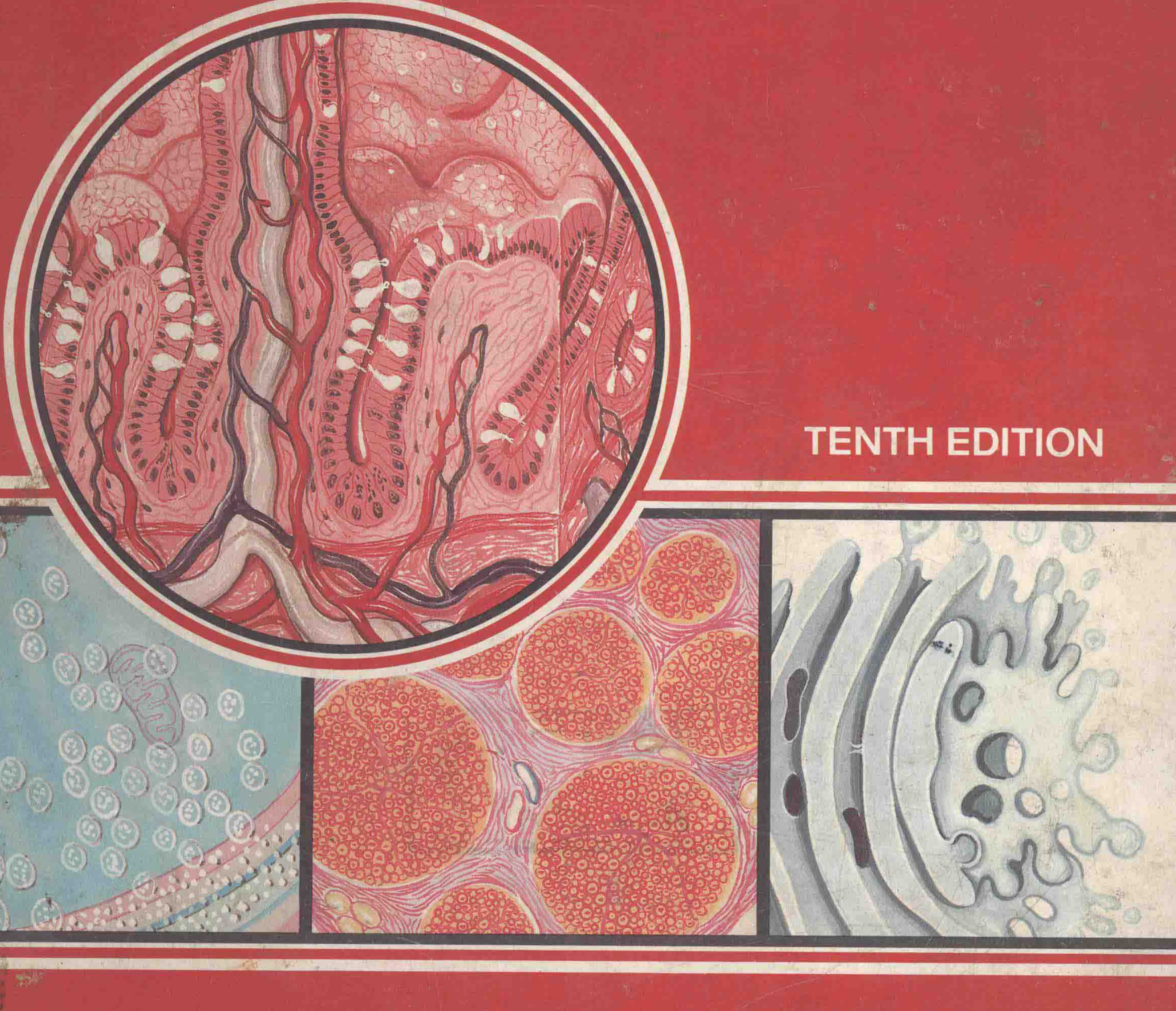
TEXTBOOK OF

Anatomy and Physiology

ANTHONY AND THIBODEAU



TENTH EDITION

TEXTBOOK OF

Anatomy and Physiology

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Preface

A textbook should above all teach. This is its reason for being. To achieve this goal the tenth edition of *Textbook of Anatomy and Physiology* presents basic, up-to-date, accurate information about the body in a concise, clear, readable style. We have tried to create a textbook that students will want to pick up, one they will not put down with a sigh, a book they will enjoy because it stimulates them, and from which, therefore, they learn eagerly and easily. We have tried to make this a book that will exhilarate instructors also and that will help make their task easier.

