3 Atoms and Radiation

Physics is mathematical not because we know so much about the physical world, but because we know so little; it is only its mathematical properties that we can discover.

Bertrand Russell

3.1 Matter and Energy

Every thing in this Universe is made up of *matter* and *energy*. Matter is anything that has *mass* and occupies space. In contrast, light and heat are forms of energy. The concept of mass is central to the discussion of matter and energy. The mass of an object depends on the quantity of matter in the object, while the mass of a body (material object) is directly associated with its weight. On earth, the weight of a body is the pull of the earth on the body, which is proportional to its mass and depends on the distance of the body from the center of the earth. Energy is the ability to produce change or the capacity to do work. Energy occurs in many forms such as electrical energy, mechanical energy, chemical energy and nuclear energy. In addition, energy can also be *potential energy* (energy due to position) or *kinetic energy* (energy due to motion). The units for the quantities of mass and energy are shown in Table 3.1.

3.1.1 Mass–Energy Relationship

The most famous relationship Einstein derived from the postulates of special relativity – concerns mass (m) and energy or the rest energy (E_0).

$$E_0 = m_0 c^2 \tag{3.1}$$

In nonrelativistic physics, kinetic energy (KE) of an object of rest mass m_0 and speed v is given by

$$KE = \frac{1}{2} m_0 v^2 \tag{3.2}$$

In relativistic physics, if the mass is moving, the total energy (E) is

$$E = mc^2 = \frac{m_0 c^2}{\sqrt{1 - v^2 / c^2}} \tag{3.3}$$

The mass of a body moving at the speed v relative to an observer is larger than its mass when at rest relative to the observer by the factor, $\sqrt{1 - v^2 / c^2}$ where "c" is the velocity of light. Relativistic mass increases are significant only at speeds approaching that of light. For particulate radiation, relativistic effects are significant. Since mass and energy are not independent entities, according to the principle of conservation of mass energy, mass can be created or destroyed. When this happens, an equivalent amount of energy simultaneously vanishes or comes into being, and vice versa. Mass and energy are different aspects of the same thing. This mass–energy relationship is key for the estimation of nuclear binding energies and for the explanation of the liberation of enormous amounts of energy during the radioactive transformation of nuclides.

3.2 Radiation

The term *radiation* simply refers to *energy in transit*. There are two specific forms of radiation:

S. Vallabhajosula, *Molecular Imaging: Radiopharmaceuticals for PET and SPECT*,
DOI: 10.1007/978-3-540-76735-0_3,

Table 3.1 Primary and derived units of quantities based on Système Internationale (SI)

Quantity	Unit	Symbol	Equivalent
Mass	kilogram	kg	1 kg = 1,000 g
Amount	Mole	mol	1 mol = Avogadro's constant, 6.02×10^{23} particles (atoms or molecules)
length	Meter	m	Distance light travels in exactly 1/299,792,458 s. 1 m = 100 cm = 1.0×10^{10} Å
time	Second	s	Duration of 9,192,631,770 cycles of microwave radiation produced by Cs-133 atoms
Temperature	Kelvin	°K	°C + 273.16
Current	Ampere	A	1 coulomb s^{-1}
Energy	Joule	J	4.184 J = 1 cal
Electric potential	Volt	V	1 eV = 1.60269×10^{-19} J

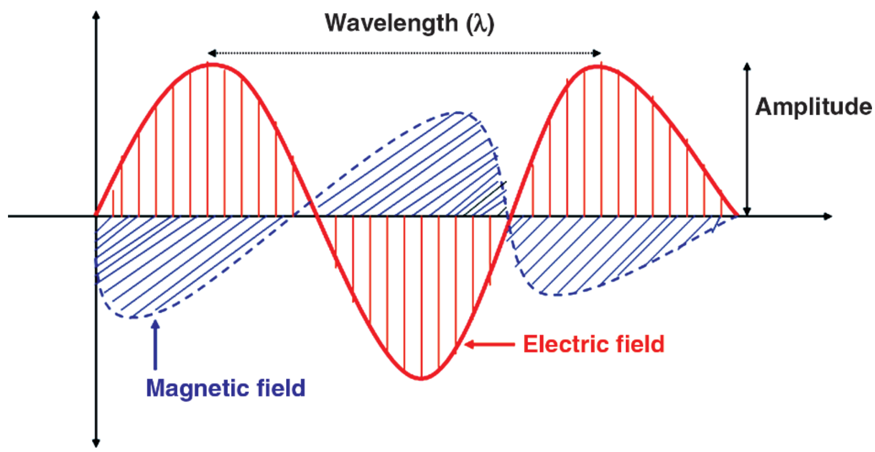

Fig. 3.1 The electric and magnetic fields in an electromagnetic wave are perpendicular to each other and to the direction of the wave

