisms, material flow, elemental composition, and medical conditions. • Ionizing radiation has been used in devices that destroy cancer tissue, kill organisms that spoil food, and control insect populations.

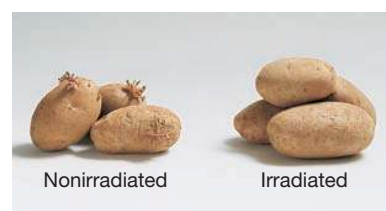


FIGURE 23.10 The increased shelf life of irradiated food.

23.6 THE INTERCONVERSION OF MASS AND ENERGY

Most of the nuclear processes we've considered so far have involved radioactive decay, in which a nucleus emits one or a few small particles or photons to become a more stable, slightly lighter nucleus. Two other nuclear processes cause much greater mass changes. In nuclear **fission**, a heavy nucleus splits into two much lighter nuclei, emitting several small particles at the same time. In nuclear **fusion**, the opposite process occurs: two lighter nuclei combine to form a heavier one. Both fission and fusion release enormous quantities of energy. Let's take a look at the origins of this energy by first examining the change in mass that accompanies the breakup of a nucleus into its nucleons and then considering the energy that is equivalent to this mass change.

The Mass Difference Between a Nucleus and Its Nucleons

For almost a century, we have known that mass and energy are interconvertible. Thus, the separate mass and energy conservation laws are combined to state that *the total quantity of mass-energy in the universe is constant*. Therefore, when any reacting system releases or absorbs energy, it also loses or gains mass.

This relation between mass and energy is not important for chemical reactions because the energy changes involved in breaking or forming chemical bonds are so small that the mass changes are negligible. For example, when 1 mol of water breaks up into its atoms, heat is absorbed and we have:



We find the mass that is equivalent to this energy from *Einstein's equation*:

$$E = mc^2 \quad \text{or} \quad \Delta E = \Delta mc^2 \quad \text{so} \quad \Delta m = \frac{\Delta E}{c^2} \quad (23.7)$$

where Δm is the mass difference between reactants and products:

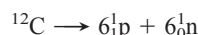
$$\Delta m = m_{\text{products}} - m_{\text{reactants}}$$

Substituting the heat of reaction (in J/mol) for ΔE and the numerical value for c ($2.9979 \times 10^8 \text{ m/s}$), we obtain

$$\Delta m = \frac{9.34 \times 10^5 \text{ J/mol}}{(2.9979 \times 10^8 \text{ m/s})^2} = 1.04 \times 10^{-11} \text{ kg/mol} = 1.04 \times 10^{-8} \text{ g/mol}$$

(Units of kg/mol are obtained because the joule includes the kilogram: $1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2$.) The mass of 1 mol of H_2O molecules (reactant) is about 10 ng *less* than the combined masses of 2 mol of H atoms and 1 mol of O atoms (products), a change difficult to measure with even the most sophisticated balance. Such minute mass changes when bonds break or form allow us to assume that, for all practical purposes, mass is conserved in *chemical* reactions.

The much larger mass change that accompanies a *nuclear* process is related to the enormous energy required to bind the nucleus together from its parts. In an analogy with the calculation above involving the water molecule, consider the change in mass that occurs when one ^{12}C nucleus breaks apart into its nucleons—six protons and six neutrons:



We calculate this mass difference in a special way. By combining the mass of six H *atoms* and six neutrons and then subtracting the mass of one ^{12}C *atom*, the masses of the electrons cancel: six e^- (in six ^1H atoms) cancel six e^- (in one ^{12}C atom). The mass of one ^1H atom is 1.007825 amu, and the mass of one neutron is 1.008665 amu, so

$$\begin{array}{r} \text{Mass of six } ^1\text{H atoms} = 6 \times 1.007825 \text{ amu} = 6.046950 \text{ amu} \\ \text{Mass of six neutrons} = 6 \times 1.008665 \text{ amu} = 6.051990 \text{ amu} \\ \hline \text{Total mass} = 12.098940 \text{ amu} \end{array}$$

The mass of the reactant, one ^{12}C atom, is 12 amu (exactly). The mass difference (Δm) we obtain is the total mass of the nucleons minus the mass of the nucleus:

$$\begin{aligned}\Delta m &= 12.098940 \text{ amu} - 12.000000 \text{ amu} \\ &= 0.098940 \text{ amu}/^{12}\text{C} = 0.098940 \text{ g/mol } ^{12}\text{C}\end{aligned}$$

First, and most important, note that *the mass of the nucleus is less than the combined masses of its nucleons*: there is *always* a mass decrease when nucleons form a nucleus. Second, note that the mass change of this nuclear process ($9.89 \times 10^{-2} \text{ g/mol}$) is nearly 10 million times that of the chemical process ($10.4 \times 10^{-9} \text{ g/mol}$) we saw earlier and easily observed on any laboratory balance.

Nuclear Binding Energy and the Binding Energy per Nucleon

Once again, Einstein's equation for the relation between mass and energy allows us to find the energy equivalent of the mass change. For ^{12}C , after converting grams to kilograms, we have

$$\begin{aligned}\Delta E &= \Delta mc^2 = (9.8940 \times 10^{-5} \text{ kg/mol})(2.9979 \times 10^8 \text{ m/s})^2 \\ &= 8.8921 \times 10^{12} \text{ J/mol} = 8.8921 \times 10^9 \text{ kJ/mol}\end{aligned}$$

This quantity of energy is called the **nuclear binding energy** for carbon-12. The nuclear binding energy is the energy required to break *1 mol of nuclei into their individual nucleons*:



Thus, the nuclear binding energy is qualitatively analogous to the sum of bond energies of a covalent compound or the lattice energy of an ionic compound. But, quantitatively, nuclear binding energies are typically several million times greater.

We use joules to express the binding energy per mole of nuclei, but the joule is much too large a unit to express the binding energy of a single nucleus. Instead, nuclear scientists use the **electron volt (eV)**, the energy an electron acquires when it moves through a potential difference of 1 volt:

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

Binding energies are commonly expressed in millions of electron volts, that is, in *mega-electron volts* (MeV):

$$1 \text{ MeV} = 10^6 \text{ eV} = 1.602 \times 10^{-13} \text{ J}$$

A particularly useful factor converts the atomic mass unit to its energy equivalent in electron volts:

$$1 \text{ amu} = 931.5 \times 10^6 \text{ eV} = 931.5 \text{ MeV} \quad (23.8)$$

Earlier we found the mass change when ^{12}C breaks apart into its nucleons to be 0.098940 amu. The binding energy per ^{12}C nucleus, expressed in MeV, is

$$\frac{\text{Binding energy}}{^{12}\text{C nucleus}} = 0.098940 \text{ amu} \times \frac{931.5 \text{ MeV}}{1 \text{ amu}} = 92.16 \text{ MeV}$$

We can compare the stability of nuclides of different elements by determining the *binding energy per nucleon*. For ^{12}C , we have

$$\text{Binding energy per nucleon} = \frac{\text{binding energy}}{\text{no. of nucleons}} = \frac{92.16 \text{ MeV}}{12 \text{ nucleons}} = 7.680 \text{ MeV/nucleon}$$

SAMPLE PROBLEM 23.6 Calculating the Binding Energy per Nucleon

Problem Iron-56 is an extremely stable nuclide. Compute the binding energy per nucleon for ^{56}Fe and compare it with that for ^{12}C (mass of ^{56}Fe atom = 55.934939 amu; mass of ^1H atom = 1.007825 amu; mass of neutron = 1.008665 amu).

