Handbook of cellular chemistry

Annabelle Cohen

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Department of Biological Sciences,
Staten Island Community College,
Staten Island, New York

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Preface

Biology is a composite science in the sense that living systems are primarily complex physicochemical systems. It is therefore customary to introduce beginning students to courses in biology and human physiology with a series of lectures on chemistry as it pertains to living structures and function. This handbook is based on such lecture material.

The study of biology on a chemical or molecular level is generally called *biochemistry*. The term *cellular chemistry* is used here in a broad sense to refer to the chemical organization and related activities of living cells. The student should realize, however, that the contents of this handbook are only a representative sample of a complicated and rapidly expanding body of knowledge. My aim in this text has been to select and explain important concepts in terms of the most recent advances in the field without at the same time overwhelming the beginner with a mass of technical detail.

For students of the life sciences, physical and chemical concepts acquire special significance when they are correlated with the function of cells, particularly the cells of human beings. I have therefore included a variety of relevant topics, ranging from cell membrane dynamics to digestive enzymes. In view of the current importance of nucleic acids in the biomedical sciences, a section on genetic defects and viruses has also been added. It is my hope that students will obtain sufficient information and be stimulated to question and to delve further into the chemical realities of life by these up-to-date, comprehensive, brief discussions.

It is a pleasure to acknowledge the helpful suggestions and continuing encouragement of my son David, a student at the Massachusetts Institute of Technology.

Annabelle Cohen

Contents

1

Introductory chemistry for biologists

If a cell or a tissue is removed from the body and analyzed in a test tube, it will be found to consist of a mixture of substances, including water, salts, proteins, sugars, and fats. Most of these substances, the nonliving chemical raw materials from which living protoplasm is made, exist in the form of *molecules*. Molecules can be broken down by conventional chemical methods into smaller units called *atoms*. In our universe, all matter (anything that occupies space and has mass or weight) is composed of atoms. Protoplasm may thus be considered a rather complex form of matter derived from chemical units that happen to be present on this particular planet. The study of the composition and properties (characteristics) of substances and the changes they undergo is the science of *chemistry*. Some knowledge of this science is essential for understanding the nature of the structure and function of living systems.

molecules

chemistry

ELEMENTS

The enormous numbers of substances found on earth, in both living and nonliving forms, are all made up of elements or combinations of elements. The term *element* is given to elementary chemical substances that cannot be further decomposed (broken down) into simpler substances by ordinary chemical means.

element

Some elements, such as gold, silver, copper, and sulfur, were known to the ancient world. As man's knowledge of chemistry advanced, more of the earth's elements were identified until what appeared to be a reasonably complete list of 89 naturally occurring elementary substances was assembled. These ranged from the lightest known element, hydrogen, to the heaviest, uranium. At present, there are 106 elements; the additional 17 elements have been made by man in various accelerator devices such as cyclotrons, synchrotrons, and nuclear reactors.

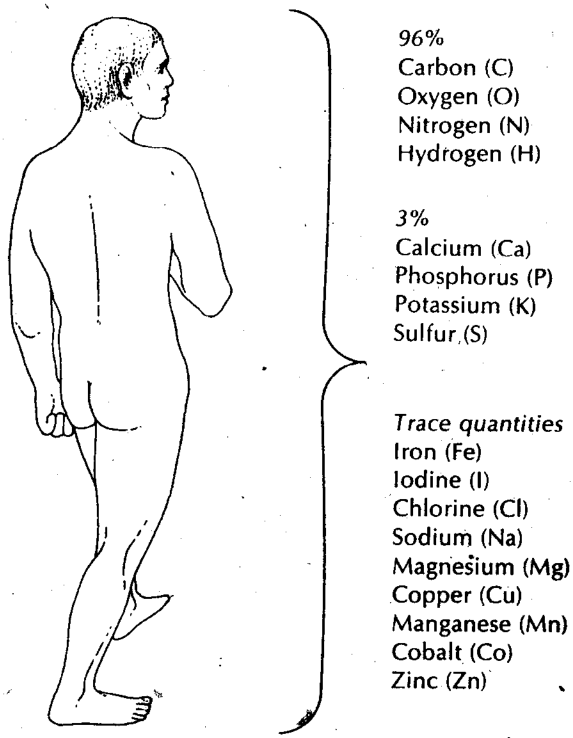


FIG. 1-1. The elements present in human tissues.

