21 NUCLEAR CHEMISTRY

WHAT'S AHEAD

21.1 | Radioactivity and Nuclear Equations

All energy that fuels life on Earth comes ultimately from sunlight. Life on Earth could not exist without energy from the sun, but where does the sun get its energy? Stars, including our sun, use **nuclear reactions** that involve changes in atomic nuclei to generate their energy. For example, the sun produces energy by fusing hydrogen atoms to form helium, releasing vast amounts of energy in the process. The fusion of hydrogen to form helium is the dominant nuclear reaction for most of a star's lifetime. Towards the end of its life, the hydrogen in the star's core is exhausted and the helium atoms fuse to form progressively heavier elements. A select few stars end their lives in dramatic supernova explosions. The nuclear reactions that occur when a star goes supernova are responsible for the existence of all naturally occurring elements heavier than nickel.

Nuclear chemistry is the study of nuclear reactions, with an emphasis on their uses in chemistry and their effects on biological systems. Nuclear chemistry affects our lives in many ways, particularly in energy and medical applications. In rad iation therapy, for example, gamma rays from a radioactive substance such as cobalt-60 are directed to cancerous tumors to destroy them. Positron emission tomography (PET) is one example of a medical diagnostic tool that relies on decay of a radioactive element injected into the body.

The emanation of gamma rays from cobalt-60 is an example of a *nuclear reaction*, in which a change in matter originates in the nucleus of an atom. When nuclei change spontaneously, emitting radiation, they are said to be **radioactive**.

Radioactivity is also used to help determine the mechanisms of chemical reactions, to trace the movement of atoms in biological systems and the environment and to date historical artifacts.

Nuclear reactions are also used to generate electricity. Roughly 15% of the electricity generated worldwide comes from nuclear power plants, though the percentage varies from one country to the next.

The use of nuclear energy for power generation is a controversial social and political issue because of the public's conceptions about the safety of nuclear reactors and, more importantly, the difficulty of disposing of nuclear reactor waste.

At the end of this section, you should be able to

- Write balanced nuclear equations

To understand nuclear reactions, we must review and develop some ideas introduced in Section 2.3:

- Two types of subatomic particles reside in the nucleus: *protons* and *neutrons*. We will refer to these particles as **nucleons**.
- All atoms of a given element have the same number of protons; this number is the element's *atomic number*.
- Atoms of a given element can have different numbers of neutrons, which means they can have different mass numbers. The *mass number* is the total number of nucleons in the nucleus.
- Atoms with the same atomic number but different mass numbers are known as *isotopes*.

The different isotopes of an element are distinguished by their mass numbers. For example, the three naturally occurring isotopes of uranium are uranium-234, uranium-235, and uranium-238, where the numerical suffixes represent the mass numbers. These isotopes are also written $^{234}_{92}U$, $^{235}_{92}U$, and $^{238}_{92}U$, where the superscript is the mass number and the subscript is the atomic number.*

Different isotopes of an element have different natural abundances. For example, 99.3% of naturally occurring uranium is uranium-238, 0.7% is uranium-235, and only a trace is uranium-234. Different isotopes of an element also have different stabilities. Indeed, the nuclear properties of any given isotope depend on the number of protons and neutrons in its nucleus.

A *nuclide* is a nucleus containing a specified number of protons and neutrons. Nuclides that are radioactive are called **radionuclides**, and atoms containing these nuclei are called **radioisotopes**.

*As noted in Section 2.3, we often do not explicitly write the atomic number of an isotope because the element symbol is specific to the atomic number. In studying nuclear chemistry, however, it is often useful to include the atomic number in order to help us keep track of changes in the nuclei.

Nuclear Equations

Most nuclei in nature are stable and remain intact indefinitely. Radionuclides, however, are unstable and spontaneously emit particles and electromagnetic radiation. Emission of radiation is one of the ways in which an unstable nucleus is transformed into a more stable one that has less energy. The emitted radiation is the carrier of the excess energy. Uranium-238, for example, is radioactive, undergoing a nuclear reaction emitting helium-4 nuclei. The helium-4 particles are known as **alpha (α) particles**, and a stream of them is called *alpha radiation*. When a ${}^{238}_{92}U$ nucleus loses an alpha particle, the remaining fragment has an atomic number of 90 and a mass number of 234. The element with atomic number 90 is Th, thorium. Therefore, the products of uranium-238 decomposition are an alpha particle and a thorium-234 nucleus. We represent this reaction by the *nuclear equation*

$$ {}^{238}_{92}\text{U} \longrightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He} \qquad [21.1] $$

When a nucleus spontaneously decomposes in this way, it is said to be *radioactive* and to have decayed or to have undergone *radioactive decay*. Because an alpha particle is involved in this reaction, scientists also describe the process as **alpha decay** or **alpha emission**.

In Equation 21.1 the sum of the mass numbers is the same on both sides of the equation ($238 = 234 + 4$). Likewise, the sum of the atomic numbers on both sides of the equation is equal ($92 = 90 + 2$). Mass numbers and atomic numbers must be balanced in all nuclear equations.

The radioactive properties of the nucleus in an atom are independent of the chemical state of the atom. In writing nuclear equations, therefore, we are not concerned with the chemical form (element or compound) of the atom in which the nucleus resides.

Sample Exercise 21.1
Predicting the Product of a Nuclear Reaction

What product is formed when radium-226 undergoes alpha emission?

SOLUTION

Analyze We are asked to determine the nucleus that results when radium-226 loses an alpha particle.

Plan We can best do this by writing a balanced nuclear reaction for the process.

Solve The periodic table shows that radium has an atomic number of 88. The complete chemical symbol for radium-226 is therefore ${}^{226}_{88}Ra$. An alpha particle is a helium-4 nucleus, and so its symbol is ${}^{4}_{2}He$. The alpha particle is a product of the nuclear reaction, and so the equation is of the form

$$ {}^{226}_{88}\text{Ra} \longrightarrow {}^{A}_{Z}\text{X} + {}^{4}_{2}\text{He} $$

where A is the mass number of the product nucleus and Z is its atomic number. Mass numbers and atomic numbers must balance, so

$$ 226 = A + 4 $$

and

$$ 88 = Z + 2 $$

Hence,

$$ A = 222 \quad \text{and} \quad Z = 86 $$

Again, from the periodic table, the element with $Z = 86$ is radon (Rn). The product, therefore, is ${}^{222}_{86}Rn$, and the nuclear equation is

$$ {}^{226}_{88}\text{Ra} \longrightarrow {}^{222}_{86}\text{Rn} + {}^{4}_{2}\text{He} $$

Practice Exercise

What product forms when plutonium-238 undergoes alpha emission?
(**a**) Plutonium-234 (**b**) Uranium-234 (**c**) Uranium-238
(**d**) Thorium-236 (**e**) Neptunium-237

Types of Radioactive Decay

The three most common kinds of radiation given off when a radionuclide decays are alpha (α), beta (β), and gamma (γ) radiation. **Table 21.1** summarizes some of the important properties of these types of radiation.

Alpha Radiation As described earlier, alpha radiation consists of a stream of helium-4 nuclei known as alpha particles, which we denote as ${}^{4}_{2}He$ or simply α.

TABLE 21.1 Properties of Alpha, Beta, and Gamma Radiation

	Type of Radiation		
Property	α	β	γ
Charge	2+	1−	0
Mass	6.64×10^{-24} g	9.11×10^{-28} g	0
Relative penetrating power	1	100	10,000
Nature of radiation	${}^{4}_{2}\text{He}$ nuclei	Electrons	High-energy photons

Beta Radiation *Beta radiation* consists of streams of **beta (β) particles**, which are high-speed electrons emitted by an unstable nucleus. Beta particles are represented in nuclear equations by ${}^{0}_{-1}\text{e}$ or more commonly by β^-. The superscript 0 indicates that the mass of the electron is exceedingly small relative to the mass of a nucleon. The subscript −1 represents the negative charge of the beta particle, which is opposite that of the proton.

Iodine-131 is an isotope that undergoes decay by **beta emission**:

$$ {}^{131}_{53}\text{I} \longrightarrow {}^{131}_{54}\text{Xe} + {}^{0}_{-1}\text{e} \qquad [21.2] $$

We see from this equation that beta decay causes the atomic number of the reactant to increase from 53 to 54, which means a proton was created. Therefore, beta emission is equivalent to the conversion of a neutron (${}^{1}_{0}\text{n}$ or simply n) to a proton (${}^{1}_{1}\text{H}$ or simply p):

$$ {}^{1}_{0}\text{n} \longrightarrow {}^{1}_{1}\text{H} + {}^{0}_{-1}\text{e} \quad \text{or} \quad \text{n} \longrightarrow \text{p} + \beta^- \qquad [21.3] $$

Just because an electron is emitted from a nucleus in beta decay, we should not think that the nucleus is composed of these particles any more than we consider a match to be composed of sparks simply because it gives them off when struck. The beta-particle electron comes into being only when the nucleus undergoes a nuclear reaction. Furthermore, the speed of the beta particle is sufficiently high that it does not end up in an orbital of the decaying atom.

Gamma Radiation *Gamma (γ) radiation* (or **gamma rays**) consists of high-energy photons (that is, electromagnetic radiation of very short wavelength). It changes neither the atomic number nor the mass number of a nucleus and is represented as either ${}^{0}_{0}\gamma$ or simply γ. Gamma radiation usually accompanies other radioactive emission because it represents the energy lost when the nucleons in a nuclear reaction reorganize into more stable arrangements. Often gamma rays are not explicitly shown when writing nuclear equations.

Positron Emission and Electron Capture Two other types of radioactive decay are positron emission and electron capture. A **positron**, represented as ${}^{0}_{+1}\text{e}$, or simply as β^+, is a particle that has the same mass as an electron (thus, we use the letter e and superscript 0 for the mass) but the opposite charge (represented by the +1 subscript).*

The isotope carbon-11 decays by **positron emission**:

$$ {}^{11}_{6}\text{C} \longrightarrow {}^{11}_{5}\text{B} + {}^{0}_{+1}\text{e} \qquad [21.4] $$

Positron emission causes the atomic number of the reactant in this equation to decrease from 6 to 5. In general, positron emission has the effect of converting a proton to a neutron, thereby decreasing the atomic number of the nucleus by 1 while not changing the mass number:

$$ {}^{1}_{1}\text{p} \longrightarrow {}^{1}_{0}\text{n} + {}^{0}_{+1}\text{e} \quad \text{or} \quad \text{p} \longrightarrow \text{n} + \beta^+ \qquad [21.5] $$

Electron capture is the capture by the nucleus of an electron from the electron cloud surrounding the nucleus, as in this rubidium-81 decay:

$$ {}^{81}_{37}\text{Rb} + {}^{0}_{-1}\text{e}\,(\text{orbital electron}) \longrightarrow {}^{81}_{36}\text{Kr} \qquad [21.6] $$

*The positron has a very short life because it is annihilated when it collides with an electron, producing gamma rays: ${}^{0}_{+1}\text{e} + {}^{0}_{-1}\text{e} \longrightarrow 2\,{}^{0}_{0}\gamma$.

TABLE 21.3 Types of Radioactive Decay

Type	Nuclear Equation	Change in Atomic Number	Change in Mass Number
Alpha emission	${}^{A}_{Z}\mathrm{X} \longrightarrow {}^{A-4}_{Z-2}\mathrm{Y} + {}^{4}_{2}\mathrm{He}$	−2	−4
Beta emission	${}^{A}_{Z}\mathrm{X} \longrightarrow {}_{Z+1}^{\;\;A}\mathrm{Y} + {}^{\;\;0}_{-1}\mathrm{e}$	+1	Unchanged
Positron emission	${}^{A}_{Z}\mathrm{X} \longrightarrow {}_{Z-1}^{\;\;A}\mathrm{Y} + {}^{\;\;0}_{+1}\mathrm{e}$	−1	Unchanged
Electron capture*	${}^{A}_{Z}\mathrm{X} + {}^{\;\;0}_{-1}\mathrm{e} \longrightarrow {}_{Z-1}^{\;\;A}\mathrm{Y}$	−1	Unchanged

*The electron captured comes from the electron cloud surrounding the nucleus.

TABLE 21.2 Particles Found in Nuclear Reactions

Particle	Symbol
Neutron	${}^{1}_{0}\mathrm{n}$ or n
Proton	${}^{1}_{1}\mathrm{H}$ or p
Electron	${}^{\;\;0}_{-1}\mathrm{e}$
Alpha particle	${}^{4}_{2}\mathrm{He}$ or α
Beta particle	${}^{\;\;0}_{-1}\mathrm{e}$ or β^-
Positron	${}^{\;\;0}_{+1}\mathrm{e}$ or β^+

Because the electron is consumed rather than formed in the process, it is shown on the reactant side of the equation. Electron capture, like positron emission, has the effect of converting a proton to a neutron:

$$ {}^{1}_{1}\mathrm{p} + {}^{\;\;0}_{-1}\mathrm{e} \longrightarrow {}^{1}_{0}\mathrm{n} \qquad [21.7]$$

Table 21.2 summarizes the symbols used to represent the particles commonly encountered in nuclear reactions. The various types of radioactive decay are summarized in **Table 21.3**.

Sample Exercise 21.2
Writing Nuclear Equations

Write nuclear equations for (**a**) mercury-201 undergoing electron capture; (**b**) thorium-231 decaying to protactinium-231.

SOLUTION

Analyze We must write balanced nuclear equations in which the masses and charges of reactants and products are equal.

Plan We can begin by writing the complete chemical symbols for the nuclei and decay particles that are given in the problem.

Solve

(**a**) The information given in the question can be summarized as

$$ {}^{201}_{\;80}\mathrm{Hg} + {}^{\;\;0}_{-1}\mathrm{e} \longrightarrow {}^{A}_{Z}\mathrm{X}$$

The mass numbers must have the same sum on both sides of the equation:

$$201 + 0 = A$$

Thus, the product nucleus must have a mass number of 201. Similarly, balancing the atomic numbers gives

$$80 - 1 = Z$$

Thus, the atomic number of the product nucleus must be 79, which identifies it as gold (Au):

$$ {}^{201}_{\;80}\mathrm{Hg} + {}^{\;\;0}_{-1}\mathrm{e} \longrightarrow {}^{201}_{\;79}\mathrm{Au}$$

(**b**) In this case, we must determine what type of particle is emitted in the course of the radioactive decay:

$$ {}^{231}_{\;90}\mathrm{Th} \longrightarrow {}^{231}_{\;91}\mathrm{Pa} + {}^{A}_{Z}\mathrm{X}$$

From $231 = 231 + A$ and $90 = 91 + Z$, we deduce $A = 0$ and $Z = -1$. According to Table 21.2, the particle with these characteristics is the beta particle (electron). We therefore write

$$ {}^{231}_{\;90}\mathrm{Th} \longrightarrow {}^{231}_{\;91}\mathrm{Pa} + {}^{\;\;0}_{-1}\mathrm{e} \quad \text{or} \quad {}^{231}_{\;90}\mathrm{Th} \longrightarrow {}^{231}_{\;91}\mathrm{Pa} + \beta^-$$

Practice Exercise

The radioactive decay of thorium-232 occurs in multiple steps, called a *radioactive decay chain*. The second product produced in this chain is actinium-228. Which of the following processes could lead to this product starting with thorium-232?

(**a**) Alpha decay followed by beta emission
(**b**) Beta emission followed by electron capture
(**c**) Positron emission followed by alpha decay
(**d**) Electron capture followed by positron emission
(**e**) More than one of these is consistent with the observed transformation.

Self-Assessment Exercise

21.1 Carbon-11 decays by positron emission and is one radioisotope used in PET scans. What is the nuclear equation for its decay?

(a) ${}^{11}_{6}\mathrm{C} \longrightarrow {}^{11}_{5}\mathrm{C} + {}^{\;\;0}_{+1}\mathrm{e}$
(b) ${}^{11}_{6}\mathrm{C} \longrightarrow {}^{11}_{5}\mathrm{B} + {}^{\;\;0}_{+1}\mathrm{e}$
(b) ${}^{11}_{6}\mathrm{C} \longrightarrow {}^{11}_{7}\mathrm{N} + {}^{\;\;0}_{+1}\mathrm{e}$

Exercises

21.2 Indicate the number of protons and neutrons in the following nuclei: (**a**) $^{214}_{83}Bi$, (**b**) $^{210}_{82}Pb$, (**c**) uranium-235.

21.3 What do these symbols stand for? (**a**) $^{0}_{0}\gamma$, (**b**) $^{4}_{2}He$, (**c**) $^{1}_{0}n$.

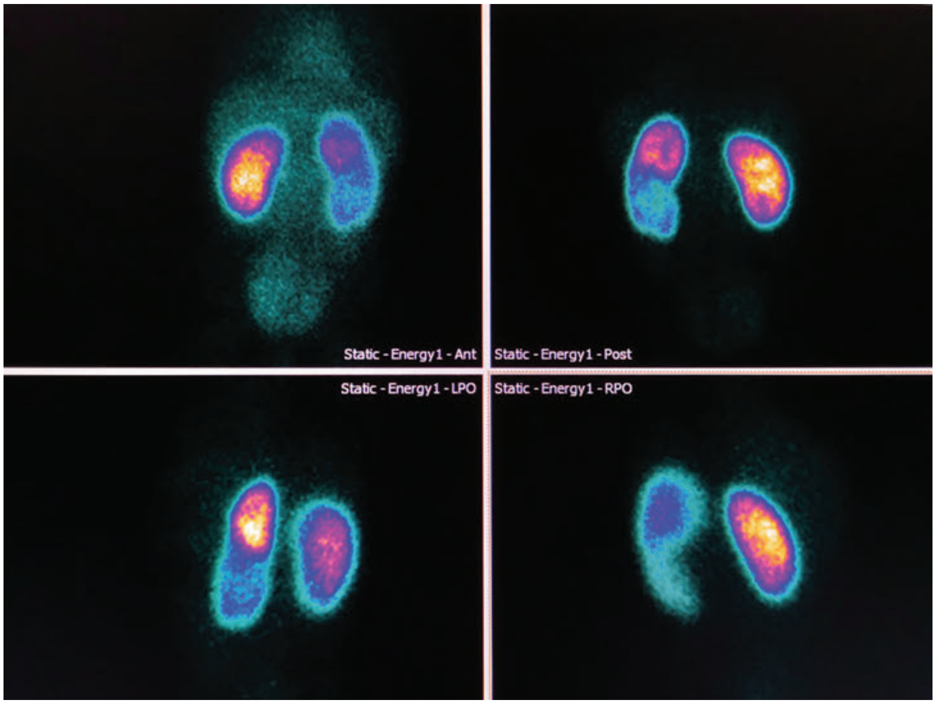
21.4 Write balanced nuclear equations for the following transformations: (**a**) polonium-210 emits alpha particle; (**b**) neptunium-235 undergoes electron capture; (**c**) fluorine-18 emits beta particle; (**d**) carbon-14 decays by beta emission.

21.5 What particle is produced during the following decay processes: (**a**) actinium-215 decays to francium-211; (**b**) boron-13 decays to carbon-13; (**c**) holmium-151 decays to terbium-147; (**d**) carbon-11 decays to boron-11?

Answers to Self-Assessment Exercise

21.1 (b)

21.2 Patterns of Nuclear Stability



One of the most useful radionuclides in diagnostic medicine is technetium-99m, ^{99m}Tc. The 'm' stands for 'metastable,' indicating that the isotope is in an excited state. The excess energy the isotope contains is emitted as easily detectable gamma rays in the absence of alpha or beta particles. The isotope can be used to study the brain, heart, thyroid, lungs, liver, kidneys (pictured), blood, and in the detection of bone cancer. This wide applicability and a short lifetime mean that nearly 95% decays to ^{99}Tc in 24 hours, by which time most of it has been flushed from the body. In turn, technetium-99 is unusual because it emits β radiation in the absence of γ radiation and this, in combination with a long half-life, makes it useful for the calibration of some types of equipment.

In this section, we look at nuclear stability. Some nuclides, such as $^{12}_{6}C$ and $^{13}_{6}C$ are stable, whereas others, such as $^{14}_{6}C$, are unstable and undergo radioactive decay. Why does a small difference in the number of neutrons affect the stability of a nuclide? No single rule allows us to predict whether a particular nucleus is radioactive and, if it is, how it might decay. However, several empirical observations can help us predict the stability of a nucleus.

By the end of this section, you should be able to

- Predict nuclear stability and expected type of nuclear decay from the neutron-to-proton ratio of an isotope.
- Write balanced nuclear equations for nuclear transmutation

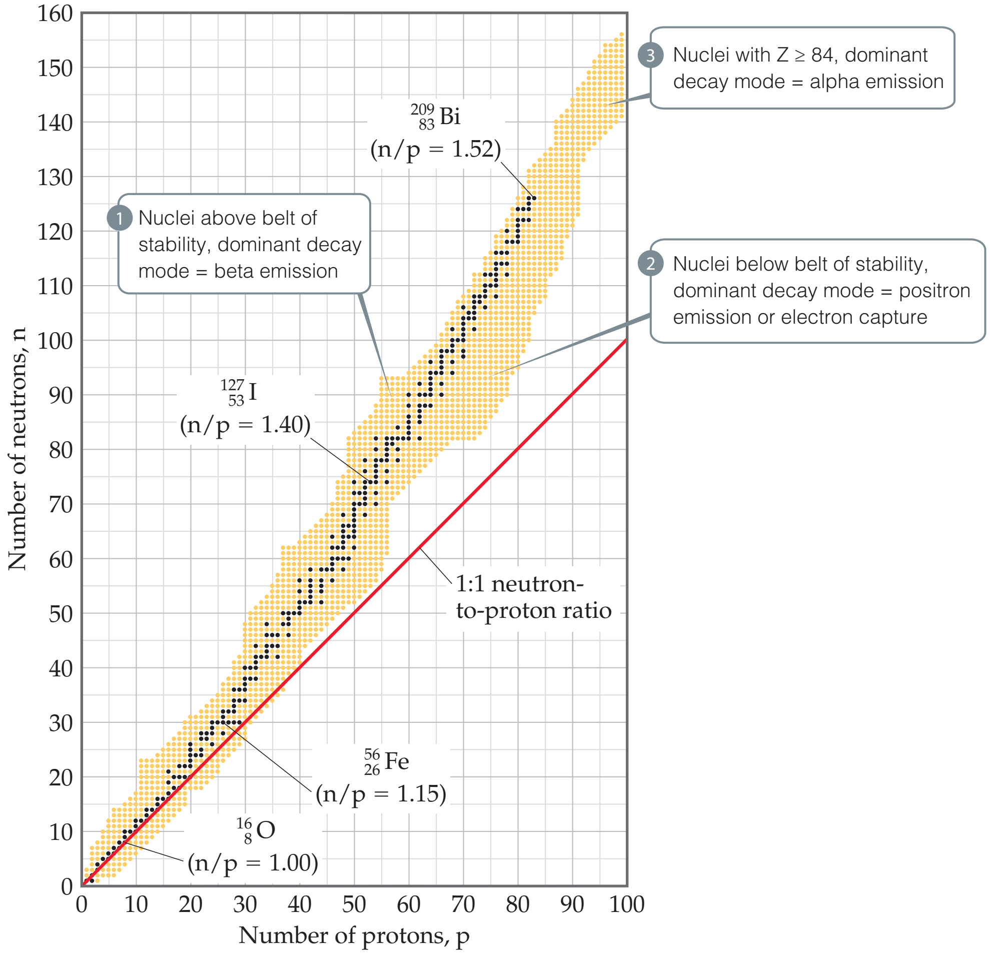
Neutron-to-Proton Ratio

Because like charges repel each other, it may seem surprising that a large number of protons can reside within the small volume of the nucleus. At close distances, however, a strong force of attraction, called the *strong nuclear force*, exists between nucleons. Neutrons are intimately involved in this attractive force.

