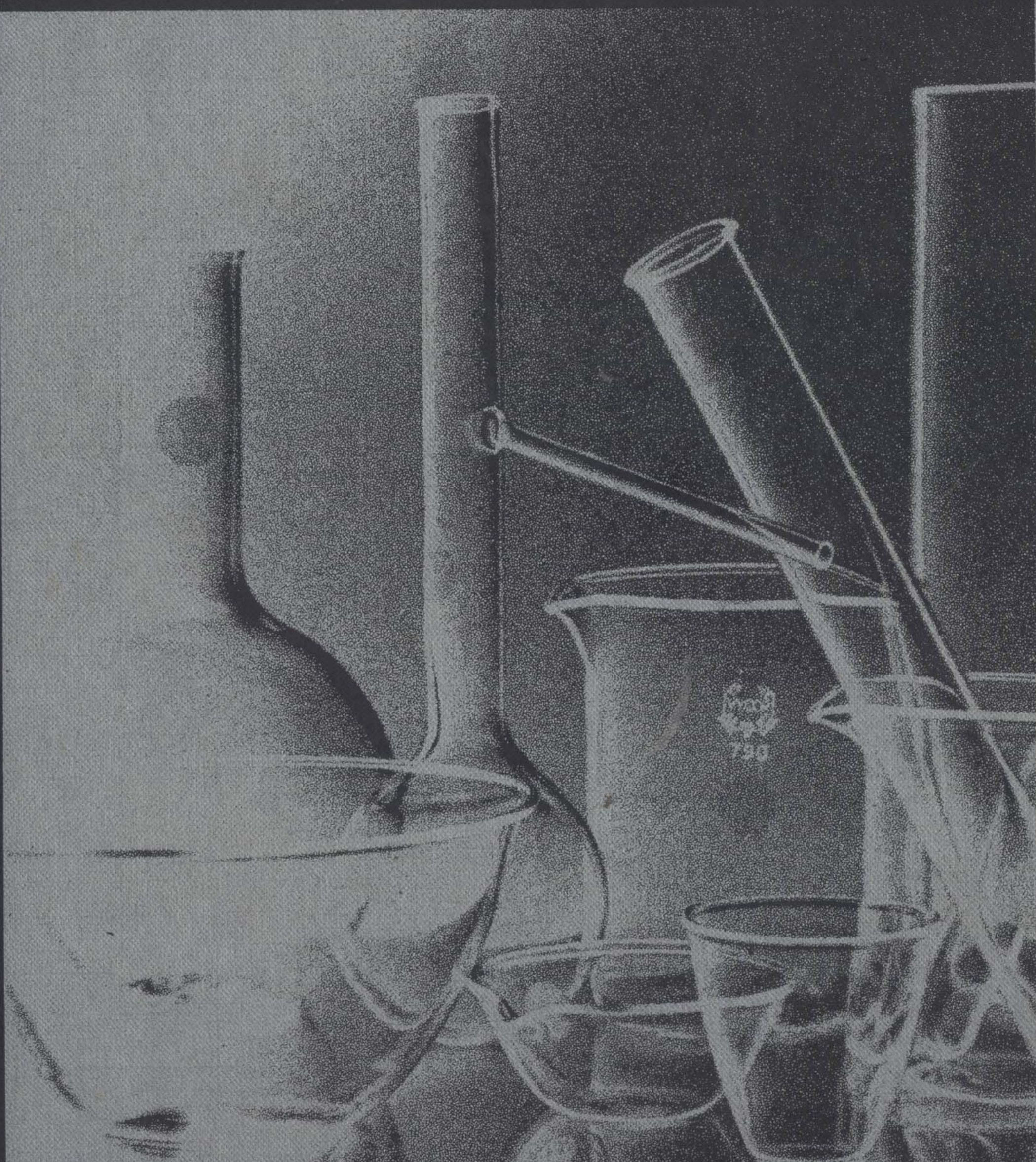
CHEMISTRY:
WITH SELECTED PRINCIPLES OF PHYSICS

Rosemary Kennelly and Raymond E. Neal



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PREFACE

This second edition is offered to students in the health professions as a more complete and up-to-date presentation of important chemical and physical principles fundamental to an understanding of life processes. Biochemical reactions have been expanded to include oxidative processes involved in the source of energy for living. Since the role of enzymes in diagnosis and treatment is becoming more important day by day, these marvelous proteins are considered in some detail. A brief discussion of the nucleic acids has also been included. As a result of valuable suggestions from students, portions of the text have been rewritten in the interests of simplification and clarification.

