

# Basic Atomic and Nuclear Physics

Radioactivity is a process involving events in individual atoms and nuclei. Before discussing radioactivity, therefore, it is worthwhile to review some of the basic concepts of atomic and nuclear physics.

## A. QUANTITIES AND UNITS

### 1. Types of Quantities and Units

Physical properties and processes are described in terms of *quantities* such as time and energy. These quantities are measured in *units* such as seconds and joules. Thus a quantity describes *what* is measured, whereas a unit describes *how much*.

Physical quantities are characterized as fundamental or derived. A *base* quantity is one that “stands alone”; that is, no reference is made to other quantities for its definition. Usually, base quantities and their units are defined with reference to standards kept at national or international laboratories. Time (s or sec), distance (m), and mass (kg) are examples of base quantities. *Derived* quantities are defined in terms of combinations of base quantities. Energy ( $\text{kg} \cdot \text{m}^2/\text{sec}^2$ ) is an example of a derived quantity.

The international scientific community has agreed to adopt so-called System International (SI) units as the standard for scientific communication. This system is based on seven base quantities in metric units, with all other quantities and units derived by appropriate definitions from them. The four quantities of mass, length, time and electrical charge are most relevant to nuclear medicine. The use of specially defined quantities (e.g., “atmospheres” of barometric pressure) is specifically discouraged. It is hoped that this will

improve scientific communication, as well as eliminate some of the more irrational units (e.g., feet and pounds). A useful discussion of the SI system, including definitions and values of various units, can be found in reference 1.

SI units or their metric subunits (e.g., centimeters and grams) are the standard for this text; however, in some instances traditional or other non-SI units are given as well (in parentheses). This is done because some traditional units still are used in the day-to-day practice of nuclear medicine (e.g., units of activity and absorbed dose). In other instances, SI units are unreasonably large (or small) for describing the processes of interest and specially defined units are more convenient and widely used. This is particularly true for units of mass and energy, as discussed in the following section.

### 2. Mass and Energy Units

Events occurring at the atomic level, such as radioactive decay, involve amounts of mass and energy that are very small when described in SI or other conventional units. Therefore they often are described in terms of specially defined units that are more convenient for the atomic scale.

The basic unit of mass is the *unified atomic mass unit*, abbreviated u. One u is defined as being equal to exactly  $\frac{1}{12}$  the mass of an unbound  $^{12}\text{C}$  atom\* at rest and in its ground state. The conversion from SI mass units to unified atomic mass units is<sup>1</sup>

$$1 \text{ u} = 1.66054 \times 10^{-27} \text{ kg} \quad (2-1)$$

\*Atomic notation is discussed in Section D.2.

The universal mass unit often is called a *Dalton* (Da) when expressing the masses of large biomolecules. The units are equivalent (i.e., 1 Da = 1 u). Either unit is convenient for expressing atomic or molecular masses, because a hydrogen atom has a mass of approximately 1 u or 1 Da.

The basic unit of energy is the *electron volt* (eV). One eV is defined as the amount of energy acquired by an electron when it is accelerated through an electrical potential of 1 V. Basic multiples are the kiloelectron volt (keV) (1 keV = 1000 eV) and the megaelectron volt (MeV) (1 MeV = 1000 keV = 1,000,000 eV). The conversion from SI energy units to the electron volt is

$$1 \text{ eV} = 1.6022 \times 10^{-19} \text{ kg} \cdot \text{m}^2/\text{sec}^2 \quad (2-2)$$

Mass  $m$  and energy  $E$  are related to each other by Einstein's equation  $E = mc^2$ , in which  $c$  is the velocity of light (approximately  $3 \times 10^8$  m/sec in vacuum). According to this equation, 1 u of mass is equivalent to 931.5 MeV of energy.

Relationships between various units of mass and energy are summarized in Appendix A. Universal mass units and electron volts are very small, yet, as we shall see, they are quite appropriate to the atomic scale.

## B. RADIATION

The term *radiation* refers to “energy in transit.” In nuclear medicine, we are interested principally in the following two specific forms of radiation:

1. Particulate radiation, consisting of atomic or subatomic particles (electrons, protons, etc.) that carry energy in the form of kinetic energy of mass in motion.

2. Electromagnetic radiation, in which energy is carried by oscillating electrical and magnetic fields traveling through space at the speed of light.

Radioactive decay processes, discussed in Chapter 3, result in the emission of radiation in both of these forms.

The wavelength,  $\lambda$ , and frequency,  $\nu$ , of the oscillating fields of electromagnetic radiation are related by:

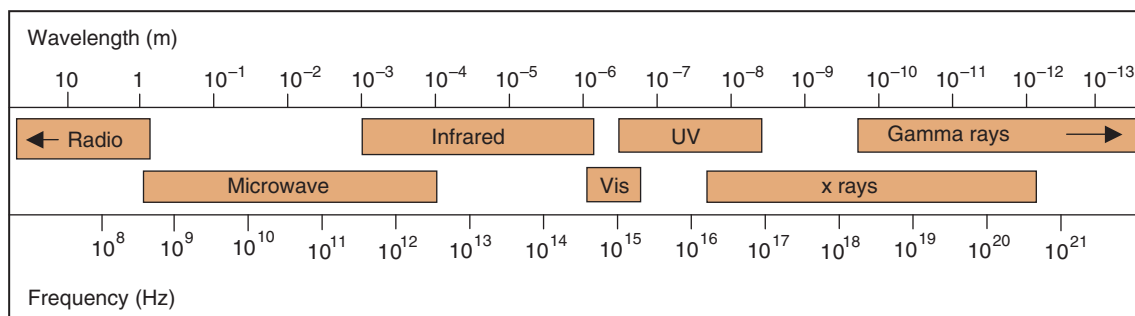
$$\lambda \times \nu = c \quad (2-3)$$

where  $c$  is the velocity of light.

