Gopal B. Saha

Fundamentals of Nuclear Pharmacy

Seventh Edition

Springer

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With my heartfelt gratitude and respect to

Late Nishi Kumar Saha (grandfather),
Late Binoy Bhushan Saha (uncle),
Late Snehangshu Bimal Saha (brother-in-law),
Late Amar Das (brother-in-law),

Whose unwavering love, inspiration, and support
have made me what I am today!

Preface

Since the sixth edition of this book in 2010, armamentariums in nuclear medicine have changed with new additions and some deletions and hence the seventh edition. Obsolete and outdated radiopharmaceuticals and techniques have been removed, and new useful ones have been added. Like the previous editions, the book is aimed to the nuclear medicine and radiology fellows, residents, and technologists in preparation for their respective professional board examinations. Also nuclear medicine physicians and radiologists should find it a useful reference in their practice.

The organization of the book remains the same as the previous editions with a total of 16 chapters with the same titles. Chapters 1, 2, and 3 remain almost unchanged except the addition of a section on Gaussian distribution in Chap. 2 and a small section on PET/MR imaging in Chap. 3. A number of radionuclides have been added in Chap. 4, and the $^{227}\text{Ac} \rightarrow {}^{227}\text{Th} \rightarrow {}^{223}\text{Ra}$ generator has been added in Chap. 5. There is no change in Chaps. 6, 8, and 12. New US FDA-approved and some non-US FDA-approved radiopharmaceuticals for clinical use are described in Chap. 7. Several radiopharmacyrelated accessories have been added to Chap. 9, and radiation doses for new radiopharmaceuticals have been included in Chap. 10. Revised and new regulations of the US FDA as well as a brief update of European Union regulations have been cited in Chap. 11. Specific clinical uses of new radiopharmaceuticals described in Chap. 7 have been presented in Chap. 13. Chapter 14 contains two new sections – theranostics and translational medicine. The use of Xofigo in the treatment of castration-resistant prostate cancer has been added in Chap. 15. Chapter 16 has been revised with new information on adverse reactions from radiopharmaceutical administration to humans. Appendix F is a new addition containing the decay factors of ^{18}F. Appropriate references and additional questions have been added in each chapter.

I would like to thank Dr. Guiyun Wu, staff of Nuclear Medicine at Cleveland Clinic, for providing a few images on neuroimaging. Many thanks are due to Ms. Stephanie Frost, developmental editor, for her skillful editing; Ms. Miranda Finch, editorial assistant, Clinical Medicine; and Ms. Margaret Moore, editor, Clinical Medicine, of Springer, for their sincere help and guidance during the production of the book. I thank Ms. Janet Foltin, former senior editor of Clinical Medicine of Springer, for her help in the initial phase of the book. Thanks and appreciation are due to Mr. Sivakumar Krishnamoorthy of SPi Global for a job well done in the production of the book. Finally, I am ever grateful to Springer Nature for its perpetual support through decades in my publication efforts.

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Contents

The Atom

1

According to Bohr's atomic theory, an atom is composed of a nucleus at the center and one or more electrons rotating around the nucleus along different energy orbits. The nucleus is primarily composed of protons and neutrons, collectively called nucleons. For an atom of a given element, the number of electrons moving around the nucleus equals the number of protons, balancing the electrical charge of the nucleus. Protons are positively charged, electrons are negatively charged and neutrons have no charge. The size of an atom is of the order of 10^{-8} cm (1 angstrom, Å), and that of a nucleus is of the order of 10^{-13} cm (1 fermi, F). The electron configuration of the atom determines the chemical properties of an element, whereas the nuclear structure characterizes the stability and radioactive decay of the nucleus of an atom.

1.1 Electronic Structure of the Atom

The Bohr atomic theory states that electrons in an atom rotate around the nucleus in discrete *energy orbits* or *shells*. These energy shells, referred to as the K shell, L shell, M shell, N shell, and so forth, are stationary and arranged in order of increasing energy. When there is a transition of an electron from an upper orbit to a lower orbit, the energy difference between the two orbits is released as the photon radiation. If the electron is raised from a lower orbit to an upper orbit, the energy difference between the two orbits is absorbed and must be supplied for the transition to occur.

According to the quantum theory, each shell is designated by a quantum number n, called the *principal quantum number*, and denoted by integers, for example, 1 for the K shell, 2 for the L shell, 3 for the M shell, 4 for the N shell, and 5 for the O shell (Table 1.1). Each energy shell is subdivided into *subshells* or *orbitals*, which are designated as s, p, d, f, and so forth. For a principal quantum number n, there are n orbitals in the main shell. These orbitals are assigned *azimuthal quantum numbers*, l, which designate the electron's angular momentum and can assume numerical values of $l = 0, 1, 2, \ldots, n - 1$. Thus for the s orbital $l = 0$, the p orbital $l = 1$, the d

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Table 1.1 Electron configuration in different energy shells