- Particulate radiation (nonpenetrating): mass (subatomic particles such as protons, neutrons, electrons) in motion carrying kinetic energy of particles
- Electromagnetic radiation: Oscillating electric and magnetic fields carrying energy and traveling through space with a constant velocity

3.2.1 Electromagnetic Radiation

Visible light is the most familiar form of electromagnetic radiation. Ultraviolet, infrared, microwave, radio waves, X-rays, γ rays, and annihilation radiation are all different types of electromagnetic radiation. All types of electromagnetic radiations exhibit "wave-like" behavior in their interactions with matter. The wavelength (λ) and frequency (ν) of the oscillating fields of electromagnetic radiation (Fig. 3.1) are related by

$$\lambda\nu = c \tag{3.4}$$

where c is the velocity of light in free space (2.998×10^8 m/s).

Maxwell was able to show that that the speed "c" of electromagnetic waves in vacuum can be deduced mathematically by

$$c = \frac{1}{\sqrt{\varepsilon_0\mu_0}} = 2.998\times10^8\,\mathrm{m/s} \tag{3.5}$$

where ε_0 is the electric permittivity of free space (8.854 × 10–12 $c^2/N\,m^2$) and μ_0 is its magnetic permeability ($4\pi \times 10^{-7}$ $T\,m/A$).

According to quantum theory, electromagnetic radiation, however, behaves as discrete packets of energy (with no mass or charge), called quanta or photons. The energy of the photon is related to the frequency of the electromagnetic radiation by

$$E = h\nu \tag{3.6}$$

where h is known as Planck's constant whose value is 6.626×10^{-34} J s.

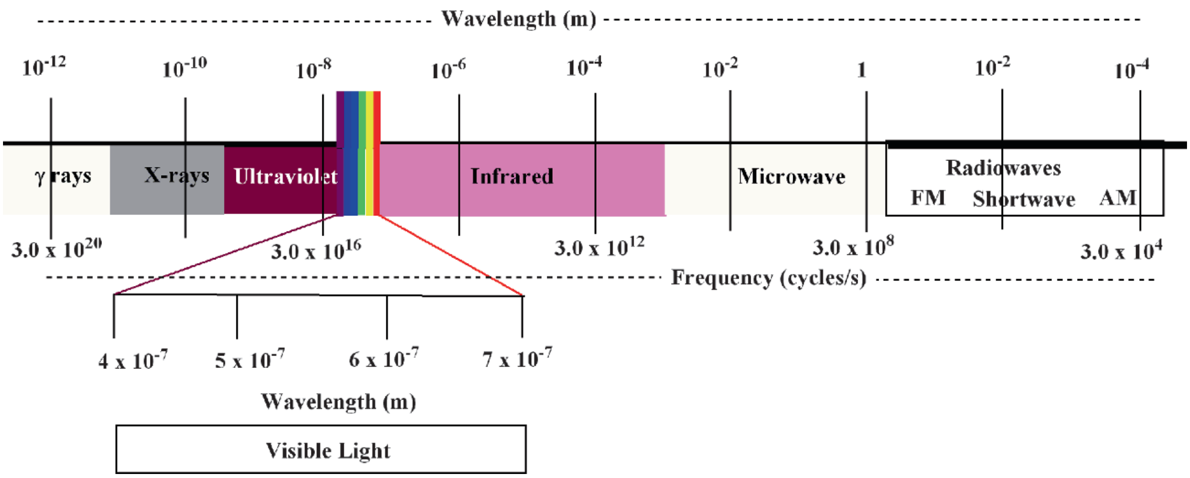

Fig. 3.2 The spectrum of electromagnetic radiation. The visible light (light waves) span only a brief frequency interval, from 4.3×10^{14} Hz for red light to about 7.4×10^{14} Hz for violet light

De Broglie suggested that a photon, an energy particle behaves in certain ways as though it has a wave nature. The wavelength of a photon is therefore specified by its momentum

$$\lambda = \frac{h}{\rho} \tag{3.7}$$

where, the momentum

$$\rho = \frac{h\nu}{c} \tag{3.8}$$

since

$$\lambda\nu = c \tag{3.9}$$

Different types of electromagnetic radiation and their corresponding photon energies, wavelengths, and frequencies are shown in Fig. 3.2.

3.3 Classification of Matter

Based on the composition of matter as a basis for classification, matter may be regarded as a pure substance or as a mixture of substances. There are two kinds of substances: *elements* and *compounds*. An element or a chemical element is a substance that cannot be broken down to simpler substances by ordinary chemical means. Compounds are substances consisting of two or more elements (some times also called molecules) combined in definite proportions by mass, having properties different from that of any of its constituent elements. A mixture consists of two or more substances and there are two kinds of mixtures; homogenous mixtures (also called solutions) and heterogeneous mixtures.