Plan Iron-56 has 26 protons and 30 neutrons. We calculate the mass difference, Δm , when the nucleus forms by subtracting the given mass of one ^{56}Fe atom from the sum of the masses of 26 ^1H atoms and 30 neutrons. To find the binding energy per nucleon, we multiply Δm by the equivalent in MeV (931.5 MeV/amu) and divide by the number of nucleons (56).

Solution Calculating the mass difference, Δm :

$$\begin{aligned}\text{Mass difference} &= [(26 \times \text{mass } ^1\text{H atom}) + (30 \times \text{mass neutron})] - \text{mass } ^{56}\text{Fe atom} \\ &= [(26)(1.007825 \text{ amu}) + (30)(1.008665 \text{ amu})] - 55.934939 \text{ amu} \\ &= 0.52846 \text{ amu}\end{aligned}$$

Calculating the binding energy per nucleon:

$$\text{Binding energy per nucleon} = \frac{0.52846 \text{ amu} \times 931.5 \text{ MeV/amu}}{56 \text{ nucleons}} = 8.790 \text{ MeV/nucleon}$$

An ^{56}Fe nucleus would require more energy per nucleon to break up into its nucleons than would ^{12}C (7.680 MeV/nucleon), so ^{56}Fe is more stable than ^{12}C .

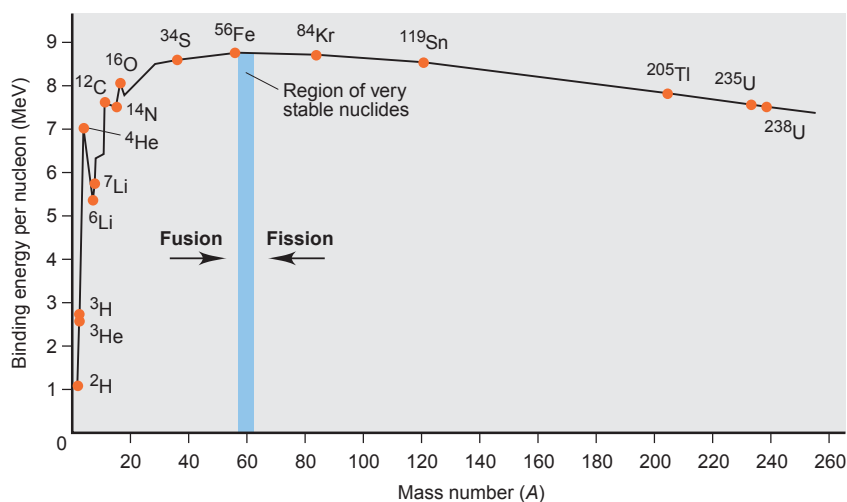
Check The answer is consistent with the great stability of ^{56}Fe . Given the number of decimal places in the values, rounding to check the math is useful only to find a *major* error. The number of nucleons (56) is an exact number, so we retain four significant figures.

FOLLOW-UP PROBLEM 23.6 Uranium-235 is an essential component of the fuel in nuclear power plants. Calculate the binding energy per nucleon for ^{235}U . Is this nuclide more or less stable than ^{12}C (mass of ^{235}U atom = 235.043924 amu)?

Fission or Fusion: Means of Increasing the Binding Energy Per Nucleon Calculations similar to those in Sample Problem 23.6 for other nuclides show that the binding energy per nucleon varies considerably. The essential point is that *the greater the binding energy per nucleon, the more stable the nuclide*.

Figure 23.11 shows a plot of the binding energy per nucleon vs. mass number. It provides information about nuclide stability and the two possible processes nuclides can undergo to form more stable nuclides. Most nuclides with fewer than 10 nucleons have a relatively small binding energy per nucleon. The ^4He nucleus is an exception—it is stable enough to be emitted intact as an α particle. Above $A = 12$, the binding energy per nucleon varies from about 7.6 to 8.8 MeV.

FIGURE 23.11 The variation in binding energy per nucleon. A plot of the binding energy per nucleon vs. mass number shows that nuclear stability is greatest in the region near ^{56}Fe . Lighter nuclei may undergo fusion to become more stable; heavier ones may undergo fission. Note the exceptional stability of ^4He among extremely light nuclei.



The most important observation is that *the binding energy per nucleon peaks at elements with $A \approx 60$* . In other words, nuclides become more stable with increasing mass number up to around 60 nucleons and then become less stable with higher numbers of nucleons. The existence of a peak of stability suggests that there are two ways nuclides can increase their binding energy per nucleon:

- **Fission.** A heavier nucleus can *split into lighter ones* (closer to $A \approx 60$) by undergoing fission. The product nuclei have greater binding energy per nucleon (are more stable) than the reactant nucleus, and the difference in energy is released. Nuclear power plants generate energy through fission, as do atomic bombs.

- **Fusion.** Lighter nuclei, on the other hand, can *combine to form a heavier one* (closer to $A \approx 60$) by undergoing fusion. Once again, the product is more stable than the reactants, and *energy is released*. The Sun and other stars generate energy through fusion, as do thermonuclear (hydrogen) bombs. In these examples and in all current research efforts for developing fusion as a useful energy source, hydrogen nuclei fuse to form the very stable helium-4 nucleus.

In Section 23.7, we examine fission and fusion and the industrial energy facilities designed to utilize them.

SECTION 23.6 SUMMARY

The mass of a nucleus is less than the sum of the masses of its nucleons. The energy equivalent to this mass difference is the nuclear binding energy, often expressed in units of MeV. • The binding energy per nucleon is a measure of nuclide stability and varies with the number of nucleons. Nuclides with $A \approx 60$ are most stable. • Lighter nuclides join (fusion) or heavier nuclides split (fission) to create more stable products.

23.7 APPLICATIONS OF FISSION AND FUSION

Of the many beneficial applications of nuclear reactions, the greatest is the potential for almost limitless amounts of energy. Our experience with nuclear energy from power plants, however, has shown that we must improve ways to tap this energy source safely and economically and deal with the waste generated. In this section, we discuss how fission and fusion occur and how we are applying them.

The Process of Nuclear Fission

During the mid-1930s, scientists bombarded uranium ($Z = 92$) with neutrons in an attempt to synthesize transuranium elements. Many of the unstable nuclides produced were tentatively identified as having $Z > 92$, but four years later, one of these was shown to be an isotope of barium ($Z = 56$). The Austrian physicist Lise Meitner and her nephew, Otto Frisch, proposed that barium resulted from the *splitting* of the uranium nucleus into *smaller* nuclei, a process that they called *fission* as an analogy to cell division in biology.