All nuclei other than $^{1}_{1}H$ contain neutrons. As the number of protons in a nucleus increases, there is an ever greater need for neutrons to counteract the proton–proton repulsions. Stable nuclei with atomic numbers up to about 20 have approximately equal numbers of neutrons and protons. For nuclei with atomic number above 20, the number of neutrons exceeds the number of protons. Indeed, the number of neutrons necessary to create a stable nucleus increases more rapidly than the number of protons. Thus, the neutron-to-proton ratios of stable nuclei increase with increasing atomic number, as illustrated by the following isotopes: $^{12}_{6}C$ (n/p = 1), manganese, $^{55}_{25}Mn$ (n/p = 1.20), and gold, $^{197}_{79}Au$ (n/p = 1.49).

Figure 21.1 shows all known isotopes of the elements up to $Z = 100$ plotted according to their numbers of protons and neutrons. Notice how the plot goes above the line for 1:1 neutron-to-proton for heavier elements. The dark blue dots in the figure represent stable (nonradioactive) isotopes, the remainder being radioactive isotopes. The region of the graph covered by these dark blue dots is known as the *belt of stability*. The belt of

Go Figure Estimate the optimal number of neutrons for a nucleus containing 70 protons.



▲ **Figure 21.1 Stable and radioactive isotopes as a function of numbers of neutrons and protons in a nucleus.** The stable nuclei (dark blue dots) define a region known as the belt of stability.

stability ends at element 83 (bismuth), which means that *all nuclei with 84 or more protons are radioactive*. For example, all isotopes of uranium, $Z = 92$, are radioactive.

Different radionuclides decay in different ways. The type of decay that occurs depends largely on a nuclide's neutron-to-proton ratio and how it compares with the ratio of nearby nuclei that lie within the belt of stability. We can envision three general situations, which are labeled 1, 2, and 3, in Figure 21.1.

1. **Nuclei above the belt of stability (high neutron-to-proton ratios).** These neutron-rich nuclei can lower their ratio and thereby move toward the belt of stability by emitting a beta particle because beta emission decreases the number of neutrons and increases the number of protons (Equation 21.3).
2. **Nuclei below the belt of stability (low neutron-to-proton ratios).** These proton-rich nuclei can increase their ratio and so move closer to the belt of stability by either positron emission or electron capture because both decays increase the number of neutrons and decrease the number of protons (Equations 21.5 and 21.7). Positron emission is more common among lighter nuclei. Electron capture becomes increasingly common as the nuclear charge increases.
3. **Nuclei with atomic numbers $\geq$ 84.** These heavy nuclei tend to undergo alpha emission, which decreases both the number of neutrons and the number of protons by two, moving the nucleus diagonally toward the belt of stability.

Sample Exercise 21.3
Predicting Modes of Nuclear Decay

Predict the mode of decay of (**a**) carbon-14, (**b**) xenon-118.

SOLUTION

Analyze We are asked to predict the modes of decay of two nuclei.

Plan To do this, we must locate the respective nuclei in Figure 21.1 and determine their positions with respect to the belt of stability in order to predict the most likely mode of decay.

Solve

(**a**) Carbon is element 6. Thus, carbon-14 has 6 protons and $14 - 6 = 8$ neutrons, giving it a neutron-to-proton ratio of 1.25. Elements with $Z < 20$ normally have stable nuclei that contain approximately equal numbers of neutrons and protons ($n/p = 1$). Thus, carbon-14 is located above the belt of stability, and we expect it to decay by emitting a beta particle to decrease the n/p ratio:

$$^{14}_{6}\text{C} \longrightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e}$$

This is indeed the mode of decay observed for carbon-14, a reaction that lowers the n/p ratio from 1.25 to 1.0.

(**b**) Xenon is element 54. Thus, xenon-118 has 54 protons and $118 - 54 = 64$ neutrons, giving it an n/p ratio of 1.18. According to Figure 21.1, stable nuclei in this region of the belt of stability have higher neutron-to-proton ratios than xenon-118. The nucleus can increase this ratio by either positron emission or electron capture:

$$^{118}_{54}\text{Xe} \longrightarrow {}^{118}_{53}\text{I} + {}^{0}_{+1}\text{e}$$

$$^{118}_{54}\text{Xe} + {}^{0}_{-1}\text{e} \longrightarrow {}^{118}_{53}\text{I}$$

In this case, both modes of decay are observed.

Comment Keep in mind that our guidelines do not always work. For example, thorium-233, which we might expect to undergo alpha decay, actually undergoes beta emission. Furthermore, a few radioactive nuclei lie within the belt of stability. Both $^{146}_{60}\text{Nd}$ and $^{148}_{60}\text{Nd}$, for example, are stable and lie in the belt of stability. $^{147}_{60}\text{Nd}$, however, which lies between them, is radioactive.

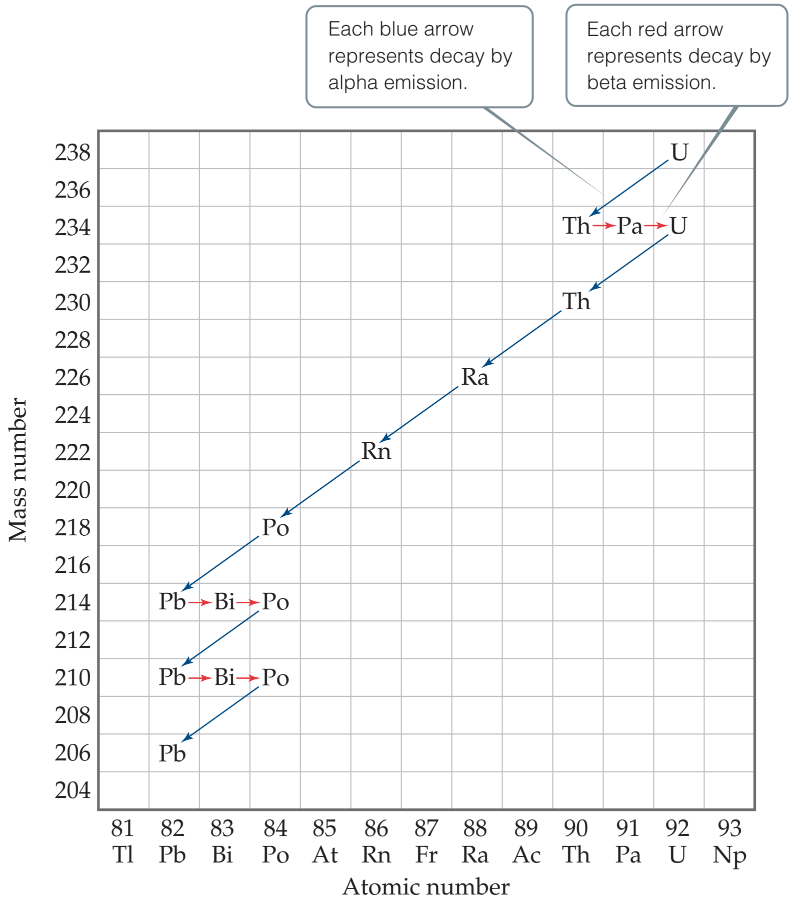
Practice Exercise

Which of the following radioactive nuclei is most likely to decay via emission of a β^- particle?
(**a**) nitrogen-13 (**b**) magnesium-23 (**c**) rubidium-83
(**d**) iodine-131 (**e**) neptunium-237

Radioactive Decay Chains

Some nuclei cannot gain stability by a single emission. Consequently, a series of successive emissions occurs as shown for uranium-238 in **Figure 21.2**. Decay continues until a stable nucleus—lead-206 in this case—is formed. A series of nuclear reactions that begins with an unstable nucleus and terminates with a stable one is known as a **radioactive decay chain** or a **nuclear disintegration series**. Three such series occur in nature: uranium-238 to lead-206, uranium-235 to lead-207, and thorium-232 to lead-208. All of the decay processes in these series are either alpha emissions or beta emissions.

Go Figure Write the nuclear equation for the step shown for the first decay of Th.



▲ **Figure 21.2 Nuclear decay chain for uranium-238.** The decay continues until the stable nucleus ^{206}Pb is formed.

Further Observations

Two further observations can help us to predict stable nuclei:

- Nuclei with the **magic numbers** of 2, 8, 20, 28, 50, or 82 protons or 2, 8, 20, 28, 50, 82, or 126 neutrons are generally more stable than nuclei that do not contain these numbers of nucleons.
- Nuclei with even numbers of protons, neutrons, or both are more likely to be stable than those with odd numbers of protons and/or neutrons. Approximately 60% of stable nuclei have an even number of both protons and neutrons, whereas less than 2% have odd numbers of both (**Table 21.4**).

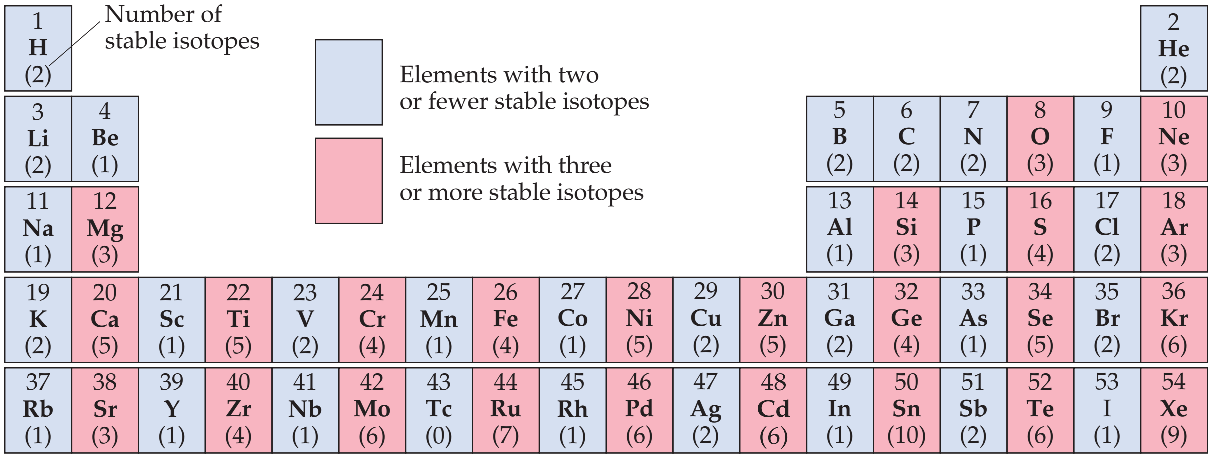
These observations can be understood in terms of the *shell model of the nucleus*, in which nucleons are described as residing in shells analogous to the shell structure for electrons in atoms. Just as certain numbers of electrons correspond to stable filled-shell electron configurations, so too do certain numbers (known as magic numbers) of nucleons represent filled shells in nuclei.

There are several examples of the stability of nuclei with magic numbers of nucleons. For example, the radioactive series in Figure 21.2 ends with the stable $^{206}_{82}Pb$ nucleus, which has a magic number of protons (82). Another example is the observation that tin, which has a magic number of protons (50), has ten stable isotopes, more than any other element.

TABLE 21.4 Number of Stable Isotopes with Even and Odd Numbers of Protons and Neutrons

Number of Stable Isotopes	Proton Number	Neutron Number
157	Even	Even
53	Even	Odd
50	Odd	Even
5	Odd	Odd

Go Figure Among the elements shown here, how many have an even number of protons and fewer than three stable isotopes? How many have an odd number of protons and more than two stable isotopes?



▲ **Figure 21.3** **Number of stable isotopes for elements 1–54.**

Evidence also suggests that pairs of protons and pairs of neutrons have a special stability, analogous to the pairs of electrons in molecules. This evidence accounts for the observation that stable nuclei with an even number of protons and/or neutrons are far more numerous than those with odd numbers. The preference for even numbers of protons is illustrated in **Figure 21.3**, which shows the number of stable isotopes for all elements up to Xe. Notice that once we move past nitrogen, the elements with an odd number of protons invariably have fewer stable isotopes than their neighbors with an even number of protons.

Nuclear Transmutations

Thus far we have examined nuclear reactions in which a nucleus decays spontaneously. A nucleus can also change identity if it is struck by a neutron or by another nucleus. Nuclear reactions induced in this way are known as **nuclear transmutations**.

In 1919, Ernest Rutherford performed the first conversion of one nucleus into another, using alpha particles emitted by radium to convert nitrogen-14 into oxygen-17:

$$^{14}_{7}\text{N} + {}^{4}_{2}\text{He} \longrightarrow {}^{17}_{8}\text{O} + {}^{1}_{1}\text{H} \quad \text{or} \quad {}^{14}_{7}\text{N} + \alpha \longrightarrow {}^{17}_{8}\text{O} + \text{p} \qquad [21.8]$$

Such reactions have allowed scientists to synthesize hundreds of radioisotopes in the laboratory.

A shorthand notation often used to represent nuclear transmutations lists the target nucleus, the bombarding particle and the ejected particle in parentheses, followed by the product nucleus. Using this condensed notation we see that, Equation 21.8 becomes

Target nucleus — Product nucleus

$$^{14}_{7}\text{N}(\alpha, \text{p})^{17}_{8}\text{O}$$

Bombarding particle — Ejected particle

Sample Exercise 21.4
Writing a Balanced Nuclear Equation

Write the balanced nuclear equation for the process summarized as ${}^{27}_{13}Al(n, \alpha){}^{24}_{11}Na$.

SOLUTION

Analyze We must go from the condensed descriptive form of the reaction to the balanced nuclear equation.

Plan We arrive at the balanced equation by writing n and α, each with its associated subscripts and superscripts.

Solve The n is the abbreviation for a neutron (${}^{1}_{0}n$), and α represents an alpha particle (${}^{4}_{2}He$). The neutron is the bombarding particle, and the alpha particle is a product. Therefore, the nuclear equation is

$$ {}^{27}_{13}Al + {}^{1}_{0}n \longrightarrow {}^{24}_{11}Na + {}^{4}_{2}He \quad \text{or} \quad {}^{27}_{13}Al + n \longrightarrow {}^{24}_{11}Na + \alpha $$

Practice Exercise

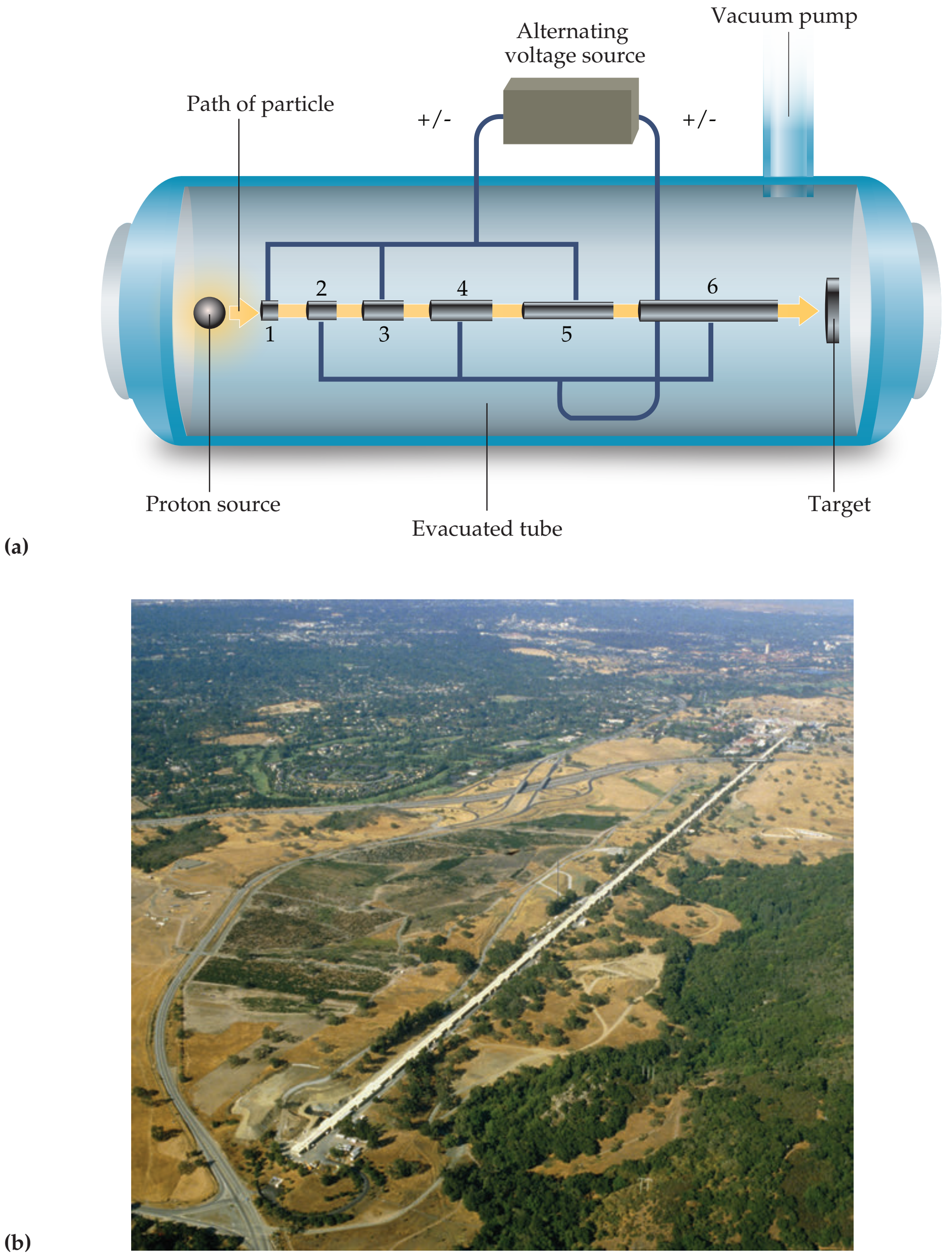
Consider the following nuclear transmutation: ${}^{238}_{92}U(n, \beta^-)X$. What is the identity of nucleus X?
(a) ${}^{238}_{93}Np$ **(b)** ${}^{239}_{92}U$ **(c)** ${}^{239}_{92}U^+$ **(d)** ${}^{235}_{90}Th$ **(e)** ${}^{239}_{93}Np$

Accelerating Charged Particles

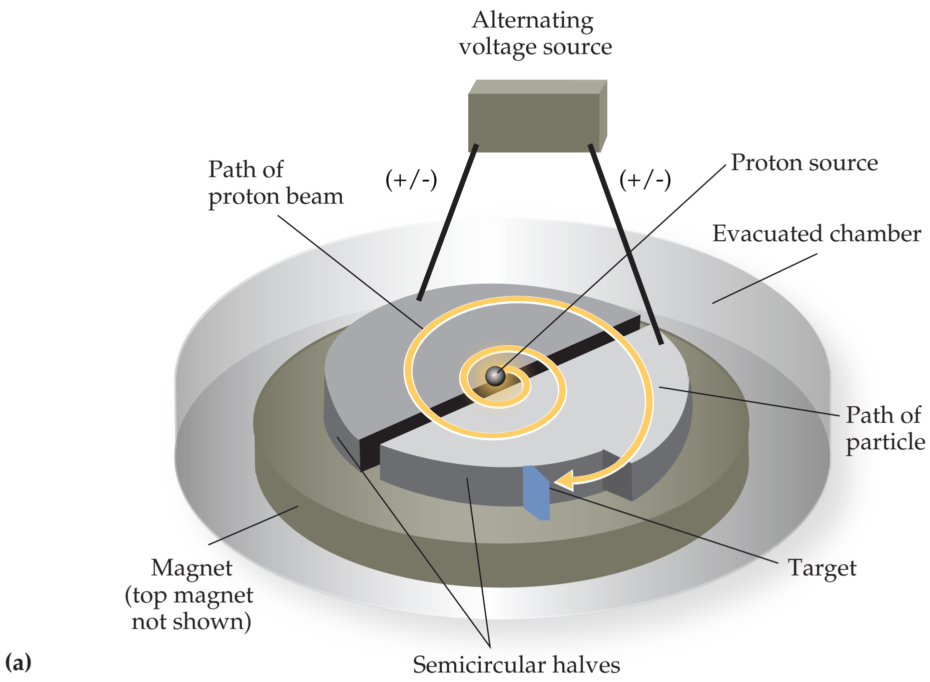
Alpha particles and other positively charged particles must move very fast to overcome the electrostatic repulsion between them and the target nucleus. The higher the nuclear charge on either the bombarding particle or the target nucleus, the faster the bombarding particle must move to bring about a nuclear reaction. Many methods have been devised to accelerate charged particles, using strong magnetic and electrostatic fields. These **particle accelerators**, popularly called "atom smashers," bear such names as *cyclotron* and *synchrotron*.

A common theme of all particle accelerators is the need to create charged particles so that they can be manipulated by electric and magnetic fields. In addition, the region through which the particles move must be kept at high vacuum so that they do not collide with any gas-phase molecules.

Figure 21.4(**a**) shows a multistage linear accelerator. A charged particle, such as a proton, is accelerated through a series of tubes of increasing length. The electrical charge on the tubes is changed from positive to negative, so that the particle is always attracted to the tube it is approaching and repelled by the one it is leaving. As a result, the particle accelerates until it has sufficient kinetic energy to smash into a target nucleus. Figure 21.4(**b**) shows the Stanford linear accelerator, which is 3.2 km in length.



▲ **Figure 21.4 The linear accelerator.**



▲ **Figure 21.5** **The cyclotron.**

Figure 21.5(**a**) shows a *cyclotron*. In this device, charged particles move in a spiral path within two D-shaped electrodes. Alternating charges on the electrodes accelerate the particles, while magnets above and below the device constrain the particles to a spiral path of increasing radius. In a *synchrotron*, the magnetic fields are synchronized so that the particle moves in a circular rather than a spiral path. Figure 21.5(**b**) shows the Fermi National Accelerator Lab at Batavia, Illinois, which has a circumference of 6.3 km.

Reactions Involving Neutrons

Most synthetic isotopes used in medicine and scientific research are made using neutrons as the bombarding particles. Because neutrons are neutral, they are not repelled by the nucleus. Consequently, they do not need to be accelerated to cause nuclear reactions. The neutrons are produced in nuclear reactors (see Section 21.5). For example, cobalt-60, which is used in cancer radiation therapy, is produced by neutron capture. Iron-58 is placed in a nuclear reactor and bombarded by neutrons to trigger the following sequence of reactions:

$$^{58}_{26}\text{Fe} + ^{1}_{0}\text{n} \longrightarrow ^{59}_{26}\text{Fe} \qquad [21.9]$$

$$^{59}_{26}\text{Fe} \longrightarrow ^{59}_{27}\text{Co} + ^{0}_{-1}\text{e} \qquad [21.10]$$

$$^{59}_{27}\text{Co} + ^{1}_{0}\text{n} \longrightarrow ^{60}_{27}\text{Co} \qquad [21.11]$$

Transuranium Elements

Nuclear transmutations have been used to produce the elements with atomic number above 92, collectively known as the **transuranium elements** because they follow uranium in the periodic table. Elements 93 (neptunium, Np) and 94 (plutonium, Pu) were produced in 1940 by bombarding uranium-238 with neutrons:

$$^{238}_{92}\text{U} + ^{1}_{0}\text{n} \longrightarrow ^{239}_{92}\text{U} \longrightarrow ^{239}_{93}\text{Np} + ^{0}_{-1}\text{e} \qquad [21.12]$$

$$^{239}_{93}\text{Np} \longrightarrow ^{239}_{94}\text{Pu} + ^{0}_{-1}\text{e} \qquad [21.13]$$

Elements with still larger atomic numbers are normally formed in small quantities in particle accelerators. Curium-242, for example, is formed when a plutonium-239 target is bombarded with accelerated alpha particles:

$$^{239}_{94}\text{Pu} + ^{4}_{2}\text{He} \longrightarrow ^{242}_{96}\text{Cm} + ^{1}_{0}\text{n} \qquad [21.14]$$

New advances in the detection of the decay patterns of single atoms have led to recent additions to the periodic table. Between 1994 and 2010, elements 110 through 118 were discovered via the nuclear reactions that occur when nuclei of much lighter elements collide with high energy. For example, in 1996 a team of European scientists based in Germany synthesized element 112, copernicium, Cn, by bombarding a lead target continuously for three weeks with a beam of zinc atoms:

$$^{208}_{82}\text{Pb} + ^{70}_{30}\text{Zn} \longrightarrow ^{277}_{112}\text{Cn} + ^{1}_{0}\text{n} \qquad [21.15]$$

Amazingly, their discovery was based on the detection of only one atom of the new element, which decays after roughly 100 μs by alpha decay to form darmstadtium-273 (element 110). Within one minute, another five alpha decays take place producing fermium-253 (element 100). The finding has been verified in both Japan and Russia.

