CHAPTER

4

Atoms and Elements

"Nothing exists except atoms and empty space; everything else is opinion."

DEMOCRITUS (460–370 B.C.)

4.1 Experiencing Atoms at Tiburon

My wife and I recently enjoyed a visit to the northern California seaside town of Tiburon. Tiburon sits next to San Francisco Bay with views of the water, the city of San Francisco, and the surrounding mountains. As we walked along a waterside path, I could feel the wind as it blew over the bay. I could hear the water splashing on the shore, and I could smell the sea air. What was the cause of these sensations? The answer is simple—atoms.

Since all matter is made of atoms, atoms are at the foundation of our sensations. The atom is the fundamental building block of everything you hear, feel, see, and experience. When you feel wind on your skin, you are feeling atoms. When you hear sounds, you are in a sense hearing atoms. When you touch a shoreside rock, you are touching atoms, and when you smell sea air, you are smelling atoms. You eat atoms, you breathe atoms, and you excrete atoms. Atoms are the building blocks of matter; they are the basic units from which nature builds. They are all around us and compose everything, including our own bodies.

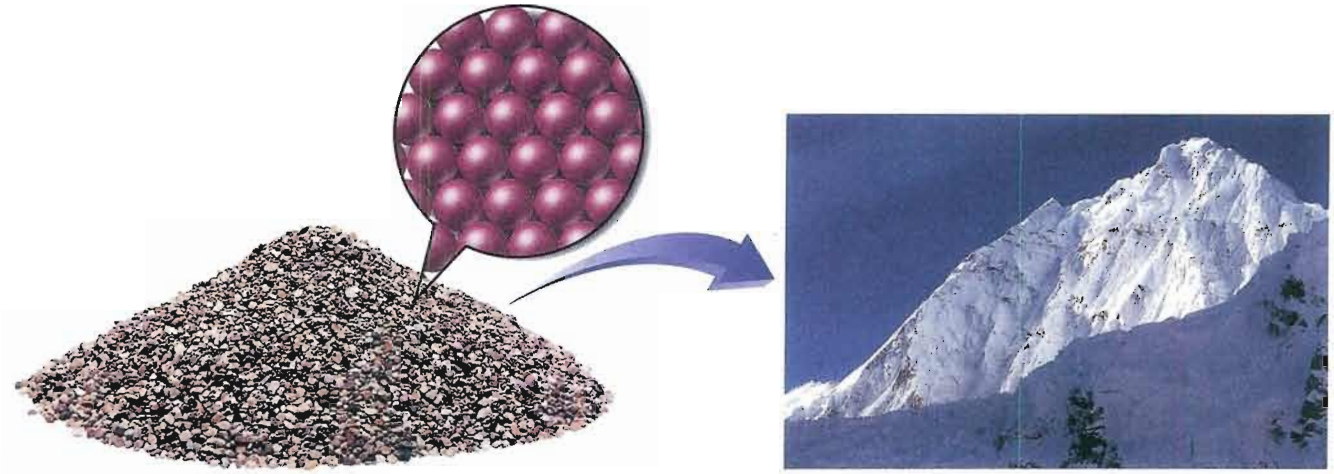
As we learned in Chapter 3, many atoms exist not as free particles but as groups of atoms bound together to form molecules. Nevertheless, all matter is ultimately made of atoms.

Atoms are incredibly small. A single pebble from the shoreline contains more atoms than you could ever count. The number of atoms in a single pebble far exceeds the number of pebbles on the bottom of San Francisco Bay. To get an idea of how small atoms are, imagine that every atom within a small pebble were the size of the pebble itself; then the pebble would be larger than Mount Everest (▶ Figure 4.1). Atoms are small—yet they compose everything.

The key to connecting the microscopic world with the macroscopic world is the atom. Atoms compose matter; their properties determine matter's properties. An *atom* is the smallest identifiable unit of an element. There are about ninety-one different elements in nature, and consequently about ninety-one different kinds of atoms. In addition, scientists have succeeded in making about twenty synthetic elements (not found in nature). In this chapter, we learn about atoms: what they are made of, how they differ from one another, and how they are structured. We also learn about the elements that atoms compose and some of the properties of those elements.

The exact number of naturally occurring elements is controversial because some elements previously considered only synthetic may actually occur in nature in very small quantities.

◀ All matter is composed of atoms. Seaside rocks are often composed of silicates, compounds of silicon and oxygen atoms. Seaside air, like all air, contains nitrogen and oxygen molecules, but it may also contain substances called amines. The amine shown here is triethylamine, which is emitted by decaying fish. Triethylamine is one of the compounds responsible for the fishy smell of the seaside.



▲ **Figure 4.1 The size of the atom** If every atom within a pebble were the size of the pebble itself, then the pebble would be larger than Mount Everest.

4.2 Indivisible: The Atomic Theory

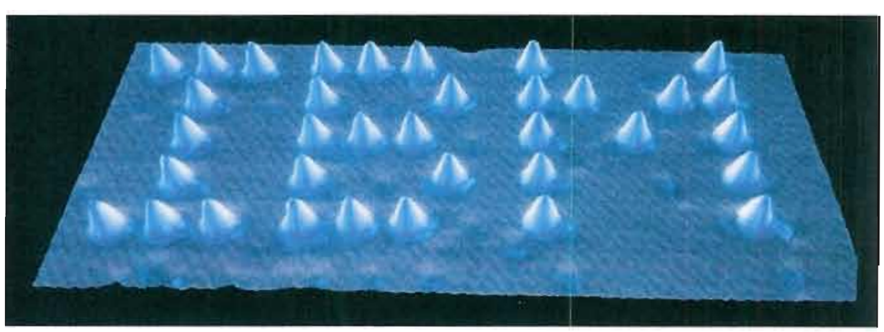
▲ Diogenes and Democritus, as imagined by a medieval artist. Democritus is the first person on record to have postulated that matter was composed of atoms.

If we simply examine matter, even under a microscope, it is not obvious that matter is composed of tiny particles. In fact, it appears to be just the opposite. If we divide a sample of matter into smaller and smaller pieces, it seems that we could divide it forever. From our perspective, matter seems continuous. The first people recorded as thinking otherwise were Leucippus (fifth century B.C., exact dates unknown) and Democritus (460–370 B.C.). These Greek philosophers theorized that matter was ultimately composed of small, indivisible particles called *atomos*, or "atoms," meaning "indivisible." Democritus suggested that if you divided matter into smaller and smaller pieces, you would eventually end up with tiny, indestructible particles—atoms.

The ideas of Leucippus and Democritus were not widely accepted, and it was not until 1808—over 2000 years later—that John Dalton formalized a theory of atoms that gained broad acceptance. Dalton's atomic theory has three parts:

1. Each element is composed of tiny indestructible particles called atoms.
2. All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements.
3. Atoms combine in simple, whole-number ratios to form compounds.

Today, the evidence for the atomic theory is overwhelming. Recent advances in microscopy have allowed scientists not only to image individual atoms but also to pick them up and move them (▼ Figure 4.2). Matter is indeed composed of atoms.



▶ **Figure 4.2 Writing with atoms** Scientists at IBM used a special microscope, called a scanning tunneling microscope (STM), to move xenon atoms to form the letters I, B, and M. The cone shape of these atoms is due to the peculiarities of the instrumentation. Atoms are, in general, spherical in shape.

Everyday Chemistry

Atoms and Humans

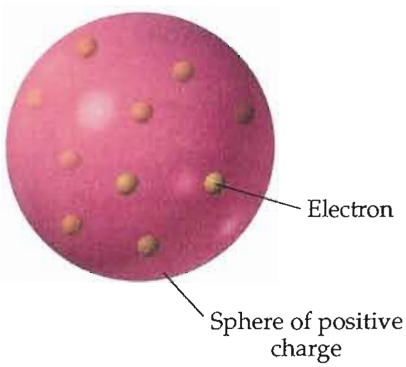
All matter is composed of atoms. What does that mean? What does it imply? It means that everything before you is composed of tiny particles too small to see. It means that even you and I are composed of these same particles. We acquired those particles from the food we have eaten over the years. The average carbon atom in our own bodies has been used by twenty other living organisms before we get to it and will be used by other organisms after we die.

The idea that all matter is composed of atoms has far-reaching implications. It implies that our bodies, our hearts, and even our brains are composed of atoms acting according to the laws of chemistry and physics. Some have viewed this as a devaluation of human life. We have always wanted to distinguish ourselves from everything else, and the idea that we are made of the same basic particles as all other matter takes something away from that distinction ... or does it?

CAN YOU ANSWER THIS? *Do you find the idea that you are made of atoms disturbing? Why or why not?*

4.3 The Nuclear Atom

Electric charge is more fully defined in Section 4.4. For now, think of it as an inherent property of electrons that causes them to interact with other charged particles.