TABLE 1-1. Chemical symbols based on Latin names of elements

Element	*Chemical symbol*	*Latin name*
Antimony	Sb	**Stibium**
Copper	Cu	**Cuprum**
Gold	Au	**Aurum**
Iron	Fe	**Ferrum**
Lead	Pb	**Plumbum**
Mercury	Hg	**Hydrargyrum***
Potassium	K	**Kalium**
Silver	Ag	**Argentum**
Sodium	Na	**Natrium**
Tin	Sn	**Stannum**
Tungsten	W	**Wolfram†**

*Literally meaning liquid silver; mercury is a liquid metallic element.
†Wolfram is the German name for tungsten.

Human tissues contain only about 20 of these elements. In fact, 4 of them, namely *carbon, hydrogen, oxygen, and nitrogen,* compose approximately 95% of the substance of the body (Fig. 1-1).

In chemical shorthand, each element is represented by a symbol. The symbol is usually the *capital* initial letter of the English or Latin name of the element and sometimes, to avoid confusion, one more *small* letter from the name. Some symbols derived from Latin names of elements are listed in Table 1-1.

Atomic structure of elements

Elements are composed of extremely small, discrete particles called atoms (the mass of an atom is in the range of 10^{-24} to 10^{-23} g; see appendix to this chapter for an explanation of the magnitude of these numbers). The *atom* is the smallest unit of an element that exhibits all the characteristic properties and undergoes the characteristic chemical changes of that element. Atoms of the same element are more or less alike and differ in their properties from atoms of other elements. Each element thus has its own characteristic atoms. It is important to remember that the symbol for an element represents not only the *name* of the element but is also the chemist's way of representing *one atom* of that element. atom

For many years it was considered that atoms were indivisible units that could not be broken down into anything smaller. However, it is now known that atoms are clusters of varying numbers of even smaller (subatomic) particles called *protons, electrons,* and *neutrons.** The arrangement of these particles in an atom resembles in some respects a solar system of planets orbiting around a sun. The mass of an atom is concentrated in a positively charged *nucleus* made up of all the protons and neutrons of that particular atom, whereas the electrons orbit the nucleus at relatively great distances and in various energy levels, or shells. nucleus

The subatomic particles are classified as follows:

1. ***Protons*** are positively charged particles (having a +1 charge) with an assigned mass of 1 amu (atomic mass unit). They are found in the nucleus of every atom. proton

2. ***Electrons*** are negatively charged particles (having a −1 charge) with a mass of about 1/1,840 amu (protons are about 1,840 times heavier than electrons). They are found moving around outside the nucleus at various distances from it. electron

3. *Neutrons* are uncharged (neutral) particles with a mass of 1 amu (approximately equal to the mass of a proton). They are found in the nucleus of every atom except the common hydrogen atom. neutron

*There are other subatomic particles, but they are not relevant to this discussion.

TABLE 1-2. Fundamental atomic particles

Name	*Symbol*	*Location in atom*	*Assigned mass*	*Charge*
Electron	e or e^-	outside of nucleus	0 amu	−1
Proton	p	nucleus	1 amu	+1
Neutron	n	nucleus	1 amu	0

amu

Note that the mass of a proton or a neutron is given as equal to approximately 1 amu. The *amu* is an arbitrary unit defined as 1/12 the mass of the most common naturally occurring form of carbon, that is, carbon 12 (^{12}C). This carbon atom has 6 protons (and 6 electrons) and 6 neutrons and has been assigned the standard atomic mass of exactly 12.0000. Because the mass of an electron is negligible when compared with that of either a proton or a neutron, it is generally considered to have no mass. The characteristics of these subatomic particles are summarized in Table 1-2.

