Atoms

2

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Image of atoms by SEM (scanning electron microscope).

Key Questions

- 2.1 What Is Matter Made Of?
- 2.2 How Do We Classify Matter?
- 2.3 What Are the Postulates of Dalton's Atomic Theory?
- 2.4 What Are Atoms Made Of?
- 2.5 What Is the Periodic Table?
- 2.6 How Are the Electrons in an Atom Arranged?
- 2.7 How Are Electron Configuration and Position in the Periodic Table Related?
- 2.8 What Is a Periodic Property?

2.1 What Is Matter Made Of?

This question was discussed for thousands of years, long before humans had any reasonable way of getting an answer. In ancient Greece, two schools of thought tried to answer this question. One group, led by a scholar named Democritus (about 460–370 BCE), believed that all matter is made of very small particles—much too small to see. Democritus called these particles atoms (Greek *atomos*, meaning "not to cut"). Some of his followers developed the idea that there were different kinds of atoms, with different properties, and that the properties of the atoms caused ordinary matter to have the properties we all know.

Not all ancient thinkers, however, accepted this idea. A second group, led by Zeno of Elea (born about 450 BCE), did not believe in atoms at all. They insisted that matter is infinitely divisible. If you took any object, such as a piece of wood or a crystal of table salt, you could cut it or otherwise divide it into two parts, divide each of these parts into two more parts, and continue the process forever. According to Zeno and his followers, you would never reach a particle of matter that could no longer be divided.

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Today we know that Democritus was right and Zeno was wrong. Atoms are the basic units of matter. Of course, there is a great difference in the way we now look at this question. Today our ideas are based on evidence. Democritus had no evidence to prove that matter cannot be divided an infinite number of times, just as Zeno had no evidence to support his claim that matter can be divided infinitely. Both claims were based not on evidence, but on visionary belief: one in unity, the other in diversity. In Section 2.3 we will discuss the evidence for the existence of atoms, but first we need to look at the diverse forms of matter.

2.2 How Do We Classify Matter?

Matter can be divided into two classes: pure substances and mixtures. Each class is then subdivided as shown in Figure 2.1.

A. Elements

An **element** is a substance (for example, carbon, hydrogen, and iron) that consists of identical atoms. At this time, 116 elements are known. Of these, 88 occur in nature; chemists and physicists have made the others in the laboratory. A list of the known elements appears on the inside front cover of this book, along with their symbols. Their symbols consist of one or two letters. Many symbols correspond directly to the name in English (for example, C for carbon, H for hydrogen, and Li for lithium), but a few are derived from the Latin or German names. Others are named for people who played significant roles in the development of science—in particular, atomic science (see Problem 2.12). Still other elements are named for geographic locations (See Problem 2.13).

B. Compounds

A **compound** is a pure substance made up of two or more elements in a fixed ratio by mass. For example, water is a compound made up of hydrogen and oxygen and table salt is a compound made up of sodium and chlorine. There are an estimated 20 million known compounds, only a few of which we will introduce in this book.

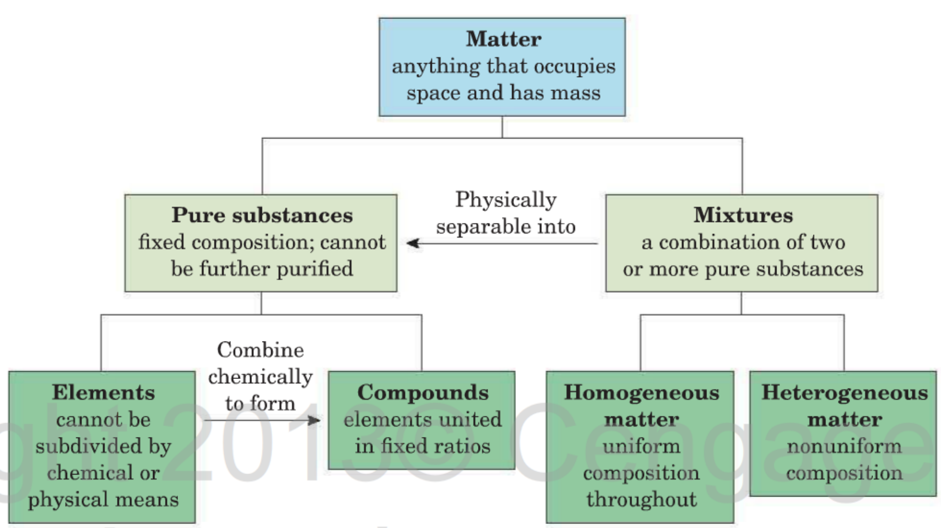


FIGURE 2.1 Classification of matter. Matter is divided into pure substances and mixtures. A pure substance may be either an element or a compound. A mixture may be either homogeneous or heterogeneous.

Chemical Connections 2A

Elements Necessary for Human Life

To the best of our knowledge, 20 of the 116 known elements are necessary for human life. The six most important of these—carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur—are the subjects of organic chemistry and biochemistry (Chapters 10–31). Carbon, hydrogen, nitrogen, and oxygen are the big four in the human body. Seven other elements are also quite important, and our bodies use at least nine additional ones (trace elements) in very small quantities. The table lists these 20 major elements and their functions in the human body. Many of these elements are more fully discussed later in the book. For the average daily requirements for these elements, their sources in foods, and symptoms of their deficiencies, see Chapter 30.

Table 2A Elements and Their Functions in the Human Body

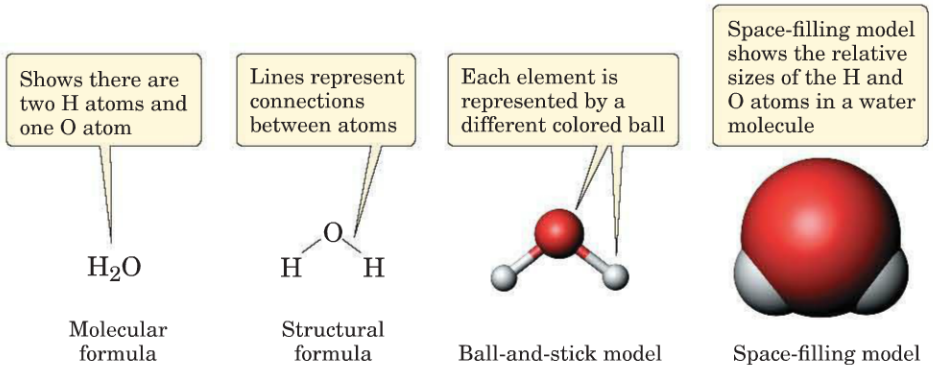
Element	Function	Element	Function
The Big Four		**The Trace Elements**	
Carbon (C)	The subject of Chapters 10–19	Chromium (Cr)	Increases effectiveness of insulin
Hydrogen (H)	(organic chemistry) and 20–31	Cobalt (Co)	Part of vitamin B_{12}
Nitrogen (N)	(biochemistry)	Copper (Cu)	Strengthens bones; assists in enzyme activity
Oxygen (O)		Fluorine (F)	Reduces the incidence of dental cavities
The Next Seven		Iodine (I)	An essential part of thyroid hormones
Calcium (Ca)	Strengthens bones and teeth; aids blood clotting	Iron (Fe)	An essential part of some proteins, such as hemoglobin, myoglobin, cytochromes, and FeS proteins
Chlorine (Cl)	Necessary for normal growth and development	Manganese (Mn)	Present in bone-forming enzymes; aids in fat and carbohydrate metabolism
Magnesium (Mg)	Helps nerve and muscle action; present in bones	Molybdenum (Mo)	Helps regulate electrical balance in body fluids
Phosphorus (P)	Present as phosphates in bone, in nucleic acids (DNA and RNA), and involved in energy storage and transfer	Zinc (Zn)	Necessary for the action of certain enzymes
Potassium (K)	Helps regulate electrical balance in body fluids; essential for nerve conduction		
Sulfur (S)	An essential component of proteins		
Sodium (Na)	Helps regulate electrical balance in body fluids		

Test your knowledge with Problem 2.69.

A compound is characterized by its formula. The formula gives us the ratios of the compound's constituent elements and identifies each element by its atomic symbol. For example, in table salt, the ratio of sodium atoms to chlorine atoms is 1:1. Given that Na is the symbol for sodium and Cl is the symbol for chlorine, the formula of table salt is NaCl. In water, the combining ratio is two hydrogen atoms to one oxygen atom. The symbol for hydrogen is H, that for oxygen is O, and the formula of water is H_2O. The subscripts following the atomic symbols indicate the ratio of the combining elements. The number 1 in these ratios is omitted from the subscript. It is understood that NaCl means a ratio of 1:1 and that H_2O represents a ratio of 2:1. You will find out more about the nature of combining elements in a compound and their names and formulas in Chapter 3.

Figure 2.2 shows four representations for a water molecule. We will have more to say about molecular models as we move through this book.

FIGURE 2.2 Four representations of a water molecule.



Example 2.1 Formula of a Compound

(a) In the compound magnesium fluoride, magnesium (atomic symbol Mg) and fluorine (atomic symbol F) combine in a ratio of 1:2. What is the formula of magnesium fluoride?

(b) The formula of perchloric acid is $HClO_4$. What are the combining ratios of the elements in perchloric acid?

Strategy

The formula gives the atomic symbol of each element combined in the compound, and subscripts give the ratio of its constituent elements.

Solution

(a) The formula is MgF_2. We do not write a subscript of 1 after Mg.

(b) Both H and Cl have no subscripts, which means that hydrogen and chlorine have a combining ratio of 1:1. The subscript on oxygen is 4. Therefore, the combining ratios in $HClO_4$ are 1:1:4.

Problem 2.1

Write the formulas of compounds in which the combining ratios are as follows:

(a) Sodium: chlorine: oxygen, 1:1:3

(b) Aluminum (atomic symbol Al): fluorine (atomic symbol F), 1:3

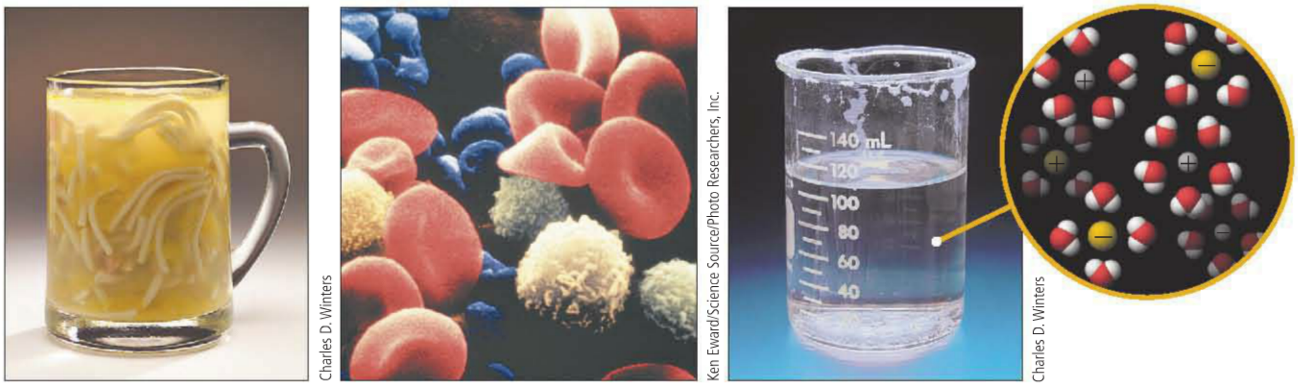
C. Mixtures

A **mixture** is a combination of two or more pure substances. Most of the matter we encounter in our daily lives (including our own bodies) consists of mixtures rather than pure substances. For example, blood, butter, gasoline, soap, the metal in a ring, the air we breathe, and the Earth we walk on are all mixtures of pure substances. An important difference between a compound and a mixture is that the ratios by mass of the elements in a compound are fixed, whereas in a mixture, the pure substances can be present in any mass ratio.

For some mixtures—blood, for example (Figure 2.3)—the texture of the mixture is even throughout. If you examine a mixture under magnification, however, you can see that it is composed of different substances.

Other mixtures are homogeneous throughout, and no amount of magnification will reveal the presence of different substances. The air we breathe, for example, is a mixture of gases, primarily nitrogen (78 percent) and oxygen (21 percent).

An important characteristic of a mixture is that it consists of two or more pure substances, each having different physical properties. If we know the physical properties of the individual substances, we can use appropriate physical means to separate the mixture into its component parts. Figure 2.4 shows one example of how a mixture can be separated.



(*a*) (*b*) (*c*)

FIGURE 2.3 Mixtures. (a) A cup of noodle soup is a heterogeneous mixture. (b) A sample of blood may look homogeneous, but examination with an optical microscope shows that it is, in fact, a heterogeneous mixture of liquid and suspended particles (blood cells). (c) A homogeneous solution of salt, NaCl, in water. The models show that the salt solution contains Na^+ and Cl^- ions as separate particles in water, with each ion being surrounded by a sphere of six or more water molecules. The particles in this solution cannot be seen with an optical microscope because they are too small.

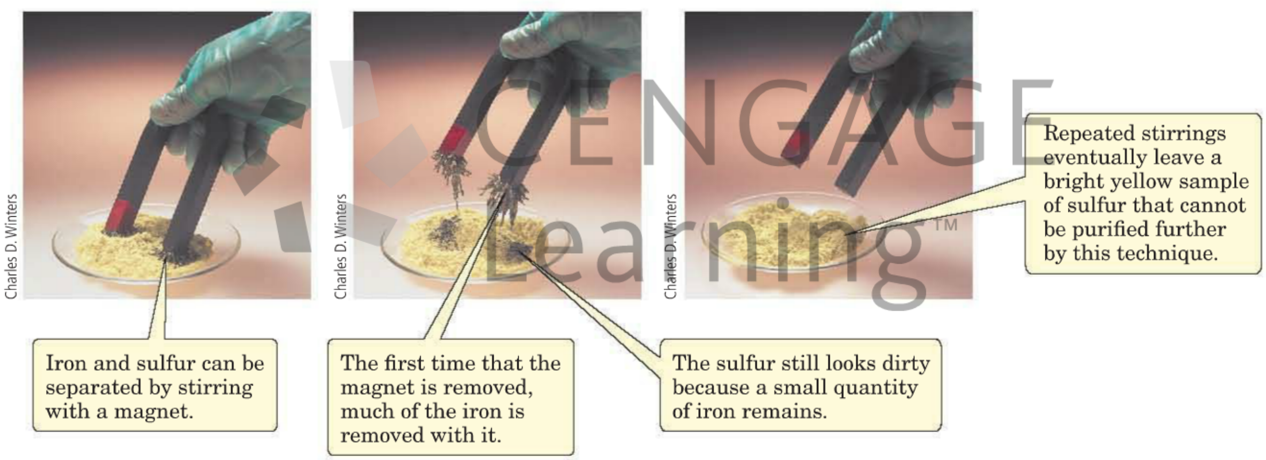


FIGURE 2.4 Separating a mixture of iron and sulfur. (a) The iron–sulfur mixture is stirred with a magnet, which attracts the iron filings. (b) Much of the iron is removed after the first stirring. (c) Stirring continues until no more iron filings can be removed.