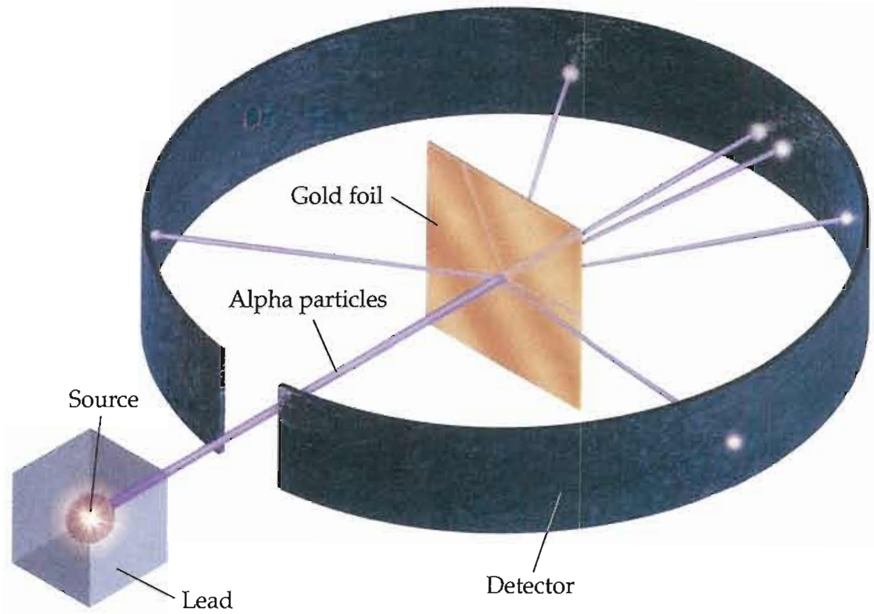


▲ **Figure 4.3 Plum pudding model of the atom** In the model suggested by J. J. Thomson, negatively charged electrons (yellow) were held in a sphere of positive charge (red).

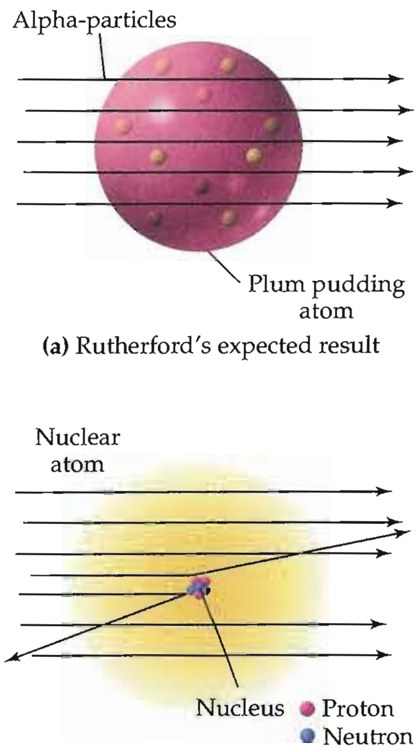
By the end of the nineteenth century, scientists were convinced that matter was composed of atoms, the permanent, indestructible building blocks from which all substances are constructed. However, an English physicist named J. J. Thomson (1856–1940) complicated the picture by discovering an even smaller and more fundamental particle called the **electron**. Thomson discovered that electrons are negatively charged, that they are much smaller and lighter than atoms, and that they are uniformly present in many different kinds of substances. The indestructible building block called the atom could apparently be "chipped."

The discovery of negatively charged particles within atoms raised the question of a balancing positive charge. Atoms were known to be charge-neutral, so they must contain positive charge that balanced the negative charge of electrons. But how did the positive and negative charges within the atom fit together? Were atoms just a jumble of even more fundamental particles? Were they solid spheres, or did they have some internal structure? Thomson proposed that the negatively charged electrons were small particles held within a positively charged sphere. This model, the most popular of the time, became known as the plum pudding model (plum pudding is an English dessert) (◀ Figure 4.3). The picture suggested by Thomson was—to those of us not familiar with plum pudding—like a blueberry muffin, where the blueberries are the electrons and the muffin is the positively charged sphere.

In 1909, Ernest Rutherford (1871–1937), who had worked under Thomson and adhered to his plum pudding model, performed an experiment in an attempt to confirm it. His experiment proved it wrong instead. In his experiment, Rutherford directed tiny, positively charged particles—called alpha-particles—at an ultrathin sheet of gold foil (▶ Figure 4.4). These particles were to act as probes of the gold atoms' structure. If the gold atoms were indeed like blueberry muffins or plum pudding—with their mass and charge spread throughout the entire volume of the atom—these speeding probes should pass right through the gold foil with minimum deflection. Rutherford performed the experiment, but the results were not as he expected. A majority of the particles did pass directly through the foil, but some particles were deflected, and some (1 in 20,000) even bounced back. The results puzzled Rutherford, who found them "about as credible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you." What must the structure of the atom be in order to explain this odd behavior?



▲ **Figure 4.4 Rutherford's gold foil experiment** Tiny particles called alpha-particles were directed at a thin sheet of gold foil. Most of the particles passed directly through the foil. A few, however, were deflected—some of them at sharp angles.



(a) Rutherford's expected result

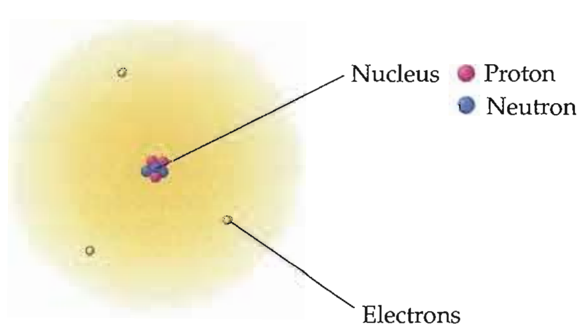
(b) Rutherford's actual result

▲ **Figure 4.5 Discovery of the atomic nucleus** **(a)** Expected result of Rutherford's gold foil experiment. If the plum pudding model were correct, the alpha-particles would pass right through the gold foil with minimal deflection. **(b)** Actual result of Rutherford's gold foil experiment. A small number of alpha-particles were deflected or bounced back. The only way to explain the deflections was to suggest that most of the mass and all of the positive charge of an atom must be concentrated in a space much smaller than the size of the atom itself—the nucleus. The nucleus itself is composed of positively charged particles (protons) and neutral particles (neutrons).

Rutherford created a new model to explain his results (◀ Figure 4.5). He concluded that matter must not be as uniform as it appears. It must contain large regions of empty space dotted with small regions of very dense matter. In order to explain the deflections he observed, the mass and positive charge of an atom must all be concentrated in a space much smaller than the size of the atom itself. Using this idea, he proposed the **nuclear theory of the atom**, which has three basic parts:

1. Most of the atom's mass and all of its positive charge are contained in a small core called the *nucleus*.
2. Most of the volume of the atom is empty space through which the tiny, negatively charged electrons are dispersed.
3. There are as many negatively charged electrons outside the nucleus as there are positively charged particles (*protons*) inside the nucleus, so that the atom is electrically neutral.

Later work by Rutherford and others demonstrated that the atom's **nucleus** contains both positively charged **protons** and neutral particles called **neutrons**. The dense nucleus makes up more than 99.9% of the mass of the atom, but occupies only a small fraction of its volume. The electrons are distributed through a much larger region, but don't have much mass (▼ Figure 4.6). For now, you can



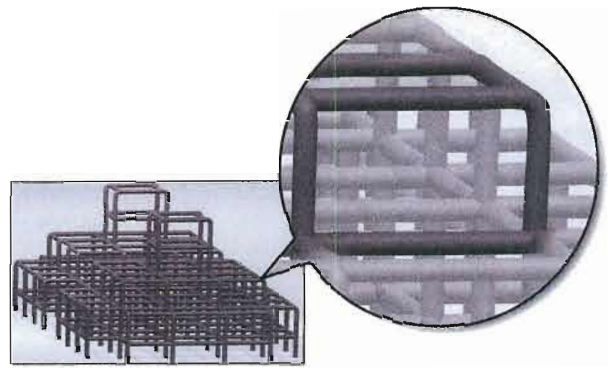
▶ **Figure 4.6 The nuclear atom** In this model, 99.9% of the atom's mass is concentrated in a small, dense nucleus that contains protons and neutrons. The rest of the volume of the atom is mostly empty space occupied by negatively charged electrons. The number of electrons outside the nucleus is equal to the number of protons inside the nucleus. In this image, the nucleus is greatly enlarged and the electrons are portrayed as particles.

Everyday Chemistry

Solid Matter?

If matter really is mostly empty space as Rutherford suggested, then why does it appear so solid? Why can I tap my knuckles on the table and feel a solid thump? Matter appears solid because the variation in the density is on such a small scale that our eyes can't see it. Imagine a jungle gym one hundred stories high and the size of a football field. It is mostly empty space. Yet if you viewed it from an airplane, it would appear as a solid mass. Matter is similar. When you tap your knuckles on the table, it is much like one giant jungle gym (your finger) crashing into another (the table). Even though they are both primarily empty space, one does not fall into the other.

CAN YOU ANSWER THIS? *Use the jungle gym analogy to explain why most of Rutherford's alpha-particles went right through the gold foil and why a few bounced back. Remember that his gold foil was extremely thin.*



▲ Matter appears solid and uniform because the variation in density is on a scale too small for our eyes to see. Just as this scaffolding appears solid at a distance, so matter appears solid to us.