Most of the more familiar types of electromagnetic radiation (e.g., visible light and radio waves) exhibit “wavelike” behavior in their interactions with matter (e.g., diffraction patterns and transmission and detection of radio signals). In some cases, however, electromagnetic radiation behaves as discrete “packets” of energy, called *photons* (also called *quanta*). This is particularly true for interactions involving individual atoms. Photons have no mass or electrical charge and also travel at the velocity of light. These characteristics distinguish them from the forms of particulate radiation mentioned earlier. The energy of the photon  $E$ , in kiloelectron volts, and the wavelength of its associated electromagnetic field  $\lambda$  (in nanometers) are related by

$$E(\text{keV}) = 1.24/\lambda(\text{nm}) \quad (2-4)$$

Figure 2-1 illustrates the photon energies for different regions of the electromagnetic spectrum. Note that x rays and  $\gamma$  rays occupy the highest-energy, shortest-wavelength end of the spectrum; x-ray and  $\gamma$ -ray photons have energies in the keV-MeV range, whereas visible light photons, for example, have



**FIGURE 2-1** Schematic representation of the different regions of the electromagnetic spectrum. Vis, visible light; UV, ultraviolet light.

energies of only a few electron volts. As a consequence of their high energies and short wavelengths, x rays and  $\gamma$  rays interact with matter quite differently from other, more familiar types of electromagnetic radiation. These interactions are discussed in detail in Chapter 6.

## C. ATOMS

### 1. Composition and Structure

All matter is composed of atoms. An atom is the smallest unit into which a chemical element can be broken down without losing its chemical identity. Atoms combine to form molecules and chemical compounds, which in turn combine to form larger, macroscopic structures.

The existence of atoms was first postulated on philosophical grounds by Ionian scholars in the 5th century BC. The concept was formalized into scientific theory early in the 19th century, owing largely to the work of the chemist, John Dalton, and his contemporaries. The exact structure of atoms was not known, but at that time they were believed to be indivisible. Later in the century (1869), Mendeleev produced the first *periodic table*, an ordering of the chemical elements according to the weights of their atoms and arrangement in a grid according to their chemical properties. For a time it was believed that completion of the periodic table would represent the final step in understanding the structure of matter.

Events of the late 19th and early 20th centuries, beginning with the discovery of x rays by Roentgen (1895) and radioactivity by Becquerel (1896), revealed that atoms had a substructure of their own. In 1910, Rutherford presented experimental evidence indicating that atoms consisted of a massive, compact, positively charged core, or *nucleus*, surrounded by a diffuse cloud of relatively light, negatively charged *electrons*. This model came to be known as the *nuclear atom*. The number of positive charges in the nucleus is called the *atomic number* of the nucleus ( $Z$ ). In the electrically neutral atom, the number of orbital electrons is sufficient to balance exactly the number of positive charges,  $Z$ , in the nucleus. The chemical properties of an atom are determined by orbital electrons; therefore the atomic number  $Z$  determines the *chemical element* to which the atom belongs. A listing of chemical elements and their atomic numbers is given in Appendix B.

According to classical theory, orbiting electrons should slowly lose energy and spiral into the nucleus, resulting in atomic “collapse.” This obviously is not what happens. The simple nuclear model therefore needed further refinement. This was provided by Niels Bohr in 1913, who presented a model that has come to be known as the *Bohr atom*. In the Bohr atom there is a set of stable electron orbits, or “shells,” in which electrons can exist indefinitely without loss of energy. The diameters of these shells are determined by *quantum numbers*, which can have only integer values ( $n = 1, 2, 3, \dots$ ). The innermost shell ( $n = 1$ ) is called the K shell, the next the L shell ( $n = 2$ ), followed by the M shell ( $n = 3$ ), N shell ( $n = 4$ ), and so forth.

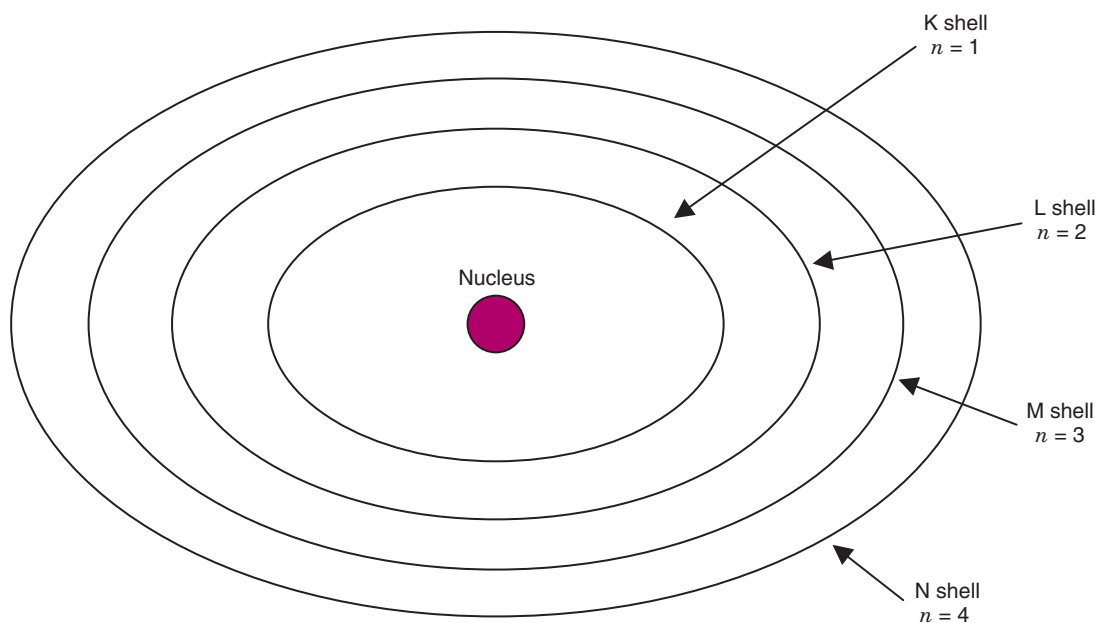
Each shell actually comprises a set of orbits, called *substates*, which differ slightly from one another. Each shell has  $2n - 1$  substates, in which  $n$  is the quantum number of the shell. Thus the K shell has only one substate; the L shell has three substates, labeled  $L_I$ ,  $L_{II}$ ,  $L_{III}$ ; and so forth. Figure 2-2 is a schematic representation of the K, L, M, and N shells of an atom.