2.3 What Are the Postulates of Dalton's Atomic Theory?

In 1808, the English chemist John Dalton (1766–1844) put forth a model of matter that underlies modern scientific atomic theory. The major difference between Dalton's theory and that of Democritus (Section 2.1) is that Dalton based his theory on evidence rather than on a belief. First, let us state his theory. We will then see what kind of evidence supported it.

1. All matter is made up of very tiny, indivisible particles, which Dalton called **atoms.**
2. All atoms of a given element have the same chemical properties. Conversely, atoms of different elements have different chemical properties.
3. In ordinary chemical reactions, no atom of any element disappears or is changed into an atom of another element.

Atom The smallest particle of an element that retains the chemical properties of the element. The interaction among atoms accounts for the properties of matter.

4. Compounds are formed by the chemical combination of two or more different kinds of atoms. In a given compound, the relative numbers of atoms of each kind of element are constant and are most commonly expressed as integers.
5. A **molecule** is a tightly bound combination of two or more atoms that acts as a single unit.

A. Evidence for Dalton's Atomic Theory

The Law of Conservation of Mass

The great French chemist Antoine Laurent Lavoisier (1743–1794) discovered the **law of conservation of mass,** which states that matter can neither be created nor destroyed. In other words, there is no detectable change in mass in an ordinary chemical reaction. Lavoisier proved this law by conducting many experiments in which he showed that the total mass of matter at the end of the experiment was exactly the same as that at the beginning. Dalton's theory explained this fact in the following way: If all matter consists of indestructible atoms (postulate 1) and if no atoms of any element disappear or are changed into an atom of a different element (postulate 3), then any chemical reaction simply changes the attachments between atoms but does not destroy the atoms themselves. Thus, mass is conserved in a chemical reaction.

In the following illustration, a carbon monoxide molecule reacts with a lead oxide molecule to give a carbon dioxide molecule and a lead atom. All of the original atoms are still present at the end; they have merely changed partners. Thus, the total mass after this chemical change is the same as the mass that existed before the reaction took place.

The Law of Constant Composition

Another French chemist, Joseph Proust (1754–1826), demonstrated the **law of constant composition,** which states that any compound is always made up of elements in the same proportion by mass. For example, if you decompose water, you will always get 8.0 g of oxygen for each 1.0 g of hydrogen. The mass ratio of oxygen to hydrogen in pure water is always 8.0 to 1.0, whether the water comes from the Atlantic Ocean or the Missouri River or is collected as rain, squeezed out of a watermelon, or distilled from urine.

This fact was also evidence for Dalton's theory. If a water molecule consists of one atom of oxygen and two atoms of hydrogen and if an oxygen atom has a mass 16 times that of a hydrogen atom, then the mass ratio of these two elements in water must always be 8.0 to 1.0. The two elements can never be found in water in any other mass ratio.

Thus, if the atom ratio of the elements in a compound is fixed (postulate 4), then their proportions by mass must also be fixed.

B. Monatomic, Diatomic, and Polyatomic Elements

Some elements—for example, helium and neon—consist of single atoms that are not connected to each other—that is, they are **monatomic elements.** In contrast, oxygen, in its most common form, contains two atoms in each

Chemical Connections 2B

Abundance of Elements Present in the Human Body and in the Earth's Crust

Table 2B shows the abundance of the elements present in the human body. As you can see, oxygen is the most abundant element by mass, followed by carbon, hydrogen, and nitrogen. If we go by number of atoms, however, hydrogen is even more abundant in the human body than oxygen.

The table also shows the abundance of elements in the Earth's crust. Although 88 elements are found in the Earth's crust (we know very little about the interior of the Earth because we have not been able to penetrate into it very far), they are not present in anything close to equal amounts. In the Earth's crust as well as the human body, the most abundant element by mass is oxygen. But there the similarity ends. Silicon, aluminum, and iron, which are the second, third, and fourth most abundant elements in the Earth's crust, respectively, are not major elements in the body, whereas carbon, the second most abundant element by mass in the human body, is present to the extent of only 0.08 percent in the Earth's crust.

Table 2B The Relative Abundance of Elements Present in the Human Body and in the Earth's Crust, Including the Atmosphere and Oceans

	Percentage in Human Body		
Element	By Number of Atoms	By Mass	Percentage in Earth's Crust by Mass
H	63.0	10.0	0.9
O	25.4	64.8	49.3
C	9.4	18.0	0.08
N	1.4	3.1	0.03
Ca	0.31	1.8	3.4
P	0.22	1.4	0.12
K	0.06	0.4	2.4
S	0.05	0.3	0.06
Cl	0.03	0.2	0.2
Na	0.03	0.1	2.7
Mg	0.01	0.04	1.9
Si	—	—	25.8
Al	—	—	7.6
Fe	—	—	4.7
Others	0.01	—	—

Test your knowledge with Problem 2.70.

molecule, connected to each other by a chemical bond. We write the formula for an oxygen molecule as O_2, with the subscript showing the number of atoms in the molecule. Six other elements also occur as diatomic molecules (that is, they contain two atoms of the same element per molecule): hydrogen (H_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). It is important to understand that under normal conditions, free atoms of O, H, N, F, Cl, Br, and I do not exist. Rather, these seven elements occur only as **diatomic elements** (Figure 2.5).

Some elements have even more atoms in each molecule. Ozone, O_3, has three oxygen atoms in each molecule. In one form of phosphorus, P_4, each molecule has four atoms. One form of sulfur, S_8, has eight atoms per molecule. Some elements have molecules that are much larger. For example, diamond has millions of carbon atoms all bonded together in a gigantic cluster. Diamond and S_8 are referred to as **polyatomic elements.**

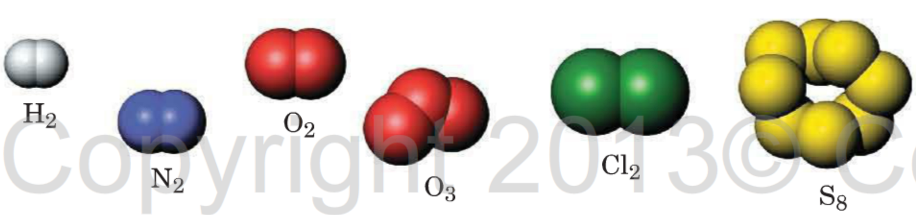


FIGURE 2.5 Some diatomic, triatomic, and polyatomic elements. Hydrogen, nitrogen, oxygen, and chlorine are diatomic elements. Ozone, O_3, is a triatomic element. One form of sulfur, S_8, is a polyatomic element.

2.4 What Are Atoms Made Of?

A. Three Subatomic Particles

There are many other subatomic particles, but we will not deal with them in this book.

Today, we know that matter is more complex than Dalton believed. A wealth of experimental evidence obtained over the last 100 years or so has convinced us that atoms are not indivisible, but rather, consist of even smaller particles called subatomic particles. Three subatomic particles make up all atoms: protons, electrons, and neutrons. Table 2.1 shows the charge, mass, and location of these particles in an atom.

Table 2.1 Properties and Location within Atoms of Protons, Neutrons, and Electrons

Subatomic Particle	Charge	Mass (g)	Mass (amu)	Mass (amu); Rounded to One Significant Figure	Location in an Atom
Proton	+1	1.6726×10^{-24}	1.0073	1	In the nucleus
Electron	−1	9.1094×10^{-28}	5.4858×10^{-4}	0.0005	Outside the nucleus
Neutron	0	1.6749×10^{-24}	1.0087	1	In the nucleus

Proton A subatomic particle with a charge of +1 and a mass of approximately 1 amu; it is found in a nucleus

Atomic mass unit (amu) A unit of the scale of relative masses of atoms: 1 amu = 1.6605×10^{-24} g. By definition, 1 amu is 1/12 the mass of a carbon atom containing 6 protons and 6 neutrons.

Electron A subatomic particle with a charge of −1 and a mass of approximately 0.0005 amu. It is found in the space surrounding a nucleus.

Neutron A subatomic particle with a mass of approximately 1 amu and a charge of zero; it is found in the nucleus

A **proton** has a positive charge. By convention we say that the magnitude of the charge is +1. Thus, one proton has a charge of +1, two protons have a charge of +2, and so forth. The mass of a proton is 1.6726×10^{-24} g, but this number is so small that it is more convenient to use another unit, called the **atomic mass unit (amu),** to describe its mass.

$$1 \text{ amu} = 1.6605 \times 10^{-24} \text{ g}$$

Thus, a proton has a mass of 1.0073 amu. For most purposes in this book, it is sufficient to round this number to one significant figure, and therefore, we say that the mass of a proton is 1 amu.

An **electron** has a charge of −1, equal in magnitude to the charge on a proton, but opposite in sign. The mass of an electron is approximately 5.4858×10^{-4} amu or 1/1837 that of the proton. It takes approximately 1837 electrons to equal the mass of one proton.

Like charges repel, and unlike charges attract. Two protons repel each other, just as two electrons also repel each other. A proton and an electron, however, attract each other.

A **neutron** has no charge. Therefore, neutrons neither attract nor repel each other or any other particle. The mass of a neutron is slightly greater than that of a proton: 1.6749×10^{-24} g or 1.0087 amu. Again, for our purposes, we round this number to 1 amu.

These three particles make up atoms, but where are they found? Protons and neutrons are found in a tight cluster in the center of an atom (Figure 2.6), which is called the **nucleus.** We will discuss the nucleus in greater detail in Chapter 9. Electrons are found as a diffuse cloud outside the nucleus.

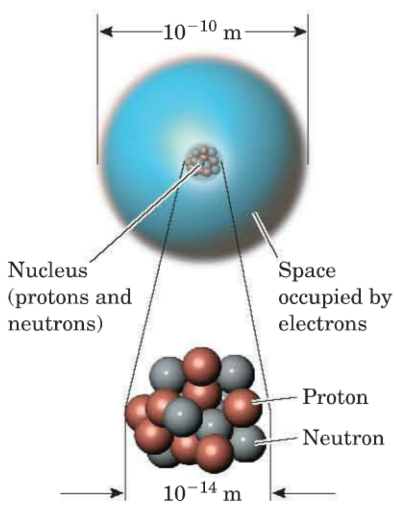


FIGURE 2.6 Relative sizes of the atomic nucleus and an atom (not to scale). The diameter of the region occupied by the electrons is approximately 10,000 times the diameter of the nucleus.

B. Mass Number

Each atom has a fixed number of protons, electrons, and neutrons. One way to describe an atom is by its **mass number** (A), which is the sum of the number of protons and neutrons in its nucleus. Note that an atom also contains electrons, but because the mass of an electron is so small compared to that of protons and neutrons (Table 2.1), electrons are not counted in determining mass number.

$$\text{Mass number (A)} = \text{the number of protons + neutrons in the nucleus of an atom}$$

For example, an atom with 5 protons, 5 electrons, and 6 neutrons has a mass number of 11.

OWL Interactive Example 2.2 Mass Number

What is the mass number of an atom containing:
(a) 58 protons, 58 electrons, and 78 neutrons?
(b) 17 protons, 17 electrons, and 20 neutrons?

Strategy

The mass number of an atom is the sum of the number of protons and neutrons in its nucleus.

Solution

(a) The mass number is 58 + 78 = 136.
(b) The mass number is 17 + 20 = 37.

Problem 2.2

What is the mass number of an atom containing:
(a) 15 protons, 15 electrons, and 16 neutrons?
(b) 86 protons, 86 electrons, and 136 neutrons?

C. Atomic Number

The **atomic number** (Z) of an element is the number of protons in its nucleus.

Atomic numbers for all the known elements are given in the atomic weight table on the inside front cover. They are also given in the Periodic Table on the inside front cover.

$$\text{Atomic number (Z)} = \text{number of protons in the nucleus of an atom}$$

Note that in a neutral atom, the number of electrons is equal to the number protons.

At the present time, 116 elements are known. These elements have atomic numbers from 1 to 116. The smallest atomic number belongs to the element hydrogen, which has only one proton and the largest (so far), to the as-yet-unnamed heaviest known element, which contains 116 protons.

If you know the atomic number and the mass number of an element, you can properly identify it. For example, the element with 6 protons, 6 electrons, and 6 neutrons has an atomic number of 6 and a mass number of 12. The element with atomic number 6 is carbon, C. Because its mass number is 12, we call this atomic nucleus carbon-12. Alternatively, we can write the symbol for this atomic nucleus as $^{12}_{6}C$. In this symbol, the mass number of the element is always written in the upper-left corner (as a superscript) of the symbol of the element and the atomic number in the lower-left corner (as a subscript).

If you know the name of the element, you can look up its atomic number and atomic weight from the atomic weight table facing the inside front cover. Conversely, if you know the atomic number of the element, you can look up its symbol from the Periodic Table on the inside front cover.

Mass number (number of protons + neutrons)
Atomic number (number of protons)
$^{12}_{6}C$ ← Symbol of the element

Example 2.3 Atomic Number

Name the elements given in Example 2.2 and write the symbols for their atomic nuclei.

Strategy

Determine the atomic number (the number of protons in the nucleus) and then locate the element in the Periodic Table on the inside front cover.

Solution

(a) This element has 58 protons. We find in the Periodic Table that the element with atomic number 58 is cerium, and its symbol is Ce. An atom of this element has 58 protons and 78 neutrons, and therefore, its mass number is 136. We call it cerium-136. Its symbol is $^{136}_{58}Ce$.
(b) This atom has 17 protons, making it a chlorine (Cl) atom. Because its mass number is 37, we call it chlorine-37. Its symbol is $^{37}_{17}Cl$.

Problem 2.3

Name the elements given in Problem 2.2. Write the symbols of their atomic nuclei.

Example 2.4 Atomic Nuclei

A number of elements have an equal number of protons and neutrons in their nuclei. Among these are oxygen, nitrogen and, neon. What are the atomic numbers of these elements? How many protons and neutrons does an atom of each have? Write the name and the symbol of each of these atomic nuclei.

Strategy

Look at the Periodic Table to determine the atomic number of each element. Mass number is the number of protons plus the number of neutrons.

Solution

Atomic numbers for these elements are found in the list of elements on the inside back cover. This table shows that oxygen (O) has atomic number 8, nitrogen (N) has atomic number 7, and neon (Ne) has atomic number 10. This means that oxygen has 8 protons and 8 neutrons. Its name is oxygen-16, and its symbol is $^{16}_{8}O$. Nitrogen has 7 protons and 7 neutrons, its name is nitrogen-14, and its symbol is $^{14}_{7}N$. Neon has 10 protons and 10 neutrons, its name is neon-20, and its symbol is $^{20}_{10}Ne$.

Problem 2.4

(a) What are the atomic numbers of mercury (Hg) and lead (Pb)?
(b) How many protons does an atom of each have?
(c) If both Hg and Pb have 120 neutrons in their nuclei, what is the mass number of each isotope?
(d) Write the name and the symbol of each.

D. Isotopes

The fact that isotopes exist means that the second statement of Dalton's atomic theory (Section 2.3) is not correct.