Because experiments to create new elements are very complicated and produce only a very small number of atoms of the new elements, they need to be carefully evaluated and reproduced before the new element is made an official part of the periodic table.

The International Union for Pure and Applied Chemistry (IUPAC) is the international body that authorizes names of new elements after their experimental discovery and confirmation. According to the IUPAC, new elements can be named after a mythological concept, a mineral, a place or country, a property, or a scientist. In 2016, IUPAC approved the following names and symbols for elements 113, 115, 117, and 118, as suggested by their discoverers: nihonium, Nh, for element 113; moscovium, Mc, for element 115; tennessine, Ts, for element 117; and organesson, Og, for element 118.

Self-Assessment Exercises

21.6 Cobalt-60 is used in radiotherapy, in industrial radiography, and in sterilizing food and equipment. Which nuclear equation represents the likely decay route of Co-60? (Hint: You may find it useful to refer to Figure 21.1)

(a) $^{60}_{27}\text{Co} \longrightarrow ^{56}_{25}\text{Mn} + ^{4}_{2}\alpha$

(b) $^{60}_{27}\text{Co} \longrightarrow ^{60}_{28}\text{Ni} + ^{0}_{-1}\text{e}$

(c) $^{60}_{27}\text{Co} \longrightarrow ^{60}_{26}\text{Fe} + ^{0}_{+1}\text{e}$

21.7 Which nuclear equation represents the nuclear transmutation: $^{11}_{6}\text{C(n,p)}^{11}_{5}\text{B}$

(a) $^{11}_{6}\text{C} + ^{1}_{0}\text{n} \longrightarrow ^{1}_{1}\text{p} + ^{11}_{5}\text{B}$

(b) $^{11}_{6}\text{C} + ^{1}_{1}\text{p} \longrightarrow ^{1}_{0}\text{n} + ^{11}_{5}\text{B}$

(c) $^{11}_{5}\text{B} + ^{1}_{1}\text{p} \longrightarrow ^{1}_{0}\text{n} + ^{11}_{6}\text{C}$

Exercises

21.8 Each of the following nuclei undergoes either beta decay or positron emission. Predict the type of emission for each: (**a**) $^{90}_{38}\text{Sr}$, (**b**) $^{85}_{38}\text{Sr}$, (**c**) potassium-40, (**d**) sulfur-30.

21.9 One nuclide in each of these pairs is radioactive. Predict which is radioactive and which is stable: (**a**) $^{40}_{20}\text{Ca}$ and $^{45}_{20}\text{Ca}$, (**b**) ^{12}C and ^{14}C, (**c**) lead-206 and thorium-230. Explain your choice in each case.

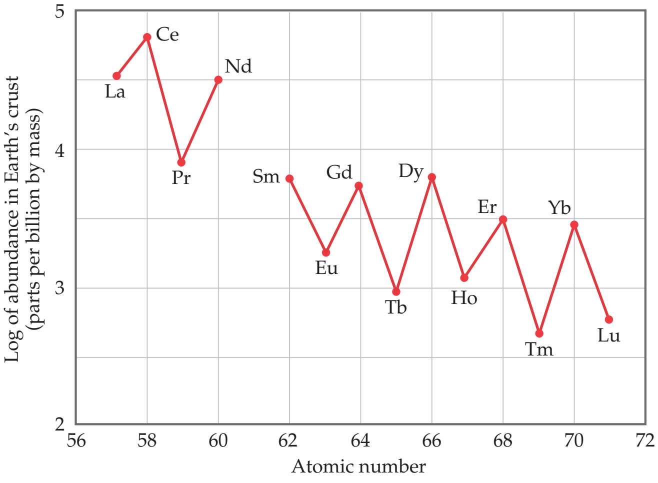
21.10 Despite the similarities in the chemical reactivity of elements in the lanthanide series, their abundances in Earth's crust vary by two orders of magnitude. This graph shows the relative abundance as a function of atomic number. Which of the following statements best explains the sawtooth variation across the series?

(**a**) The elements with an odd atomic number lie above the belt of stability.

(**b**) The elements with an odd atomic number lie below the belt of stability.

(**c**) The elements with an even atomic number have a magic number of protons.

(**d**) Pairs of protons have a special stability.



21.11 Which of the following nuclides would you expect to be radioactive: $^{58}_{26}\text{Fe}$, $^{60}_{27}\text{Co}$, $^{92}_{41}\text{Nb}$, mercury-202, radium-226? Justify your choices.

21.12 A radioactive decay series that begins with $^{232}_{90}Th$ ends with formation of the stable nuclide $^{208}_{82}Pb$. How many alpha-particle emissions and how many beta-particle emissions are involved in the sequence of radioactive decays?

21.13 In 1930 the American physicist Ernest Lawrence designed the first cyclotron in Berkeley, California. In 1937 Lawrence bombarded a molybdenum target with deuterium ions, producing for the first time an element not found in nature. What was this element? Starting with molybdenum-96 as your reactant, write a nuclear equation to represent this process.

21.14 Complete and balance the following nuclear equations by supplying the missing particle:

(**a**) $^{106}_{47}Ag + ^{0}_{-1}e \longrightarrow ?$

(**b**) $^{263}_{106}Sg \longrightarrow ^{4}_{2}He + ?$

(**c**) $^{216}_{84}Po \longrightarrow ^{212}_{82}Pb + ?$

(**d**) $^{10}_{5}B + ? \longrightarrow ^{7}_{3}Li + ^{4}_{2}He$

(**e**) $^{220}_{86}Rn \longrightarrow ^{4}_{2}He + ?$

21.15 Write balanced equations for each of the following nuclear reactions: (**a**) $^{238}_{92}U(n,\gamma)^{239}_{92}U$, (**b**) $^{16}_{8}O(p,\alpha)^{13}_{7}N$, (**c**) $^{18}_{8}O(n,\beta^-)^{19}_{9}F$.

Answers to Self-Assessment Exercises

21.6 (b) **21.7** (a)

21.3 | Rates of Radioactive Decay

The discovery of radiocarbon dating around 1950 had a profound effect on archaeology. For the first time, it was possible to obtain the date of an organic object (for example, cloth, bone, parchment, or charcoal from a fire) without reference to stratigraphy. It also made possible the comparison of sites that were distant from one another showing that the same innovation could arise independently in two or more places. It had previously been thought that all innovations spread slowly across the pre-historic landscape as people migrated, traded, and invaded neighboring regions.

By the end of this section, you should be able to

- Use the half-life of selected radionuclides to calculate the ages of objects

Some radioisotopes, such as uranium-238, are found in nature even though they are not stable. Other radioisotopes do not exist in nature but can be synthesized in nuclear reactions. To understand this distinction, we must realize that different nuclei undergo radioactive decay at different rates. Many radioisotopes decay essentially completely in fractions of a second, so we do not find them in nature. Uranium-238, on the other hand, decays very slowly. Therefore, despite its instability, we can still observe what remains from its formation in the early history of the universe.

Radioactive decay is a first-order kinetic process. Recall that a first-order process has a characteristic **half-life**, which is the time required for half of any given quantity of a substance to react. Nuclear decay rates are commonly expressed in terms of half-lives. Each radioisotope has its own characteristic half-life. For example, strontium-90 has a half-life of 28.8 yr:

$$^{90}_{38}Sr \longrightarrow ^{90}_{39}Y + ^{0}_{-1}e \qquad t_{1/2} = 28.8\text{ yr} \qquad [21.16]$$

TABLE 21.5 The Half-Lives and Type of Decay for Several Radioisotopes

	Isotope	Half-Life (yr)	Type of Decay
Natural radioisotopes	$^{238}_{92}U$	4.5×10^{9}	Alpha
	$^{235}_{92}U$	7.0×10^{8}	Alpha
	$^{232}_{90}Th$	1.4×10^{10}	Alpha
	$^{40}_{19}K$	1.3×10^{9}	Beta
	$^{14}_{6}C$	5700	Beta
Synthetic radioisotopes	$^{239}_{94}Pu$	24,000	Alpha
	$^{137}_{55}Cs$	30.2	Beta
	$^{90}_{38}Sr$	28.8	Beta
	$^{131}_{53}I$	0.022	Beta

Thus, if we start with 10.0 g of strontium-90, only 5.0 g of that isotope remains after 28.8 yr, 2.5 g remains after another 28.8 yr, and so on (**Figure 21.6**).

Half-lives as short as millionths of a second and as long as billions of years are known. The half-lives of some radioisotopes are listed in **Table 21.5**. One important feature of half-lives for nuclear decay is that they are unaffected by external conditions such as temperature, pressure, or state of chemical combination. Unlike toxic chemicals, therefore, radioactive atoms cannot be rendered harmless by chemical reaction or by any other practical treatment.

Radiometric Dating

Because the half-life of any particular nuclide is constant, the half-life can serve as a "nuclear clock" to determine the age of objects. The method of dating objects based on their isotopes and isotope abundances is called *radiometric dating*.

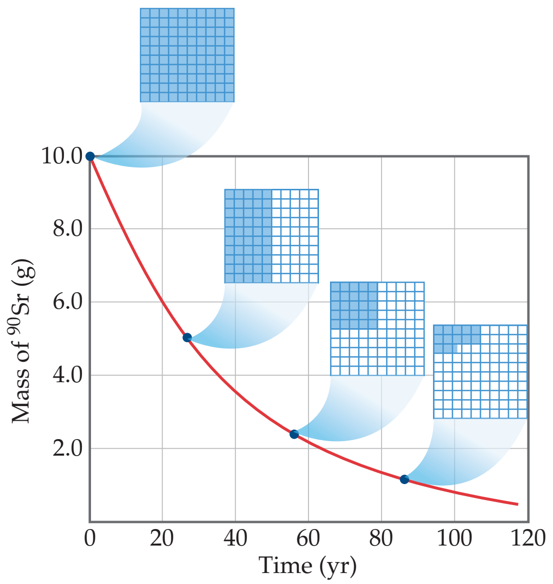
When carbon-14 is used in radiometric dating, the technique is known as *radiocarbon dating*. The procedure is based on the formation of carbon-14 as neutrons created by cosmic rays in the upper atmosphere convert nitrogen-14 into carbon-14 (**Figure 21.7**). The ^{14}C reacts with oxygen to form $^{14}CO_2$ in the atmosphere, and this "labeled" CO_2 is taken up by plants and introduced into the food chain through photosynthesis. This process provides a small but reasonably constant source of carbon-14, which is radioactive and undergoes beta decay with a half-life of 5700 yr (to two significant figures):

$$^{14}_{6}C \longrightarrow {}^{14}_{7}N + {}^{0}_{-1}e \qquad [21.17]$$

Because a living plant or animal has a constant intake of carbon compounds, it is able to maintain a ratio of carbon-14 to carbon-12 that is nearly identical with that of the atmosphere.

Go Figure

If we start with a 50.0-g sample, how much remains after three half-lives have passed?



▲ **Figure 21.6 Decay of a 10.0 g sample of strontium-90 ($t_{1/2}$ = 28.8 yr).** The 10 × 10 grids show how much of the radioactive isotope remains after various amounts of time.

Sample Exercise 21.5
Calculation Involving Half-Lives

The half-life of cobalt-60 is 5.27 yr. How much of a 1.000 mg sample of cobalt-60 is left after 15.81 yr?

SOLUTION

Analyze We are given the half-life for cobalt-60 and asked to calculate the amount of cobalt-60 remaining from an initial 1.000 mg sample after 15.81 yr.

Plan We will use the fact that the amount of a radioactive substance decreases by 50% for every half-life that passes.

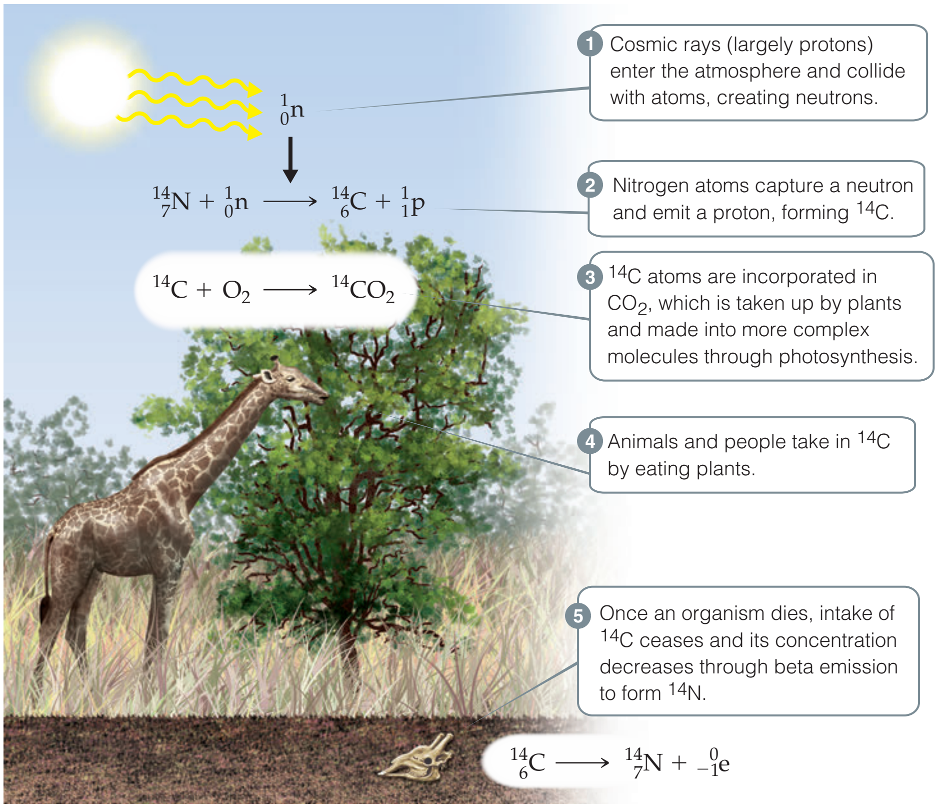
Solve Because $5.27 \times 3 = 15.81$, 15.81 yr is three half-lives for cobalt-60. At the end of one half-life, 0.500 mg of cobalt-60 remains, 0.250 mg at the end of two half-lives, and 0.125 mg at the end of three half-lives.

Practice Exercise

A radioisotope of technetium is useful in medical imaging techniques. A sample initially contains 80.0 mg of this isotope. After 24.0 h, only 5.0 mg of the technetium isotope remains. What is the half-life of the isotope?
(a) 3.0 h **(b)** 6.0 h **(c)** 12.0 h **(d)** 16.0 h **(e)** 24.0 h

Once the organism dies, however, it no longer ingests carbon compounds to replenish the carbon-14 lost through radioactive decay. The ratio of carbon-14 to carbon-12 therefore decreases. By measuring this ratio and comparing it with that of the atmosphere, we can estimate the age of an object. For example, if the ratio diminishes to half that of the atmosphere, we can conclude that the object is one half-life, or 5700 yr, old.

Go Figure How does $^{14}CO_2$ become incorporated into the mammalian food chain?



▲ **Figure 21.7 Creation and distribution of carbon-14.** The ratio of carbon-14 to carbon-12 in a dead animal or plant is related to the time since death occurred.

This method cannot be used to date objects older than about 50,000 yr because after this length of time the radioactivity is too low to be measured accurately.

In radiocarbon dating, a reasonable assumption is that the ratio of carbon-14 to carbon-12 in the atmosphere has been relatively constant for the past 50,000 yr. However, because variations in solar activity control the amount of carbon-14 produced in the atmosphere, that ratio can fluctuate. We can correct for this effect by using other kinds of data. Recently, scientists have compared carbon-14 data with data from tree rings, corals, lake sediments, ice cores, and other natural sources to correct variations in the carbon-14 "clock" back to 26,000 yr.

Other isotopes can be similarly used to date other types of objects. For example, it takes 4.5×10^9 yr for half of a sample of uranium-238 to decay to lead-206. The age of rocks containing uranium can therefore be determined by measuring the ratio of lead-206 to uranium-238. If the lead-206 had somehow become incorporated into the rock by normal chemical processes instead of by radioactive decay, the rock would also contain large amounts of the more abundant isotope lead-208. In the absence of large amounts of this "geonormal" isotope of lead, it is assumed that all of the lead-206 was at one time uranium-238.

The oldest rocks found on Earth are approximately 3×10^9 yr old. This age indicates that Earth's crust has been solid for at least this length of time. Scientists estimate that it required 1×10^9 to 1.5×10^9 yr for Earth to cool and its surface to become solid, making the age of Earth 4.0 to 4.5×10^9 yr.

Calculations Based on Half-Life

The rate at which a sample decays is called its **activity**, and it is often expressed as number of disintegrations per unit time. The **becquerel** (Bq) is the SI unit for expressing activity. A becquerel is defined as one nuclear disintegration per second. An older, but still widely used, unit of activity is the **curie** (Ci), defined as 3.7×10^{10} disintegrations per second, which is the rate of decay of 1 g of radium. Thus, a 4.0 mCi sample of cobalt-60 undergoes

$$4.0 \times 10^{-3}\ \cancel{\text{Ci}} \times \frac{3.7 \times 10^{10}\ \text{disintegrations/s}}{1\ \cancel{\text{Ci}}} = 1.5 \times 10^{8}\ \text{disintegrations/s}$$

and so has an activity of 1.5×10^8 Bq.

As a radioactive sample decays, the amount of radiation emanating from the sample decays as well. For example, the half-life of cobalt-60 is 5.27 yr. The 4.0 mCi sample of cobalt-60 would, after 5.27 yr, have a radiation activity of 2.0 mCi, or 7.5×10^7 Bq.

Because radioactive decay is a first-order kinetic process, its rate is proportional to the number of radioactive nuclei N in a sample:

$$\text{Rate} = kN \qquad [21.18]$$

The first-order rate constant, k, is called the *decay constant* and is related to the half-life:

$$k = \frac{0.693}{t_{1/2}} \qquad [21.19]$$

Thus, if we know the value of either the half-life or the decay constant, we can calculate the value of the other.

As we saw in Section 14.4, a first-order rate law can be expressed in the following form:

$$\ln \frac{N_t}{N_0} = -kt \qquad [21.20]$$

In this equation t is the time interval of decay, k is the decay constant, N_0 is the initial number of nuclei (at time zero), and N_t is the number remaining after the time interval. Both the mass of a particular radioisotope and its activity are proportional to the number of radioactive nuclei. Thus, either the ratio of the mass at any time t to the mass at time $t = 0$ or the ratio of the activities at time t and $t = 0$ can be substituted for N_t/N_0 in Equation 21.20.

Sample Exercise 21.6
Calculating the Age of Objects Using Radioactive Decay

A rock contains 0.257 mg of lead-206 for every milligram of uranium-238. The half-life for the decay of uranium-238 to lead-206 is 4.5×10^9 yr. How old is the rock?

SOLUTION

Analyze We are asked to calculate the age of a rock containing uranium-238 and lead-206, given the half-life of the uranium-238 and the relative amounts of the uranium-238 and lead-206.

Plan Lead-206 is the product of the radioactive decay of uranium-238. We will assume that the only source of lead-206 in the rock is from the decay of uranium-238, which has a known half-life. To apply first-order kinetics expressions (Equations 21.19 and 21.20) to calculate the time elapsed since the rock was formed, we first need to calculate how much initial uranium-238 there was for every 1 mg that remains today.

Solve Let's assume that the rock currently contains 1.000 mg of uranium-238 and therefore 0.257 mg of lead-206. The amount of uranium-238 in the rock when it was first formed therefore equals 1.000 mg plus the quantity that has decayed to lead-206. Because the mass of lead atoms is not the same as the mass of uranium atoms, we cannot just add 1.000 mg and 0.257 mg. We have to multiply the present mass of lead-206 (0.257 mg) by the ratio of the mass number of uranium to that of lead, into which it has decayed. Therefore, the original mass of ${}^{238}_{92}\text{U}$ was

$$\text{Original } {}^{238}_{92}\text{U} = 1.000 \text{ mg} + \frac{238}{206}(0.257 \text{ mg})$$

$$= 1.297 \text{ mg}$$

Using Equation 21.19, we can calculate the decay constant for the process from its half-life:

$$k = \frac{0.693}{4.5 \times 10^9 \text{ yr}} = 1.5 \times 10^{-10} \text{ yr}^{-1}$$

Rearranging Equation 21.20 to solve for time, t, and substituting known quantities gives

$$t = -\frac{1}{k}\ln\frac{N_t}{N_0} = -\frac{1}{1.5 \times 10^{-10} \text{ yr}^{-1}}\ln\frac{1.000}{1.297} = 1.7 \times 10^9 \text{ yr}$$

Practice Exercise

Cesium-137, which has a half-life of 30.2 yr, is a component of the radioactive waste from nuclear power plants. If the activity due to cesium-137 in a sample of radioactive waste has decreased to 35.2% of its initial value, how old is the sample?
(**a**) 1.04 yr (**b**) 15.4 yr (**c**) 31.5 yr (**d**) 45.5 yr (**e**) 156 yr

Sample Exercise 21.7
Calculations Involving Radioactive Decay and Time

If we start with 1.000 g of strontium-90, 0.953 g will remain after 2.00 yr. (**a**) What is the half-life of strontium-90? (**b**) How much strontium-90 will remain after 5.00 yr?

SOLUTION

Analyze (**a**) We are asked to calculate a half-life, $t_{1/2}$, based on data that tell us how much of a radioactive nucleus has decayed in a time interval $t = 2.00$ yr and the information $N_0 = 1.000$ g, $N_t = 0.953$ g. (**b**) We are asked to calculate the amount of a radionuclide remaining after a given period of time.

Plan (**a**) We first calculate the rate constant for the decay, k, and then we use that to compute $t_{1/2}$. (**b**) We need to calculate N_t, the amount of strontium present at time t, using the initial quantity, N_0, and the rate constant for decay, k, calculated in part (a).

Solve

(**a**) Equation 21.20 is solved for the decay constant, k, and then Equation 21.19 is used to calculate half-life, $t_{1/2}$:

$$k = -\frac{1}{t}\ln\frac{N_t}{N_0} = -\frac{1}{2.00 \text{ yr}}\ln\frac{0.953 \text{ g}}{1.000 \text{ g}}$$

$$= -\frac{1}{2.00 \text{ yr}}(-0.0481) = 0.0241 \text{ yr}^{-1}$$

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0241 \text{ yr}^{-1}} = 28.8 \text{ yr}$$

Continued

(**b**) Again using Equation 21.20, with $k = 0.0241\ \text{yr}^{-1}$, we have:

$$\ln \frac{N_t}{N_0} = -kt = -(0.0241\ \cancel{\text{yr}^{-1}})(5.00\ \cancel{\text{yr}}) = -0.120$$

N_t/N_0 is calculated from $\ln(N_t/N_0) = -0.120$ using the e^x or INV LN function of a calculator:

$$\frac{N_t}{N_0} = e^{-0.120} = 0.887$$

Because $N_0 = 1.000$ g, we have:

$$N_t = (0.887)N_0 = (0.887)(1.000\ \text{g}) = 0.887\ \text{g}$$

Practice Exercise

As mentioned in the previous Practice Exercise, cesium-137, a component of radioactive waste, has a half-life of 30.2 yr. If a sample of waste has an initial activity of 15.0 Ci due to cesium-137, how long will it take for the activity due to cesium-137 to drop to 0.250 Ci?
(**a**) 0.728 yr (**b**) 60.4 yr (**c**) 78.2 yr (**d**) 124 yr (**e**) 178 yr

Self-Assessment Exercise

21.16 A one-gram sample of carbon from peat moss has an activity of 0.350 mCi. A reference or modern standard sample yields 0.446 mCi. What is the radiocarbon age of the peat moss given a half-life of ^{14}C = 5700 years?