The ^{235}U nucleus can split in many different ways, giving rise to various daughter nuclei, but all routes have the same general features. Figure 23.12 depicts one of these fission patterns. Neutron bombardment results in a highly excited ^{236}U nucleus, which splits apart in 10^{-14} s. The products are two nuclei of unequal

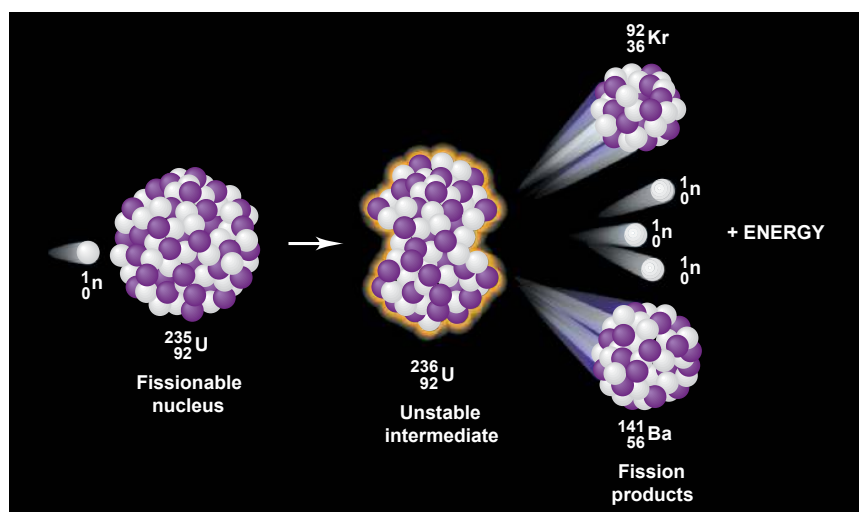
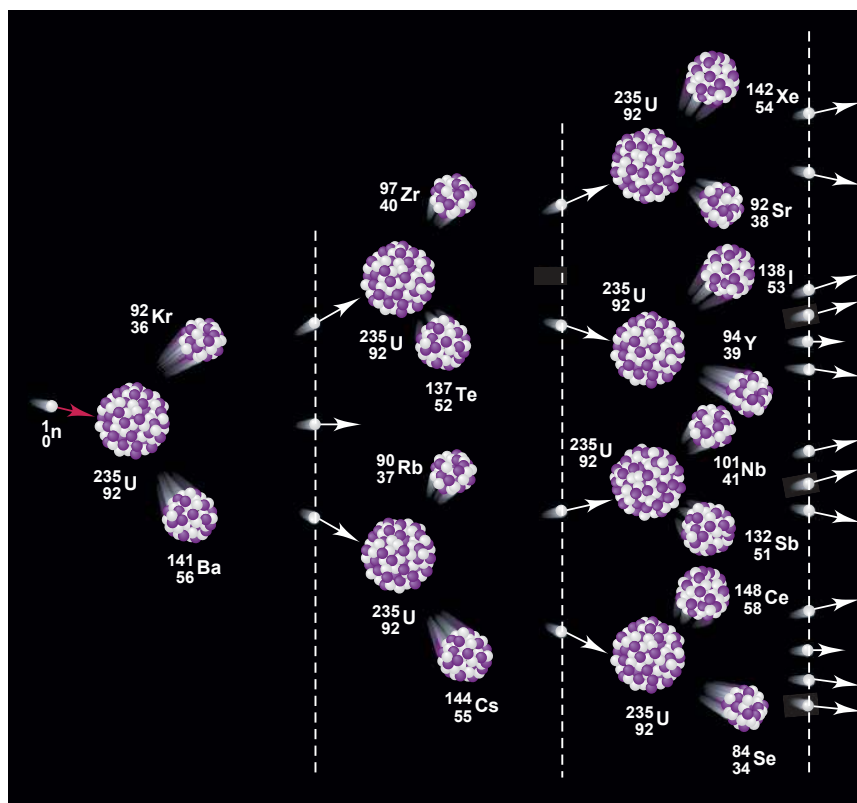


FIGURE 23.12 Induced fission of ^{235}U .

A neutron bombarding a ^{235}U nucleus results in an extremely unstable ^{236}U nucleus, which becomes distorted in the act of splitting. In this case, which shows one of many possible splitting patterns, the products are ^{92}Kr and ^{141}Ba . Three neutrons and a great deal of energy are also released.

FIGURE 23.13 A chain reaction of ^{235}U .

If a sample exceeds the critical mass, neutrons produced by the first fission event collide with other nuclei, causing their fission and the production of more neutrons to continue the process. Note that various product nuclei form. The vertical dashed lines identify succeeding “generations” of neutrons.



mass, two or three neutrons (average of 2.4), and a large quantity of energy. A single ^{235}U nucleus releases $3.5 \times 10^{-11} \text{ J}$ when it splits; 1 mol of ^{235}U (about $\frac{1}{2}$ lb) releases $2.1 \times 10^{13} \text{ J}$ —a billion times as much energy as burning $\frac{1}{2}$ lb of coal (about $2 \times 10^4 \text{ J}$)!

We harness the energy of nuclear fission, much of which eventually appears as heat, by means of a **chain reaction**, illustrated in Figure 23.13: the two to three neutrons that are released by the fission of one nucleus collide with other fissionable nuclei and cause them to split, releasing more neutrons, which then collide with other nuclei, and so on, in a self-sustaining process. In this manner, the energy released increases rapidly because each fission event in a chain reaction releases two to three times as much energy as the preceding one.

Whether a chain reaction occurs depends on the mass (and thus the volume) of the fissionable sample. If the piece of uranium is large enough, the product neutrons strike another fissionable nucleus *before* flying out of the sample, and a chain reaction takes place. The mass required to achieve a chain reaction is called the **critical mass**. If the sample has less than the critical mass (a *subcritical mass*), too many product neutrons leave the sample before they collide with and cause the fission of another ^{235}U nucleus, and thus a chain reaction does not occur.

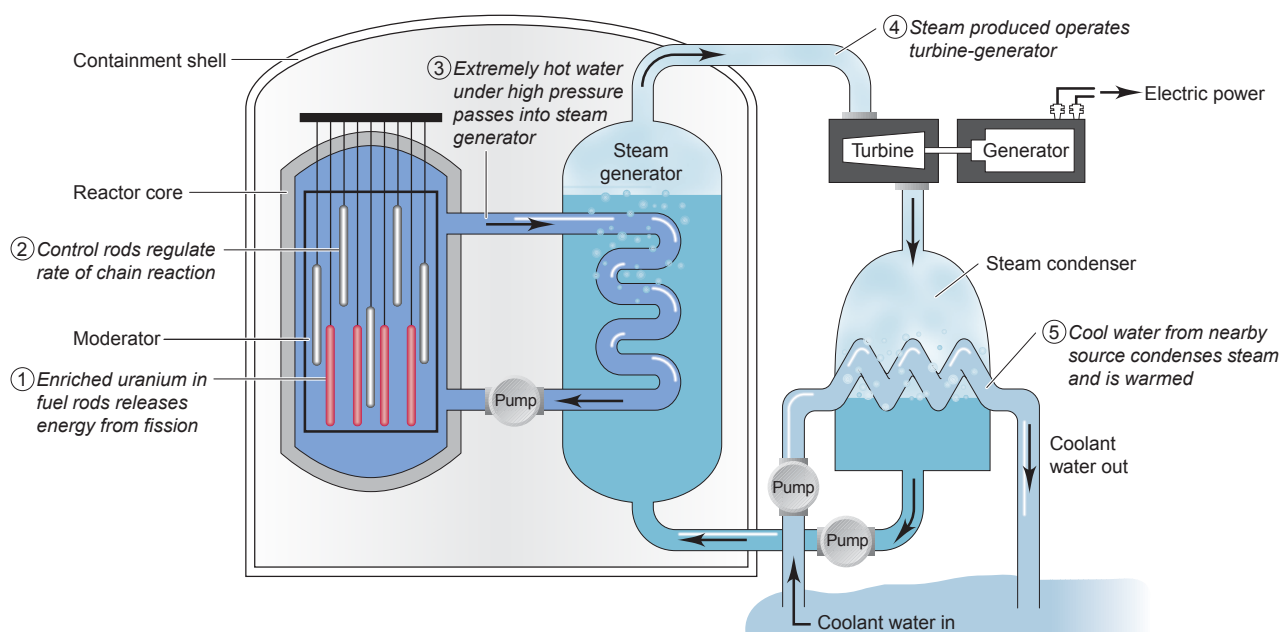
Nuclear Energy Reactors Like a coal-fired power plant, a *nuclear power plant generates heat to produce steam, which turns a turbine attached to an electric generator*. But a nuclear plant has the potential to produce electric power much more cleanly than can the combustion of coal.

