Fourth Edition

Essentials of Nuclear Medicine Physics, Instrumentation, and Radiation Biology

R. A. Powsner | M. R. Palmer E. R. Powsner

WILEY Blackwell

Essentials of Nuclear Medicine Physics, Instrumentation, and Radiation Biology

Dedication

In memory of my parents, Rhoda and Edward Powsner, for all of their love, support, and guidance throughout the years.

R.A.P.

To the memory of my father, Ivan George Palmer, who would have said, "Well, what's all this then?" M.R.P.

Essentials of Nuclear Medicine Physics, Instrumentation, and Radiation Biology

FOURTH EDITION

Rachel A. Powsner, MD

Clinical Professor of Radiology Boston University School of Medicine Director, Division of Nuclear Medicine Department of Radiology Boston Veterans Administration Healthcare System Boston, MA, USA

Matthew R. Palmer, PhD

Director of Medical Imaging Physics Department of Radiology Beth Israel Deaconess Medical Center Assistant Professor of Radiology Harvard Medical School Boston, MA, USA

Edward R. Powsner, MD*

Former Chief, Nuclear Medicine Service Veterans Administration Hospital Allen Park, MI, USA Former Professor and Associate Chairman Department of Pathology Michigan State University East Lansing, MI, USA *Deceased

WILEY Blackwell

This edition first published 2022 © 2022 John Wiley & Sons Ltd

Edition History John Wiley & Sons, Ltd (3e, 2013)

All rights reserved. No part of this publication may be reproduced, stored in a retrieval system, or transmitted, in any form or by any means, electronic, mechanical, photocopying, recording or otherwise, except as permitted by law. Advice on how to obtain permission to reuse material from this title is available at http://www.wiley.com/go/permissions.

The right of Rachel A. Powsner and Matthew R. Palmer to be identified as the authors of this work has been asserted in accordance with law.

Registered Office(s)

John Wiley & Sons, Inc., 111 River Street, Hoboken, NJ 07030, USA John Wiley & Sons Ltd, The Atrium, Southern Gate, Chichester, West Sussex, PO19 8SQ, UK

Editorial Office

9600 Garsington Road, Oxford, OX4 2DQ, UK

For details of our global editorial offices, customer services, and more information about Wiley products visit us at www.wiley.com.

Wiley also publishes its books in a variety of electronic formats and by print-on-demand. Some content that appears in standard print versions of this book may not be available in other formats.

Limit of Liability/Disclaimer of Warranty

The contents of this work are intended to further general scientific research, understanding, and discussion only and are not intended and should not be relied upon as recommending or promoting scientific method, diagnosis, or treatment by physicians for any particular patient. In view of ongoing research, equipment modifications, changes in governmental regulations, and the constant flow of information relating to the use of medicines, equipment, and devices, the reader is urged to review and evaluate the information provided in the package insert or instructions for each medicine, equipment, or device for, among other things, any changes in the instructions or indication of usage and for added warnings and precautions. While the publisher and authors have used their best efforts in preparing this work, they make no representations or warranties with respect to the accuracy or completeness of the contents of this work and specifically disclaim all warranties, including without limitation any implied warranties of merchantability or fitness for a particular purpose. No warranty may be created or extended by sales representatives, written sales materials or promotional statements for this work. The fact that an organization, website, or product is referred to in this work as a citation and/or potential source of further information does not mean that the publisher and authors endorse the information or services the organization, website, or product may provide or recommendations it may make. This work is sold with the understanding that the publisher is not engaged in rendering professional services. The advice and strategies contained herein may not be suitable for your situation. You should consult with a specialist where appropriate. Further, readers should be aware that websites listed in this work may have changed or disappeared between when this work was written and when it is read. Neither the publisher nor authors shall be liable for any loss of profit or any other commercial damages, including but not limited to special, incidental, consequential, or other damages.