As in the first edition, the emphasis is on application of basic science principles to situations arising in working with patients. A sound understanding of the chemistry of electrolytes is fundamental to an understanding of fluid and electrolyte balance and replacement fluid therapy. A knowledge of gas and water pressure is basic to any understanding of the mysteries of suctioning apparatus.

The first six chapters contain the background material necessary for an understanding of any of the subsequent chapters. By using this design, it is easy to correlate the subsequent chapters with other areas in the curriculum. Thus having completed the first six chapters, the instructor may choose the order of presentation of the rest of the material.

Equation writing has been deemphasized except in those cases where equations serve a major role in clarification as in oxidations, chemical digestion, and electrolytes and electrolyte balance. Technical and mathematical material has been kept to a minimum and, when included, has been translated into terms and explanations that can be readily understood. As far as possible, examples, applications, and illustrative materials have been selected from situations in the health fields.

Rosemary Kennelly

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chapter one THE SHAPE OF THINGS

As scientific instruments, tools, and measuring devices become more refined and precise, scientists are able to probe deeper and deeper into the composition and structure of matter, coming ever closer to a scientific explanation of life and life processes. The elucidation of the structures of DNA (deoxyribonucleic acid) and RNA (ribonucleic acid), their role in heredity, the structure and methods of synthesis of enzymes, and the application of this knowledge to the study of health and disease have opened vast areas of research, promising ever greater comprehension of the relationships between chemical transformations in health and disease.

On a fundamental level, the old distinctions between chemistry, biology, and physics have broken down. While we may still say that chemistry is the study of matter and its transformations, biology is the study of life, and physics is the study dealing chiefly with the interactions of matter and energy and the transformation of energy, increasingly studies are being done in molecular biology, biophysics, biochemistry, and biophysical chemistry. These interdisciplinary approaches have greatly expanded our knowledge and have emphasized the fact that we cannot study only matter or energy, but must consider also the ways substances are organized or arranged on the atomic level.

MATTER

Matter is anything that has mass. We recognize matter by its properties. One property of matter is inertia, namely, the ability of material to remain at rest or in uniform motion in a straight line unless acted upon by some external force. Mass enables an object to have weight in a gravitational field and imparts to the object its inertia. Weight is a measure of the gravitational attraction between the earth and an object. The weight of an object depends upon the mass of the object and its distance from the earth. We usually think of weight in relation to the earth but the law of gravity applies to the attraction between any two objects.[1] As the distance between an object and the earth increases, the gravitational pull or weight decreases. Thus we have the problem of weightlessness faced by astronauts. The mass, however, remains unchanged.

Substances

Any sample of matter has in addition to inertia, other properties. These are classified

[1]The law of gravity states that *the gravitational force is equal to the product of the masses of the bodies divided by the square of the distance between them.*

$$F = \frac{M_1 M_2}{D^2}$$

as physical and chemical properties. Physical properties can be detected or measured without altering the composition of the sample. They include color, odor, state, solubility, boiling point, freezing point, etc. Chemical properties, on the other hand, cannot be observed without altering the composition of the material. Combustibility is a chemical property. If a sample of material does burn, you no longer have your original starting material, but the products of the combustion.

If a sample of material has definite, recognizable, unvarying properties, it is classified as a pure substance. Gold, mercury, table salt, and glucose are examples of pure substances. Two or more substances mixed together can give a variety of properties depending upon the amounts of each substance present. These combinations are, logically enough, called mixtures. Milk, mayonnaise, flour, tincture of iodine, and sea water are examples of mixtures.

Pure substances are subdivided into elementary substances and compound substances, or simply elements and compounds.