3.3.1 Chemical Element

All matter in nature is made up of pure substances, known as chemical elements. In the periodic table (Fig. 3.3), 112 known chemical elements have been listed; 22 of these elements have been man made. Oxygen is the most abundant element in the earth's crust. Air is mostly a mixture of nitrogen (78%) and oxygen (21%). Since ancient times, the following ten elements have been known to be pure substances (carbon, sulfur, copper, silver, gold, iron, tin, mercury, tin and antimony) were known as pure substances. The rest of the elements were discovered by man. In 1869, based on the chemical properties and relative atomic weights of approximately 70 natural elements, Mendeleyev, a Russian chemist arranged the elements in a table of groups and periods, known as the periodic table. The long form of the modern periodic table with all the 112 elements is shown in Fig. 3.3.

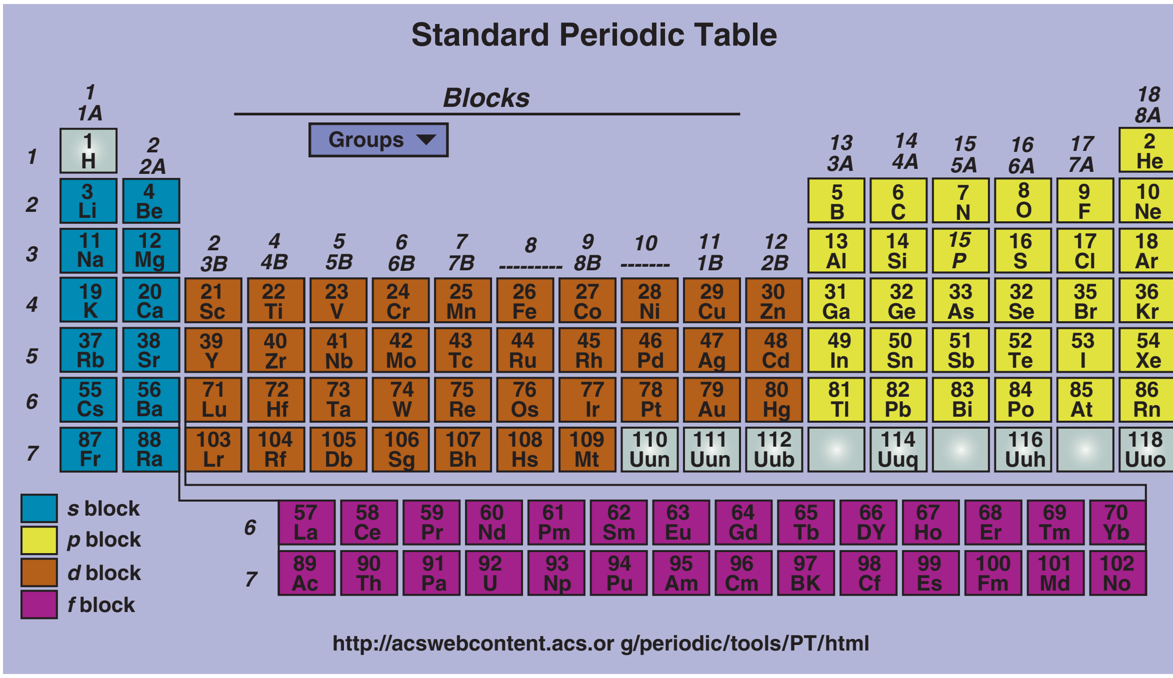

Fig. 3.3 The modern standard periodic table showing 18 groups

3.4 Atoms

In 1804, John Dalton proposed the existence of atoms and attributed certain properties to the *atoms* based on experimental studies. Every chemical element is composed of individual particles, called atoms. An atom by definition is the smallest unit into which a chemical element can be broken down, without losing its chemical identity. Dalton assumed that all the atoms of a given element have the same *atomic weight* or *atomic mass*. By 1914, it was realized that atoms of the same element can vary in weight, but may have the same chemical properties. The word *isotope* was coined by Frederick Soddy to represent atoms of an element that have different weights, but which can still be placed in the same position in the periodic table. An isotope can be stable or unstable (radioactive).

Some of the elements in nature such as Na, F, Al, and Bi have no isotopes, while 81 of the elements listed in the periodic table have at least one stable isotope. Natural carbon exists in two stable isotopic forms (^{12}C and ^{13}C) while tin has ten stable isotopes. All elements heavier than Bi are unstable; no stable isotope exists for these elements. In 1947, the American chemist Truman Paul Kohman suggested that the word *nuclide* is a more appropriate term to represent both, stable and unstable atoms (or radionuclide) of an element.

