FACETS OF CHEMISTRY 1.1

Experiments Leading to the Discovery of Subatomic Particles

Our current knowledge of atomic structure was pieced together from facts obtained from experiments by scientists that began in the nineteenth century. In 1834, Michael Faraday discovered that the passage of electricity through aqueous solutions could cause chemical changes, which was the first hint that matter was electrical in nature. Later in that century, scientists began to experiment with *gas discharge tubes* in which a high-voltage electric current was passed through a gas at low pressure in a glass tube (Figure 1). Such a tube is fitted with a pair of metal *electrodes,* and when the electricity begins to flow between them, the gas in the tube glows. This flow of electricity is called an *electric discharge,* which is how the tubes got their names. (Modern neon signs work this way.)

The physicists who first studied this phenomenon did not know what caused the tube to glow, but tests soon revealed that negatively charged particles were moving from the negative electrode (the *cathode*) to the positive electrode (the *anode*). The physicists called these emissions *rays,* and because the rays came from the cathode, they were called *cathode rays*.

In 1897, the British physicist J. J. Thomson constructed a special gas discharge tube to make quantitative measurements of the properties of cathode rays. In some ways, the *cathode ray tube* he used was similar to a TV picture tube, as Figure 2 shows. In Thomson's tube, a beam of cathode rays was focused on a glass surface coated with a phosphor that glows when the cathode rays strike it (point 1). The cathode ray beam passed between the poles of a magnet and between a pair of metal electrodes that could be given electrical charges. The magnetic field tends to bend the beam in one direction (to point 2) while the charged electrodes bend the beam in the opposite direction (to point 3). By adjusting the charge on the electrodes, the two effects can be made to cancel, and from the amount of charge on the electrodes required to balance the effect of the magnetic field, Thomson was able to calculate the first bit of quantitative information about a cathode ray particle—the ratio of its charge to its mass (often expressed as e/m, where e stands for charge and m stands for mass). The charge-to-mass ratio has a value of -1.76×10^8 coulombs/gram, where the coulomb (C) is a standard unit of electrical charge and the negative sign reflects the negative charge on the particle.

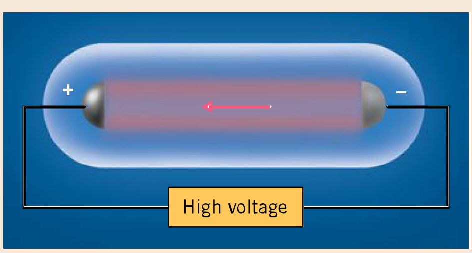


FIGURE 1 ***A gas discharge tube.*** Cathode rays flow from the negatively charged cathode to the positively charged anode.

Many experiments were performed using the cathode ray tube and they demonstrated that cathode ray particles are in all matter. They are, in fact, *electrons*.

Measuring the Charge and Mass of the Electron

In 1909, a researcher at the University of Chicago, Robert Millikan, designed a clever experiment that enabled him to measure the electron's charge (Figure 3). During an experiment he would spray a fine mist of oil droplets above a pair of parallel metal plates, the top one of which had a small hole in it. As the oil drops settled, some would pass through this hole into the space between the plates, where he would irradiate them briefly with X rays. The X rays knocked electrons off molecules in the air, and the electrons became attached to the oil drops, which thereby were given an electrical charge. By observing the rate of fall of the charged drops both when the metal plates were electrically charged and when they were not, Millikan was able to calculate the amount of charge carried by each drop. When he examined his results, he found that all the values he obtained were whole-number multiples of -1.60×10^{-19} C. He reasoned that since a drop could only pick up whole numbers of electrons, this value must be the charge carried by each individual electron.

Once Millikan had measured the electron's charge, its mass could then be calculated from Thomson's charge-to-mass ratio. This mass was calculated to be 9.09×10^{-28} g.

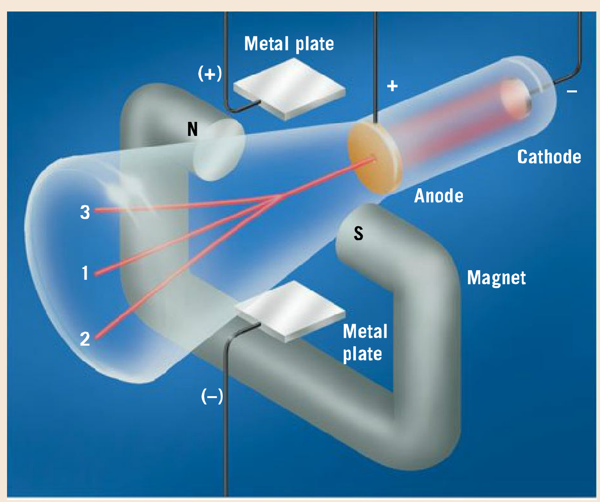


FIGURE 2 ***Thomson's cathode ray tube,*** which was used to measure the charge-to-mass ratio for the electron.