Principal shell	Principal quantumnumber (n)	Orbital (l)	No. of electrons = $2(2l + 1)$ in each orbital	$2n^2$
K	1	$s(0)$	2	2
L	2	$s(0)$	2	
		$p(1)$	6	8
M	3	$s(0)$	2	
		$p(1)$	6	
		$d(2)$	10	18
N	4	$s(0)$	2	
		$p(1)$	6	
		$d(2)$	10	
		$f(3)$	14	32
O	5	$s(0)$	2	
		$p(1)$	6	
		$d(2)$	10	
		$f(3)$	14	
		$g(4)$	18	50

orbital $l = 2$, and so forth. According to the above description, the K shell has one orbital, designated as $1s$; the L shell has two orbitals, designated as $2s$ and $2p$, and so forth. The orientation of the electron's magnetic moment in a magnetic field is described by the *magnetic quantum number*, m. The values of m can be $m = -l, -(l - 1), \ldots, \ldots, (l - 1), l$. Another quantum number, the *spin quantum number*, $s(s = -1/2$ or $+1/2)$, is assigned to each electron in order to specify its rotation about its own axis. Each orbital can accommodate a maximum of $2(2l + 1)$ electrons, and the total number of electrons in a given shell is $2n^2$. Thus, the K shell can contain only 2 electrons, the next L shell 8 electrons, the M shell 18 electrons, the N shell 32 electrons, and the O shell 50 electrons. In atoms, the orbitals are filled in order of increasing energy; that is, the lowest energy orbital is filled in first. However, when d orbitals start filling, there are some exceptions to this rule as seen in the examples of the electron configurations of some elements given below:

$_{11}$Na	$1s^22s^22p^63s^1$
$_{18}$Ar	$1s^22s2^2p^63s^2\ 3p^6$
$_{26}$Fe	$s^22s^22p^63s^2\ 3p^63d^64s^2$
$_{43}$Tc	$1s^22s^22p^63s^2\ 3p^63d^{10}4s^24p^64d^65s^1$
$_{49}$In	$1s^22s^22p^63s^2\ 3p^63d^{10}4s^24p^64d^{10}5s^25p^1$

The electron configuration in different orbitals and shells is illustrated in Table 1.1, and, for example, the structure of $_{28}$Ni is shown in Fig. 1.1.

1.2 Chemical Bonds

The electronic structure of the atom of an element determines to a large degree the chemical properties of the element. The periodic table has been devised to arrange the groups of elements of similar chemical properties in order of increasing atomic number. In the periodic table (Fig. 1.2), nine groups are presented vertically, and

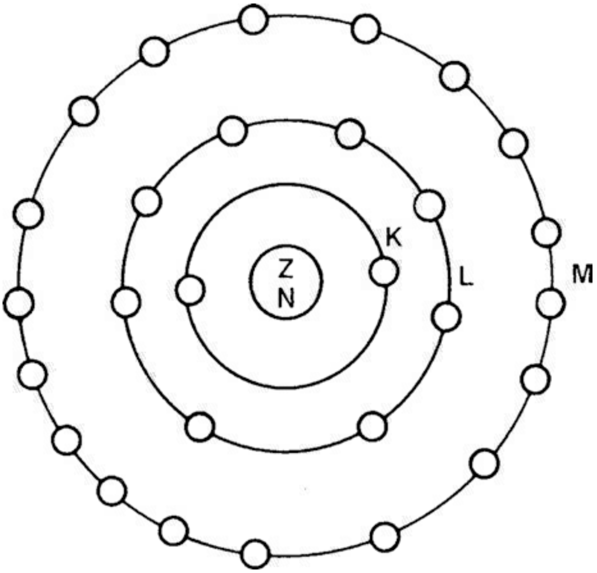
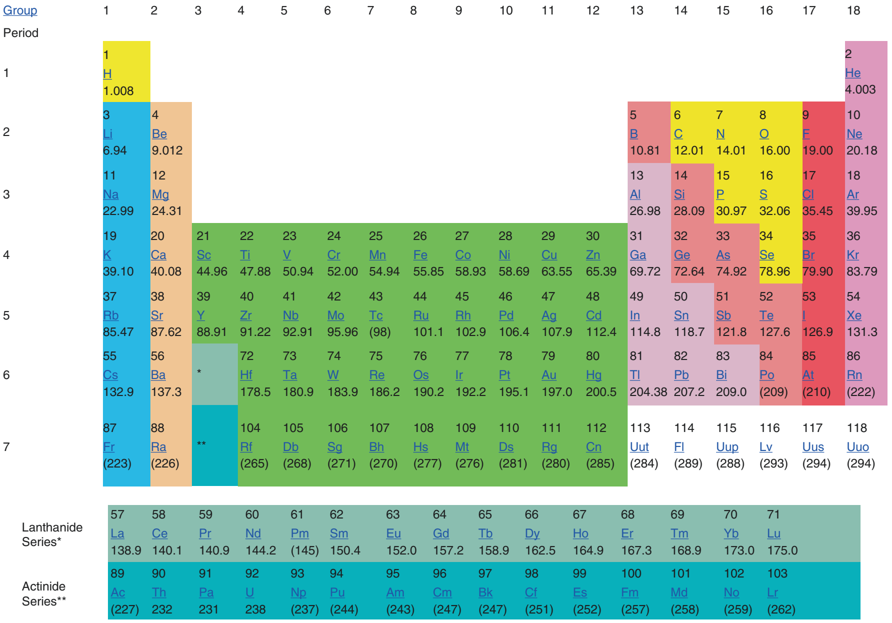


Fig. 1.1 Schematic electron configuration of *K*, *L*, and *M* shells in a nickel atom