Heat generation takes place in the **reactor core** of a nuclear plant (Figure 23.14). The core contains the *fuel rods*, which consist of fuel enclosed in tubes of a corrosion-resistant zirconium alloy. The fuel is uranium(IV) oxide (UO_2) that has been *enriched* from 0.7% ^{235}U , the natural abundance of this fissionable isotope, to the 3% to 4% ^{235}U required to sustain a chain reaction in a practical volume.



A

FIGURE 23.14 A light-water nuclear reactor. **A**, Photo of a facility showing the concrete containment shell and nearby water source. **B**, Schematic of a light-water reactor.



B

Sandwiched between the fuel rods are movable *control rods* made of cadmium or boron (or, in nuclear submarines, hafnium), substances that absorb neutrons very efficiently. When the control rods are moved between the fuel rods, the chain reaction slows because fewer neutrons are available to bombard uranium atoms; when they are removed, the chain reaction speeds up. Neutrons that leave the fuel-rod assembly collide with a *reflector*, usually made of a beryllium alloy, which absorbs very few neutrons. Reflecting the neutrons back to the fuel rods speeds the chain reaction.

Flowing around the fuel and control rods in the reactor core is the *moderator*, a substance that slows the neutrons, making them much better at causing fission than the fast ones emerging directly from the fission event. In most modern reactors, the moderator also acts as the *coolant*, the fluid that transfers the released heat to the steam-producing region. *Light-water reactors* use H_2O as the moderator because ^1H absorbs neutrons to some extent. Heavy-water reactors use D_2O because it absorbs very few neutrons, leaving more available for fission. Thus, heavy-water reactors can use uranium that has been *less enriched*. As the coolant flows around the encased fuel, pumps circulate it through coils that transfer its heat to the water reservoir. Steam formed in the reservoir turns the turbine that

runs the generator. The steam is then condensed in large cooling towers, using water from a lake or river to absorb heat, and returned to the water reservoir.

Some major accidents at nuclear plants have caused decidedly negative public reactions. In 1979, malfunctions of coolant pumps and valves at the Three-Mile Island facility in Pennsylvania led to melting of some of the fuel and damage to the reactor core, but the release of only a very small amount (about 1 Ci) of radioactive gases into the atmosphere. In 1986, a million times as much radioactivity (1 MCi) was released when a cooling system failure at the Chernobyl plant in Ukraine caused a much greater melting of fuel and an uncontrolled reaction. High-pressure steam and ignited graphite moderator rods caused the reactor building to explode and expel radioactive debris. Carried by prevailing winds, the radioactive particles contaminated vegetables and milk in much of Europe.

Despite potential safety problems, nuclear power remains an important source of electricity. Since the late 1990s, nearly every European country has employed nuclear power, and it is the major power source in some countries—50% of the electricity in Sweden and almost 80% in France. Currently, the United States obtains about 20% of its electricity from nuclear power, and Canada slightly less.

However, even a smoothly operating plant has certain inherent problems. The problem of *thermal pollution* is common to all power plants. Water used to condense the steam is several degrees warmer when returned to its source, which can harm aquatic organisms (Section 13.3). A more serious problem is *nuclear waste disposal*. Many of the fission products formed in nuclear reactors have long half-lives, and no satisfactory plan for their permanent disposal has yet been devised. Proposals to place the waste in containers and bury them in deep bedrock cannot possibly be field-tested for the thousands of years the material will remain harmful. Leakage of radioactive material into groundwater is a danger, and earthquakes can occur even in geologically stable regions. Despite studies indicating that the proposed disposal site at Yucca Mountain, Nevada, is too geologically active, the U.S. government approved the site. It remains to be seen whether we can operate fission reactors *and* dispose of the waste safely and economically.

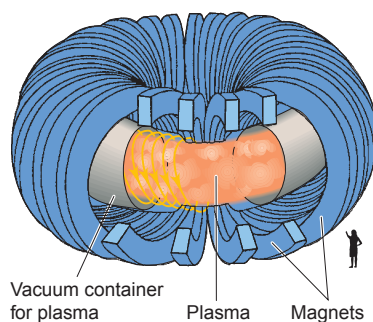
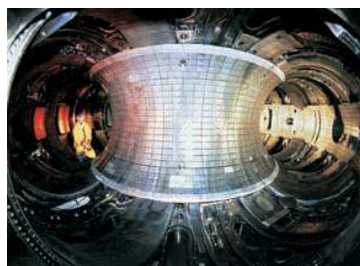
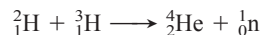


FIGURE 23.15 The tokamak design for magnetic containment of a fusion plasma. The donut-shaped chamber of the tokamak (photo, top; schematic, bottom) contains the plasma within a helical magnetic field.

The Promise of Nuclear Fusion

Nuclear fusion is the ultimate source of nearly all the energy on Earth because almost all other sources depend, directly or indirectly, on the energy produced by nuclear fusion in the Sun. But the Sun and other stars generate more than energy; in fact, *all the elements larger than hydrogen were formed in fusion and decay processes within stars.*

Much research is being devoted to making nuclear fusion a practical, direct source of energy on Earth. To understand the advantages of fusion, let's consider one of the most discussed fusion reactions, in which deuterium and tritium react:



This reaction produces 1.7×10^9 kJ/mol, an enormous quantity of energy with no radioactive byproducts. Moreover, the reactant nuclei are relatively easy to come by. Thus, in principle, fusion seems very promising and may represent an ideal source of power. However, some extremely difficult technical problems remain. Fusion requires enormous energy in the form of heat to give the positively charged nuclei enough kinetic energy to force themselves together. The fusion of deuterium and tritium, for example, occurs at practical rates at about 10^8 K, a temperature hotter than the Sun's core! How can such conditions be achieved?

Two research approaches hold promise. In one, atoms are stripped of their electrons at high temperatures, resulting in a gaseous *plasma*, a neutral mixture of positive nuclei and electrons. Because extreme temperatures are needed for fusion, no *material* can contain the plasma. The most successful approach to date has been to enclose the plasma within a magnetic field. Figure 23.15 shows the

tokamak design: a donut-shaped container in which a helical magnetic field confines the plasma and prevents it from contacting the walls. Scientists at the Princeton University Plasma Physics facility have achieved some success in generating energy from fusion this way. In another approach, the high temperature is reached by using many focused lasers to compress and heat the fusion reactants. In any event, one or more major breakthroughs are needed before fusion will be realized as a practical, everyday source of energy.