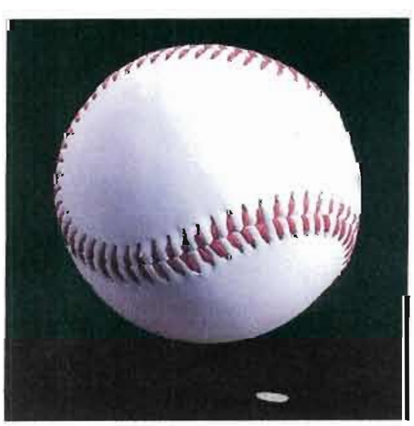
think of these electrons like the water droplets that make up a cloud—they are dispersed throughout a large volume but weigh almost nothing.

Rutherford's nuclear theory was a success and is still valid today. The revolutionary part of this theory is the idea that matter—at its core—is much less uniform than it appears. If the nucleus of the atom were the size of this dot •, the average electron would be about 10 m away. Yet the dot would contain almost the entire mass of the atom. Imagine what matter would be like if atomic structure broke down. What if matter were composed of atomic nuclei piled on top of each other like marbles? Such matter would be incredibly dense; a single grain of sand composed of solid atomic nuclei would have a mass of 5 million kg (or a weight of about 10 million lb). Astronomers believe there are some places in the universe where such matter exists: neutron stars and black holes.

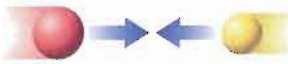
4.4 The Properties of Protons, Neutrons, and Electrons

Protons and neutrons have very similar masses. In SI units, the mass of the proton is 1.67262×10^{-27} kg, and the mass of the neutron is a close 1.67493×10^{-27} kg. A more common unit to express these masses, however, is the **atomic mass unit (amu)**, defined as one-twelfth of the mass of a carbon atom containing six protons and six neutrons. In this unit, a proton has a mass of 1.0073 amu and a neutron has a mass of 1.0087 amu. Electrons, by contrast, have an almost negligible mass of 0.00091×10^{-27} kg, or approximately 0.00055 amu.

The proton and the electron both have electrical **charge**. The proton's charge is 1+ and the electron's charge is 1−. The charge of the proton and the electron are equal in magnitude but opposite in sign, so that when the two particles are paired, the charges exactly cancel. The neutron has no charge.

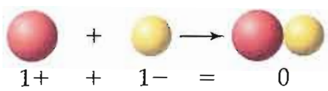


◄ If a proton had the mass of a baseball, an electron would have the mass of a rice grain. The proton is nearly 2000 times as massive as an electron.



Positive (red) and negative (yellow) charges attract.

Positive–positive and negative–negative charges repel.



Positive and negative charges cancel.

▲ **Figure 4.7 The properties of electrical charge**

What is electrical charge? Electrical charge is a fundamental property of protons and electrons, just as mass is a fundamental property of matter. Most matter is charge-neutral because protons and electrons occur together and their charges cancel. However, you have probably experienced excess electrical charge when brushing your hair on a dry day. The brushing action results in the accumulation of electrical charge on the hair strands, which then repel each other, causing your hair to stand on end.

We can summarize the nature of electrical charge as follows (◀ Figure 4.7)

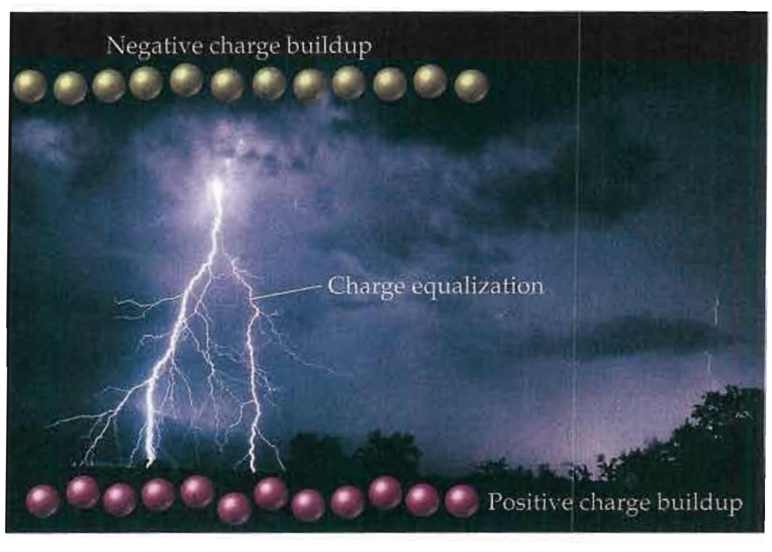
- Electrical charge is a fundamental property of protons and electrons.
- Positive and negative electrical charges attract each other.
- Positive–positive and negative–negative charges repel each other.
- Positive and negative charges cancel each other so that a proton and an electron, when paired, are charge-neutral.

Notice that matter is usually charge-neutral due to the canceling effect of protons and electrons. When matter does acquire charge imbalances, these imbalances usually equalize quickly, often in dramatic ways. For example, the shock you receive when touching a doorknob during dry weather is the equalization of a charge imbalance that developed as you walked across the carpet. Lightning is an equalization of charge imbalances that develop during electrical storms.

If you had a sample of matter—even a tiny sample, such as a sand grain—that was composed of only protons or only electrons, the forces around that matter would be extraordinary, and the matter would be unstable. Fortunately, matter is not that way—protons and electrons exist together, canceling each other's charge and making matter charge-neutral. The properties of protons, neutrons, and electrons are summarized in Table 4.1.

TABLE 4.1 Subatomic Particles

	Mass (kg)	Mass (amu)	Charge
proton	1.67262×10^{-27}	1.0073	1+
neutron	1.67493×10^{-27}	1.0087	0
electron	0.00091×10^{-27}	0.00055	1−



▲ Matter is normally charge-neutral, having equal numbers of positive and negative charges that exactly cancel. When the charge balance of matter is disturbed, as in an electrical storm, it quickly rebalances, often in dramatic ways such as lightning.

CONCEPTUAL CHECKPOINT 4.1

An atom with which of these compositions would have a mass of approximately 12 amu and be charge-neutral?

(a) 6 protons and 6 electrons
(b) 3 protons, 3 neutrons, and 6 electrons
(c) 6 protons, 6 neutrons, and 6 electrons
(d) 12 neutrons and 12 electrons

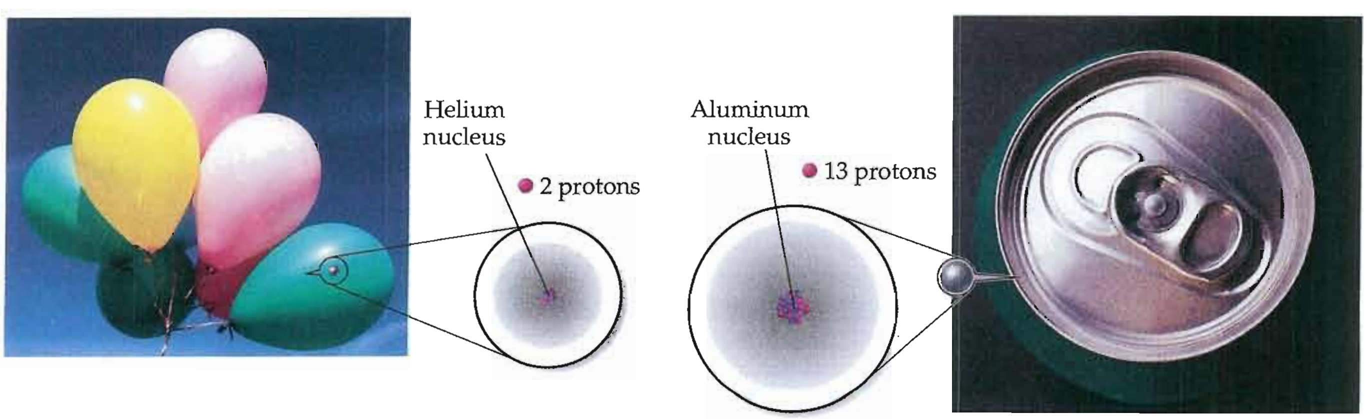
4.5 Elements: Defined by Their Numbers of Protons

We have seen that atoms are composed of protons, neutrons, and electrons. However, it is the number of protons in the nucleus of an atom that identifies it as a particular element. For example, atoms with 2 protons in their nucleus are helium atoms, atoms with 13 protons in their nucleus are aluminum atoms, and atoms with 92 protons in their nucleus are uranium atoms. The number of protons in an atom's nucleus defines the element (▼ Figure 4.8). Every aluminum atom has 13 protons in its nucleus; if it had a different number of protons, it would be a different element. The number of protons in the nucleus of an atom is called the **atomic number** and is given the symbol Z.

The periodic table of the elements (▶ Figure 4.9) lists all known elements according to their atomic numbers. Each element is represented by a unique **chemical symbol**, a one- or two-letter abbreviation for the element that appears directly below its atomic number on the periodic table. The chemical symbol for helium is He; for aluminum, Al; and for uranium, U. The chemical symbol and the atomic number always go together. If the atomic number is 13, the chemical symbol must be Al. If the atomic number is 92, the chemical symbol must be U. This is just another way of saying that the number of protons defines the element.