(a) Age = 880 yr old
(b) Age = 1400 yr old
(c) Age = 2000 yr old

Exercises

21.17 It has been suggested that strontium-90 (generated by nuclear testing) deposited in the hot desert will undergo radioactive decay more rapidly because it will be exposed to much higher average temperatures. (**a**) Is this a reasonable suggestion? (**b**) Does the process of radioactive decay have an activation energy, like the Arrhenius behavior of many chemical reactions (Section 14.5)?

21.18 It takes 180 minutes for a 200 mg sample of an unknown radioactive substance to decay to 112 mg. What is the half-life of this substance?

21.19 How much time is required for a 5.00 g sample of ^{233}Pa to decay to 0.625 g if the half-life for the beta decay of ^{233}Pa is 27.4 days?

21.20 Iodine-131, which undergoes beta decay, has a half-life of 8.02 days. (**a**) How many beta particles are emitted in 1 min by a 5.00 mg sample of ^{131}I? (**b**) What is the activity of the sample in Bq?

21.21 A wooden artifact from an Indian temple has a ^{14}C activity of 42 counts per minute as compared with an activity of 58.2 counts per minute for a standard zero age. From the half-life of ^{14}C decay, 5715 years, calculate the age of the artifact.

21.22 Iodine-131 is used as a nuclear medicine to treat hyperthyroidism. The half-life of ^{131}I is 8.04 days. How long will it take for a 500 mg sample of ^{131}I to decay into 1% of its original mass?

Answers to Self-Assessment Exercise

21.16 (c)

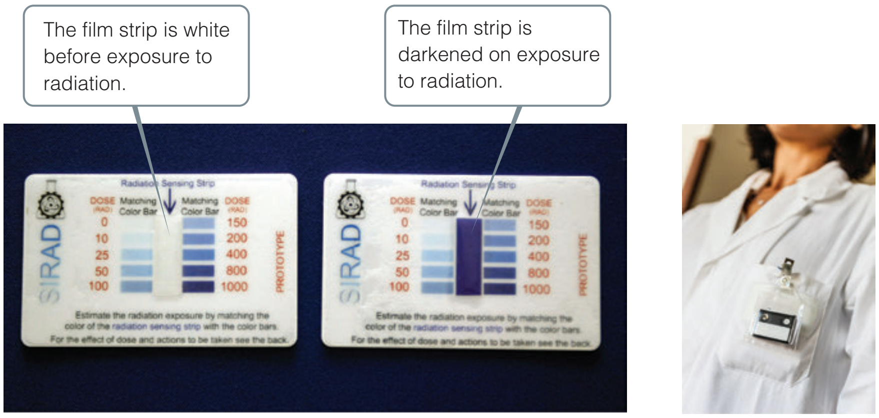
21.4 Detection of Radioactivity

Domestic smoke alarms of the ‘ionization’ type contain a small amount of radioactive material (usually Am-241). This ionizes the air in a chamber in the alarm allowing a small current to flow. If there is smoke present, this disrupts the flow of ions, which triggers an alarm to sound. This is just one common device in which radionuclides are present. In order to ensure such devices are safe to have in the home the amount of radiation they emit is carefully regulated. In this section, we look at the ways to quantify radiation. By the end of this section, you should

- Appreciate the two main ways to detect radioactivity

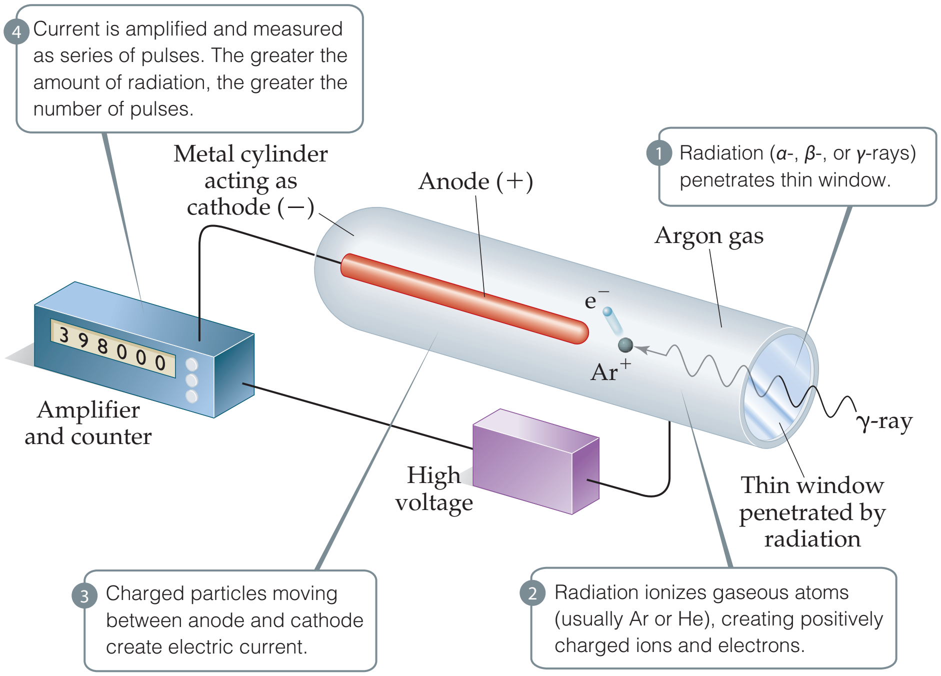
A variety of methods have been devised to detect emissions from radioactive substances. Henri Becquerel discovered radioactivity because radiation caused fogging of photographic plates, and since that time photographic plates and film have been used to detect radioactivity. The radiation affects photographic film in much the same way as X-rays do. The greater the extent of exposure to radiation, the darker the area of the developed negative. People who work with radioactive substances carry film badges to record the extent of their exposure to radiation (**Figure 21.8**).

Go Figure Which type of radiation—alpha, beta, or gamma—is likely to fog a film that is sensitive to X rays?



▲ **Figure 21.8 Badge dosimeters monitor the extent to which the individual has been exposed to high-energy radiation.** The radiation dose is determined from the extent of darkening of the film in the dosimeter.

Go Figure Which property of the atoms of gas inside a Geiger counter is most relevant to the operation of the device?



▲ **Figure 21.9 Schematic drawing of a Geiger counter.**

Radioactivity can also be detected and measured by a Geiger counter. The operation of this device is based on the fact that radiation is able to ionize matter. The ions and electrons produced by the ionizing radiation permit conduction of an electrical current. The basic design of a Geiger counter is shown in **Figure 21.9**. A current pulse between the anode and the metal cylinder occurs whenever entering radiation produces ions. Each pulse is counted in order to estimate the amount of radiation.

Some substances, called *phosphors*, emit light when radiation strikes or passes through them. The radioactivity excites the atoms, ions, or molecules of the phosphor to a higher energy state, and they release this energy as light as they return to their ground states. For example, ZnS responds this way to alpha radiation. An instrument called a *scintillation counter* detects and counts the flashes of light produced when radiation strikes the phosphor. The flashes of light are magnified electronically and counted to measure the amount of radiation.

Radiotracers

Because radioisotopes can be detected readily, they can be used to follow an element through its chemical reactions. The incorporation of carbon atoms from CO_2 into glucose during photosynthesis, for example, has been studied using CO_2 enriched in carbon-14:

$$6\,{}^{14}CO_2 + 6\,H_2O \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} {}^{14}C_6H_{12}O_6 + 6\,O_2 \qquad [21.21]$$

Use of the carbon-14 label provides direct experimental evidence that carbon dioxide in the environment is chemically converted to glucose in plants. Analogous labeling experiments using oxygen-18 show that the O_2 produced during photosynthesis comes from water, not carbon dioxide. When it is possible to isolate and purify intermediates and products from reactions, detection devices such as scintillation counters can be used to “follow” the radioisotope as it moves from starting material through intermediates to final product. These types of experiments are useful for identifying elementary steps in a reaction mechanism.

The use of radioisotopes is possible because all isotopes of an element have essentially identical chemical properties. When a small quantity of a radioisotope is mixed with the naturally occurring stable isotopes of the same element, all the isotopes go through the same reactions together. The element's path is revealed by the radioactivity of the radioisotope. Because the radioisotope can be used to trace the path of the element, it is called a **radiotracer**.

CHEMISTRY AND LIFE Medical Applications of Radiotracers

Radiotracers have found wide use as diagnostic tools in medicine. Table 21.6 lists some radiotracers and their uses. These radioisotopes are incorporated into a compound that is administered to the patient, usually intravenously. The diagnostic use of these isotopes is based on the ability of the radioactive compound to localize and concentrate in the organ or tissue under investigation. Iodine-131, for example, has been used to test the activity of the thyroid gland. This gland is the only place in which iodine is incorporated significantly in the body. The patient drinks a solution of NaI containing iodine-131. Only a very small amount is used so that the patient does not receive a harmful dose of radioactivity. A Geiger counter placed close to the thyroid, in the neck region, determines the ability of the thyroid to take up the iodine. A normal thyroid will absorb about 12% of the iodine within a few hours.

TABLE 21.6 Some Radionuclides Used as Radiotracers

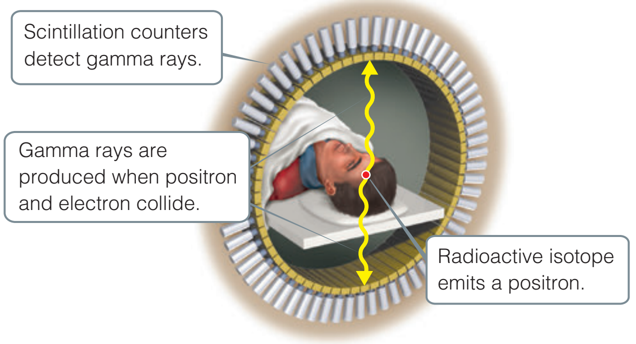
Nuclide	Half-Life	Area of the Body Studied
Iodine-131	8.04 days	Thyroid
Iron-59	44.5 days	Red blood cells
Phosphorus-32	14.3 days	Eyes, liver, tumors
Technetium-99[a]	6.0 hours	Heart, bones, liver, and lungs
Thallium-201	73 hours	Heart, arteries
Sodium-24	14.8 hours	Circulatory system

[a]The isotope of technetium is actually a special isotope of Tc-99 called Tc-99*m*, where the *m* indicates a so-called *metastable* isotope.

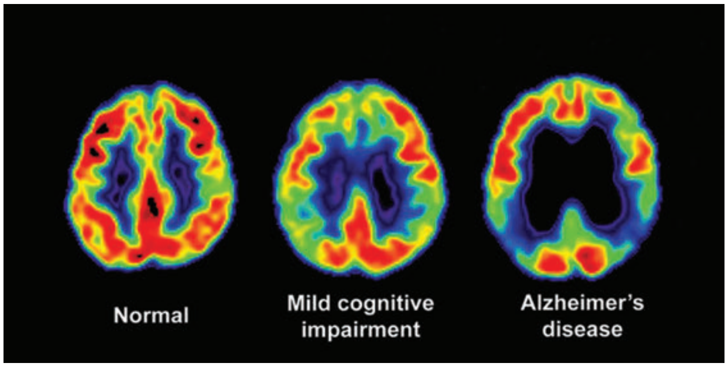
The medical applications of radiotracers are further illustrated by *positron emission tomography* (PET). PET is used for clinical diagnosis of many diseases. In this method, compounds containing radionuclides that decay by positron emission are injected into a patient. These compounds are chosen to enable researchers to monitor blood flow, oxygen and glucose metabolic rates, and other biological functions. Some of the most interesting work involves study of the brain, which depends on glucose for most of its energy. Changes in how this sugar is metabolized or used by the brain may signal a disease such as cancer, epilepsy, Parkinson's disease, or schizophrenia.

The radionuclides that are most widely used in PET are carbon-11 ($t_{1/2} = 20.4$ min), fluorine-18 ($t_{1/2} = 110$ min), oxygen-15 ($t_{1/2} = 2$ min), and nitrogen-13 ($t_{1/2} = 10$ min). Glucose, for example, can be labeled with carbon-11. Because the half-lives of positron emitters are so short, they must be generated on site using a cyclotron and the chemist must quickly incorporate the radionuclide into the sugar (or other appropriate) molecule and inject the compound immediately. The patient is placed in an instrument that measures the positron emission and constructs a computer-based image of the organ in which the emitting compound is localized. When the element decays, the emitted positron quickly collides with an electron. The positron and electron are annihilated in the collision, producing two gamma rays that move in opposite directions. The gamma rays are detected by an encircling ring of scintillation counters (Figure 21.10). Because the rays move in opposite directions but were created in the same place at the same time, it is possible to accurately locate the point in the body where the radioactive isotope decayed. The nature of this image provides clues to the presence of disease or other abnormality and helps medical researchers understand how a particular disease affects the functioning of the brain. For example, the images shown in Figure 21.11 reveal that levels of activity in brains of patients with Alzheimer's disease are different from the levels in those without the disease.

Related Exercises: 21.24, 21.64, 21.89



▲ **Figure 21.10 Schematic representation of a positron emission tomography (PET) scanner.**



▲ **Figure 21.11 Positron emission tomography (PET) scans showing glucose metabolism levels in the brain.** Red and yellow colors show higher levels of glucose metabolism.

Self-Assessment Exercise

21.23 The luminous dial of an old watch contains a radioisotope and a material called a phosphor, which glows when energized. What type of radiation detector uses a similar mechanism?

(a) A Geiger counter
(b) A scintillation counter

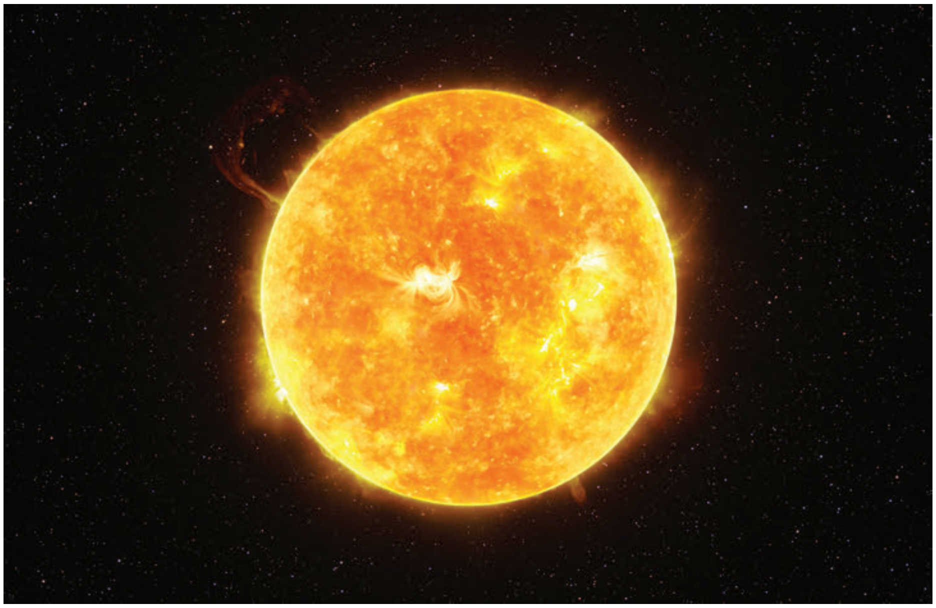
Exercise

21.24 Iodine-131 is a convenient radioisotope to monitor thyroid activity in humans. It is a beta emitter with a half-life of 8.02 days. The thyroid is the only gland in the body that uses iodine. A person undergoing a test of thyroid activity drinks a solution of NaI, in which only a small fraction of the iodide is radioactive. (**a**) Why is NaI a good choice for the source of iodine? (**b**) If a Geiger counter is placed near the person's thyroid (which is near the neck) right after the sodium iodide solution is taken, what will the data look like as a function of time? (**c**) A normal thyroid will take up about 12% of the ingested iodide in a few hours. How long will it take for the radioactive iodide taken up and held by the thyroid to decay to 0.01% of the original amount?

Answers to Self-Assessment Exercises

21.23 (b)

21.5 Energy Changes in Nuclear Reactions



The sun is a huge sphere so hot that nuclei and electrons move independently. It accounts for 99.86% of the mass of our solar system and is composed of 73.8% hydrogen, 24.8% helium, and 1.4% other elements. Most of the Sun's energy is generated in its core by the fusion of hydrogen nuclei to form helium nuclei. Such nuclear reactions can release enormous amounts of energy—far more than the amounts involved in even the most energetic chemical reactions. The Sun's surface releases this energy as electromagnetic radiation accompanied by a stream of charged particles called solar wind. Bursts of radiation and particles continuously erupt from the surface, producing solar flares.

In this section, we discuss the processes involved in nuclear reactions and our attempts to harness them. By the end of the section, you should be able to

- Calculate mass and energy changes for nuclear reactions.
- Describe the difference between fission and fusion.

Why are the energies associated with nuclear reactions so large, in many cases orders of magnitude larger than those associated with nonnuclear chemical reactions? The answer to this question begins with Einstein's celebrated equation from the theory of relativity that relates mass and energy:

$$E = mc^2 \qquad [21.22]$$

In this equation E stands for energy, m for mass, and c for the speed of light, 2.9979×10^8 m/s. This equation states that mass and energy are equivalent and can be converted into one another. If a system loses mass, it loses energy; if it gains mass, it gains energy. Because the proportionality constant between energy and mass, c^2, is such a large number, even small changes in mass are accompanied by large changes in energy.

The mass changes in chemical reactions are too small to detect. For example, the mass change associated with the combustion of 1 mol of CH_4 (an exothermic process) is -9.9×10^{-9} g. Because the mass change is so small, it is possible to treat chemical reactions as though mass is conserved.

The mass changes and the associated energy changes in nuclear reactions are much greater than those in chemical reactions. The mass change accompanying the radioactive decay of 1 mol of uranium-238, for example, is 50,000 times greater than that for the combustion of 1 mol of CH_4. Let's examine the energy change for the nuclear reaction

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

The masses of the nuclei are $^{238}_{92}$U, 238.0003 u; $^{234}_{90}$Th, 233.9942 u; and $^{4}_{2}$He, 4.0015 u. The mass change, Δm, is the total mass of the products minus the total mass of the reactants. The mass change for the decay of 1 mol of uranium-238 can then be expressed in grams:

$$233.9942\text{ g} + 4.0015\text{ g} - 238.0003\text{ g} = -0.0046\text{ g}$$

The fact that the system has lost mass indicates that the process is exothermic. All spontaneous nuclear reactions are exothermic.

The energy change per mole associated with this reaction is

$$\begin{aligned}\Delta E &= \Delta(mc^2) = c^2\Delta m \\ &= (2.9979 \times 10^8\text{ m/s})^2(-0.0046\text{ g})\left(\frac{1\text{ kg}}{1000\text{ g}}\right) \\ &= -4.1 \times 10^{11}\frac{\text{kg-m}^2}{\text{s}^2} = -4.1 \times 10^{11}\text{ J}\end{aligned}$$

Notice that Δm must be converted to kilograms, the SI unit of mass, to obtain ΔE in joules, the SI unit of energy. The negative sign for the energy change indicates that energy is released in the reaction—in this case, over 400 billion joules per mole of uranium! This energy would provide the average annual household electricity for about 10,000 homes.

Sample Exercise 21.8

Calculating Mass Change in a Nuclear Reaction

How much energy is lost or gained when 1 mol of cobalt-60 undergoes beta decay, ${}^{60}_{27}\text{Co} \longrightarrow {}^{60}_{28}\text{Ni} + {}^{0}_{-1}\text{e}$? The mass of a ${}^{60}_{27}\text{Co}$ atom is 59.933819 u, and that of a ${}^{60}_{28}\text{Ni}$ atom is 59.930788 u.

SOLUTION

Analyze We are asked to calculate the energy change in a nuclear reaction.

Plan We must first calculate the mass change in the process. We are given atomic masses, but we need the masses of the nuclei in the reaction. We calculate these by taking account of the masses of the electrons that contribute to the atomic masses.

Solve

A ${}^{60}_{27}\text{Co}$ atom has 27 electrons. The mass of an electron is 5.4858×10^{-4} u. (See the list of fundamental constants in the back inside cover.) We subtract the mass of the 27 electrons from the mass of the ${}^{60}_{27}\text{Co}$ *atom* to find the mass of the ${}^{60}_{27}\text{Co}$ *nucleus*:

$$59.933819\ \text{u} - (27)(5.4858 \times 10^{-4}\ \text{u}) = 59.919007\ \text{u (or 59.919007 g/mol)}$$

Likewise, for ${}^{60}_{28}\text{Ni}$, the mass of the nucleus is:

$$59.930788\ \text{u} - (28)(5.4858 \times 10^{-4}\ \text{u}) = 59.915428\ \text{u (or 59.915428 g/mol)}$$

The mass change in the nuclear reaction is the total mass of the products minus the mass of the reactant:

$$\begin{aligned} \Delta m &= \text{mass of electron} + \text{mass } {}^{60}_{28}\text{Ni nucleus} - \text{mass of } {}^{60}_{27}\text{Co nucleus} \\ &= 0.00054858\ \text{u} + 59.915428\ \text{u} - 59.919007\ \text{u} \\ &= -0.003030\ \text{u} \end{aligned}$$

Thus, when a mole of cobalt-60 decays,

$$\Delta m = -0.003030\ \text{g}$$

Because the mass decreases ($\Delta m < 0$), energy is released ($\Delta E < 0$). The quantity of energy released *per mole* of cobalt-60 is calculated using Equation 21.22:

$$\begin{aligned} \Delta E &= c^2 \Delta m \\ &= (2.9979 \times 10^8\ \text{m/s})^2(-0.003030\ \text{g})\left(\frac{1\ \text{kg}}{1000\ \text{g}}\right) \\ &= -2.723 \times 10^{11}\ \frac{\text{kg-m}^2}{\text{s}^2} = -2.723 \times 10^{11}\ \text{J} \end{aligned}$$

Practice Exercise

Positron emission from ${}^{11}\text{C}$, ${}^{11}_{6}\text{C} \longrightarrow {}^{11}_{5}\text{B} + {}^{0}_{+1}\text{e}$, occurs with release of 2.87×10^{11} J per mole of ${}^{11}\text{C}$. What is the mass change per mole of ${}^{11}\text{C}$ in this nuclear reaction? The masses of ${}^{11}\text{B}$ and ${}^{11}\text{C}$ are 11.009305 and 11.011434 u, respectively.

