CHAPTER 1

1 Basic nuclear medicine physics

Properties and Structure of Matter

Matter has several fundamental properties. For our purposes the most important are **mass** and **charge** (**electric**). We recognize mass by the force gravity exerts on a material object (commonly referred to as its weight) and by the object's inertia, which is the "resistance" we encounter when we attempt to change the position or motion of a material object.

Similarly, we can, at least at times, recognize charge by the direct effect it can have on us or that we can observe it to have on inanimate objects. For example, we may feel the presence of a strongly charged object when it causes our hair to move or even to stand on end. More often than not, however, we are insensitive to charge. But whether grossly detectable or not, its effects must be considered here because of the role charge plays in the structure of matter.

Charge is generally thought to have been recognized first by the ancient Greeks. They noticed that some kinds of matter, an amber rod for example, can be given an electric charge by rubbing it with a piece of cloth. Their experiments convinced them that there are two kinds of charge: opposite charges, which attract each other, and like charges, which repel. One kind of charge came to be called positive, the other negative. We now know that the negative charge is associated with electrons. The rubbing transferred some of the electrons from the atoms of the matter in the rod to the cloth. In a similar fashion, electrons can be transferred to the shoes of a person walking across a carpet. The carpet will then have a net positive charge and the shoes (and wearer) a net negative charge (Fig. 1-1). With these basic properties in mind, we can look at matter in more detail.

Matter is composed of **molecules**. In any chemically pure material, the molecules are the smallest units that retain the characteristics of the material itself. For example, if a block of salt were to be broken into successively smaller pieces,

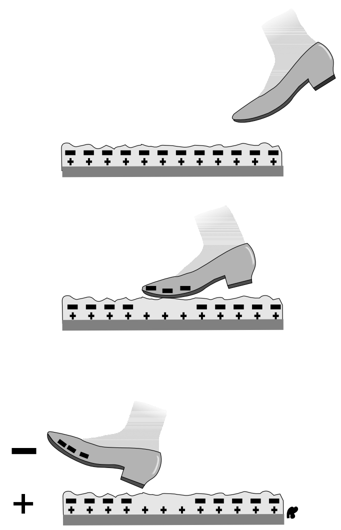

Figure 1-1 Electrostatic charge.

the smallest fragment with the properties of salt would be a single salt molecule (Fig. 1-2). With further fragmentation the molecule would no longer be salt. Molecules, in turn, are composed of **atoms**. Most molecules consist of more than one kind of atom—salt, for example, is made up of atoms of chlorine and sodium. The atoms themselves are composed of smaller particles, the **subatomic particles**, which are discussed later.

The molecule is held together by the chemical bonds among its atoms. These bonds are formed by the force of electrical attraction between oppositely charged parts of the molecule. This force is often referred to as the Coulomb force after Charles A. de Coulomb, the physicist who characterized it. This is the force involved in chemical reactions such as the combining of hydrogen and oxygen to form water. The electrons of the atom are held by the electrical force between them and the positive nucleus. The nucleus of the atom is held together by another type of force—nuclear force—which is involved in the release of atomic energy. Nuclear forces are of greater magnitudes than electrical forces.

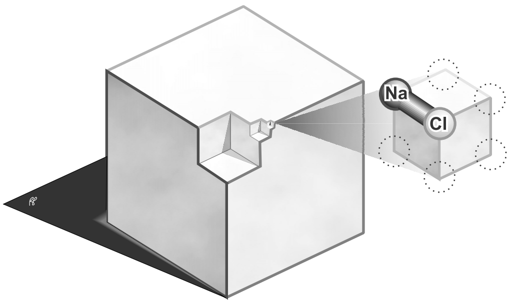

Figure 1-2 The NaCl molecule is the smallest unit of salt that retains the characteristics of salt.

Elements

There are more than 100 species of atoms. These species are referred to as **elements**. Most of the known elements—for example, mercury, helium, gold, hydrogen, and oxygen—occur naturally on earth; others are not usually found in nature but are made by humans—for example, europium and americium. A reasonable explanation for the absence of some elements from nature is that if and when they were formed they proved too unstable to survive in detectable amounts into the present.

All the elements have been assigned symbols or abbreviated chemical names: gold—Au, mercury—Hg, helium—He. Some symbols are obvious abbreviations of the English name; others are derived from the original Latin name of the element, for example, Au is from aurum, the Latin word for gold.

All of the known elements, both natural and those made by humans, are organized in the **periodic table**. In Figure 1-3, the elements that have a

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
			72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	
			104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt	110 ?	111 ?							

Figure 1-3 Periodic table.

stable state are shown in white boxes; those that occur only in a radioactive form are shown in gray boxes. Elements 104 to 111 have not been formally named (proposed names are listed). When necessary, the chemical symbol shown in the table for each element can be expanded to include three numbers to describe the composition of its nucleus (Fig. 1-4).

Atomic Structure

Atoms initially were thought of as no more than small pieces of matter. Our understanding that they have an inner structure has it roots in the observations of earlier physicists that the atoms of which matter is composed contain **electrons** of negative charge. In as much as the atom as a whole is electrically neutral, it seemed obvious that it must also contain something with a positive charge to balance the negative charge of the electrons. Thus, early attempts to picture the atom, modeled on our solar system, showed the negatively charged electrons orbiting a central group of particles, the positively charged **nucleus** (Fig. 1-5).

Electrons

In our simple solar-system model of the atom, the electrons are viewed as orbiting the nucleus at high speeds. They have a negative charge and are drawn toward the positively charged nucleus. The electrical charges of the atom are "balanced," that is, the total negative charge of the electrons equals the positive charge of the nucleus. As we shall see in a moment, this is simply another way to point out that the number of orbital electrons equals the number of nuclear protons.

Although each electron orbits at high speed, it remains in its orbit because the electrical force draws it toward the positively charged nucleus. This attraction keeps the moving electron in its orbit in much the same way as a string tied to a ball will hold it in its path as you swing it rapidly around your head (Fig. 1-6).

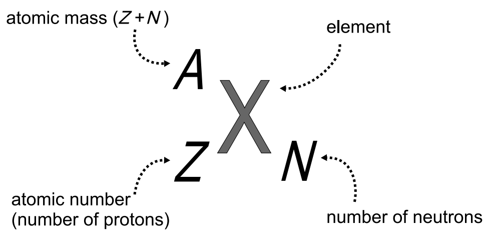

Figure 1-4 Standard atomic notation.

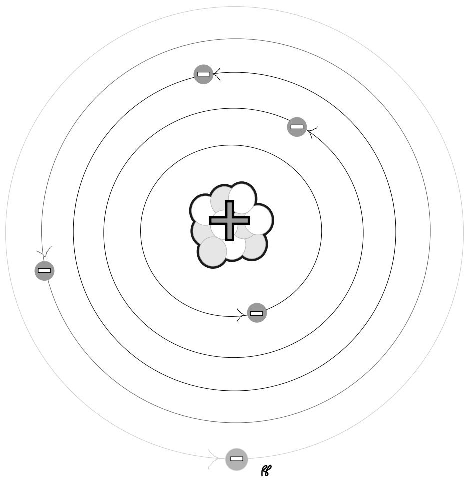

Figure 1-5 Flat atom. The standard two-dimensional drawing of atomic structure.

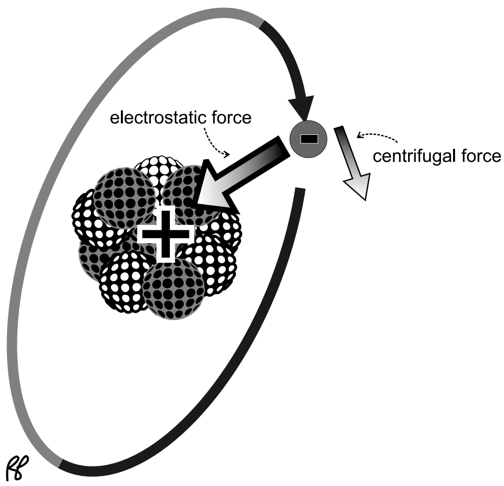

Figure 1-6 The Coulomb force between the negative electrons and the positive protons keeps the electron in orbit. Without this electric force the electron would fly off into space.