Although we can say that an atom of carbon always has 6 protons and 6 electrons, we cannot say that an atom of carbon must have any particular number of neutrons. Some of the carbon atoms found in nature have 6 neutrons; the mass number of these atoms is 12, they are written as carbon-12, and their symbol is ${}^{12}_{6}C$. Other carbon atoms have 6 protons and 7 neutrons and, therefore, a mass number of 13; they are written as carbon-13, and their symbol is ${}^{13}_{6}C$. Still other carbon atoms have 6 protons and 8 neutrons; they are written as carbon-14 or ${}^{14}_{6}C$. Atoms with the same number of protons but different numbers of neutrons are called **isotopes.** All isotopes of carbon contain 6 protons and 6 electrons (or they wouldn't be carbon atoms). Each isotope, however, contains a different number of neutrons and, therefore, has a different mass number.

The properties of isotopes of the same element are almost identical, and for most purposes, we regard them as identical. They differ, however, in radioactive properties, which we discuss in Chapter 9.

Example 2.5 Isotopes

How many neutrons are in each isotope of oxygen? Write the symbol of each isotope.

(a) Oxygen-16 (b) Oxygen-17 (c) Oxygen-18

Strategy

Each oxygen atom has 8 protons. The difference between the mass number and the number of protons gives the number of neutrons.

Solution

(a) Oxygen-16 has $16 - 8 = 8$ neutrons. Its symbol is ${}^{16}_{8}O$.
(b) Oxygen-17 has $17 - 8 = 9$ neutrons. Its symbol is ${}^{17}_{8}O$.
(c) Oxygen-18 has $18 - 8 = 10$ neutrons. Its symbol is ${}^{18}_{8}O$.

Problem 2.5

Two iodine isotopes are used in medical treatments: iodine-125 and iodine-131. How many neutrons are in each isotope? Write the symbol for each isotope.

Most elements are found on Earth as mixtures of isotopes, in a more or less constant ratio. For example, all naturally occurring samples of the element chlorine contain 75.77% chlorine-35 (18 neutrons) and 24.23% chlorine-37 (20 neutrons). Silicon exists in nature in a fixed ratio of three isotopes, with 14, 15, and 16 neutrons, respectively. For some elements, the ratio of isotopes may vary slightly from place to place, but for most purposes, we can ignore these slight variations. The atomic masses and isotopic abundances are determined using an instrument called a mass spectrometer.

E. Atomic Weight

The **atomic weight** of an element given in the Periodic Table is a weighted average of the masses (in amu) of its isotopes found on the Earth. As an example of the calculation of atomic weight, let us examine chlorine. As we have just seen, two isotopes of chlorine exist in nature, chlorine-35 and chlorine-37. The mass of a chlorine-35 atom is 34.97 amu, and the mass of a chlorine-37 atom is 36.97 amu. Note that the atomic weight of each chlorine isotope (its mass in amu) is very close to its mass number (the number of protons and neutrons in its nucleus). This statement holds true for the isotopes of chlorine and those of all elements, because protons and neutrons have a mass of approximately (but not exactly) 1 amu.

Atomic weight The weighted average of the masses of the naturally occurring isotopes of the element. The units of atomic weight are atomic mass units (amu).

The atomic weight of chlorine is a weighted average of the masses of the two naturally occurring chlorine isotopes:

$$\left(\frac{75.77}{100}\times \overset{\text{Chlorine-35}}{34.97\text{ amu}}\right)+\left(\frac{24.23}{100}\times \overset{\text{Chlorine-37}}{36.97\text{ amu}}\right)=35.45\text{ amu}$$

17
Cl
35.4527

Atomic weight in the Periodic Table is given to four decimal places using more precise data than given here

Some elements—for example, gold, fluorine, and aluminum—occur naturally as only one isotope. The atomic weights of these elements are close to whole numbers (gold, 196.97 amu; fluorine, 18.998 amu; aluminum, 26.98 amu). A table of atomic weights is found facing the inside front cover of this book.

Example 2.6 Atomic Weight

The natural abundances of the three stable isotopes of magnesium are 78.99% magnesium-24 (23.98504 amu), 10.00% magnesium-25 (24.9858 amu), and 11.01% magnesium-26 (25.9829 amu). Calculate the atomic weight of magnesium and compare your value with that given in the Periodic Table.

Strategy

To calculate the weighted average of the masses of the isotopes, multiply each atomic mass by its abundance and then add.

Solution

$$\left(\frac{78.99}{100}\times \overset{\text{Magnesium-24}}{23.99\text{ amu}}\right)+\left(\frac{10.00}{100}\times \overset{\text{Magnesium-25}}{24.99\text{ amu}}\right)+\left(\frac{11.01}{100}\times \overset{\text{Magnesium-26}}{25.98\text{ amu}}\right)=$$

$$18.95 \quad + \quad 2.499 \quad + \quad 2.860 \quad = 24.31\text{ amu}$$

The atomic weight of magnesium given in the Periodic Table to four decimal places is 24.3050.

Problem 2.6

The atomic weight of lithium is 6.941 amu. Lithium has only two naturally occurring isotopes: lithium-6 and lithium-7. Estimate which isotope of lithium is in greater natural abundance.

F. The Mass and Size of an Atom

A typical heavy atom (although not the heaviest) is lead-208, a lead atom with 82 protons, 82 electrons, and 208 − 82 = 126 neutrons. It has a mass of 3.5×10^{-22} g. You would need 1.3×10^{24} atoms (a very large number) of lead-208 to make 1 lb. of lead. There are approximately 7 billion people on Earth right now. If you divided 1 lb. of these atoms among all the people on Earth, each person would get about 2.2×10^{14} atoms.

An atom of lead-208 has a diameter of about 3.1×10^{-10} m. If you could line them up with the atoms just touching, it would take 82 million lead atoms to make a line 1 inch long. Despite their tiny size, we can actually see atoms, in certain cases, by using a special instrument called a scanning tunneling microscope (Figure 2.7).

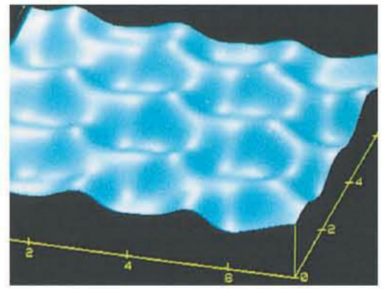
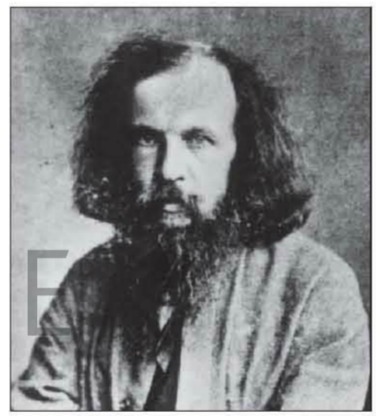


FIGURE 2.7 The surface of graphite is revealed with a scanning tunneling microscope. The contours represent the arrangement of individual carbon atoms on a crystal surface.

Virtually all of the mass of an atom is concentrated in its nucleus (because the nucleus contains the protons and neutrons). The nucleus of a lead-208 atom, for example, has a diameter of about 1.6×10^{-14} m. When you compare this with the diameter of a lead-208 atom, which is about 3.1×10^{-10} m, you see that the nucleus occupies only a tiny fraction of the total volume of the atom. If the nucleus of a lead-208 atom were the size of a baseball, then the entire atom would be much larger than a baseball stadium. In fact, it would be a sphere about one mile in diameter. Because a nucleus has such a relatively large mass concentrated in such a relatively small volume, a nucleus has a very high density. The density of a lead-208 nucleus, for example, is 1.6×10^{14} g/cm^3. Nothing in our daily life has a density anywhere near as high. If a paper clip had this density, it would weigh about 10 million (10^7) tons.

2.5 What Is the Periodic Table?

A. Origin of the Periodic Table



Dmitri Mendeleyev.

In the 1860s, the Russian scientist Dmitri Mendeleyev (1834–1907), then professor of chemistry at the University of St. Petersburg, produced one of the first Periodic Tables, the form of which we still use today. Mendeleyev started by arranging the known elements in order of increasing atomic weight beginning with hydrogen. He soon discovered that when the elements are arranged in the order of increasing atomic weight, certain sets of properties recur periodically. Mendeleyev then arranged those elements with recurring properties into **periods** (horizontal rows) by starting a new row each time he came to an element with properties similar to hydrogen. In this way, he discovered that lithium, sodium, potassium, and so forth, each start new rows. All are metallic solids at room temperature, all form ions with a charge of +1 (Li^+, Na^+, K^+, and so on), and all react with water to form metal hydroxides (LiOH, NaOH, KOH, and so on). Mendeleyev also discovered that elements in other vertical columns (families) have similar properties.

Period of the Periodic Table A horizontal row of the Periodic Table

Family in the Periodic Table The elements in a vertical column of the Periodic Table

For example, the elements fluorine (atomic number 9), chlorine (17), bromine (35), and iodine (53) all fall in the same column of the table. These elements, which are called **halogens,** are all colored substances, with the color deepening as we go down the table (Figure 2.8). All form compounds with sodium that have the general formula NaX (for example, NaCl and NaBr), but not NaX_2, Na_2X, Na_3X, or anything else. Only the elements in this column share this property.

"X" is a commonly used symbol for a halogen.

At this point, we must say a word about the numbering of the columns (families or groups) of the Periodic Table. Mendeleyev gave them numerals and added the letter A for some columns and B for others. This numbering pattern remains in common use in the United States today. In 1985, an alternative pattern was recommended by the International Union of Pure and Applied Chemistry (IUPAC). In this system, the groups are numbered 1 to 18, without added letters, beginning on the left. Thus, in Mendeleyev's numbering system, the halogens are in Group 7A; in the new international numbering system, they are in Group 17. Although this book uses the Mendeleyev numbering system, both patterns are shown on the Periodic Table on the inside front cover. The A group elements (Groups 1A and 2A on the left side of the table and Groups 3A through 8A at the right) are known collectively as **main-group elements.**

Main-group element An element in the A groups (Groups 1A, 2A, and 3A–8A) of the Periodic Table

The elements in the B columns (Groups 3 to 12 in the new numbering system) are called **transition elements.** Notice that elements 58 to 71 and

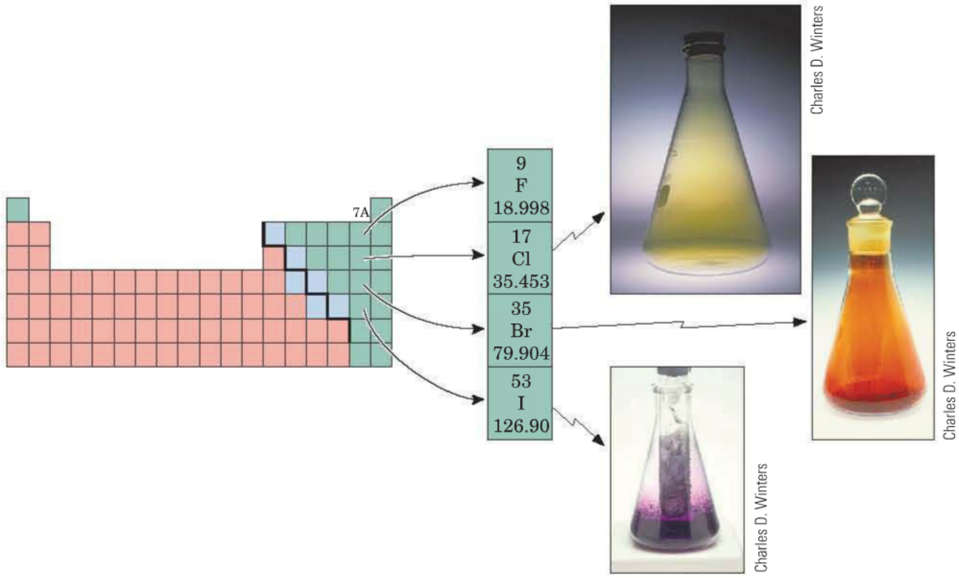


FIGURE 2.8 Four halogens. Fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid.

90 to 103 are not included in the main body of the table but rather are shown separately at the bottom. These sets of elements, called **inner transition elements,** actually belong in the main body of the Periodic Table, between columns 3 and 4 (between La and Hf and Ac and Rf). As is customary, we put them outside the main body solely to make a more compact presentation. If you like, you may mentally take a pair of scissors, cut through the heavy line between columns 3B and 4B, move them apart, and insert the inner transition elements. You will now have a table with 32 columns.

B. Classification of the Elements

There are three classes of elements: metals, nonmetals, and metalloids. The majority of elements are **metals**—only 24 are not. Metals are solids at room temperature (except for mercury, which is a liquid), shiny, conductors of electricity, ductile (they can be drawn into wires), and malleable (they can be hammered and rolled into sheets). In their reactions, metals tend to give up electrons. They also form alloys, which are solutions of one or more metals dissolved in another metal. Brass, for example, is an alloy of copper and zinc. Bronze is an alloy of copper and tin, and pewter is an alloy of tin, antimony, and lead. In their chemical reactions, metals tend to give up electrons (Section 3.2). Figure 2.9 shows a form of the Periodic Table in which the elements are classified by type.

Metal An element that is a solid at room temperature (except for mercury, which is a liquid), shiny, conducts electricity, is ductile and malleable, and forms alloys. In their reactions, metals tend to give up electrons.

FIGURE 2.9 Classification of the elements.

Metals | Metalloids | Nonmetals

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Te	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	≠	≠	≠	≠		

≠ Not yet named

Chemical Connections 2C

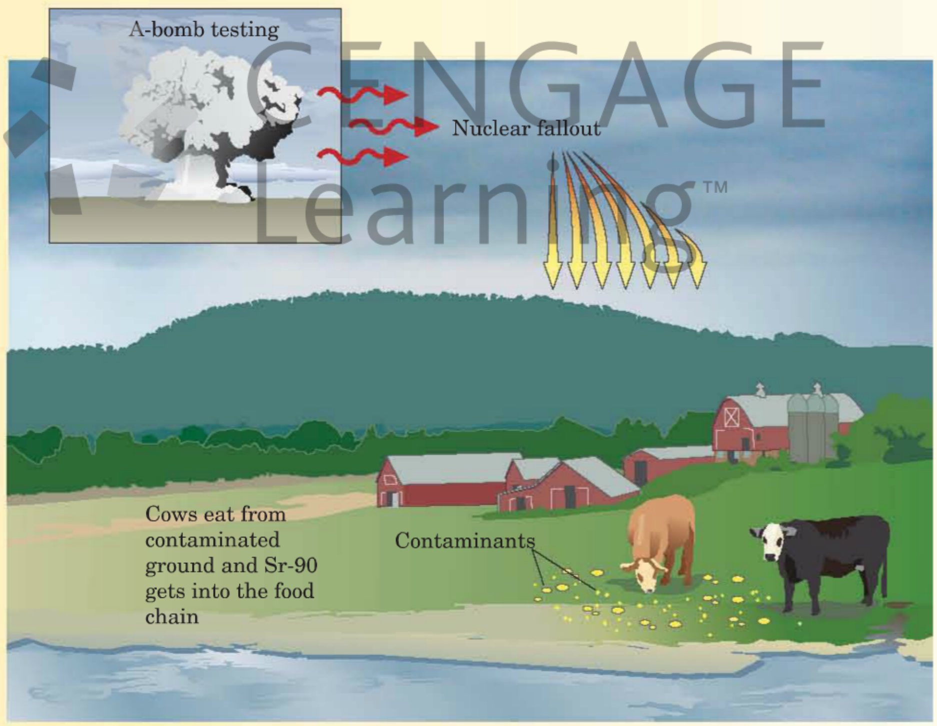
Strontium-90

Elements in the same column of the Periodic Table show similar properties. One important example is the similarity of strontium (Sr) and calcium (strontium is just below calcium in Group 2A). Calcium is an important element for humans because our bones and teeth consist largely of calcium compounds. We need some of this mineral in our diet every day, and we get it mostly from milk, cheese, and other dairy products.