Group Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.008																	2 He 4.003
2	3 Li 6.94	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.64	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.79
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.96	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	*	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.5	81 Tl 204.38	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	**	104 Rf (265)	105 Db (268)	106 Sg (271)	107 Bh (270)	108 Hs (277)	109 Mt (276)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (284)	114 Fl (289)	115 Uup (288)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

Lanthanide Series*	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.2	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
Actinide Series**	89 Ac (227)	90 Th 232	91 Pa 231	92 U 238	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Fig. 1.2 Periodic table of elements (Reproduced from http://periodic.lanl.gov/index.shtml under the rule of fair use of material)

seven periods are shown horizontally. Each group contains elements of similar chemical properties, whereas the periods consist of elements having the same number of electron shells but dissimilar chemical properties. As can be seen in Fig. 1.2, group VIIB consists of manganese, technetium, and rhenium, whose chemical properties are very similar. Period 2 contains lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine, and neon, all of which have the *K* shell and *L* shell in common, but are widely different in their chemical behavior.

The *valence* of an element is the tendency of the atom to lose or gain electrons to achieve a stable electron configuration. It is primarily determined by the number of electrons present in the outermost shell, referred to as the *valence shell*. In the most stable and chemically inert elements, such as neon, argon, krypton, and xenon, the valence shell has the electron configuration $\mathrm{n}s^2\mathrm{n}p^6$. Helium, although a noble gas, has the configuration $1s^2$. The electrons in the valence shell are termed the *valence electrons*. To achieve the stable electron configurations $\mathrm{n}s^2\mathrm{n}p^6$, electrons can be added to or given up from the valence shell for chemical bond formation between the atoms of appropriate elements. All chemical bond formation is governed by the *octet rule*, which states that the electronic structure of each atom in a chemical bond assumes $\mathrm{n}s^2\mathrm{n}p^6$ containing eight electrons, with the exception of hydrogen and lithium atoms, which essentially assume the structure $1s^2$. The energy involved in chemical bond formation is of the order of a few electron volts (eV). An electron volt is the energy acquired by an electron accelerated through a potential difference of 1 V. There are three main types of chemical bonds, described below.

1.2.1 Electrovalent or Ionic Bond

An electrovalent or ionic bond is formed by the complete transfer of an electron from the valence shell of one atom to that of another atom. In ionic bonds, two oppositely charged ions are held together in the compound by coulombic forces. The compound NaCl is formed as follows:

$$\mathrm{Na}^+ + \mathrm{Cl}^- \rightarrow \mathrm{NaCl}$$

The sodium atom has the structure $1s^22s^22p^63s^1$, which can spare the $1s^1$ electron to achieve the stable structure of neon, $1s^22s^22p^6$. On the other hand, the chlorine atom has the structure $1s^22s^22p^63s^23p^5$, which is short of one electron in achieving the electronic structure of argon, $1s^22s^22p^63s^23p^6$. Thus, in the formation of NaCl, the sodium atom loses one electron to become Na^+, and the chlorine atom receives the electron to become Cl^-. Both ions are then held by an electrovalent bond. Because of their ionic properties, compounds with electrovalent bonds conduct electricity in the solid state as well as in solution.

1.2.2 Covalent Bond

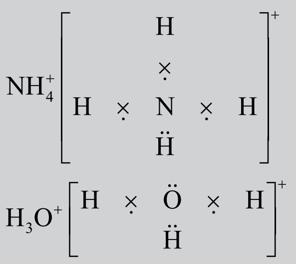
In covalent bonds, each of the two atoms participating in bond formation contributes one electron to the bond. Both electrons are shared equally by each atom and, unlike electrovalent bonds, do not belong exclusively to one atom alone. The shared electrons are localized in the region between the two atoms and the molecules are nonionic. The following molecules are examples of covalent bonds:

H_2	$H \underset{\cdot}{\times} H$
HCl	$H \underset{\cdot}{\times} Cl$
$BeCl_2$	$Cl \underset{\cdot}{\times} Be \underset{\cdot}{\times} Cl$

Here the symbols × and · represent electrons from separate partners in the bond. Because the compounds with covalent bonds are nonionic, they are poor conductors of electricity.

1.2.3 Coordinate Covalent Bond

In a coordinate covalent bond, the pair of electrons required for bond formation is donated by only one atom to another that can accommodate two electrons in octet formation. These bonds are also called semipolar bonds, because only a partial positive charge is generated on the donor atom and a partial negative charge on the acceptor atom. Some donor atoms with a lone pair of electrons are $\dot{\underset{\cdot\cdot}{N}}:$, $\ddot{\underset{\cdot\cdot}{O}}:$, $\ddot{\underset{\cdot\cdot}{S}}:$, and so forth, and they can form coordinate covalent bonds with various metal ions to form metal complexes. The following molecules are examples of coordinate covalent bonds:



$$NH_4^+ \left[\begin{array}{ccccc} & & H & & \\ & & \underset{\cdot}{\times} & & \\ H & \underset{\cdot}{\times} & N & \underset{\cdot}{\times} & H \\ & & \ddot{H} & & \end{array}\right]^+$$

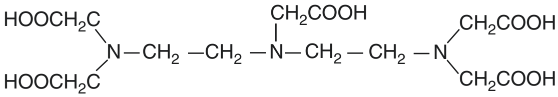
$$H_3O^+ \left[\begin{array}{ccccc} H & \underset{\cdot}{\times} & \ddot{O} & \underset{\cdot}{\times} & H \\ & & \ddot{H} & & \end{array}\right]^+$$

In these examples, nitrogen and oxygen atoms have donated their lone pair of electrons to a hydrogen ion.