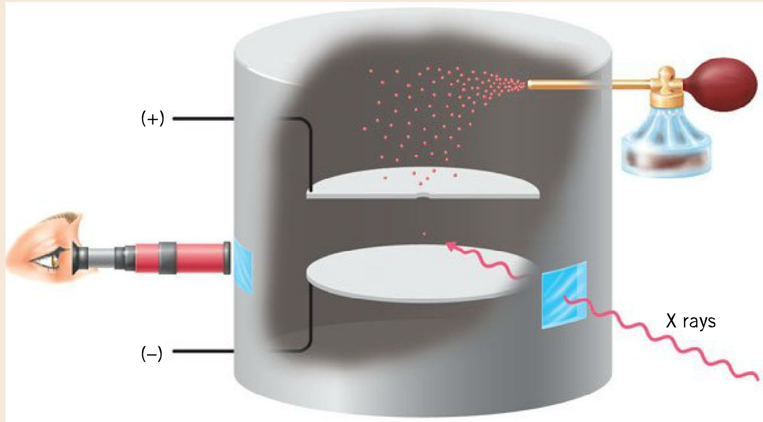


FIGURE 3 ***Millikan's oil drop experiment.*** Electrons, which are ejected from air molecules by the X rays, are picked up by very small drops of oil falling through the tiny hole in the upper metal plate. By observing the rate of fall of the charged oil drops, with and without electrical charges on the metal plates, Millikan was able to calculate the charge carried by an electron.

More precise measurements have since been made, and the mass of the electron is currently reported to be $9.1093897 \times 10^{-28}$ g.

Discovery of the Proton

The removal of electrons from an atom gives a positively charged particle (called an *ion*). To study these, a modification was made in the construction of the cathode ray tube to produce a new device called a *mass spectrometer*. This apparatus is described in Facets of Chemistry 1.2 and was used to measure the charge-to-mass ratios of positive ions. These ratios were found to vary, depending on the chemical nature of the gas in the discharge tube, showing that their masses also varied. The lightest positive particle observed was produced when hydrogen was in the tube, and its mass was about 1800 times as heavy as an electron. When other gases were used, their masses always seemed to be whole-number multiples of the mass observed for hydrogen atoms. This suggested the possibility that clusters of the positively charged particles made from hydrogen atoms made up the positively charged particles of other gases. The hydrogen atom, minus an electron, thus seemed to be a fundamental particle in all matter and was named the *proton,* after the Greek word *proteios,* meaning "of first importance."

Discovery of the Atomic Nucleus

Early in the twentieth century, Hans Geiger and Ernest Marsden, working under Ernest Rutherford at Great Britain's Manchester University, studied what happened when *alpha rays* hit thin metal foils. Alpha rays are composed of particles having masses four times those of the proton and bearing two positive charges; they are emitted by certain unstable atoms in a phenomenon called *radioactivity.* Most of the alpha particles sailed right on through as if the foils were virtually empty space (Figure 4). A significant number of alpha particles, however, were deflected at very large angles. Some were even deflected backward, as if they had hit stone walls. Rutherford was so astounded that he compared the effect to that of firing a 15 in. artillery shell at a piece of tissue paper and having it come back and hit the gunner! From studying the angles of deflection of the particles, Rutherford reasoned that only something extraordinarily massive and positively charged could cause such an occurrence. Since most of the alpha particles went straight through, he further reasoned

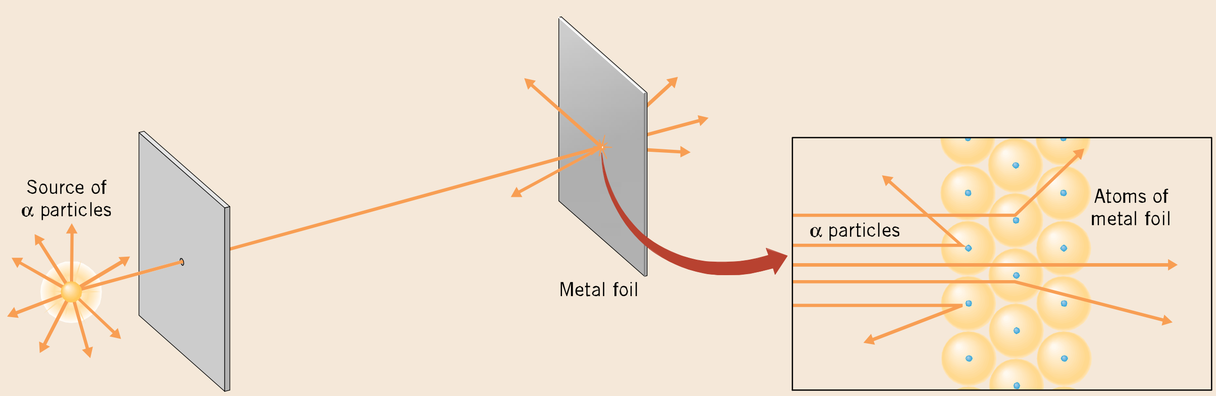


FIGURE 4 ***Alpha particles are scattered in all directions by a thin metal foil.*** Some hit something very massive head-on and are deflected backward. Many sail through. Some, making near misses with the massive "cores" (nuclei), are still deflected, because alpha particles have the same kind of charge (+) as these cores.