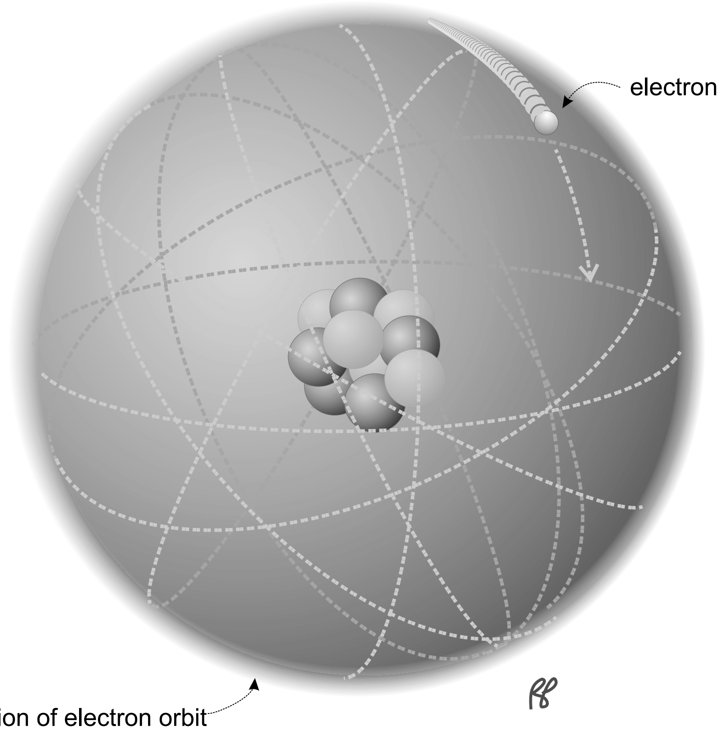

Figure 1-7 An electron shell is a representation of the energy level in which the electron moves.

Electron Shells

By adding a third dimension to our model of the atom, we can depict the electron orbits as the surfaces of spheres (called **shells**) to suggest that, unlike the planets orbiting the sun, electrons are not confined to a circular orbit lying in a single plane but may be more widely distributed (Fig. 1-7). Of course, neither the simple circular orbits nor these electron shells are physical entities; rather, they are loose representations of the "distances" the orbital electrons are from the nucleus (Fig. 1-8). Although it is convenient for us to talk about distances and diameters of the shells, distance on the atomic scale does not have quite the same meaning it does with everyday objects. The more significant characteristic of a shell is the energy it signifies.

The closer an electron is to the nucleus, the more tightly it is held by the positive charge of nucleus. In saying this, we mean that more work (energy) is required to remove an inner-shell electron than an outer one. The energy that must be put into the atom to separate an electron is called the **electron binding energy**. It is usually expressed in **electron volts** (**eV**). The electron binding energy varies from a few thousand electron volts (keV) for inner-shell electrons to just a few eV for the less tightly bound outer-shell electrons.

ELECTRON VOLT

The electron volt is a special unit defined as the energy required to move one electron against a potential difference of one volt. It is a small unit on the everyday scale, at only 1.6×10^{-19} joules (J), but a very convenient unit on the atomic scale. One joule is the Système International (SI) unit of work or energy. For comparison, 1 J equals 0.24 small calories (as opposed to the kcal used to measure food intake).

Figure 1-8 Cut-away model of a medium-sized atom such as argon.

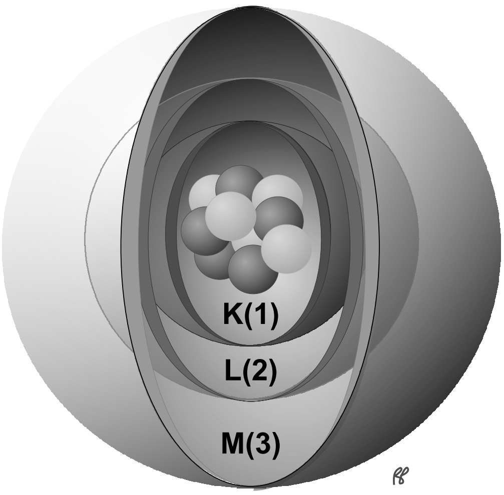

Figure 1-9 K, L, and M electron shells.

Quantum Number

The atomic electrons in their shells are usually described by their **quantum numbers**, of which there are four types. The first is the **principal** quantum number (**n**), which identifies the energy shell. The first three shells (K, L, and M) are depicted in Figure 1-9. The electron binding energy is greatest for the innermost shell (K) and is progressively less for the outer shells. Larger atoms have more shells.

The second quantum number is the **azimuthal** quantum number (**l**), which can be thought of as a **subshell** within the shell. Technically **l** is the angular momentum of the electron and is related to the product of the mass of the electron, its velocity, and the radius of its orbit. Each subshell is assigned a letter designation: s, p, d, f, and so on. For completeness, the full label of a subshell includes the numeric designation of its principal shell, which for L is the number 2; thus 2s and 2p.

The third number, the **magnetic** quantum number ($\mathbf{m_l}$), describes the direction of rotation of the electron and the orientation of the subshell orbit. The fourth quantum number is the **spin** quantum number ($\mathbf{m_s}$), which refers to the direction the electron spins on its axis. Both the third and fourth quantum numbers contribute to the **magnetic moment** (or magnetic field) created by the moving electron. The four quantum numbers are outlined in Table 1-1.

QUANTUM NUMBERS

The term quantum means, literally, amount. It acquired its special significance in physics when Bohr and others theorized that physical quantities such as energy and light could not have a range of values as on a continuum, but rather could have only discrete, step-like values. The individual steps are so small that their existence escaped the notice of physicists until Bohr postulated them to explain his theory of the atom. We now refer to Bohr's theory as **quantum theory** and the resulting explanations of motion in the atomic scale as **quantum mechanics** to distinguish it from the classical mechanics described by Isaac Newton, which is still needed for everyday engineering.

The innermost or K shell has only one subshell (called the s subshell). This subshell has a magnetic quantum number of zero and two possible values for the spin quantum number, $\mathbf{m_s}$; these are $+\frac{1}{2}$ and $-\frac{1}{2}$. The neutral atom with a full K shell, that is to say, with two electrons "circling" the nucleus, is the helium atom.

Table 1-1 Quantum Numbers and Values

Quantum Number	Corresponding Names	Range of Values
Principal (**n**)	K, L, M, ...	1, 2, 3, ...
Azimuthal (**l**)	s, p, d, f, g, ...	0, 1, 2, 3, ... (**n** − **1**)
Magnetic ($\mathbf{m}_l$)	None	−**l**, −(**l** − 1), ... 0, ... (**l** − 1), **l**
Spin ($\mathbf{m}_s$)	Down, up	$-\frac{1}{2}, +\frac{1}{2}$

The next shell, the L shell, has available an s subshell and a second subshell (called the 2p subshell). The 2s subshell in the L shell is similar to the s subshell of the K shell and can accommodate two electrons (Fig. 1-10A). The 2p subshell has three possible magnetic quantum numbers (−1, 0, and 1) or subshells, and for each of these quantum numbers there are the two available spin quantum numbers, which allows for a total of six electrons. Each 2p subshell