One of the products released by test nuclear explosions in the 1950s and 1960s was the isotope strontium-90. This isotope is radioactive, with a half-life of 28.1 years. (Half-life is discussed in Section 9.4.) Strontium-90 was present in the fallout from aboveground nuclear test explosions. It was carried all over the Earth by winds and slowly settled to the ground, where it was eaten by cows and other animals. Strontium-90 got into milk and eventually into human bodies as well. If it were not so similar to calcium, our bodies would eliminate it within a few days. Because it is similar, however, some of the strontium-90 became deposited in bones and teeth (especially in children), subjecting all of us to a small amount of radioactivity for long periods of time.

In 1958, pathologist Walter Bauer helped start the St. Louis Baby Tooth Survey to study the effects of nuclear fallout on children. The study helped establish an early '60s ban on aboveground A-bomb testing and led to similar surveys across the United States and the rest of the world. By 1970, the team had collected 300,000 shed primary teeth, which they discovered had absorbed nuclear waste from the milk of cows that were fed contaminated grass.

A 1963 treaty between the United States and the former Soviet Union banned aboveground nuclear testing. Although a few other countries still conducted occasional aboveground tests, there is reason to hope that such testing will be completely halted in the future.



Test your knowledge with Problem 2.71.

Nonmetals are the second class of elements. With the exception of hydrogen, the 18 nonmetals appear to the right side of the Periodic Table. With the exception of graphite, which is one form of carbon, nonmetals do not conduct electricity. At room temperature, nonmetals such as phosphorus and iodine are solids. Bromine is a liquid, and the elements of Group 8A (the noble gases)—helium through radon—are gases. In their chemical reactions, nonmetals tend to accept electrons (Section 3.2). Virtually all of the

Nonmetal An element that does not have the characteristic properties of a metal and, in its reactions, tends to accept electrons. Eighteen elements are classified as nonmetals.

(*a*) Metals (*b*) Nonmetals (*c*) Metalloids

FIGURE 2.10 Representative elements. (a) Magnesium, aluminum, and copper are metals. All can be drawn into wires and conduct electricity. (b) Only 18 or so elements are classified as nonmetals. Shown here are liquid bromine and solid iodine. (c) Only six elements are generally classified as metalloids. This photograph is of solid silicon in various forms, including a wafer on which electronic circuits are printed.

compounds we will encounter in our study of organic and biochemistry are built from just six nonmetals: H, C, N, O, P, and S.

Metalloid An element that displays some of the properties of metals and some of the properties of nonmetals. Six elements are classified as metalloids.

Six elements are classified **metalloids:** boron, silicon, germanium, arsenic, antimony, and tellurium.

B	Si	Ge	As	Sb	Te
Boron	Silicon	Germanium	Arsenic	Antimony	Tellurium

These elements have some properties of metals and some of nonmetals. For example, some metalloids are shiny like metals, but do not conduct electricity. One of these metalloids, silicon, is a semiconductor—that is, it does not conduct electricity under certain applied voltages, but becomes a conductor at higher applied voltages. This semiconductor property of silicon makes it a vital element for Silicon Valley–based companies and the entire electronics industry (Figure 2.10).

Although hydrogen (H) appears in Group 1A, it is not an alkali metal; it is a nonmetal. Hydrogen is placed in Group 1A because of its electron configuration (Section 2.7).

C. Examples of Periodicity in the Periodic Table

Halogen An element in Group 7A of the Periodic Table

Alkali metal An element, except hydrogen, in Group 1A of the Periodic Table

Not only do the elements in any particular column (group or family) of the Periodic Table share similar properties, but the properties also vary in some fairly regular ways as we go up or down a column (family). For instance, Table 2.2 shows that the melting and boiling points of the **halogens** regularly increase as we go down a column.

Another example involves the Group 1A elements, also called the **alkali metals.** All alkali metals are soft enough to be cut with a knife, and their softness increases going down the column. They have relatively low melting and boiling points, which decrease going down the columns (Table 2.3).

All alkali metals react with water to form hydrogen gas, H_2, and a metal hydroxide with the formula MOH, where "M" stands for the alkali metal. The violence of their reaction with water increases in going down the column.

$$\underset{\text{Sodium}}{2Na} + \underset{\text{Water}}{2H_2O} \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH} + \underset{\text{Hydrogen gas}}{H_2}$$

Sodium metal can be cut with a knife.

Chemical Connections 2D

The Use of Metals as Historical Landmarks

The malleability of metals played an important role in the development of human society. In the Stone Age, tools were made from stone, which has no malleability. Then, about 11,000 BCE, it was discovered that the pure copper found on the surface of the Earth could be hammered into sheets, which made it suitable for use in vessels, utensils, and religious and artistic objects. This period became known as the Copper Age. Pure copper on the surface of the Earth, however, is scarce. Around 5000 BCE, humans found that copper could be obtained by putting malachite, $Cu_2CO_3(OH)_2$, a green copper-containing stone, into a fire. Malachite yielded pure copper at the relatively low temperature of 200°C.

Copper is a soft metal made of layers of large copper crystals. It can easily be drawn into wires because the layers of crystals can slip past one another. When hammered, the large crystals break into smaller ones with rough edges and the layers can no longer slide past one another. Therefore, hammered copper sheets are harder than drawn copper. Using this knowledge, the ancient profession of coppersmith was born, and beautiful plates, pots, and ornaments were produced.

Werner Forman/Art Resource, NY

Bronze Age artifact.

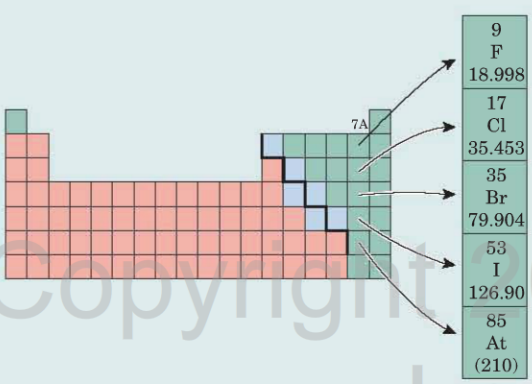
Around 4000 BCE, it was discovered that an even greater hardness could be achieved by mixing molten copper with tin. The resulting alloy is called bronze. The Bronze Age was born somewhere in the Middle East and quickly spread to China and all over the world. Because hammered bronze takes an edge, knives and swords could be manufactured using it.

An even harder metal was soon to come. The first raw iron was found in meteorites. (The ancient Sumerian name of iron is "metal from heaven.") Around 2500 BCE, it was discovered that iron could be recovered from its ore by smelting, the process of recovering a metal from its ore by heating the ore. Thus began the Iron Age. More advanced technology was needed for smelting iron ores because iron melts only at a high temperature (about 1500°C). For this reason, it took a longer time to perfect the smelting process and to learn how to manufacture steel, which is about 90–95% iron and 5–10% carbon. Steel objects appeared first in India around 100 BCE.

Modern anthropologists and historians look back at ancient cultures and use the discovery of a new metal as a landmark for that age.

Test your knowledge with Problems 2.72 and 2.73.

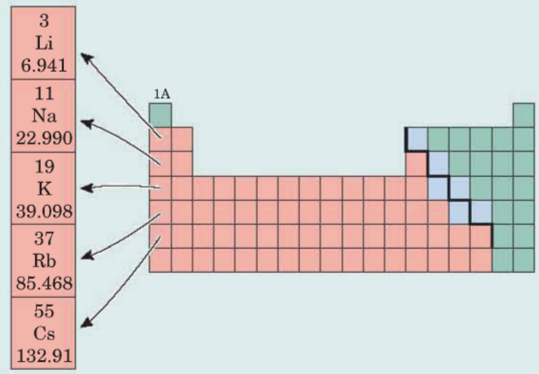
Table 2.2 Melting and Boiling Points of the Halogens (Group 7A Elements)



Element	Melting point (°C)	Boiling point (°C)
Fluorine	−220	−188
Chlorine	−101	−35
Bromine	−7	59
Iodine	114	184
Astatine	302	337

Table 2.3 Melting and Boiling Points of the Alkali Metals (Group 1A Elements)

Element	Melting point (°C)	Boiling point (°C)
Lithium	180	1342
Sodium	98	883
Potassium	63	760
Rubidium	39	686
Cesium	28	669



They also form compounds with the halogens with the formula MX, where "X" stands for the halogen.

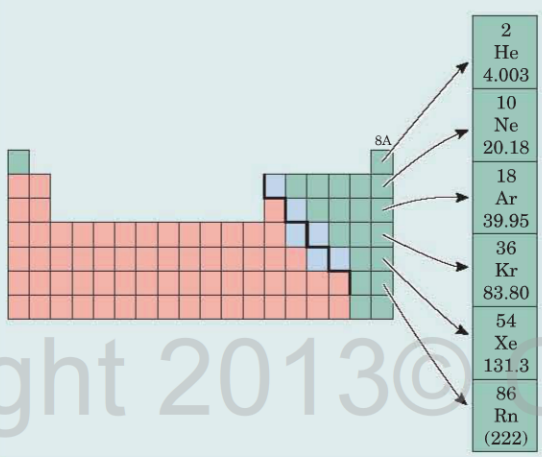
$$2Na + Cl_2 \longrightarrow 2NaCl$$

Sodium Chlorine Sodium chloride

The elements in Group 8A, often called the **noble gases,** provide yet another example of how the properties of elements change gradually within a column. Group 8A elements are gases under normal temperature and pressure, and they form either no compounds or very few compounds. Notice how close the melting and boiling points of the elements in this series are to one another (Table 2.4).

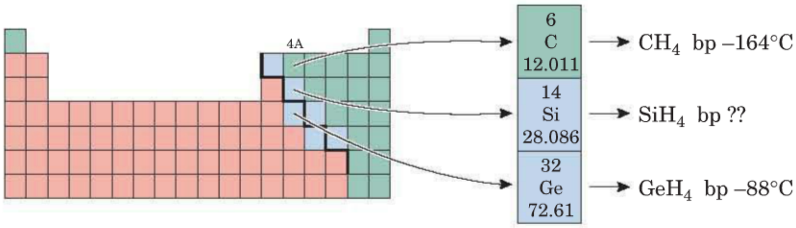
The Periodic Table is so useful that it hangs in nearly every chemistry classroom and chemical laboratory throughout the world. What makes it so useful is that it correlates a vast amount of data about the elements and their compounds and allows us to make many predictions about both chemical and physical properties. For example, if you were told that the boiling point of germane (GeH_4) is −88°C and that of methane (CH_4) is −164°C,

Table 2.4 Melting and Boiling Points of the Noble Gases (Group 8A Elements)



Element	Melting point (°C)	Boiling point (°C)
Helium	−272	−269
Neon	−249	−246
Argon	−189	−186
Krypton	−157	−152
Xenon	−112	−107
Radon	−71	−62

could you predict the boiling point of silane (SiH_4)? The position of silicon in the table, between germanium and carbon, might lead you to a prediction of about −125°C. The actual boiling point of silane is −112°C, not far from this prediction.



2.6 How Are the Electrons in an Atom Arranged?

We have seen that the protons and neutrons of an atom are concentrated in the atom's very small nucleus and that the electrons of an atom are located in the considerably larger space outside the nucleus. We can now ask how the electrons of an atom are arranged in this extranuclear space. Are they arranged randomly like seeds in a watermelon, or are they organized into layers like the layers of an onion?

Let us begin with hydrogen because it has only one electron and is the simplest atom. Before we do so, however, it is necessary to describe a discovery made in 1913 by the Danish physicist Niels Bohr (1885–1962). At the time, it was known that an electron is always moving around the nucleus and so possesses kinetic energy. Bohr discovered that only certain values are possible for this energy. This was a very surprising discovery. If you were told that you could drive your car at 23.4 mi/h or 28.9 mi/h or 34.2 mi/h, but never at any speed in between these values, you wouldn't believe it. Yet, that is just what Bohr discovered about electrons in atoms. The lowest possible energy level is the **ground state.**

Niels Bohr was awarded the 1922 Nobel Prize for physics. In addition, element 107 was named Bohrium in his honor.

Ground-state electron configuration The electron configuration of the lowest energy state of an atom

If an electron is to have more energy than it has in the ground state, only certain values are allowed; values in between are not permitted. Bohr was unable to explain why these energy levels of electrons exist in atoms, but the accumulated evidence forced him to the conclusion that they do. We say that the energy of electrons in atoms is quantized. We can liken quantization to walking up a ramp compared with walking up a flight of stairs (Figure 2.11). You can put your foot on any stair step, but you cannot stand any place between two steps. You can stand only on steps.

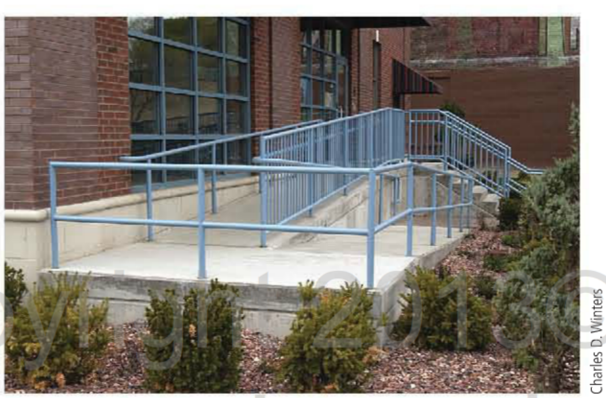


FIGURE 2.11 An energy stairway. A ramp, foreground (not quantized), and stair steps, background (quantized).

A. Electrons Are Distributed in Shells, Subshells, and Orbitals

Principal energy level An energy level containing orbitals of the same number (1, 2, 3, 4, and so forth)

Shell All orbitals of a principal energy level of an atom

One conclusion reached by Bohr is that electrons in atoms do not move freely in the space around the nucleus, but rather remain confined to specific regions of space called **principal energy levels,** or more simply, **shells.** These shells are numbered 1, 2, 3, and 4, and so on, from the inside out. Table 2.5 gives the number of electrons that each of the first four shells can hold.