Library of Congress Cataloging-in-Publication Data

Names: Powsner, Rachel A. author. | Palmer, Matthew R., 1958- author. | Powsner, Edward R., 1926- author. Title: Essentials of nuclear medicine physics, instrumentation, and radiation biology / Rachel A. Powsner, Matthew R. Palmer, Edward R. Powsner. Other titles: Essentials of nuclear medicine physics Description: Fourth edition. | Hoboken, NJ : Wiley-Blackwell, 2022. | Preceded by: Essentials of nuclear medicine physics and instrumentation / Rachel A. Powsner, Matthew R. Palmer, Edward R. Powsner. Third edition. Chichester, West Sussex, UK : Wiley-Blackwell, 2013. | Includes bibliographical references and index. Identifiers: LCCN 2021032438 (print) | LCCN 2021032439 (ebook) | ISBN 9781119620990 (paperback) | ISBN 9781119621003 (adobe pdf) | ISBN 9781119621010 (epub) Subjects: MESH: Nuclear Medicine | Nuclear Medicine-instrumentation | Nuclear Physics | Radiation, Ionizing | Radiation Effects | Radioactive Hazard Release-prevention & control Classification: LCC R896.7 (print) | LCC R896.7 (ebook) | NLM WN 440 | DDC 616.07/575-dc23 LC record available at https://lccn.loc.gov/2021032438 LC ebook record available at https://lccn.loc.gov/2021032439 Cover Design: Wiley Cover Image: © Courtesy of Rachel A. Powsner

Set in 9/12pt PhotinaMTStd by Straive, Pondicherry, India

Contents

Preface, vii

Acknowledgments, ix

- 1. Basic Nuclear Medicine Physics, 1
- 2. Interaction of Radiation with Matter, 21
- 3. Formation of Radionuclides, 33
- 4. Nonscintillation Detectors, 42
- 5. Scintillation Detectors, 62
- 6. Imaging Instrumentation, 74
- 7. Single-photon Emission Computed Tomography (SPECT), 92
- 8. Positron Emission Tomography (PET), 103
- 9. X-ray Computed Tomography (CT), 117
- 10. Magnetic Resonance Imaging (MRI), 128
- 11. Hybrid Imaging Systems: PET-CT, SPECT-CT, and PET-MRI, 145
- 12. Image Reconstruction, Processing, and Display, 151

- 13. Information Technology, 181
- 14. Quality Control, 189
- 15. Radiation Biology, 209
- 16. Radiation Dosimetry, 222
- 17. Radiation Safety, 231
- 18. Radiopharmaceutical Therapy, 239
- 19. Management of Nuclear Event Casualties, 256
- Appendix A: Common Nuclides, 269
- Appendix B: Major Dosimetry for Common Pharmaceuticals, 271
- Appendix C: Guide to Nuclear Regulatory Commission (NRC) Publications, 273
- Appendix D: Recommended Reading by Topic, 276

Index, 278

Preface

After many years of postgraduate training, many physicians have forgotten some (or most) of their undergraduate and high school physics and may find submersion into nuclear physics somewhat daunting. This book begins with a very basic introduction to nuclear physics and the interactions of radiation and matter. It then proceeds with a discussion of the methods for production of nuclides and the instrumentation used for dose measurement, surveying radioactivity, and imaging. The imaging section has been expanded to cover MRI and PET-MRI in addition to SPECT, PET, and PET-CT. The final chapters of the book focus on radiation biology, radiation safety, radiopharmaceutical therapy, and radiation accidents.

Numerous illustrations are included. They are highly schematic and are designed to illustrate concepts rather than represent scale models of their subjects. This text is intended for radiology residents, cardiology fellows, nuclear medicine residents and fellows, nuclear medicine technology students, and others interested in an introduction to concepts in nuclear medicine physics and instrumentation.