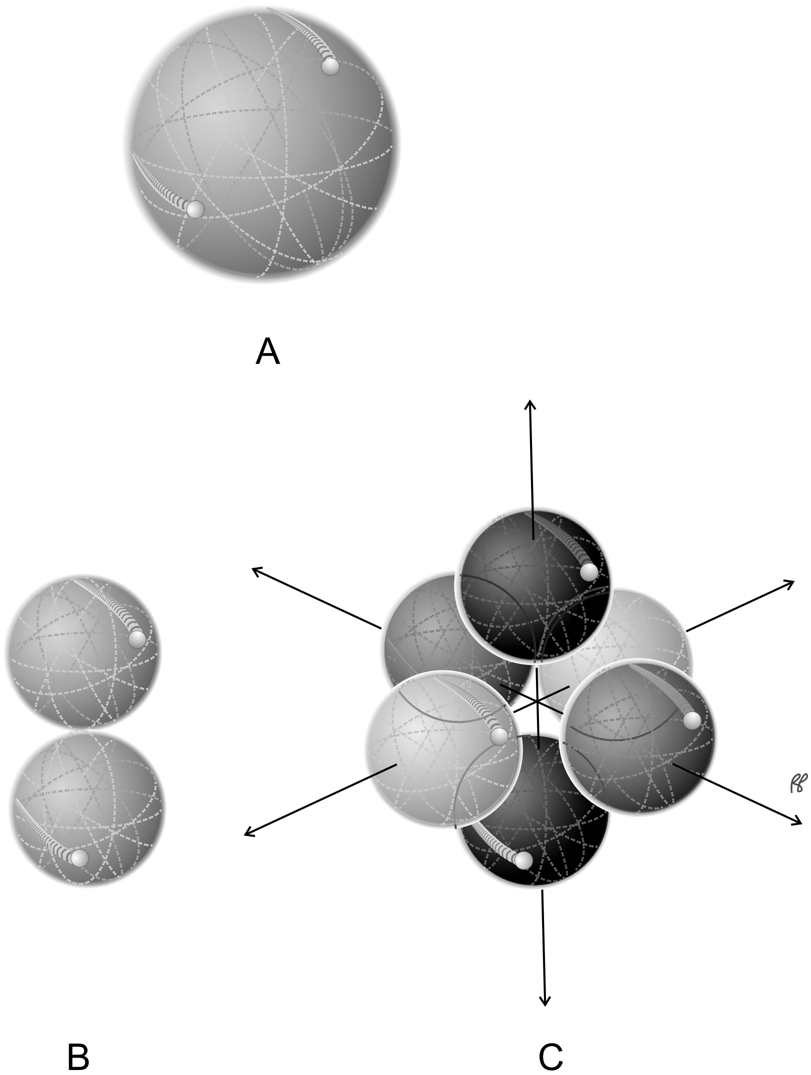

Figure 1-10 Subshells of the L shell.

Table 1-2 Electron Quantum States

Quantum Number Designations		**Quantum States**									
Principal or radial (n)		**1**		**2**							
Azimuthal (l)		**0**		**0**		**1**					
Magnetic (m_l)		**0**		**0**		**−1**		**0**		**1**	
Spin (m_s)		$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$
Element											
Atomic Number	**Chemical Name**	**Number of Electrons in Each State**									
1	Hydrogen	1	0	0	0	0	0	0	0	0	0
2	Helium	1	1	0	0	0	0	0	0	0	0
3	Lithium	1	1	1	0	0	0	0	0	0	0
4	Beryllium	1	1	1	1	0	0	0	0	0	0
5	Boron	1	1	1	1	1	0	0	0	0	0
6	Carbon	1	1	1	1	1	1	0	0	0	0
7	Nitrogen	1	1	1	1	1	1	1	0	0	0
8	Oxygen	1	1	1	1	1	1	1	1	0	0
9	Flourine	1	1	1	1	1	1	1	1	1	0
10	Neon	1	1	1	1	1	1	1	1	1	1

is depicted as two adjacent spheres, a kind of three-dimensional figure eight (Fig. 1-10B). The arrangement of all three subshells is shown in Figure 1-10C. The L shell can accommodate a total of eight electrons. The neutral atom containing all ten electrons in the K and L shells is neon.

The number of electrons in each set of shells for the light elements, hydrogen through neon, forms a regular progression, as shown in Table 1-2. For the third and subsequent shells, the ordering and filling of the subshells, as dictated by the rules of quantum mechanics, is less regular and will not be covered here.

Stable Electron Configuration

Just as it takes energy to remove an electron from its atom, it takes energy to move an electron from an inner shell to an outer shell, which can also be thought of as the energy required to pull a negative electron away from the positively charged nucleus. Any vacancy in an inner shell creates an unstable condition often referred to as an **excited state**.

The electrical charges of the atom are balanced, that is, the total negative charge of the electrons equals the total positive charge of the nucleus. This is simply another way of pointing out that the number of orbital electrons equals the number of nuclear protons. Furthermore, the electrons must fill the shells with the highest binding energy first. At least in the elements of low atomic number, electrons in the inner shells have the highest binding energy.

If the arrangement of the electrons in the shells is not in the stable state, they will undergo rearrangement in order to become stable, a process often referred to as **de-excitation**. Because the stable configuration of the shells always has less energy than any unstable configuration, the de-excitation releases energy as photons, often as **x-rays**.

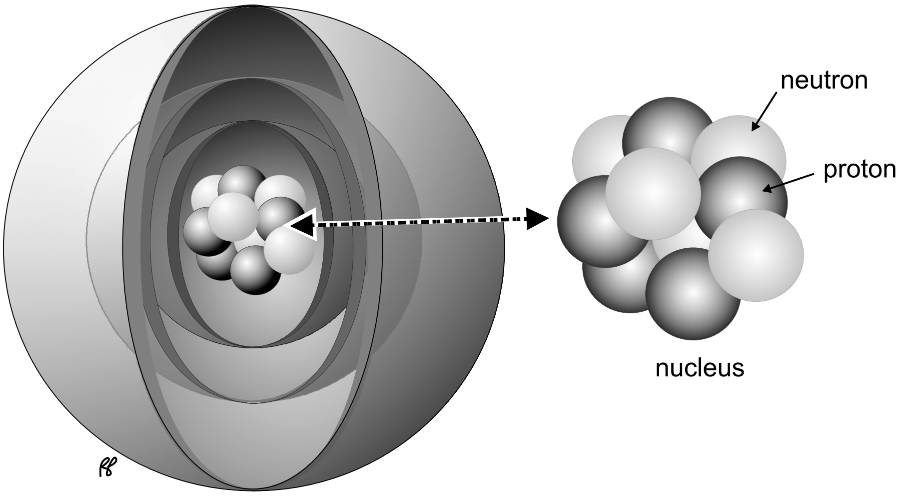

Figure 1-11 The nucleus of an atom is composed of protons and neutrons.

Table 1-3 The Subatomic Particles

Name	Symbol	Location	Mass[a]	Charge
Neutron	N	Nucleus	1840	None
Proton	P	Nucleus	1836	Positive (+)
Electron	e^-	Shell	1	Negative (−)

[a] Relative to an electron.

Nucleus

Like the atom itself, the atomic nucleus also has an inner structure (Fig. 1-11). Experiments showed that the nucleus consists of two types of particles: **protons**, which carry a positive charge, and **neutrons**, which carry no charge. The general term for protons and neutrons is **nucleons**. The nucleons, as shown in Table 1-3, have a much greater mass than electrons. Like electrons, nucleons have quantum properties including spin. The nucleus has a spin value equal to the sum of the nucleon spin values.

A simple but useful model of the nucleus is a tightly bound cluster of protons and neutrons. Protons naturally repel each other since they are positively charged; however, there is a powerful binding force called the **nuclear force** that holds the nucleons together very tightly

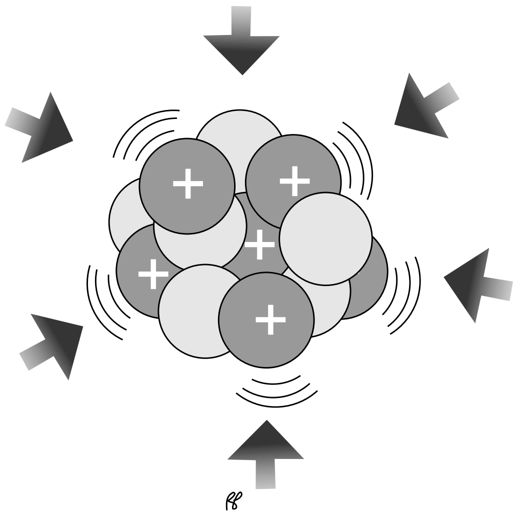

Figure 1-12 Nuclear binding force is strong enough to overcome the electrical repulsion between the positively charged protons.