Table 2.5 Distribution of Electrons in Shells

Shell	Number of Electrons Shell Can Hold	Relative Energies of Electrons in Each Shell
4	32	Higher
3	18	↑
2	8	
1	2	Lower

Electrons in the first shell are closest to the positively charged nucleus and are held most strongly by it; these electrons are said to be the lowest in energy (hardest to remove). Electrons in higher-numbered shells are farther from the nucleus and are held less strongly to it; these electrons are said to be higher in energy (easier to remove).

Subshell All of the orbitals of an atom having the same principal energy level and the same letter designation (either *s*, *p*, *d*, or *f*)

Orbital A region of space around a nucleus that can hold a maximum of two electrons

Shells are divided into **subshells** designated by the letters *s*, *p*, *d*, and *f*. Within these subshells, electrons are grouped in **orbitals.** An orbital is a region of space and can hold two electrons (Table 2.6). The first shell contains a single *s* orbital and can hold two electrons. The second shell contains one *s* orbital and three *p* orbitals. All *p* orbitals come in sets of three and can hold six electrons. The third shell contains one *s* orbital, three *p* orbitals, and five *d* orbitals. All *d* orbitals come in sets of five and can hold ten electrons. The fourth shell also contains a set of *f* orbitals. All *f* orbitals come in sets of seven and can hold 14 electrons.

Table 2.6 Distribution of Orbitals within Shells

Shell	Orbitals Contained in Each Shell	Maximum Number of Electrons Shell Can Hold
4	One 4*s*, three 4*p*, five 4*d*, and seven 4*f* orbitals	2 + 6 + 10 + 14 = 32
3	One 3*s*, three 3*p*, and five 3*d* orbitals	2 + 6 + 10 = 18
2	One 2*s* and three 2*p* orbitals	2 + 6 = 8
1	One 1*s* orbital	2

B. Orbitals Have Definite Shapes and Orientations in Space

All *s* orbitals have the shape of a sphere with the nucleus at the center of the sphere. Figure 2.12 shows the shapes of the 1*s* and 2*s* orbitals. Of the *s* orbitals, the 1*s* is the smallest sphere, the 2*s* is a larger sphere, and the 3*s* (not shown) is a still larger sphere. Figure 2.12 also shows the three-dimensional shapes of the three 2*p* orbitals. Each 2*p* orbital has the shape of a dumbbell with the nucleus at the midpoint of the dumbbell. The three 2*p* orbitals are at right angles to each other, with one orbital on the *x*-axis, the second on

The *d* and *f* orbitals are less important to us, so we will not discuss their shapes.

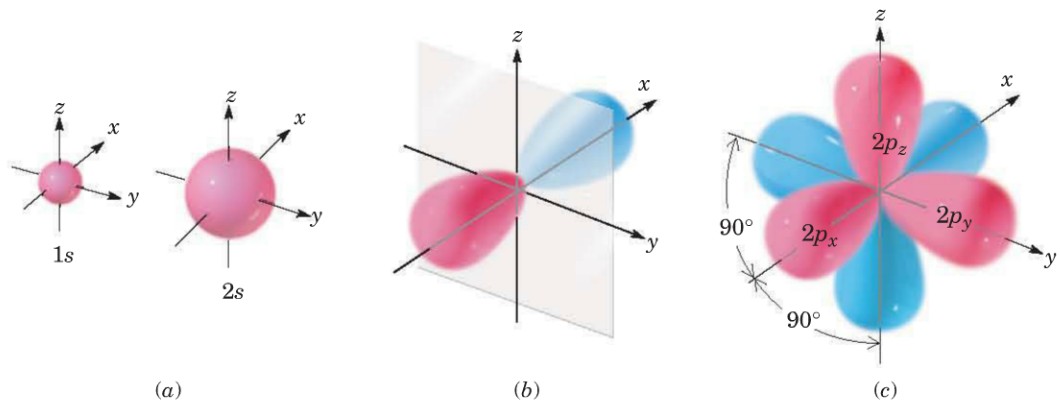


FIGURE 2.12 The 1*s*, 2*s*, and 2*p* orbitals. (a) A 1*s* orbital has the shape of a sphere, with the nucleus at the center of the sphere. A 2*s* orbital is a larger sphere than a 1*s* orbital, and a 3*s* orbital (not shown) is larger still. (b) A 2*p* orbital has the shape of a dumbbell, with the nucleus at the midpoint of the dumbbell. (c) Each 2*p* orbital is perpendicular to the other two. The 3*p* orbitals are similar in shape but larger. To make it easier for you to see the two lobes of each 2*p* orbital, one lobe is colored red and the other is colored blue.

the *y*-axis, and the third on the *z*-axis. The shapes of 3*p* orbitals are similar, but larger.

Because the vast majority of organic compounds and biomolecules consist of the elements H, C, N, O, P, and S, which use only 1*s*, 2*s*, 2*p*, 3*s*, and 3*p* orbitals for bonding, we will concentrate on just these and other elements of the first, second, and third periods of the Periodic Table.

C. Electron Configurations of Atoms Are Governed by Three Rules

The **electron configuration** of an atom is a description of the orbitals that its electrons occupy. The orbitals available to all atoms are the same—namely, 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, and so on. In the ground state of an atom, only the lowest-energy orbitals are occupied; all other orbitals are empty. We determine the ground-state electron configuration of an atom using the following rules:

Electron configuration A description of the orbitals of an atom or ion occupied by electrons

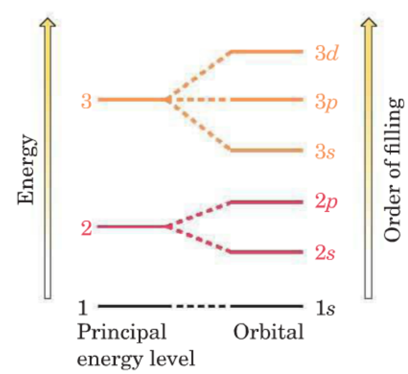


FIGURE 2.13 Energy levels for orbitals through the third shell.

Rule 1. Orbitals fill in the order of increasing energy from lowest to highest.

Example: In this book, we are concerned primarily with elements of the first, second, and third periods of the Periodic Table. Orbitals in these elements fill in the order 1*s*, 2*s*, 2*p*, 3*s*, and 3*p*. Figure 2.13 shows the order of filling through the third period.

Rule 2: Each orbital can hold up to two electrons with spins paired.

Example: With four electrons, the 1*s* and 2*s* orbitals are filled and we write them as $1s^2 2s^2$. With an additional six electrons, the three 2*p* orbitals are filled and we write them either in the expanded form of $2p_x^2\ 2p_y^2\ 2p_z^2$, or in the condensed form $2p^6$. Spin pairing means that the electrons spin in opposite directions (Figure 2.14).

Rule 3: When there is a set of orbitals of equal energy, each orbital becomes half-filled before any of them becomes completely filled.

Example: After the 1*s* and 2*s* orbitals are filled, a fifth electron is put into the $2p_x$ orbital, a sixth into the $2p_y$ orbital, and a seventh into the $2p_z$ orbital. Only after each 2*p* orbital has one electron is a second added to any 2*p* orbital.

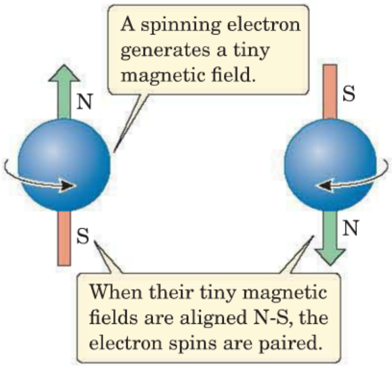


FIGURE 2.14 The pairing of electron spins.

D. Showing Electron Configurations: Orbital Box Diagrams

To illustrate how these rules are used, let us write the ground-state electron configurations for several of the elements in periods 1, 2, and 3. In the following **orbital box diagrams,** we use a box to represent an orbital, an arrow with its head up to represent a single electron, and a pair of arrows with heads in opposite directions to represent two electrons with paired spins. In addition, we show both expanded and condensed electron configurations. Table 2.7 gives the complete condensed ground-state electron configurations for elements 1 through 18.

Hydrogen (H) The atomic number of hydrogen is 1, which means that its neutral atoms have a single electron. In the ground state, this electron is placed in the 1*s* orbital. Shown first is its orbital box diagram and then its electron configuration. A hydrogen atom has one unpaired electron.

There is one electron in this orbital

H (1) [↑] Electron configuration: $1s^1$
1*s*

Helium (He) The atomic number of helium is 2, which means that its neutral atoms have two electrons. In the ground state, both electrons are placed in the 1*s* orbital with paired spins, which fill the 1*s* orbital. All electrons in helium are paired.

This orbital is now filled with two electrons

He (2) [↑↓] Electron configuration: $1s^2$
1*s*

Lithium (Li) Lithium has atomic number 3, which means that its neutral atoms have three electrons. In the ground state, two electrons are placed in the 1*s* orbital with paired spins and the third electron is placed in the 2*s* orbital. A lithium atom has one unpaired electron.

Li has one unpaired electron

Li (3) [↑↓] [↑] Electron configuration: $1s^22s^1$
1*s* 2*s*

Carbon (C) Carbon, atomic number 6, has six electrons in its neutral atoms. Two electrons are placed in the 1*s* orbital with paired spins and two are placed in the 2*s* orbital with paired spins. The fifth and sixth electrons are placed one each in the $2p_x$ and $2p_y$ orbitals. The ground state of a carbon atom has two unpaired electrons.

All orbitals of equal energy have at least one electron before any of them is filled

In a condensed electron configuration, orbitals of equal energy are grouped together

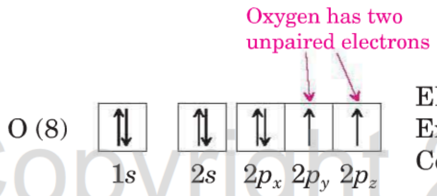
C (6) [↑↓] [↑↓] [↑][↑][ ]
1*s* 2*s* $2p_x$ $2p_y$ $2p_z$

Electron configuration
Expanded: $1s^22s^22p_x^{\,1}2p_y^{\,1}$
Condensed: $1s^22s^22p^2$

Table 2.7 Ground-State Electron Configurations of the First 18 Elements

	Orbital Box Diagram									Electron Configuration (Condensed)	Noble Gas Notation
	1s	2s	2px	2py	2pz	3s	3px	3py	3pz		
H (1)	↑									$1s^1$	
He (2)	⇅									$1s^2$	
Li (3)	⇅	↑								$1s^2\,2s^1$	[He] $2s^1$
Be (4)	⇅	⇅								$1s^2\,2s^2$	[He] $2s^2$
B (5)	⇅	⇅	↑							$1s^2\,2s^2\,2p^1$	[He] $2s^2\,2p^1$
C (6)	⇅	⇅	↑	↑						$1s^2\,2s^2\,2p^2$	[He] $2s^2\,2p^2$
N (7)	⇅	⇅	↑	↑	↑					$1s^2\,2s^2\,2p^3$	[He] $2s^2\,2p^3$
O (8)	⇅	⇅	⇅	↑	↑					$1s^2\,2s^2\,2p^4$	[He] $2s^2\,2p^4$
F (9)	⇅	⇅	⇅	⇅	↑					$1s^2\,2s^2\,2p^5$	[He] $2s^2\,2p^5$
Ne (10)	⇅	⇅	⇅	⇅	⇅					$1s^2\,2s^2\,2p^6$	[He] $2s^2\,2p^6$
Na (11)	⇅	⇅	⇅	⇅	⇅	↑				$1s^2\,2s^2\,2p^6\,3s^1$	[Ne] $3s^1$
Mg (12)	⇅	⇅	⇅	⇅	⇅	⇅				$1s^2\,2s^2\,2p^6\,3s^2$	[Ne] $3s^2$
Al (13)	⇅	⇅	⇅	⇅	⇅	⇅	↑			$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$	[Ne] $3s^2\,3p^1$
Si (14)	⇅	⇅	⇅	⇅	⇅	⇅	↑	↑		$1s^2\,2s^2\,2p^6\,3s^2\,3p^2$	[Ne] $3s^2\,3p^2$
P (15)	⇅	⇅	⇅	⇅	⇅	⇅	↑	↑	↑	$1s^2\,2s^2\,2p^6\,3s^2\,3p^3$	[Ne] $3s^2\,3p^3$
S (16)	⇅	⇅	⇅	⇅	⇅	⇅	⇅	↑	↑	$1s^2\,2s^2\,2p^6\,3s^2\,3p^4$	[Ne] $3s^2\,3p^4$
Cl (17)	⇅	⇅	⇅	⇅	⇅	⇅	⇅	⇅	↑	$1s^2\,2s^2\,2p^6\,3s^2\,3p^5$	[Ne] $3s^2\,3p^5$
Ar (18)	⇅	⇅	⇅	⇅	⇅	⇅	⇅	⇅	⇅	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$	[Ne] $3s^2\,3p^6$

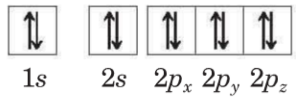
Oxygen (O) Oxygen, atomic number 8, has eight electrons in its neutral atoms. The first four electrons fill the $1s$ and $2s$ orbitals. The next three electrons are placed in the $2p_x$, $2p_y$, and $2p_z$ orbitals so that each $2p$ orbital has one electron. The remaining electron now fills the $2p_x$ orbital. The ground state of an oxygen atom has two unpaired electrons.



Electron configuration
Expanded: $1s^2 2s^2 2p_x^{\,2} 2p_y^{\,1} 2p_z^{\,1}$
Condensed: $1s^2 2s^2 2p^4$

Neon (Ne) Neon, atomic number 10, has ten electrons in its neutral atoms, which completely fill all orbitals of the first and second shells. The ground state of a neon atom has no unpaired electrons.

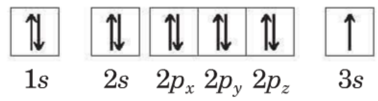
Ne (10)



Electron configuration
Expanded: $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2$
Condensed: $1s^2 2s^2 2p^6$

Sodium (Na) Sodium, atomic number 11, has 11 electrons in its neutral atoms. The first 10 fill the 1*s*, 2*s*, and 2*p* orbitals. The 11th electron is placed in the 3*s* orbital. The ground state of a sodium atom has one unpaired electron.

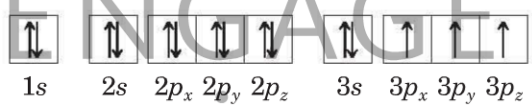
Na (11)



Electron configuration
Expanded: $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2 3s^1$
Condensed: $1s^2 2s^2 2p^6 3s^1$

Phosphorus (P) Phosphorus, atomic number 15, has 15 electrons in its neutral atoms. The first 12 fill the 1*s*, 2*s*, 2*p*, and 3*s* orbitals. Electrons 13, 14, and 15 are placed one each in the $3p_x$, $3p_y$, and $3p_z$ orbitals. The ground state of a phosphorus atom has three unpaired electrons.