The tenth edition of *Textbook of Anatomy and Physiology* is designed for college-level introductory courses in human anatomy and physiology. It should serve students in liberal arts, nursing, and other health-related programs equally well.

Begin your examination of this book by noting features of its organization. Outline surveys of the chapters that compose each unit appear immediately below the unit titles. Thus students are alerted, almost at a glance, to chapter contents. To help reinforce learning, summarizing outlines and thought-provoking questions complete each chapter. Summarizing tables and diagrams throughout the text serve a similar purpose.

But probably most important of all for understanding and learning the material in this book—and certainly for enjoying it—are its many illustrations. One of our most illustrious contemporary medical artists, Ernest W. Beck, has created them. To see whether you share our excitement over their unsurpassed clarity, accuracy, and beauty, please look at some of the bone and joint illustrations in Chapters 5 and 6 and at Figs. 7-7, 7-17, 8-9, 14-1, 16-6, 25-14, and 25-17. In addition to the numerous artist's drawings, this revision includes many electron micrographs.

A great many changes in content have been made in this edition. Outdated or inaccurate material has been deleted, and significant new material has been added. Examples: a new section on basic chemistry in the first chapter, new chapters on articulations and the immune system, and a new four-chapter unit on transportation. The latter includes new material about indirect methods of measuring blood volume, anemias, platelet functions, autonomic control of the heart, the cardiac cycle, and vasomotor control mechanisms. Some other changes you will find in this book are new concepts about fever, an expanded discussion of renal function, new information about prostaglandins, pancreatic polypeptide, thymosin, recombinant DNA, embryology, antenatal diagnosis and treatment, and many more clinical applications.

The design of this book excels, we think, that of all previous editions. For example, look at the table starting on p. 120. Notice how the small red boxes attract your attention and cause your eye to automatically pick out major bone markings. As another example, glance at p. 676, and observe how quickly its design enables you to rapidly select supplementary readings for any chapter.

A host of people have produced this book. We wish to thank each and every one of them, and to thank, too, Dr. Kathleen Prezbindowski for revising the chapters on digestion and metabolism. With warmest appreciation, we acknowledge a deep indebtedness to the artist, Ernest W. Beck.

Catherine Parker Anthony
Gary A. Thibodeau

TEXTBOOK OF

Anatomy and Physiology

Contents

unit one

The body as a whole

chapter 1

Organization of the body

What a piece of work is man! how noble in reason! how infinite in faculty! in form and moving how express and admirable! in action how like an angel! in apprehension how like a god! the beauty of the world! the paragon of animals!

Hamlet, in *Hamlet, Prince of Denmark,* Act II, Scene 2

This is a book about that most wondrous of all structures—the human body. It presents information about both the body's structure and its functions. The science of the structure of organisms is called *anatomy* (from the Greek *ana*, up, and *temnein*, to cut). *Physiology* (from the Greek *physis*, nature, and *logos*, discourse) is the science of the functions of organisms. Many facets of anatomy and physiology are based on another science, that of chemistry. Therefore we shall begin this introductory chapter with some basic concepts of chemistry. If you are already familiar with these, you may wish to skim through this first section—or perhaps skip it entirely. The remaining pages of the chapter introduce some important ideas about the structure and functions of the body as a whole. They also contain definitions of terms used in describing the body's structures.

Basic chemistry

Definitions

matter—anything that occupies space, that has mass, whether it is large or small, living or nonliving, is matter.

elements—an element is a simple form of matter, a substance that cannot be broken down into two or more different substances. Twenty-four elements are present in the body. Of these, by far the most abundant are hydrogen, oxygen, carbon, and nitrogen.

atoms—atoms are the structural units that make up each element.

chemical symbols—chemical symbols are abbreviations for the names of elements. Symbols for 12 of the 24 elements in the body follow:

H hydrogen
O oxygen
C carbon
N nitrogen
P phosphorous
S sulfur
K potassium (Latin, *kalium*)
Na sodium (Latin, *natrium*)
Fe iron (Latin, *ferrum*)
Ca calcium
Co cobalt
Mg magnesium

Atoms

The concept that matter is made up of atoms was proposed early in the 19th century by an English chemist, J. Dalton. He conceived atoms as solid, indivisible particles of matter. After a lapse of about 100 years, Lord Rutherford, another English scientist, disproved this idea and showed that still smaller particles compose atoms. Even the largest atoms are so small that a lineup of 100 million or more of them would measure barely an inch. A staggering fact, then, that even smaller particles can and do exist within such infinitesimal structures!