Electrons and energy levels (shells)

electron shells

The atom, as described by the Danish physicist Niels Bohr in the early years of the twentieth century, was visualized as consisting of a very small, dense, positively charged nucleus with electrons moving in circular or elliptical orbits at fixed distances around the nucleus. According to this model, an atom can have as many as 7 *electron shells* (energy levels). A shell is defined as a discrete volume around a nucleus in which a given electron or set of electrons moves. Each shell can hold only a given maximum number of electrons. The first shell (nearest the nucleus) can hold a maximum of 2 electrons, the second a maximum of 8, the third up to 18, the fourth and fifth a maximum of 32, and so on.

Regardless of the total number of electron shells, the outermost shell never contains more than 8 electrons; if there is only 1 shell (as in hydrogen and helium), then it is filled by only 2 electrons.

Diagrams of several representative atoms are shown in Fig. 1-2. The number of neutrons (n) and protons (p) making up the nucleus of each atom are indicated in the central circle. The electrons (e) are shown as numbers of e at the appropriate energy levels outside the nucleus. The energy levels, or shells, are numbered at the right side of the atoms from the nucleus outward. The electron shells should be visualized as three-dimensional spheres rather than two-dimensional circles; electron orbits are apparently not flat, as are the orbits of the planets.

Note that hydrogen has only 1 electron in only 1 shell; thus

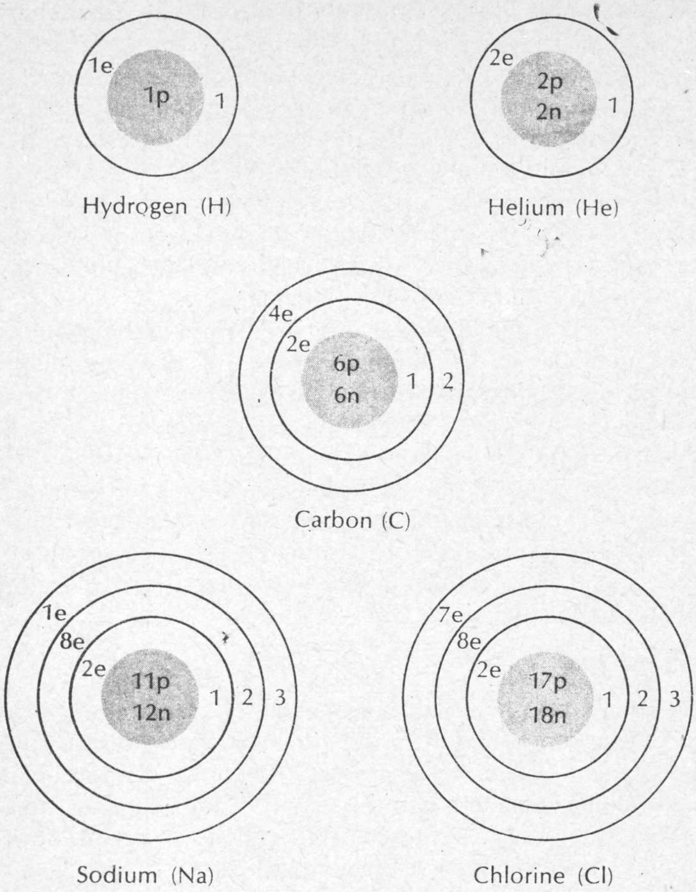


FIG. 1-2. Atomic structures of selected elements, indicating numbers and locations of protons (p), electrons (e), and neutrons (n). These drawings represent cross sections through the nuclei and three-dimensional, spherical electron shells. The sizes of the nuclei are greatly exaggerated; on this scale, each nucleus would actually be much smaller than a pencil point.

its outermost shell is unfilled. Helium has 2 electrons in its only shell, which is therefore completely filled with the maximum number of electrons (2) for that shell. Carbon has 2 shells, the first of which is filled with the maximum 2 electrons; the second, the outermost, is incompletely filled with 4 electrons. The sodium atom has 11 electrons in 3 shells. The first is filled with 2 electrons, the second with 8, and the third shell, the outermost, is unfilled with only 1 electron. Chlorine has 17 electrons, likewise distributed in its 3 shells, with 2 electrons in the first, 8 in the second, and 7 in the incomplete outer third shell. The electrons in an unfilled, or unsatisfied, outermost shell are extremely significant because these are the *valence electrons* that make an atom chemically reactive (see Chapter 2).