Nuclear Binding Energies

Scientists discovered in the 1930s that the masses of nuclei are always less than the masses of the individual nucleons of which they are composed. For example, the helium-4 nucleus (an alpha particle) has a mass of 4.00150 u. The mass of a proton is 1.00728 u and that of a neutron is 1.00866 u. Consequently, two protons and two neutrons have a total mass of 4.03188 u:

$$\begin{aligned} \text{Mass of two protons} &= 2(1.00728\ \text{u}) = 2.01456\ \text{u} \\ \text{Mass of two neutrons} &= 2(1.00866\ \text{u}) = \underline{2.01732\ \text{u}} \\ \text{Total mass} &= 4.03188\ \text{u} \end{aligned}$$

TABLE 21.7 Mass Defects and Binding Energies for Three Nuclei

Nucleus	Mass of Nucleus (u)	Mass of Individual Nucleons (u)	Mass Defect (u)	Binding Energy (J)	Binding Energy per Nucleon (J)
${}^{4}_{2}He$	4.00150	4.03188	0.03038	4.53×10^{-12}	1.13×10^{-12}
${}^{56}_{26}Fe$	55.92068	56.44914	0.52846	7.90×10^{-11}	1.41×10^{-12}
${}^{238}_{92}U$	238.00031	239.93451	1.93420	2.89×10^{-10}	1.21×10^{-12}

The mass of the individual nucleons is 0.03038 u greater than that of the helium-4 nucleus:

$$\begin{aligned} \text{Mass of two protons and two neutrons} &= 4.03188\text{ u} \\ \text{Mass of } {}^{4}_{2}\text{He nucleus} &= \underline{4.00150\text{ u}} \\ \text{Mass difference } \Delta m &= 0.03038\text{ u} \end{aligned}$$

The mass difference between a nucleus and its constituent nucleons is called the **mass defect**. The origin of the mass defect is readily understood if we consider that energy must be added to a nucleus to break it into separated protons and neutrons:

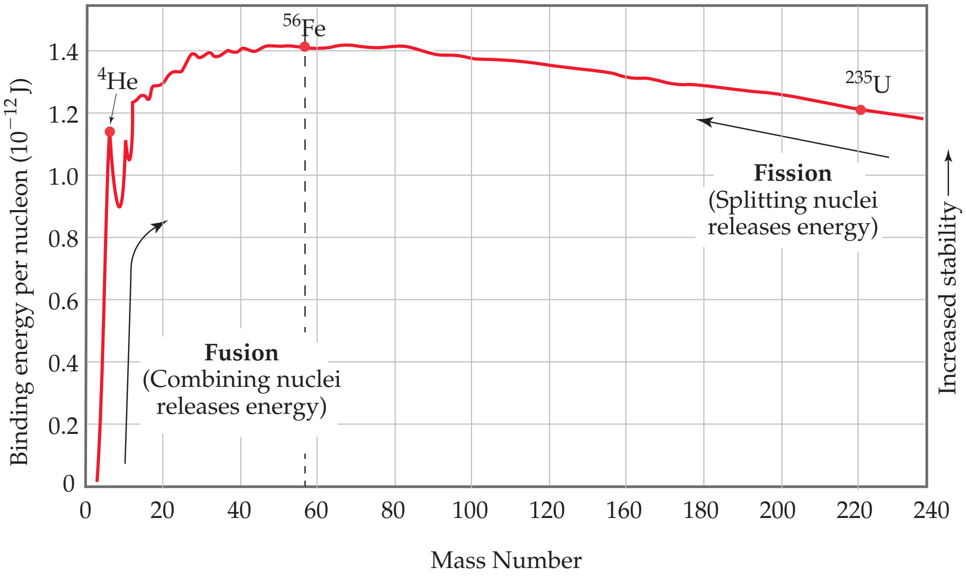
$$\text{Energy} + {}^{4}_{2}\text{He} \longrightarrow 2\,{}^{1}_{1}\text{H} + 2\,{}^{1}_{0}\text{n} \qquad [21.23]$$

By Einstein's relation, the addition of energy to a system must be accompanied by a proportional increase in mass. The mass change we just calculated for the conversion of helium-4 into separated nucleons is $\Delta m = 0.03038$ u. Therefore, the energy required for this process is

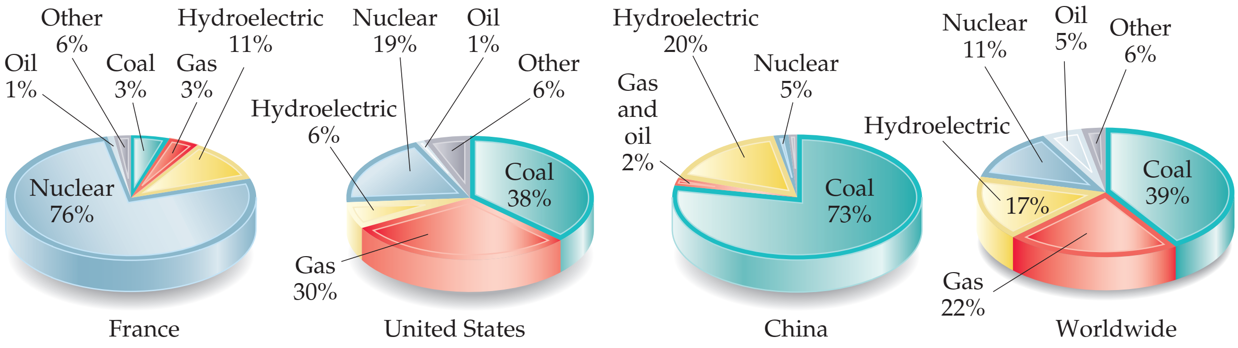
$$\begin{aligned} \Delta E &= c^2 \Delta m \\ &= (2.9979 \times 10^8\,\text{m/s})^2 (0.03038\,\cancel{\text{u}}) \left(\frac{1\,\cancel{\text{g}}}{6.022 \times 10^{23}\,\cancel{\text{u}}}\right)\left(\frac{1\,\text{kg}}{1000\,\cancel{\text{g}}}\right) \\ &= 4.534 \times 10^{-12}\,\text{J} \end{aligned}$$

The energy required to separate a nucleus into its individual nucleons is called the **nuclear binding energy**. The mass defect and nuclear binding energy for three elements are compared in **Table 21.7**.

Values of binding energies per nucleon can be used to compare the stabilities of different combinations of nucleons (such as two protons and two neutrons arranged either as ${}^{4}_{2}$He or as 2 ${}^{2}_{1}$H). **Figure 21.12** shows average binding energy per nucleon plotted



▲ **Figure 21.12 Nuclear binding energies.** The average binding energy per nucleon increases initially as the mass number increases and then decreases slowly. Because of these trends, fusion of light nuclei and fission of heavy nuclei are exothermic processes.



▲ **Figure 21.13 Sources of electricity generation, worldwide and for select countries.**
(***Sources:*** The Shift Project and the World Bank, 2014 data)

against mass number. Binding energy per nucleon at first increases in magnitude as mass number increases, reaching about 1.4×10^{-12} J for nuclei whose mass numbers are in the vicinity of iron-56. It then decreases slowly to about 1.2×10^{-12} J for very heavy nuclei. *This trend indicates that nuclei of intermediate mass numbers are more tightly bound (and therefore more stable) than those with either smaller or larger mass numbers.*

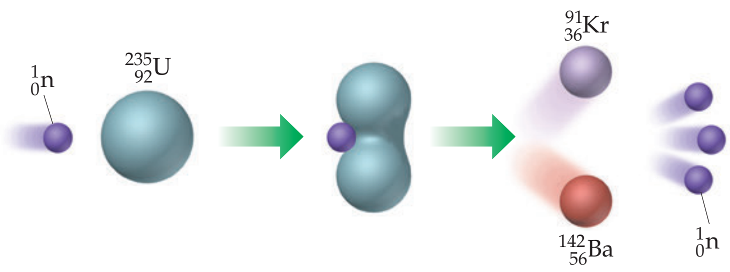
This trend has two significant consequences: First, heavy nuclei gain stability and therefore give off energy if they are fragmented into two midsized nuclei. This process, known as **fission**, is used to generate energy in nuclear power plants. Second, because of the sharp increase in the graph for small mass numbers, even greater amounts of energy are released if very light nuclei are combined, or fused together, to give more massive nuclei. This **fusion** process is the essential energy-producing process in the Sun and other stars.

Nuclear Power: Fission

Nuclear fission is the process used to generate energy in nuclear power plants. Over 11% of the electricity generated worldwide comes from nuclear power plants, though the percentage varies from one country to the next, as **Figure 21.13** shows. There are 440 commercial nuclear power plants in operation in 30 countries, and approximately another 65 are under construction.

Most nuclear reactors rely on the fission of uranium-235. This was the first nuclear fission reaction to be discovered. This nucleus, as well as those of uranium-233 and plutonium-239, undergoes fission when struck by a slow-moving neutron (**Figure 21.14**).*

Go Figure What is the relationship between the sum of the mass numbers on the two sides of this reaction?

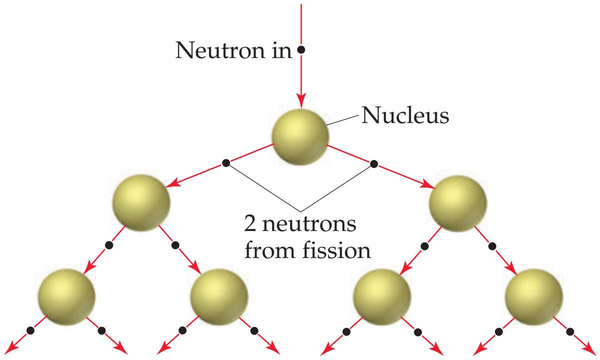


▲ **Figure 21.14 Uranium-235 fission.** This is just one of many fission patterns. In this reaction, 3.5×10^{-11} J of energy is released per ^{235}U nucleus that is split.

*Other heavy nuclei can also undergo fission. However, these three are the only ones of practical importance.

Go Figure

If this figure were extended one more “generation” down, how many neutrons would be produced?



▲ **Figure 21.15 Fission chain reaction.**

A heavy nucleus can split in many ways, giving rise to a variety of smaller nuclei. Two ways that the uranium-235 nucleus splits, for instance, are

$$^{1}_{0}n + ^{235}_{92}U \rightarrow ^{137}_{52}Te + ^{97}_{40}Zr + 2\,^{1}_{0}n \qquad [21.24]$$

$$^{1}_{0}n + ^{235}_{92}U \rightarrow ^{142}_{56}Ba + ^{91}_{36}Kr + 3\,^{1}_{0}n \qquad [21.25]$$

The nuclei produced in Equations 21.24 and 21.25—called the *fission products*—are themselves radioactive and undergo further nuclear decay. More than 200 isotopes of 35 elements have been found among the fission products of uranium-235. Most of them are radioactive.

Slow-moving neutrons are required for the fission of uranium-235 because the process involves initial absorption of the neutron by the nucleus. The resulting more massive nucleus is extremely unstable and spontaneously undergoes fission. Fast neutrons tend to bounce off the nucleus, and little fission occurs.

Note that the coefficients of the neutrons produced in Equations 21.24 and 21.25 are 2 and 3, respectively. On average, 2.4 neutrons are produced by every fission of a uranium-235 nucleus. If one fission produces two neutrons, the two neutrons can cause two additional fissions, each producing two neutrons. The four neutrons thereby released can produce four fissions, and so forth, as shown in Figure 21.15. The number of fissions and the energy released quickly escalate, and if the process is unchecked, the result is a violent explosion. Reactions that multiply in this fashion are called **chain reactions**.

For a fission chain reaction to occur, the sample of fissionable material must have a certain minimum mass. Otherwise, neutrons escape from the sample before they have the opportunity to strike other nuclei and cause additional fission. The amount of fissionable material large enough to maintain a chain reaction with a constant rate of fission is called the **critical mass**. When a critical mass of material is present, one neutron on average from each fission is subsequently effective in producing another fission and the fission continues at a constant, controllable rate. The critical mass of uranium-235 is about 50 kg for a bare sphere of the metal.*

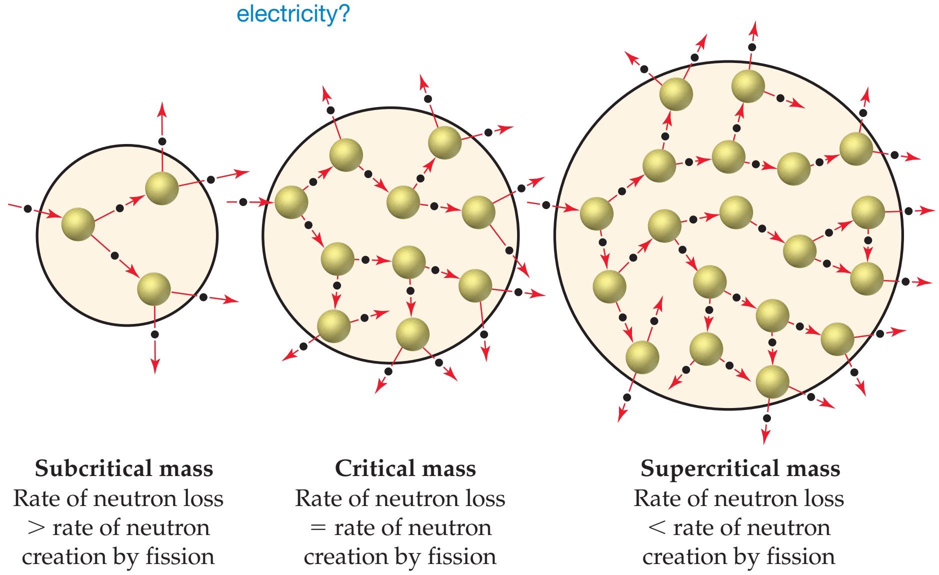
If more than a critical mass of fissionable material is present, very few neutrons escape. The chain reaction thus multiplies the number of fissions, which can lead to a nuclear explosion. A mass in excess of a critical mass is referred to as a **supercritical mass**. The effect of mass on a fission reaction is illustrated in Figure 21.16.

Figure 21.17 shows a schematic diagram of the first atomic bomb used in warfare, the bomb, code-named “Little Boy,” that was dropped on Hiroshima, Japan, on August 6, 1945. The bomb contained about 64 kg of uranium-235, which had been separated

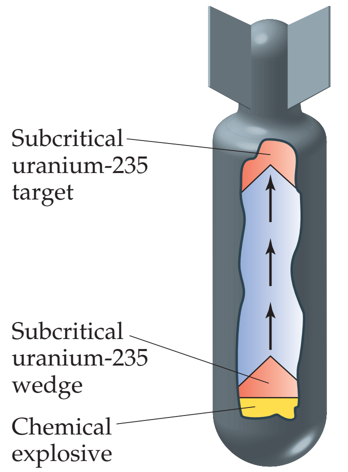
*The exact value of the critical mass depends on the shape of the radioactive substance. The critical mass can be reduced if the radioisotope is surrounded by a material that reflects some neutrons.

Go Figure

Which of these criticality scenarios—subcritical, critical, or supercritical—is desirable in a nuclear power plant that generates electricity?



▲ **Figure 21.16 Subcritical, critical, and supercritical nuclear fission.**



▲ **Figure 21.17 Schematic drawing of an atomic bomb.** A conventional explosive is used to bring two subcritical masses together to form a supercritical mass.

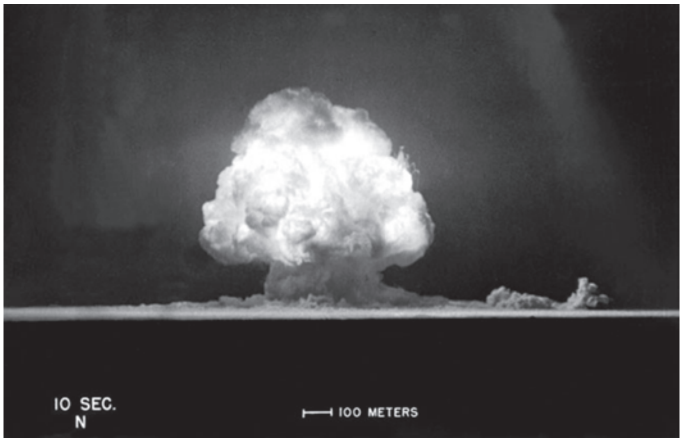
A CLOSER LOOK The Dawning of the Nuclear Age

Uranium-235 fission was first achieved during the late 1930s by Enrico Fermi and coworkers in Rome and shortly thereafter by Otto Hahn and coworkers in Berlin. Both groups were trying to produce transuranium elements. In 1938, Hahn identified barium among his reaction products. He was puzzled by this observation and questioned the identification because the presence of barium was so unexpected. He sent a letter describing his experiments to Lise Meitner, a former coworker who had been forced to leave Germany because of the anti-Semitism of the Third Reich and had settled in Sweden. She surmised that Hahn's experiment indicated a nuclear process was occurring in which the uranium-235 split. She called this process *nuclear fission*.

Meitner passed word of this discovery to her nephew, Otto Frisch, a physicist working at Niels Bohr's institute in Copenhagen. Frisch repeated the experiment, verifying Hahn's observations, and found that tremendous energies were involved. In January 1939, Meitner and Frisch published a short article describing the reaction. In March 1939, Leo Szilard and Walter Zinn at Columbia University discovered that more neutrons are produced than are used in each fission. As we have seen, this result allows a chain reaction to occur.

News of these discoveries and an awareness of their potential use in explosive devices spread rapidly within the scientific community. Several scientists finally persuaded Albert Einstein, the most famous physicist of the time, to write a letter to President Franklin D. Roosevelt explaining the implications of these discoveries. Einstein's letter, written in August 1939, outlined the possible military applications of nuclear fission and emphasized the danger that weapons based on fission would pose if they were developed by the Nazis. Roosevelt judged it imperative that the United States investigate the possibility of such weapons. Late in 1941, the decision was made to build a bomb based on the fission reaction. An enormous research project, known as the Manhattan Project, began.

On December 2, 1942, the first artificial self-sustaining nuclear fission chain reaction was achieved in an abandoned squash court at the University of Chicago. This accomplishment led to the development of the first atomic bomb, at Los Alamos National Laboratory in New Mexico in July 1945 (**Figure 21.18**). In August 1945 the United States dropped atomic bombs on two Japanese cities, Hiroshima and Nagasaki. The nuclear age had arrived, albeit in a sadly destructive fashion. Humanity has struggled with the conflict between the positive potential of nuclear energy and its terrifying potential as a weapon ever since.



▲ **Figure 21.18 The Trinity test for the atom bomb developed during World War II.** The first human-made nuclear explosion took place on July 16, 1945, on the Alamogordo test range in New Mexico.

from the nonfissionable uranium-238 primarily by gaseous diffusion of uranium hexafluoride, UF_6. To trigger the fission reaction, two subcritical masses of uranium-235 were slammed together using chemical explosives. The combined masses of the uranium formed a supercritical mass, which led to a rapid, uncontrolled chain reaction and, ultimately, a nuclear explosion. The energy released by the bomb dropped on Hiroshima was equivalent to that of 16,000 tons of TNT (it therefore is called a *16-kiloton* bomb). Unfortunately, the basic design of a fission-based atomic bomb is quite simple, and the fissionable materials are potentially available to any nation with a nuclear reactor. The combination of design simplicity and materials availability has generated international concerns about the proliferation of atomic weapons.

Nuclear Reactors

Nuclear power plants use nuclear fission to generate energy. The core of a typical nuclear reactor consists of four principal components: fuel elements, control rods, a moderator, and a primary coolant (**Figure 21.19**). The fuel is a fissionable substance, such as uranium-235. The natural isotopic abundance of uranium-235 is only 0.7%, too low to sustain a chain reaction in most reactors. Therefore, the ^{235}U content of the fuel must be enriched to 3–5% for use in a reactor. The *fuel elements* contain enriched uranium in the form of UO_2 pellets encased in zirconium or stainless steel tubes.

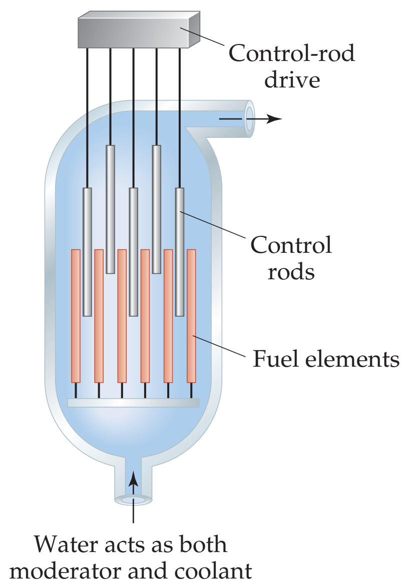
The *control rods* are composed of materials that absorb neutrons, such as boron-10 or an alloy of silver, indium, and cadmium. These rods regulate the flux of neutrons to keep the reaction chain self-sustaining and also prevent the reactor core from overheating.*

The probability that a neutron will trigger fission of a ^{235}U nucleus depends on the speed of the neutron. The neutrons produced by fission have high speeds (typically in excess of 10,000 km/s). The function of the *moderator* is to slow down the neutrons (to speeds of a few kilometers per second) so that they can be captured more readily by the fissionable nuclei. The moderator is typically either water or graphite.

The *primary coolant* is a substance that transports the heat generated by the nuclear chain reaction away from the reactor core. In a *pressurized water reactor*, which is the most common commercial reactor design, water acts as both the moderator and the primary coolant.

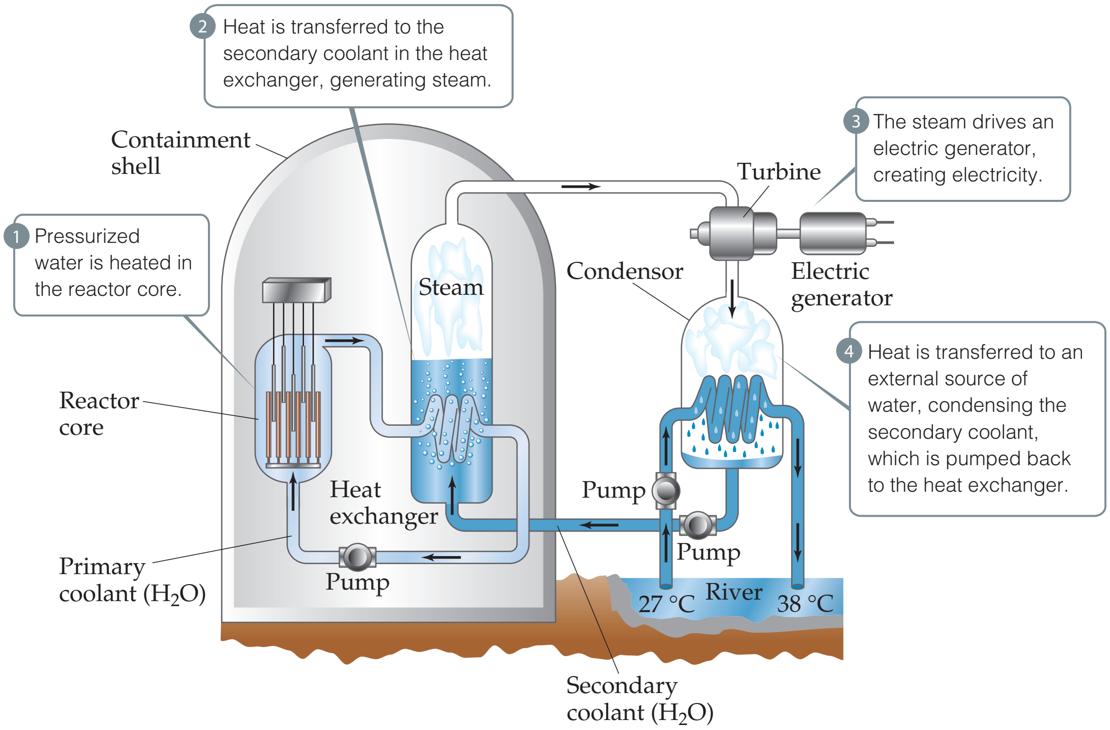
The design of a nuclear power plant is basically the same as that of a power plant that burns fossil fuel (except that the burner is replaced by a reactor core). The nuclear power plant design shown in **Figure 21.20**, a pressurized water reactor, is currently the most popular. The primary coolant passes through the core in a closed system, which lessens the chance that radioactive products could escape the core. As an added safety precaution, the reactor is surrounded by a reinforced concrete *containment shell* to shield personnel and nearby residents from radiation and to protect the reactor from external forces. After passing through the reactor core, the very hot primary coolant passes through a heat exchanger where much of its heat is transferred to a *secondary coolant*, converting the latter to high-pressure steam that is used to drive a turbine. The secondary coolant is then condensed by transferring heat to an external source of water, such as a river or lake. The cooling systems of nuclear power plants are necessary for proper and safe operation. The nuclear disaster in Fukushima, Japan, in March 2011 occurred when a tsunami damaged the reactor cooling systems, resulting in a large-scale release of radioactive materials.