Most chemical symbols are based on the English name of the element. For example, the symbol for carbon is C; for silicon, Si; and for bromine, Br. Some elements, however, have symbols based on their Latin names. For example, the symbol for potassium is K, from the Latin *kalium*, and the symbol for sodium is Na, from the Latin *natrium*. Other elements with symbols based on their Greek or Latin names include the following:

lead	Pb	*plumbum*
mercury	Hg	*hydrargyrum*
iron	Fe	*ferrum*
silver	Ag	*argentum*
tin	Sn	*stannum*
copper	Cu	*cuprum*



▲ **Figure 4.8 The number of protons in the nucleus defines the element**

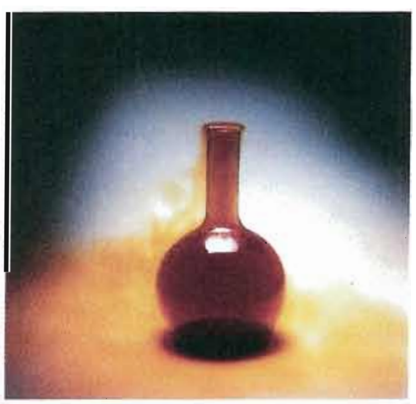
4 — Atomic number (Z)
Be — Chemical symbol
9.01 — Atomic mass (defined in section 4.9)
beryllium — Name

Periods	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H 1.01 hydrogen																	2 He 4.00 helium
2	3 Li 6.94 lithium	4 Be 9.01 beryllium											5 B 10.81 boron	6 C 12.01 carbon	7 N 14.01 nitrogen	8 O 16.00 oxygen	9 F 19.00 fluorine	10 Ne 20.18 neon
3	11 Na 22.99 sodium	12 Mg 24.31 magnesium											13 Al 26.98 aluminum	14 Si 28.09 silicon	15 P 30.97 phosphorus	16 S 32.07 sulfur	17 Cl 35.45 chlorine	18 Ar 39.95 argon
4	19 K 39.10 potassium	20 Ca 40.08 calcium	21 Sc 44.96 scandium	22 Ti 47.88 titanium	23 V 50.94 vanadium	24 Cr 52.00 chromium	25 Mn 54.94 manganese	26 Fe 55.85 iron	27 Co 58.93 cobalt	28 Ni 58.69 nickel	29 Cu 63.55 copper	30 Zn 65.39 zinc	31 Ga 69.72 gallium	32 Ge 72.61 germanium	33 As 74.92 arsenic	34 Se 78.96 selenium	35 Br 79.90 bromine	36 Kr 83.80 krypton
5	37 Rb 85.47 rubidium	38 Sr 87.62 strontium	39 Y 88.91 yttrium	40 Zr 91.22 zirconium	41 Nb 92.91 niobium	42 Mo 95.94 molybdenum	43 Tc (99) technetium	44 Ru 101.07 ruthenium	45 Rh 102.91 rhodium	46 Pd 106.42 palladium	47 Ag 107.87 silver	48 Cd 112.41 cadmium	49 In 114.82 indium	50 Sn 118.71 tin	51 Sb 121.75 antimony	52 Te 127.60 tellurium	53 I 126.90 iodine	54 Xe 131.29 xenon
6	55 Cs 132.91 cesium	56 Ba 137.33 barium	57 La 138.91 lanthanum	72 Hf 178.49 hafnium	73 Ta 180.95 tantalum	74 W 183.85 tungsten	75 Re 186.21 rhenium	76 Os 190.2 osmium	77 Ir 192.22 iridium	78 Pt 195.08 platinum	79 Au 196.97 gold	80 Hg 200.59 mercury	81 Tl 204.38 thallium	82 Pb 207.2 lead	83 Bi 208.98 bismuth	84 Po (209) polonium	85 At (210) astatine	86 Rn (222) radon
7	87 Fr (223) francium	88 Ra (226) radium	89 Ac (227) actinium	104 Rf (261) rutherfordium	105 Db (262) dubnium	106 Sg (263) seaborgium	107 Bh (262) bohrium	108 Hs (265) hassium	109 Mt (266) meitnerium	110 Ds (281) darmstadtium	111 Rg (280) roentgenium	112 — (285)	113 — (284)	114 — (289)	115 — (288)	116 — (292)		118 — (294)

Lanthanides	58 Ce 140.12 cerium	59 Pr 140.91 praseodymium	60 Nd 144.24 neodymium	61 Pm (147) promethium	62 Sm 150.36 samarium	63 Eu 151.97 europium	64 Gd 157.25 gadolinium	65 Tb 158.93 terbium	66 Dy 162.50 dysprosium	67 Ho 164.93 holmium	68 Er 167.26 erbium	69 Tm 168.93 thulium	70 Yb 173.04 ytterbium	71 Lu 174.97 lutetium
Actinides	90 Th (232) thorium	91 Pa (231) protactinium	92 U (238) uranium	93 Np (237) neptunium	94 Pu (244) plutonium	95 Am (243) americium	96 Cm (247) curium	97 Bk (247) berkelium	98 Cf (251) californium	99 Es (252) einsteinium	100 Fm (257) fermium	101 Md (258) mendelevium	102 No (259) nobelium	103 Lr (260) lawrencium

▲ **Figure 4.9 The periodic table of the elements**

The names of elements were often given to describe their properties. For example, *argon* originates from the Greek word *argos*, meaning "inactive," referring to argon's chemical inertness (it does not react with other elements). *Bromine* originates from the Greek word *bromos*, meaning "stench," referring to bromine's strong odor. Other elements were named after countries. For example, polonium was named after Poland, francium after France, and americium after the United States of America. Still other elements were named after scientists. Curium was named after Marie Curie, and mendelevium after Dmitri Mendeleev. Every element's name, symbol, and atomic number are listed in the periodic table (inside front cover) and in an alphabetical listing (inside back cover) in this book.



◀ The name *bromine* originates from the Greek word *bromos*, meaning "stench." Bromine vapor, seen as the red-brown gas in this photograph, has a strong odor.

▶ Curium is named after Marie Curie, a chemist who helped discover radioactivity and also discovered two new elements. Curie won two Nobel Prizes for her work.

Curium
96
Cm
(247)

EXAMPLE 4.1 Atomic Number, Atomic Symbol, and Element Name

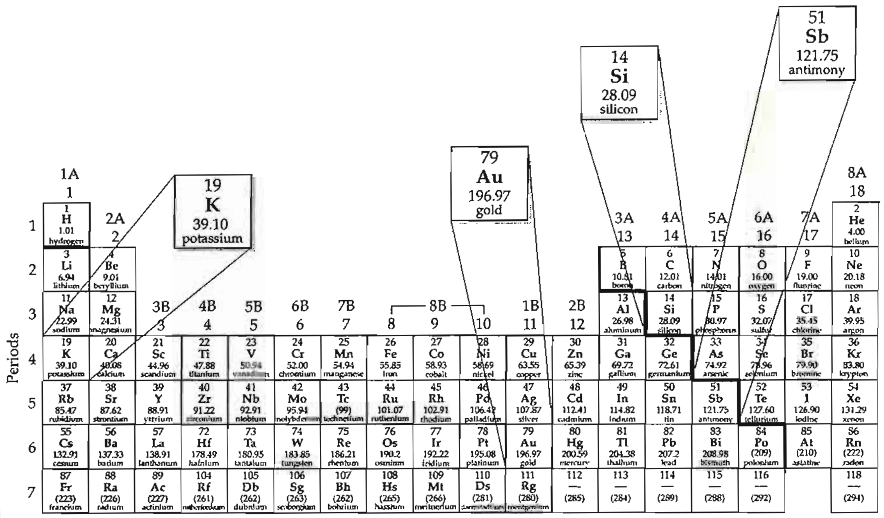
Find the atomic symbol and atomic number for each of the following elements.

(a) silicon
(b) potassium
(c) gold
(d) antimony

Solution:

As you become familiar with the periodic table, you will be able to quickly locate elements on it. For now, it might be easier to find them in the alphabetical listing on the inside back cover of this book, but you should also find their position in the periodic table.

Element	Symbol	Atomic Number
silicon	Si	14
potassium	K	19
gold	Au	79
antimony	Sb	51



SKILLBUILDER 4.1 Atomic Number, Atomic Symbol, and Element Name

Find the name and atomic number for each of the following elements.

(a) Na
(b) Ni
(c) P
(d) Ta

FOR MORE PRACTICE Problems 43, 44, 47, 48, 49, 50, 51, 52.

4.6 Looking for Patterns: The Periodic Law and the Periodic Table

The organization of the periodic table has its origins in the work of Dmitri Mendeleev (1834–1907), a nineteenth-century Russian chemistry professor. In his time, about sixty-five different elements had been discovered. Through the work of a number of chemists, much was known about each of these elements, including their relative masses, chemical activity, and some of their physical properties. However, there was no systematic way of organizing them.

▲ Dmitri Mendeleev, a Russian chemistry professor who proposed the periodic law and arranged early versions of the periodic table, shown on a Russian postage stamp.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
H	He	Li	Be	B	C	N	O	F	Ne	Na	Mg	Al	Si	P	S	Cl	Ar	K	Ca

▲ **Figure 4.10 Recurring properties** The elements shown are listed in order of increasing atomic number (Mendeleev used relative mass, which is similar). The color of each element represents its properties. Notice that the properties (colors) of these elements form a repeating pattern.

In 1869, Mendeleev noticed that certain groups of elements had similar properties. Mendeleev found that if he listed the elements in order of increasing relative mass, those similar properties recurred in a regular pattern (▲ Figure 4.10). Mendeleev summarized these observations in the **periodic law**:

> When the elements are arranged in order of increasing relative mass, certain sets of properties recur periodically.