Elements

An element is a substance that cannot be decomposed into simpler substances or built up from them. This definition is satisfactory for the present. However, radioactive elements are continually decomposing, forming daughter elements, and many of these continue to decompose until a stable element is formed. New elements can also be made from other elements in the cyclotron, betatron, or as the result of atomic detonations. We can say that elements are simple or fundamental substances composed of only one kind of atoms.

It is hardly safe today to state the number of elements. Through the study of the structure of atoms, it is possible to say with certainty that there are at least 103 elements. At room temperature, 2 of these are liquids: mercury and bromine; 11 are gases: hydrogen, helium, nitrogen, oxygen, fluorine, neon, chlorine, argon, krypton, xenon, and radon. The others are solids. Of the elements, only about 20 are normally found in the human body.

The elements are divided into two classes, the metals and the nonmetals. The metals are the more common. We meet them in our daily experiences: the iron of the beds, the mercury of thermometers, and the silver of silverware. The metals are characterized by the property of luster. This, obviously, can be seen only while observing the clean surface of the metal. Many of the metals are so active that the exposed surface soon becomes covered with a layer of oxide, which hides its true appearance. The luster of metallic sodium, for instance, can be seen only on a freshly cut surface. In some instances the oxide is transparent—e.g., aluminum oxide—and while this reduces the luster somewhat, its presence protects the metal from further action. The metals exhibit high conductivity of heat. Silver is notable in this regard, as everyone knows who has used a solid silver spoon in a cup of hot tea. They also have high electrical conductivity.

The nonmetals are not so well known. It is true that carbon and sulfur are familiar, but silicon and selenium are merely words to most people. Iodine, on the other hand, is well known because of the general use of tincture of iodine, an alcoholic solution of the nonmetal and potassium iodide.

Atoms

According to the theory postulated by John Dalton, an English schoolmater, elements are made up of atoms. He defined the atom as the smallest subdivision of an element that can exist. This means that as we continue to reduce the size of the particles of an element, we finally arrive at a submicroscopic, indivisible particle—the atom—which retains all the properties of the element. Today we know that further subdivision is possible, but we obtain particles that no longer possess

Table 1-1 APPROXIMATE COMPOSITION OF THE HUMAN BODY*

Element	*Symbol*	*Per Cent*	*Element*	*Symbol*	*Per Cent*
Oxygen	O	65.0	Iron	Fe	
Carbon	C	18.0	Iodine	I	
Hydrogen	H	10.0	Fluorine	F	
Nitrogen	N	3.0	Silicon	Si	
Calcium	Ca	2.0	Cobalt	Co	
Phosphorus	P	1.0	Zinc	Zn	Traces
Potassium	K	0.4	Copper	Cu	
Sulfur	S	0.3	Manganese	Mn	
Sodium	Na	0.2	Molybdenum	Mo	
Chlorine	Cl	0.2	Chromium	Cr	
Magnesium	Mg	0.1	Selenium	Se	

*Nursing students should memorize the symbols for the elements in the body.

the properties of the element. These particles are called *subatomic particles*, the most familiar of which are protons, neutrons, and electrons.

Symbols

When Dalton postulated the atomic theory, he proposed a set of geometric figures to represent the atoms. Thus ○ represented oxygen; ⊙, hydrogen; ●, carbon; ⦶, nitrogen; and ⊕, sulfur. These symbols were clumsy to use, and they easily gave way to the modern symbols of Berzelius, who used, as far as possible, the capitalized initial letter of the name of the element. Thus O became the symbol for oxygen; H, for hydrogen; and C, for carbon. In cases where two or more names had the same initial letter, two letters were used. In this way, Ca became the symbol for calcium; Cd, for cadmium; Co, for cobalt; and Cr, for chromium. The symbols, however, were to be used internationally, and the names of some elements begin with different letters in different languages; e.g., "iron" is *fer* in French and *eisen* in German. In such cases, the symbols were derived from their Latin names. The symbol for iron therefore became Fe from *ferrum*, that of silver became Ag from *argentum*, etc. Oddly enough, however, Na for sodium and K for potassium are still derived from the German names *natrium* and *kalium*. A symbol may represent the element, an atom of the element, or the atomic weight of the element. Atomic weights will be explained later. The symbol O can represent the element oxygen or 1 atom of oxygen or 16 grams of oxygen. A table showing the symbols of the elements will be found on the *inside front* cover.