The Bohr model of the atom was further refined with the statement of the *Pauli Exclusion Principle* in 1925. According to this principle, no two orbital electrons in an atom can move with exactly the same motion. Because of different possible electron “spin” orientations, more than one electron can exist in each substate; however, the number of electrons that can exist in any one shell or its substates is limited. For a shell with quantum number  $n$ , the maximum number of electrons allowed is  $2n^2$ . Thus the K shell ( $n = 1$ ) is limited to two electrons, the L shell ( $n = 2$ ) to eight electrons, and so forth.

The Bohr model is actually an oversimplification. According to modern theories, the orbital electrons do not move in precise circular orbits but rather in imprecisely defined “regions of space” around the nucleus, sometimes actually passing through the nucleus; however, the Bohr model is quite adequate for the purposes of this text.

### 2. Electron Binding Energies and Energy Levels

In the most stable configuration, orbital electrons occupy the innermost shells of an atom, where they are most “tightly bound” to the nucleus. For example, in carbon, which has a total of six electrons, two electrons (the maximum number allowed) occupy the K



**FIGURE 2-2** Schematic representation of the Bohr model of the atom;  $n$  is the quantum number of the shell. Each shell has multiple substates, as described in the text.

shell, and the four remaining electrons are found in the L shell. Electrons can be moved to higher shells or completely removed from the atom, but doing so requires an energy input to overcome the forces of attraction that “bind” the electron to the nucleus. The energy may be provided, for example, by a particle or a photon striking the atom.

The energy required to completely remove an electron from a given shell in an atom is called the *binding energy* of that shell. It is symbolized by the notation  $K_B$  for the K shell,\*  $L_B$  for the L shell ( $L_{IB}$ ,  $L_{IIB}$ ,  $L_{IIIB}$  for the L shell substates), and so forth. Binding energy is greatest for the innermost shell, that is,  $K_B > L_B > M_B$ . Binding energy also increases with the positive charge (atomic number  $Z$ ) of the nucleus, because a greater positive charge exerts a greater force of attraction on an electron. Therefore binding energies are greatest for the heaviest elements. Values of K-shell binding energies for the elements are listed in Appendix B.

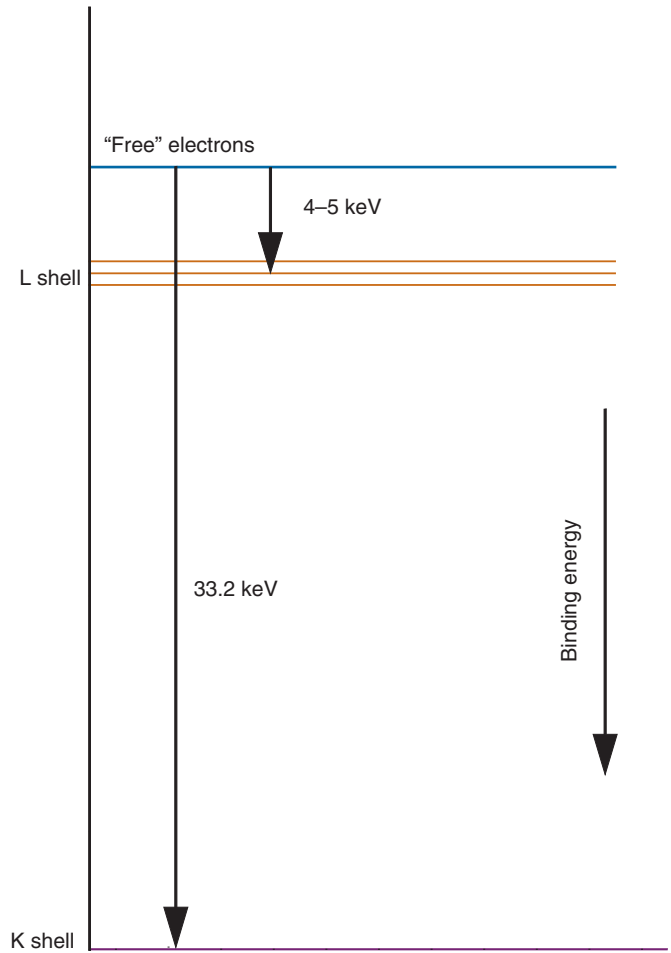
The energy required to move an electron from an inner to an outer shell is exactly equal to the difference in binding energies between the two shells. Thus the energy required to move an electron from the K shell to the L shell in an atom is  $K_B - L_B$  (with slight differences for different L shell substates).

\*Sometimes the notation  $K_{ab}$  also is used.

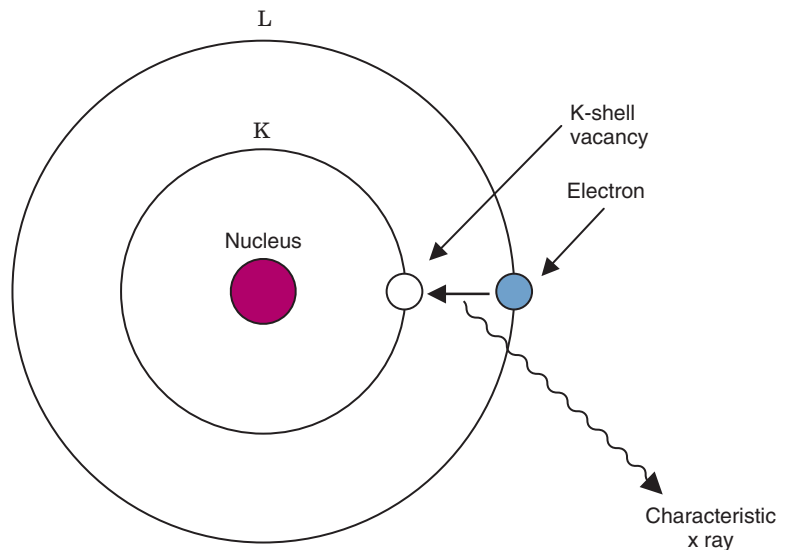
Binding energies and energy differences are sometimes displayed on an *energy-level diagram*. Figure 2-3 shows such a diagram for the K and L shells of the element iodine. The top line represents an electron completely separated from the parent atom (“unbound” or “free” electron). The bottom line represents the most tightly bound electrons, that is, the K shell. Above this are lines representing substates of the L shell. (The M shell and other outer shell lines are just above the L shell lines.) The distance from the K shell to the top level represents the K-shell binding energy for iodine (33.2 keV). To move a K-shell electron to the L shell requires approximately  $33 - 5 = 28$  keV of energy.