(Fig. 1-12). The work (energy) required to overcome the nuclear force, the work to remove a nucleon from the nucleus, is called the **nuclear binding energy**. Typical binding energies are in the range of 6 million to 9 million electron volts (MeV) (approximately one thousand to one million times the electron binding force). The magnitude of the binding energy is related to

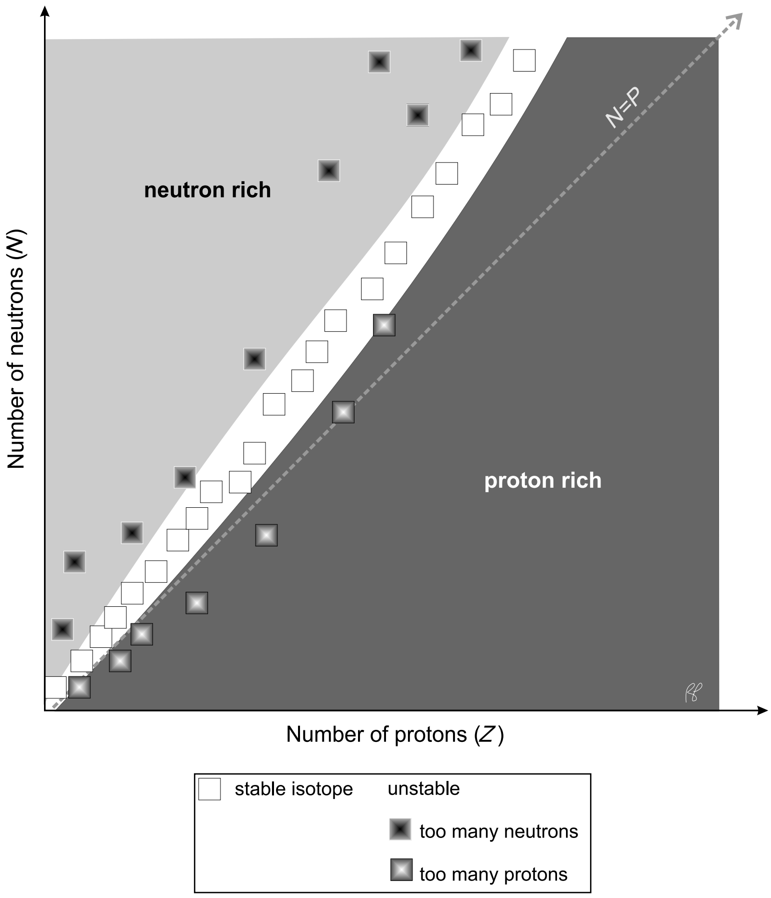

Figure 1-13 All combinations of neutrons and protons that can coexist in a stable nuclear configuration lie within the broad white band.

another fact of nature: the measured mass of a nucleus is always less than the mass expected from the sum of the masses of its neutrons and protons. The "missing" mass is called the **mass defect**, the energy equivalent of which is equal to the nuclear binding energy. This interchangeability of mass and energy was immortalized in Einstein's equation $E = mc^2$.

The Stable Nucleus

Not all elements have stable nuclei; they do exist for most of the light and mid-weight elements, those with atomic numbers up to and including bismuth ($Z = 83$). The exceptions are technetium ($Z = 43$) and promethium ($Z = 61$). All those with atomic numbers higher than 83, such as radium ($Z = 88$) and uranium ($Z = 92$), are inherently unstable because of their large size.

For those nuclei with a stable state there is an optimal ratio of neutrons to protons. For the lighter elements this ratio is approximately 1 : 1; for increasing atomic weights, the number of neutrons exceeds the number of protons. A plot depicting the number of neutrons as a function of the number of protons is called the **line of stability,** depicted as a broad white band in Figure 1-13.

Isotopes, Isotones, and Isobars

Each atom of any sample of an element has the same number of protons (the same **Z**: atomic number) in its nucleus. Lead found anywhere in the world will always be composed of atoms with 82 protons. The same does not apply, however, to the number of neutrons in the nucleus.

An **isotope** of an element is a particular variation of the nuclear composition of the atoms of that element. The number of protons (***Z***: atomic number) is unchanged, but the number of neutrons (***N***) varies. Since the number of neutrons changes, the total number of neutrons and protons (***A***: the atomic mass) changes. Two related entities are **isotones** and **isobars**. Isotones are atoms of different elements that contain identical numbers of neutrons but varying numbers of protons. Isobars are atoms of different elements with identical numbers of nucleons. Examples of these are illustrated in Figure 1-14.

Radioactivity

The Unstable Nucleus and Radioactive Decay

A nucleus not in its stable state will adjust itself until it is stable either by ejecting

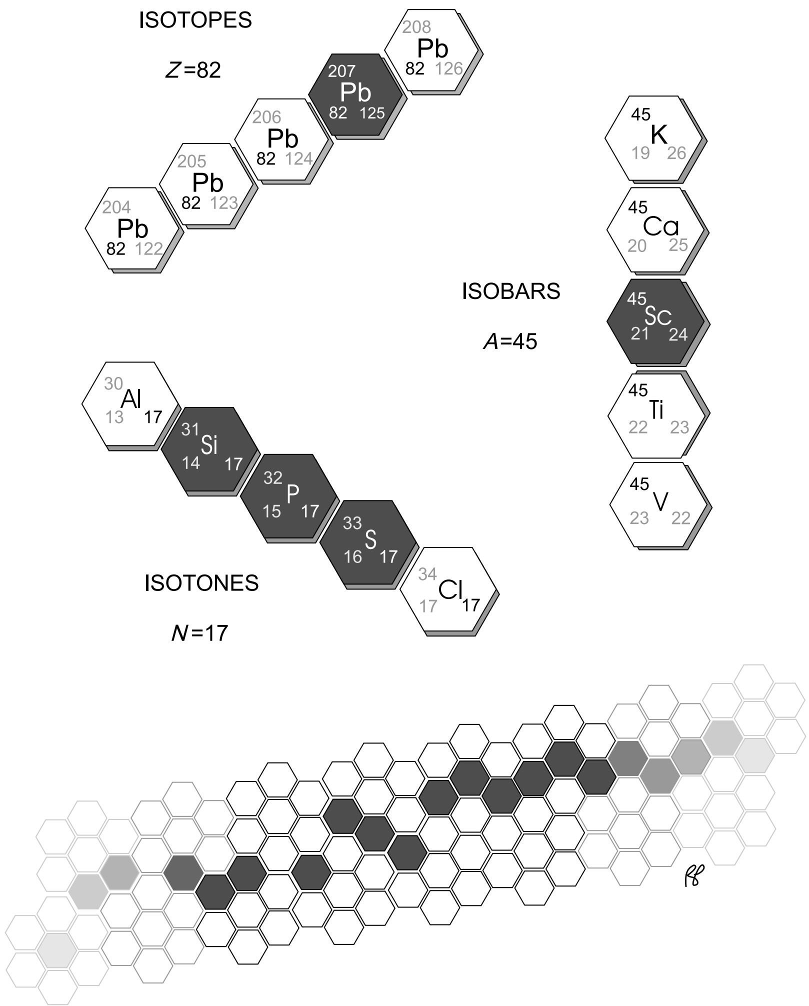

Figure 1-14 Nuclides of the same atomic number but different atomic mass are called isotopes, those of an equal number of neutrons are called isotones, and those of the same atomic mass but different atomic number are called isobars.