Molecules

Two or more atoms can combine chemically to form a molecule. If the atoms are the same, we have a molecule of an elementary substance. This is the case with ordinary oxygen, which is composed of 2 atoms. If the combining atoms are from different elements, the resulting molecule represents a compound. Each molecule of carbon dioxide contains 1 atom of carbon and 2 atoms of oxygen. A molecule of glucose contains 6 carbon, 12 hydrogen, and 6 oxygen atoms.

Formulas

Just as the symbol represents 1 atom of an element, the formula signifies 1 molecule of an element or compound. The formula tells us what elements are contained in the molecule and how many atoms of each element are present. The formula for helium, He, indicates that there is 1 atom in the molecule. Helium along with the other inert gases, neon, argon, krypton, xenon, and radon, are monatomic because, being inert, their atoms will not react with those of other elements or with each other. The formulas for the elementary gases, oxygen, hydrogen, nitrogen, and chlorine are O_2, H_2, N_2, and Cl_2, indicating that there are 2 atoms per molecule. These gases are *diatomic.* The formula for carbon dioxide is CO_2 and for glucose, $C_6H_{12}O_6$.

Compounds

The second type of pure substance we call a compound. Compounds are composed of two or more different elements, chemically combined. When a chemical combination occurs, the starting materials of which the compound is made lose their original properties and the resulting compound can have radically different properties. Consider common table salt, sodium chloride (NaCl). Sodium is a soft, silvery metal and is extremely active. If a small piece of sodium is dropped on the surface of water in a dish or beaker, a violent reaction occurs. Hydrogen gas is given off and the piece of sodium is propelled about on the surface of the water. If it is held in one spot by some obstruction, it often bursts into flame and burns with a bright yellow color. This metal that acts so violent is part of table salt. The other part is the nonmetal chlorine, a greenish-yellow poison gas with a disagreeable odor. Yet not only does no violent reaction occur when sodium chloride is added to water, but sodium chloride is essential for life and a constituent of body fluids and tissues. Many compounds found in the body contain hundreds of atoms per molecule. These giant molecules or macromolecules include the proteins such as hemoglobin in red blood cells and the enzymes involved in all body functions. In these and other compounds, not only must the proper number and kinds of atoms be present in the molecule but these atoms must be arranged in a definite order and the molecule itself must have a definite shape. If there is a deviation in either of these conditions, the molecule cannot function normally and metabolic functions will be impaired. An example of this kind of impaired function is sickle cell anemia, a disease resulting from the abnormal arrangement of the atoms in the hemoglobin molecule.

We see that a knowledge of the arrangement of atoms in molecules is important and that different arrangements form compounds with different properties. Compounds formed from the same numbers and kinds of atoms but with different arrangements are called *isomers.* In the formation of compounds in living organisms, DNA and RNA control not only the composition of the new substances but also their shapes. We shall recognize the importance of the shape of things as we learn more about substances.

Mixtures

Unlike compounds, in which the constituent elements exist in definite proportion by weight, mixtures are highly variable. Mixtures may be composed of elements, as a mixture of iron filings and sulfur. Flour is essentially a mixture of compounds, namely, starch, protein, and salts. Lugol's solution, a medication that is taken internally, is a mixture of the element iodine, the compound potassium iodide, and the compound water. We may even have mixtures of mixtures, as we have in an eggnog.

Frequently mixtures may be quickly recognized because they are heterogeneous. Many mixtures, however, appear homogeneous to the naked eye. Almost anyone would suppose merely from looking at baking powder that it was homogeneous. Howeven, if it is examined under the microscope at least three different substances will be evident.

Compounds have a fixed composition by weight. It is impossible to alter this composition and still have the same compound. On the other hand, it is impossible to have two mixtures just alike. It is true that it is possible to make two mixtures sufficiently alike to serve a practical purpose; e.g., a prescription filled by the pharmacist may be refilled and be practically the same as the original. Baking powder from one batch has the same practical effect as a sample from a previous mixing. Nevertheless, although they seem alike, they cannot be exactly the same. Further, the composition of a mixture, other than a solution, may vary over the widest range, and yet the substance is still a mixture.