Although about two-thirds of all commercial reactors are pressurized water reactors, there are several variations on this basic design, each with advantages and disadvantages. A *boiling water reactor* generates steam by boiling the primary coolant; thus, no secondary coolant is needed. The reactors at Fukushima, Japan, were boiling water reactors. Pressurized water reactors and boiling water reactors are collectively referred to as *light water reactors* because they use H_2O as moderator and primary coolant. A *heavy water reactor* uses D_2O (D = deuterium, ^2H) as moderator and primary coolant, and a *gas-cooled reactor* uses a gas, typically CO_2, as primary coolant and graphite as the moderator.



▲ **Figure 21.19 Schematic diagram of a pressurized water reactor core.**

*The reactor core cannot reach supercritical levels and explode with the violence of an atomic bomb because the concentration of uranium-235 is too low. However, if the core overheats, sufficient damage can lead to release of radioactive materials into the environment.



▲ **Figure 21.20 Basic design of a pressurized water reactor nuclear power plant.**

Use of either D_2O or graphite as the moderator has the advantage that both substances absorb fewer neutrons than H_2O. Consequently, the uranium fuel does not need to be as enriched.

Nuclear Waste

The fission products that accumulate as a reactor operates decrease the efficiency of the reactor by capturing neutrons. For this reason, commercial reactors must be stopped periodically to either replace or reprocess the nuclear fuel. When the fuel elements are removed from the reactor, they are initially very radioactive. It was originally intended that they be stored for several months in pools at the reactor site to allow decay of short-lived radioactive nuclei. They were then to be transported in shielded containers to reprocessing plants where the unspent fuel would be separated from the fission products. Reprocessing plants have been plagued with operational difficulties, however, and there is intense opposition in some countries to the transport of nuclear wastes on the nation's roads and rails. Spent fuel is reprocessed, however, in France, Russia, the United Kingdom, India, and Japan.

Storage of spent nuclear fuel poses a major problem because the fission products are extremely radioactive. It is estimated that 10 half-lives are required for their radioactivity to reach levels acceptable for biological exposure. Based on the 28.8-yr half-life of strontium-90, one of the longer-lived and most dangerous of the products, the wastes must be stored for nearly 300 yr. Plutonium-239 is one of the by-products present in spent fuel elements. It is formed by absorption of a neutron by uranium-238, followed by two successive beta emissions. (Remember that most of the uranium in the fuel elements is uranium-238.) If the elements are reprocessed, the plutonium-239 is largely recovered because it can be used as a nuclear fuel. However, if the plutonium is not removed, spent elements must be stored for a very long time because plutonium-239 has a half-life of 24,000 yr.

A *fast breeder reactor* offers one approach to getting more power out of existing uranium sources and potentially reducing radioactive waste. This type of reactor is so named because it creates ("breeds") more fissionable material than it consumes. The reactor operates without a moderator, which means the neutrons used are not slowed down.

In order to capture the fast neutrons, the fuel must be highly enriched with both uranium-235 and plutonium-239. Water cannot be used as a primary coolant because it would moderate the neutrons, and so a liquid metal, usually sodium, is used. The core is surrounded by a blanket of uranium-238 that captures neutrons that escape the core, producing plutonium-239 in the process. The plutonium can later be separated by reprocessing and used as fuel in a future cycle.

Because fast neutrons are more effective at decaying many radioactive nuclides, the material separated from the uranium and plutonium during reprocessing is less radioactive than waste from other reactors. However, generation of relatively high levels of plutonium coupled with the need for reprocessing is problematic in terms of nuclear nonproliferation. Thus, political factors coupled with increased safety concerns and higher operational costs make fast breeder reactors quite rare.

A considerable amount of research is being devoted to disposal of radioactive wastes. At present, the most attractive possibilities appear to be formation of glass, ceramic, or synthetic rock from the wastes as a means of immobilizing them. These solid materials would then be placed in containers of high corrosion resistance and durability and buried deep underground. The process of selecting appropriate deep repositories for high-level waste and spent fuel is now under way in several countries.

In spite of these difficulties, nuclear power is making a modest comeback as an energy source. Concerns about climate change caused by escalating atmospheric CO_2 levels have increased support for nuclear power as a major energy source in the future. Increasing demand for power in rapidly developing countries, particularly China, has sparked a rise in construction of new nuclear power plants in those parts of the world.

Nuclear Power: Fusion

Energy is produced when light nuclei fuse into heavier ones. Reactions of this type are responsible for the energy produced by the Sun. The following reactions are among the numerous fusion processes believed to occur in the Sun:

$$ {}^{1}_{1}\mathrm{H} + {}^{1}_{1}\mathrm{H} \longrightarrow {}^{2}_{1}\mathrm{H} + {}^{0}_{+1}\mathrm{e} \qquad [21.26]$$

$$ {}^{1}_{1}\mathrm{H} + {}^{2}_{1}\mathrm{H} \longrightarrow {}^{3}_{2}\mathrm{He} \qquad [21.27]$$

$$ {}^{3}_{2}\mathrm{He} + {}^{3}_{2}\mathrm{He} \longrightarrow {}^{4}_{2}\mathrm{He} + 2\,{}^{1}_{1}\mathrm{H} \qquad [21.28]$$

$$ {}^{3}_{2}\mathrm{He} + {}^{1}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{0}_{+1}\mathrm{e} \qquad [21.29]$$

Fusion is appealing as an energy source because of the availability of light isotopes on Earth and because fusion products are generally not radioactive. Despite this fact, fusion is not presently used to generate energy. The problem is that extremely high temperatures and pressures are needed to overcome the electrostatic repulsion between nuclei in order to fuse them. The lowest temperature required for any fusion is about 40,000,000 K, the temperature needed to fuse deuterium and tritium:

$$ {}^{2}_{1}\mathrm{H} + {}^{3}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{1}_{0}\mathrm{n} \qquad [21.30]$$

Fusion reactions are therefore also known as **thermonuclear reactions**.

Such high temperatures have been achieved by using an atomic bomb to initiate fusion. This is the operating principle behind a thermonuclear, or hydrogen, bomb. This approach is obviously unacceptable, however, for a power generation plant.*

Numerous problems must be overcome before fusion becomes a practical energy source. In addition to the high temperatures necessary to initiate the reaction, there is the problem of confining the reaction. No known structural material is able to withstand the enormous temperatures necessary for fusion. Much research has centered on the use of an apparatus called a *tokamak*, which uses strong magnetic fields to contain and to heat the reaction. Temperatures of over 100,000,000 K have been achieved in a tokamak. Unfortunately, scientists have not yet been able to generate more power than is consumed over a sustained period of time.

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*Historically, a nuclear weapon that relies solely on a fission process to release energy is called an atomic bomb, whereas one that also releases energy via a fusion reaction is called a hydrogen bomb.

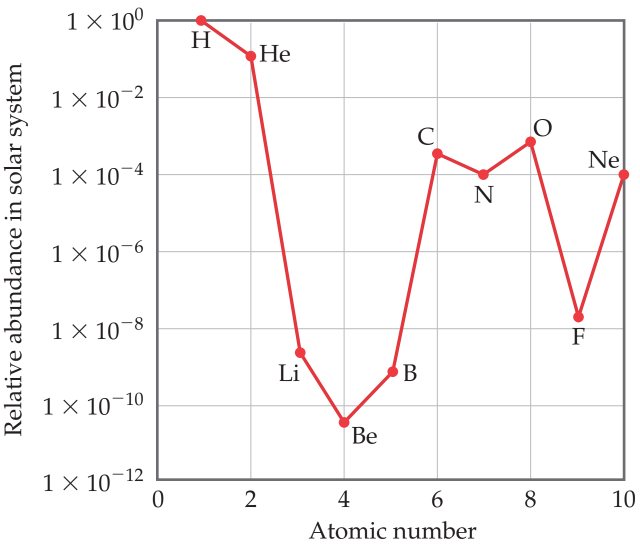
A CLOSER LOOK Nuclear Synthesis of the Elements

The lightest elements—hydrogen and helium along with very small amounts of lithium and beryllium—were formed as the universe expanded in the moments following the Big Bang. All the heavier elements owe their existence to subsequent nuclear reactions that occur in stars. These heavier elements are not all created in equal amounts, however. In our solar system, for example, carbon and oxygen are a million times more abundant than lithium and boron, and over 100 million times more abundant than beryllium (**Figure 21.21**)! In fact, of the elements heavier than helium, carbon and oxygen are the most abundant. This is more than an academic curiosity given the fact that these elements, together with hydrogen, are the most important elements for life on Earth. Let's look at the factors responsible for the relatively high abundance of carbon and oxygen in the universe.

A star is born from a cloud of gas and dust called a *nebula*. When conditions are right, gravitational forces collapse the cloud, and its core density and temperature rise until nuclear fusion commences. Hydrogen nuclei fuse to form deuterium, ${}^2_1H$, and eventually ${}^4_2He$ through the reactions shown in Equations 21.26—21.29. Because ${}^4_2He$ has a larger binding energy than any of its immediate neighbors (Figure 21.12), these reactions release an enormous amount of energy. This process, called *hydrogen burning*, is the dominant process for most of a star's lifetime.

Once a star's supply of hydrogen is nearly exhausted, several important changes occur as the star enters the next the phase of its life, and is transformed into a *red giant*. The decrease in nuclear fusion causes the core to contract, triggering an increase in core temperature and pressure. At the same time, the outer regions expand and cool enough to make the star emit red light (thus, the name *red giant*). The star now must use ${}^4_2He$ nuclei as its fuel. The simplest reaction that can occur in the He-rich core, fusion of two alpha particles to form a ${}^8_4Be$ nucleus, does occur. The binding energy per nucleon for ${}^8_4Be$ is very slightly smaller than that for ${}^4_2He$, so this fusion process is very slightly endothermic. The ${}^8_4Be$ nucleus is highly unstable (half-life of 7×10^{-17} s) and so falls apart almost immediately. In a tiny fraction of cases, however, a third ${}^4_2He$ collides with a ${}^8_4Be$ nucleus before it decays, forming carbon-12:

$$ {}^4_2He + {}^4_2He \longrightarrow {}^8_4Be $$
$$ {}^8_4Be + {}^4_2He \longrightarrow {}^{12}_6C $$



▲ **Figure 21.21 Relative abundance of elements 1–10 in the solar system.** Note the logarithmic scale used for the *y*-axis.

Some of the ${}^{12}_6C$ nuclei go on to react with alpha particles to form oxygen-16:

$$ {}^{12}_6C + {}^4_2He \longrightarrow {}^{16}_8O $$

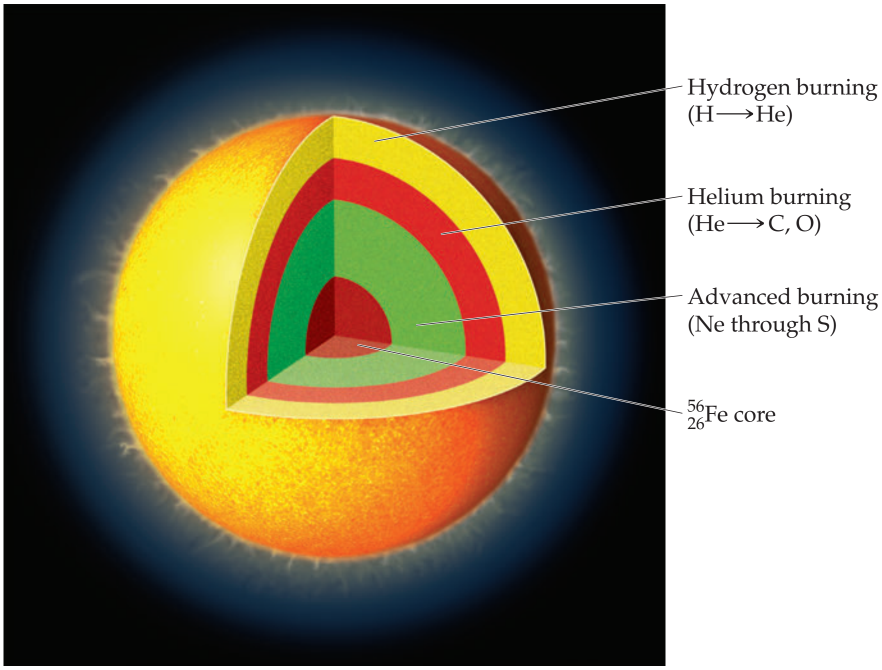
This stage of nuclear fusion is called *helium burning*. Notice that carbon, element 6, is formed without prior formation of elements 3, 4, and 5, explaining in part their unusually low abundance. Nitrogen is relatively abundant because it can be produced from carbon through a series of reactions involving proton capture and positron emission.

Most stars gradually cool and dim as the helium is converted to carbon and oxygen, ending their lives as *white dwarfs*, a phase in which stars become incredibly dense—generally about one million times denser than the Sun. The extreme density of white dwarfs is accompanied by much higher temperatures and pressures at the core, where a variety of fusion processes lead to synthesis of the elements from neon to sulfur. These fusion reactions are collectively called *advanced burning*.

Eventually, progressively heavier elements form at the core until it becomes predominantly ${}^{56}Fe$, as shown in **Figure 21.22**. Because this is such a stable nucleus, further fusion to heavier nuclei consumes energy rather than releasing it. When this happens, the fusion reactions that power the star diminish, and immense gravitational forces lead to a dramatic collapse called a supernova *explosion*. Neutron capture coupled with subsequent radioactive decays in the dying moments of such a star are responsible for the presence of all elements heavier than iron and nickel.

Without these dramatic supernova events in the past history of the universe, heavier elements that are so familiar to us, such as silver, gold, iodine, lead, and uranium, would not exist.

Related Exercises: 21.81, 21.83



▲ **Figure 21.22 Fusion processes going on in a red giant just prior to a supernova explosion.**

Self-Assessment Exercise

21.25 Based on the following atomic mass values, calculate the energy released per mole in the following nuclear reaction:

^{1}H 1.00782 u; ^{2}H, 2.01410 u; ^{3}He, 3.01603 u; ^{4}He, 4.00260 u

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

(a) -5.909×10^{3} J

(b) -1.771×10^{12} J

(b) -1.771×10^{15} J

21.26 Would fusing two stable nuclei that have mass numbers in the vicinity of 100 be an energy-releasing process?

(a) Yes

(b) No

Exercises

21.27 How much energy must be supplied to break a single ^{21}Ne nucleus into separated protons and neutrons if the nucleus has a mass of 20.98846 u? What is the nuclear binding energy for 1 mol of ^{21}Ne?

21.28 The atomic masses of nitrogen-14, titanium-48, and xenon-129 are 13.999234 u, 47.935878 u, and 128.904779 u, respectively. For each isotope, calculate (**a**) the nuclear mass, (**b**) the nuclear binding energy, (**c**) the nuclear binding energy per nucleon.

21.29 The energy from solar radiation falling on Earth is 1.07×10^{16} kJ/min. (**a**) How much loss of mass from the Sun occurs in one day from just the energy falling on Earth? (**b**) If the energy released in the reaction

$$^{235}U + ^{1}_{0}n \longrightarrow ^{141}_{56}Ba + ^{92}_{36}Kr + 3^{1}_{0}n$$

(^{235}U nuclear mass, 234.9935 u; ^{141}Ba nuclear mass, 140.8833 u; ^{92}Kr nuclear mass, 91.9021 u) is taken as typical of that occurring in a nuclear reactor, what mass of uranium-235 is required to equal 0.10% of the solar energy that falls on Earth in 1.0 day?

21.30 Using Figure 21.12, predict which of the following nuclei are likely to have the largest mass defect per nucleon: (**a**) ^{27}Al, (**b**) ^{55}Mn, (**c**) ^{165}Ho, (**d**) ^{209}Bi?

21.31 Which of the following statements about the uranium used in nuclear reactors is or are true? (i) Natural uranium has too little ^{235}U to be used as a fuel. (ii) ^{238}U cannot be used as a fuel because it forms a supercritical mass too easily. (iii) To be used as fuel, uranium must be enriched so that it is more than 50% ^{235}U in composition. (iv) The neutron-induced fission of ^{235}U releases more neutrons per nucleus than fission of ^{238}U.

21.32 (**a**) What is the function of the moderator in a nuclear reactor? (**b**) What substance acts as the moderator in a pressurized water generator? (**c**) What other substances are used as a moderator in nuclear reactor designs?

21.33 Complete and balance the nuclear equations for the following fission reactions:

(**a**) $^{239}_{94}Pu + ^{1}_{0}n \longrightarrow ^{137}_{52}Te + ^{100}_{42}Mo + \_$

(**b**) $^{256}_{100}Fm + ^{1}_{0}n \longrightarrow ^{113}_{46}Pd + \_ + 4\,^{1}_{0}n$

21.34 Which type or types of nuclear reactors have these characteristics?

(**a**) Can use natural uranium as a fuel

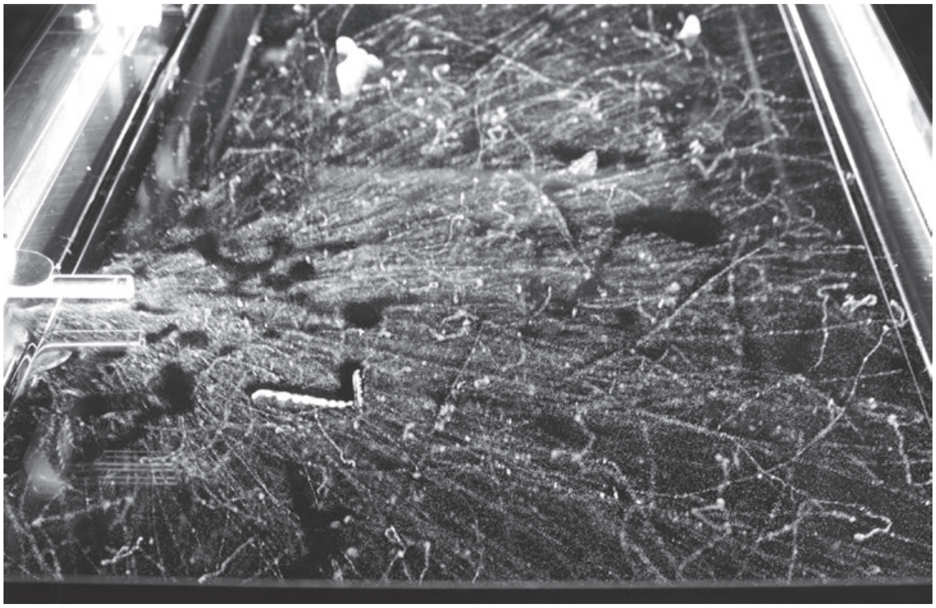
(**b**) Does not use a moderator

(**c**) Can be refueled without shutting down

Answers to Self-Assessment Exercises

21.25 (b) **21.26** (b)

21.6 Radiation in the Environment and Living Systems



A cloud chamber is a device that detects particles that have been ionized by ionizing radiation. Our exposure to natural radiation is from cosmic rays and terrestrial sources. Radon is produced by the radioactive decay of uranium and is one of the most important contributors to natural radiation from the earth. This noble gas seeps out of the ground and is dispersed by the wind. Being a heavy gas, it can accumulate in basements that are poorly sealed from the ground and have little ventilation. While radon has a short half-life, some of its decay products are more persistent. In this section, we look at a few of the effects of radiation on our bodies. By the end of this section, you should be able to

- Compare different measurements and units of radiation dosage and describe the biological effects of radiation

We are continuously bombarded by radiation from both natural and artificial sources. We are exposed to infrared, ultraviolet, and visible radiation from the Sun; radio waves from radio and television stations; microwaves from microwave ovens; X rays from medical procedures; and radioactivity from natural materials (Table 21.8). Understanding the

TABLE 21.8 Average Abundances and Activities of Natural Radionuclides†

	Potassium-40	Rubidium-87	Thorium-232	Uranium-238
Land elemental abundance (ppm)	28,000	112	10.7	2.8
Land activity (Bq/kg)	870	102	43	35
Ocean elemental concentration (mg/L)	339	0.12	1×10^{-7}	0.0032
Ocean activity (Bq/L)	12	0.11	4×10^{-7}	0.040
Ocean sediments elemental abundance (ppm)	17,000	—	5.0	1.0
Ocean sediments activity (Bq/kg)	500	—	20	12
Human body activity (Bq)	4000	600	0.08	0.4‡

†Data from "Ionizing Radiation Exposure of the Population of the United States," Report 93, 1987, and Report 160, 2009, National Council on Radiation Protection.

‡Includes lead-210 and polonium-210, daughter nuclei of uranium-238.

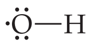
different energies of these various kinds of radiation is necessary in order to understand their different effects on matter.

When matter absorbs radiation, the radiation energy can cause atoms in the matter to be either excited or ionized. In general, radiation that causes ionization, called **ionizing radiation**, is far more harmful to biological systems than radiation that does not cause ionization. The latter, called **nonionizing radiation**, is generally of lower energy, such as radiofrequency electromagnetic radiation or slow-moving neutrons.

Most living tissue contains at least 70% water by mass. When living tissue is irradiated, water molecules absorb most of the energy of the radiation. Thus, it is common to define ionizing radiation as radiation that can ionize water, a process requiring a minimum energy of 1216 kJ/mol. Alpha, beta, and gamma rays (as well as X rays and higher-energy ultraviolet radiation) possess energies in excess of this quantity and are therefore forms of ionizing radiation.

When ionizing radiation passes through living tissue, electrons are removed from water molecules, forming highly reactive H_2O^+ ions. An H_2O^+ ion can react with another water molecule to form an H_3O^+ ion and a neutral OH molecule:

$$H_2O^+ + H_2O \longrightarrow H_3O^+ + OH \qquad [21.31]$$

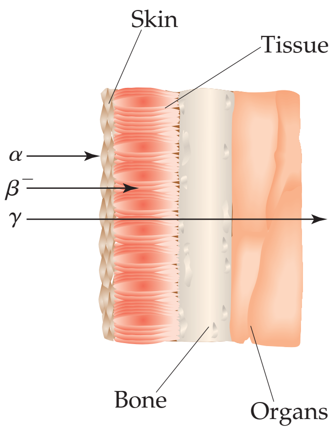


The unstable and highly reactive OH molecule is a **free radical**, a substance with one or more unpaired electrons, as seen in the Lewis structure shown in the margin. The OH molecule is also called the *hydroxyl radical*, and the presence of the unpaired electron is often emphasized by writing the species with a single dot, $\cdot$OH. In cells and tissues, hydroxyl radicals can attack biomolecules to produce new free radicals, which in turn attack yet other biomolecules. Thus, the formation of a single hydroxyl radical via Equation 21.31 can initiate a large number of chemical reactions that are ultimately able to disrupt the normal operations of cells.

The damage produced by radiation depends on the activity and energy of the radiation, the length of exposure, and whether the source is inside or outside the body. Gamma rays are particularly harmful outside the body because they penetrate human tissue very effectively, just as X rays do. Consequently, their damage is not limited to the skin. In contrast, most alpha rays are stopped by skin, and beta rays are able to penetrate only about 1 cm beyond the skin surface (**Figure 21.23**). Therefore, neither alpha rays nor beta rays are as dangerous as gamma rays, *unless* the radiation source somehow enters the body. Within the body, alpha rays are particularly dangerous because they transfer their energy efficiently to the surrounding tissue, causing considerable damage.

Go Figure

Why are alpha rays much more dangerous when the source of radiation is located inside the body?