valence electrons

It must be mentioned at this point that the Bohr atom model cannot account for certain aspects of the physical behavior of electrons in atoms. The modern theory of atomic structure is

antum mechanics

based on *quantum mechanics*. Although the quantum mechanical model of the atom is similar to the Bohr model, it embodies a considerably more complicated three-dimensional spatial arrangement of electrons in energy levels, energy sublevels, and orbitals. The orbitals define a region in the space outside the atomic nucleus, where any particular electron has a 90% probability of being found at any given moment.

Electrons are in rapid motion and are viewed in the quantum mechanical model of the atom as clouds of negative charge (electron clouds) rather than particles in specific positions revolving around the nucleus. The details of the modern version of electron configuration in atoms are outside the scope of this text. However, students should visualize electrons, protons, and neutrons, not in terms of tiny static bodies fixed in time and space but as swift-moving packets of flashing energy. The subatomic particles are, in fact, entities on the borderline between energy and matter in our universe.

Atomic number

atomic number

The atoms of a given element contain a specific number of *protons* in their nuclei. This proton count is the *atomic number* of that element. There are at present 106 known elements. The range of atomic numbers is from 1 to 106; for example, atoms of the lightest element, hydrogen (H), contain *1* proton; atoms of the most recent man-made element (the heaviest, so far), contain 106 protons.

In a neutral atom (an atom in an uncombined state) the number of protons inside the nucleus (the atomic number) *is the same* as the number of electrons outside the nucleus; that is, the number of positive charges is balanced by the number of negative charges. The importance of this fact is that the basic physical and chemical properties of an atom depend on the number of protons in its nucleus and the number of orbiting electrons. Atomic number is therefore the most fundamental property of an atom. If the atomic number changes for some reason (as it does in certain radioactive elements), the properties of that atom change quite drastically.

Mass number

mass number

The *mass number* of an atom is the total number of particles in the *nucleus* (protons plus neutrons). It is called the mass number because protons and neutrons have about the same mass (1 amu), and the sum of their masses accounts for most of the mass of an atom; the mass of electrons in an atom is negligible.

(One atomic mass unit [amu] is defined as 1/12 the mass of an atom of carbon 12.)

The complete chemical notation for an element gives its mass number and atomic number:

$$^{A}_{Z}E$$

where

A = mass number (in amu)
E = symbol of the element
Z = atomic number of the element

Examples of this type of notation are:

1. Most atoms of hydrogen (H) have 1 proton and no neutrons in the nucleus. Therefore, the atomic number of H is 1; the mass number of H is 1 amu. An atom of this element is thus represented as

$$^{1}_{1}H$$

2. Most atoms of helium (He) have 2 protons and 2 neutrons in the nucleus. The atomic number of He is 2; the mass number of He is 4 amu. An atom of this element can be shown as

$$^{4}_{2}He$$

3. Most atoms of chlorine (Cl) have 17 protons and 18 neutrons. The atomic number of Cl is 17; the mass number of Cl is 35 amu. Thus,

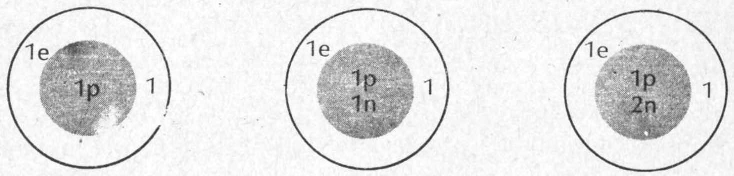
$$^{35}_{17}Cl$$

It will also be obvious from the above that an element's mass number (A) minus its atomic number (Z) equals the number of neutrons in the atoms of that element.