1.2.4 Complex Formation

Metal complexes are produced by coordinate covalent bonds that are formed by the electrons donated by the chemical species having a lone pair of electrons. These complexes can be cationic, anionic, or neutral, examples of which are $[Co(NH_3)_6]^{3+}$, $[Fe(CN)_6]^{3-}$, and $[Ni(CO)_4]$, respectively.

Fig. 1.3 Molecular structure of DTPA



Diethylenetriaminepentaacetic Acid (DTPA; MW = 393)

The molecules containing NH_3, −CN, −SH, −COO, $-NH_2$, CO, etc. are called *ligands*, which may be neutral or ionic in structure. The common characteristic of the ligands is that they all possess an unshared pair of electrons that can be donated to a metal ion to form a complex. These ligands are firmly attached to the metal ion, and the number of ligands in a complex is called the *coordination number* of the complex. For example, Co in $[Co(NH_3)_6]^{3+}$ has the coordination number 6.

A single ligand molecule can possess more than one donor atom and can donate more than one pair of electrons in the complex, provided spatial configuration permits. In such cases, more than one coordinate covalent bond is formed in the complex, and the mechanism of bond formation is called *chelation* (from Greek, meaning "clawlike configuration"). Such ligands are called *chelating agents*. Ethylenediaminetetraacetic acid (EDTA) and diethylenetriaminepentaacetic acid (DTPA) are typical examples of chelating agents; the structure of the latter is shown in Fig. 1.3. Donor atoms are nitrogen in the amino groups and oxygen in carboxyl groups. Depending on the number of electron pair donating groups in the molecule, the ligands are named unidentate, bidentate, tridentate, and so on.

The stability of a metal complex is influenced by the sizes of the metal ion and the ligand and the dipole moment of the ligand molecule. The smaller the size of the metal ion and the ligand, the more stable the coordinate covalent bond. Ligands with larger dipole moments form more stable complexes. The stability of a complex is also increased by chelation and the number of electron donor atoms in the chelating agent.

Various ^{99m}Tc-radiopharmaceuticals, such as ^{99m}Tc-DTPA and ^{99m}Tc-dimercaptosuccinate (DMSA), are complexes formed by coordinate covalent bonds between ^{99m}Tc and the chelating compounds. The coordination number of technetium in these complexes varies between 4 and 9 (see in Chap. 6).

1.3 Structure of the Nucleus

The nucleus of an atom is composed of protons and neutrons, collectively called *nucleons*. The characteristics of nucleons and electrons are summarized in Table 1.2. The number of protons in a nucleus is called the *atomic number* of the atom, denoted by *Z*. The number of neutrons is denoted by *N*. The total number of nucleons in a nucleus is referred to as the *mass number*, denoted by *A*. Thus, *A* is equal to $Z + N$. An elemental atom *X* having a mass number *A*, atomic number *Z*, and neutron

Table 1.2 Characteristics of electrons and nucleons

Particle	Charge	Mass (amu)[a]	Mass (kg)	Mass (MeV)[b]
Electron	−1	0.000549	0.9108×10^{-30}	0.511
Proton	+1	1.00728	1.6721×10^{-27}	938.78
Neutron	0	1.00867	1.6744×10^{-27}	939.07

[a] amu, 1 atomic mass unit = 1.66×10^{-27} kg = one twelfth of the mass of ^{12}C
[b] 1 atomic mass unit = 931 MeV

number *N* is represented by $^{A}_{Z}X_{N}$. For example, the stable aluminum nucleus has 13 protons (*Z*) and 14 neutrons (*N*), and therefore its mass number is 27. Thus it is designated as $^{27}_{13}\mathrm{Al}_{14}$. Since all the aluminum atoms have the same atomic number, and the neutron number can be calculated as *A*–*Z*, both the atomic number 13 and the neutron number 14 are omitted from its representation. Thus, the aluminum nucleus is normally designated as ^{27}Al. Alternatively, it is written as Al-27.

Different models have been postulated for the arrangement of the nucleons in a nucleus to explain various experimental observations. According to the Bohr liquid drop model, the nucleus is assumed to be spherical and composed of closely packed nucleons, and particle emission by the nucleus resembles evaporation of molecules from a liquid drop. This theory explains various phenomena, such as nuclear density, binding energy, energetics of particle emission by radioactive nuclei, and fission of heavy nuclei.

In the shell model, nucleons are arranged in discrete energy shells similar to the electron shells of the atom in the Bohr atomic theory. Nuclei containing 2, 8, 20, 50, 82, or 126 protons or neutrons are very stable, and these nucleon numbers are called *magic numbers*.

Nuclei are less stable if they contain an odd number of protons or neutrons, whereas nuclei with even number of protons and neutrons are more stable. The ratio of the number of neutrons to the number of protons (*N*/*Z*) is also an approximate index of the stability of an atom. This ratio equals 1 in the stable nuclei with a lower atomic number, such as $^{12}_{6}\mathrm{C}$, $^{16}_{8}\mathrm{O}$, and $^{14}_{7}\mathrm{N}$, and the ratio increases with the increasing atomic number of the nucleus. For example, it is 1.40 for $^{127}_{53}\mathrm{I}$ and 1.54 for $^{208}_{82}\mathrm{Pb}$. Nuclei with *N*/*Z* different from that of stable nuclei are unstable and decay by β-particle emission or electron capture. The shell model explains various nuclear characteristics such as the angular momentum, magnetic moment, and parity of the nucleus.