Mixtures differ from compounds in that they may be separated into their components by physical means. If we examine a sample of coarse sand even with the naked eye, we find it made up of grains of different colors. If we use a reading glass, a pair of tweezers, sufficient time, and enough patience, we may separate this mixture into several piles of colored grains—surely a physical process.

We may make use of other physical processes. The components of a mixture of iron filings and sulfur have very different densities, and if the mixture is shaken with water in a test tube, the heavy iron settles to the bottom and the lighter sulfur rises to the surface. This same mixture might have been separated by making use of the magnetic property of iron. A horseshoe magnet held over the mixture would draw out the iron filings and leave the sulfur behind.

Often a mixture may be separated into its components by making use of differences in solubility. If a mixture of sugar and white sand is shaken with water, the soluble sugar dissolves and the insoluble sand settles to the bottom. Now, the solution may be separated from the sand by filtration or decantation, the water allowed to evaporate, and the crystals of sugar obtained.

Solutions present a problem to the beginner when he tries to decide whether they are mixtures or compounds. A solution appears homogeneous regardless of the magnification used in its examination. In this respect it resembles a compound. Quantitatively, however, although a given solution is uniform in composition down to the last drop, its composition may be varied by dilution while qualitatively the same homogeneous solution remains. In this it resembles a mixture. Finally, the solution may be separated into its components by a *physical process*, since usually the solvent may be evaporated, leaving the solute. This establishes the fact that a solution is a mixture but a homogeneous mixture.

The properties of the components of a mixture are still evident in the mixture. Since each component is still present as a separate substance, it continues to exhibit its individual characteristics even though it is in a mixture. The properties are additive, which means that the properties of the mixture are the sum of the properties of the components. This principle is well illustrated by a mixture of sugar and butter, the familiar hard sauce. Alone, the sugar is white, tastes sweet, and is in the form of hard crystals large enough to be felt in the mouth. The butter, on the other hand, is yellow and greasy and has a characteristic flavor. Let us mix the two and consider the properties of the mixture. The color is now a lighter yellow. The taste is sweet, but the flavor of the butter is also still apparent. The granular nature of the sugar is still felt, reduced somewhat by the greasiness of the butter. Thus it is seen that each component has contributed its own characteristics to the mixture.

Often a prescription calls for a mixture of drugs. Each one retains its identity and exercises its specific effect in the body, but together they produce the accumulated effect, the cure or relief intended.

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Study Exercises

1. Which of the following are pure substances and which are mixtures? table salt, vinegar, baking soda, baking powder, air, milk, 14-carat gold, platinum, tincture of iodine, and boric acid.
2. What are the differences between a mixture and a compound?
3. Define the following: element, compound, isomer, symbol, and formula.
4. How could the following mixtures be separated into their components? (*a*) sugar and white sand, (*b*) iron filings and charcoal.
5. Compare the properties of the elements sodium and chlorine with the properties of the salt sodium chloride.
6. Mention three properties typical of metals.
7. Write the sumbols for sodium, sulfur, carbon, calcium, chlorine, iron, iodine, potassium, phosphorus, nitrogen, magnesium, manganese, molybdenum, fluorine, and zinc.

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chapter two SOME FUNDAMENTAL CONCEPTS

Both chemistry and physics are exact sciences, in which measurement plays an important part. In medicine, too, measurements are important and tell much more than descriptive terms. If fluids are to be forced on a patient, which statement by the nurse is more useful? "The patient drank a good deal of water and took a little orange juice." "The patient drank 1,500 ml water and 500 ml orange juice." It is obviously more informative to report that a patient's temperature is 104.2°F than to say that he has a high fever.

THE METRIC SYSTEM

The metric system is used in all scientific work and, in all the civilized countries of the world except the United States and the British Commonwealth of Nations, for everyday measurements. In the United States, many of the newer drugs are being marketed with labels describing dosage in metric units. Perhaps the day will come soon when one will no longer need to know drachms, grains, fluid ounces, etc., because of the widespread use of the metric system, which has appropriate units to describe any quantity. The metric system is easier to understand and to use than other systems. To convert from one unit to another, we simply divide or multiply by 10 or 100 or 1,000.