Isotopes

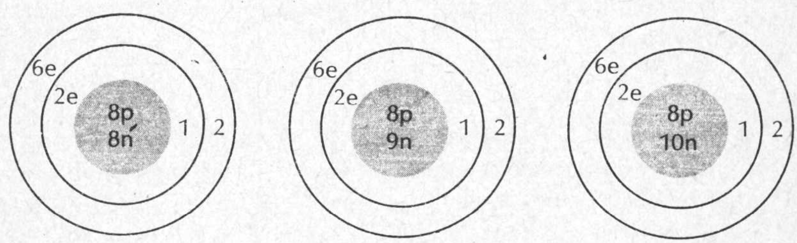
When the mass numbers of the atoms of a given element are compared, it is found that all of its atoms do not have the same mass number. For example, while the nuclei of 99% of all hydrogen atoms consist of 1 proton only, a small percentage will be found to have 1 or 2 neutrons as well. Atoms of the same element having the *same* atomic number but with *different* mass numbers (due to different numbers of neutrons in the nucleus), are called *isotopes* of that element. Generally speaking, the isotopes of a given element have the same chemical properties. This factor is controlled by the numbers of protons and electrons, which are constant for all isotopes of that element. **isotope**

Most elements found in nature have several isotopes, but a great many isotopes have also been made by man. As a result,



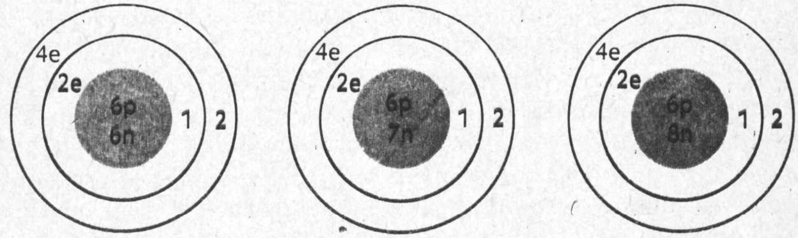
Common hydrogen ^1_1H | Heavy hydrogen, or deuterium ^2_1H | Triple-weight hydrogen, or tritium ^3_1H

Three isotopes of hydrogen are shown. All have 1 electron and 1 proton but different numbers of neutrons. Over 99% of naturally occurring hydrogen atoms are ^1_1H. Tritium is a radioactive isotope of hydrogen.



Oxygen $^{16}_8O$ | Oxygen $^{17}_8O$ | Oxygen $^{18}_8O$

Oxygen has three naturally occurring isotopes, with 8, 9, and 10 neutrons respectively. All atoms of this element have 8 protons and 8 electrons. Over 99% of oxygen atoms are of the first type (mass number = 16).



Common carbon $^{12}_6C$ | Carbon $^{13}_6C$ | Carbon $^{14}_6C$

Over 98% of naturally occurring carbon atoms are carbon 12. Carbon 14 is a radioactive isotope.

FIG. 1-3. **Isotopes of common elements.**

although there are just over 100 different known elements, there are more than 1,400 different varieties of atoms. Some examples of isotopes of elements are shown in Fig. 1-3.

Atomic weight

When classifying elements by their mass numbers, it is inconvenient to include the different mass numbers of all the isotopes of that element. Chemists usually refer to the mass numbers of the atoms of an element in terms of the *atomic weight* of the element. This is not really a weight but an *average* of the mass number of the isotopes, expressed in amu. The standard for atomic weights is the mass number of the most common isotope of carbon (mass number, 12 amu). Carbon 12 has been assigned an atomic weight of exactly 12.0000. To say that a given atom has an atomic weight of 36 indicates that it is three times as heavy as an atom of carbon 12. Atomic weight is thus a relative weight. As a general rule, *the atomic weight of any element is approximately equal to the mass number of its most*

atomic weight

TABLE 1-3. Some important elements, their symbols, atomic numbers, and atomic weights