3.4.1 Atomic Structure

In 1910, Rutherford presented a model of an atomic structure known as *nuclear atom*. According to this nuclear model, atoms consist of a massive, compact, positively charged core, or *nucleus* surrounded by a diffuse cloud of relatively light, negatively charged electrons. In 1913, Moseley formulated the property of *atomic number* (Z), which is the number of positive charges in the nucleus. In 1914, Rutherford named this positively charged particle in the nucleus, *proton* (from the Greek word first). Proton has an electrical charge, equal to that of an electron; however; a proton is positive and has a mass 1836 times that of the electron. In the electrically neutral atom, the number of orbiting electrons is sufficient to balance the number of positive protons. In the 1920s, Rutherford sumrised that the existence of a neutral particle (the same size as a proton) within the nucleus accounts for the total mass of the atom. Such a neutral particle was discovered in 1932 by James Chadwick eventually discovered these

Table 3.2 Fundamental particles of matter

	Mass[a]		Energy[b]	Charge	
Particle	U	kg	MeV	Elementary	Coulombs (C)
Electron, e^-	0.0005486	9.11×10^{-31}	0.511	−1	-1.602×10^{-19}
Positron, e^+	0.0005486	9.11×10^{-31}	0.511	+1	$+1.602 \times 10^{-19}$
Proton, p/H^+	1.007825	1.673×10^{-27}	938.78	+1	$+1.602 \times 10^{-19}$
Neutron, n°	1.008665	1.675×10^{-27}	939.56	0	–

[a] Mass is expressed in international mass unit (u), which is equal to 1/12th the mass of carbon-12 atom (1.66054×10^{-24} g)
[b] The energy given here is the rest mass energy of the particle

neutral particles. Subsequently, Heisenberg proposed that the nuclei of different elements consist of protons and neutrons which are held together by strong exchange forces known as nuclear forces. The total number of protons and neutrons within the nucleus is called the *mass number* (A), which is very close to the atomic weight of an element. The mass number (A) of ^{12}C nuclide is 12 and its atomic mass is considered as 12 atomic mass units (AMU or u). The atomic weight of natural carbon, however, is 12.011 since it is a mixture of ^{12}C (98.89%) and ^{13}C (1.11%) nuclides or isotopes. The mass–energy relationships of the fundamental subatomic particles are summarized in Table 3.2.

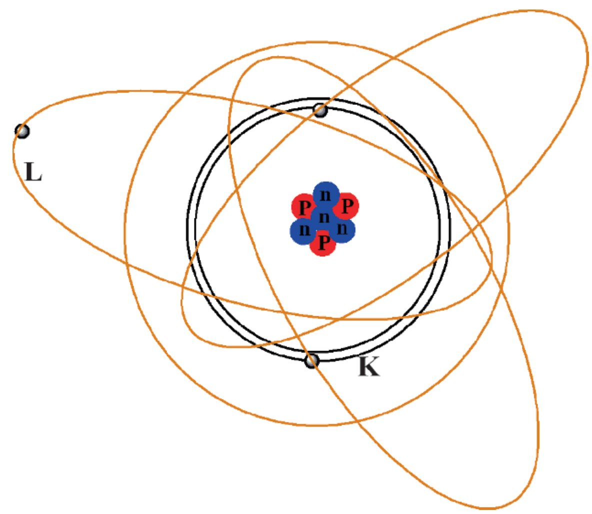

Fig. 3.4 Schematic representation of the Bohr model of atom showing circular (*K* shell) and elliptical (*L* shell) orbits

3.4.2 The Bohr Model of an Atom

The atom consists of an extremely dense, small positively charged nucleus surrounded by a cloud of electrons with small size and mass, each carrying a single negative charge equal but opposite to that of a proton in the nucleus. Because the number of electrons and protons in an atom, is the same the whole atom is electrically neutral. The *atomic volume* (10^{-8} cm diameter) with the cloud of electrons is significantly larger than the volume of the nucleus (10^{-13} cm diameter). Based on Rutherford's discovery of the atomic nucleus and Plank's discovery of the energy quantum ($E = h\nu$), Neils Bohr developed the first quantum model of an atom in 1913 (Fig. 3.4). According to this model, an electron in a hydrogen atom rotates around the nucleus at high speeds in closed *circular orbits* associated with a characteristic *quantum number (n)*. The electron in general exists in a low energy orbit (*ground state*). Gain or loss of a quantum of energy occurs only when an electron moves from one orbit to one of greater or lesser energy. As a result, an atom can absorb or emit energy in discrete units or quanta.

In 1916, Arnold Sommerfeld suggested that the electron orbits can be elliptical, and to different degrees. Consequently, the orbital *quantum number (l)* was introduced. The circular and elliptical orbits are assumed to be in a single plane. When atoms are in a magnetic field, however, the orbits may be tipped. These orbits can best be represented in a three dimensional spherical space around the nucleus. To account for this magnetic effect, a *magnetic quantum number (m)* was introduced. Subsequently, Wolfgang Pauli introduced the *spin quantum number* (*s*) to represent the direction of electron spin (clockwise or anticlockwise).