SECTION 23.7 SUMMARY

In nuclear fission, neutron bombardment causes a nucleus to split into two smaller nuclei and release neutrons that split other nuclei, giving rise to a chain reaction. • A nuclear power plant controls the rate of the chain reaction to produce heat that creates steam, which is used to generate electricity. • Potential hazards, such as radiation leaks, thermal pollution, and disposal of nuclear waste, remain current concerns. • Nuclear fusion holds great promise as a source of clean abundant energy, but it requires extremely high temperatures and is not yet practical.

CHAPTER REVIEW GUIDE

The following sections provide many aids to help you study this chapter. (Numbers in parentheses refer to pages, unless noted otherwise.)

• LEARNING OBJECTIVES *These are concepts and skills to review after studying this chapter.*

Related section (§), sample problem (SP), and end-of-chapter problem (EP) numbers are listed in parentheses.

1. Describe the differences between nuclear and chemical changes; identify the three types of radioactive emissions and the types of radioactive decay, and know how each changes A and Z ; explain how a decay series leads to a stable nuclide; write and balance nuclear equations; use the N/Z ratio to predict nuclear stability and the type of decay a nuclide undergoes (§ 23.1) (SPs 23.1–23.3) (EPs 23.1–23.16)
2. Understand why radioactive decay is a first-order process and the meaning of half-life; convert among units of radioactivity, and calculate specific activity, decay constant, half-life, and number of nuclei; estimate the age of an object from its specific activity (§ 23.2) (SPs 23.4, 23.5) (EPs 23.17–23.30)
3. Describe how particle accelerators are used to synthesize new nuclides and write balanced equations for nuclear transmutations (§ 23.3) (EPs 23.31–23.35)

4. Distinguish between excitation and ionization, and describe their effects on matter; convert among units of radiation dose, and understand the penetrating power of emissions and how ionizing radiation is used beneficially (§ 23.4) (EPs 23.36–23.42)
5. Describe how radioisotopes are used in research, elemental analysis, and diagnosis (§ 23.5) (EPs 23.43–23.45)
6. Explain the mass difference and how it is related to nuclear binding energy; understand how nuclear stability is related to binding energy per nucleon and why unstable nuclides undergo either fission or fusion; use Einstein's equation to find mass-energy equivalence in J and eV; compare nuclide stability from binding energy per nucleon (§ 23.6) (SP 23.6) (EPs 23.46–23.52)
7. Discuss the pros and cons of power generation by nuclear fission, and evaluate the potential of nuclear fusion (§ 23.7) (EPs 23.53–23.59)

• KEY TERMS *These important terms appear in boldface in the chapter and are defined again in the Glossary.*

Section 23.1

radioactivity (785)
nucleon (786)
nuclide (786)
isotope (786)
alpha (α) particle (786)
beta (β) particle (786)
gamma (γ) ray (786)
alpha (α) decay (788)
beta (β) decay (788)
 β^- decay (788)
positron (β^+) emission (788)
positron (788)

electron (e^-) capture (EC) (788)
gamma (γ) emission (789)
band of stability (790)
decay (disintegration) series (792)

Section 23.2

activity ($\mathcal{A}$) (793)
becquerel (Bq) (793)
curie (Ci) (793)
decay constant (794)
half-life ($t_{1/2}$) (794)
radioisotopic dating (796)
radioisotope (796)

Section 23.3

nuclear transmutation (797)
deuteron (797)
particle accelerator (797)
transuranium element (798)

Section 23.4

ionization (799)
ionizing radiation (799)
gray (Gy) (799)
rad (radiation-absorbed dose) (799)
rem (roentgen equivalent for man) (799)

sievert (Sv) (799)
background radiation (800)

Section 23.5

tracer (801)

Section 23.6

fission (804)
fusion (804)
nuclear binding energy (805)
electron volt (eV) (805)

Section 23.7

chain reaction (808)
critical mass (808)
reactor core (808)

• **KEY EQUATIONS AND RELATIONSHIPS** *Numbered and screened concepts are listed for you to refer to or memorize.*

23.1 Balancing a nuclear equation (786):

$$\frac{\text{Total } A}{\text{Total } Z} \text{ Reactants} = \frac{\text{Total } A}{\text{Total } Z} \text{ Products}$$

23.2 Defining the unit of radioactivity (curie, Ci) (793):

$$1 \text{ Ci} = 3.70 \times 10^{10} \text{ disintegrations per second (d/s)}$$

23.3 Expressing the decay rate (activity) for radioactive nuclei (794):

$$\text{Decay rate } (\mathcal{A}) = -\frac{\Delta N}{\Delta t} = kN$$

23.4 Finding the number of nuclei remaining after a given time, N_t (794):

$$N_t = N_0 e^{-kt} \quad \text{and} \quad \ln \frac{N_0}{N_t} = kt$$

23.5 Finding the half-life of a radioactive nuclide (794):

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

23.6 Calculating the time to reach a given specific activity (age of an object in radioisotopic dating) (796):

$$t = \frac{1}{k} \ln \frac{\mathcal{A}_0}{\mathcal{A}_t}$$

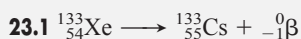
23.7 Adapting Einstein's equation to calculate mass difference and/or nuclear binding energy (804):

$$\Delta m = \frac{\Delta E}{c^2} \quad \text{or} \quad \Delta E = \Delta mc^2$$

23.8 Relating the atomic mass unit to its energy equivalent in MeV (805):

$$1 \text{ amu} = 931.5 \times 10^6 \text{ eV} = 931.5 \text{ MeV}$$

• **BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS** *Compare your own solutions to these calculation steps and answers.*



23.2 Phosphorus-31 has a slightly higher N/Z ratio and an even N (16).

23.3 (a) $N/Z = 1.35$; too high for this region of band: β decay
(b) Mass too high for stability: α decay

23.4 $\ln \mathcal{A}_t = -kt + \ln \mathcal{A}_0$

$$= -\left(\frac{\ln 2}{15 \text{ h}} \times 4.0 \text{ days} \times \frac{24 \text{ h}}{1 \text{ day}}\right) + \ln (2.5 \times 10^9)$$

$$= 17.20$$

$$\mathcal{A}_t = 3.0 \times 10^7 \text{ d/s}$$

23.5 $t = \frac{1}{k} \ln \frac{\mathcal{A}_0}{\mathcal{A}_t} = \frac{5730 \text{ yr}}{\ln 2} \ln \left(\frac{15.3 \text{ d/min} \cdot \text{g}}{9.41 \text{ d/min} \cdot \text{g}}\right) = 4.02 \times 10^3 \text{ yr}$

The mummy case is about 4000 years old.

23.6 ${}^{235}_{92}\text{U}$ has 92 ${}^1_1\text{p}$ and 143 ${}^1_0\text{n}$.

$$\Delta m = [(92 \times 1.007825 \text{ amu}) + (143 \times 1.008665 \text{ amu})] - 235.043924 \text{ amu} = 1.9151 \text{ amu}$$

$$\frac{\text{Binding energy}}{\text{nucleon}} = \frac{1.9151 \text{ amu} \times \frac{931.5 \text{ MeV}}{1 \text{ amu}}}{235 \text{ nucleons}}$$

$$= 7.591 \text{ MeV/nucleon}$$

Therefore, ${}^{235}_{92}\text{U}$ is less stable than ${}^{12}\text{C}$.