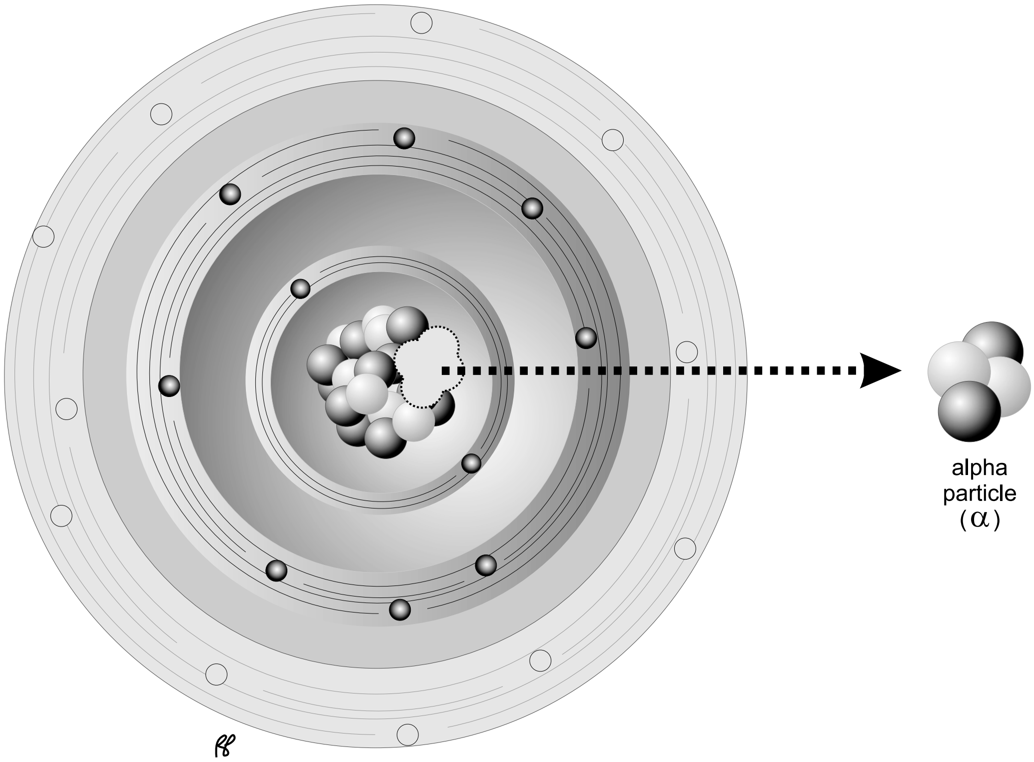

Figure 1-15 Alpha decay.

portions of its nucleus or by emitting energy in the form of photons (**gamma rays**). This process is referred to as **radioactive decay**. The type of decay depends on which of the following rules for nuclear stability is violated.

Excessive Nuclear Mass

Alpha Decay

Very large unstable atoms, atoms with high atomic mass, may split into nuclear fragments. The smallest stable nuclear fragment that is emitted is the particle consisting of two neutrons and two protons, equivalent to the nucleus of a helium atom. Because it was one of the first types of radiation discovered, the emission of a helium nucleus is called **alpha radiation**, and the emitted helium nucleus is called an **alpha particle** (Fig. 1-15).

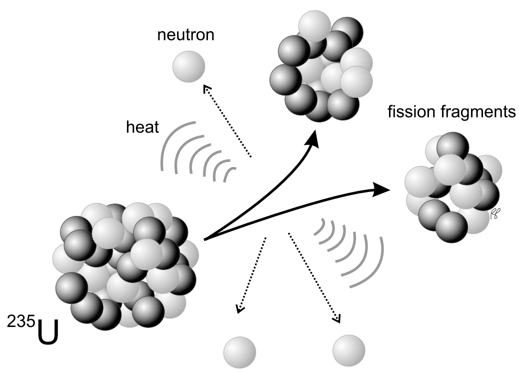

Figure 1-16 Fission of a ^{235}U nucleus.

Fission

Under some circumstances, the nucleus of the unstable atom may break into larger fragments, a process usually referred to as **nuclear fission**. During fission two or three neutrons and heat are emitted (Fig. 1-16).

Unstable Neutron–Proton Ratio

Too Many Neutrons: Beta Decay

Nuclei with excess neutrons can achieve stability by a process that amounts to the conversion of a neutron into a proton and an electron. The proton remains in the nucleus, but the electron is emitted. This is called **beta radiation**, and the electron itself is called a **beta particle** (Fig. 1-17). The process and the emitted electron were given these names to contrast with the alpha particle before the physical nature of either was discovered. The beta particle generated in this decay will become a free electron until it finds a vacancy in an electron shell either in the atom of its origin or in another atom.

Careful study of beta decay suggested to physicists that the conversion of neutron to proton involved more than the emission of a beta particle (electron). Beta emission satisfied the rule for conservation of charge in that the neutral neutron yielded one positive proton and one negative electron; however, it did not appear to satisfy the equally important rule for conservation of energy. Measurements showed that most of the emitted electrons simply did not have all the energy expected. To explain this apparent discrepancy, the emission of a second particle was postulated and that particle was later identified experimentally. Called an **antineutrino** (neutrino for small and neutral), it carries the "missing" energy of the reaction.

Too Many Protons: Positron Decay and Electron Capture

In a manner analogous to that for excess neutrons, an unstable nucleus with too many protons can undergo a decay that has the effect of converting a proton into a neutron. There are two ways this can occur: positron decay and electron capture.

Positron decay: A proton can be converted into a neutron and a **positron**, which is an electron with a positive, instead of negative, charge (Fig. 1-18). The positron is also referred to as a positive beta particle or positive electron or anti-electron. In positron decay, a **neutrino** is also emitted. In many ways, positron decay is the mirror image of beta decay: positive electron instead of negative electron, neutrino instead of antineutrino. Unlike the negative electron, the positron itself survives only briefly. It quickly encounters an electron (electrons are plentiful in matter), and both are **annihilated** (see Fig. 8-1). This is why it is considered an anti-electron.

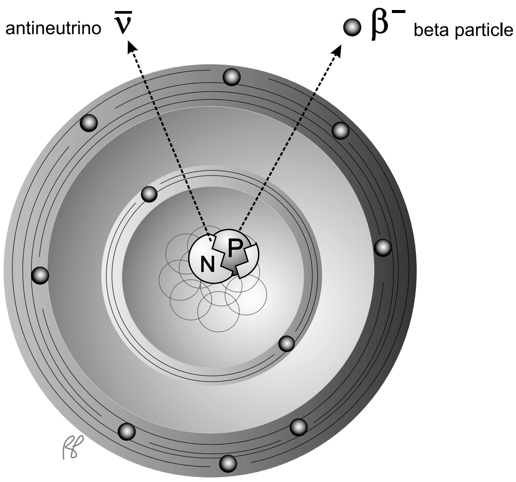

Figure 1-17 β^- (negatron) decay.

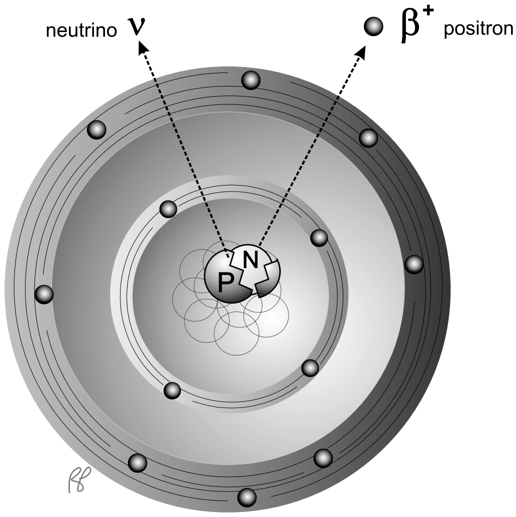

Figure 1-18 β^+ (positron) decay.

Generally speaking, antiparticles react with the corresponding particle to annihilate both.

During the annihilation reaction, the combined mass of the positron and electron is converted into two photons of energy equivalent to the mass destroyed. Unless the difference between the masses of the parent and daughter atoms is at least equal to the mass of one electron plus one positron, a total equivalent to 1.02 MeV, there will be insufficient energy available for positron emission.