that the metal atoms in the foils must be mostly empty space. Rutherford's ultimate conclusion was that virtually all of the mass of an atom must be concentrated in a particle having a very small volume located in the center of the atom. He called this massive particle the atom's *nucleus.*

Discovery of the Neutron

From the way alpha particles were scattered by a metal foil, Rutherford and his students were able to estimate the number of positive charges on the nucleus of an atom of the metal. This had to be equal to the number of protons in the nucleus, of course. But when they computed the nuclear mass based on this number of protons, the value always fell short of the actual mass. In fact, Rutherford found that only about half of the nuclear mass could be accounted for by protons. This led him to suggest that there were other particles in the nucleus that had a mass close to or equal to that of a proton, but with no electrical charge. This suggestion initiated a search that finally ended in 1932 with the discovery of the *neutron* by Sir James Chadwick, a British physicist.

trons in an atom surround the nucleus and fill the remaining volume of the atom. (*How* the electrons are distributed around the nucleus is the subject of Chapter 8.) The properties of the subatomic particles are summarized in Table 1.3, and the general structure of the atom is illustrated in Figure 1.13.

Physicists have discovered a large number of subatomic particles, but protons, neutrons, and electrons are the only ones that will concern us at this time.

Protons are in all nuclei. Except for ordinary hydrogen, all nuclei also contain neutrons.

Notice that two of the subatomic particles carry electrical charges. Protons carry a single unit of **positive charge,** which is one type of electrical charge. Electrons carry one unit of the opposite charge, a **negative charge.** These electrical charges have the property that like charges repel each other and opposite charges attract, so negatively charged electrons are attracted to positively charged protons. In fact, it is this attraction that holds the electrons around the nucleus. Neutrons have no charge and are said to be electrically neutral (hence the name *neutron*).

Because of their identical charges, electrons repel each other. The repulsions between the electrons keep them spread out throughout the volume of the atom, and it is the *balance* between the attractions the electrons feel toward the nucleus and the repulsions they feel toward each other that controls the sizes of atoms.

Protons also repel each other, but they are able to stay together in the small volume of the nucleus because their repulsions are apparently offset by powerful nuclear forces that involve other subatomic particles we will not study.

Matter as we generally find it in nature appears to be electrically neutral, which means that it contains equal numbers of positive and negative charges. Therefore, *in a neutral atom, the number of electrons must equal the number of protons.*

The proton and neutron are much more massive than the electron, so in any atom almost all of the atomic mass is contributed by the particles that are found in the nucleus. (The mass of an electron is only about 1/1800 of that of a proton or neutron.) It is also interesting to note, however, that the diameter of the atom is approximately 10,000 times the diameter of its nucleus, so almost all the *volume* of an atom is occupied by its electrons, which fill the space around the nucleus. (To place this on a more meaningful scale, if the nucleus was 1 ft in diameter, it would lie at the center of an atom with a diameter of approximately 1.9 miles!)

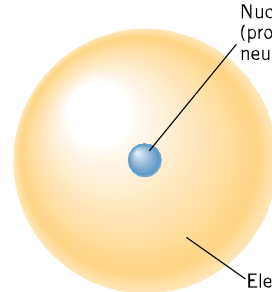


FIGURE 1.13 ***The internal structure of an atom.*** An atom is composed of a tiny nucleus that holds all the protons and neutrons, plus electrons that fill the space outside the nucleus.

TABLE 1.3 PROPERTIES OF SUBATOMIC PARTICLES

Particle	Mass (g)	Mass (u)	Electrical Charge	Symbol
Electron	$9.1093897 \times 10^{-28}$	$5.48579903 \times 10^{-4}$	1−	$^{0}_{-1}e$
Proton	$1.6726231 \times 10^{-24}$	1.007276470	1+	$^{1}_{1}H^{+}, ^{1}_{1}p$
Neutron	$1.6749286 \times 10^{-24}$	1.008664904	0	$^{1}_{0}n$

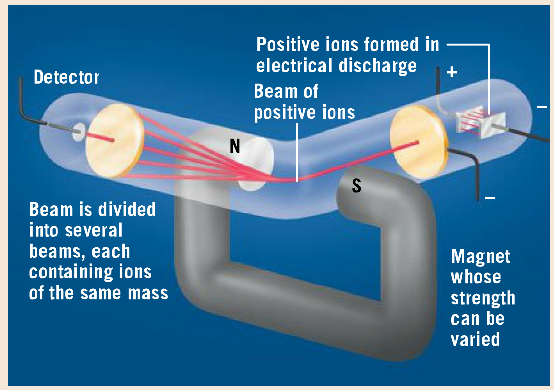
FACETS OF CHEMISTRY 1.2

The Mass Spectrometer and the Experimental Measurement of Atomic Masses

When a spark is passed through a gas, electrons are knocked off the gas molecules. Because electrons are negatively charged, the particles left behind carry positive charges; they are called *positive ions.* These positive ions have different masses, depending on the masses of the molecules from which they are formed. Thus, some molecules have large masses and give heavy ions, while others have small masses and give light ions.