PROBLEMS

Problems with **colored** numbers are answered in Appendix E. Sections match the text and provide the numbers of relevant sample problems. Bracketed problems are grouped in pairs (indicated by a short rule) that cover the same concept. Comprehensive Problems are based on material from any section or previous chapter.

Radioactive Decay and Nuclear Stability

(Sample Problems 23.1 to 23.3)

23.1 How do chemical and nuclear reactions differ in

- Magnitude of the energy change?
- Effect on rate of increasing temperature?
- Effect on rate of higher reactant concentration?
- Effect on yield of higher reactant concentration?

23.2 Which of the following types of radioactive decay produce an atom of a *different* element: (a) alpha; (b) beta; (c) gamma; (d) positron; (e) electron capture? Show how Z and N change, if at all, with each type.

23.3 Why is ${}^3_2\text{He}$ stable but ${}^2_2\text{He}$ so unstable that it has never been detected?

23.4 How do the modes of decay differ for a neutron-rich nuclide and a proton-rich nuclide?

23.5 Why can't you use the position of a nuclide's N/Z ratio relative to the band of stability to predict whether it is more likely to decay by positron emission or by electron capture?

23.6 Write balanced nuclear equations for the following:

- Alpha decay of ${}^{234}_{92}\text{U}$
- Electron capture by neptunium-232
- Positron emission by ${}^{12}_7\text{N}$

23.7 Write balanced nuclear equations for the following:

- β^- decay of sodium-26
- β^- decay of francium-223
- α decay of ${}^{212}_{83}\text{Bi}$

23.8 Write balanced nuclear equations for the following:

- Formation of ${}^{48}_{22}\text{Ti}$ through positron emission
- Formation of silver-107 through electron capture
- Formation of polonium-206 through alpha decay

23.9 Write balanced nuclear equations for the following:

- (a) Production of $^{241}_{95}\text{Am}$ through β^- decay
- (b) Formation of $^{228}_{89}\text{Ac}$ through β^- decay
- (c) Formation of $^{203}_{83}\text{Bi}$ through α decay

23.10 Which nuclide(s) would you predict to be stable? Why?

- (a) $^{20}_8\text{O}$ (b) $^{59}_{27}\text{Co}$ (c) ^9_3Li

23.11 Which nuclide(s) would you predict to be stable? Why?

- (a) $^{146}_{60}\text{Nd}$ (b) $^{114}_{48}\text{Cd}$ (c) $^{88}_{42}\text{Mo}$

23.12 What is the most likely mode of decay for each?

- (a) $^{238}_{92}\text{U}$ (b) $^{48}_{24}\text{Cr}$ (c) $^{50}_{25}\text{Mn}$

23.13 What is the most likely mode of decay for each?

- (a) $^{61}_{26}\text{Fe}$ (b) $^{41}_{17}\text{Cl}$ (c) $^{110}_{44}\text{Ru}$

23.14 Why is $^{52}_{24}\text{Cr}$ the most stable isotope of chromium?

23.15 Why is $^{40}_{20}\text{Ca}$ the most stable isotope of calcium?

23.16 Neptunium-237 is the parent nuclide of a decay series that starts with α emission, followed by β^- emission, and then two more α emissions. Write a balanced nuclear equation for each step.

The Kinetics of Radioactive Decay

(Sample Problems 23.4 and 23.5)

23.17 What is the reaction order of radioactive decay? Explain.

23.18 After 1 minute, half the radioactive nuclei remain from an original sample of six nuclei. Is it valid to conclude that $t_{1/2}$ equals 1 minute? Would this conclusion be valid if the original sample contained 6×10^{12} nuclei? Explain.

23.19 Radioisotopic dating depends on the constant rate of decay and formation of various nuclides in a sample. How is the proportion of ^{14}C kept relatively constant in living organisms?

23.20 What is the specific activity (in Ci/g) if 1.55 mg of an isotope emits 1.66×10^6 α particles per second?

23.21 What is the specific activity (in Bq/g) if 8.58 μg of an isotope emits 7.4×10^4 α particles per minute?

23.22 If 1.00×10^{-12} mol of ^{135}Cs emits 1.39×10^5 β particles in 1.00 yr, what is the decay constant?

23.23 If 6.40×10^{-9} mol of ^{176}W emits 1.07×10^{15} positrons in 1.00 h, what is the decay constant?

23.24 The isotope $^{212}_{83}\text{Bi}$ has a half-life of 1.01 yr. What mass (in mg) of a 2.00-mg sample will not have decayed after 3.75×10^3 h?

23.25 The half-life of radium-226 is 1.60×10^3 yr. How many hours will it take for a 2.50-g sample to decay to the point where 0.185 g of the isotope remains?

23.26 A rock contains 270 μmol of ^{238}U ($t_{1/2} = 4.5 \times 10^9$ yr) and 110 μmol of ^{206}Pb . Assuming that all the ^{206}Pb comes from decay of the ^{238}U , estimate the rock's age.

23.27 A fabric remnant from a burial site has a $^{14}\text{C}/^{12}\text{C}$ ratio of 0.735 of the original value. How old is the fabric?

23.28 Due to decay of ^{40}K , cow's milk has a specific activity of about 6×10^{-11} mCi per milliliter. How many disintegrations of ^{40}K nuclei are there per minute in 1.0 qt of milk?

23.29 Plutonium-239 ($t_{1/2} = 2.41 \times 10^4$ yr) represents a serious nuclear waste disposal problem. If seven half-lives are required to reach a tolerable level of radioactivity, how long must ^{239}Pu be stored?

23.30 A volcanic eruption melts a large area of rock, and all gases are expelled. After cooling, ^{40}Ar accumulates from the ongoing decay of $^{40}_{19}\text{K}$ in the rock ($t_{1/2} = 1.25 \times 10^9$ yr). When a piece of rock is analyzed, it is found to contain 1.38 mmol of ^{40}K and 1.14 mmol of ^{40}Ar . How long ago did the rock cool?

Nuclear Transmutation: Induced Changes in Nuclei

23.31 Why must the electrical polarity of the tubes in a linear accelerator be reversed at very short time intervals?

23.32 Why does bombardment with protons usually require higher energies than bombardment with neutrons?

23.33 Name the unidentified species in each transmutation, and write a full nuclear equation:

- (a) Bombardment of ^{10}B with an α particle yields a neutron and a nuclide.
- (b) Bombardment of ^{28}Si with ^2H yields ^{29}P and another particle.
- (c) Bombardment of a nuclide with an α particle yields two neutrons and ^{244}Cf .

23.34 Name the unidentified species in each transmutation, and write a full nuclear equation:

- (a) Bombardment of a nuclide with a γ photon yields a proton, a neutron, and ^{29}Si .
- (b) Bombardment of ^{252}Cf with ^{10}B yields five neutrons and a nuclide.
- (c) Bombardment of ^{238}U with a particle yields three neutrons and ^{239}Pu .