> Rachel A. Powsner Matthew R. Palmer

Acknowledgments

The authors would like to thank the following experts for their help with this edition: Larry Panych, PhD for his in-depth critique and suggestions for the MRI chapter, Annick Van Den Abbeele, MD, FACR, and Kun Huang, MD, for their review and corrections of the chapters on radiopharmaceutical therapy and radiation biology, respectively, and Chad Smith, PhD for his review of radiation safety and dosimetry. We are grateful for the assistance of Arda Konik, PhD for his guidance on interpretation of GE PET QC images as well as Mr. Brandon Clifton and Jeanette Y. Kleier, RT (R)(N), CNMT for help with understanding Philips PET and SPECT QC images and Stephen Baker for his aid with Siemens PET QC images. In addition, Mr. Mitchell Kruzel from Nuclear Fields provided helpful information on current utility of collimator designs.

Since this edition is built upon information included in prior editions, the authors would like to thank the following individuals for their help on the third edition: Anupma Jati, MD for critiquing the CT dosimetry section of the text, Gary Murphy, RT, for help with CT QC questions and Kandace Craft, RTN and Chris Lindsey, FSE for information about PET-CT. P. Satish Nair, PhD generously made comments on the dosimetry chapter. David Drum, MD answered numerous questions about radiation safety and dosimetry. J. Anthony Parker, MD was helpful on the topic of cancer induction from low-dose ionizing radiation as well as serving as a reference for an assortment of other specific questions. For the second edition the following individuals were most generous with information: Stephen Moore, PhD on the topic of SPECT processing including iterative reconstruction, Fred Fahey, DSc on PET instrumentation, and Robert Zimmerman, MSEE on gamma camera quality control and the physics of crystal scintillators. In addition, Dr. Frank Masse generously reviewed the material on radiation accidents and Mark Walsh, CHP critiqued the radiation safety text.

We would also like to thank the following individuals for their help in reviewing portions of the first edition during its preparation: David Rockwell, MD, Maura Dineen-Burton, CNMT, Dipa Patel, MD, Alfonse Taghian, MD, Hernan Jara, PhD, Susan Gussenhoven, PhD, John Shaw, MS, Michael Squillante, PhD, Kevin Buckley, CHP, Jayne Caruso, Victor Lee, MD, Toby Wroblicka, MD, Dan Winder, MD, Dennis Atkinson, MD, and Inna Gazit, MD. Thanks to Peter Shomphe, ARRT, CNMT, Bob Dann, PhD, and Laura Partriquin, MD for wading through the manuscript in its entirety. We greatly appreciate the patience shown at that time by Robert Zimmerman, MSEE, Kevin Buckley, CHP, John Widman, PhD, CHP, Peter Waer, PhD, Stephen Moore, PhD, Bill Worstell, PhD, and Hernan Jara, PhD while answering our numerous questions. Thanks to Delia Edwards, Milda Pitter, and Paul Guidone, MD for taking time to pose as models. The authors would also like to thank Rhoda M. Powsner, MD for her assistance in reviewing sections of the text and for proofreading the review questions.

CHAPTER 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 from a cat's fur to a hand. After petting, the cat will have a net positive charge and the person a net negative charge (Figure 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, the smallest fragment with the properties of salt would be a single salt molecule (Figure 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 atoms of 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 magnitudes greater than electrical forces.

Elements

There are more than 100 species of atoms. These species are referred to as **elements**. Most of the

Essentials of Nuclear Medicine Physics, Instrumentation, and Radiation Biology, Fourth Edition.

Rachel A. Powsner, Matthew R. Palmer, and Edward R. Powsner.

^{© 2022} John Wiley & Sons Ltd. Published 2022 by John Wiley & Sons Ltd.



Figure 1.2 The NaCl molecule is the smallest unit of salt that retains the characteristics of salt.