The arrangement of electrons around the nucleus (Fig. 3.4) in an atom can be described based on the four quantum numbers. A more detailed representation of electron orbits based on quantum mechanics is described in the chemistry section Chap. 7.

3.4.2.1 Electron Binding Energy

When the atom is in a ground state, electrons in an atom occupy the innermost shells of an atom. In the most stable configuration, electrons are most tightly bound to the nucleus. Electrons can move to higher energy or higher shells or can even be completely removed from the atom by providing energy to the electrons. The amount of energy required to remove an electron from a given shell (and to overcome the force of attraction of the nucleus), is called the *binding energy* of the shell. The binding energy is greatest for the innermost shell (*K* shell, $n = 1$) and increases as the atomic number (Z) increases (Fig. 3.5). The energy required for an electron to jump from a lower shell (*K* shell) to a higher shell (*L*, *M*, *N*, etc) is exactly equal to the difference between the shells.

Calculations of the energy levels of electrons in an atom involve the following physical principles. Based on the electron rest mass (m), velocity (v) and radius of the orbit (r), the angular momentum (mvr) of the electron about the nucleus is given by

$$mvr = n\frac{h}{2\pi} \tag{3.10}$$

where n is the principal quantum number and h is Planck's constant (6.626×10^{-34} J s).

The force experienced by an electron (mv^2/r) in its orbit is supplied by the Coulomb attraction between the electron ($-e$) and the nuclear charge ($+Ze$), Therefore the equation for the motion of electron is given by

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi\varepsilon_0 r^2} \tag{3.11}$$

where ε_0 is the *permittivity of free space*, and $1/4\pi\varepsilon_0 = k_0 = 8.98755 \times 10^{-9}$ N m^2 C^{-2}.

The total energy of the electron in any orbit (nth orbit) is given by the sum of its kinetic and potential energy

$$KE = \frac{1}{2}mv^2 + PE = -\frac{e^2}{4\pi\varepsilon_0 r} \tag{3.12}$$

$$E_n = \frac{k_0^2 Z^2 e^4 m}{2n^2\left(\frac{h}{2\pi}\right)^2} = -\frac{13.6 Z^2}{n^2} eV \tag{3.13}$$

The lowest energy occurs when $n = 1$. For a hydrogen atom, the ground state energy or the binding energy is -13.6 eV (Fig. 3.5). The energy required to remove an electron from its ground state is called the ionization

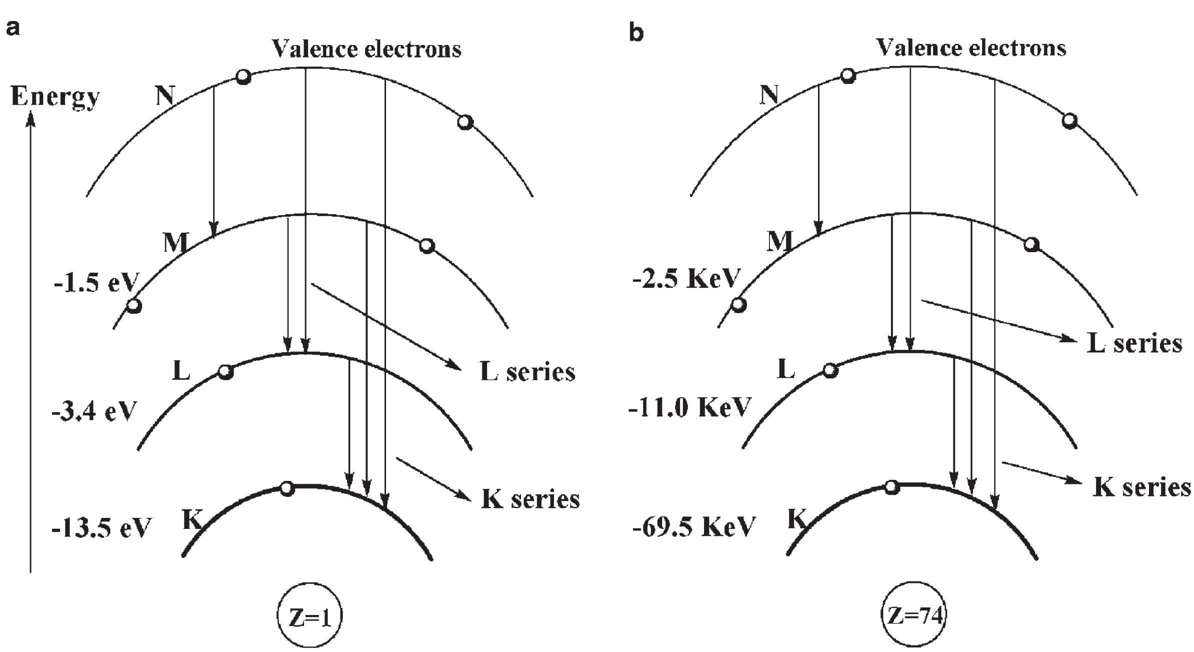

Fig. 3.5 The electron binding energy level diagram for hydrogen (**a**) and tungsten (**b**)

potential, which is 13.6 eV. In comparison, for lead, the electron binding energy of a *K* shell electron is 88,005 eV.