The device that is used to study the positive ions produced from gas molecules is called a *mass spectrometer* (illustrated in the figure at the right). In a mass spectrometer, positive ions are created by passing an electrical spark (called an *electric discharge*) through a sample of the particular gas being studied. As the positive ions are formed, they are attracted to a negatively charged metal plate that has a small hole in its center. Some of the positive ions pass through this hole and travel onward through a tube that passes between the poles of a powerful magnet.

One of the properties of charged particles, both positive and negative, is that their paths become curved as they pass through a magnetic field. This is exactly what happens to the positive ions in the mass spectrometer as they pass between the poles of the magnet. However, the extent to which their paths are bent depends on the masses of the ions. This is because the path of a heavy ion, like that of a speeding cement truck, is difficult to change, but the path of a light ion, like that of a motorcycle, is influenced more easily. As a result, heavy ions emerge from between the magnet's poles along different lines than the lighter ions. In effect, an entering beam containing ions of different masses is sorted by the magnet into a number of beams, each containing ions of the same mass. This spreading out of the ion beam thus produces an array of different beams called a *mass spectrum.*



In practice, the strength of the magnetic field is gradually changed, which sweeps the beams of ions across a detector located at the end of the tube. As a beam of ions strikes the detector, its intensity is measured and the masses of the particles in the beam are computed based on the strength of the magnetic field and the geometry of the apparatus.

Among the benefits derived from measurements using the mass spectrometer are very accurate isotopic masses and relative isotopic abundances. These serve as the basis for the very precise values of the atomic masses that you find in the periodic table.

Atomic numbers and mass numbers describe isotopes

What distinguishes one element from another is the number of protons in the nuclei of its atoms, because *all the atoms of a particular element have an identical number of protons.* In fact, this allows us to redefine an **element** as *a substance whose atoms all contain the identical number of protons.* Thus, each element has associated with it a unique number, which we call its **atomic number (*Z*),** that equals the number of protons in the nuclei of any of its atoms.

$$\text{atomic number, } Z = \text{number of protons}$$

What makes isotopes of the same element different are the numbers of neutrons in their nuclei. *The* **isotopes** *of a given element have atoms with the same number of protons but different numbers of neutrons.* The numerical sum of the protons and neutrons in the atoms of a particular isotope is called the **mass number (*A*)** of the isotope.

$$\text{mass number, } A = (\text{number of protons}) + (\text{number of neutrons})$$

Therefore, every isotope is fully defined by two numbers, its atomic number and its mass number. Sometimes these numbers are added to the left of the chemical symbol as a subscript and a superscript, respectively. Thus, if X stands for the chemical symbol for the element, an isotope of X is represented as

$$^{A}_{Z}X$$

The isotope of uranium used in nuclear reactors, for example, can be symbolized as follows:

$$\text{mass number (protons + neutrons)} \rightarrow {}^{235}_{\ 92}\text{U} \leftarrow \text{atomic number (number of protons)}$$

uranium-235

As indicated, the name of this isotope is uranium-235. Each neutral atom contains 92 protons and $(235 - 92) = 143$ neutrons as well as 92 electrons. In writing the symbol for the isotope, the atomic number is often omitted because it is really redundant. Every atom of uranium has 92 protons, and every atom that has 92 protons is an atom of uranium. Therefore, this uranium isotope can be represented simply as ^{235}U.

It is useful to remember that for a neutral atom, the atomic number equals both the number of protons and the number of electrons.

In naturally occurring uranium, a more abundant isotope is ^{238}U. Atoms of this isotope also have 92 protons, but the number of neutrons is 146. Thus, atoms of ^{235}U and ^{238}U have the identical number of protons but differ in the numbers of neutrons.

In general, the mass number of an isotope differs slightly from the atomic mass of the isotope. For instance, the isotope ^{35}Cl has an atomic mass of 34.968852 u. In fact, the *only* isotope that has an atomic mass equal to its mass number is ^{12}C; *by definition* the mass of this atom is exactly 12 u.

PRACTICE EXERCISE 6: Write the symbol for the isotope of plutonium (Pu) that contains 146 neutrons.

PRACTICE EXERCISE 7: How many protons, neutrons, and electrons are in each atom of $^{35}_{17}Cl$?

1.7 The periodic table is used to organize and correlate facts

When we study different kinds of substances, we find that some are elements and others are compounds. Among compounds, some are composed of discrete molecules. Others, as you will learn, are made up of atoms that have acquired electrical charges. For elements such as sodium and chlorine, we mentioned metallic and nonmetallic properties. If we were to continue on this way, without attempting to build our subject around some central organizing structure, it would not be long before we became buried beneath a mountain of information in the form of seemingly unconnected facts.

The need for organization was recognized by many early chemists, and there were numerous attempts to discover relationships among the chemical and physical properties of the elements. A number of different sequences of elements were tried in the search for some sort of order or pattern. A few of these arrangements came quite close, at least in some respects, to our current periodic table, but either they were flawed in some way or they were presented to the scientific community in a manner that did not lead to their acceptance.