An atom has two main parts—a nucleus and one or more shells of space surrounding it. The nucleus consists of closely packed, positively charged particles called *protons* and uncharged or neutral particles named *neutrons*—therefore the nucleus as a whole carries a positive charge equal to the number of its protons. Moving around the nucleus in orbitals in each of the atom's shells of space are negatively charged particles called *electrons*. The number of electrically charged electrons in any one atom al-

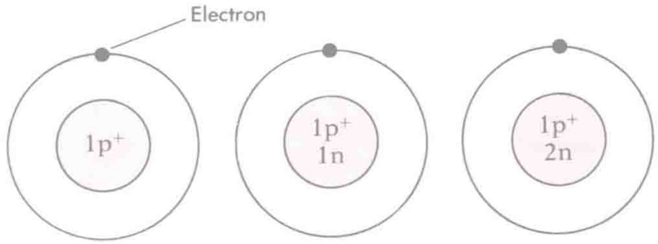

Fig. 1-1 Structure of hydrogen atoms. Left, ordinary hydrogen atom; nucleus consists of 1 proton, and its only shell contains 1 unpaired electron. Note differences in the other two atoms. Middle, rare type of hydrogen atom called deuterium. Right, another rare type of hydrogen atom called tritium.

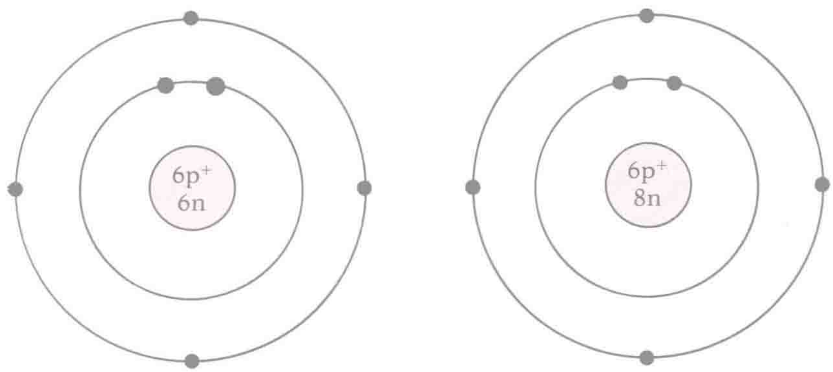

Fig. 1-2 Atomic structure of carbon isotopes. Left, ordinary carbon (atomic number, 6; atomic weight, 12; symbol C^{12}). Right, carbon isotope (atomic number, 6; atomic weight, 14; symbol C^{14}).

ways equals the number of positively charged protons. The opposite charges cancel or neutralize each other; all atoms, therefore, are electrically neutral structures.

The number of protons in an atom's nucleus identifies the kind of element it is and is known as an element's *atomic number.* Hydrogen, for example, has an atomic number of 1; this means that all hydrogen atoms—and only hydrogen atoms—have 1 proton in their nuclei. All carbon atoms, and only carbon atoms, contain 6 protons, so have an atomic number of 6. All oxygen atoms, and only oxygen atoms, have 8 protons and an atomic number of 8. In short, each element is identified by its own unique number of protons, that is, by its own unique atomic number. If two atoms contain a different number of protons, they necessarily have different atomic numbers and are different elements.

Although all atoms of the same element have the same number of protons, they do not necessarily have the same number of neutrons in their nuclei. The nuclei of ordinary hydrogen atoms, for instance, contain no neutrons, but about 1 in 5,000 hydrogen atoms has 1 neutron in its nucleus, and a still smaller proportion of hydrogen atoms contains 2 neutrons. All hydrogen atoms have an atomic number of 1, so all contain 1 proton and 1 electron.