According to the classical electrostatic theory, a nucleus should not hold as a single entity because of the electrostatic repulsive forces among the protons in the nucleus. However, its stable existence has been explained by the postulation of a binding force, referred to as the *nuclear force*, which is much stronger than the electrostatic force and binds equally protons and neutrons within the nucleus. The nuclear force exists only in the nucleus and has no influence outside the nucleus.

The short range of nuclear force results in the very small size (~10^{-13} cm) and very high density (~10^{14} g/cm^3) of the nucleus.

The mass *M* of a nucleus is always less than the combined masses of the nucleons *A* in the nucleus. This difference in mass (*M*–*A*) is termed the *mass defect*, which has been used as energy in binding all the nucleons in the nucleus. This energy is the *binding energy* of the nucleus and needs to be supplied to separate all nucleons completely from each other. The binding energy of an individual nucleon has a definite value depending on the shell it occupies; the average energy is approximately equal to the total binding energy divided by the number of nucleons. This energy is about 6–9 MeV and has to be supplied to remove a single nucleon from the nucleus.

1.3.1 Nomenclature

Several nomenclatures are important and need to be mentioned here. An exact nuclear composition including the mass number *A*, atomic number *Z*, and arrangement of nucleons in the nucleus identifies a distinct species, called the *nuclide*. Some nuclides (~288 or so) are stable and naturally occurring of which nearly 54 nuclides have half-life of >50 million and are considered stable. There are about 3400 artificially produced nuclides which are unstable or radioactive and decay by spontaneous fission or emission of α-particle, β-particle, and γ-ray. This group of nuclides are termed *radionuclides*. Nuclides of the same atomic number are called *isotopes* and exhibit the same chemical properties. Examples of oxygen isotopes are ${}^{15}_{8}\mathrm{O}$, ${}^{16}_{8}\mathrm{O}$, ${}^{17}_{8}\mathrm{O}$, and ${}^{18}_{8}\mathrm{O}$. Nuclides having the same number of neutrons but different atomic numbers are called isotones. Examples are ${}^{59}_{26}\mathrm{Fe}$, ${}^{60}_{27}\mathrm{Co}$, and ${}^{62}_{29}\mathrm{Cu}$ each having 33 neutrons. *Isobars* are nuclides with the same number of nucleons, that is, the same mass number, but a different number of protons and neutrons. For example, ${}^{67}_{29}\mathrm{Cu}$, ${}^{67}_{30}\mathrm{Zn}$, ${}^{67}_{31}\mathrm{Ga}$, and ${}^{67}_{32}\mathrm{Ge}$ are isobars having the same mass number 67. Nuclides having the same number of protons and neutrons but differing in energy states and spins are called *isomers*. ${}^{99}\mathrm{Tc}$ and ${}^{99\mathrm{m}}\mathrm{Tc}$ are isomers of the same nuclide. The lifetime of the isomeric states ranges from picoseconds to years, and those with long half-life are represented by “m” as in ${}^{99\mathrm{m}}\mathrm{Tc}$.

The nuclides, both stable and radioactive, are arranged in the form of a chart, referred to as the *chart of the nuclides*, a section of which is presented in Fig. 1.4. Each nuclide is represented by a square containing various information such as the half-life, type and energy of radiations, and so forth, of the radionuclide and the neutron capture cross-section of the stable nuclides (see Chap. 4). The nuclides are arranged in increasing neutron number horizontally and in increasing proton number vertically. Each horizontal bar contains all isotopes of the same element; for example, all silicon isotopes are grouped in the horizontal block designated by the proton number 14. All isotones are grouped vertically; for example, ${}^{26}_{14}\mathrm{Si}$, ${}^{25}_{13}\mathrm{Al}$, and ${}^{24}_{12}\mathrm{Mg}$ are isotones with 12 neutrons and are positioned in the vertical column identified by the neutron number 12. The diagonal nuclides in the chart are isobars, for example, ${}^{27}_{14}\mathrm{Si}$, ${}^{27}_{13}\mathrm{Al}$, and ${}^{27}_{12}\mathrm{Mg}$. The radionuclides ${}^{24}_{13}\mathrm{Al}$ and ${}^{26}_{13}\mathrm{Al}$ each have an isomer.