3.5 Nuclear Structure

3.5.1 Composition and Nuclear Families

The size or the diameter of the atomic nucleus is very small compared to the diameter of an atom (10^{-13} cm versus 101^{-8} cm). The nucleus is composed of protons and neutrons and, collectively, these elementary particles are known as *nucleons*. The mass of nucleons is almost 2,000 times that of electron mass. Therefore the density of the nucleus is very high compared to that of an atom.

For a given chemical element, the total number of protons (*Z*) and neutrons in the nucleus is known as the mass number (*A*), which is almost equal to the atomic weight (A_W). The difference between *A* and *Z*, is the neutron number (*N*).

A nuclide with a life span of greater than 10^{-12} s is generally characterized by an exact nuclear composition, given by *A*, *Z*, and the arrangement of nucleons within the nucleus. Even though 112 chemical elements are known, more than 3,000 nuclides have been discovered. The physical characteristics of the subatomic particles are summarized in Table 3.2.

Different combinations of protons and neutrons create different nuclear families. Atomic nuclei of the same element have the same number of protons (*Z*) but can have different numbers of neutrons (*N*) and different mass numbers (*A*). Such nuclides of an element are known as *isotopes* of that element. Nuclides with the same *A*, but different *Z* are called *isobars* while the nuclides with same *N* but different *A* and *Z* are called *isotones*. *Nuclear isomers* are nuclides with the same *A* and *Z*, but different nuclear energies. Examples of different nuclear families are shown below (Fig. 3.6).

Isotopes of Carbon	${}^{11}_{6}C_{5}$	${}^{12}_{6}C_{6}$	${}^{13}_{6}C_{7}$	${}^{14}_{6}C_{8}$
Isobars		${}^{11}_{6}C_{5}$	${}^{11}_{5}B_{6}$	
Isotones		${}^{14}_{6}C_{8}$	${}^{15}_{7}N_{8}$	
Isomers		${}^{99m}_{43}Tc_{56}$		${}^{99}_{43}Tc_{56}$

Fig. 3.6 Various nuclides or nuclear species are grouped into four major families having certain common characteristics. For example, isotopes of an element have the same number of protons (*Z*), but different mass number (*A*)

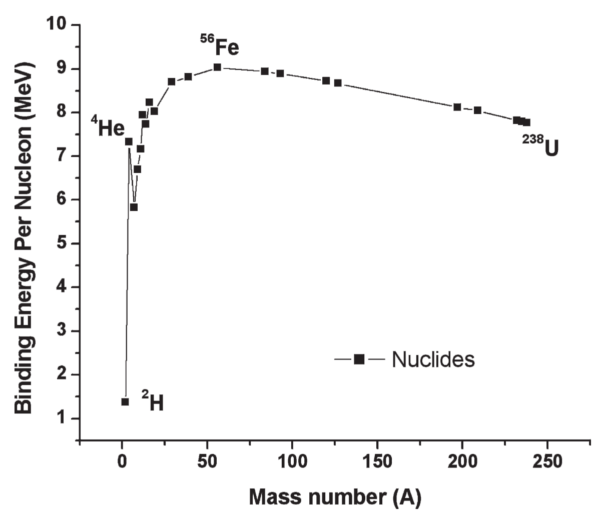

Fig. 3.7 Binding energy per nucleon versus mass number (*A*) for several stable nuclides. One of the isotopes of iron (^{56}Fe) has the highest nuclear binding energy

3.5.2 Nuclear Binding Energy

The binding energy is the minimum amount of energy required to overcome the forces holding the nucleus together and to separate the nucleus completely into the individual components. Most of the binding energy, however, is nuclear binding energy, since energy required to remove electrons from an atom is relatively small. Mass defect or deficiency (Δm) is the difference between the mass of the atom and the sum of the masses of individual components. Binding energy (E_B) is the energy equivalent of Δm. For example, in the case of ^{12}C atom, the Δm is 0.09906 u, which is equal to 92.22 MeV. The binding energy per nucleon (E_B/A) for ^{12}C, is 7.685 MeV. The binding energies of several nuclides as a function of mass number are shown in Fig. 3.7.

3.5.3 Nuclear Stability

Protons and neutrons within the nucleus are subject to two kinds of forces. Since like charges repel, electrical forces of repulsion exist between protons. In contrast, the strong exchange forces which can operate only at small distances (<3 fm), keep the protons and neutrons together in the nucleus. The shell model of the nucleus suggests that nucleons move in orbits about one another, while the liquid-drop model suggests that the nucleus is more like a drop of liquid.