Protons, neutrons, and electrons are particles of extremely small *mass* or weight. (Mass equals weight under ordinary conditions.) A proton weighs almost exactly the same as a neutron, and, compared with an electron, they are heavy particles indeed—each weighs 1,836 times as much as an electron. For all practical purposes, therefore, protons plus neutrons constitute an atom's weight. Look now at the three atoms shown in Fig. 1-1. Note that each contains 1 proton, a fact that identifies each one as a hydrogen atom. How many neutrons does each of these hydrogen atoms contain? Apply the formula Atomic weight = $(p^+ + n)$ to each of these three types of hydrogen atoms. You will quickly discover that they have three different atomic weights. Why? Because each nucleus contains a different number of neutrons. Atoms such as these that contain the same number of protons but a different number of neutrons are called *isotopes*. Because isotopic atoms contain the same number of protons, they are necessarily atoms of the same element and possess identical chemical properties. But because the atoms of isotopes contain different numbers of neutrons, they also have different atomic weights. Fig. 1-2 shows how these facts apply to the structure of carbon isotopes.

Each shell of space around an atom's nucleus can accommodate a certain maximum number of electrons moving in orbitals in the shell. The innermost shell can hold the fewest electrons—1 single electron in the hydrogen atom, but 1 pair of electrons in all other elements. Both the second and the outermost shells can hold a maximum of 4 pairs of electrons.

The number and arrangement of the electrons orbiting in the outermost shell of space around an atom's nucleus determines whether or not the atom is chemically active. When the outer shell contains 1 or more single, that is, unpaired, electrons, the atom is chemically active; it is able to take part in chemical reactions. But when the outer shell contains no single electrons but only pairs of electrons, the atom is chemically inactive or stable.

Chemical reactions

Chemical reactions involve unpaired electrons in the outer shells of atoms. A chemical reaction consists of either the transfer of unpaired electrons from the outer orbital of one atom to the outer orbital of another or the sharing of one atom's unpaired electrons with those of another atom. Fig. 1-3 illustrates an electron transfer type of chemical reaction. In examining this figure, pay attention first to the number and arrangement of electrons in the outer shells of the sodium and chlorine atoms. Note that sodium has only one unpaired electron in its outer shell in contrast to chlorine, which has 1 unpaired electron plus 3 paired electrons, or a total of 7 electrons. Sodium transfers or donates its 1 unpaired electron to chlorine. Chlorine accepts it and pairs it with its 1 unpaired electron, thereby filling its outer shell with the maximum of 4 electron pairs. The formation of this new electron pair in chlorine's outer orbit creates an attractive force that binds the sodium and chlorine atoms to each other. In short, it creates a *chemical bond* between them. The electron transfer constitutes a chemical reaction between the two atoms and produces one molecule of the compound sodium chloride (ordinary table salt). A chemical bond formed by the transfer of electrons is called an *electrovalent* or *ionic bond*. A *covalent bond* is formed by two atoms sharing a pair of electrons (Fig. 1-4). A *compound* is a substance made up of atoms of one or more elements joined to each other by chemical bonds to form units called molecules. *Molecules* are the structural units of compounds; *atoms* are the structural units of elements.

Valence is the number of unpaired electrons in an atom's outer shell, and therefore valence indicates the number of chemical bonds an atom can form with other atoms. For example, sodium and chlorine atoms each have 1 unpaired electron in their outer shells, so each has a valence of 1 and can form one chemical bond with another atom. Hydrogen also has 1 unpaired electron in its only shell, so it, too, has a valence of 1 and can form one chemical bond with another atom. Look again at Fig. 1-2. How many unpaired electrons are in the carbon atom? What