### 3. Atomic Emissions

When an electron is removed from one of the inner shells of an atom, an electron from an outer shell promptly moves in to fill the vacancy and energy is released in the process. The energy released when an electron drops from an outer to an inner shell is exactly equal to the difference in binding energies between the two shells. The energy may appear as a photon of electromagnetic radiation (Fig. 2-4). Electron binding energy differences have exact characteristic values for different elements; therefore the photon emissions are called *characteristic radiation* or *characteristic x rays*. The notation used to



**FIGURE 2-3** Electron energy-level diagram for an iodine atom. *Vertical axis* represents the energy required to remove orbital electrons from different shells (binding energy). Removing an electron from the atom, or going from an inner (e.g., K) to an outer (e.g., L) shell, requires an energy input, whereas an electron moving from an outer to an inner shell results in the emission of energy from the atom.



**FIGURE 2-4** Emission of characteristic x rays occurs when orbital electrons move from an outer shell to fill an inner-shell vacancy. ( $K_{\alpha}$  x-ray emission is illustrated.)

identify characteristic x rays from various electron transitions is summarized in Table 2-1. Note that some transitions are not allowed, owing to the selection rules of quantum mechanics.

As an alternative to characteristic x-ray emission, the atom may undergo a process known as the *Auger* (pronounced oh-zhay) effect. In the Auger effect, an electron from an outer shell again fills the vacancy, but the energy released in the process is transferred to another orbital electron. This electron then is emitted from the atom instead of

TABLE 2-1  
SOME NOTATION USED FOR  
CHARACTERISTIC X RAYS

Shell with Vacancy	Shell from Which Filled	Notation
K	L <sub>I</sub>	Not allowed
K	L <sub>II</sub>	K <sub>α2</sub>
K	L <sub>III</sub>	K <sub>α1</sub>
K	M <sub>I</sub>	Not allowed
K	M <sub>II</sub>	K <sub>β3</sub>
K	M <sub>III</sub>	K <sub>β1</sub>
K	N <sub>I</sub>	Not allowed
K	N <sub>II</sub> , N <sub>III</sub>	K <sub>β2</sub>
L <sub>II</sub>	M <sub>IV</sub>	L <sub>β1</sub>
L <sub>III</sub>	M <sub>IV</sub>	L <sub>α2</sub>
L <sub>III</sub>	M <sub>V</sub>	L <sub>α1</sub>

characteristic radiation. The process is shown schematically in Figure 2-5. The emitted electron is called an *Auger electron*.

The kinetic energy of an Auger electron is equal to the difference between the binding energy of the shell containing the original vacancy and the sum of the binding energies of the two shells having vacancies at the end. Thus the kinetic energy of the Auger electron emitted in Figure 2-5 is  $K_B - 2L_B$  (ignoring small differences in L-substate energies).

Two orbital vacancies exist after the Auger effect occurs. These are filled by electrons from the other outer shells, resulting in the emission of additional characteristic x rays or Auger electrons.

The number of vacancies that result in emission of characteristic x rays versus Auger electrons is determined by probability values that depend on the specific element and orbital shell involved. The probability that a vacancy will yield characteristic x rays is called the *fluorescent yield*, symbolized by  $\omega_K$  for the K shell,  $\omega_L$  for the L shell, and so forth. Figure 2-6 is a graph of  $\omega_K$  versus  $Z$ . Both characteristic x rays and Auger electrons are emitted by all elements, but heavy elements are more likely to emit x rays (large  $\omega$ ), whereas light elements are more likely to emit electrons (small  $\omega$ ).

The notation used to identify the shells involved in Auger electron emission is  $e_{abc}$ , in which  $a$  identifies the shell with the original vacancy,  $b$  the shell from which the electron dropped to fill the vacancy, and  $c$  the shell from which the Auger electron was emitted.