Periodic means "recurring regularly." The properties of the elements, when listed in order of increasing relative mass, formed a *repeating pattern.*

1 H							2 He
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca						

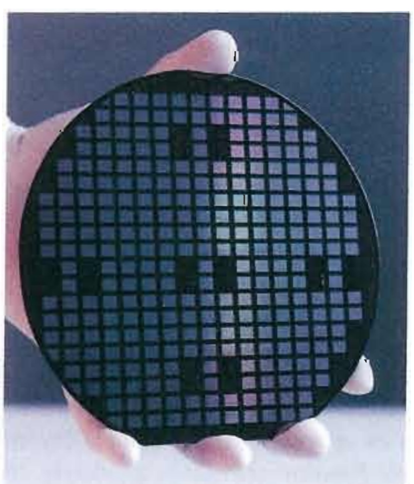
▲ **Figure 4.11 Making a periodic table** If we place the elements from Figure 4.10 in a table, we can arrange them in rows so that similar properties align in the same vertical columns. This is similar to Mendeleev's first periodic table.

Mendeleev then organized all the known elements in a table in which relative mass increased from left to right and elements with similar properties were aligned in the same vertical columns (◀ Figure 4.11). Since many elements had not yet been discovered, Mendeleev's table contained some gaps, which allowed him to predict the existence of yet-undiscovered elements. For example, Mendeleev predicted the existence of an element he called *eka-silicon,* which fell below silicon on the table and between gallium and arsenic. In 1886, eka-silicon was discovered by German chemist Clemens Winkler (1838–1904) and was found to have almost exactly the properties that Mendeleev had anticipated. Winkler named the element germanium, after his home country.

Mendeleev's original listing has evolved into the modern **periodic table.** In the modern table, elements are listed in order of increasing atomic number rather than increasing relative mass. The modern periodic table also contains more elements than Mendeleev's original table because many more have been discovered since his time.

Mendeleev's periodic law was based on observation. Like all scientific laws, the periodic law summarized many observations but did not give the underlying reason for the observation—only theories do that. For now, we accept the periodic law as it is, but in Chapter 9 we examine a powerful theory that explains the law and gives the underlying reasons for it.

The elements in the periodic table can be broadly classified as metals, nonmetals, and metalloids (▶ Figure 4.12). **Metals** occupy the left side of the periodic table and have similar properties: They are good conductors of heat and electricity; they can be pounded into flat sheets (malleability); they can be drawn into wires (ductility); they are often shiny; and they tend to lose electrons when they undergo chemical changes. Good examples of metals include iron, magnesium, chromium, and sodium.



▲ Silicon is a metalloid used extensively in the computer and electronics industries.

Nonmetals occupy the upper right side of the periodic table. The dividing line between metals and nonmetals is the zigzag diagonal line running from boron to astatine (see Figure 4.12). Nonmetals have more varied properties—some are solids at room temperature, others are gases—but as a whole they tend to be poor conductors of heat and electricity, and they all tend to gain electrons when they undergo chemical changes. Good examples of nonmetals include oxygen, nitrogen, chlorine, and iodine.

Most of the elements that lie along the zigzag diagonal line dividing metals and nonmetals are called **metalloids**, or semimetals, and show mixed properties. Metalloids are also called **semiconductors** because of their intermediate electrical conductivity, which can be changed and controlled. This property makes semiconductors useful in the manufacture of the electronic devices that are central to computers, cell phones, and many other modern gadgets. Good examples of metalloids include silicon, arsenic, and germanium.

Metals

Nonmetals

Metalloids

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H**																	2 **He**
2	3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
3	11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
4	19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
5	37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
6	55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
7	87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112	113	114	115	116		118

Lanthanides	58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
Actinides	90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

▲ **Figure 4.12** **Metals, nonmetals, and metalloids** The elements in the periodic table can be broadly classified as metals, nonmetals, and metalloids.

EXAMPLE 4.2 Classifying Elements as Metals, Nonmetals, or Metalloids

Classify each of the following elements as a metal, nonmetal, or metalloid.

(a) Ba
(b) I
(c) O
(d) Te

Solution:

(a) Barium is on the left side of the periodic table; it is a metal.
(b) Iodine is on the right side of the periodic table; it is a nonmetal.
(c) Oxygen is on the right side of the periodic table; it is a nonmetal.
(d) Tellurium is in the middle-right section of the periodic table, along the line that divides the metals from the nonmetals; it is a metalloid.

SKILLBUILDER 4.2 Classifying Elements as Metals, Nonmetals, or Metalloids

Classify each of the following elements as a metal, nonmetal, or metalloid.
(a) S
(b) Cl
(c) Ti
(d) Sb

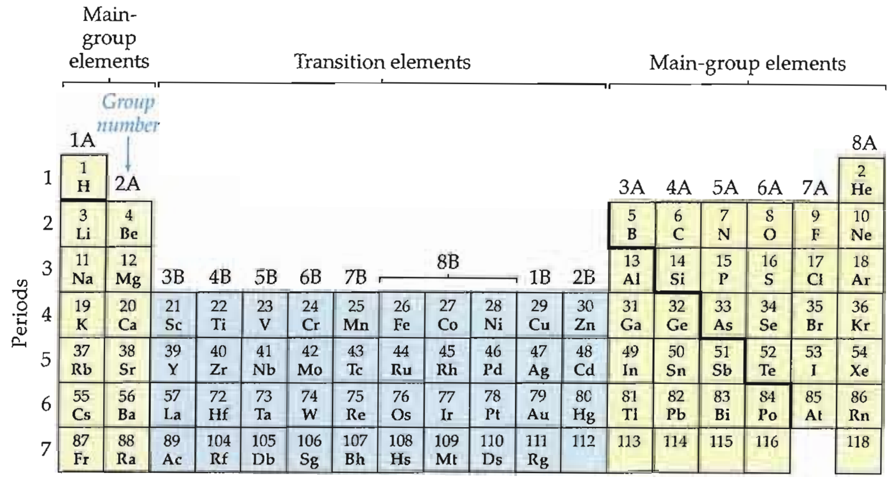
FOR MORE PRACTICE Problems: 53, 54, 55, 56.

Main-group elements are in columns labeled with a number and the letter A. Transition elements are in columns labeled with a number and the letter B.

A competing numbering system does not use letters, but only the numbers 1–18. Both numbering systems are shown in Figure 4.12.

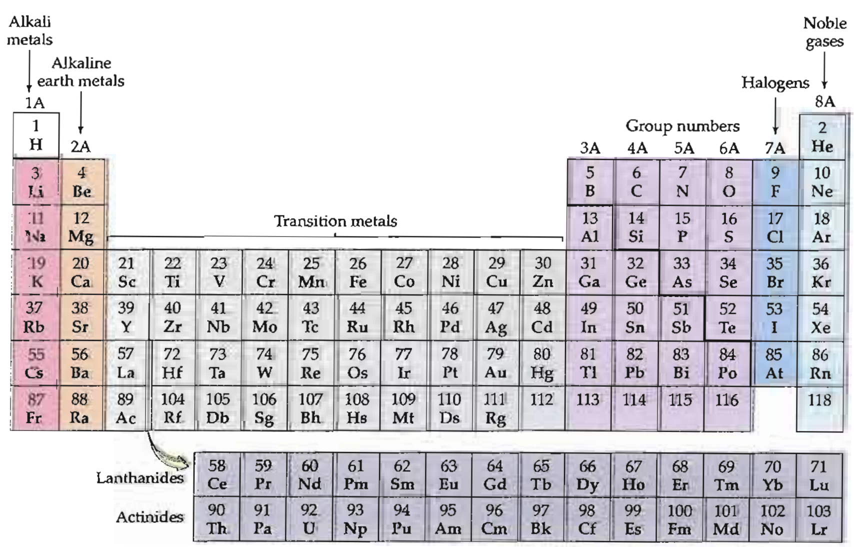
The periodic table can also be broadly divided into **main-group elements**, whose properties tend to be more predictable based on their position in the periodic table, and **transition elements** or **transition metals**, whose properties are less easily predictable based simply on their position in the periodic table (▶ Figure 4.13). Each column within the main-group elements in the periodic table is called a **family** or **group** of elements and is designated with a number and a letter printed directly above the column.

Figure 4.13 Main-group and transition elements The periodic table can be broadly divided into main-group elements, whose properties can generally be predicted based on their position, and transition elements, whose properties tend to be less predictable based on their position.



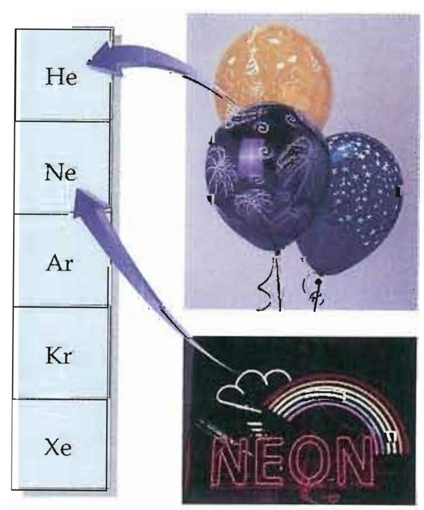
The noble gases are inert (or unreactive) compared to other elements. However, some noble gases, especially the heavier ones, will form a limited number of compounds with other elements under special conditions.