ENERGY OF BETA PARTICLES AND POSITRONS

Although the total energy emitted from an atom during beta decay or positron emission is constant, the relative distribution of this energy between the beta particle and antineutrino (or positron and neutrino) is variable. For example, the total amount of available energy released during beta decay of a phosphorus-32 atom is 1.7 MeV. This energy can be distributed as 0.5 MeV to the beta particle and 1.2 MeV to the antineutrino, or 1.5 MeV to the beta particle and 0.2 MeV to the antineutrino, or 1.7 MeV to the beta particle and no energy to the antineutrino, and so on. In any group of atoms the likelihood of occurrence of each of such combinations is not equal. It is very uncommon, for example, that all of the energy is carried off by the beta particle. It is much more common for the particle to receive less than half of the total amount of energy emitted. This is illustrated by Figure 1-19, a plot of the number of beta particles emitted at each energy from zero to the maximum energy released in the decay. $E_{\beta\text{max}}$ is the maximum possible energy that a beta particle can receive during beta decay of any atom, and $\bar{E}_\beta$ is the average energy of all beta particles for decay of a group of such atoms. The average energy is approximately one-third of the maximum energy

$$\bar{E}_\beta \cong \tfrac{1}{3} E_{\beta\text{max}} \qquad \textit{(Eq. 1-1)}$$

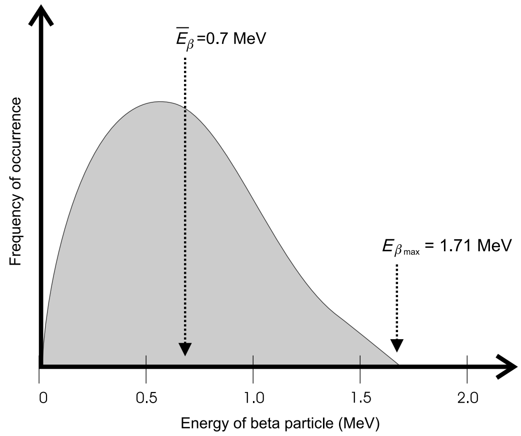

Figure 1-19 Beta emissions (both β^- and β^+) are ejected from the nucleus with energies between zero and their maximum possible energy ($E_{\beta\,\text{max}}$). The average energy ($\bar{E}_\beta$) is equal to approximately one third of the maximum energy. This is an illustration of the spectrum of emissions for 32P.

Electron capture: Through a process that competes with positron decay, a nucleus can combine with one of its inner orbital electrons to achieve the net effect of converting one of the protons in the nucleus into a neutron (Fig. 1-20). An outer-shell electron then fills the vacancy in the inner shell left by the captured electron. The energy lost by the "fall" of the outer-shell electron to the inner shell is emitted as an x-ray.

Appropriate Numbers of Nucleons, but Too Much Energy

If the number of nucleons and the ratio of neutrons to protons are both within their stable ranges, but the energy of the nucleus is greater than its resting level (an excited state), the excess energy is shed by **isomeric transition**. This may occur by either of the competing reactions, gamma emission or internal conversion.

Gamma Emission

In this process, excess nuclear energy is emitted as a **gamma ray** (Fig. 1-21). The name gamma was given to this radiation, before its physical nature was understood, because it was the third (alpha, beta, gamma) type of radiation discovered. A gamma ray is a photon (energy) emitted by an excited nucleus. Despite its unique name, it cannot be distinguished from photons of the

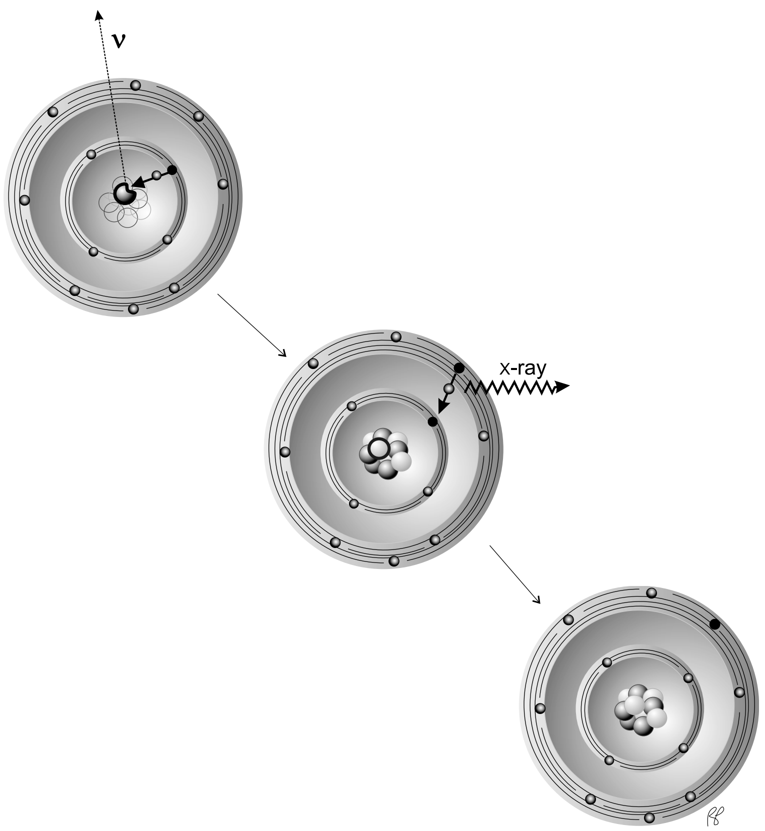

Figure 1-20 Electron capture.

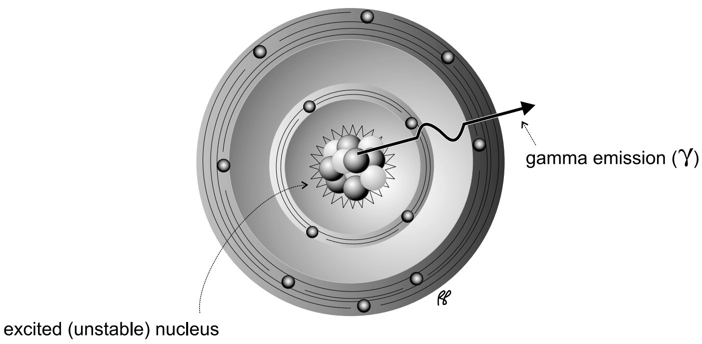

Figure 1-21 Isomeric transition. Excess nuclear energy is carried off as a gamma ray.

same energy from different sources, for example x-rays.

Internal Conversion

The excited nucleus can transfer its excess energy to an orbital electron (generally an inner-shell electron) causing the electron to be ejected from the atom. This can only occur if the excess energy is greater than the binding energy of the electron. This electron is called a **conversion electron** (Fig. 1-22). The resulting inner orbital vacancy is rapidly filled with an outer-shell electron (as the atom assumes a more stable state, inner orbitals are filled before outer orbitals).

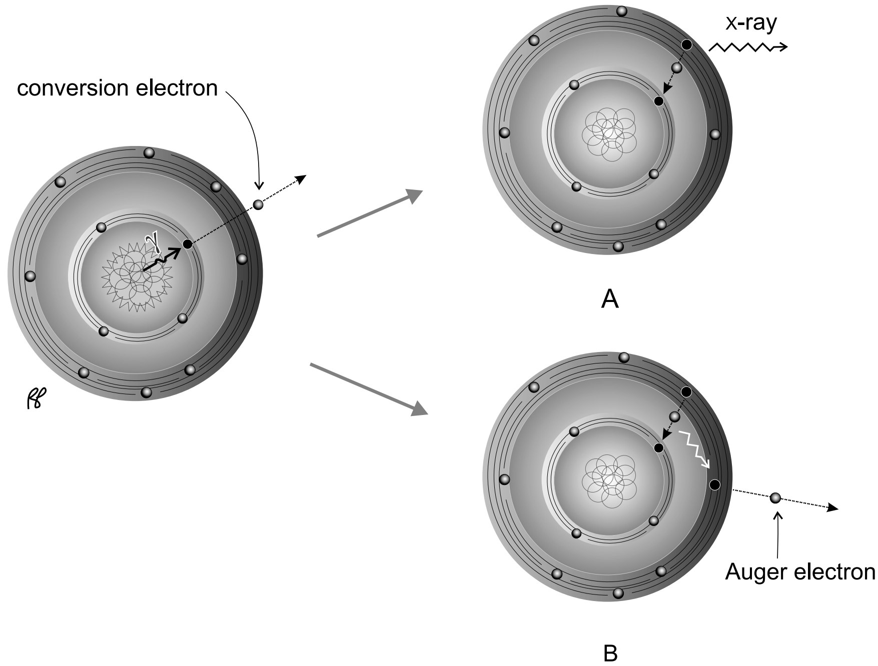

Figure 1-22 Internal conversion. As an alternative to gamma emission, it can lead to emission of either an x-ray (A) or an Auger electron (B).