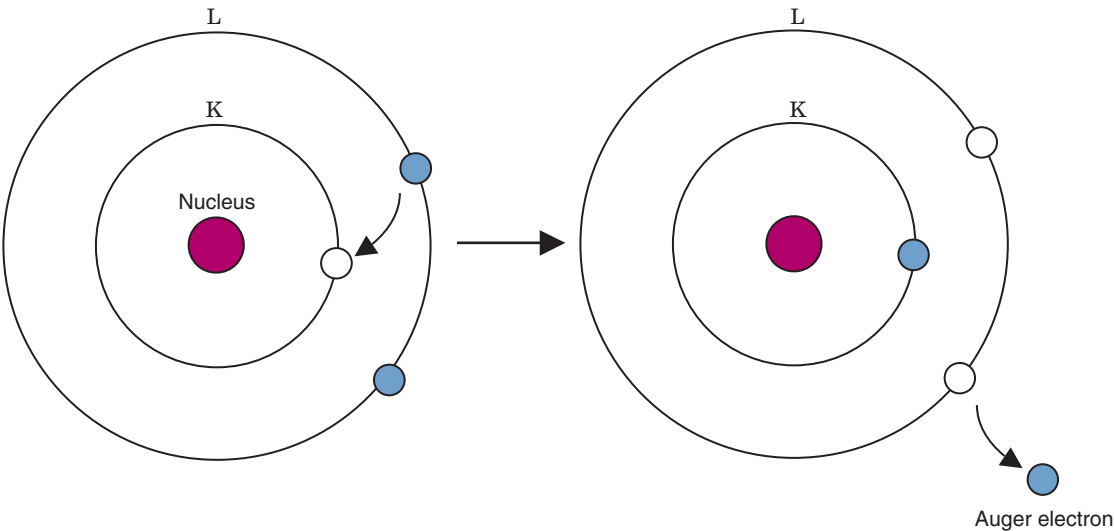
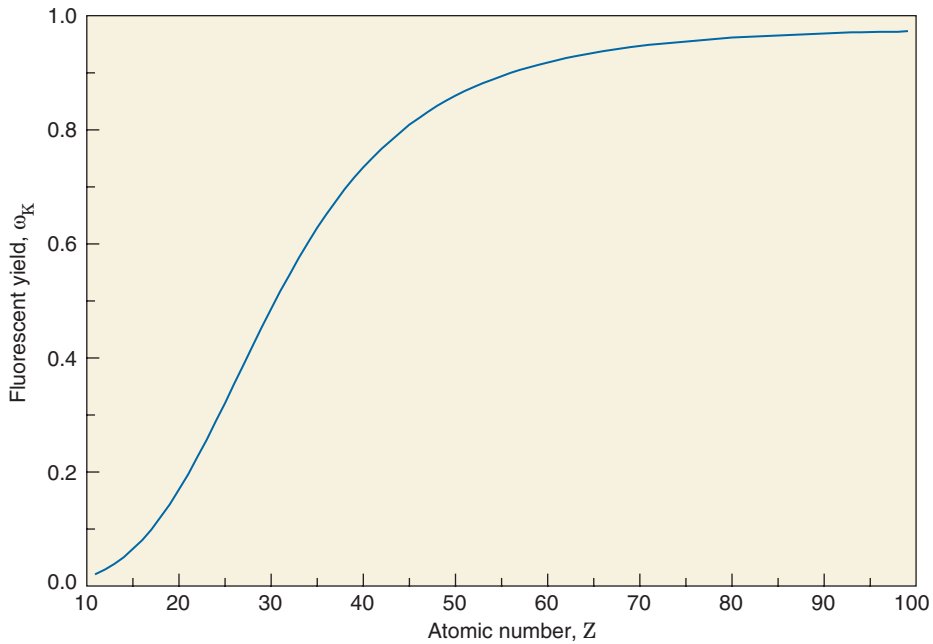


FIGURE 2-5 Emission of an Auger electron as an alternative to x-ray emission. No x ray is emitted.





**FIGURE 2-6** Fluorescent yield,  $\omega_K$ , or probability that an orbital electron shell vacancy will yield characteristic x rays rather than Auger electrons, versus atomic number  $Z$  of the atom. (Data from Hubbell JH, Trehan PN, Singh N, et al: *A review, bibliography, and tabulation of K, L, and higher atomic shell x-ray fluorescence yields*. J Phys Chem Ref Data 23:339-364, 1994.)

Thus the electron emitted in Figure 2-5 is a *KLL* Auger electron, symbolized by  $e_{KLL}$ . In the notation  $e_{Kxx}$ , the symbol  $x$  is inclusive, referring to all Auger electrons produced from initial K-shell vacancies.

## D. THE NUCLEUS

### 1. Composition

The atomic nucleus is composed of *protons* and *neutrons*. Collectively, these particles are known as *nucleons*. The properties of nucleons and electrons are summarized in Table 2-2.

**TABLE 2-2**  
**BASIC PROPERTIES OF NUCLEONS AND ELECTRONS<sup>\*</sup>**

Particle	Charge <sup>*</sup>	Mass	
		u	MeV
Proton	+1	1.007276	938.272
Neutron	0	1.008665	939.565
Electron	-1	0.000549	0.511

<sup>\*</sup>One unit of charge is equivalent to  $1.602 \times 10^{-19}$  coulombs.

Nucleons are much more massive than electrons (by nearly a factor of 2000). Conversely, nuclear diameters are very small in comparison with atomic diameters ( $10^{-13}$  vs.  $10^{-8}$  cm). Thus it can be deduced that the density of nuclear matter is very high ( $\sim 10^{14}$  g/cm<sup>3</sup>) and that the rest of the atom (electron cloud) is mostly empty space.

### 2. Terminology and Notation

An atomic nucleus is characterized by the number of neutrons and protons it contains. The number of protons determines the *atomic number* of the atom,  $Z$ . As mentioned earlier, this also determines the number of orbital electrons in the electrically neutral atom and therefore the *chemical element* to which the atom belongs.

The total number of nucleons is the *mass number* of the nucleus,  $A$ . The difference,  $A - Z$ , is the *neutron number*,  $N$ . The mass number  $A$  is approximately equal to, but not the same as, the *atomic weight* (AW) used in chemistry. The latter is the average weight of an atom of an element in its natural abundance (see Appendix B).

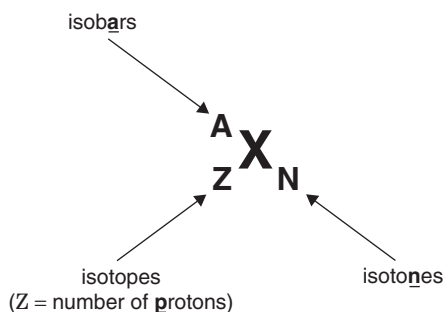
The notation now used to summarize atomic and nuclear composition is  ${}^A_ZX_N$ , in which  $X$  represents the chemical element to which the atom belongs. For example, an

atom composed of 53 protons, 78 neutrons (and thus 131 nucleons), and 53 orbital electrons represents the element iodine and is symbolized by  $^{131}_{53}\text{I}_{78}$ . Because all iodine atoms have atomic number 53, the “I” and the “53” are redundant and the “53” can be omitted. The neutron number, 78, can be inferred from the difference,  $131 - 53$ , so this also can be omitted. Therefore a shortened but still complete notation for this atom is  $^{131}\text{I}$ . An acceptable alternative in terms of medical terminology is I-131. Obsolete forms (sometimes found in older texts) include  $\text{I}^{131}$ ,  $^{131}\text{I}$ , and  $\text{I}_{131}$ .

### 3. Nuclear Families

Nuclear species sometimes are grouped into families having certain common characteristics. A *nuclide* is characterized by an exact nuclear composition, including the mass number  $A$ , atomic number  $Z$ , and arrangement of nucleons within the nucleus. To be classified as a nuclide, the species must have a “measurably long” existence, which for current technology means a lifetime greater than about  $10^{-12}$  sec. For example,  $^{12}\text{C}$ ,  $^{16}\text{O}$ , and  $^{131}\text{I}$  are nuclides.