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

1 H																	² He
Li	Be											5 B	⁶ C	N	⁸ 0	F	Ne
Na 11	Mg											13 Al	Si	15 P	16 S	17 Cl	Ar
19 K	Ca	21 Sc	²² Ti	23 V	Cr	²⁵ Mn	Fe	27 Co	28 Ni	Cu ²⁹	³⁰ Zn	Ga	Ge	As	Se	Br	³⁶ Kr
Rb	Sr	39 Y	Zr ⁴⁰	⁴¹ Nb	42 Mo	⁴³ Tc	Ru	⁴⁵ Rh	Pd	A7 Ag	⁴⁸ Cd	49 In	Sn 50	Sb	Te	53	54 Xe
Cs	Ba	Lu 71	⁷² Hf	⁷³ Ta	74 W	⁷⁵ Re	Os	⁷⁷ Ir	78 Pt	⁷⁹ Au	⁸⁰ Hg	81 TI	82 Pb	⁸³ Bi	⁸⁴ Po	At	⁸⁶ Rn
Fr	⁸⁸ Ra	¹⁰³ Lr	¹⁰⁴ Rf	¹⁰⁵ Ha	¹⁰⁶ Sg	107 Ns	108 Hs	¹⁰⁹ Mt	110 Ds	111 Rg	112 Ds	113 ?	114 ?	115 ?	116 ?	117 ?	118 ?
																	ß
		57 La	Ce	⁵⁹ Pr	⁶⁰ Nd	61 Pm	Sm	Eu	G4 Gd	⁶⁵ Tb	⁶⁶ Dy	67 Ho	Er	⁶⁹ Tm	70 Yb		
		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		

Figure 1.3 Periodic table.

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 stable state are shown in white boxes; those that occur only in a radioactive form are shown in gray boxes. The number appearing above each element's abbreviation is referred to as the atomic number, which will be discussed later in this chapter.

The elements in the periodic table are arranged in columns (called groups) and rows (called periods). In general, elements within groups demonstrate similar properties. This is because elements in a group often have similar numbers of electrons in their outer shell; outer shell electron configurations are more important in determining how an atom interacts with other elemental atoms. The lanthanides and actinides are special groups of elements, conventionally shown in rows, separated and placed below the table. These two groups have the same number of outer-shell electrons and share many common properties.

Atomic structure

Atoms initially were thought of as no more than small pieces of matter. Our understanding that



Figure 1.4 Flat atom. The standard two-dimensional drawing of atomic structure.

they have an inner structure has its 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** (Figure 1.4).

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 the nucleus has a positive charge. 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.

Electron shells and binding energy: 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 (Figure 1.5). 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 most significant characteristic of a shell is its energy level. The "closer" an electron is to the nucleus, the more tightly it is bound to the 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. Conversely it is also the amount of kinetic (motion) energy an electron acquires if it "falls" through a potential difference of one volt. It is a very 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 1.5 An electron shell is a representation of the energy level associated with an atomic electron.



Figure 1.6 K, L, and M electron shells.

Quantum numbers: 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 shell. The first three shells (K, L, and M) are depicted in Figure 1.6. 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 (azimuthal), third (magnetic), and fourth (spin) quantum numbers refer to other physical properties of the electron. Each electron within an atom has a unique combination of the four quantum numbers.

The maximum number of electrons associated with each energy shell is $2n^2$, where n is the shell number. The first shell (the K shell) can contain a maximum of two electrons, the second shell (the L shell) can contain a maximum of eight electrons, the third shell (the M shell) can contain a maximum of 18 electrons, and so on.

Representation of electron distribution: Most of the diagrams (for example Figure 1.6) in this chapter reflect what is referred to as the **Bohr model** of the atom and as such all electrons within each shell are depicted as moving along the surface of a sphere, each shell represented as one such sphere with a distinct radial distance from a centrally located nucleus. The radius of these spheres increases with principal quantum number. This model of the atom

is frequently used for teaching purposes because the radial distance of an electron from the nucleus is used to depict with how tightly bound it is to the atom—the closest electrons being most tightly bound.