The energy released as a result of the "fall" of an outer-shell electron to an inner shell is emitted as an x-ray or as a free electron (**Auger electron**).

Table 1-4 reviews the properties of the various subatomic particles.

Decay Notation

Decay of a nuclide from an unstable (excited) to a stable (ground) state can occur in a series of steps, with the production of particles and photons characteristic of each step. A standard notation is used to describe these steps (Fig. 1-23). The uppermost level of the schematic is the state with the greatest energy. As the nuclide decays by losing energy and/or particles, lower horizontal levels represent states of relatively lower energy. Directional arrows from one level to the next indicate the type of decay. By convention, an oblique line angled downward and to the left indicates electron capture; downward and to the right, beta emission; and a vertical arrow, an isomeric transition. The dogleg is used for positron emission. Notice that a pathway ending to the left, as in electron capture or positron emission, corresponds to a decrease in atomic number. On the other hand, a line ending to the right, as in beta emission, corresponds to an increase in atomic number.

Figure 1-24 depicts specific decay schemes for ^{99m}Tc, ^{111}In, and ^{131}I. The "m" in ^{99m}Tc stands for **metastable**, which refers to an excited nucleus with an appreciable lifetime ($>10^{-12}$ seconds) prior to undergoing isomeric transition.

Half-Life

It is not possible to predict when an individual nuclide atom will decay, just as in preparing popcorn one cannot determine when any particular kernel of corn will open. However, the average behavior of a large number of the popcorn kernels is predictable. From experience with microwave popcorn, one knows that half of the kernels will pop within 2 min and most of the

Table 1-4 Properties of the Subatomic Particles

Name(s)	Symbol	Mass[a]	Charge
Neutron	N	1840	None
Proton	P	1836	Positive (+)
Electron	e^-	1	Negative (−)
Beta particle (beta minus particle, electron)[b]	β^-	1	Negative (−)
Positron (beta plus particle, positive electron)	β^+	1	Positive (+)
Gamma ray (photon)	γ	None	None
X-ray	X-ray	None	None
Neutrino	ν	Near zero	None
Antineutrino	$\bar{\nu}$	Near zero	None

[a] Relative to an electron.
[b] There is no physical difference between a beta particle and an electron; the term beta particle is applied to an electron that is emitted from a radioactive nucleus. The symbol β without a minus or plus sign attached always refers to a beta minus particle or electron.

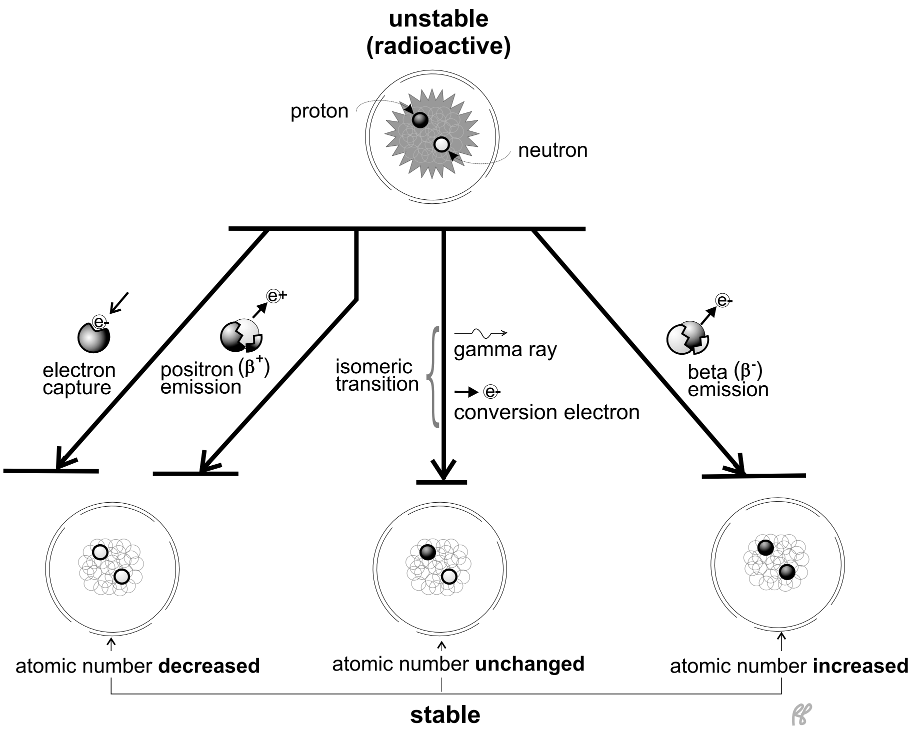

Figure 1-23 Decay schematics.

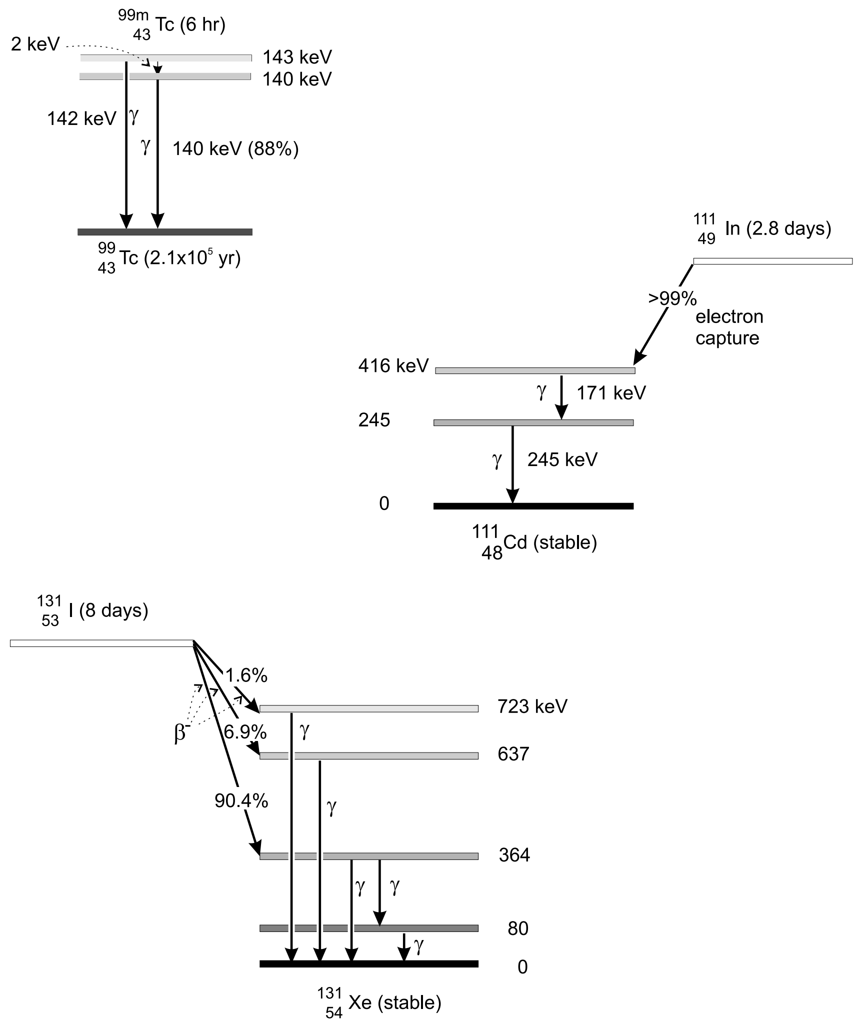

Figure 1-24 Decay schemes showing principal transitions for technetium-99m, indium-111, iodine-131. Energy levels are rounded to three significant figures.

bag will be done in 4 min. In a like manner, the average behavior of a radioactive sample containing billions of atoms is predictable. The time it takes for half of these atoms to decay is called (appropriately enough) the **half-life**, or in scientific notation $T_{1/2}$ (pronounced "T one-half"). It is not surprising that the time it takes for half of the remaining atoms to decay is also $T_{1/2}$. This process continues until the number of nuclide atoms eventually comes so close to zero that we can consider the process complete. A plot of $\boldsymbol{A}(t)$, the activity remaining, is shown in Figure 1-25. This curve, and therefore the average behavior of the sample of radioactivity, can be described by the **decay equation**:

$$A(t) = A(0)e^{-0.693t/T_{1/2}} \qquad \textit{(Eq. 1-2)}$$

where $A(0)$ is the initial number of radioactive atoms.