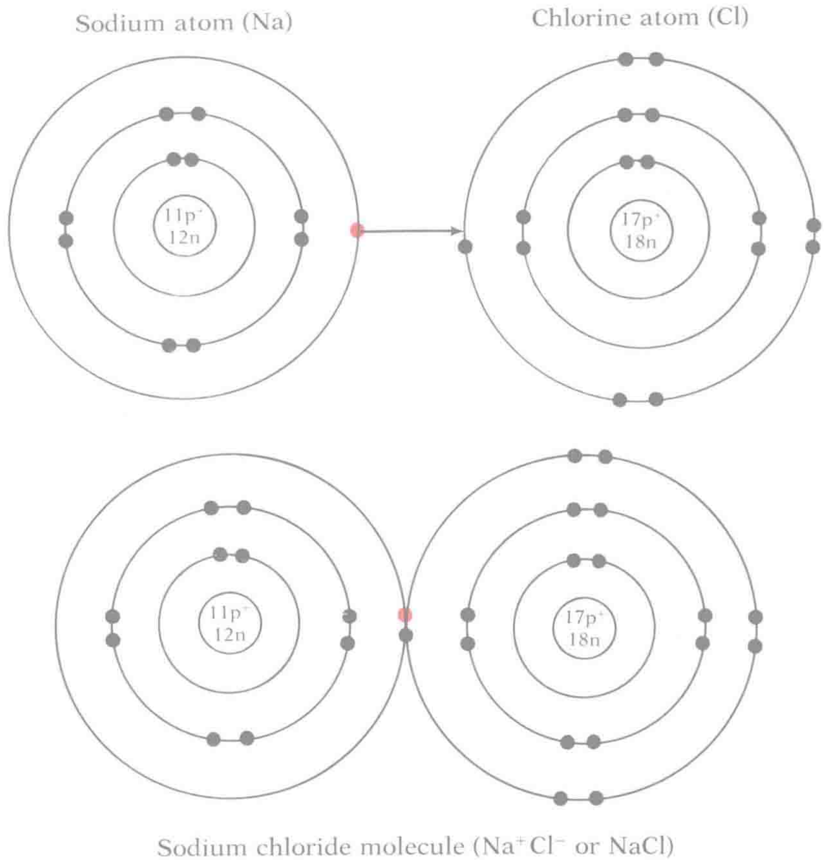

Fig. 1-3 Diagram of sodium and chlorine atoms and electron transfer to form an electrovalent or ionic bond between them and produce a molecule of sodium chloride.

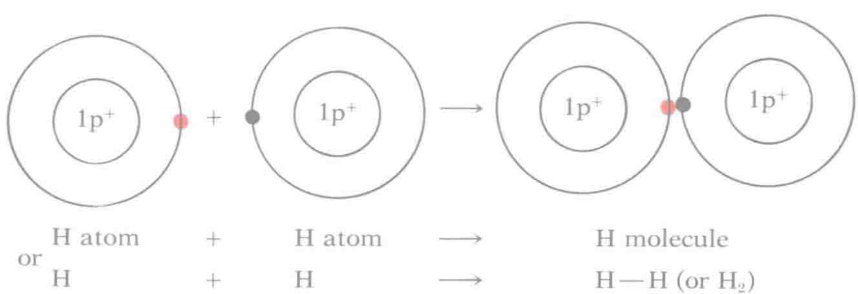

Fig. 1-4 Formation of covalent bond, that is, the sharing of a pair of electrons, between 2 hydrogen atoms produces 1 hydrogen molecule.

is carbon's valence? How many chemical bonds can it form with other atoms? Answers: The carbon atom contains 4 unpaired electrons in its outer shell and so has a valence of 4 and can form chemical bonds with four other atoms. For example, one carbon atom can combine with four hydrogen atoms to form one molecule of methane (CH_4).

Four basic types of chemical reactions that you will learn to recognize as you study physiology are named synthesis, decomposition, exchange, and reversible reactions. *Synthesis* (from the Greek *syn*, together, and *thesis*, putting) reactions put together or combine two or more substances to form a more complex substance, and in so doing, synthesis reactions form new chemical bonds. Many such reactions occur in the body. Every one of its cells, for example, combines amino acid molecules to form complex protein compounds. The name *decomposition* suggests what these reactions do. They decompose or break down a substance into two or more simpler substances, and in so doing, decomposition reactions break chemical bonds. Decomposition and synthesis, in a word, are opposites. Synthesis builds up, decomposition breaks down. Synthesis forms chemical bonds, decomposition breaks chemical bonds. One of the many examples of decomposition reactions in the body is the digestion or breakdown of large fat molecules into two simpler substances, fatty acids and glycerol. The nature of *exchange reactions* is also suggested by their name. They break down or decompose two compounds and in exchange synthesize two new compounds. Certain exchange reactions take place in the blood. One example is the reaction between lactic acid and sodium bicarbonate. The decomposition of both substances is exchanged for the synthesis of sodium lactate and carbonic acid. Perhaps you can see these changes more easily in an equation.

$$H \cdot \text{lactate} + NaHCO_3 \rightarrow Na \cdot \text{lactate} + H \cdot HCO_3$$

The formula H · lactate represents lactic acid; $NaHCO_3$ is the formula for sodium bicarbonate; Na · lactate represents sodium lactate; and H · HCO_3 represents carbonic acid. *Reversible reactions*, as the name suggests, proceed in both directions. A great many chemical reactions are reversible, and we shall cite a number of them in later chapters of the book.