[1]Great Britain is in the process of converting to a basically metric system called the International System of Units or Système International, abbreviated S.I.

The standard unit of length is the meter; of weight, the kilogram; and of time, the second. This system is known as the *mks system* (meter, kilogram, second system). A related system is called the *cgs system* (centimeter, gram, second system).

In measurements, one chooses a measuring unit appropriate to the size of the object or distance to be measured. In English units, to measure long distances, one would use the mile; for shorter distances, the yard; still shorter, the foot; and for still smaller sizes, the inch or fraction of an inch. Other units such as leagues, fathoms, spans, and hands are unfamiliar to most of us. The metric system has appropriate units for measuring all kinds of things, including very small objects such as viruses and molecules.

The meter, used to measure length or distance, is equal to 39.37 inches, slightly longer than a yard. To measure long distances, a unit 1,000 times larger than the meter is used. This is the kilometer, which is approximately 0.62 miles. Other units, the hectometer, 100 meters, and the decameter, 10 meters, are rarely used. Now we consider units smaller than the meter. The unit one-tenth the size of the meter is the decimeter. The prefix deci- means one-tenth. Hence, 10 decimeters equal 1 meter. The prefix centi- means one-hundredth, so 100 centimeters equal 1 meter, and the prefix milli- indicates one-thousandth, so 1,000 millimeters equal 1 meter. The decimeter, which is rarely used in scientific measurements, is approximately equal to 3.94 inches, the centimeter is 0.39 inches, and the millimeter 0.04 inches. To measure the size of small things such as microbes, a unit 1,000 times smaller than a millimeter is used. This is the micron.[1] A millimicron is 1,000 times smaller than a micron. A unit used to measure the length of X rays and average-size molecules is one-tenth the size of a millimicron and is called the angstrom unit. We are now considering objects too small to be seen, even with the electron microscope. To summarize:

1,000 meters (m) = 1 kilometer (km)
10 decimeters (dm) = 1 meter
100 centimeters (cm) = 1 meter
1,000 millimeters (mm) = 1 meter
1,000,000 micra (μ) * = 1 meter
or
1,000 micra = 1 mm
1,000,000 millimicra (mμ) = 1 mm
10 angstrom units (Å) = 1 mμ

With these relationships in mind, the systems for weights and volumes are simple. The unit of weight is the kilogram (kg), which consists of 1,000 grams (Gm) and is equal to 2.2 pounds.

10 decigrams (dg) = 1 gram (Gm)
100 centigrams (cg) = 1 Gm
1,000 milligrams (mg) = 1 Gm
1,000 micrograms (μg) = 1 mg

One gram is equal to 0.035 ounces.

In measuring volumes, only three units are commonly used: the kiloliter, the liter, and the milliliter. The liter equals 1.06 quarts.

1,000 liters (l) = 1 kiloliter (kl)
1,000 milliliters (ml) = 1 liter

Summarizing the units most commonly used in medicine:

1 cm = 0.01 meter
1 mm = 0.001 meter
1 μ = 0.001 mm
1 mμ = 0.000,001 mm
1 mμ = 0.001 μ
1 Å = 0.1 mμ
1 Å = 0.000,000,1 mm (1×10^{-7})
1 Å = 0.000,000,01 cm (1×10^{-8})

1,000 liters = 1 kl
1 ml = 0.001 liter
1,000 Gm = 1 kg
1 mg = 0.001 Gm

and also the meter, the liter, and the gram.

Area is measured in square units derived from linear units; e.g., square centimeters may be written sq cm or cm^2. The same procedure applies to cubic measurements; we have cubic centimeters, or cm^3, or the very common cc.

To recap:

100 cm = 1 meter
10 mm = 1 cm
1,000 mm = 1 meter
1,000 mg = 1 Gm
1,000 ml = 1 liter

In converting from one unit to another within a system, small units are changed into larger units by dividing by the number of small units in the larger unit. To change larger units into smaller units, multiply the larger unit by the number of smaller units that it contains. Since all the conversion factors are multiples of 10, this amounts to moving the decimal point to the right or to the left.