The elements within a group usually have similar properties. For example, the Group 8A elements, called the **noble gases**, are chemically inert gases. The most familiar noble gas is probably helium, used to fill balloons. Helium, like the other noble gases, is chemically stable—it won't combine with other elements to form compounds—and is therefore safe to put into balloons. Other noble gases include neon, often used in neon signs; argon, which makes up a small percentage of our atmosphere; krypton; and xenon. The Group 1A elements, called the **alkali metals**, are all very reactive metals. A marble-sized piece of sodium can explode when dropped into water. Other alkali metals include lithium, potassium, and rubidium. The Group 2A elements, called the **alkaline earth metals**, are also fairly reactive, although not quite as reactive as the alkali metals. Calcium, for example, reacts fairly vigorously when dropped into water but will not explode as easily as sodium. Other alkaline earth metals include magnesium, a common low-density structural metal; strontium; and barium. The Group 7A elements, called the **halogens**, are very reactive nonmetals. The most familiar halogen is probably chlorine, a greenish-yellow gas with a pungent odor. Because of its reactivity, chlorine is often used as a sterilizing and disinfecting agent. Other halogens include bromine, a red-brown liquid that easily evaporates into a gas; iodine, a purple solid; and fluorine, a pale yellow gas.



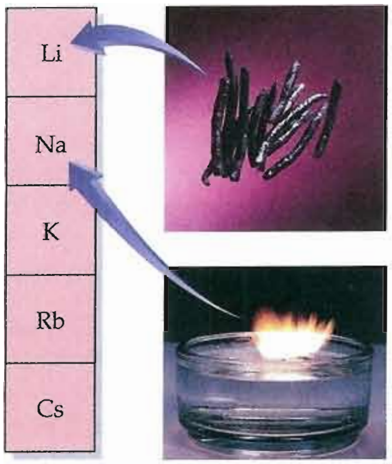
▲The periodic table with Groups 1A, 2A, 7A, and 8A highlighted.

Noble gases



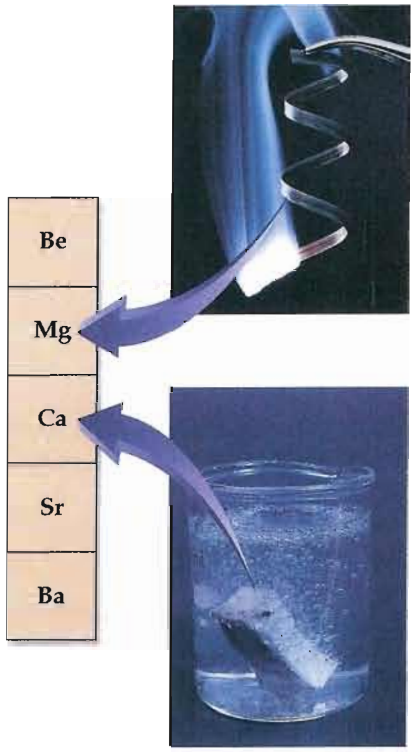
▲The noble gases include helium (used in balloons), neon (found in neon signs), argon, krypton, and xenon.

Alkali Metals



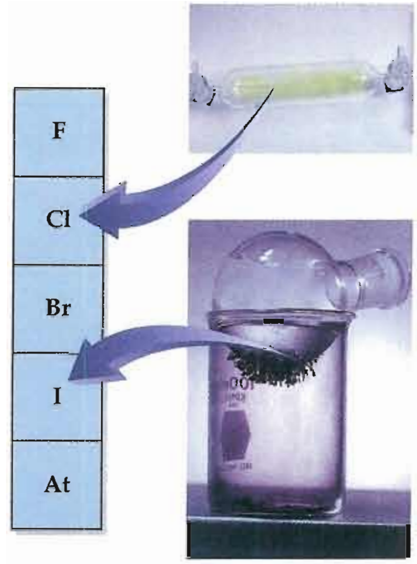
▲ The alkali metals include lithium (shown in the first photo), sodium (shown in the second photo reacting with water), potassium, rubidium, and cesium.

Alkaline Earth Metals



▲ The alkaline earth metals include beryllium, magnesium (shown burning in the first photo), calcium (shown reacting with water in the second photo), strontium, and barium.

Halogens



▲ The halogens include fluorine, chlorine (shown in the first photo), bromine, iodine (shown in the second photo), and astatine.

EXAMPLE 4.3 Groups and Families of Elements

To which group or family of elements does each of the following elements belong?

(a) Mg
(b) N
(c) K
(d) Br

Solution:

(a) Mg is in Group 2A; it is an alkaline earth metal.
(b) N is in Group 5A.
(c) K is in Group 1A; it is an alkali metal.
(d) Br is in Group 7A; it is a halogen.

SKILLBUILDER 4.3 Groups and Families of Elements

To which group or family of elements does each of the following elements belong?
(a) Li
(b) B
(c) I
(d) Ar

FOR MORE PRACTICE Problems 59, 60, 61, 62, 63, 64, 65, 66

CHECKPOINT 4.2

Which of these statements can NEVER be true?

(a) An element can be both a transition element and a metal.
(b) An element can be both a transition element and a metalloid.
(c) An element can be both a metalloid and a halogen.
(d) An element can be both a main-group element and a halogen.

4.7 Ions: Losing and Gaining Electrons

The charge of an ion is shown in the upper right corner of the symbol.

In chemical reactions, atoms often lose or gain electrons to form charged particles called **ions**. For example, neutral lithium (Li) atoms contain 3 protons and 3 electrons; however, in reactions, lithium atoms lose one electron (e^-) to form Li^+ ions.

$$Li \longrightarrow Li^+ + e^-$$

The Li^+ *ion* contains 3 protons but only 2 electrons, resulting in a net charge of 1+. Ion charges are usually written with the magnitude of the charge first followed by the sign of the charge. For example, a positive two charge is written as 2+ and a negative two charge is written as 2−. The charge of an ion depends on how many electrons were gained or lost and is given by the following formula:

$$\begin{aligned} \text{Ion charge} &= \text{number of protons} - \text{number of electrons} \\ &= \#p - \#e^- \end{aligned}$$

where p stands for *proton* and e^- stands for *electron.*

For the Li^+ ion:

$$\text{Ion charge} = 3 - 2 = 1+$$

Neutral fluorine (F) atoms contain 9 protons and 9 electrons; however, in chemical reactions fluorine atoms gain one electron to form F^- ions:

$$F + e^- \longrightarrow F^-$$

The F^- *ion* contains 9 protons and 10 electrons, resulting in a 1− charge.

$$\begin{aligned} \text{Ion charge} &= 9 - 10 \\ &= 1- \end{aligned}$$

Positively charged ions, such as Li^+, are called **cations**, and negatively charged ions, such as F^-, are called **anions**. Ions behave very differently than the atoms from which they are formed. Neutral sodium atoms, for example, are extremely reactive, reacting violently with most things they contact. Sodium cations (Na^+), on the other hand, are relatively inert—we eat them all the time in sodium chloride (table salt). In nature, cations and anions always occur together so that, again, matter is charge-neutral. For example, in table salt, the sodium cation occurs together with the chloride anion (Cl^-).

EXAMPLE 4.4 Determining Ion Charge from Numbers of Protons and Electrons

Determine the charge of each of the following ions.

(a) a magnesium ion with 10 electrons
(b) a sulfur ion with 18 electrons
(c) an iron ion with 23 electrons

Solution:

To determine the charge of each ion, we use the ion charge equation.

$$\text{Ion charge} = \#\text{p} - \#\text{e}^-$$

The number of electrons is given in the problem. The number of protons is obtained from the element's atomic number in the periodic table.

(a) magnesium with atomic number 12

$$\text{Ion charge} = 12 - 10 = 2+ \ (Mg^{2+})$$

(b) sulfur with atomic number 16

$$\text{Ion charge} = 16 - 18 = 2- \ (S^{2-})$$

(c) iron with atomic number 26

$$\text{Ion charge} = 26 - 23 = 3+ \ (Fe^{3+})$$

SKILLBUILDER 4.4 Determining Ion Charge from Numbers of Protons and Electrons

Determine the charge of each of the following ions.

(a) a nickel ion with 26 electrons
(b) a bromine ion with 36 electrons
(c) a phosphorus ion with 18 electrons

FOR MORE PRACTICE Example 4.10; Problems 75, 76.

EXAMPLE 4.5 Determining the Number of Protons and Electrons in an Ion

Find the number of protons and electrons in the Ca^{2+} ion.

From the periodic table, find that the atomic number for calcium is 20, so calcium has 20 protons. The number of electrons can be found using the ion charge equation.

Solution:

$\text{Ion charge} = \#\text{p} - \#\text{e}^-$
$2+ = 20 - \#\text{e}^-$
$\#\text{e}^- = 20 - 2 = 18$

Therefore the number of electrons is 18. The Ca^{2+} ion has 20 protons and 18 electrons.

SKILLBUILDER 4.5 Determining the Number of Protons and Electrons in an Ion

Find the number of protons and electrons in the S^{2-} ion.

FOR MORE PRACTICE Example 4.11; Problems 77, 78.