P (15)



Electron configuration
Expanded: $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2 3s^2 3p_x^1 3p_y^1 3p_z^1$
Condensed: $1s^2 2s^2 2p^6 3s^2 3p^3$

E. Showing Electron Configurations: Noble Gas Notations

In an alternative way of writing ground-state electron configurations, we use the symbol of the noble gas immediately preceding the particular atom to indicate the electron configuration of all filled shells. The first shell of lithium, for example, is abbreviated [He] and the single electron in its 2*s* shell is indicated by $2s^1$. Thus, the electron configuration of a lithium atom is $[He]2s^1$ (right column of Table 2.7).

F. Showing Electron Configurations: Lewis Dot Structures

Valence electron An electron in the outermost occupied (valence) shell of an atom

Valence shell The outermost occupied shell of an atom

When discussing the physical and chemical properties of an element, chemists often focus on the outermost shell of its electrons because electrons in this shell are the ones involved in the formation of chemical bonds (Chapter 3) and in chemical reactions (Chapter 4). Outer-shell electrons are called **valence electrons,** and the energy level in which they are found is called the **valence shell.** Carbon, for example, with a ground-state electron configuration of $1s^2 2s^2 2p^2$, has four valence (outer-shell) electrons.

To show the outermost electrons of an atom, we commonly use a representation called a **Lewis dot structure,** named after the American chemist Gilbert N. Lewis (1875–1946), who devised this notation. A Lewis structure shows the symbol of the element surrounded by a number of dots equal to the number of electrons in the outer (valence) shell of an atom of that element. In a Lewis structure, the atomic symbol represents the nucleus and all filled inner shells. Table 2.8 shows Lewis structures for the first 18 elements of the Periodic Table.

Lewis dot structure The symbol of the element surrounded by a number of dots equal to the number of electrons in the valence shell of an atom of that element

Table 2.8 Lewis Dot Structures for Elements 1–18 of the Periodic Table

1A	2A	3A	4A	5A	6A	7A	8A
H·							He:
Li·	Be:	$\dot{\mathrm{B}}$:	·$\dot{\mathrm{C}}$:	·$\dot{\mathrm{N}}$:	:$\dot{\mathrm{O}}$:	:$\ddot{\mathrm{F}}$:	:$\ddot{\mathrm{Ne}}$:
Na·	Mg:	$\dot{\mathrm{Al}}$:	·$\dot{\mathrm{Si}}$:	·$\dot{\mathrm{P}}$:	:$\dot{\mathrm{S}}$:	:$\ddot{\mathrm{Cl}}$:	:$\ddot{\mathrm{Ar}}$:

Each dot represents one valence electron.

The noble gases helium and neon have filled valence shells. The valence shell of helium is filled with two electrons ($1s^2$); that of neon is filled with eight electrons ($2s^2 2p^6$). Neon and argon have in common an electron configuration in which the *s* and *p* orbitals of their valence shells are filled with eight electrons. The valence shells of all other elements shown in Table 2.8 contain fewer than eight electrons.

At this point, let us compare the Lewis structures given in Table 2.8 with the ground-state electron configurations given in Table 2.7. The Lewis structure of boron (B), for example, is shown in Table 2.8 with three valence electrons; these are the paired $2s$ electrons and the single $2p_x$ electron shown in Table 2.7. The Lewis structure of carbon (C) is shown in Table 2.8 with four valence electrons; these are the two paired $2s$ electrons and the unpaired $2p_x$ and $2p_y$ electrons shown in Table 2.7.

Example 2.7 Electron Configuration

The Lewis dot structure for nitrogen shows five valence electrons. Write the expanded electron configuration for nitrogen and show to which orbitals its five valence electrons are assigned.

Strategy

Locate nitrogen in the Periodic Table and determine its atomic number. In an electrically neutral atom, the number of negatively charged extranuclear electrons is the same as the number of positively charged protons in its nucleus. The order of filling of orbitals is $1s$ $2s$ $2p_x$ $2p_y$ $2p_z$ $3s$, etc.

Solution

Nitrogen, atomic number 7, has the following ground-state electron configuration:

$$1s^2\,2s^2\,2p_x^1\,2p_y^1\,2p_z^1$$

The five valence electrons of the Lewis dot structure are the two paired electrons in the $2s$ orbital and the three unpaired electrons in the $2p_x$, $2p_y$, and $2p_z$ orbitals.

Problem 2.7

Write the Lewis dot structure for the element that has the following ground-state electron configuration. What is the name of this element?

$$1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2\ 3s^2\ 3p_x^1$$

2.7 How Are Electron Configuration and Position in the Periodic Table Related?

When Mendeleyev published his first Periodic Table in 1869, he could not explain why it worked—that is, why elements with similar properties became aligned in the same column. Indeed, no one had a good explanation for this phenomenon. It was not until the discovery of electron configurations that chemists finally understood why the Periodic Table works. The answer, they discovered, is very simple: Elements in the same column have the same ground-state electron configuration in their outer shells. Figure 2.15 shows the relationship between shells (principal energy levels) and orbitals being filled.

All main-group elements (those in A columns) have in common the fact that either their *s* or *p* orbitals are being filled. Notice that the 1*s* shell is filled with two electrons; there are only two elements in the first period. The 2*s* and 2*p* orbitals are filled with eight electrons; there are eight elements in period 2. Similarly, the 3*s* and 3*p* orbitals are filled with eight electrons; there are eight elements in period 3.

To create the elements of period 4, one 4*s*, three 4*p*, and five 3*d* orbitals are available. These orbitals can hold a total of 18 electrons; there are 18 elements in period 4. Similarly, there are 18 elements in period 5. Inner transition elements are created by filling *f* orbitals, which come in sets of seven and can hold a total of 14 electrons; there are 14 inner transition elements in the lanthanide series and 14 in the actinide series.

To see the similarities in electron configurations within the Periodic Table, let us look at the elements in column 1A. We already know the configurations for lithium, sodium, and potassium (Table 2.7). To this list we can add rubidium and cesium. All elements in column 1A have one electron in their valence shell (Table 2.9).

All Group 1A elements are metals, with the exception of hydrogen, which is a nonmetal. The properties of elements largely depend on the electron configuration of their outer shell. As a consequence, it is not surprising that

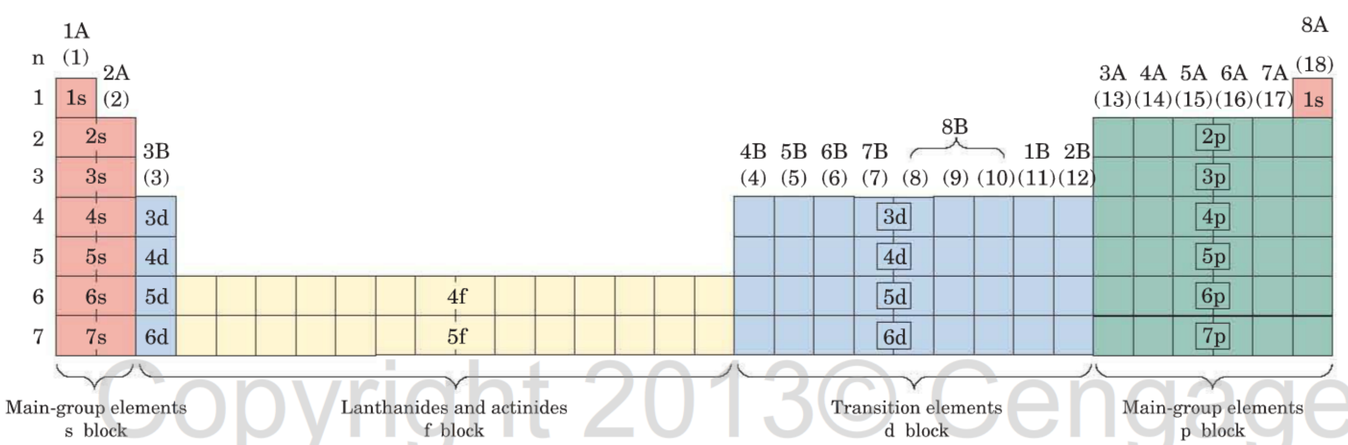
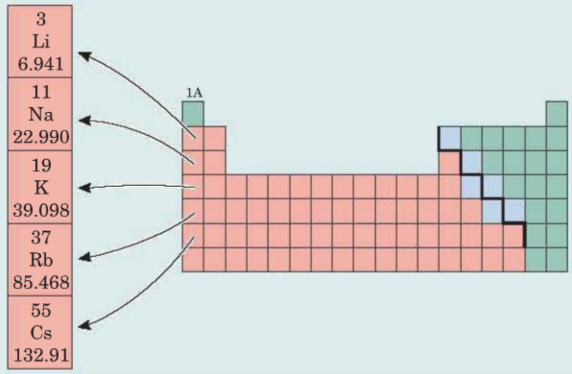


FIGURE 2.15 Electron configuration and the Periodic Table.

Table 2.9 Noble Gas Notation and Lewis Dot Structures for the Alkali Metals (Group 1A Elements)

Noble Gas Notation	Lewis Dot Structure
$[He]2s^1$	Li•
$[Ne]3s^1$	Na•
$[Ar]4s^1$	K•
$[Kr]5s^1$	Rb•
$[Xe]6s^1$	Cs•



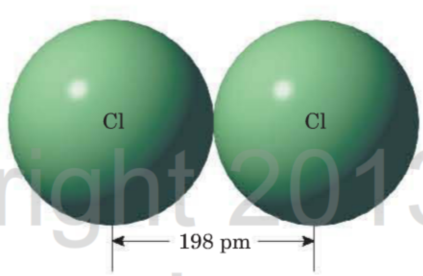
Group 1A elements, all of which have the same outer-shell configuration, are metals (except for hydrogen) and have such similar physical and chemical properties.

2.8 What Is a Periodic Property?

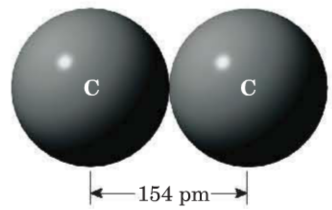
As we have now seen, the Periodic Table originally was constructed on the basis of trends (periodicity) in physical and chemical properties. With an understanding of electron configurations, chemists realized that the periodicity in chemical properties could be understood in terms of the periodicity in ground-state electron configuration. As we noted in the opening of Section 2.7, The Periodic Table works because "elements in the same column have the same ground-state electron configuration in their outer shells." Thus, chemists could now explain why certain chemical and physical properties of elements changed in predictable ways in going down a column or going across a row of the Periodic Table, In this section, we will examine the periodicity of one physical property (atomic size) and one chemical property (ionization energy) to illustrate how periodicity is related to position in the Periodic Table.

A. Atomic Size

The size of an atom is determined by the size of its outermost occupied orbital. The size of a sodium atom, for example, is the size of its singly occupied 3*s* orbital. The size of a chlorine atom is determined by the size of its three 3*p* orbitals ($3s^23p^5$). The simplest way to determine the size of an atom is to determine the distance between bonded nuclei in a sample of the element. A chlorine molecule, for example, has a diameter of 198 pm (pm = picometer; 1 pm = 10^{-12} meter). The radius of a chlorine atom is thus 99 pm, which is one-half of the distance between two bonded chlorine nuclei in Cl_2.



Similarly, the distance between bonded carbon nuclei in diamond is 154 pm, and so the radius of a carbon atom is 77 pm.



From measurements such as these, we can assemble a set of atomic radii (Figure 2.16).

From the information in this figure, we can see that for main group elements, (1) atomic radii increase going down a group and (2) decrease going from left to right across a period. Let us examine the correlation between each of these trends and ground-state electron configuration.

1. The size of an atom is determined by the size of its outermost electrons. In going down a column, the outermost electrons are assigned to higher and higher principal energy levels. The electrons of lower principal energy levels (those lying below the valence shell) must occupy some space, so the outer-shell electrons must be farther and farther from the nucleus, which rationalizes the increase in size in going down a column.

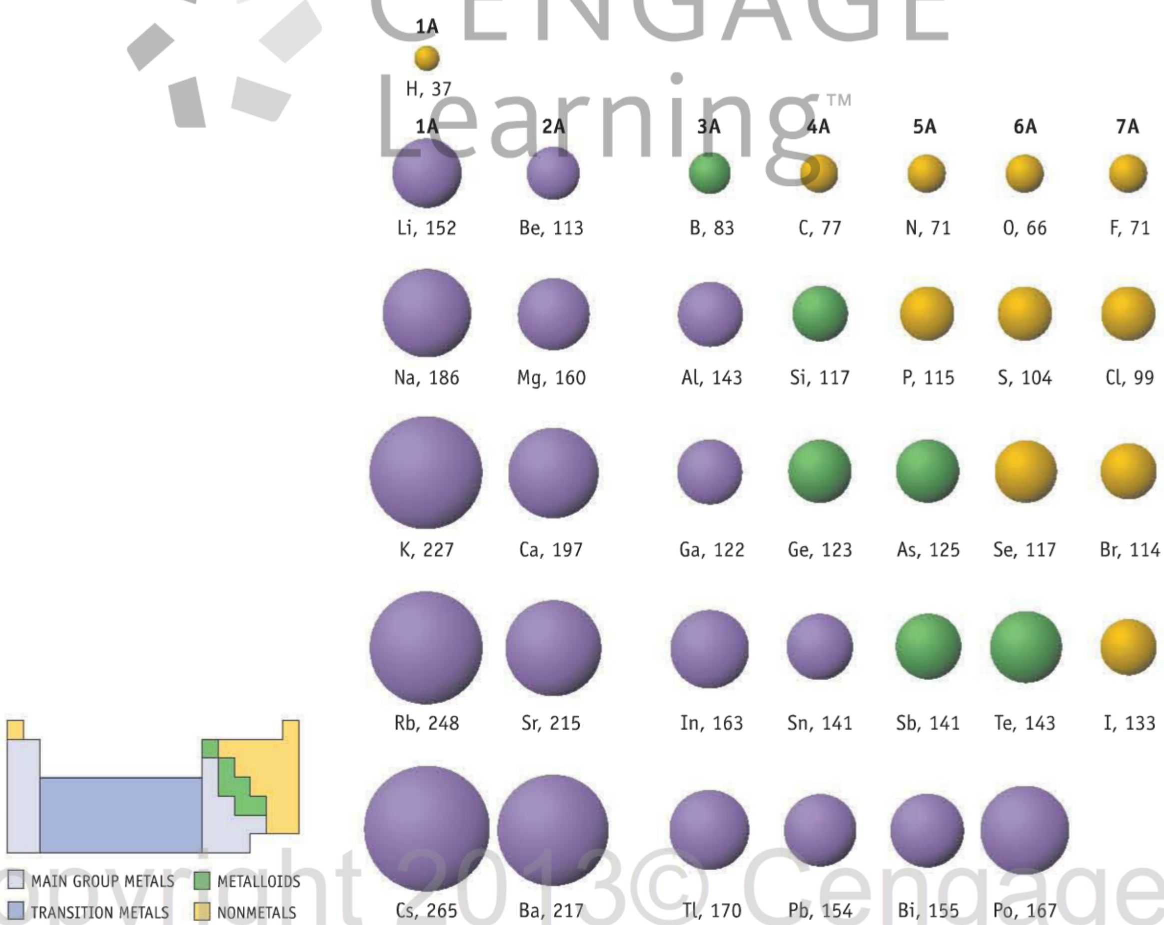


FIGURE 2.16 Atomic radii of main-group elements (in picometers, 1 pm = 10^{-12} m).