Before we leave the subject of chemical reactions, we must state the following important principle about them: Every chemical change is coupled with an energy change. The formation of chemical bonds is coupled with energy use. The breaking of chemical bonds is coupled with energy release. Synthesis reactions, therefore, use energy, and decomposition reactions release energy. Example in living cells of synthesis with energy use: reactions that combine simple food compounds to form complex compounds. Example in living cells of decomposition with release of energy: reactions that break food compounds down into simpler compounds.

Radioactivity

Some atoms have the property not only of chemical activity but also of radioactivity. Whereas chemical activity consists of a change in an atom's outer shell—a transfer or a sharing of its single, unpaired electrons—radioactivity involves a change in an atom's nucleus. Specifically, *radioactivity* is the emission from the nucleus of particles or rays known as alpha, beta, and gamma rays. Alpha and beta rays are actually minute particles. *Alpha rays* consist of 2 protons and 2 neutrons ejected from the nucleus. *Beta rays* are electrons, not from an atom's shell, but from its nucleus, where they originate by a neutron breaking down into a proton and an electron. The proton remains behind in the nucleus, and the electron shoots out of it as a beta ray. *Gamma rays* are not particles but electromagnetic radiations emitted from an atom's nucleus and presumably produced by the shifting of particles in the nucleus after the ejection of alpha or beta particles.

How does radioactivity, the emission of alpha, beta, or gamma rays, change an atom? To answer this question, consider the following example. A synthetic radioactive isotope of iodine used in certain diagnostic tests and commonly called iodine-131, is represented by the symbol $_{53}I^{131}$. Interpreted, this symbol indicates iodine atoms that have an atomic number of 53 (53 protons in their nuclei) and an atomic weight of 131. By rewriting the formula Atomic weight = $(p^+ + n)$ as (Atomic weight $- p^+) = n$, we can determine that there are $(131 - 53)$ or 78 neutrons in the nucleus of iodine-131. Ordinary iodine atoms have an atomic weight of 127, with 53 protons in their nuclei but only 74 neutrons. Radioactive iodine-131 emits beta and gamma rays and thereby changes the nuclei of its atoms. One of its neutrons converts to 1 proton and 1 electron. The electron leaves the nucleus as a beta ray, but the proton stays in the nucleus. Result? The nucleus now contains one less neutron—it converted to a proton and electron—and one more proton than it had before its beta ray emission. Changing the number of protons in an atom's nucleus changes the atom into another element. A basic principle about atoms, you will recall, is that all atoms of the same element contain the same number of protons. A corollary stems from this principle: Atoms that contain different numbers of protons are necessarily different elements. Iodine-131 atoms that have emitted beta rays contain 54, not 53, protons, as do all iodine atoms. Atoms of the element xenon contain 54 protons. Radioactivity, therefore, changes iodine into xenon. An iodine atom that has participated in a chemical reaction is still an iodine atom; only the number of electrons in its outer shell has changed—the number of protons in its nucleus remains the same. We started this paragraph with a question and shall end it with an answer. Radioactivity, that is, the emission of alpha, beta, or gamma rays from an atom's nucleus, changes one element into another by changing the number of protons in the atomic nucleus. One more question and answer. Why do you, a student of anatomy and physiology, need to know anything about radioactivity? One reason that seems valid, at least to us, is that radioactive isotopes have become widely used in studies of various aspects of physiology and in the diagnosis and treatment of diseases. The rays emitted by radioactive atoms serve as labels of the atoms, which can be detected by Geiger counters.

Selected facts from biochemistry

Definition of biochemistry

Biochemistry is the chemistry of living organisms. It is a young but extensive science—so extensive that some biochemistry textbooks contain more than a thousand pages. Out of this abundance, we have tried to select for inclusion in this chapter the facts and principles you will need most to know in order to understand the rest of the book. We have organized this information around three topics: biomolecules, bioenergy, and biosynthesis.