Mendeleev created the first periodic table

The periodic table we use today is based primarily on the efforts of a Russian chemist, Dmitri Ivanovich Mendeleev (1834–1907), and a German physicist, Julius Lothar Meyer (1830–1895). Working independently, these scientists developed similar periodic tables only a few months apart in 1869. Mendeleev is usually given the credit, however, because he had the good fortune to publish first.

Mendeleev was preparing a chemistry textbook for his students at the University of St. Petersburg. Looking for some pattern among the properties of the elements, he found that when he arranged them in order of increasing atomic mass, similar chemical properties were repeated over and over again at regular intervals. For instance, the elements lithium (Li), sodium (Na), potassium (K), rubidium (Rb), and cesium (Cs) are soft metals that are very reactive toward water. They form compounds with chlorine that have a 1-to-1 ratio of metal to chlorine. Similarly, the elements that immediately follow each of these also constitute a set with similar chemical properties. Thus, beryllium (Be) follows lithium, magnesium (Mg) follows sodium, calcium (Ca) follows potassium, strontium (Sr) follows rubidium, and barium (Ba) follows cesium. All of these elements form a water-soluble chlorine compound with a 1-to-2 metal to chlorine atom ratio. Mendeleev used such observations to construct his **periodic table,** which is illustrated in Figure 1.14.

Periodic refers to the recurrence of properties at regular intervals.

The elements in Mendeleev's table are arranged in order of increasing atomic mass. When the sequence is broken at the right places and stacked, the elements fall naturally into columns, called **groups,** in which the elements of a given group have similar chemical properties. The rows themselves are called **periods.**

Mendeleev's genius rested on his placing elements with similar properties in the same group even when this left occasional gaps in the table. For example, he placed arsenic (As) in Group V under phosphorus because its chemical properties were similar to those of phosphorus, even though this left gaps in Groups III and IV. Mendeleev reasoned, correctly, that the elements that belonged in these gaps had simply not yet been discovered. In fact, on the basis of the location of these

Periods	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H 1							
2	Li 7	Be 9.4	B 11	C 12	N 14	O 16	F 19	
3	Na 23	Mg 24	Al 27.3	Si 28	P 31	S 32	Cl 35.5	
4	K 39	Ca 40	—44	Ti 48	V 51	Cr 52	Mn 55	Fe 56, Co 59 Ni 59, Cu 63
5	(Cu 63)	Zn 65	—68	—72	As 75	Se 78	Br 80	
6	Rb 85	Sr 87	?Yt 88	Zr 90	Nb 94	Mo 96	—100	Ru 104, Rh 104 Pd 105, Ag 108
7	(Ag 108)	Cd 112	In 113	Sn 118	Sb 122	Te 128	I 127	
8	Cs 133	Ba 137	?Di 138	?Ce 140	—	—	—	——
9	—	—	—	—	—	—	—	
10	—	—	?Er 178	?La 180	Ta 182	W 184	—	Os 195, Ir 197 Pt 198, Au 199
11	(Au 199)	Hg 200	Tl 204	Pb 207	Bi 208	—		
12	—	—	—	Th 231	—	U 240	—	—— ——

FIGURE 1.14 ***The first periodic table.*** Mendeleev's periodic table roughly as it appeared in 1871. The numbers next to the symbols are atomic masses.

gaps Mendeleev was able to predict with remarkable accuracy the properties of these yet-to-be-found substances. His predictions helped serve as a guide in the search for the missing elements.

The elements tellurium (Te) and iodine (I) caused Mendeleev some problems. According to the best estimates at that time, the atomic mass of tellurium was greater than that of iodine. Yet if these elements were placed in the table according to their atomic masses, they would not fall into the proper groups required by their properties. Therefore, Mendeleev switched their order and in so doing violated his ordering sequence. (Actually, he believed that the atomic mass of tellurium had been incorrectly measured, but this wasn't so.)

The table that Mendeleev developed is in many ways similar to the one we use today. One of the main differences, though, is that Mendeleev's table lacks the column containing the elements helium (He) through radon (Rn). In Mendeleev's time, none of these elements had yet been found because they are relatively rare and because they have virtually no tendency to undergo chemical reactions. When these elements were finally discovered, beginning in 1894, another problem arose. Two more elements, argon (Ar) and potassium (K), did not fall into the groups required by their properties if they were placed in the table in the order required by their atomic masses. Another switch was necessary and another exception had been found. It became apparent that atomic mass was not the true basis for the periodic repetition of the properties of the elements. To determine what the true basis was, however, scientists had to await the discoveries of the atomic nucleus, the proton, and atomic numbers.

The modern periodic table arranges elements by atomic number

Once atomic numbers had been discovered, it was soon realized that the elements in Mendeleev's table are arranged in precisely the order of increasing atomic number. In other words, if we take atomic numbers as the basis for arranging the elements in sequence, no annoying switches are required and the elements Te and I or Ar and K are no longer a problem. The fact that it is the atomic number—the number of protons in the nucleus of an atom—that determines the order of elements in the table is very significant. We will see later that this has important implications with regard to the relationship between the number of electrons in an atom and the atom's chemical properties.