2. For elements in the same period, the principal energy level remains the same (for example, the valence electrons of all second-period elements occupy the second principal energy level). But in going from one element to the next across a period, one more proton is added to the nucleus, thus increasing the nuclear charge by one unit for each step from left to right. The result is that the nucleus exerts an increasingly stronger pull on the valence electrons and atomic radius decreases.

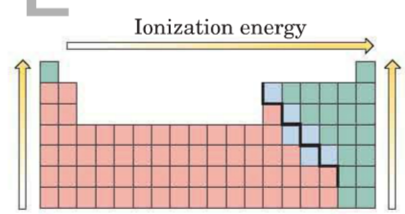
B. Ionization energy

Atoms are electrically neutral—the number of electrons outside the nucleus of an atom is equal to the number of protons inside the nucleus. Atoms do not normally lose or gain protons or neutrons, but they can lose or gain electrons. When a lithium atom, for example, loses one electron, it becomes a lithium **ion.** A lithium atom has three protons in its nucleus and three electrons outside the nucleus. When a lithium atom loses one of these electrons, it still has three protons in its nucleus (and, therefore, is still lithium), but now it has only two electrons outside the nucleus. The two remaining electrons cancel the charge of two of the protons, but there is no third electron to cancel the charge of the third proton. Therefore, a lithium ion has a charge of +1 and we write it as Li^+. The ionization energy for a lithium ion in the gas phase is 0.52 kJ/mol.

Ion An atom with an unequal number of protons and electrons

$$\underset{\text{Lithium atom}}{Li} + \underset{\text{Ionization energy}}{\text{energy}} \longrightarrow \underset{\text{Lithium ion}}{Li^+} + \underset{\text{Electron}}{e^-}$$

Ionization energy The energy required to remove the most loosely held electron from an atom in the gas phase



Ionization energy is a measure of how difficult it is to remove the most loosely held electron from an atom in the gaseous state. The more difficult it is to remove the electron, the higher the ionization energy. Ionization energies are always positive because energy must be supplied to overcome the attractive force between the electron and the positively charged nucleus. Figure 2.17 shows the ionization energies for the atoms of main-group elements 1 through 37 (hydrogen through rubidium).

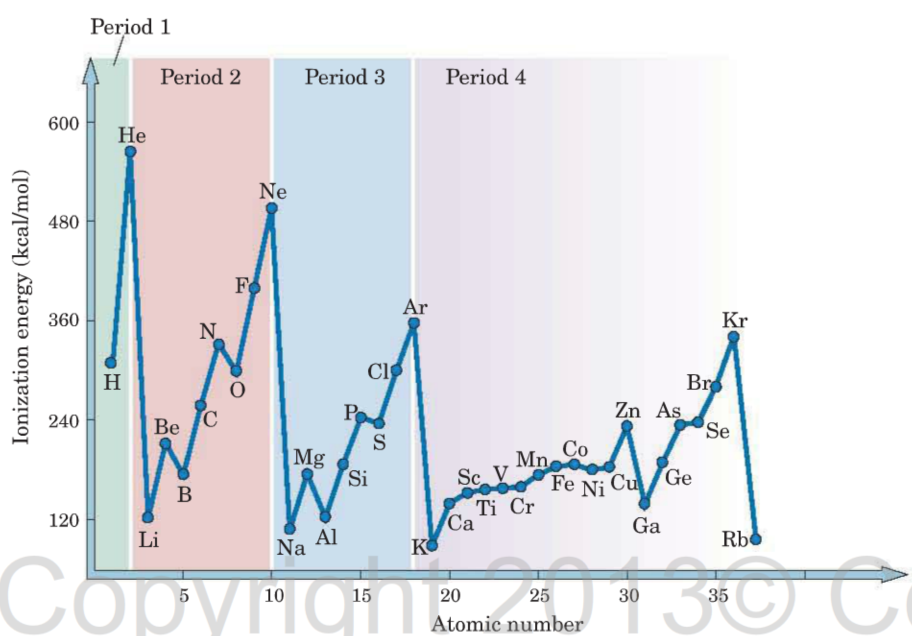


FIGURE 2.17 Ionization energy versus atomic number for elements 1–37.

As we see in Figure 2.17, ionization energy generally increases as we go up a column of the Periodic Table and, with a few exceptions, generally increases as we go from left to right across a row. For example, within the Group 1A metals, rubidium gives up its 5*s* electron most easily and lithium gives up its 2*s* electron least easily.

We explain this trend by saying that the 5*s* electron of rubidium is farther from the positively charged nucleus than is the 4*s* electron in potassium, which in turn is farther from the positively charged nucleus than is the 3*s* electron of sodium, and so forth. Furthermore, the 5*s* electron of rubidium is more "shielded" by inner-shell electrons from the attractive force of the positive nucleus than is the 4*s* electron of potassium, and so forth. The greater the shielding, the lower the ionization energy. Thus, going down a column of the Periodic Table, the shielding of an atom's outermost electrons increases and the element's ionization energies decrease.

We explain the increase in ionization energy across a row by the fact that the valence electrons across a row are in the same shell (principal energy level). Because the number of protons in the nucleus increases regularly across a row, the valence electrons experience an increasingly stronger pull by the nucleus, which makes them more difficult to remove. Thus, ionization energy increases from left to right across a row of the Periodic Table.

Summary

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Section 2.1 What Is Matter Made Of?

- The Greek philosopher Democritus (circa 460–370 BCE) was the first person to propose an atomic theory of matter. He stated that all matter is made of very tiny particles, which he called atoms.

Section 2.2 How Do We Classify Matter?

- We classify matter as **elements, compounds,** or **mixtures.**

Section 2.3 What Are the Postulates of Dalton's Atomic Theory?

- (1) All matter is made up of atoms; (2) all atoms of a given element are identical, and the atoms of any one element are different from those of any other element; (3) compounds are formed by the chemical combination of atoms; and (4) a molecule is a cluster of two or more atoms that acts as a single unit.
- Dalton's theory is based on the **law of conservation of mass** (matter can be neither created nor destroyed) and the **law of constant composition** (any compound is always made up of elements in the same proportion by mass).

Section 2.4 What Are Atoms Made Of?

- Atoms consist of protons and neutrons found inside the nucleus and electrons located outside it. An **electron** has a mass of approximately 0.0005 amu and a charge of -1. A **proton** has a mass of approximately 1 amu and a charge of $+1$. A **neutron** has a mass of approximately 1 amu and no charge.
- The **mass number** of an atom is the sum of the number of its protons and neutrons.
- The **atomic number** of an element is the number of protons in the nucleus of an atom of that element.
- **Isotopes** are atoms with the same atomic number but different mass numbers; that is, they have the same number of protons but different numbers of neutrons in their nuclei.
- The **atomic weight** of an element is a weighted average of the masses (in amu) of its isotopes as they occur in nature.
- Atoms are very tiny, with a very small mass, almost all of which is concentrated in the nucleus. The nucleus is tiny, with an extremely high density.

Section 2.5 What Is the Periodic Table?

- The **Periodic Table** is an arrangement of elements with similar chemical properties into columns; the properties gradually change as we move down a column.
- **Metals** are solids (except for mercury, which is a liquid), shiny, conductors of electricity, ductile, malleable, and form alloys, which are solutions of one or more metals dissolved in another metal. In their chemical reactions, metals tend to give up electrons.
- With the exception of hydrogen, the **nonmetals** appear on the right side of the Periodic Table. With the exception of graphite, they do not conduct electricity. In their chemical reactions, nonmetals tend to accept electrons.

- Six elements are classified as **metalloids:** boron, silicon, germanium, arsenic, antimony, and tellurium. These elements have some properties of metals and some properties of nonmetals.

Section 2.6 How Are the Electrons in an Atom Arranged?

- Electrons in atoms exist in **principal energy levels** or **shells.**
- All principal energy levels except the first are divided into **subshells** designated by the letters *s, p, d,* and *f*. Within each subshell, electrons are grouped into **orbitals.** An orbital is a region of space that can hold two electrons with paired spins. All *s* orbitals are spherical and can hold two electrons. All *p* orbitals come in sets of three, and each is shaped like a dumbbell, with the nucleus at the center of the dumbbell. A set of three *p* orbitals can hold six electrons. A set of five *d* orbitals can hold ten electrons, and a set of 7 *f* orbitals can hold fourteen electrons.
- Electrons are arranged in orbitals according to the following rules.
  (1) Orbitals fill in order of increasing energy; (2) each orbital can hold a maximum of two electrons with paired spins; (3) when filling orbitals of equivalent energy, each orbital adds one electron before any orbital adds a second electron.
- The electron configuration of an atom may be shown by an orbital notation, an orbital box diagram, or a noble gas notation.
- Electrons in the outermost or **valence shell** of an atom are called **valence electrons.** In a **Lewis dot structure** of an atom, the symbol of the element is surrounded by a number of dots equal to the number of its valence electrons.

Section 2.7 How Are Electron Configuration and Position in the Periodic Table Related?

- The Periodic Table works because elements in the same column have the same outer-shell electron configuration.

Section 2.8 What Is a Periodic Property?

- **Ionization energy** is the energy necessary to remove the most loosely held electron from an atom in the gas phase to form an **ion.** Ionization energy increases from bottom to top within a column of the Periodic Table because the valence shell of the atom becomes closer to the positively charged nucleus. It increases from left to right within a row because the positive charge on the nucleus increases in this direction.
- The **size of an atom (atomic radius)** is determined by the size of its outermost occupied orbital. Atomic size is a periodic property. For main-group elements, atomic size increases going down a group and decreases going from left to right across a period. In going down a column, the outermost electrons are assigned to higher and higher principal energy levels. For elements in the same period, the principal energy level remains the same from one element to the next, but the nuclear charge increases by one unit (by one proton). As a result of this increase in nuclear charge across a period, the nucleus exerts an increasingly stronger pull on the valence electrons and atomic size decreases.

Problems

OWL Interactive versions of these problems may be assigned in OWL.

Orange-numbered problems are applied.

Section 2.1 What Are Atoms Made Of?

2.8 In what way(s) was Democritus's atomic theory similar to that of Dalton's atomic theory?

Section 2.2 How Do We Classify Matter?

2.9 Answer true or false.

(a) Matter is divided into elements and pure substances.

(b) Matter is anything that has mass and volume (occupies space).

(c) A mixture is composed of two or more pure substances.

(d) An element is a pure substance.

(e) A heterogeneous mixture can be separated into pure substances, but a homogeneous mixture cannot.

(f) A compound consists of elements combined in a fixed ratio.

(g) A compound is a pure substance.

(h) All matter has mass.

(i) All of the 118 known elements occur naturally on Earth.

(j) The first six elements in the Periodic Table are the most important for human life.

(k) The combining ratio of a compound tells you how many atoms of each element are combined in the compound.

(l) The combining ratio of 1:2 in the compound CO_2 tells you that this compound is formed by the combination of one gram of carbon is with two grams of oxygen.

2.10 Classify each of the following as an element, a compound, or a mixture:

(a) Oxygen (b) Table salt
(c) Sea water (d) Wine
(e) Air (f) Silver
(g) Diamond (h) A pebble
(i) Gasoline (j) Milk
(k) Carbon dioxide (l) Bronze

2.11 Name these elements (try not to look at a Periodic Table):

(a) O (b) Pb (c) Ca (d) Na
(e) C (f) Ti (g) S (h) Fe
(i) H (j) K (k) Ag (l) Au

2.12 The elements game, Part I. Name and give the symbol of the element that is named for each person.

(a) Niels Bohr (1885–1962), Nobel Prize for physics in 1922
(b) Pierre and Marie Curie, Nobel Prize for chemistry in 1903
(c) Albert Einstein (1879–1955), Nobel Prize for physics in 1921
(d) Enrico Fermi (1901–1954), Nobel Prize for physics in 1938
(e) Ernest Lawrence (1901–1958), Nobel Prize for physics in 1939
(f) Lise Meitner (1868–1968), codiscoverer of nuclear fission
(g) Dmitri Mendeleyev (1834–1907), first person to formulate a workable Periodic Table
(h) Alfred Nobel (1833–1896), discoverer of dynamite
(i) Ernest Rutherford (1871–1937), Nobel Prize for chemistry in 1908
(j) Glen Seaborg (1912–1999), Nobel Prize for chemistry in 1951

2.13 The elements game, Part II. Name and give the symbol of the element that is named for each geographic location.

(a) The Americas
(b) Berkeley, California
(c) The state and University of California
(d) Dubna, location in Russia of the Joint Institute of Nuclear Research
(e) Europe
(f) France
(g) Gallia, the Latin name for ancient France
(h) Germany
(i) Hafnia, the Latin name for ancient Copenhagen
(j) Hesse, a German state
(k) Holmia, the Latin name for ancient Stockholm
(l) Lutetia, the Latin name for ancient Paris
(m) Magnesia, a district in Thessaly
(n) Poland, the native country of Marie Curie
(o) Rhenus, the Latin name for the river Rhine
(p) Ruthenia, the Latin name for ancient Russia
(q) Scandia, the Latin name for ancient Scandinavia
(r) Strontian, a town in Scotland
(s) Ytterby, a village in Sweden (three elements)
(t) Thule, the earliest name for Scandinavia

2.14 The elements game, Part III. Give the names and symbols for the two elements named for planets. Note that the element plutonium was named for Pluto, which is no longer classified as a planet.

2.15 Write the formulas of compounds in which the combining ratios are as follows:

(a) Potassium: oxygen, 2:1
(b) Sodium: phosphorus: oxygen, 3:1:4
(c) Lithium: nitrogen: oxygen, 1:1:3

2.16 Write the formulas of compounds in which the combining ratios are as follows:

(a) Sodium: hydrogen: carbon: oxygen, 1:1:1:3
(b) Carbon: hydrogen: oxygen, 2:6:1
(c) Potassium: manganese: oxygen, 1:1:4

Section 2.3 What Are Postulates of Dalton's Atomic Theory?

2.17 How does Dalton's atomic theory explain:

(a) the law of conservation of mass?
(b) the law of constant composition?

2.18 When 2.16 g of mercuric oxide is heated, it decomposes to yield 2.00 g of mercury and 0.16 g of oxygen. Which law is supported by this experiment?

2.19 The compound carbon monoxide contains 42.9% carbon and 57.1% oxygen. The compound carbon dioxide contains 27.3% carbon and 72.7% oxygen. Does this disprove Proust's law of constant composition?

2.20 Calculate the percentage of hydrogen and oxygen in water, H_2O, and hydrogen peroxide, H_2O_2.

Section 2.4 What Are Atoms Made Of?