▲ **Figure 21.23 Relative penetrating abilities of alpha, beta, and gamma radiation.**

In general, the tissues damaged most by radiation are those that reproduce rapidly, such as bone marrow, blood-forming tissues, and lymph nodes. The principal effect of extended exposure to low doses of radiation is to cause cancer. Cancer is caused by damage to the growth-regulation mechanism of cells, inducing the cells to reproduce uncontrollably. Leukemia, which is characterized by excessive growth of white blood cells, is probably the major type of radiation-caused cancer.

In light of the biological effects of radiation, it is important to determine whether any levels of exposure are safe. Unfortunately, we are hampered in our attempts to set realistic standards because we do not fully understand the effects of long-term exposure. Scientists concerned with setting health standards have used the hypothesis that the effects of radiation are proportional to exposure. *Any* amount of radiation is assumed to cause some finite risk of injury, and the effects of high dosage rates are extrapolated to those of lower ones. Other scientists believe, however, that there is a threshold below which there are no radiation risks. Until scientific evidence enables us to settle the matter with some confidence, it is safer to assume that even low levels of radiation present some danger.

Radiation Doses

Two units are commonly used to measure exposure to radiation. The **gray** (Gy), the SI unit of absorbed dose, corresponds to the absorption of 1 J of energy per

TABLE 21.9 Effects of Short-Term Exposures to Radiation

Dose (Sv)	Effect
0–0.25	No detectable clinical effects
0.25–0.50	Slight, temporary decrease in white blood cell counts
1–2	Nausea; marked decrease in white blood cell counts
5	Death of half the exposed population within 30 days

kilogram of tissue. The **rad** (*r*adiation *a*bsorbed *d*ose) corresponds to the absorption of 1×10^{-2} J of energy per kilogram of tissue. Thus, 1 Gy = 100 rad. The rad is the unit most often used in medicine.

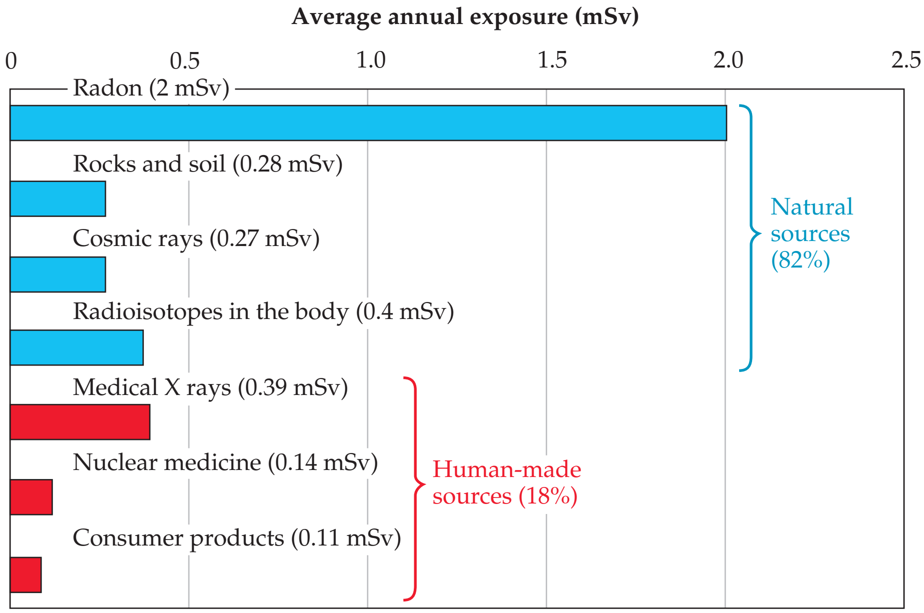
Not all forms of radiation harm biological materials to the same extent even at the same level of exposure. For example, 1 rad of alpha radiation can produce more damage than 1 rad of beta radiation. To correct for these differences, the radiation dose is multiplied by a factor that measures the relative damage caused by the radiation. This multiplication factor is known as the *relative biological effectiveness*, *RBE*. The RBE is approximately 1 for gamma and beta radiation, and 10 for alpha radiation.

The exact value of the RBE varies with dose rate, total dose, and type of tissue affected. The product of the radiation dose in rads and the RBE of the radiation give the *effective dosage* in **rem** (*r*oentgen *e*quivalent for *m*an):

$$\text{Number of rem} = (\text{number of rad})(\text{RBE}) \qquad [21.32]$$

The SI unit for effective dose is the *sievert* (Sv), obtained by multiplying the RBE times the SI unit for radiation dose, the gray; because a gray is 100 times larger than a rad, 1 Sv = 100 rem. The rem is the unit of radiation damage usually used in medicine.

The effects of short-term exposure to radiation appear in **Table 21.9**. An exposure of 6 Sv is fatal to most humans. To put this number in perspective, a typical dental X ray entails an exposure of about 5 μSv. The average exposure for a person in 1 yr due to all natural sources of ionizing radiation (called *background radiation*) is about 3.6 mSv (**Figure 21.24**).



▲ **Figure 21.24** **Sources of U.S. average annual exposure to high-energy radiation.** The total average annual exposure is 3.6 mSv.
Data from "Ionizing Radiation Exposure of the Population of the United States," Report 93, 1987 and Report 160, 2009, National Council on Radiation Protection.

CHEMISTRY AND LIFE Radiation Therapy

Healthy cells are either destroyed or damaged by high-energy radiation, leading to physiological disorders. This radiation can also destroy *unhealthy* cells, however, including cancerous cells. All cancers are characterized by runaway cell growth that can produce *malignant tumors*. These tumors can be caused by the exposure of healthy cells to high-energy radiation. Paradoxically, however, they can be destroyed by the same radiation that caused them because the rapidly reproducing cells of the tumors are very susceptible to radiation damage. Thus, cancerous cells are more susceptible to destruction by radiation than healthy ones, allowing radiation to be used effectively in the treatment of cancer. As early as 1904, physicians used the radiation emitted by radioactive substances to treat tumors by destroying the mass of unhealthy tissue. The treatment of disease by high-energy radiation is called *radiation therapy*.

Many radionuclides are currently used in radiation therapy. Some of the more commonly used ones are listed in Table 21.10. Most of them have short half-lives, meaning that they emit a great deal of radiation in a short period of time.

The radiation source used in radiation therapy may be inside or outside the body. In almost all cases, radiation therapy uses gamma radiation emitted by radioisotopes. Any alpha or beta radiation that is emitted concurrently can be blocked by appropriate packaging.

TABLE 21.10 Some Radioisotopes Used in Radiation Therapy

Isotope	Half-Life	Isotope	Half-Life
^{32}P	14.3 days	^{137}Cs	30 yr
^{60}Co	5.27 yr	^{192}Ir	74.2 days
^{90}Sr	28.8 yr	^{198}Au	2.7 days
^{125}I	60.25 days	^{222}Rn	3.82 days
^{131}I	8.04 days	^{226}Ra	1600 yr

For example, ^{192}Ir is often administered as "seeds" consisting of a core of radioactive isotope coated with 0.1 mm of platinum metal. The platinum coating stops the alpha and beta rays, but the gamma rays penetrate it readily. The radioactive seeds can be surgically implanted in a tumor.

In some cases, human physiology allows a radioisotope to be ingested. For example, most of the iodine in the human body ends up in the thyroid gland, so thyroid cancer can be treated by using large doses of ^{131}I. Radiation therapy on deep organs, where a surgical implant is impractical, often uses a ^{60}Co "gun" outside the body to shoot a beam of gamma rays at the tumor. Particle accelerators are also used as an external source of high-energy radiation for radiation therapy.

Because gamma radiation is so strongly penetrating, it is nearly impossible to avoid damaging healthy cells during radiation therapy. Many cancer patients undergoing radiation treatment experience unpleasant and dangerous side effects such as fatigue, nausea, hair loss, a weakened immune system, and occasionally even death. However, if other treatments such as *chemotherapy* (the use of drugs to combat cancer) fail, radiation therapy can be a good option.

Much current research in radiation therapy is engaged in developing new drugs that specifically target tumors using a method called *neutron capture therapy*. In this technique, a nonradioactive isotope, usually boron-10, is concentrated in the tumor by using specific tumor-seeking reagents. The boron-10 is then irradiated with neutrons, where it undergoes the following nuclear reaction producing alpha particles:

$$^{10}_{5}B + ^{1}_{0}n \longrightarrow ^{7}_{3}Li + ^{4}_{2}He$$

Tumor cells are killed or damaged by exposure to the alpha particles. Healthy tissue farther away from the tumor is unaffected because of the short-range penetrating power of alpha particles. Thus, neutron-capture therapy has the promise to be a "silver bullet" that specifically targets unhealthy cells for exposure to radiation.

Related Exercises: 21.24, 21.60, 21.64

Sample Integrative Exercise

Putting Concepts Together

Potassium ion is present in foods and is an essential nutrient in the human body. One of the naturally occurring isotopes of potassium, potassium-40, is radioactive. Potassium-40 has a natural abundance of 0.0117% and a half-life $t_{1/2} = 1.28 \times 10^9$ yr. It undergoes radioactive decay in three ways: 98.2% is by electron capture, 1.35% is by beta emission, and 0.49% is by positron emission. (**a**) Why should we expect ^{40}K to be radioactive? (**b**) Write the nuclear equations for the three modes by which ^{40}K decays. (**c**) How many $^{40}K^+$ ions are present in 1.00 g of KCl? (**d**) How long does it take for 1.00% of the ^{40}K in a sample to undergo radioactive decay?

SOLUTION

(**a**) The ^{40}K nucleus contains 19 protons and 21 neutrons. There are very few stable nuclei with odd numbers of both protons and neutrons (Section 21.2).

(**b**) Electron capture is capture of an inner-shell electron by the nucleus:

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

Beta emission is loss of a beta particle ($^{0}_{-1}e$) by the nucleus:

$$^{40}_{19}K \longrightarrow ^{40}_{20}Ca + ^{0}_{-1}e$$

Positron emission is loss of a positron ($^{0}_{+1}e$) by the nucleus:

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

Continued

(**c**) The total number of K^+ ions in the sample is:

$$(1.00\,\cancel{\text{g KCl}})\left(\frac{1\,\cancel{\text{mol KCl}}}{74.55\,\cancel{\text{g KCl}}}\right)\left(\frac{1\,\cancel{\text{mol K}^+}}{1\,\cancel{\text{mol KCl}}}\right)\left(\frac{6.022\times10^{23}\,\text{K}^+}{1\,\cancel{\text{mol K}^+}}\right) = 8.08\times10^{21}\,\text{K}^+\ \text{ions}$$

Of these, 0.0117% are $^{40}K^+$ ions:

$$(8.08\times10^{21}\,\cancel{\text{K}^+\ \text{ions}})\left(\frac{0.0117\ ^{40}\text{K}^+\ \text{ions}}{100\,\cancel{\text{K}^+\ \text{ions}}}\right) = 9.45\times10^{17}\ ^{40}\text{K}^+\ \text{ions}$$

(**d**) The decay constant (the rate constant) for the radioactive decay can be calculated from the half-life, using Equation 21.19:

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{1.28\times10^{9}\ \text{yr}} = (5.41\times10^{-10})/\text{yr}$$

The rate equation, Equation 21.20, then allows us to calculate the time required:

$$\ln\frac{N_t}{N_0} = -kt$$

$$\ln\frac{99}{100} = -[(5.41\times10^{-10})/\text{yr}]t$$

$$-0.01005 = -[(5.41\times10^{-10})/\text{yr}]t$$

$$t = \frac{-0.01005}{(-5.41\times10^{-10})/\text{yr}} = 1.86\times10^{7}\ \text{yr}$$

That is, it would take 18.6 million years for just 1.00% of the ^{40}K in a sample to decay.

Self-Assessment Exercise

21.35 If a person is uniformly irradiated by 0.010 J/kg alpha radiation, what is the effective dosage in rem?

(a) 0.010 rem

(b) 1.0 rem

(c) 10 rem

Exercises

21.36 Hydroxyl radicals can pluck hydrogen atoms from molecules ("hydrogen abstraction"), and hydroxide ions can pluck protons from molecules ("deprotonation"). Write the reaction equations and Lewis dot structures for the hydrogen abstraction and deprotonation reactions for the generic carboxylic acid R—COOH with hydroxyl radical and hydroxide ion, respectively. Why is hydroxyl radical more toxic to living systems than hydroxide ion?

21.37 A 65-kg person is accidentally exposed for 240 s to a 15-mCi source of beta radiation coming from a sample of ^{90}Sr. (**a**) What is the activity of the radiation source in disintegrations per second? In becquerels? (**b**) Each beta particle has an energy of 8.75×10^{-14} J. and 7.5% of the radiation is absorbed by the person. Assuming that the absorbed radiation is spread over the person's entire body, calculate the absorbed dose in rads and in grays. (**c**) If the RBE of the beta particles is 1.0, what is the effective dose in mrem and in sieverts? (**d**) Is the radiation dose equal to, greater than, or less than that for a typical mammogram (3 mSv)?

Answers to Self-Assessment Exercises

21.35 (c)

Chapter Summary and Key Terms

INTRODUCTION TO RADIOACTIVITY AND NUCLEAR EQUATIONS (SECTION 21.1) The nucleus of an atom contains protons and neutrons, both of which are called **nucleons**. Reactions that involve changes in atomic nuclei are called **nuclear reactions**. Nuclei that spontaneously change by emitting radiation are said to be **radioactive**. Radioactive nuclei are called **radionuclides**, and the atoms containing them are called **radioisotopes**. Radionuclides spontaneously change through a process called radioactive decay. The three most important types of radiation given off as a result of radioactive decay are **alpha (α) particles** ($^{4}_{2}He$ or α), **beta (β)**

particles ($_{-1}^{0}e$ or β^-), and **gamma (γ) radiation** ($_{0}^{0}\gamma$ or γ). **Positrons** ($_{+1}^{0}e$ or β^+), which are particles with the same mass as an electron but the opposite charge, can also be produced when a radioisotope decays.

In nuclear equations, reactant and product nuclei are represented by giving their mass numbers and atomic numbers, as well as their chemical symbol. The totals of the mass numbers on both sides of the equation are equal; the totals of the atomic numbers on both sides are also equal. There are four common modes of radioactive decay: **alpha emission**, which reduces the atomic number by 2 and the mass number by 4; **beta emission**, which increases the atomic number by 1 and leaves the mass number unchanged; and **positron emission** and **electron capture**, both of which reduce the atomic number by 1 and leave the mass number unchanged.

PATTERNS OF NUCLEAR STABILITY (SECTION 21.2) The neutron-to-proton ratio is an important factor determining nuclear stability. By comparing a nuclide's neutron-to-proton ratio with those of stable nuclei, we can predict the mode of radioactive decay. In general, neutron-rich nuclei tend to emit beta particles; proton-rich nuclei tend to either emit positrons or undergo electron capture; and heavy nuclei tend to emit alpha particles. The presence of **magic numbers** of nucleons and an even number of protons and neutrons also help determine the stability of a nucleus. A nuclide may undergo a series of decay steps before a stable nuclide forms. This series of steps is called a **radioactive decay chain** or a **nuclear disintegration series**.

Nuclear transmutations, induced conversions of one nucleus into another, can be brought about by bombarding nuclei with either charged particles or neutrons. **Particle accelerators** increase the kinetic energies of positively charged particles, allowing these particles to overcome their electrostatic repulsion by the nucleus. Nuclear transmutations are used to produce the **transuranium elements**, those elements with atomic numbers greater than that of uranium.

RADIOACTIVE DECAY RATES AND DETECTION OF RADIOACTIVITY (SECTIONS 21.3 AND 21.4) The SI unit for the activity of a radioactive source is the **becquerel** (Bq), defined as one nuclear disintegration per second. A related unit, the **curie** (Ci), corresponds to 3.7×10^{10} disintegrations per second. Nuclear decay is a first-order process. The decay rate (**activity**) is therefore directly proportional to the number of radioactive nuclei. The **half-life** of a radionuclide, which is a constant independent of temperature, is the time needed for one-half of the nuclei to decay. Some radioisotopes can be used to date objects; ^{14}C, for example, is used to date organic objects. Geiger counters and scintillation counters count the emissions from radioactive samples. The ease of detection of radioisotopes also permits their use as **radiotracers** to follow elements through reactions.

ENERGY CHANGES IN NUCLEAR REACTIONS (SECTION 21.5) The energy produced in nuclear reactions is accompanied by measurable changes of mass in accordance with Einstein's relationship, $\Delta E = c^2\Delta m$. The difference in mass between nuclei and the nucleons of which they are composed is known as the **mass defect**. The mass defect of a nuclide makes it possible to calculate its **nuclear binding energy**, the energy required to separate the nucleus into individual nucleons. Because of trends in the nuclear binding energy with atomic number, energy is produced when heavy nuclei split (**fission**) and when light nuclei fuse (**fusion**).

Uranium-235, uranium-233, and plutonium-239 undergo fission when they capture a neutron, splitting into lighter nuclei and releasing more neutrons. The neutrons produced in one fission can cause further fission reactions, which can lead to a nuclear **chain reaction**. A reaction that maintains a constant rate is said to be critical, and the mass necessary to maintain this constant rate is called a **critical mass**. A mass in excess of the critical mass is termed a **supercritical mass**.

In nuclear reactors, the fission rate is controlled to generate a constant power. The reactor core consists of fuel elements containing fissionable nuclei, control rods, a moderator, and a primary coolant. A nuclear power plant resembles a conventional power plant except that the reactor core replaces the fuel burner. There is concern about the disposal of highly radioactive nuclear wastes that are generated in nuclear power plants.

Nuclear fusion requires high temperatures because nuclei must have large kinetic energies to overcome their mutual repulsions. Fusion reactions are therefore called **thermonuclear reactions**. It is not yet possible to generate power on Earth through a controlled fusion process.

NUCLEAR CHEMISTRY AND LIVING SYSTEMS (SECTION 21.6) **Ionizing radiation** is energetic enough to remove an electron from a water molecule; radiation with less energy is called **nonionizing radiation**. Ionizing radiation generates **free radicals**, reactive substances with one or more unpaired electrons. The effects of long-term exposure to low levels of radiation are not completely understood, but there is evidence that the extent of biological damage varies in direct proportion to the level of exposure.

The amount of energy deposited in biological tissue by radiation is called the radiation dose and is measured in units of gray or rad. One **gray** (Gy) corresponds to a dose of 1 J/kg of tissue. It is the SI unit of radiation dose. The **rad** is a smaller unit; 100 rad = 1 Gy. The effective dose, which measures the biological damage created by the deposited energy, is measured in units of rem or sievert (Sv). The **rem** is obtained by multiplying the number of rad by the relative biological effectiveness (RBE); 100 rem = 1 Sv.

Learning Outcomes After studying this chapter, you should be able to:

- Write balanced nuclear equations. (Section 21.1) *Related Exercises: 21.4, 21.49*
- Predict nuclear stability and expected type of nuclear decay from the neutron-to-proton ratio of an isotope. (Section 21.2) *Related Exercises: 21.38, 21.51*
- Write balanced nuclear equations for nuclear transmutations. (Section 21.3) *Related Exercises: 21.56, 21.57*
- Calculate ages of objects and/or the amount of a radionuclide remaining after a given period of time using the half-life of the radionuclide in question. (Section 21.4) *Related Exercises: 21.43, 21.60*
- Calculate mass and energy changes for nuclear reactions. (Section 21.6) *Related Exercises: 21.29, 21.65*
- Calculate the binding energies for nuclei. (Section 21.6) *Related Exercises: 21.67, 21.68*
- Describe the difference between fission and fusion and explain how a nuclear power plant operates. (Sections 21.7 and 21.8) *Related Exercises: 21.71, 21.76*
- Compare different measurements and units of radiation dosage and describe the biological effects of radiation. (Section 21.9) *Related Exercises: 21.36, 21.78*

Key Equations

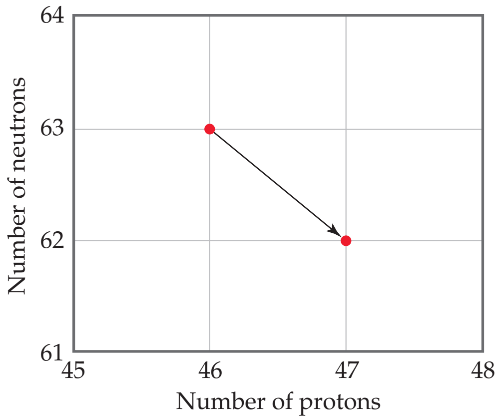
- $k = \dfrac{0.693}{t_{1/2}}$ [21.19] Relationship between nuclear decay constant and half-life; this is derived from the following equation at $N_t = \frac{1}{2}N_0$
- $\ln\dfrac{N_t}{N_0} = -kt$ [21.20] First-order rate law for nuclear decay
- $E = mc^2$ [21.22] Einstein's equation that relates mass and energy

Exercises

Visualizing Concepts

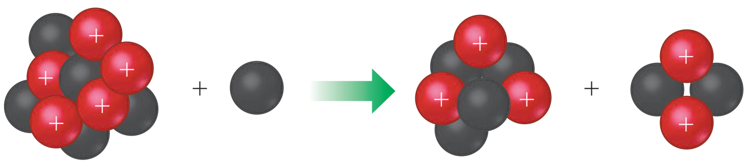
21.38 Indicate whether each of the following nuclides lies within the belt of stability in Figure 21.2: (**a**) neon-24, (**b**) chlorine-32, (**c**) tin-108, (**d**) polonium-216. For any that do not, describe a nuclear decay process that would alter the neutron-to-proton ratio in the direction of increased stability. [Section 21.2]

21.39 Write the balanced nuclear equation for the reaction represented by the diagram shown here. [Section 21.2]

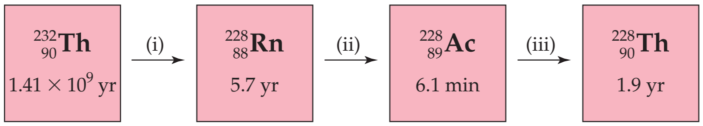


21.40 Draw a diagram similar to that shown in Exercise 21.39 that illustrates the nuclear reaction ${}^{211}_{83}\text{Bi} \longrightarrow {}^{4}_{2}\text{He} + {}^{207}_{81}\text{Tl}$. [Section 21.2]

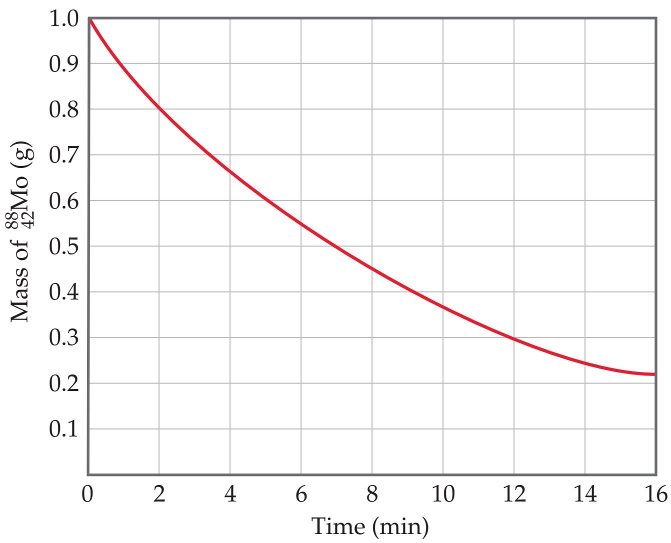
21.41 In this sketch, the red spheres represent protons and the gray spheres represent neutrons. (**a**) What are the identities of the four particles involved in the reaction depicted? (**b**) Write the transformation represented here using condensed notation. (**c**) Based on its atomic number and mass number, do you think the product nucleus is stable or radioactive? [Section 21.2]



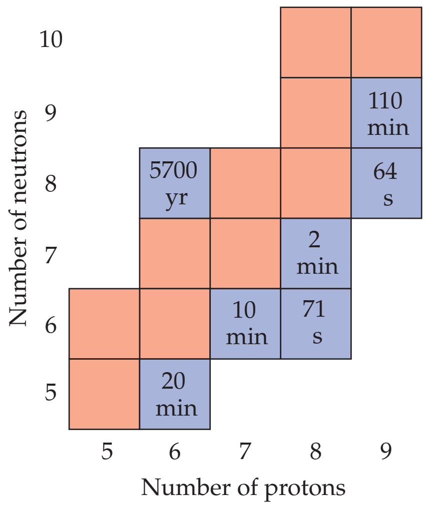
21.42 The steps drawn here show three of the steps in the radioactive decay chain for ${}^{232}_{90}\text{Th}$. The half-life of each isotope is shown below the symbol of the isotope. (**a**) Identify the type of radioactive decay for each of the steps (i), (ii), and (iii). (**b**) Which of the isotopes shown has the highest activity? (**c**) Which of the isotopes shown has the lowest activity? (**d**) The next step in the decay chain is an alpha emission. What is the next isotope in the chain? [Sections 21.2 and 21.3]



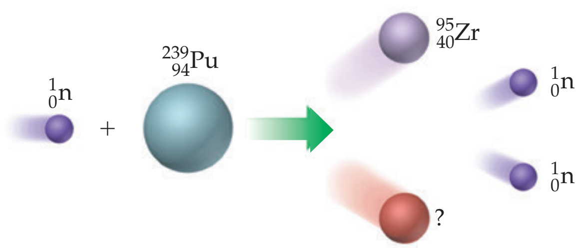
21.43 The accompanying graph illustrates the decay of ${}^{88}_{42}\text{Mo}$, which decays via positron emission. (**a**) What is the half-life of the decay? (**b**) What is the rate constant for the decay? (**c**) What fraction of the original sample of ${}^{88}_{42}\text{Mo}$ remains after 12 min? (**d**) What is the product of the decay process? [Section 21.3]



21.44 All the stable isotopes of boron, carbon, nitrogen, oxygen, and fluorine are shown in the accompanying chart (in red), along with their radioactive isotopes with $t_{1/2} > 1$ min (in blue). (**a**) Write the chemical symbols, including mass and atomic numbers, for all of the stable isotopes. (**b**) Which radioactive isotopes are most likely to decay by beta emission? (**c**) Some of the isotopes shown are used in positron emission tomography. Which ones would you expect to be most useful for this application? (**d**) Which isotope would decay to 12.5% of its original concentration after 1 hour? [Sections 21.2, 21.3, and 21.4]



21.45 The diagram shown here illustrates a fission process. (**a**) What is the unidentified product of the fission? (**b**) Use Figure 21.2 to predict whether the nuclear products of this fission reaction are stable. [Section 21.5]



Radioactivity and Nuclear Equations (Section 21.1)

21.46 Indicate the number of protons and neutrons in the following nuclei: (**a**) ${}^{239}_{94}Pu$, (**b**) ${}^{142}_{56}Ba$, (**c**) potassium-41.