The modern periodic table is shown in Figure 1.15 and also appears on the inside front cover of the book. We will refer to the table frequently, so it is important for you to become familiar with it and with some of the terminology applied to it.

Special terminology is associated with the periodic table

As in Mendeleev's table, the elements are arranged in rows that we call **periods,** but here they are arranged in order of increasing atomic number. For identification purposes, the periods are numbered. We will find these numbers useful later on. Below the main body of the table are two long rows of 14 elements each. These actually belong in the main body of the table following La ($Z = 57$) and Ac ($Z = 89$), as shown in Figure 1.16. They are almost always placed below the table simply to conserve space. If the fully spread-out table is printed on one page, the type is so small that it's difficult to read. Notice that in the fully extended form of the table, with all the elements arranged in their proper locations, there is a great deal of empty space. An important requirement of a detailed atomic theory, which we will get to in Chapter 8, is that it must explain not only the repetition of properties, but also why there is so much empty space in the table.

Recall that the symbol Z stands for atomic number.

Again, as in Mendeleev's table, the vertical columns are called **groups.** However, there is not uniform agreement among chemists on how they should be

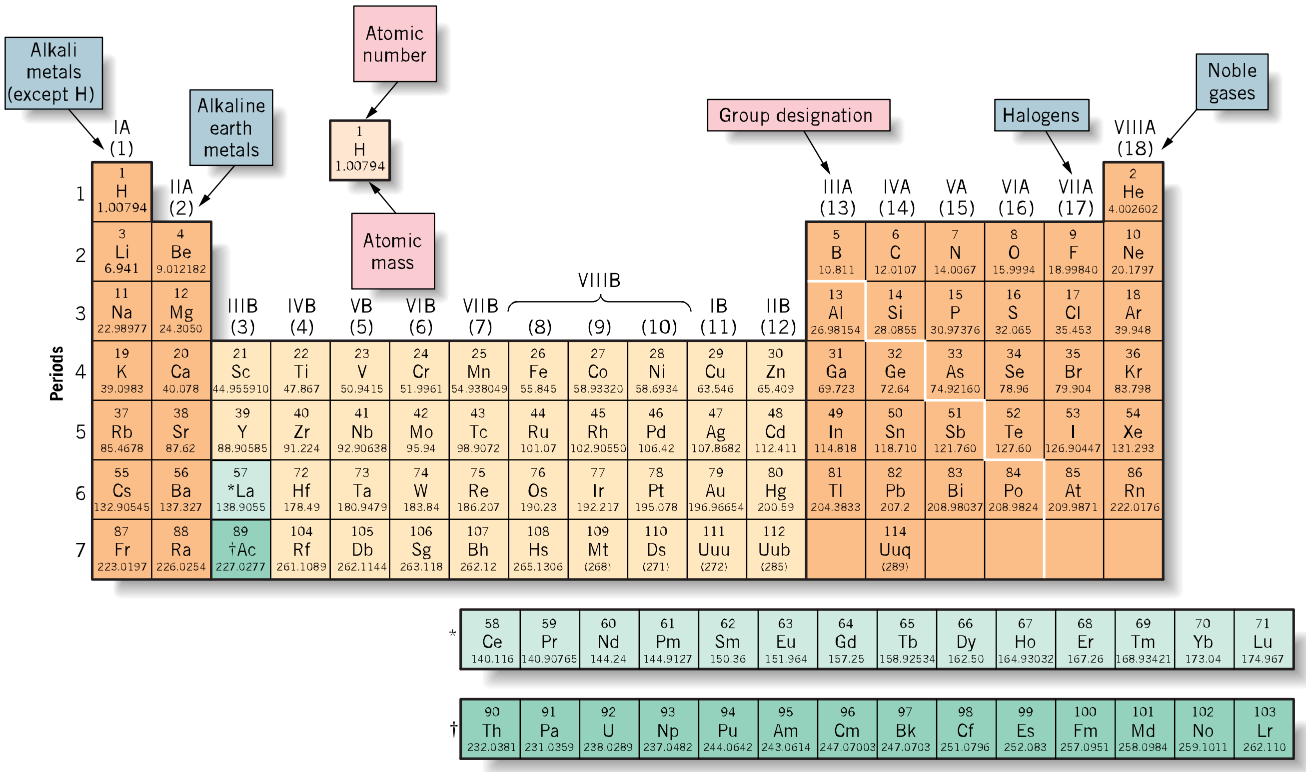


FIGURE 1.15 ***The modern periodic table.***

numbered. In the past, the groups were labeled with Roman numerals and divided into A groups and B groups, separated by the three short columns headed by Fe, Co, and Ni (Group VIIIB), as indicated in Figure 1.15. This corresponds quite closely to Mendeleev's original designations and is preferred by many chemists in North America. However, another version of the table, popular in Europe, has the first seven groups from left to right labeled A, followed by the three short columns headed by Fe, Co, and Ni, and then the next seven groups labeled B. In an attempt to standardize the table, the International Union of Pure and Applied Chemistry (the IUPAC), an international body of scientists responsible for setting standards in chemistry, officially adopted a third system in which the groups are simply numbered sequentially from left to right using Arabic numerals. Thus, Group IA in the North American table is Group 1 in the IUPAC