23.35 Elements 104, 105, and 106 have been named rutherfordium (Rf), dubnium (Db), and seaborgium (Sg), respectively. These elements are synthesized from californium-249 by bombardment with carbon-12, nitrogen-15, and oxygen-18 nuclei, respectively. Four neutrons are formed in each reaction as well. Write balanced nuclear equations for the formation of these elements.

The Effects of Nuclear Radiation on Matter

23.36 The effects on matter of γ rays and α particles differ. Explain.

23.37 Why is ionizing radiation more dangerous to children than to adults?

23.38 A 135-lb person absorbs 3.3×10^{-7} J of energy from radioactive emissions. (a) How many rads does she receive? (b) How many grays (Gy) does she receive?

23.39 A 3.6-kg laboratory animal receives a single dose of 8.92×10^{-4} Gy. (a) How many rads did the animal receive? (b) How many joules did the animal absorb?

23.40 A 70.-kg person exposed to ^{90}Sr absorbs 6.0×10^5 β^- particles, each with an energy of 8.74×10^{-14} J. (a) How many grays does the person receive? (b) If the RBE is 1.0, how many millirems is this? (c) What is the equivalent dose in sieverts (Sv)?

23.41 A laboratory rat weighs 265 g and absorbs 1.77×10^{10} β^- particles, each with an energy of 2.20×10^{-13} J. (a) How many rads does the animal receive? (b) What is this dose in Gy? (c) If the RBE is 0.75, what is the equivalent dose in Sv?

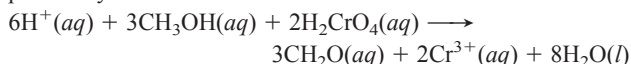
23.42 A small region of a cancer patient's brain is exposed for 24.0 min to 475 Bq of radioactivity from ^{60}Co for treatment of a tumor. If the brain mass exposed is 1.858 g and each β^- particle emitted has an energy of 5.05×10^{-14} J, what is the dose in rads?

Applications of Radioisotopes

23.43 Describe two ways that radioactive tracers are used in organisms.

23.44 Why is neutron activation analysis (NAA) useful to art historians and criminologists?

23.45 The oxidation of methanol to formaldehyde can be accomplished by reaction with chromic acid:



The reaction can be studied with the stable isotope tracer ^{18}O and mass spectrometry. When a small amount of $\text{CH}_3^{18}\text{OH}$ is present in the alcohol reactant, CH_2^{18}O forms. When a small amount of $\text{H}_2\text{Cr}^{18}\text{O}_4$ is present, H_2^{18}O forms. Does chromic acid or methanol supply the O atom to the aldehyde? Explain.

The Interconversion of Mass and Energy

(Sample Problem 23.6)

Note: Use the following data to solve the problems in this section: mass of ^1H atom = 1.007825 amu; mass of neutron = 1.008665 amu.

23.46 What is a mass difference, and how does it arise?

23.47 What is the binding energy per nucleon? Why is the binding energy per nucleon, rather than per nuclide, used to compare nuclide stability?

23.48 A ^3H nucleus decays with an energy of 0.01861 MeV. Convert this energy into (a) electron volts; (b) joules.

23.49 Arsenic-84 decays with an energy of 1.57×10^{-15} kJ per nucleus. Convert this energy into (a) eV; (b) MeV.

23.50 Cobalt-59 is the only stable isotope of this transition metal. One ^{59}Co atom has a mass of 58.933198 amu. Calculate the binding energy (a) per nucleon in MeV; (b) per atom in MeV; (c) per mole in kJ.

23.51 Iodine-131 is one of the most important isotopes used in the diagnosis of thyroid cancer. One atom has a mass of 130.906114 amu. Calculate the binding energy (a) per nucleon in MeV; (b) per atom in MeV; (c) per mole in kJ.

23.52 The ^{80}Br nuclide decays by either β^- decay or e^- capture. (a) What is the product of each process? (b) Which process releases more energy? (Masses of atoms: ^{80}Br = 79.918528 amu; ^{80}Kr = 79.916380 amu; ^{80}Se = 79.916520 amu; neglect the mass of the electron involved.)

Applications of Fission and Fusion

23.53 In what main way is fission different from radioactive decay? Are all fission events in a chain reaction identical? Explain.

23.54 What is the purpose of enrichment in the preparation of fuel rods?

23.55 Describe the nature and purpose of these components of a nuclear reactor: (a) control rods; (b) moderator.

23.56 State an advantage and a disadvantage of heavy-water reactors compared to light-water reactors.

23.57 What are the expected advantages of fusion reactors over fission reactors?

23.58 The reaction that will probably power the first commercial fusion reactor is $^3_1\text{H} + ^2_1\text{H} \longrightarrow ^4_2\text{He} + ^1_0\text{n}$. How much energy

would be produced per mole of reaction? (Masses of atoms: ^3_1H = 3.01605 amu; ^2_1H = 2.0140 amu; ^4_2He = 4.00260 amu; mass of ^1_0n = 1.008665 amu.)

23.59 Write balanced nuclear equations for the following:

(a) β^- decay of sodium-26

(b) β^- decay of francium-223

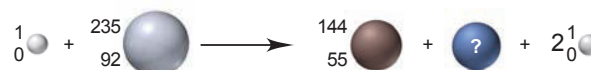
(c) α decay of $^{212}_{83}\text{Bi}$

Comprehensive Problems

Problems with an asterisk (*) are more challenging.

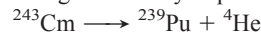
23.60 Some $^{243}_{95}\text{Am}$ was present when Earth formed, but it all decayed in the next billion years. The first three steps in this decay series are emission of an α particle, a β^- particle, and another α particle. What other isotopes were present on the young Earth in a rock that contained some $^{243}_{95}\text{Am}$?

23.61 The scene below depicts a neutron bombarding $^{235}_{92}\text{U}$:



(a) Is this an example of fission or of fusion? (b) Identify the other nuclide formed. (c) What is the most likely mode of decay of the nuclide with $Z = 55$?

23.62 Curium-243 undergoes α decay to plutonium-239:



(a) Calculate the change in mass, Δm (in kg). (Masses: ^{243}Cm = 243.0614 amu; ^{239}Pu = 239.0522 amu; ^4He = 4.0026 amu; 1 amu = 1.661×10^{-24} g.)

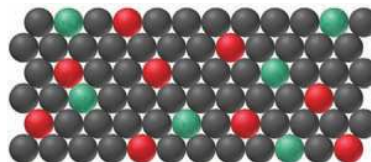
(b) Calculate the energy released in joules.

(c) Calculate the energy released in kJ/mol of reaction, and comment on the difference between this value and a typical heat of reaction for a chemical change of a few hundred kJ/mol.

23.63 Plutonium “triggers” for nuclear weapons were manufactured at the Rocky Flats plant in Colorado. An 85-kg worker inhaled a dust particle containing 1.00 μg of $^{239}_{94}\text{Pu}$, which resided in his body for 16 h ($t_{1/2}$ of ^{239}Pu = 2.41×10^4 yr; each disintegration released 5.15 MeV). (a) How many rads did he receive? (b) How many grays?