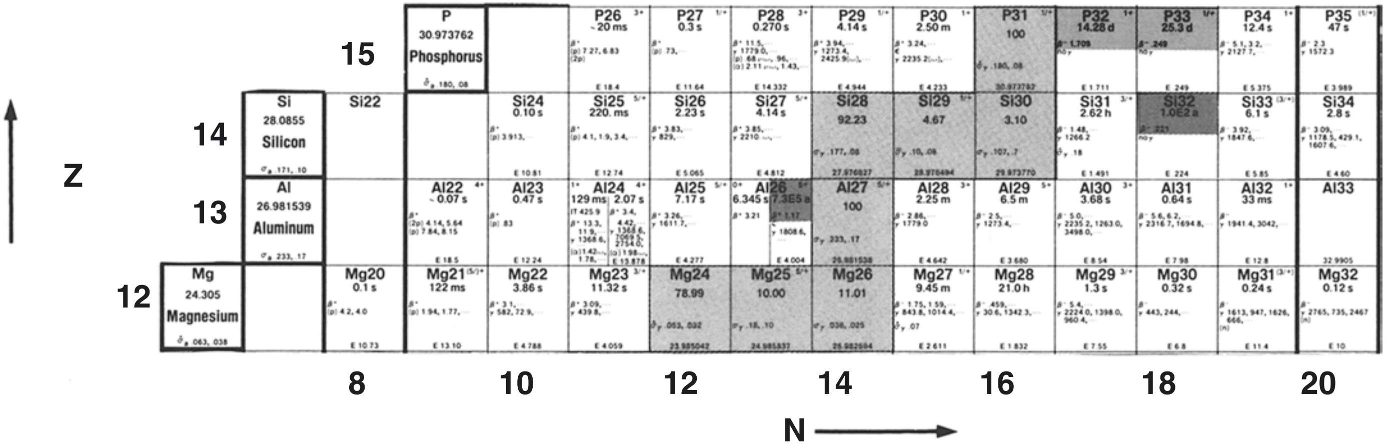


Fig. 1.4 A section of the chart of the nuclides (Courtesy of Knolls Atomic Power Laboratory, Schenectady, New York, operated by the General Electric Company for the United States Department of Energy Naval Reactor Branch)

Questions

1. Describe the basic concept and significance of the Bohr atomic theory.
2. Write the electron configuration of $_{6}$C, $_{17}$Cl, $_{54}$Xe, $_{37}$Rb, $_{43}$Tc, and $_{49}$In.
3. What is the octet rule? Why is it necessary that the electron configuration of the atoms be ns^2np^6 in a chemical bond?
4. The compounds with electrovalent bonds dissociate mostly into ions in water, whereas those with covalent bonds rarely do so. Explain.
5. What are ligands and chelating agents? Define coordination number and explain complex formation.
6. Group the following nuclides into isotopes, isobars, isotones, and isomers: $^{13}_{6}\mathrm{C}$, $^{12}_{6}\mathrm{C}$, $^{14}_{6}\mathrm{C}$, $^{14}_{5}\mathrm{B}$, $^{17}_{7}\mathrm{N}$, $^{17}_{8}\mathrm{O}$, $^{19}_{10}\mathrm{Ne}$, $^{113}_{49}\mathrm{In}$, $^{113}_{49}\mathrm{In}$, $^{57}_{27}\mathrm{Co}$, $^{57}_{26}\mathrm{Fe}$, and $^{57}_{28}\mathrm{Ni}$.
7. Define mass defect and magic number. What does the mass defect account for?
8. Explain why the nuclear force differs from the electrostatic force in the nucleus of an atom.
9. Write the following nuclides in order of increasing stability: $^{88}_{39}\mathrm{Y}$, $^{88}_{38}\mathrm{Sr}$, and $^{87}_{39}\mathrm{Y}$.
10. What are the sizes of an atom and a nucleus? What is responsible for this size difference? What is the difference in magnitude between the chemical and nuclear binding energies?
11. The mass of $^{67}_{31}\mathrm{Ga}$ is 66.9858. (a) Calculate the mass defect in MeV. (b) Calculate the average binding energy in MeV of each nucleon in $^{67}_{31}\mathrm{Ga}$.

Suggested Reading

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Radioactive Decay

2

2.1 Decay of Radionuclides

As mentioned in Chap. 1, some 3700 nuclides have been discovered thus far, and the majority of them are unstable. Unstable nuclei decay by spontaneous fission, α-particle, β-particle, γ-ray emission, or electron capture (EC) in order to achieve stability. The stability of a nuclide is governed by the structural arrangement and binding energy of the nucleons in the nucleus. One criterion of stability is the neutron-to-proton ratio (N/Z) of the stable nuclides; the radionuclides decay to achieve the N/Z of the nearest possible stable nuclide. Radioactive decay by particle emission or electron capture changes the atomic number of the radionuclide, whereas decay by γ-ray emission does not.

Radionuclides may decay by any one or a combination of six processes: spontaneous fission, α-decay, β^--decay, β^+-decay, electron capture, and isomeric transition (IT). In radioactive decay, particle emission or electron capture may be followed by isomeric transition. In all decay processes, the energy, mass, and charge of radionuclides must be conserved. Each of these decay processes is briefly described below.

2.1.1 Spontaneous Fission

Fission is a process in which a heavy nucleus breaks down into two fragments typically in the ratio of 60:40. This process is accompanied by the emission of two or three neutrons with a mean energy of 1.5 MeV and a release of nearly 200-MeV energy, which appears mostly as heat.

Fission in heavy nuclei can occur spontaneously or by bombardment with energetic particles. Spontaneous fission occurs due to dumbbell-shape structure of the heavy nucleus, the middle neck part of which is prone to breakage by molecular vibration causing fission. The probability of spontaneous fission is low and increases with mass number of the heavy nuclei. The half-life for spontaneous fission is 2×10^{17} years for ^{235}U and only 55 days for ^{254}Cf. It should be noted that

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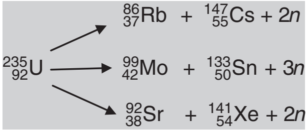
spontaneous fission is an alternative to α-decay or γ-emission. An example of spontaneous fission of ^{235}U is given below.