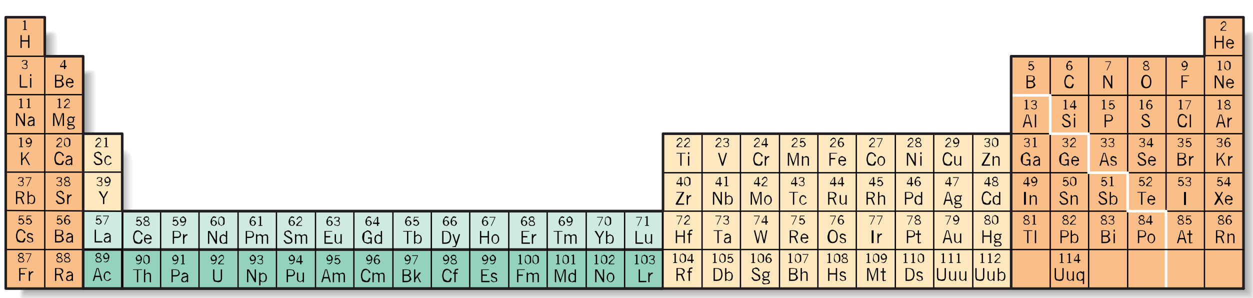
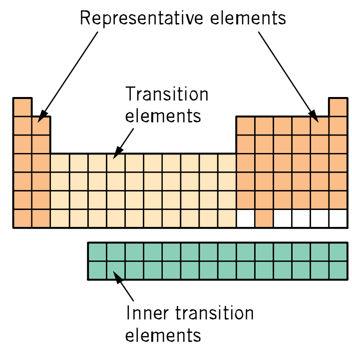


FIGURE 1.16 ***Extended form of the periodic table.*** The two long rows of elements below the main body of the table in Figure 1.15 are placed in their proper places in this table.

table, and Group VIIA in the North American table is Group 17 in the IUPAC table. In Figure 1.15 and on the inside front cover of the book, we have used both the North American labels as well as those preferred by the IUPAC. Because of the lack of uniform agreement among chemists on how the groups should be specified, we will use the A-group/B-group designations in Figure 1.15 when we wish to specify a particular group.

As we have already noted, the elements in a given group bear similarities to each other. Because of such similarities, groups are sometimes referred to as **families of elements.** The elements in the longer columns (the A groups) are known as the **representative elements** or **main group elements.** Those that fall into the B groups in the center of the table are called **transition elements.** The elements in the two long rows below the main body of the table are the **inner transition elements,** and each row is named after the element that it follows in the main body of the table. Thus, elements 58–71 are called the **lanthanide elements** because they follow lanthanum ($Z = 57$), and elements 90–103 are called the **actinide elements** because they follow actinium ($Z = 89$).



Some of the groups have acquired common names. For example, except for hydrogen, the Group IA elements are metals. They form compounds with oxygen that dissolve in water to give solutions that are strongly alkaline, or caustic. As a result, they are called the **alkali metals** or simply the *alkalis.* The Group IIA elements are also metals. Their oxygen compounds are alkaline, too, but many compounds of the Group IIA elements are unable to dissolve in water and are found in deposits in the ground. Because of their properties and where they occur in nature, the Group IIA elements became known as the **alkaline earth metals.**

On the right side of the table, in Group VIIIA, are the **noble gases.** They used to be called the inert gases until it was discovered that the heavier members of the group show a small degree of chemical reactivity. The term *noble* is used when we wish to suggest a very limited degree of chemical reactivity. Gold, for instance, is often referred to as a noble metal because so few chemicals are capable of reacting with it.

Finally, the elements of Group VIIA are called the **halogens,** derived from the Greek word meaning "sea" or "salt." Chlorine (Cl), for example, is found in familiar table salt, a compound that accounts in large measure for the salty taste of seawater.

PRACTICE EXERCISE 8: Circle the correct choices.

(a) Representative elements are: K, Cr, Pr, Ar, Al.
(b) A halogen is: Na, Fe, O, Cl, Cu.
(c) An alkaline earth metal is: Rb, Ba, La, As, Kr.
(d) A noble gas is: H, Ne, F, S, N.
(e) An alkali metal is: Zn, Ag, Br, Ca, Li.
(f) An inner transition element is: Ce, Pb, Ru, Xe, Mg.

Elements can be classified as metals, nonmetals, or metalloids

The periodic table organizes all sorts of chemical and physical information about the elements and their compounds. It allows us to study systematically the way properties vary with an element's position within the table and, in turn, makes the similarities and differences among the elements easier to understand and remember.

TOOLS

Periodic table and trends in the metallic character of the elements.