Why are some combinations of protons and neutrons more stable than others? ^{12}C with six protons and six neutrons is stable forever, while ^{11}C with six protons and five neutrons is very unstable and decays with a half-life of only 20 min. In general, elements with a low atomic number (Z <10) contain equal number of protons and neutrons and, therefore, are more stable. In contrast, as the atomic number Z increases, more and more neutrons are needed in the nucleus to keep the elements stable (Fig. 3.8). In general nuclides with even number of protons and neutrons are more stable than nuclides with an odd number of protons and neutrons. The nuclear stability, therefore depends on

- Neutron/proton ratio (*n/p*): Among the isotopes of every element, there is an optimal ratio favoring stability. The number of neutrons required to maintain stability is approximately 1.5 Z for heavy elements. All elements with Z > 83 and A > 209 are unstable and spontaneously decay or transform into more stable combinations of protons and neutrons.
- The binding energy/nucleon (E_B/A): The binding energy is greatest (>8 MeV) for nuclides with $A \approx 60$. It decreases slowly with increasing A, indicating the tendency toward instability for heavy elements.

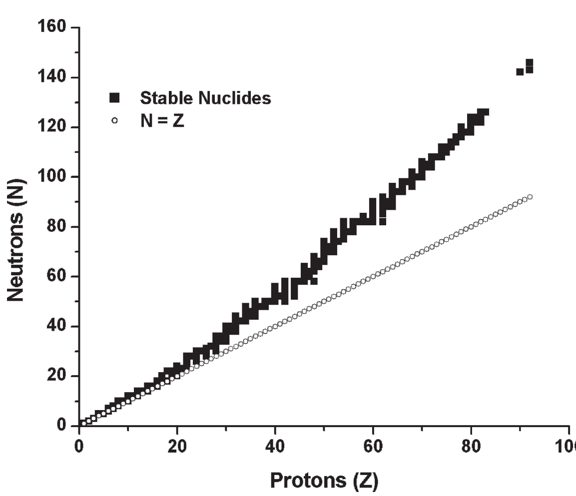

Fig. 3.8 Atomic number (Z) versus the neutron number (N) for the stable isotopes of elements. The neutron/proton ratio of unity represents a theoretical line of stability. The elements with higher Z, however, prefer *n*/*p* >1 in order to achieve stability

The most stable arrangement of nucleons within the nuc-leus is called the ground state. *Excited states* are unstable and have higher energies. They also have a transient existence (10^{-12} s) and tend to transform into stable ground states. *Metastable states* exist longer (>10^{-12} s), but, eventually, also transform to ground states.

3.6 Atomic and Nuclear Emissions

3.6.1 Emissions from Electron Shells

When atoms are excited, electrons in the inner shells gain energy and jump to higher energy outer shells, but with lower electron binding energies. When these electrons jump back to more stable lower orbits, energy is emitted in the form of electromagnetic radiation. All types of electromagnetic radiations, visible, UV, infrared, radio waves, and X-rays are emitted from atomic shells (Fig. 3.2). Also, electrons (*photoelectrons* or *Auger electrons*) are emitted from atoms under certain circumstances. The emission of photoelectrons following interaction of energetic photons is discussed in Chap. 6.

3.6.1.1 X-rays

In the photoelectric effect, photons of light transfer energy to the electrons in an atom. In the inverse photoelectric effect, the kinetic energy of an electron is converted into a photon. In 1895, Röntgen discovered that when cathode rays (negatively charged particles, electrons) strike the atoms of a glass vacuum tube, they produce penetrating x-rays that cause a salt (barium platinocyanide) to glow. Soon it was realized that x-rays are high energy electromagnetic (em) radiation. An accelerated electric charge will radiate em waves. As the fast moving cathode rays approach the positive charge in the nuclei of atoms in the glass, electrons are brought to rest (accelerated), and the kinetic energy of electrons is converted to em waves. Radiation produced under these circumstances was given the name *bremsstrahlung* (breaking radiation). It was also realized that when denser, more massive atoms in a metal are used to stop accelerated electrons, x-rays of higher energies can be generated.

3.6.1.2 Characteristic X-rays

In 1911, the British physicist Charles Barkla noticed that each metal produces x-rays of a particular wavelength, depending on the metal (Fig. 3.9) and called the more penetrating beam *K* x-rays and the less penetrating beam *L* x-rays. The wavelength of x-rays decreases (energy increases) as the atomic number of elements increased. Emissions from transitions >100 eV (0.1 keV) are called *characteristic or fluorescent x-rays*. Emission of characteristic x-rays occurs when orbital electrons move from an outer shell to fill an inner shell vacancy (such as $L \rightarrow K$ or $M \rightarrow K$ shell). The energy of the characteristic x-rays is the difference in binding energies of these shells. For example, de-excitation (e.g., $M \rightarrow K$ transition = K_β) of a tungsten atom, results in the emission of characteristic x-rays of 67 keV (Fig. 1.5 in Chap. 1); the difference in binding energy between *M* and *K* shells. The probability that the electron transition will result in the emission of characteristic x-rays is called *fluorescent yield* (ω), which is essentially zero for elements with low *Z* (<10) and increases as *Z* increases.