A more accurate quantum mechanical description of electron distribution uses a sequence of **orbitals**. Orbitals are mathematical functions that describe the probability of finding an electron in a region of space near the nucleus. For each principal quantum number (each shell) there is a spherical orbital denoted by the principal quantum number followed by the letter "s". This orbital contains two electrons (Figure 1.7a). This is the only orbital for the K shell which contains at maximum two electrons and this orbital is called the **1s orbital**. The neutral atom with a full K shell is the helium atom.

The next shell, the L shell (n = 2), also has a spherical orbital, denoted 2s (also depicted as Figure 1.7a) which contains two electrons, as well as three **sub-orbitals**, denoted $2p_x$, $2p_y$, $2p_z$. Each sub-orbital has a shape like a dumb-bell or three-dimensional figure eight (see Figure 1.7b). The three sub-orbitals are oriented along three orthogonal axes as shown in Figure 1.7c. Each sub-orbital is filled by two electrons and the neutral atom with completely filled orbitals for n = 1 and n = 2 is Neon.

For the higher order orbitals, n > 2, the suborbitals associated with higher azimuthal quantum number become even more complicated in structure and will not be discussed here.

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.



Figure 1.7 Electron orbitals and sub-orbitals. (a) s orbital, (b) p suborbital, (c) p suborbitals, p, p,, p,,

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 within 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 X-rays and electrons (this will be discussed in more detail later in this chapter in the section on internal conversion).

Nucleus

Like the atom itself, the atomic nucleus also has an inner structure (Figure 1.8). 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 have a much greater mass than electrons. Table 1.1 reviews the properties of the various subatomic particles.

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 1.9).



Figure 1.8 The nucleus of an atom is composed of protons and neutrons.

Symbol	Mass ^a	Charge
N	1839	None
Р	1836	Positive (+)
e-	1	Negative (-)
B-	1	Negative (-)
β+	1	Positive (+)
γ	None	None
X-ray	None	None
ν	Near zero	None
\overline{V}	Near zero	None
	Symbol N P e^- B- β + γ X-ray ν $\overline{\nu}$	SymbolMassaN1839P1836 e^- 1B-1 β +1 γ NoneX-rayNone ν Near zero $\overline{\nu}$ Near zero

 Table 1.1 Properties of the subatomic particles

^a Relative to an electron.

 $^{\rm b}$ 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.

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 2 million



Figure 1.9 Nuclear binding force is strong enough to overcome the electrical repulsion between the positively charged protons.

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 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 $E = mc^2$.

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



Figure 1.10 Standard atomic notation.

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 (\mathbf{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. The chemical symbol for each element can be expanded to include these three numbers (Figure 1.10).

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 1.11.



Figure 1.11 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. Stable nuclear configurations are shaded gray, radioactive configurations are white. (Adapted from Brucer, M. Trilinear Chart of the Nuclides, Mallinkrodt Inc, 1979.)

Nuclide is a general term for the composition of a nucleus and includes isotopes, isotones, isobars, and other nuclear configurations.

The stable nucleus: Not all elements have stable isotopes; they do exist for most of the light and midweight elements, those with atomic numbers (number of protons) up to and including bismuth (Z = 83). However, there are no stable isotopes of technetium (Z = 43), promethium (Z = 61), or for all elements with atomic numbers higher than 83. Prominent examples are radium (Z = 88) and uranium (Z = 92), which are found naturally as a mix of isotopic forms that are all radioactive.

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, stability is more likely when

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** (Figure 1.12).

Stability

Strictly speaking, stability is a relative term. We call a nuclide stable when its half-life is so long as to be practically immeasurable—say greater than 100 years. An isotope of potassium, ⁴⁰K for example, which makes up about 1% of the potassium found in nature is considered stable but actually has a half-life of 10^9 years.



Figure 1.12 Combinations of neutrons and protons that can coexist in a stable nuclear configuration all lie within the gray shaded regions.