Element	*Symbol*	*Atomic number*	*Atomic weight**	*Element*	*Symbol*	*Atomic number*	*Atomic weight**
Aluminum	Al	13	26.9815	Lithium	Li	3	6.94
Antimony	Sb	51	121.7	Magnesium	Mg	12	24.305
Argon	Ar	18	39.94	Manganese	Mn	25	54.9380
Arsenic	As	33	74.9216	Mercury	Hg	80	200.5
Barium	Ba	56	137.3	Neon	Ne	10	20.17
Beryllium	Be	4	9.01218	Nickel	Ni	28	58.7
Bismuth	Bi	83	208.9806	Nitrogen	N	7	14.0067
Boron	B	5	10.81	Oxygen	O	8	15.999
Bromine	Br	35	79.904	Phosphorus	P	15	30.9738
Calcium	Ca	20	40.08	Platinum	Pt	78	195.0
Carbon	C	6	12.011	Potassium	K	19	39.10
Chlorine	Cl	17	35.453	Radium	Ra	88	226.0254
Chromium	Cr	24	51.996	Selenium	Se	34	78.9
Cobalt	Co	27	58.9332	Silicon	Si	14	28.08
Copper	Cu	29	63.54	Silver	Ag	47	107.868
Fluorine	F	9	18.9984	Sodium	Na	11	22.9898
Gold	Au	79	196.9665	Strontium	Sr	38	87.62
Helium	He	2	4.00260	Sulfur	S	16	32.06
Hydrogen	H	1	1.008	Tin	Sn	50	118.6
Iodine	I	53	126.9045	Tungsten	W	74	183.8
Iron	Fe	26	55.84	Uranium	U	92	238.029
Lead	Pb	82	207.2	Zinc	Zn	30	65.3

*Based on the assigned relative atomic mass of ^{12}C = 12.0000.

common isotope (the average of the mass numbers of the isotopes of an element is weighted in favor of the most abundantly occurring isotope). Thus:

1. Oxygen (O) has an assigned atomic weight of 15.999, indicating that its most common isotope is ^{16}O
2. Chlorine (Cl), a mixture of 2 isotopes, about 75% ^{35}Cl and 25% ^{37}Cl, has an atomic weight of 35.453
3. The atomic weight of carbon (C) is 12.011. Naturally occurring carbon is a mixture of about 99% carbon 12 and about 1% carbon 13.

Table 1-3 lists some important elements, their symbols, atomic numbers, and atomic weights.

Radioactivity

The heavier isotopes of a number of elements are unstable; their nuclei decay spontaneously to more stable forms. This decay is called *radioactivity*. The earth is slightly radioactive from the presence of a few naturally *radioactive elements*, for example, radium (Ra), polonium (Po), and uranium (U), and from radioactive isotopes, or *radioisotopes*, of *stable elements*, for example, lead 204 (^{204}Pb) and potassium 40 (^{40}K). Additionally, about 1,000 artificial radioisotopes have been made in nuclear reactors. The disintegration of a radioactive nucleus may be described as a tiny explosion in which one or more particles are ejected at high speeds. The particle emitted may be either an *alpha particle*, which consists of 2 neutrons and 2 protons, or a *beta particle* (actually an electron resulting from the disintegration of a neutron). In some instances, alpha and beta emission is accompanied by high-energy electromagnetic rays called *gamma rays*.

radioactivity

radioisotope

alpha and beta radiation

It is interesting that the decay of a radioactive nucleus confirms the idea that the atomic number of an atom determines its properties and thus its identity. If a radioactive nucleus decays by emitting an alpha particle, a new nucleus is formed with a mass number that is four less than the original nucleus and an atomic number that is two less. If a nucleus decays by the disintegration of a neutron into an electron (emitted as a beta particle) and a proton that stays behind, then a new nucleus is formed. The new nucleus has the same mass number as the original, but an atomic number that is greater by one. In either case, a nuclear *transformation* has taken place, and a *different* atom, that is, a *different* element, is formed. For example, it can be readily demonstrated that when radioactive phosphorus (^{32}P) decays by emitting a beta particle, it is transformed into the element sulfur:

$$^{32}_{15}P \xrightarrow[\text{decay}]{\text{radioactive}} {}^{32}_{16}S + \text{energy as beta radiation}$$