3.6.1.3 Auger Electrons

Just as in the production of characteristic x-rays, an electron from the outer shell drops down to fill the vacancy in a lower shell, and energy is released. As an alternative to the emission of x-rays, the energy is transferred to another orbital electron, which is ejected from the atom (Fig.3.9). The emission of the electron from the inner shell is known as the *Auger effect*, which in turn creates two vacancies in the shell. For example, the deexcitation (e.g., $M \rightarrow K$ transition = K_β) of a tungsten atom, results in the emission of an Auger electron with 64.5 keV (Fig. 1.5 in Chap. 1): the difference in binding energy between *M* and *K* shells minus the binding energy of another electron in *M* shell, which is ejected. Again, this vacancy is filled by outer electrons resulting in characteristic x-rays or Auger electrons. Elements with lower *Z* are more likely to eject Auger electrons while elements with higher *Z* are more likely to emit characteristic x-rays.

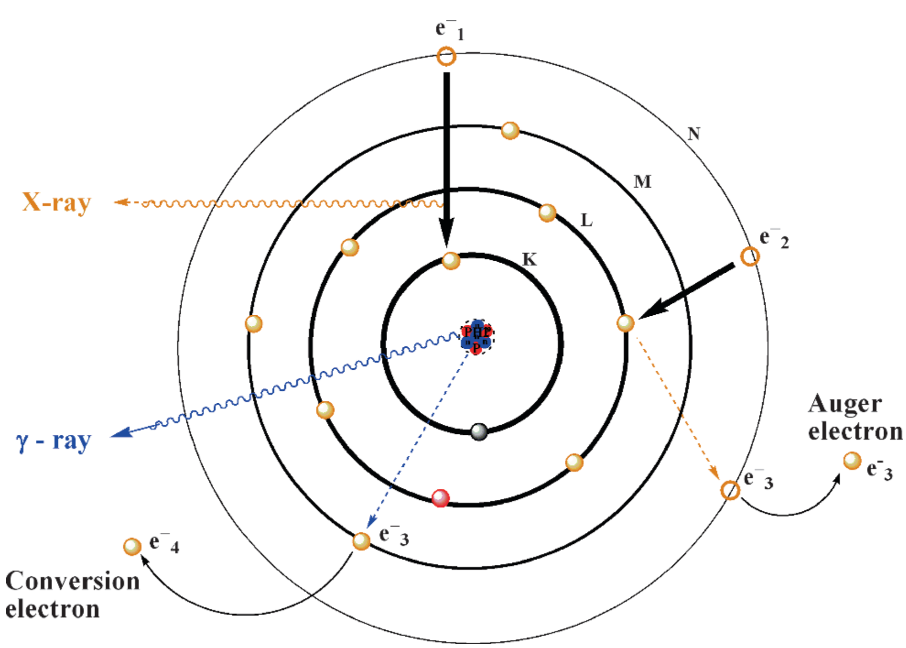

Fig. 3.9 Atomic emissions: When an electron from an outer higher energy shell moves in to fill a vacancy in a lower shell, the energy released may appear as electromagnetic radiations, such as x-rays or as an alternative, the atom may undergo Auger effect and may emit electrons, known as Auger electrons. In contrast, the energy from the nucleus may be emitted in the form of γ-rays or promote emission of conversion electrons

3.6.2 Nuclear Emissions

3.6.2.1 Gamma Rays

Like an excited atom, an excited nucleus can emit electromagnetic radiation in order to come to the ground state (Fig. 3.9). The photons emitted by the nuclei range in energy up to several MeV, and are traditionally called *γ rays*. The energy of a γ photon corresponds to the energy difference between the various initial and final states in the transitions involved. Deexcitation of the excited nuclei may emit only γ rays (as in the case of nuclear isomers) or emit γ rays in addition to nonpenetrating radiation such as β-rays, α-rays, protons and neutrons emitted during the radioactive decay process.

3.6.2.2 Internal Conversion

As an alternative to γ rays, an excited nucleus in some cases may return to its ground state by transferring the excitation energy to one of the orbital electrons, which is then ejected from the atom. This process of converting nuclear energy to eject an orbital electron (Fig. 3.9) is called *internal conversion*, which is analogous to the photoelectric effect. The *KE* of conversion electron is equal to the difference between the lost nuclear excitation energy and the binding energy of the electron from the ejected atomic shell.

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