Even a casual inspection of samples of the elements reveals that some are familiar metals and that others, equally familiar, are not metals. Most of us are already familiar with metals such as lead, iron, or gold and nonmetals such as oxygen or nitrogen. A closer look at the nonmetallic elements, though, reveals that some of them, silicon and arsenic to name two, have properties that lie between

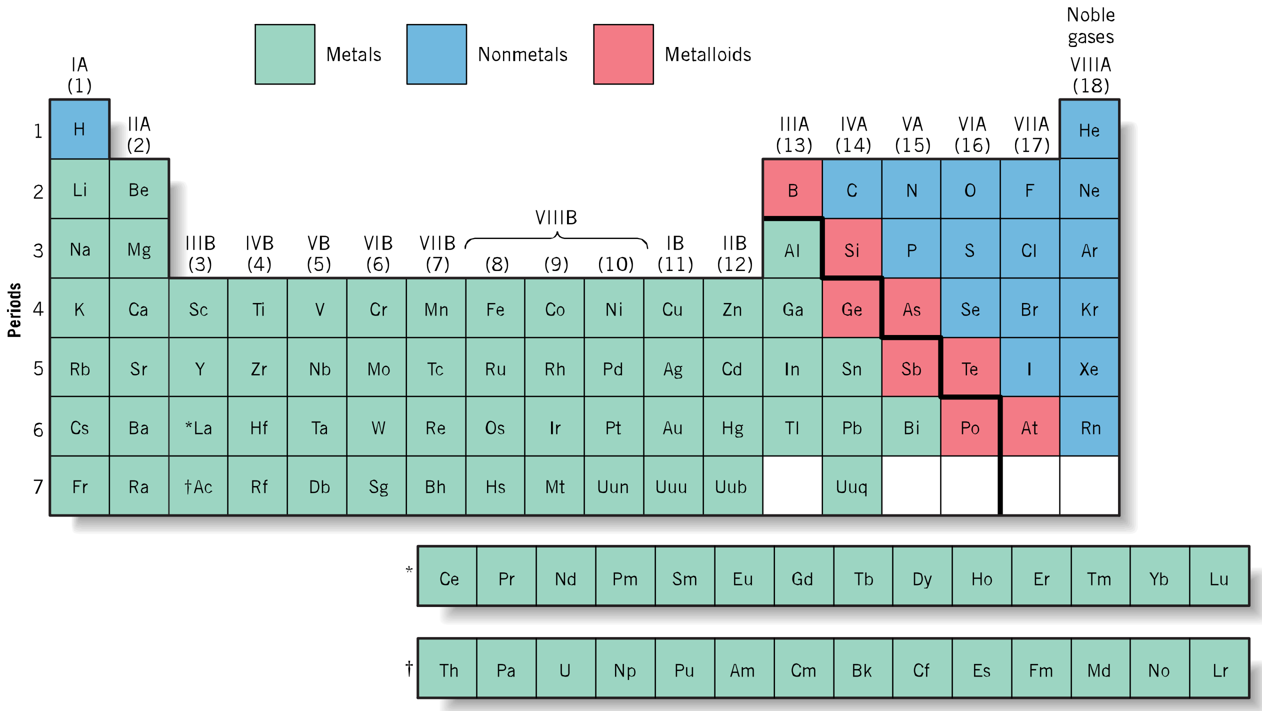


FIGURE 1.17 ***Distribution of metals, nonmetals, and metalloids among the elements in the periodic table.***

Notice that the metalloids are grouped around the bold stair-step line that is drawn diagonally from boron (B) to astatine (At).

those of true metals and true nonmetals. These elements are called **metalloids.** Division of the elements into the categories of metals, nonmetals, and metalloids is not even, however (see Figure 1.17). Most elements are metals, slightly over a dozen are nonmetals, and only a handful are metalloids.

Metals

You probably know a metal when you see one. Metals tend to have a shine so unique that it's called a *metallic luster.* For example, the silvery sheen of the freshly exposed surface of sodium in Figure 1.18 would most likely lead you to identify sodium as a metal even if you had never seen or heard of it before. We also know that metals conduct electricity. Few of us would hold an iron nail in our hand and poke it into an electrical outlet. In addition, we know that metals conduct heat very well. On a cool day, metals always feel colder to the touch than do neighboring nonmetallic objects because metals conduct heat away from your hand very rapidly. Nonmetals seem less cold because they can't conduct heat away as quickly and therefore their surfaces warm up faster.

Thin lead sheets are used for sound deadening because the easily deformed lead absorbs the sound vibrations.

Other properties that metals possess, to varying degrees, are **malleability**—the ability to be hammered or rolled into thin sheets—and **ductility**—the ability to be drawn into wire. The ability of a blacksmith to make horseshoes from a bar of iron (Figure 1.19) depends on the malleability of iron and steel, and the manufacture of electrical wire (Figure 1.20) is based on the ductility of copper.

The statue of Prometheus that stands above the skating rink in Rockefeller Center in New York City is covered by a thin decorative layer of gold leaf.

Hardness is another physical property that we usually think of for metals. Some, such as chromium or iron, are indeed quite hard, but others, like copper and lead, are rather soft. The alkali metals such as sodium (Figure 1.18) are so soft they can be cut with a knife, but they are also so chemically reactive that we rarely get to see them as free elements.

All the metallic elements, except mercury, are solids at room temperature (Figure 1.21). Mercury's low freezing point (-39 °C) and fairly high boiling