2.1.2 Alpha Decay (α-Decay)

Usually heavy nuclei such as radon, uranium, neptunium, and so forth decay by α-particle emission. The α-particle is a helium ion with two electrons stripped off the atom and contains two protons and two neutrons bound together in the nucleus. In α-decay, the atomic number of the parent nuclide is therefore reduced by 2 and the mass number by 4. An example of α-decay is

$$^{235}_{92}\mathrm{U} \rightarrow {}^{231}_{90}\mathrm{Th} + \alpha\left[{}^{4}_{2}\mathrm{He}^{+}\right]$$

An α-transition may be followed by β^{-}-emission or γ-ray emission or both. The α-particles are monoenergetic, and their range in matter is very short (on the order of 10^{-6} cm) and is approximately 0.03 mm in body tissue.

2.1.3 Beta Decay (β⁻-Decay)

When a nucleus is "neutron rich" (i.e., has a higher *N*/*Z* ratio compared to the stable nucleus), it decays by β^{-}-particle emission along with an antineutrino. An *antineutrino* ($\overline{\nu}$) is an entity almost without mass and charge and is primarily needed to conserve energy in the decay. In β^{-}-decay, a neutron (n) essentially decays into a proton (p) and a β^{-}-particle; for example,

$$\mathrm{n} \rightarrow \mathrm{p} + \beta^{-} + \overline{\nu}$$

The β^{-}-particle is emitted with variable energy from zero up to the decay energy. The *decay or transition energy* is the difference in energy between the parent and daughter nuclides. An antineutrino carries away the difference between the β^{-}-particle energy and the decay energy. The β^{-}-decay may be followed by γ-ray emission, if the daughter nuclide is in an excited state, and the number of γ-rays emitted depends on the excitation energy. After β^{-}-decay, the atomic number of the daughter nuclide is one more than that of the parent nuclide; however, the mass number remains the same for both nuclides.

Some examples of β^{-}-decay are

$$^{131}_{53}\mathrm{I} \rightarrow {}^{131}_{54}\mathrm{Xe} + \beta^{-} + \overline{\nu}$$

$$^{59}_{26}\mathrm{Fe} \rightarrow ^{59}_{27}\mathrm{Co} + \beta^- + \bar{\nu}$$

$$^{99}_{42}\mathrm{Mo} \rightarrow ^{99m}_{43}\mathrm{Tc} + \beta^- + \bar{\nu}$$

$$^{60}_{27}\mathrm{Co} \rightarrow ^{60}_{28}\mathrm{Ni} + \beta^- + \bar{\nu}$$

The radioactive decay of nuclides is represented schematically by decay schemes, and examples of the decay schemes of ^{131}I and ^{99}Mo are given in Figs. 2.1 and 2.2, respectively.

The β^--particles emitted by radionuclides can produce what is called *bremsstrahlung* by interaction with surrounding medium. Electrons passing through matter are decelerated in the Coulomb field of atomic nuclei, and as a result, the loss in

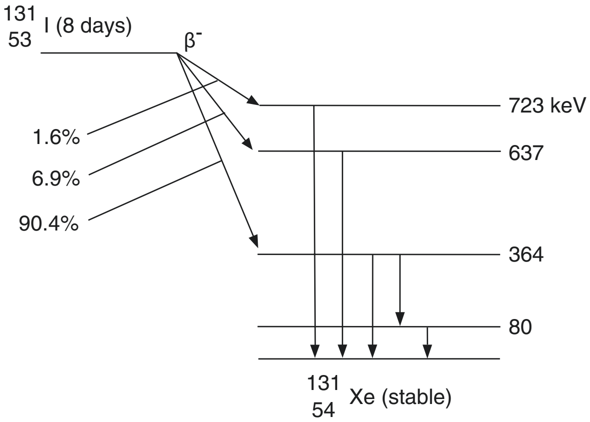


Fig. 2.1 Decay scheme of ^{131}I. Eighty one percent of the total ^{131}I disintegrations decay by 364-keV γ-ray emission. The half-life of ^{131}I is shown in parentheses

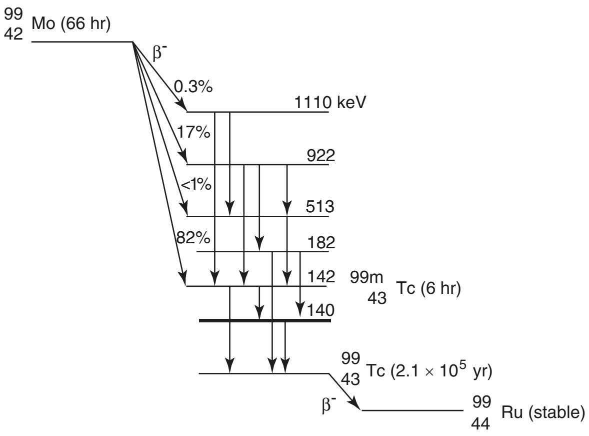


Fig. 2.2 Decay scheme of ^{99}Mo. There is a 2-keV isomeric transition from the 142-keV level to the 140-keV level, which occurs by internal conversion. Approximately 87% of the total ^{99}Mo ultimately decays to ^{99m}Tc, and the remaining 13% decays to ^{99}Tc. (The energy levels are not shown in scale.)

electron energy appears as continuous X-rays. These X-rays are called bremsstrahlung (German for "braking" or "slowing down" radiation) and are used in radiographic procedures. The probability of producing bremsstrahlung increases with increasing electron energy and increasing atomic number of the medium. In tungsten, for example, a 10-MeV electron loses about 50% of its energy by bremsstrahlung, whereas a 100-MeV electron loses more than 90% of its energy by this process.