23.64 Archeologists removed some charcoal from a Native American campfire, burned it in O_2 , and bubbled the CO_2 formed into $\text{Ca}(\text{OH})_2$ solution (limewater). The CaCO_3 that precipitated was filtered and dried. If 4.58 g of the CaCO_3 had a radioactivity of 3.2 d/min, how long ago was the campfire?

23.65 $^{238}_{92}\text{U}$ ($t_{1/2}$ = 4.5×10^9 yr) begins a decay series that ultimately forms $^{206}_{82}\text{Pb}$. The scene below depicts the relative number of ^{238}U atoms (red) and ^{206}Pb atoms (green) in a mineral. If all the Pb comes from ^{238}U , calculate the age of the sample.



23.66 A 5.4- μg sample of $^{226}_{88}\text{RaCl}_2$ has a radioactivity of 1.5×10^5 Bq. Calculate $t_{1/2}$ of ^{226}Ra .

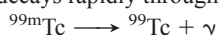
23.67 The major reaction taking place during hydrogen fusion in a young star is $4^1_1\text{H} \longrightarrow ^4_2\text{He} + 2^1_0\text{p} + 2^0_0\text{n} + \text{energy}$. How much

energy (in MeV) is released per He nucleus formed? Per mole of He? (Masses: ^1H atom = 1.007825 amu; ^4He atom = 4.00260 amu; positron = 5.48580×10^{-4} amu.)

23.68 A sample of AgCl emits 175 nCi/g. A saturated solution prepared from the solid emits 1.25×10^{-2} Bq/mL due to radioactive Ag^+ ions. What is the molar solubility of AgCl?

23.69 In the 1950s, radioactive material was spread over the land from aboveground nuclear tests. A woman drinks some contaminated milk and ingests 0.0500 g of ^{90}Sr , which is taken up by bones and teeth and not eliminated. (a) How much ^{90}Sr ($t_{1/2} = 29$ yr) is present in her body after 10 yr? (b) How long will it take for 99.9% of the ^{90}Sr to decay?

* **23.70** Technetium-99m is a metastable nuclide used in numerous cancer diagnostic and treatment programs. It is prepared just before use because it decays rapidly through γ emission:



Use the data below to determine (a) the half-life of $^{99\text{m}}\text{Tc}$, and (b) the percentage of the isotope that is lost if it takes 2.0 h to prepare and administer the dose.

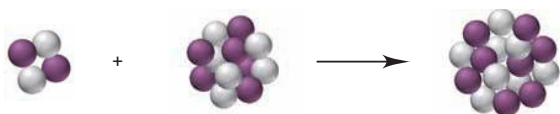
Time (h)	γ Emission (photons/s)
0	5000.
4	3150.
8	2000.
12	1250.
16	788
20	495

* **23.71** What volume of radon will be produced per hour at STP from 1.000 g of ^{226}Ra ($t_{1/2} = 1599$ yr; 1 yr = 8766 h; mass of one ^{226}Ra atom = 226.025402 amu)?

23.72 Which isotope in each pair would you predict to be more stable? Why?

- (a) $^{140}_{55}\text{Cs}$ or $^{133}_{55}\text{Cs}$ (b) $^{79}_{35}\text{Br}$ or $^{78}_{35}\text{Br}$
 (c) $^{28}_{12}\text{Mg}$ or $^{24}_{12}\text{Mg}$ (d) $^{14}_7\text{N}$ or $^{18}_7\text{N}$

23.73 The scene below represents a reaction (with neutrons gray and protons purple) that occurs during the lifetime of a star. (a) Write a balanced nuclear equation for the reaction. (b) If the mass difference is 7.7×10^{-2} amu, find the energy (kJ) released.



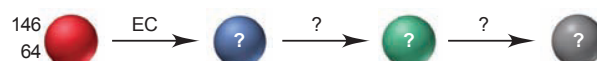
23.74 The 23rd-century starship *Enterprise* uses a substance called “dilithium crystals” as its fuel.

- (a) Assuming this material is the result of fusion, what is the product of the fusion of two ^6Li nuclei?
 (b) How much energy is released per kilogram of dilithium formed? (Mass of one ^6Li atom is 6.015121 amu.)
 (c) When four ^1H atoms fuse to form ^4He , how many positrons are released?
 (d) To determine the energy potential of the fusion processes in parts (b) and (c), compare the changes in mass per kilogram of dilithium and of ^4He .
 (e) Compare the change in mass per kilogram in part (b) to that for the formation of ^4He by the method used in current fusion reactors (Section 23.7). (For masses, see Problem 23.58.)

(f) Using early 21st-century fusion technology, how much tritium can be produced per kilogram of ^6Li in the following reaction: $^6_3\text{Li} + ^1_0\text{n} \longrightarrow ^4_2\text{He} + ^3_1\text{H}$? When this amount of tritium is fused with deuterium, what is the change in mass? How does this quantity compare with that for dilithium in part (b)?

23.75 Nuclear disarmament could be accomplished if weapons were not “replenished.” The tritium in warheads decays to helium with a half-life of 12.26 yr and must be replaced or the weapon is useless. What fraction of the tritium is lost in 5.50 yr?

23.76 Gadolinium-146 undergoes electron capture. Identify the product, and use Figure 23.2 to find the modes of decay and the other two nuclides in the series below:



23.77 A decay series starts with the synthetic isotope $^{239}_{92}\text{U}$. The first four steps are emissions of a β^- particle, another β^- , an α particle, and another α . Write a balanced nuclear equation for each step. Which natural series could be started by this sequence?

23.78 The approximate date of a San Francisco earthquake is to be found by measuring the ^{14}C activity ($t_{1/2} = 5730$ yr) of parts of a tree uprooted during the event. The tree parts have an activity of 12.9 d/min·g C, and a living tree has an activity of 15.3 d/min·g C. How long ago did the earthquake occur?

23.79 Carbon from the most recent remains of an extinct Australian marsupial, called *Diprotodon*, has a specific activity of 0.61 pCi/g. Modern carbon has a specific activity of 6.89 pCi/g. How long ago did the *Diprotodon* apparently become extinct?

* **23.80** With our early 21st-century technology, hydrogen fusion requires temperatures around 10^8 K, but lower temperatures can be used if the hydrogen is compressed. In the late 24th century, the starship *Leinad* uses such methods to fuse hydrogen at 10^6 K.

- (a) What is the kinetic energy of an H atom at 1.00×10^6 K?
 (b) How many H atoms are heated to 1.00×10^6 K from the energy of one H and one anti-H atom annihilating each other?
 (c) If these H atoms fuse into ^4He atoms (with the loss of two positrons per ^4He formed), how much energy (in J) is generated?
 (d) How much more energy is generated by the fusion in (c) than by the hydrogen-antihydrogen collision in (b)?
 (e) Should the captain of the *Leinad* change the technology and produce ^3He (mass = 3.01603 amu) instead of ^4He ?

23.81 Seaborgium-263 (Sg; $Z = 106$) was the first isotope of this element synthesized. It was made, together with four neutrons, by bombarding californium-249 with oxygen-18. The nuclide then decayed by three α emissions. Write balanced equations for the synthesis and three decay steps of ^{263}Sg .

23.82 Representations of three nuclei (with neutrons gray and protons purple) are shown below. Nucleus 1 is stable, but 2 and 3 are not. (a) Write the symbol for each isotope. (b) What is (are) the most likely mode(s) of decay for 2 and 3?



1



2



3