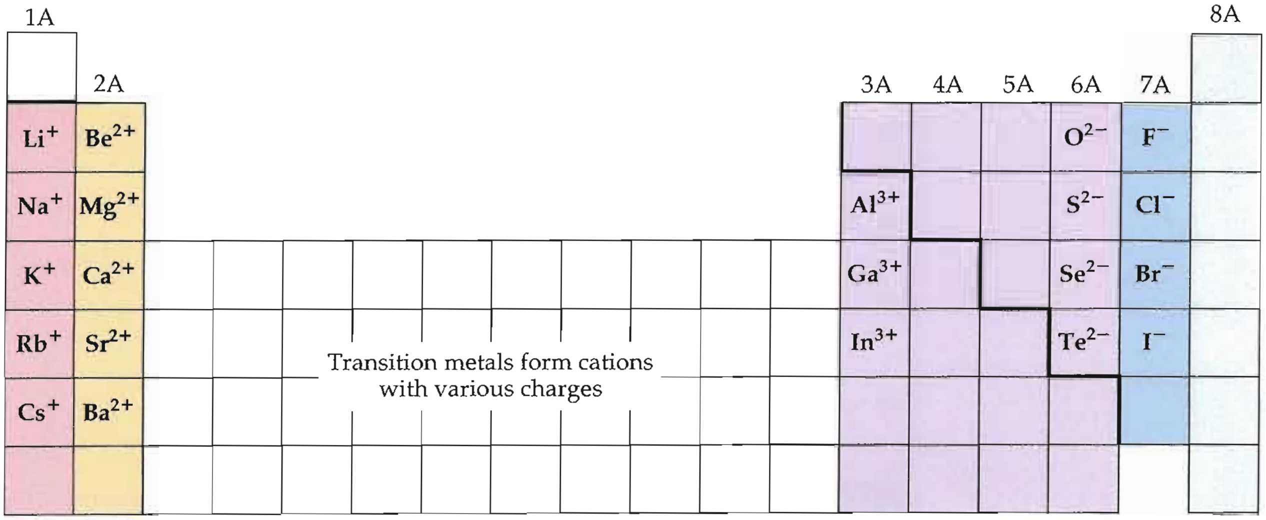
IONS AND THE PERIODIC TABLE

We have learned that in chemical reactions, metals have a tendency to lose electrons and nonmetals have a tendency to gain them. For many main-group elements, we can use the periodic table to predict how many electrons tend to be lost or gained. The number above each *main-group* column in the periodic table—1 through 8—gives the number of *valence electrons* for the elements in that column. We discuss the concept of valence electrons more fully in Chapter 9, but for now, think of valence electrons as the outermost electrons in an atom. For example, since oxygen is in column 6A, we know it has 6 valence electrons; since magnesium is in column 2A, we know it has 2 valence electrons, and so on. Valence electrons are particularly important because, as we shall see in Chapter 10, it is these electrons that take part in chemical bonding.

An important exception to this rule is helium—it is in column 8A, but has only 2 valence electrons.

The key to predicting the charge acquired by a particular metal or nonmetal when it ionizes is as follows:

> Main-group elements tend to form ions that have the same number of valence electrons as the nearest noble gas.



▲ **Figure 4.14** **Elements that form predictable ions**

For example, the closest noble gas to oxygen is neon. When oxygen ionizes, it *acquires* two additional electrons for a total of 8 valence electrons—the same number as neon. The closest noble gas to magnesium is also neon. Therefore magnesium *loses* its 2 valence electrons to attain the same number of valence electrons as neon.

In accordance with this principle, the alkali metals (Group 1A) tend to lose 1 electron and therefore form 1+ ions, while the alkaline earth metals (Group 2A) tend to lose 2 electrons and therefore form 2+ ions. The halogens (Group 7A) tend to gain 1 electron and therefore form 1− ions. The groups in the periodic table that form predictable ions are shown in ▲ Figure 4.14. Be familiar with these groups and the ions they form. In Chapter 9, we examine a theory that more fully explains why these groups form ions as they do.

EXAMPLE 4.6 Charge of Ions from Position in Periodic Table

Based on their position in the periodic table, what ions will barium and iodine tend to form?

Solution:

Since barium is in Group 2A, it will tend to form a cation with a 2+ charge (Ba^{2+}). Since iodine is in Group 7A, it will tend to form an anion with a 1− charge (I^-).

SKILLBUILDER 4.6 Charge of Ions from Position in Periodic Table

Based on their position in the periodic table, what ions will potassium and sulfur tend to form?

FOR MORE PRACTICE Problems 81, 82.

CONCEPTUAL CHECKPOINT 4.3

Which of these pairs of ions would have the same total number of electrons?

(a) Na^+ and Mg^{2+}
(b) F^- and Cl^-
(c) O^- and O^{2-}
(d) Ga^{3+} and Fe^{3+}

4.8 Isotopes: When the Number of Neutrons Varies

All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons. Since neutrons and protons have nearly the same mass (1 amu), and since the number of neutrons in the atoms of a given element can vary, all atoms of a given element *do not* have the same mass (contrary to what John Dalton originally proposed in his atomic theory). For example, all neon atoms in nature contain 10 protons, but they may have 10, 11, or 12 neutrons (▼ Figure 4.15). All three types of neon atoms exist, and each has a slightly different mass. Atoms with the same number of protons but different numbers of neutrons are called **isotopes**. Some elements, such as beryllium (Be) and aluminum (Al), have only one naturally occurring isotope, while other elements, such as neon (Ne) and chlorine (Cl), have two or more.

Fortunately, for nearly all elements, the relative amounts of each different isotope in a naturally occurring sample are the same. For example, in any natural sample of neon atoms, 90.48% of them are the isotope with 10 neutrons, 0.27% are the isotope with 11 neutrons, and 9.25% are the isotope with 12 neutrons as summarized in Table 4.2. This means that out of 10,000 neon atoms, for example, 9048 have 10 neutrons, 27 have 11 neutrons, and 925 have 12 neutrons. These percentages are called the **percent natural abundance** of the isotopes. The preceding numbers are for neon, but all elements have their own unique percent natural abundance of isotopes.

Percent means "per hundred." 90.48% means that 90.48 atoms out of 100 are the isotope with 10 neutrons.

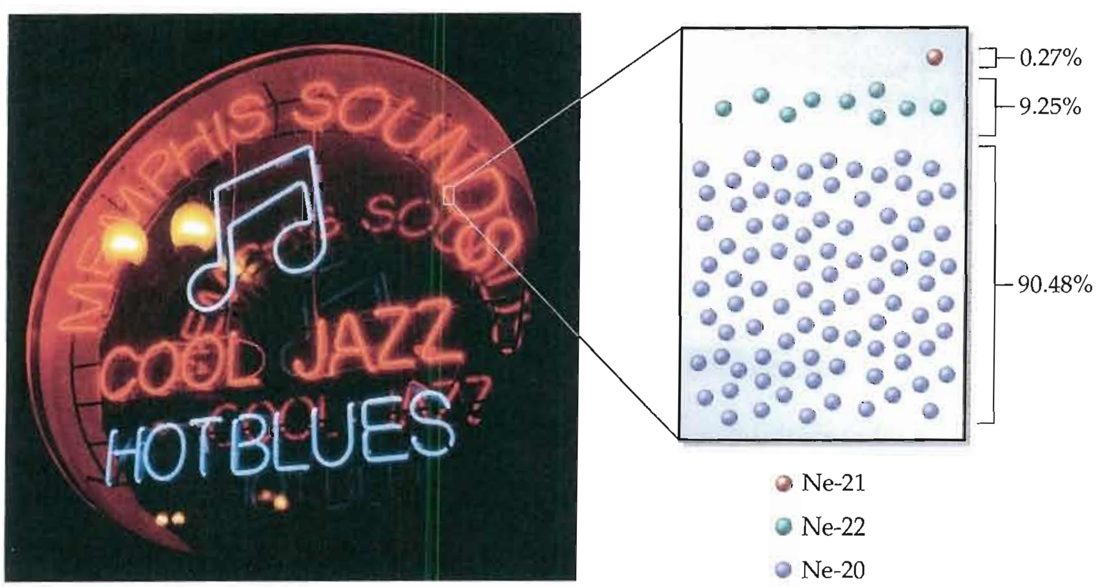
The sum of the number of neutrons and protons in an atom is called the **mass number** and is given the symbol **A**.

$$\mathbf{A = Number\ of\ protons + Number\ of\ neutrons}$$

For neon, which has 10 protons, the mass numbers of the three different naturally occurring isotopes are 20, 21, and 22, corresponding to 10, 11, and 12 neutrons, respectively.

TABLE 4.2 Neon Isotopes

Symbol	Number of Protons	Number of Neutrons	A (Mass Number)	Percent Natural Abundance
Ne-20 or $^{20}_{10}Ne$	10	10	20	90.48%
Ne-21 or $^{21}_{10}Ne$	10	11	21	0.27%
Ne-22 or $^{22}_{10}Ne$	10	12	22	9.25%



▼ **Figure 4.15 Isotopes of neon** Naturally occurring neon contains three different isotopes, Ne-20 (with 10 neutrons), Ne-21 (with 11 neutrons), and Ne-22 (with 12 neutrons).

Isotopes are often symbolized in the following way:

Mass number

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

Atomic number

where X is the chemical symbol, A is the mass number, and Z is the atomic number.