Problems

The barometric pressure is expressed as 760 mm of mercury. You wish to change this measurement to centimeters. Millimeter is the smaller unit. There are 10 mm in 1 cm. Dividing 760 mm by 10 gives us the answer, 76 cm. In a whole number there is no need for a decimal point, but one is understood to be at the right of the number, as 760.0. To divide

[1]The micron is designated by the lower-case Greek letter mu which is written μ. Micra is often used as the plural of micron.

Table 2-1 METRIC UNITS

Weight	*Length*	*Volume*
Gram	Meter	Liter
1,000 Gm = 1 kg	100 cm = 1 meter	1,000 liters = 1 kl
1,000 mg = 1 Gm	1,000 mm = 1 meter	1,000 ml = 1 liter

*Chemists and physicists use the abbreviation g for gram, but in medicine, to avoid confusion between gram and grain (apothecaries' system), the abbreviation Gm is used.

by 10, one moves the decimal point one space to the left.

An individual is 169 cm tall. To change this measurement to meters, we realize that centimeter is the smaller unit and that there are 100 cm in 1 meter. Moving our imaginary decimal point 2 places to the left gives the answer, 1.69 meters.

A patient voids 1,250 ml of urine in a 24-hr period. Since there are 1,000 ml in 1 liter, move the imaginary decimal point three places to the left to find the volume in liters, 1.250 liters.

In changing from smaller units to larger units, divide. Move the decimal point the appropriate number of places to the left. Change 50 mg to Gm. Divide by 1,000. Move the decimal point three places to the left. Answer, .050 Gm. As a precaution and to alert the reader to the decimal, if a number starts with a decimal, a zero is placed before the decimal thus, 0.050.

If a medication order was written as .5 Gm and the decimal was blurred or indistinct, one might read this order as 5 Gm. This would be a gross error and the patient would receive 10 times the intended dose. By writing 0.5, the reader expects a decimal and looks for it. The possible error is avoided.

To change larger units to smaller ones, multiply by moving the decimal point (real or imagined) the appropriate number of places to the right.

A baby is 55 cm long. Change this measurement to millimeters. Imagine the decimal point as 55.0. There are 10 mm in 1 cm. Multiply by 10 by moving the decimal point one place to the right. Answer, 550 mm.

An object is 0.25 meter wide. To express as centimeters, multiply by the number of centimeters in a meter, 100. Moving the decimal point two places to the right gives us the answer, 25 cm.

Express 2.65 Gm as milligrams. There are 1,000 mg in 1 Gm. Multiply by 1,000. Move the decimal point three places to the right to get 2,650 mg.

In these problems, one can only use the same types of units. Grams cannot be changed to volume units without additional information.

To describe volume of solids, a unit called the cubic centimeter (cc) is used in place of the milliliter. This is an older measurement in science, which is gradually disappearing. One could go to a dairy and ask for a cubic foot of milk. If the clerk filled a container whose capacity was 1 cu ft, it would hold 1 cu ft milk. However, it is more acceptable to sell milk by the gallon than by the cubic foot.

Since the cubic centimeter and the milliliter are almost exactly the same size, these units are often used interchangeably.

When we start learning a foreign language, we think in our native tongue and translate; after a while we can think in the foreign language. When one first starts using the metric system, one will probably refer to English, or apothecaries', units. Here are some useful equivalents:

2.2 lb = 1 kg
1 lb = 454 Gm
1 in. = 2.5 cm
39.3 in. = 1 meter
1 cm = 0.39 in.

1 qt, or 32 fl oz ≅ 1 liter, or 1,000 ml[1]
1 pt, or 16 fl oz ≅ 500 ml
1 fl oz = 30 ml
1 oz ≅ 30 gm

Since dosages are sometimes prescribed in terms of body weight, particularly for infants, it may be necessary to change weights from English to metric units in order to compute the proper dosage. Since 2.2 lb = 1 kg, to change pounds to kilograms divide by 2.2 and to change kilograms to pounds, multiply by 2.2. For instance, an infant weighs 5½ lbs; 5.5 divided by 2.2 gives 2.5 kg.