Figure 2-7 summarizes the notation used for identifying a particular nuclear species, as well as the terminology used for nuclear families. Nuclides that have the same atomic number  $Z$  are called *isotopes*. Thus  $^{125}\text{I}$ ,  $^{127}\text{I}$ , and  $^{131}\text{I}$  are isotopes of the element iodine. Nuclides with the same mass number  $A$  are *isobars* (e.g.,  $^{131}\text{I}$ ,  $^{131}\text{Xe}$ , and  $^{131}\text{Cs}$ ). Nuclides with the same neutron number  $N$  are *isotones* (e.g.,  $^{131}_{53}\text{I}_{78}$ ,  $^{132}_{54}\text{Xe}_{78}$ , and  $^{133}_{55}\text{Cs}_{78}$ ). A mnemonic device for remembering these relationships is that *isotopes* have the same number of protons, *isotones* the same number of neutrons, and *isobars* the same mass number ( $A$ ).



**FIGURE 2-7** Notation and terminology for nuclear families.

### 4. Forces and Energy Levels within the Nucleus

Nucleons within the nucleus are subject to two kinds of forces. Repulsive *coulombic* or *electrical forces* exist between positively charged protons. These are counteracted by very strong forces of attraction, called *nuclear forces* (sometimes also called *exchange forces*), between any two nucleons. Nuclear forces are effective only over very short distances, and their effects are seen only when nucleons are very close together, as they are in the nucleus. Nuclear forces hold the nucleus together against the repulsive coulombic forces between protons.

Nucleons move about within the nucleus in a very complicated way under the influence of these forces. One model of the nucleus, called the *shell model*, portrays the nucleons as moving in “orbits” about one another in a manner similar to that of orbital electrons moving about the nucleus in the Bohr atom. Only a limited number of motions are allowed, and these are determined by a set of nuclear quantum numbers.

The most stable arrangement of nucleons is called the *ground state*. Other arrangements of the nucleons fall into the following two categories:

1. *Excited states* are arrangements that are so unstable that they have only a transient existence before transforming into some other state.
2. *Metastable states* also are unstable, but they have relatively long lifetimes before transforming into another state. These also are called *isomeric states*.

The dividing line for lifetimes between excited and metastable states is approximately  $10^{-12}$  sec. This is not a long time according to everyday standards, but it is “relatively long” by nuclear standards. (The prefix *meta* derives from the Greek word for “almost.”) Some metastable states are quite long-lived; that is, they have average lifetimes of several hours. Because of this, metastable states are considered to have separate identities and are themselves classified as nuclides. Two nuclides that differ from one another in that one is a metastable state of the other are called *isomers*.

In nuclear notation, excited states are identified by an asterisk ( $^{\text{A}}\text{X}^*$ ) and metastable states by the letter m ( $^{\text{A}}\text{X}^{\text{m}}$  or  $\text{X-}\text{Am}$ ).<sup>†</sup> Thus

<sup>†</sup>The notation  $^{\text{A}}\text{X}^{\text{m}}$  is sometimes used in Europe (e.g.,  $^{99\text{m}}\text{Tc}^{\text{m}}$ ).

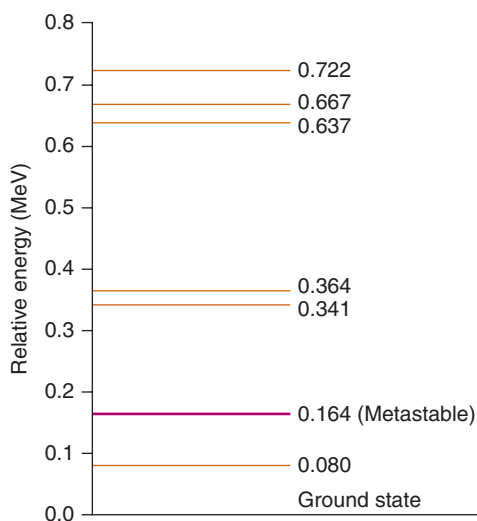


$^{99\text{m}}\text{Tc}$  (or Tc-99m) represents a metastable state of  $^{99}\text{Tc}$ , and  $^{99\text{m}}\text{Tc}$  and  $^{99}\text{Tc}$  are isomers.

Nuclear transitions between different nucleon arrangements involve discrete and exact amounts of energy, as do the rearrangements of orbital electrons in the Bohr atom. A *nuclear energy-level diagram* is used to identify the various excited and metastable states of a nuclide and the energy relationships among them. Figure 2-8 shows a partial diagram for  $^{131}\text{Xe}$ .<sup>\*</sup> The bottom line represents the ground state, and other lines represent excited or metastable states. Metastable states usually are indicated by somewhat heavier lines. The vertical distances between lines are proportional to the energy differences between levels. A transition from a lower to a higher state requires an energy input of some sort, such as a photon or particle striking the nucleus. Transitions from higher to lower states result in the release of energy, which is given to emitted particles or photons.

## 5. Nuclear Emissions

Nuclear transformations can result in the emission of particles (primarily electrons or  $\alpha$  particles) or photons of electromagnetic radiation. This is discussed in detail in Chapter 3.



**FIGURE 2-8** Partial nuclear energy-level diagram for the  $^{131}\text{Xe}$  nucleus. The *vertical axis* represents energy differences between nuclear states (or “arrangements” of nucleons). Going up the scale requires energy input. Coming down the scale results in the emission of nuclear energy. Heavier lines indicate metastable states.

<sup>\*</sup>Actually, these are the excited and metastable states formed during radioactive decay by  $\beta^-$  emission of  $^{131}\text{I}$  (see Chapter 3, Section D, and Appendix C).

Photons of nuclear origin are called  $\gamma$  rays (*gamma rays*). The energy difference between the states involved in the nuclear transition determines the  $\gamma$ -ray energy. For example, in Figure 2-8 a transition from the level marked 0.364 MeV to the ground state would produce a 0.364-MeV  $\gamma$  ray. A transition from the 0.364-MeV level to the 0.080-MeV level would produce a 0.284-MeV  $\gamma$  ray.