For example, the symbols for the neon isotopes are:

$^{20}_{10}Ne$ $^{21}_{10}Ne$ $^{22}_{10}Ne$

Notice that the chemical symbol, Ne, and the atomic number, 10, are redundant: If the atomic number is 10, the symbol must be Ne, and vice versa. The mass numbers, however, are different, reflecting the different number of neutrons in each isotope.

A second common notation for isotopes is the chemical symbol (or chemical name) followed by a hyphen and the mass number of the isotope.

X—A

Chemical symbol or name Mass number

In this notation, the neon isotopes are:

Ne-20	neon-20
Ne-21	neon-21
Ne-22	neon-22

In general, mass number increases with increasing atomic number.

Notice that all isotopes of a given element have the same number of protons (otherwise they would be a different element). Notice also that the mass number is the *sum* of the number of protons and the number of neutrons. The number of neutrons in an isotope is the difference between the mass number and the atomic number.

EXAMPLE 4.7 Atomic Numbers, Mass Numbers, and Isotope Symbols

What are the atomic number (Z), mass number (A), and symbols of the carbon isotope with 7 neutrons?

Solution:

From the periodic table, we find that the atomic number (Z) of carbon is 6, so carbon atoms have 6 protons. The mass number (A) for the isotope with 7 neutrons is the sum of the number of protons and the number of neutrons.

$$A = 6 + 7 = 13$$

So, $Z = 6$, $A = 13$, and the symbols for the isotope are C-13 and $^{13}_{6}C$.

SKILLBUILDER 4.7 Atomic Numbers, Mass Numbers, and Isotope Symbols

What are the atomic number, mass number, and symbols for the chlorine isotope with 18 neutrons?

FOR MORE PRACTICE Example 4.12; Problems: 87, 89, 91, 92.

EXAMPLE 4.8 Numbers of Protons and Neutrons from Isotope Symbols

How many protons and neutrons are in the following chromium isotope?

$^{52}_{24}\mathrm{Cr}$

The number of protons is equal to Z (lower left number).	**Solution:** $\#p = Z = 24$
The number of neutrons is equal to A (upper left number) − Z (lower left number).	$\#n = A - Z$ $= 52 - 24$ $= 28$

SKILLBUILDER 4.8 Numbers of Protons and Neutrons from Isotope Symbols

How many protons and neutrons are in the following potassium isotope?

$^{39}_{19}\mathrm{K}$

FOR MORE PRACTICE Example 4.13; Problems 93, 94.

CONCEPTUAL CHECKPOINT 4.4

If an atom with a mass number of 27 has 14 neutrons, it must be an isotope of which element?

(a) silicon
(b) aluminum
(c) cobalt
(d) niobium

CONCEPTUAL CHECKPOINT 4.5

Throughout this book, we represent atoms as spheres. For example, a carbon atom is represented by a black sphere as shown below. In light of the nuclear theory of the atom, would C-12 and C-13 look different in this representation of atoms? Why or why not?

Carbon

4.9 Atomic Mass

The mass of an atom is always a bit less than the sum of the masses of its protons and neutrons because some mass is converted to energy when a nucleus forms. This energy then holds the nucleus together.

An important part of Dalton's atomic theory was that all atoms of a given element have the same mass. However, we just learned that because of isotopes, the atoms of a given element may have different masses, so Dalton was not completely correct. We can, however, calculate an average mass—called the **atomic mass**—for each element. The atomic mass of each element is listed in the periodic table directly beneath the element's symbol; it represents the average mass of the atoms that compose that element. For example, the periodic table lists the atomic mass of chlorine as 35.45 amu. Naturally occurring chlorine consists of 75.77% chlorine-35 (mass 34.97 amu) and 24.23% chlorine-37 (mass 36.97 amu). Its atomic mass is:

Some books call this *average atomic mass* or atomic weight instead of simply *atomic mass*.

$$\begin{aligned}\text{Atomic mass} &= (0.7577 \times 34.97\ \text{amu}) + (0.2423 \times 36.97\ \text{amu}) \\ &= 35.45\ \text{amu}\end{aligned}$$

Notice that the atomic mass of chlorine is closer to 35 than 37, because naturally occurring chlorine contains more chlorine-35 atoms than chlorine-37 atoms. Notice also that when percentages are used in these calculations, they must always be converted to their decimal value. To convert a percentage to its decimal value, simply divide by 100. For example:

$$75.77\% = 75.77/100 = 0.7577$$

$$24.23\% = 24.23/100 = 0.2423$$

Chemistry in the Environment

Radioactive Isotopes at Hanford, Washington

Nuclei of the isotopes of a given element are not all equally stable. For example, naturally occurring lead is composed primarily of Pb-206, Pb-207, and Pb-208. Other isotopes of lead also exist, but their nuclei are unstable. Scientists can make some of these other isotopes, such as Pb-185, in the laboratory. However, within seconds Pb-185 atoms emit a few energetic subatomic particles from their nuclei and change into different isotopes of different elements (which are themselves unstable). These emitted subatomic particles are called **nuclear radiation**, and the isotopes that emit them are termed **radioactive**. Nuclear radiation, always associated with unstable nuclei, can be harmful to humans and other living organisms because the energetic particles interact with and damage biological molecules. Some isotopes, such as Pb-185, emit significant amounts of radiation only for a very short time. Others, however, remain radioactive for a long time—in some cases millions or even billions of years.

The nuclear power and nuclear weapons industries both produce by-products containing unstable isotopes of several different elements. Many of these isotopes emit nuclear radiation for a long time, and their disposal is an environmental problem. For example, in Hanford, Washington, which for fifty years produced fuel for nuclear weapons, 177 underground storage tanks contain 55 million gallons of high-level nuclear waste. Certain radioactive isotopes within that waste will produce nuclear radiation for the foreseeable future. Unfortunately, some of the underground storage tanks are aging, and leaks have allowed some of the waste to seep into the environment. While the danger from short-term external exposure to this waste is minimal, ingestion of the waste through contamination of drinking water or food supplies would pose significant health risks. Consequently, Hanford is now the site of the largest environmental cleanup project in U.S. history. The U.S. government expects the project to take more than twenty years and cost about $10 billion.

Radioactive isotopes are not always harmful, however, and many have beneficial uses. For example, technetium-99 (Tc-99) is often given to patients to diagnose disease. The radiation emitted by Tc-99 helps doctors image internal organs or detect infection.

CAN YOU ANSWER THIS? *Give the number of neutrons in each of the following isotopes: Pb-206, Pb-207, Pb-208, Pb-185, Tc-99.*

▲ Storage tanks at Hanford, Washington, contain 55 million gallons of high-level nuclear waste. Each tank pictured here holds 1 million gallons.

In general, the atomic mass is calculated according to the following equation:

$$\begin{aligned}\textbf{Atomic mass} = &\ \textbf{(Fraction of isotope 1} \times \textbf{Mass of isotope 1)} + \\ &\ \textbf{(Fraction of isotope 2} \times \textbf{Mass of isotope 2)} + \\ &\ \textbf{(Fraction of isotope 3} \times \textbf{Mass of isotope 3)} + \ldots\end{aligned}$$

where the fractions of each isotope are the percent natural abundances converted to their decimal values. Atomic mass is useful because it allows us to assign a characteristic mass to each element, and as we will see in Chapter 6, it allows us to quantify the number of atoms in a sample of that element.

EXAMPLE 4.9 Calculating Atomic Mass

Gallium has two naturally occurring isotopes: Ga-69 with mass 68.9256 amu and a natural abundance of 60.11%, and Ga-71 with mass 70.9247 amu and a natural abundance of 39.89%. Calculate the atomic mass of gallium.

Convert the percent natural abundances into decimal form by dividing by 100.	**Solution:** $\text{Fraction Ga-69} = \frac{60.11}{100} = 0.6011$ $\text{Fraction Ga-71} = \frac{39.89}{100} = 0.3989$
Use the fractional abundances and the atomic masses of the isotopes to compute the atomic mass according to the atomic mass definition given earlier.	$\text{Atomic mass} = (0.6011 \times 68.9256\ \text{amu}) + (0.3989 \times 70.9247\ \text{amu})$ $= 41.4\underline{3}12\ \text{amu} + 28.2\underline{9}19\ \text{amu}$ $= 69.7\underline{2}31 = 69.72\ \text{amu}$

SKILLBUILDER 4.9 Calculating Atomic Mass

Magnesium has three naturally occurring isotopes with masses of 23.99, 24.99, and 25.98 amu and natural abundances of 78.99%, 10.00%, and 11.01%. Calculate the atomic mass of magnesium.

FOR MORE PRACTICE Example 4.14; Problems 97, 98.

Chapter in Review

Chemical Principles

The Atomic Theory: Democritus and Leucippus, ancient Greek philosophers, were the first to assert that matter is ultimately composed of small, indestructible particles. It was not until 2000 years later, however, that John Dalton introduced a formal atomic theory stating that matter is composed of atoms; atoms of a given element have unique properties that distinguish them from atoms of other elements; and atoms combine in simple, whole-number ratios to form compounds.

Relevance

The Atomic Theory: The concept of atoms is important because it explains the physical world. You and everything you see are made of atoms. To understand the physical world, we must begin by understanding atoms. Atoms are the key concept—they determine the properties of matter.