Radioactivity

The unstable nucleus and radioactive decay

A nucleus which is not in its stable state will adjust itself until it is more stable either by ejecting 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 an **alpha particle** (Figure 1.13).

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 are emitted (Figure 1.14).

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 a **beta particle** (Figure 1.15). 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.











Figure 1.15 β – (negatron) decay.

Neutrino ν Φ

Figure 1.16 β + (positron) decay.

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. In general, these proton rich nuclei decay by a combination of these two processes.

Positron decay: A proton can be converted into a neutron and a **positron**, which is an electron with a positive, instead of negative, charge (Figure 1.16).

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 Chapter 8, Figure 8.1). This is why it is considered an anti-electron. 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, each with an energy of 511 keV or a total of 1.022 MeV. Following ejection of a positron from a nucleus the atom must also shed an orbital electron to keep the overall charge of the atom neutral. So, in essence, the atom is losing the mass equivalent of two electrons (remember positrons are basically positively charged electrons). Positron emission will only occur when the difference in mass between the parent (original) and daughter atoms is at minimum the mass of two electrons, which, as we will see in Chapter 2, Figure 2.12 is equal to 1.02 MeV of energy.

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 1.17, a plot of the number of beta particles emitted at each energy from zero to the maximum energy released in the decay. $E_{\rm \beta max}$ is the maximum possible energy that a beta particles can receive during beta decay of any atom, \overline{E}_{β} 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 or

$$\overline{E_{\beta}} \cong \frac{1}{3} E_{\beta \max}$$



Figure 1.17 Beta emissions (both β^- and β^+) are ejected from the nucleus with energies between 0 and their maximum possible energy ($E_{\alpha_{max}}$). The average energy (\overline{E}_{β}) is equal to approximately one third of the maximum energy.

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 (Figure 1.18). 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 outershell electron to the inner shell isemitted as an X-ray.



Figure 1.18 Electron capture.

Appropriate numbers of nucleons, but too much energy

Isomeric transition: Following alpha and beta decay and electron capture, the nucleus has a more favorable physical configuration of nucleons but usually contains an excess of energy. The nucleus is said to be in an excited state when the energy of the nucleus is greater than its resting level. This excess energy is shed by **isomeric transition**. This may occur by either or both of two competing reactions: gamma emission or internal conversion. Most isomeric transitions occur as a combination of these two reactions.

Gamma emission: In this process, excess nuclear energy is emitted as a gamma ray (Figure 1.19). 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 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. 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). The energy released as a result of the "fall" of an outer-shell electron to an inner shell is emitted as an X-ray (Figure 1.20a) or as a free electron, an Auger electron (Figure 1.20b). The emitted X-ray is called a characteristic X-ray because its energy always equals the difference in binding energies between the electron shells.

Decay notation

Decay from an unstable parent nuclide to a more stable daughter nuclide 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 (Figure 1.21). The uppermost level of the schematic is the state with 14 Essentials of Nuclear Medicine Physics, Instrumentation, and Radiation Biology



Figure 1.19 Isomeric transition. Excess nuclear energy is carried off as a gamma ray.



Figure 1.20 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 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. A dogleg with a "Z" denotes alpha decay. 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 1.22 depicts specific decay schemes for ^{99m}Tc, ¹¹¹In, ¹³¹I, and ²²⁶Ra (this is not the isotope used for treatment in nuclear medicine, ²²³Ra, which will be discussed in detail in Chapter 18). 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



Figure 1.21 Decay schematics.

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 minutes and most of the bag will be done in 4 minutes. 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 A(t), the activity remaining, is shown in Figure 1.23.

This curve, and therefore the average behavior of the sample of radioactivity, can be described by the **decay equation:**

$$A(t) = A(0)e^{-0.693t/T_{1/2}}$$

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 ($T_{1/2}$):

$$\lambda = 0.693/T_{1/2}$$

The decay equation can be rewritten as

$$A(t) = A(0)e^{-\lambda t}$$

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 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 1.2.