2.1.4 Positron or β^+-Decay

Nuclei that are "neutron deficient" or "proton rich" (i.e., have an *N*/*Z* ratio less than that of the stable nuclei) can decay by β^+-particle emission accompanied by the emission of a neutrino (v), which is an opposite entity of the antineutrino. After β^+-particle emission, the daughter nuclide has an atomic number that is 1 less than that of the parent. The range of positrons is short in matter. At the end of the path of β^+-particles, positrons combine with electrons and are thus annihilated, each event giving rise to two photons of 511 keV that are emitted in opposite directions. These photons are referred to as *annihilation radiations*.

In β^+-decay, a proton transforms into a neutron by emitting a β^+-particle and a neutrino; for example,

$$\mathrm{p} \rightarrow \mathrm{n} + \beta^+ + v$$

Since a β^+-particle can be emitted with energy between zero and decay energy, the neutrino carries away the difference between decay energy and β^+-energy. We know that a neutron is equivalent to one proton plus an electron. Therefore, in β^+-decay, a mass equivalent of two electrons is created by the conversion of a proton to a neutron, as shown in the above equations, i.e., 1.02 MeV is needed to create these two particles. So positron emission takes place only when the energy difference between the parent and daughter nuclides is equal to or greater than 1.02 MeV. Some examples of β^+-decay are

$${}^{64}_{29}\mathrm{Cu} \rightarrow {}^{64}_{28}\mathrm{Ni} + \beta^+ + v$$

$${}^{18}_{9}\mathrm{F} \rightarrow {}^{18}_{8}\mathrm{O} + \beta^+ + v$$

$${}^{15}_{8}\mathrm{O} \rightarrow {}^{15}_{7}\mathrm{N} + \beta^+ + v$$

$${}^{52}_{26}\mathrm{Fe} \rightarrow {}^{52}_{25}\mathrm{Mn} + \beta^+ + v$$

The decay scheme of ${}^{18}\mathrm{F}$ is presented in Fig. 2.3.

2.1.5 Electron Capture

When a nucleus has a smaller *N*/*Z* ratio compared to the stable nucleus, as an alternative to β^+-decay, it may also decay by the so-called electron capture

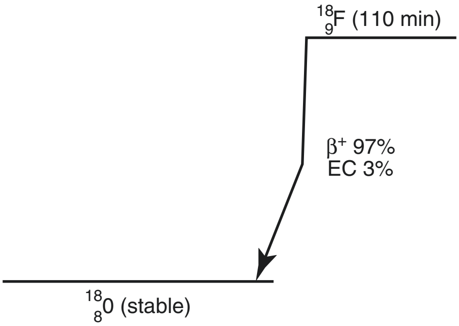


Fig. 2.3 Decay scheme of ^{18}F. The positrons are annihilated in a medium to give rise to two 511-keV γ-rays emitted in opposite directions

process, in which an electron is captured from the extranuclear electron shells, thus transforming a proton into a neutron and emitting a neutrino. For this process to occur, the energy difference between the parent and daughter nuclides is usually, but not necessarily, less than 1.02 MeV. Nuclides having an energy difference greater than 1.02 MeV may also decay by electron capture. The larger the energy difference, the higher the positron decay. The atomic number of the parent is reduced by 1 in this process. Some examples of electron capture decay are

$$^{67}_{31}\mathrm{Ga} + \mathrm{e}^- \rightarrow ^{67}_{30}\mathrm{Zn} + \nu$$

$$^{111}_{49}\mathrm{In} + \mathrm{e}^- \rightarrow ^{111}_{48}\mathrm{Cd} + \nu$$

$$^{57}_{27}\mathrm{Co} + \mathrm{e}^- \rightarrow ^{57}_{26}\mathrm{Fe} + \nu$$

Usually the *K*-shell electrons are captured because of their proximity to the nucleus; the process is then called *K* capture. Thus, in *L* capture, an *L*-shell electron is captured and so on. The vacancy created in the *K* shell after electron capture is filled by the transition of electrons from an upper level (probably the *L* shell and possibly the *M* or *N* shell). The difference in energies of the electron shells will appear as an X-ray that is characteristic of the daughter nucleus. These X-rays are termed characteristic *K* X-rays, *L* X-rays, and so on belonging to the daughter nuclide. The probability of electron capture increases with increasing atomic number, because electron shells in these nuclei are closer to the nucleus. The decay scheme of ^{111}In is given in Fig. 2.4.

2.1.6 Isomeric Transition

A nucleus can remain in several excited energy states above the ground state that are defined by quantum mechanics. All these excited states are referred to as *isomeric states* and decay to the ground state, with a lifetime of fractions of picoseconds to many years. The decay of an upper excited state to a lower excited state is called the