As an alternative to emitting a  $\gamma$  ray, the nucleus may transfer the energy to an orbital electron and emit the electron instead of a photon. This process, which is similar to the Auger effect in x-ray emission (see Section C.3, earlier in this chapter), is called *internal conversion*. It is discussed in detail in Chapter 3, Section E.

## 6. Nuclear Binding Energy

When the mass of an atom is compared with the sum of the masses of its individual components (protons, neutrons, and electrons), it always is found to be less by some amount,  $\Delta m$ . This mass deficiency, expressed in energy units, is called the *binding energy*  $E_B$  of the atom:

$$E_B = \Delta mc^2 \quad (2-5)$$

For example, consider an atom of  $^{12}\text{C}$ . This atom is composed of six protons, six electrons, and six neutrons, and its mass is precisely 12 u (by definition of the universal mass unit u). The sum of the masses of its components is

electrons	$6 \times 0.000549 \text{ u} = 0.003294 \text{ u}$
protons	$6 \times 1.007276 \text{ u} = 6.043656 \text{ u}$
neutrons	$6 \times 1.008665 \text{ u} = 6.051990 \text{ u}$
total	12.098940 u

Thus  $\Delta m = 0.098940 \text{ u}$ . Because  $1 \text{ u} = 931.5 \text{ MeV}$ , the binding energy of a  $^{12}\text{C}$  atom is  $0.098940 \times 931.5 \text{ MeV} = 92.16 \text{ MeV}$ .

The binding energy is the minimum amount of energy required to overcome the forces holding the atom together to separate it completely into its individual components. Some of this represents the binding energy of orbital electrons, that is, the energy required to strip the orbital electrons away from the nucleus; however, comparison of the total binding energy of a  $^{12}\text{C}$  atom with the K-shell binding energy of carbon (see Appendix B) indicates that most of this energy is *nuclear binding energy*, that is, the energy required to separate the nucleons.

Nuclear processes that result in the release of energy (e.g.,  $\gamma$ -ray emission) always *increase*

the binding energy of the nucleus. Thus a nucleus emitting a 1-MeV  $\gamma$  ray would be found to weigh *less* (by the mass equivalent of 1 MeV) after the  $\gamma$  ray was emitted than before. In essence, mass is converted to energy in the process.

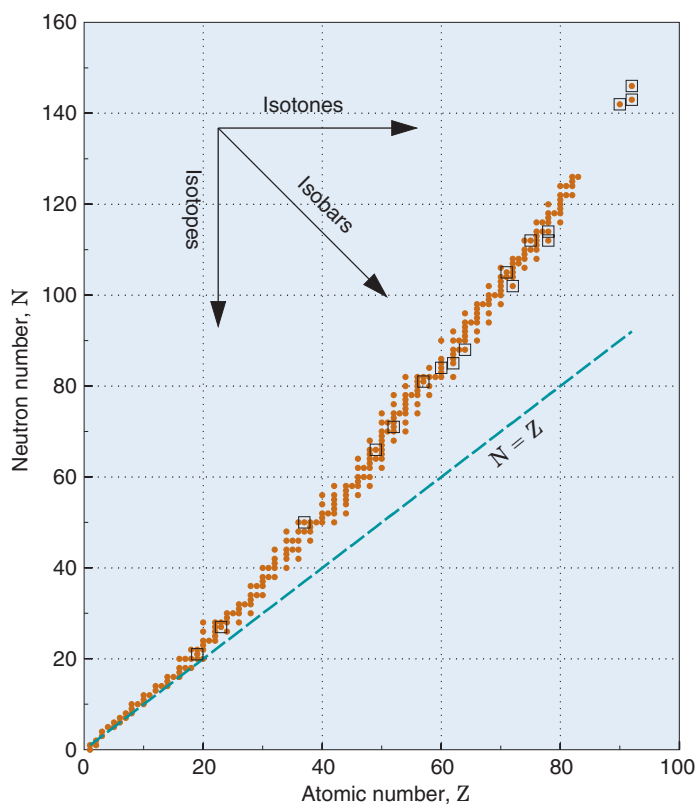
## 7. Characteristics of Stable Nuclei

Not all combinations of protons and neutrons produce stable nuclei. Some are unstable, even in their ground states. An unstable nucleus emits particles or photons to transform itself into a more stable nucleus. This is the process of *radioactive disintegration* or *radioactive decay*, discussed in Chapter 3. A survey of the general characteristics of naturally occurring *stable nuclides* provides clues to the factors that contribute to nuclear instability and thus to radioactive decay.

Figure 2-9 is a plot of the nuclides found in nature, according to their neutron and proton numbers. For example, the nuclide  $^{12}_6\text{C}$  is represented by a dot at the point  $Z = 6$ ,  $N = 6$ . Most of the naturally occurring nuclides are stable; however, 17 very long-lived but unstable (radioactive) nuclides that still are present from the creation of the elements also are shown.

A first observation is that there are favored neutron-to-proton ratios among the naturally occurring nuclides. They are clustered around an imaginary line called the *line of stability*. For light elements, the line corresponds to  $N \approx Z$ , that is, approximately equal numbers of protons and neutrons. For heavy elements, it corresponds to  $N \approx 1.5 Z$ , that is, approximately 50% more neutrons than protons. The line of stability ends at  $^{209}\text{Bi}$  ( $Z = 83$ ,  $N = 126$ ). All heavier nuclides are unstable.

In general, there is a tendency toward instability in atomic systems composed of large numbers of identical particles confined in a small volume. This explains the instability of very heavy nuclei. It also explains why, for light elements, stability is favored by more or less equal numbers of neutrons and protons rather than grossly unequal numbers. A moderate excess of neutrons is favored among heavier elements because neutrons provide only exchange forces (attraction), whereas protons provide both exchange forces and coulombic forces (repulsion). Exchange forces are effective over very short distances and thus affect only “close neighbors” in the nucleus, whereas the repulsive coulombic forces are effective over much greater distances. Thus



an excess of neutrons is required in heavy nuclei to overcome the long-range repulsive coulombic forces between a large number of protons.