21.47 What do these symbols stand for? (**a**) ${}^{1}_{1}p$, (**b**) ${}^{0}_{-1}e$, (**c**) ${}^{0}_{+1}e$

21.48 Write balanced nuclear equations for the following processes: (**a**) radon-198 undergoes alpha emission; (**b**) thorium-234 undergoes beta emission; (**c**) copper-61 undergoes positron emission; (**d**) silver-106 undergoes electron capture.

21.49 Decay of which nucleus will lead to the following products: (**a**) uranium-235 by alpha decay; (**b**) aluminum-26 by positron emission; (**c**) deuterium by alpha decay; (**d**) yttrium-90 by beta decay?

21.50 The naturally occurring radioactive decay series that begins with ${}^{235}_{92}U$ stops with formation of the stable ${}^{207}_{82}Pb$ nucleus. The decays proceed through a series of alpha-particle and beta-particle emissions. How many of each type of emission are involved in this series?

Patterns of Nuclear Stability (Section 21.2)

21.51 Predict the type of radioactive decay process for the following radionuclides: (**a**) ${}^{15}_{8}O$, (**b**) ${}^{41}_{21}Sc$, (**c**) uranium-237, (**d**) sulphur-35.

21.52 One of the nuclides in each of the following pairs is radioactive. Predict which is radioactive and which is stable: (**a**) ${}^{92}_{44}Ru$ and ${}^{102}_{44}Ru$, (**b**) ${}^{138}_{56}Ba$ and ${}^{139}_{56}Ba$, (**c**) tin-109 and tin-120.

21.53 Which of the following nuclides have magic numbers of both protons and neutrons: (**a**) beryllium-10, (**b**) silicon-28, (**c**) chromium-52, (**d**) nickel-56, (**e**) krypton-84?

21.54 Which of the following statements best explains why alpha emission is relatively common, but proton emission is extremely rare?

(**a**) Alpha particles are very stable because of magic numbers of protons and neutrons.

(**b**) Alpha particles occur in the nucleus.

(**c**) Alpha particles are the nuclei of an inert gas.

(**d**) An alpha particle has a higher charge than a proton.

21.55 Which statement best explains why nuclear transmutations involving neutrons are generally easier to accomplish than those involving protons or alpha particles?

(**a**) Neutrons are not a magic number particle.

(**b**) Neutrons do not have an electrical charge.

(**c**) Neutrons are smaller than protons or alpha particles.

(**d**) Neutrons are attracted to the nucleus even at long distances, whereas protons and alpha particles are repelled.

21.56 Complete and balance the following nuclear equations by supplying the missing particle:

(**a**) ${}^{239}_{94}Pu + {}^{1}_{0}n \longrightarrow {}^{0}_{-1}e + ?$

(**b**) ${}^{238}_{92}U + {}^{4}_{2}He \longrightarrow 3\,{}^{1}_{0}n + ?$

(**c**) ${}^{218}_{85}At \longrightarrow {}^{0}_{-1}e + ?$

(**d**) ${}^{146}_{62}Sm \longrightarrow {}^{142}_{60}Nd + ?$

(**e**) ${}^{118}_{53}I + {}^{0}_{-1}e \longrightarrow ?$

21.57 Write balanced equations for (**a**) ${}^{238}_{92}U(\alpha, n){}^{241}_{94}Pu$, (**b**) ${}^{14}_{7}N(\alpha, p){}^{17}_{8}O$, (**c**) ${}^{56}_{26}Fe(\alpha, \beta^-){}^{60}_{29}Cu$.

Rates of Radioactive Decay (Section 21.3)

21.58 Each statement that follows refers to a comparison between two radioisotopes, A and X. Indicate whether each of the following statements is true or false.

(**a**) If the half-life for A is shorter than the half-life for X, A has a larger decay rate constant.

(**b**) If X is "not radioactive," its half-life is essentially zero.

(**c**) If A has a half-life of 10 yr, and X has a half-life of 10,000 yr, A would be a more suitable radioisotope to measure processes occurring on the 40-yr time scale.

21.59 Some watch dials are coated with a phosphor, like ZnS, and a polymer in which some of the ${}^{1}H$ atoms have been replaced by ${}^{3}H$ atoms, tritium. The phosphor emits light when struck by the beta particle from the tritium decay, causing the dials to glow in the dark. The half-life of tritium is 12.3 yr. If the light given off is assumed to be directly proportional to the amount of tritium, by how much will a dial be dimmed in a watch that is 50 yr old?

21.60 Cobalt-60 is a strong gamma emitter that has a half-life of 5.26 yr. The cobalt-60 in a radiotherapy unit must be replaced when its radioactivity falls to 75% of the original sample. If an original sample was purchased in June 2016, when will it be necessary to replace the cobalt-60?

21.61 Radium-226, which undergoes alpha decay, has a half-life of 1600 yr. (**a**) How many alpha particles are emitted in 5.0 min by a 10.0-mg sample of ^{226}Ra? (**b**) What is the activity of the sample in mCi?

21.62 A 10.00 g plant fossil from an archaeological site is found to have a ^{14}C activity of 3094 disintegrations over a period of ten hours. A living plant is found to have a ^{14}C activity of 9207 disintegrations over the same period of time for an equivalent amount of sample with respect to the total contents of carbon. Given that the half-life of ^{14}C is 5715 years, how old is the plant fossil?

21.63 Phosphorus-32 is commonly used in nuclear medicine for the identification of malignant tumors. It decays to sulphur-32 with a half-life of 14.29 days. If a patient is given 3.5 mg of phosphorus-32, how much phosphorus-32 will remain after 1 month (i.e. 30 days)?

Detection of Radioactivity (Section 21.4)

21.64 Why is it important that radioisotopes used as diagnostic tools in nuclear medicine produce gamma radiation when they decay? Why are alpha emitters not used as diagnostic tools?

Energy Changes in Nuclear Reactions (Section 21.5)

21.65 An analytical laboratory balance typically measures mass to the nearest 0.1 mg. What energy change would accompany the loss of 0.1 mg in mass?

21.66 The thermite reaction, $Fe_2O_3(s) + 2\,Al(s) \longrightarrow 2\,Fe(s) + Al_2O_3(s)$, $\Delta H^\circ = -851.5$ kJ/mol, is one of the most exothermic reactions known. Because the heat released is sufficient to melt the iron product, the reaction is used to weld metal under the ocean. How much heat is released per mole of Al_2O_3 produced? How does this amount of thermal energy compare with the energy released when 2 mol of protons and 2 mol of neutrons combine to form 1 mol of alpha particles?

21.67 How much energy must be supplied to break a single aluminum-27 nucleus into separated protons and neutrons if an aluminum-27 atom has a mass of 26.9815386 u? How much energy is required for 100.0 g of aluminum-27? (The mass of an electron is given on the inside back cover.)

21.68 The atomic masses of hydrogen-2 (deuterium), helium-4, and lithium-6 are 2.014102 u, 4.002602 u, and 6.0151228 u, respectively. For each isotope, calculate (**a**) the nuclear mass, (**b**) the nuclear binding energy, (**c**) the nuclear binding energy per nucleon. (**d**) Which of these three isotopes has the largest nuclear binding energy per nucleon? Does this agree with the trends plotted in Figure 21.12?

21.69 Based on the following atomic mass values = ^{2}H, 2.01410 u; ^{3}H, 3.01605 u; ^{3}He, 3.01603 u; ^{4}He, 4.00260 u—and the mass of the neutron given in the text, calculate the energy released per mole in each of the following nuclear reactions, all of which are possibilities for a controlled fusion process.

(**a**) ${}^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n$

(**b**) ${}^{2}_{1}H + {}^{2}_{1}H \longrightarrow {}^{3}_{2}He + {}^{1}_{0}n$

21.70 The isotope ${}^{62}_{28}Ni$ has the largest binding energy per nucleon of any isotope. Calculate this value from the atomic mass of nickel-62 (61.928345 u) and compare it with the value given for iron-56 in Table 21.7.

21.71 (**a**) Which of the following are required characteristics of an isotope to be used as a fuel in a nuclear power reactor? (i) It must emit gamma radiation. (ii) On decay, it must release two or more neutrons. (iii) It must have a half-life less than one hour. (iv) It must undergo fission upon the absorption of a neutron. (**b**) What is the most common fissionable isotope in a commercial nuclear power reactor?

21.72 What is the function of the control rods in a nuclear reactor? What substances are used to construct control rods? Why are these substances chosen?

21.73 Complete and balance the nuclear equations for the following fission or fusion reactions:

(**a**) ${}^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{92}_{36}Kr + {}^{141}_{56}Zn + \_$

(**b**) ${}^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + \_$

21.74 A portion of the Sun's energy comes from the reaction

$$4\,{}^{1}_{1}H \longrightarrow {}^{4}_{2}He + 2\,{}^{0}_{1}e$$

which requires a temperature of 10^6 to 10^7 K. Use the mass of the helium-4 nucleus given in Table 21.7 to determine how much energy is released per mol of hydrogen atoms.

21.75 The spent fuel elements from a fission reactor are much more intensely radioactive than the original fuel elements. (**a**) What does this tell you about the products of the fission process in relationship to the belt of stability, Figure 21.2? (**b**) Given that only two or three neutrons are released per fission event and knowing that the nucleus undergoing fission has a neutron-to-proton ratio characteristic of a heavy nucleus, what sorts of decay would you expect to be dominant among the fission products?

21.76 Which type or types of nuclear reactors have these characteristics?

(**a**) Does not use a secondary coolant

(**b**) Creates more fissionable material than it consumes

(**c**) Uses a gas, such as He or CO_2, as the primary coolant

21.77 Which are not classified as ionizing radiation: gamma rays, beta particles, radio waves used in radio and television, and infrared radiation from sun?

Radiation in the Environment (Sections 21.6)

21.78 A laboratory rat is exposed to an alpha-radiation source whose activity is 14.3 mCi. (**a**) What is the activity of the radiation in disintegrations per second? In becquerels? (**b**) The rat has a mass of 385 g and is exposed to the radiation for 14.0 s, absorbing 35% of the emitted alpha particles, each having an energy of 9.12×10^{-13} J. Calculate the absorbed dose in millirads and grays. (**c**) If the RBE of the radiation is 9.5, calculate the effective absorbed dose in mrem and Sv.

Additional Exercises

21.79 The table to the right gives the number of protons (p) and neutrons (n) for four nuclides. (**a**) Write the symbol for each of the isotopes based on the information given in the table. (**b**) Which of the isotopes is most likely to be unstable? (**c**) Which of the isotopes involves a magic number of protons and/or neutrons? (**d**) Which isotope will yield iron-58 following positron emission?

	(i)	(ii)	(iii)	(iv)
p	27	27	28	28
n	30	31	28	30

21.80 Assume that Bismuth-213 decays to a stable nucleus by a series of two alpha and two beta emissions. What is the stable nucleus that is formed?

21.81 Equation 21.28 is the nuclear reaction responsible for much of the helium-4 production in our Sun. How much energy is released in this reaction?

21.82 Chlorine has two stable nuclides, ^{35}Cl and ^{37}Cl. In contrast, ^{36}Cl is a radioactive nuclide that decays by beta emission. (**a**) What is the product of decay of ^{36}Cl? (**b**) Based on the empirical rules about nuclear stability, explain why the nucleus of ^{36}Cl is less stable than either ^{35}Cl or ^{37}Cl.

21.83 When two protons fuse in a star, the product is ^{2}H plus a positron. Write the nuclear equation for this process.

21.84 Nuclear scientists have synthesized new elements and isotopes, which are not known in nature using heavy-ion bombardment techniques in high-energy particle accelerators. Complete and balance the following reactions:

(**a**) ${}^{12}_{6}C + {}^{12}_{6}C \longrightarrow ? + {}^{4}_{2}He$

(**b**) ${}^{6}_{3}Li + {}^{63}_{28}Ni \longrightarrow ?$

(**c**) ${}^{252}_{98}Cf + {}^{10}_{5}B \longrightarrow ?$

(**d**) ${}^{238}_{92}U + {}^{12}_{6}C \longrightarrow ? + 4\,{}^{1}_{0}n$

21.85 In 2002, a team of scientists from Russia and the United States reported the creation of the first atom of element 118, which is named oganesson, and whose symbol is Og. The synthesis involved the collision of californium-249 atoms with accelerated ions of an atom which we will denote X. In the synthesis, an oganesson-294 is formed together with three neutrons.

$$ {}^{249}_{98}Cf + X \longrightarrow {}^{294}_{118}Og + 3\,{}^{1}_{0}n $$

(**a**) What are the identities of isotopes X? (**b**) Isotope X is unusual in that it is very long-lived (its half-life is on the order of 10^{19} yr) in spite of having an unfavorable neutron-to-proton ratio (Figure 21.1). Can you propose a reason for its unusual stability? (**c**) Oganesson-294 decays into livermorium-290 by alpha decay. Write a balanced equation for this.

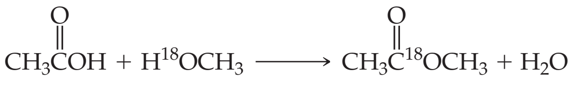
21.86 The synthetic radioisotope, phosphorus-32, which decays by beta emission, is used as radioactive labelled for DNA molecules. The following data were collected on a sample of phosphorus:

Ci	Time (days)
80	0
49	10
30	20
18	30
10.5	40
6.3	50
4.2	60

Using these data, make an appropriate graph and curve fit to determine the half-life.

21.87 According to current regulations, the maximum permissible dose of strontium-90 in the body of an adult is 1 μCi (1×10^{-6} Ci). Using the relationship rate $= kN$, calculate the number of atoms of strontium-90 to which this dose corresponds. To what mass of strontium-90 does this correspond? The half-life for strontium-90 is 28.8 yr.

21.88 Methyl acetate (CH_3COOCH_3) is formed by the reaction of acetic acid with methanol. If the methanol is labeled with oxygen-18, the oxygen-18 ends up in the methyl acetate:



(**a**) Do the C—OH bond of the acid and the O—H bond of the alcohol break in the reaction, or do the O—H bond of the acid and the C—OH bond of the alcohol break? (**b**) Imagine a similar experiment using the radioisotope ^{3}H, which is called *tritium* and is usually denoted T. Would the reaction between CH_3COOH and $TOCH_3$ provide the same information about which bond is broken as does the experiment with $H^{18}OCH_3$?

21.89 Each of the following transmutations produces a radionuclide used in positron emission tomography (PET). (**a**) In equations (i) and (ii), identify the species signified as "X." (**b**) In equation (iii), one of the species is indicated as "d." What do you think it represents?

(i) $^{14}N(p, \alpha)X$

(ii) $^{18}O(p, X)^{18}F$

(iii) $^{14}N(d, n)^{15}O$

21.90 The nuclear masses of ^{7}Be, ^{9}Be, and ^{10}Be are 7.0147, 9.0100, and 10.0113 u, respectively. Which of these nuclei has the largest binding energy per nucleon?

21.91 A 26.00 g sample of water containing tritium, ${}^{3}_{1}H$, emits 1.50×10^{3} beta particles per second. Tritium is a weak beta emitter with a half-life of 12.3 yr. What fraction of all the hydrogen in the water sample is tritium?

21.92 The Sun radiates energy into space at the rate of 3.9×10^{26} J/s. (**a**) Calculate the rate of mass loss from the Sun in kg/s. (**b**) How does this mass loss arise? (**c**) It is estimated that the Sun contains 9×10^{56} free protons. How many protons per second are consumed in nuclear reactions in the Sun?

21.93 The average energy released in the fission of a single uranium-235 nucleus is about 3×10^{-11} J. If the conversion of this energy to electricity in a nuclear power plant is 40% efficient, what mass of uranium-235 undergoes fission in a year in a plant that produces 1000 megawatts? Recall that a watt is 1 J/s.

21.94 Tests on human subjects in Boston in 1965 and 1966, following the era of atomic bomb testing, revealed average quantities of about 2 pCi of plutonium radioactivity in the average person. How many disintegrations per second does this level of activity imply? If each alpha particle deposits 8×10^{-13} J of energy and if the average person weighs 75 kg, calculate the number of grays and sieverts of radiation in 1 yr from such a level of plutonium.

Integrative Exercises

21.95 A 0.53 g sample of fludeoxyglucose ($C_6H_{11}FO_5$) contains radioactive fluorine-18 (whose atomic mass is 18.0 u). If 68.3% of the fluorine atoms in the sample are fluorine-18 and the remainder are naturally occurring nonradioactive fluorine-19 atoms, how many disintegrations per second are produced by this sample? The half-life of fluorine-18 is 110 min.

21.96 Calculate the mass of methane, $CH_4(g)$, that must be burned in air to evolve the same quantity of energy as produced by the fusion of 1.0 g of deuterium and 1.5 g of tritium in the following fusion reaction:

$$^{2}_{1}H + ^{3}_{1}H \longrightarrow ^{4}_{2}He + ^{1}_{0}n$$

Assume that all the products of the combustion of methane are in their gas phases. Use data from Exercise 21.28, Appendix C, and the inside covers of the text. The standard enthalpy of formation of methane is -74.8 kJ/mol.

21.97 Naturally found uranium consists of 99.274% ^{238}U, 0.720% ^{235}U, and 0.006% ^{233}U. As we have seen, ^{235}U is the isotope that can undergo a nuclear chain reaction. Most of the ^{235}U used in the first atomic bomb was obtained by gaseous diffusion of uranium hexafluoride, $UF_6(g)$. (**a**) What is the mass of UF_6 in a 30.0 L vessel of UF_6 at a pressure of 92.7 kPa at 350 K? (**b**) What is the mass of ^{235}U in the sample described in part (a)? (**c**) Now suppose that the UF_6 is diffused through a porous barrier and that the change in the ratio of ^{238}U and ^{235}U in the diffused gas can be described by Equation 10.23. What is the mass of ^{235}U in a sample of the diffused gas analogous to that in part (a)? (**d**) After one more cycle of gaseous diffusion, what is the percentage of $^{235}UF_6$ in the sample?

21.98 Polonium-210 is a powerful alpha emitter. A sample of polonium-210 having an activity of 85.2 Ci is stored in a 25.0-mL sealed container at 30 °C for 15 hours. (**a**) How many alpha particles are formed during this time? (**b**) Assuming that each alpha particle is converted to a helium atom, what is the partial pressure, in kPa, of helium gas in the container after this 15-hour period?

21.99 Charcoal samples from Stonehenge in England were burned in O_2, and the resultant CO_2 gas bubbled into a solution of $Ca(OH)_2$ (limewater), resulting in the precipitation of $CaCO_3$. The $CaCO_3$ was removed by filtration and dried. A 788 mg sample of the $CaCO_3$ had a radioactivity of 1.5×10^{-2} Bq due to carbon-14. By comparison, living organisms undergo 15.3 disintegrations per minute per gram of carbon. Using the half-life of carbon-14, 5700 yr, calculate the age of the charcoal sample.

21.100 A 2.5 mL sample of 0.188 *M* silver nitrate solution was mixed with 2.5 mL of 0.188 *M* sodium chloride solution labeled with radioactive chlorine-36. The activity of the initial sodium chloride solution was 2.46×10^6 Bq/mL. After the resultant precipitate was removed by filtration, the remaining filtrate was found to have an activity of 175 Bq/mL. (**a**) Write a balanced chemical equation for the reaction that occurred. (**b**) Calculate the K_{sp} for the precipitate under the conditions of the experiment.

Design an Experiment

Because radioactivity can have harmful effects on human health, very stringent experimental procedures and precautions are required when undertaking experiments on radioactive materials. As such, we typically do not have experiments involving radioactive substances in general chemistry laboratories. We can nevertheless ponder the design of some hypothetical experiments that would allow us to explore some of the properties of radium, which was discovered by Marie and Pierre Curie in 1898.

(**a**) A key aspect of the discovery of radium was Marie Curie's observation that *pitchblende*, a natural ore of uranium, had greater radioactivity than pure uranium metal. Design an experiment to reproduce this observation and to obtain a ratio of the activity of pitchblende relative to that of pure uranium.

(**b**) Radium was first isolated as halide salts. Suppose you had pure samples of radium metal and radium bromide. The sample sizes are on the order of milligrams and are not amenable to the usual forms of elemental analysis. Could you use a device that measures radioactivity quantitatively to determine the empirical formula of radium bromide? What information must you use that the Curies may not have had at the time of their discovery?

(**c**) Suppose you had a 1-yr time period in order to measure the half-life of radium and related elements. You have some pure samples and a device that measures radioactivity quantitatively. Could you determine the half-life of the elements in the samples? Would you have different experimental constraints depending on whether the half-life were 10 yr or 1000 yr?

(**d**) Before its negative health effects were better understood, small amounts of radium salts were used in "glow in the dark" watches, such as the one shown here. The glow is not due to the radioactivity of radium directly; rather, the radium is combined with a luminescent substance, such as zinc sulfide, which glows when it is exposed to radiation. Suppose you had pure samples of radium and zinc sulfide. How could you determine whether the glow of zinc sulfide is due to alpha, beta, or gamma radiation? What type of device could you design to use the glow as a quantitative measure of the amount of radioactivity in a sample?