A commonly used alternative form of the decay equation employs the **decay constant** (λ), which is approximately 0.693 divided by the half-life ($T_{1/2}$):

$$\lambda = 0.693/T_{1/2} \qquad \textit{(Eq. 1-3)}$$

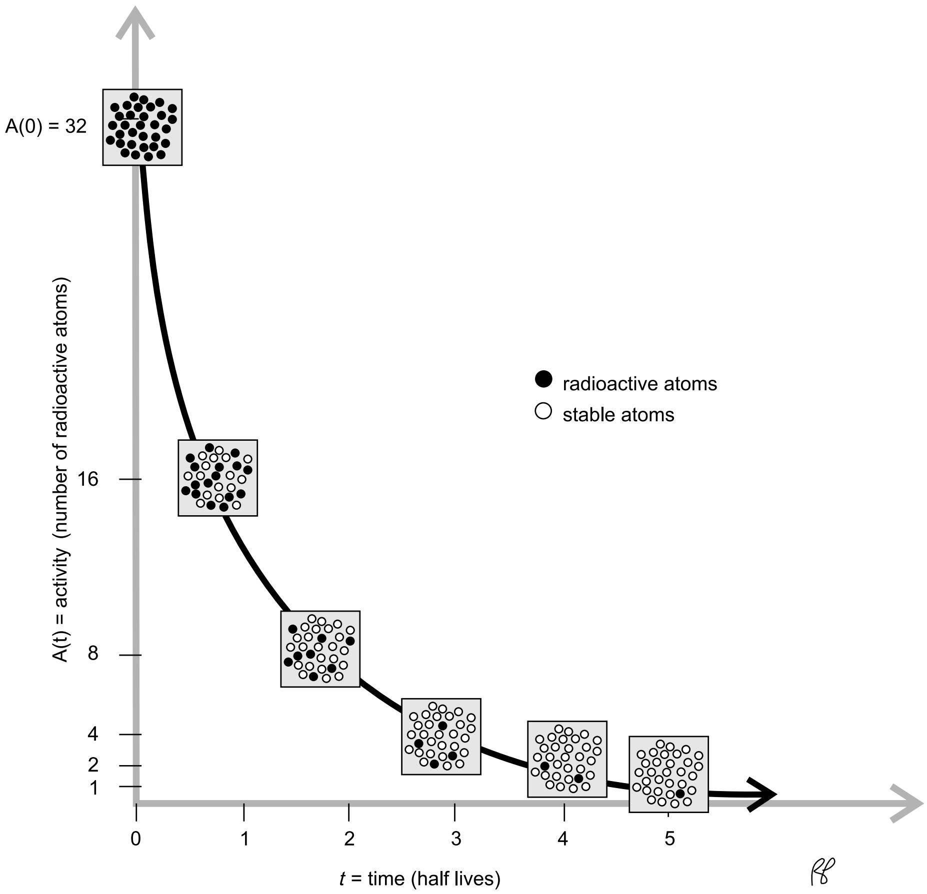

Figure 1-25 Decay curve. Note the progressive replacement of radioactive atoms by stable atoms as shown schematically in each box.

The decay equation can be rewritten as

$$A(t) = A(0)e^{-\lambda t} \quad \textit{(Eq. 1-4)}$$

The amount of activity of any radionuclide may be expressed as the number of decays per unit time. Common units for measuring radioactivity are the **curie** (after Marie Curie) or the newer SI unit, the **becquerel** (after another nuclear pioneer, Henri Becquerel). One becquerel is defined as one radioactive decay per second. Nuclear medicine doses are generally a million times greater and are more easily expressed in megabecquerels (MBq). One curie (Ci) is defined as 3.7×10^{10} decays per second (this was picked because it is approximately equal to the radioactivity emitted by 1 g of radium in equilibrium with its daughter nuclides). A partial list of conversion values is provided in Table 1-5.

A related term that is frequently confused with decay is the **count**, which refers to the registration of a single decay by a detector such as a Geiger counter. Most of the detectors used in nuclear medicine detect only a fraction of the decays, principally because the radiation from many of the decays is directed away from the detector. Count rate refers to the number of decays actually counted in a given time, usually counts per minute. All things being equal, the count rate will be proportional to the decay rate, and

Table 1-5 Conversion Values for Units of Radioactivity

One curie (Ci)	One millicurie (mCi)	One microcurie (μCi)	One bequerel (Bq)[a]	One megabecquerel (MBq)
	10^{-3} Ci	10^{-6} Ci	27×10^{-12} Ci	27×10^{-6} Ci
1×10^3 mCi		10^{-3} mCi	27×10^{-9} mCi	27×10^{-3} mCi
1×10^6 μCi	1×10^3 μCi		27×10^{-6} μCi	27 μCi
37×10^9 Bq	37×10^6 Bq	37×10^3 Bq		1×10^6 Bq
37×10^3 MBq	37 MBq	37×10^{-3} MBq	1×10^{-6} MBq	

[a] One becquerel equals one decay per second.

it is a commonly used, if inexact, measure of radioactivity.

Test Yourself

1 For each of the five terms below, choose the best definition

(1) Isobars
(2) Isoclines
(3) Isomers
(4) Isotones
(5) Isotopes

- **(a)** Atoms of the same element (equal Z) with different numbers of neutrons (N)
- **(b)** Atoms of different elements (different Z) with equal numbers of neutrons (N)
- **(c)** Atoms of different elements with equal atomic mass (A).
- **(d)** None of the above, usually used as a geological term.
- **(e)** Atoms of equal atomic mass (A) and equal atomic number (Z), but with unstable nuclei which exist in different energy states.

2 Which of the following statements are correct?

- **(a)** There is a stable isotope of technetium.
- **(b)** Atoms with atomic numbers (Z) > 83 are inherently unstable.
- **(c)** For light elements nuclear stability is achieved with equal numbers of protons and neutrons; for heavier elements the number of neutrons exceeds the number of protons.

3 For internal conversion to occur, the excess energy of the excited nucleus must equal or exceed:

- **(a)** 0.551 eV
- **(b)** 1.102 eV
- **(c)** the internal conversion coefficient
- **(d)** the average energy of the Auger electrons
- **(e)** the binding energy of the emitted electron.

4 For an atom undergoing beta decay, the average energy of the emitted beta particles is approximately:

- **(a)** 0.551 eV
- **(b)** 0.551 times the loss of atomic mass
- **(c)** one half of the total energy released for the individual event
- **(d)** one third of the maximum energy of the emitted beta particles
- **(e)** equal to the average energy of the accompanying antineutrinos.

5 You receive a dose of ^{99m}Tc measuring 370 MBq from the radiopharmacy at 10 AM. Your patient does not arrive in the department until 2 PM. How much activity, in millicurie, remains? (The $T_{1/2}$ of ^{99m}Tc is 6 hours. $e = 2.718$).