Nuclides that are not close to the line of stability are likely to be unstable. Unstable nuclides lying above the line of stability are said to be “proton deficient,” whereas those lying below the line are “neutron deficient.” Unstable nuclides generally undergo radioactive decay processes that transform them into nuclides lying closer to the line of stability, as discussed in Chapter 3.

Figure 2-9 demonstrates that there often are many stable isotopes of an element. Isotopes fall on vertical lines in the diagram. For example, there are ten stable isotopes of tin (Sn,  $Z = 50$ )\*. There may also be several stable isotones. These fall along horizontal lines. In relatively few cases, however, is there more than one stable isobar (isobars fall along descending 45-degree lines on the graph).

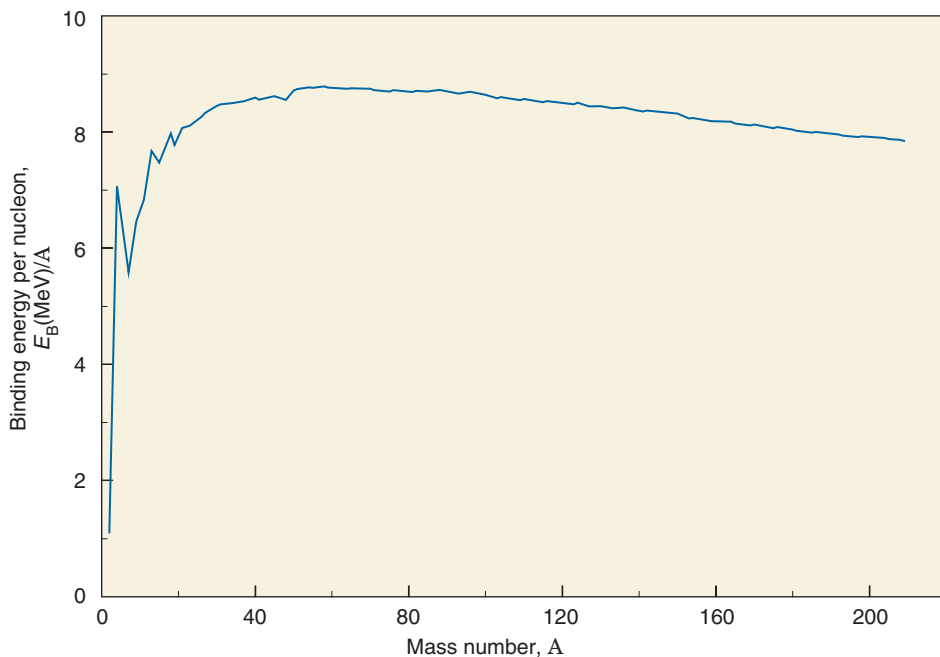
\*Although most element symbols are simply one- or two-letter abbreviations of their (English) names, ten symbols derive from Latin or Greek names of metals known for more than 2 millennia: antimony (stibium, Sb), copper (cuprum, Cu), gold (aurum, Au), iron (ferrum, Fe), lead (plumbum, Pb), mercury (hydrargyrum, Hg), potassium (kalium, K), silver (argentum, Ag), sodium (natrium, Na), and tin (stannum, Sn). The symbol for tungsten, W, derives from the German “wolfram,” the name it was first given in medieval times.

This reflects the existence of several modes of “isobaric” radioactive decay that permit nuclides to transform along isobaric lines until the most stable isobar is reached. This is discussed in detail in Chapter 3.

One also notes among the stable nuclides a tendency to favor even numbers. For example, there are 165 stable nuclides with both even numbers of protons and even numbers of neutrons. Examples are  ${}^4_2\text{He}$  and  ${}^{12}_6\text{C}$ . There are 109 “even-odd” stable nuclides, with even numbers of protons and odd numbers of neutrons or vice versa. Examples are  ${}^9_4\text{Be}$  and  ${}^{11}_5\text{B}$ . However, there are only four stable “odd-odd” nuclides:  ${}^2_1\text{H}$ ,  ${}^6_3\text{Li}$ ,  ${}^{10}_5\text{B}$ , and  ${}^{14}_7\text{N}$ . The stability of even numbers reflects the tendency of nuclei to achieve stable arrangements by the “pairing up” of nucleons in the nucleus.

Another measure of relative nuclear stability is nuclear binding energy, because this represents the amount of energy required to break the nucleus up into its separate components. Obviously, the greater the number of nucleons, the greater the total binding energy. Therefore a more meaningful parameter is the *binding energy per nucleon*,  $E_B/A$ . Higher values of  $E_B/A$  are indicators of greater nuclear stability.

Figure 2-10 is a graph of  $E_B/A$  versus  $A$  for the stable nuclides. Binding energy is greatest ( $\approx 8$  MeV per nucleon) for nuclides of mass number  $A \approx 60$ . It decreases slowly with increasing  $A$ , indicating the tendency toward



**FIGURE 2-10** Binding energy per nucleon ( $E_B/A$ ) versus mass number ( $A$ ) for the stable nuclides.

instability for very heavy nuclides. Finally, there are a few peaks in the curve representing very stable light nuclides, including  ${}^4_2\text{He}$ ,  ${}^{12}_6\text{C}$ , and  ${}^{16}_8\text{O}$ . Note that these are all even-even nuclides.

## REFERENCES

1. National Institute of Standards and Technology (NIST): Fundamental Physics Constants. Available at <http://physics.nist.gov/cuu/Constants/index.html> [accessed July 4, 2011].

## BIBLIOGRAPHY

Fundamental quantities of physics and mathematics, as well as constants and conversion factors, can be found in reference 1.

**Recommended texts for in-depth discussions of topics in atomic and nuclear physics are the following:**

- Evans RD: *The Atomic Nucleus*, New York, 1972, McGraw-Hill.  
 Jelley NA: *Fundamentals of Nuclear Physics*, New York, 1990, Cambridge University Press.  
 Yang F, Hamilton JH: *Modern Atomic and Nuclear Physics*, New York, 1996, McGraw-Hill.