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.



Figure 1.22 Decay schemes showing principal transitions for technetium-99m, indium-111, iodine-131 and radium-226. Energy levels are rounded to three significant figures.



Figure 1.23 Decay curve. Note the progressive replacement of radioactive atoms (parent) by relatively more stable atoms (daughter) as shown schematically in each box.

Table 1.2 Conversion values for units of radioac	tivity
--	--------

One curie (Ci) =		$1 \times 10^3 mCi$	$1 \times 10^6 \mu Ci$	$37 \times 10^9 Bq$	$37 \times 10^3 MBq$
One millicurie (mCi) =	1×10^{-3} Ci		$1 \times 10^3 \mu Ci$	$37 \times 10^{6} \mathrm{Bq}$	37 MBq
One microcurie (μ Ci) =	1×10^{-6} Ci	$1 \times 10^{-3} \mathrm{mCi}$		37×10^3 Bq	$37 \times 10^{-3} \text{ MBq}$
One bequerel (Bq)* =	27×10^{-12} Ci	27 × 10 ⁻⁹ mCi	$27 \times 10^{-6} \mu\text{Ci}$		$1 \times 10^{-6} \mathrm{MBq}$
One megabequerel (MBq) =	27×10^{-6} Ci	27×10^{-3} mCi	27 μCi	$1 \times 10^{6} \mathrm{Bq}$	

* One bequerel = 1 decay per second.

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

Questions

- 1. The chemical interactions between various elements are mainly determined by:
 - (a) The number of protons.
 - (b) The number of neutrons.
 - (c) The number of electrons in the outermost shell.
 - (d) The number of protons minus the number of electrons.
- 2. 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 the same element (equal *Z*) with different numbers of protons.
 - (c) Atoms of different elements (different *Z*) with equal numbers of neutrons (*N*).
 - (d) Atoms of different elements with equal atomic mass (*A*).
- 3. 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 lighter elements nuclear stability is achieved with equal numbers of protons and neutrons; for heavier elements the number of neutrons exceeds the number of protons.
- 4. For internal conversion to occur, the excess energy of the excited nucleus must equal or exceed:
 - (a) 511 keV.
 - (b) 1.022 MeV.
 - (c) The internal conversion coefficient.
 - (d) The average energy of the Auger electrons.
 - (e) The binding energy of the emitted electron.
- 5. For an atom undergoing beta decay, the average energy of the emitted beta particles is approximately:
 - (a) 511 keV.
 - (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.
- 6. 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 mCi, remains? (The $T_{1/2}$ of 99m Tc is 6 hours. The constant e = 2.718).
- 7. Rank the following binding energies from greatest to least:
 - (1) Electron binding energy for outer shell electrons.
 - (2) Nuclear binding energy.
 - (3) Electron binding energy for inner shell electrons.
- 8. True or false: The term metastable refers to an intermediate state of nuclear decay lasting longer than 10^{-12} seconds prior to undergoing isomeric transition.
- 9. Which of the following is true regarding beta decay of a specific radioisotope:
 - (1) The energy of the emitted beta particle is always the same.
 - (2) The energy of the emitted antineutrino is always the same.
 - (3) The summed energy of the emitted beta particle and antineutrino is always the same.
- 10. Which unit of measurement for radioactivity is defined as one radioactive decay per second?
 - (1) Bequerel.
 - (2) Millicurie.
 - (3) Megabequerel.
- 11. 10 mCi equals how many MBq?
 - (1) 2.7 MBq.
 - (2) 37 MBq.
 - (3) 270 MBq.
 - (4) 370 MBq.
- 12. Lighter nuclides (Z < 83) with an excess of neutrons tend to decay by:
 - (1) Gamma emission.
 - (2) Beta minus decay.
 - (3) Isomeric transition.
 - (4) Positron emission.
 - (5) Alpha emission.