ATOMIC WEIGHTS

When we realize that atoms are too small to be seen with the electron microscope, we also realize that an individual atom is too small to be weighed. It is however possible to determine relative weights. One could weigh *x* trillion atoms of hydrogen and *x* trillion atoms of helium and discover that helium is four times heavier than hydrogen. Relative weights of atoms are extremely useful and may be obtained with great accuracy. In expressing relative weights, it is necessary to set up a standard or basis of comparison. In the early days, hydrogen, the lightest element, was assigned a weight of 1 and was used as the standard or reference point from which the weights of the other elements were computed. In other words, if hydrogen has a weight of 1, then carbon, which is 12 times heavier than hydrogen, will have a weight of 12, and oxygen, which is 16 times heavier than hydrogen, will have a weight of 16, etc. Subsequently, oxygen 16 was used as the standard, but at present carbon, atomic weight 12, is the official standard against which the atomic weights are computed. These relative weights have no units; they are simply comparisons.

Having established relative weights, the next step was to translate these relative weights into actual weights. How much hydrogen will weigh one gram? The number of hydrogen atoms weighing one gram is 602,000,000,000,000,000,000,000 or 6.02×10^{23} atoms. We have now the gram-atomic weight of hydrogen, which is 1 gram. The gram-atomic weights of the other elements are the weights in grams of an equal number of the atoms of that element.

The symbol of an element can represent the name of the element, an atom of the element, or the atomic weight of the element. In the formula H_2O, H taken twice can mean 2 atoms of hydrogen or it can mean 2 Gm hydrogen; O can mean 1 atom of oxygen or 16 Gm oxygen.

MOLECULAR WEIGHTS

The molecular weight of a substance is the sum of the atomic weights of the atoms contained in the molecule. The molecular weight of carbon dioxide, CO_2, is 44, for the weight of 1 atom of carbon is 12, and the weight of 2 atoms of oxygen is 32 (2×16).

The molecular weight of water, H_2O, is 18. Hydrogen, weight 1 times 2, plus oxygen 16. Since the atomic weights are numbers that show the relative weights of the atoms referred to carbon 12, the molecular weight of a substance is a number showing the weight of the molecule compared to carbon 12.

The gram-molecular weight or the *mole* of a substance is the number of *grams* of a substance equal to the sum of its gram-atomic weights or to its molecular weight; the gram-molecular weight of CO_2 is 44 grams. The weight of one mole of water is 18 gm.

law of definite composition

Every time we analyze a given compound, regardless of its source, if it is a pure sample,

[1]The symbol ≅ means approximately equal.

we always find that it contains the same elements and in the same proportion by weight. That is, no matter how we obtain pure sodium chloride, table salt, whether it is formed from the direct union of sodium and chloride, mined from a salt mine, crystallized from brine, or obtained as a by-product in a chemical reaction, it will always contain sodium and chloride and the elements will be present in the ratio of 23 parts of sodium to 35.3 parts of chlorine. Furthermore, if sodium and chloride react directly, 23 Gm sodium will react exactly with 35.5 Gm chlorine to form 58.5 Gm sodium chloride. These observations are summarized in the law of definite composition which states that *in every pure sample of a compound, the elements are always present in a definite proportion by weight*. By consulting the table of atomic weights, one can see that these proportions are related to the atomic weights of the elements involved.

Study Exercises

1. Distinguish between mass and weight.
2. An infant was reported to be 38 cm long and having a weight of 3,000 Gm. Convert these figures into units of inches and pounds.
3. A student is 5 ft 7 in. tall and weighs 125 lb. Express these figures in metric units.
4. The average diameter of a human red blood cell is 7.5 μ, platelets are 2 to 3 μ in diameter, and the average diameter of pathogenic bacilli is 0.5 μ. Express these measurements as millimeters.
5. The "recommended daily dietary allowances" for the average female, age 18 to 22, for iodine, iron, and magnesium are, respectively, 100 μg, 18 mg, and 300 mg. Convert these figures to grams.
6. Define the following: atomic weight, molecular weight, gram-atomic weight, gram-molecular weight, and mole.
7. What is the law of definite composition?

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