Biomolecules

A textbook by the distinguished biochemist Albert L. Lehninger opens with the sentence, "Living things are composed of lifeless molecules." Biomolecules are these lifeless molecules. By definition then, biomolecules are compounds that compose living organisms. Most abundant of them all is water. The thousands of other biomolecules fall into seven classes of compounds: acids, bases, salts, proteins, carbohydrates, lipids, and nucleic acids. To learn some basic facts about each of these kinds of biomolecules, study the following paragraphs.

Water serves a host of vital functions in the body, many of which are related to the fact that it dissolves so many substances. By dissolving oxygen and food substances, for instance, water enables these essential materials to enter and leave the blood capillaries and to enter cells. Also, by dissolving substances, water makes possible the thousands of chemical reactions that keep us alive. Dry substances are virtually nonreactive. Another important function of water, that of enabling the body to maintain a relatively constant temperature, stems from the fact that water both absorbs and gives up heat slowly.

Acids, bases, and salts belong to a large group of compounds called electrolytes. *Electrolytes*, by definition, are compounds whose molecules consist of positive ions (cations) and negative ions (anions) and that dissociate when dissolved in water into cations and anions. In short, electrolytes are substances that ionize in solution. Formation of an electrolyte takes place by electron transfer, as shown in Fig. 1-3. Here you see a sodium atom donating the single electron in its valence shell to the valence shell of a chlorine stem. Note again what this accomplishes. The transferred electron fills chlorine's outer shell by forming a fourth electron pair in it; this pair constitutes an electrovalent bond between the sodium and the chlorine. It also does something else. It converts both the sodium and chlorine atoms into ions or electrically charged particles. Since the sodium atom has donated one of its electrons, that is, one of its negative charges, it now contains one more positive charge than negative (11 p^+ and 10 e^-). It has become a *positive ion*, a *cation*, designated by the symbol Na^+. Chlorine, after accepting the electron, bears one more negative than positive charge. It has become a *negative ion*, an *anion*. The electrovalent bond between the two ions creates out of them one molecule of the ionic compound sodium chloride. The formula Na^+Cl^- represents sodium chloride (ordinary table salt). Because chemically bound ions compose its molecules, it is an *ionic compound*. Because solutions of ionic compounds conduct an electric current, another name for them is electrolytes. Chapters 21 and 22 give more information about these essential biomolecules.

Acids, like all electrolytes, ionize when dissolved in water to form cations and anions. By definition, an *acid* is a compound that yields hydrogen ions (H^+) in solution. A *base* is a compound that yields hydroxyl or OH ions (OH^-) in solution. A *salt* is a compound that in water yields the positive ions of a base and the negative ions of an acid. A salt is formed when an acid reacts with a base. Here is an example:

HCl	+	NaOH	=	NaCl + Water
Hydrochloric acid		Sodium hydroxide, a base		Salt

Proteins (from the Greek *proteios*, of the first rank) are the most abundant of the carbon-containing or organic compounds in the body, and as their name implies, their functions are of first-rank importance. The units or building blocks that compose proteins are *amino acids*. Fig. 1-5 diagrams the basic structure of an amino acid. What four elements compose it? The same elements are present in all amino acids, plus sulfur in many of them and phosphorus or iron in some. Note that bonded to the first, or alpha, carbon are an amino group (NH_2), a carboxyl group (COOH), a hydrogen atom, and R, which stands for a side chain of one or more atoms. In 19 of the 20 different amino acids, R is a group of atoms—a different group in each one—but in the simplest amino acid (glycine), it is a hydrogen atom. The nature of the side chain identifies the amino acid.

Amino acids become linked together by *peptide bonds*, that is, chemical bonds that join the carboxyl group of one amino acid to the amino group of another by eliminating water from

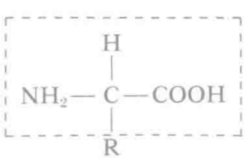

Fig. 1-5 Basic structural formula for an amino acid. All amino acids contain the part of the molecule enclosed in the rectangle; *R* represents the part unique to each of the 20 different amino acids. Examples: In glycine, the simplest amino acid, *R* is a hydrogen atom; in alanine, it is the group CH_3.