2.21 Answer true or false.

(a) A proton and an electron have the same mass but opposite charges.
(b) The mass of an electron is considerably smaller than that of a neutron.
(c) An atomic mass unit (amu) is a unit of mass.
(d) One amu is equal to 1 gram.
(e) The protons and neutrons of an atom are found in the nucleus.
(f) The electrons of an atom are found in the space surrounding the nucleus.
(g) All atoms of the same element have the same number of protons.
(h) All atoms of the same element have the same number of electrons.
(i) Electrons and protons repel each other.
(j) The size of an atom is approximately the size of its nucleus.
(k) The mass number of an atom is the sum of the numbers of protons and neutrons in the nucleus of that atom.
(l) For most atoms, their mass number is the same as their atomic number.
(m) The three isotopes of hydrogen (hydrogen-1, hydrogen-2, and hydrogen-3) differ only in the number of neutrons in the nucleus.
(n) Hydrogen-1 has one neutron in its nucleus, hydrogen-2 has two neutrons in its nucleus, and hydrogen-3 has three neutrons.

(o) All isotopes of an element have the same number of electrons.

(p) Most elements found on Earth are mixtures of isotopes.

(q) The atomic weight of an element given in the Periodic Table is the weighted average of the masses of its isotopes found on Earth.

(r) The atomic weights of most elements are whole numbers.

(s) Most of the mass of an atom is found in its nucleus.

(t) The density of a nucleus is its mass number expressed in grams.

2.22 Where in an atom are these subatomic particles located?

(a) Protons (b) Electrons (c) Neutrons

2.23 It has been said, "The number of protons determines the identity of the element." Do you agree or disagree with this statement? Explain.

2.24 What is the mass number of an atom with:

(a) 22 protons, 22 electrons, and 26 neutrons?

(b) 76 protons, 76 electrons, and 114 neutrons?

(c) 34 protons, 34 electrons, and 45 neutrons?

(d) 94 protons, 94 electrons, and 150 neutrons?

2.25 Name and give the symbol for each element in Problem 2.24.

2.26 Given these mass numbers and number of neutrons, what is the name and symbol of each element?

(a) Mass number 45; 24 neutrons

(b) Mass number 48; 26 neutrons

(c) Mass number 107; 60 neutrons

(d) Mass number 246; 156 neutrons

(e) Mass number 36; 18 neutrons

2.27 If each atom in Problem 2.26 acquired two more neutrons, what element would each then be?

2.28 How many neutrons are in:

(a) a carbon atom of mass number 13?

(b) a germanium atom of mass number 73?

(c) an osmium atom of mass number 188?

(d) a platinum atom of mass number 195?

2.29 How many protons and how many neutrons does each of these isotopes of radon contain?

(a) Rn-210 (b) Rn-218 (c) Rn-222

2.30 How many neutrons and protons are in each isotope?

(a) ^{22}Ne (b) ^{104}Pd

(c) ^{35}Cl (d) Tellurium-128

(e) Lithium-7 (f) Uranium-238

2.31 Tin-118 is one of the isotopes of tin. Name the isotopes of tin that contain two, three, and six more neutrons than tin-118.

2.32 What is the difference between atomic number and mass number?

2.33 Define:

(a) Ion (b) Isotope

2.34 There are only two naturally occurring isotopes of antimony: ^{121}Sb (120.90 amu) and ^{123}Sb (122.90 amu). The atomic weight of antimony given in the Periodic Table is 121.75. Which of the two isotopes has the greater natural abundance?

2.35 The two most abundant naturally occurring isotopes of carbon are carbon-12 (98.90%, 12.000 amu) and carbon-13 (1.10%, 13.003 amu). From these abundances, calculate the atomic weight of carbon and compare your calculated value with that given in the Periodic Table.

2.36 Another isotope of carbon, carbon-14, occurs in nature but in such small amounts relative to carbon-12 and carbon-13 that it does not contribute to the atomic weight of carbon as recorded in the Periodic Table. Carbon-14 is invaluable in the science of radiocarbon dating (see Chemical Connections 9A). Give the number of protons, neutrons, and electrons in an atom of carbon-14.

2.37 The isotope carbon-11 does not occur in nature but has been made in the laboratory. This isotope is used in a medical imaging technique called positron emission tomography (PET, see Section 9.7A). Give the number of protons, neutrons, and electrons in an atom of carbon-11.

2.38 Other isotopes used in PET imaging are fluorine-18, nitrogen-13, and oxygen-15. None of these isotopes occurs in nature; all must be produced in the laboratory. Give the number of protons, neutrons, and electrons in an atom of each of these artificial isotopes.

2.39 Americium-241 is used in household smoke detectors. This element has 11 known isotopes, none of which occurs in nature, but must be made in the laboratory. Give the number of protons, neutrons, and electrons in an atom of americium-241.

2.40 In dating geological samples, scientists compare the ratio of rubidium-87 to strontium-87. Give the number of protons, neutrons, and electrons in an atom of each element.

Section 2.5 What Is the Periodic Table?

2.41 Answer true or false.

(a) Mendeleyev discovered that when elements are arranged in order of increasing atomic weight, certain sets of properties recur periodically.

(b) Main-group elements are those in the columns 3A to 8A of the Periodic Table.

(c) Nonmetals are found at the top of the Periodic Table, metalloids in the middle, and metals at the bottom.

(d) Among the 118 known elements, there are approximately equal numbers of metals and nonmetals.

(e) A horizontal row in the Periodic Table is called a group.

(f) The Group 1A elements are called the "alkali metals."

(g) The alkali metals react with water to give hydrogen gas and a metal hydroxide, MOH, where "M" is the metal.

(h) The halogens are Group 7A elements.

(i) The boiling points of noble gases (Group 8A elements) increase going from top to bottom of the column.

2.42 How many metals, metalloids, and nonmetals are there in the third period of the Periodic Table?

2.43 Which group(s) of the Periodic Table contain(s):

(a) Only metals? (b) Only metalloids?

(c) Only nonmetals?

2.44 Which period(s) in the Periodic Table contain(s) more nonmetals than metals? Which contain(s) more metals than nonmetals?

2.45 Group the following elements according to similar properties (look at the Periodic Table): As, I, Ne, F, Mg, K, Ca, Ba, Li, He, N, P.

2.46 Which are transition elements?

(a) Pd (b) K (c) Co

(d) Ce (e) Br (f) Cr

2.47 Which element in each pair is more metallic?

(a) Silicon or aluminum (b) Arsenic or phosphorus

(c) Gallium or germanium (d) Gallium or aluminum

2.48 Classify these elements as metals, nonmetals, or metalloids:

(a) Argon (b) Boron (c) Lead

(d) Arsenic (e) Potassium (f) Silicon

(g) Iodine (h) Antimony (i) Vanadium

(j) Sulfur (k) Nitrogen

Section 2.6 How Are the Electrons in an Atom Arranged?

2.49 Answer true or false.

(a) To say that "energy is quantized" means that only certain energy values are allowed.

(b) Bohr discovered that the energy of an electron in an atom is quantized.

(c) Electrons in atoms are confined to regions of space called "principal energy levels."

(d) Each principal energy level can hold a maximum of two electrons.

(e) An electron in a 1*s* orbital is held closer to the nucleus than an electron in a 2*s* orbital.

(f) An electron in a 2*s* orbital is harder to remove from an atom than an electron in a 1*s* orbital.

(g) An *s* orbital has the shape of a sphere, with the nucleus at the center of the sphere.

(h) Each 2*p* orbital has the shape of a dumbbell, with the nucleus at the midpoint of the dumbbell.

(i) The three 2*p* orbitals in an atom are aligned parallel to each other.

(j) An orbital is a region of space that can hold two electrons.

(k) The second shell contains one *s* orbital and three *p* orbitals.

(l) In the ground-state electron configuration of an atom, only the lowest-energy orbitals are occupied.

(m) A spinning electron behaves as a tiny bar magnet, with a North Pole and a South Pole.

(n) An orbital can hold a maximum of two electrons with their spins paired.

(o) Paired electron spins means that the two electrons are aligned with their spins North Pole to North Pole and South Pole to South Pole.

(p) An orbital box diagram puts all of the electrons of an atom in one box with their spins aligned.

(q) An orbital box diagram of a carbon atom shows two unpaired electrons.

(r) A Lewis dot structure shows only the electrons in the valence shell of an atom of the element.

(s) A characteristic of Group 1A elements is that each has one unpaired electron in its outermost occupied (valence) shell.

(t) A characteristic of Group 6A elements is that each has six unpaired electrons in its valence shell.

2.50 How many periods of the Periodic Table have two elements? How many have eight elements? How many have 18 elements? How many have 32 elements?

2.51 What is the correlation between the group number of the main-group elements (those in the A columns of the Mendeleyev system) and the number of valence electrons in an element in the group?

2.52 Given your answer to Problem 2.51, write the Lewis dot structure for each of the following elements using no information other than the number of the group in the Periodic Table to which the element belongs.

(a) Carbon (4A) (b) Silicon (4A)

(c) Oxygen (6A) (d) Sulfur (6A)

(e) Aluminum (3A) (f) Bromine (7A)

2.53 Write the condensed ground-state electron configuration for each of the following elements. The element's atomic number is given in parentheses.

(a) Li (3) (b) Ne (10) (c) Be (4)

(d) C (6) (e) Mg (12)

2.54 Write the Lewis dot structure for each element in Problem 2.53.

2.55 Write the condensed ground-state electron configuration for each of the following elements. The element's atomic number is given in parentheses.

(a) He (2) (b) Na (11) (c) Cl (17)

(d) P (15) (e) H (1)

2.56 Write the Lewis dot structure for each element in Problem 2.55.

2.57 What is the same and what is different in the electron configurations of:

(a) Na and Cs? (b) O and Te? (c) C and Ge?

2.58 Silicon, atomic number 14, is in Group 4A. How many orbitals are occupied by the valence electrons of Si in its ground state?

2.59 You are presented with a Lewis dot structure of element X as X:. To which two groups in the Periodic Table might this element belong?

2.60 The electron configurations for the elements with atomic numbers higher than 36 follow the same rules as given in the text for the first 36 elements. In fact, you can arrive at the correct order of filling of orbitals from Figure 2.15 by starting with H and reading the orbitals from left to right across the first row, then the second row, and so on. Write the condensed ground-state electron configuration for:

(a) Rb (b) Sr (c) Br

Section 2.7 How Are Electron Configuration and Position in the Periodic Table Related?

2.61 Answer true or false.

(a) Elements in the same column of the Periodic Table have the same outer-shell electron configuration.

(b) All Group 1A elements have one electron in their valence shell.

(c) All Group 6A elements have six electrons in their valence shell.

(d) All Group 8A elements have eight electrons in their valence shell.

(e) Period 1 of the Periodic Table has one element, period 2 has two elements, period 3 has three elements, and so forth.

(f) Period 2 results from filling the 2*s* and 2*p* orbitals, and therefore, there are eight elements in period 2.

(g) Period 3 results from filling the 3*s*, 3*p*, and 3*d* orbitals, and therefore, there are nine elements in period 3.

(h) The main-group elements are *s* block and *p* block elements.

2.62 Why do the elements in column 1A of the Periodic Table (the alkali metals) have similar but not identical properties?

Section 2.8 What Is a Periodic Property?

2.63 Answer true or false.

(a) Ionization energy is the energy required to remove the most loosely held electron from an atom in the gas phase.

(b) When an atom loses an electron, it becomes a positively charged ion.

(c) Ionization energy is a periodic property because ground-state electron configuration is a periodic property.

(d) Ionization energy generally increases going from left to right across a period of the Periodic Table.

(e) Ionization energy generally increases in going from top to bottom within a column in the Periodic Table.

(f) The sign of an ionization energy is always positive; the process is always endothermic.

2.64 Consider the elements B, C, and N. Using only the Periodic Table, predict which of these three elements has:

(a) the largest atomic radius.

(b) the smallest atomic radius.

(c) the largest ionization energy.

(d) the smallest ionization energy.

2.65 Account for the following observations.

(a) The atomic radius of an anion is always larger than that of the atom from which it is derived. Examples: Cl 99 pm and Cl^- 181 pm; O 73 pm and O^{2-} 140 pm.

(b) The atomic radius of a cation is always smaller than that of the atom from which it is derived. Examples: Li 152 pm and Li^+ 76 pm; Na 156 pm and Na^+ 98 pm.

2.66 Using only the Periodic Table, arrange the elements in each set in order of increasing ionization energy:

(a) Li, Na, K (b) C, N, Ne

(c) O, C, F (d) Br, Cl, F

2.67 Account for the fact that the first ionization energy of oxygen is less than that of nitrogen.

2.68 Every atom except hydrogen has a series of ionization energies (IE) because they have more than one electron that can be removed. Following are the first three ionization energies for magnesium:

$$Mg(g) \longrightarrow Mg^{+}(g) + e^{-}(g) \quad IE_1 = 738 \text{ kJ/mol}$$
$$Mg^{+}(g) \longrightarrow Mg^{2+}(g) + e^{-}(g) \quad IE_2 = 1450 \text{ kJ/mol}$$
$$Mg^{2+}(g) \longrightarrow Mg^{3+}(g) + e^{-}(g) \quad IE_3 = 7734 \text{ kJ/mol}$$

(a) Write the ground-state electron configuration for Mg, Mg^+, Mg^{2+}, and Mg^{3+}.

(b) Account for the large increase in ionization energy for the removal of the third electron compared with the ionization energies for removal of the first and second electrons.

Chemical Connections

2.69 (Chemical Connections 2A) Why does the body need sulfur, calcium, and iron?

2.70 (Chemical Connections 2B) Which are the two most abundant elements, by weight, in:

(a) the Earth's crust? (b) the human body?

2.71 (Chemical Connections 2C) Why is strontium-90 more dangerous to humans than most other radioactive isotopes that were present in the Chernobyl fallout?

2.72 (Chemical Connections 2D) Bronze is an alloy of which two metals?

2.73 (Chemical Connections 2D) Copper is a soft metal. How can it be made harder?

Additional Problems

2.74 Give the designations of all subshells in the:
(a) 1 shell (b) 2 shell
(c) 3 shell (d) 4 shell

2.75 Tell whether metals or nonmetals are more likely to have each of the following characteristics:
(a) Conduct electricity and heat
(b) Accept electrons
(c) Be malleable
(d) Be a gas at room temperature
(e) Be a transition element
(f) Lose electrons

2.76 Explain why:
(a) atomic radius decreases going across a period in the Periodic Table.
(b) energy is required to remove an electron from an atom.

2.77 Name and give the symbol of the element with the given characteristic.
(a) Largest atomic radius in Group 2A.
(b) Smallest atomic radius in Group 2A.
(c) Largest atomic radius in the second period.
(d) Smallest atomic radius in the second period.
(e) Largest ionization energy in Group 7A.
(f) Lowest ionization energy in Group 7A.

2.78 What is the outer-shell electron configuration of the elements in:
(a) Group 3A? (b) Group 7A?
(c) Group 5A?

2.79 Determine the number of protons, electrons, and neutrons present in:
(a) ^{32}P (b) ^{98}Mo (c) ^{44}Ca
(d) ^{3}H (e) ^{158}Gd (f) ^{212}Bi

2.80 What percentage of the mass of each element do neutrons contribute?
(a) Carbon-12 (b) Calcium-40
(c) Iron-55 (d) Bromine-79
(e) Platinum-195 (f) Uranium-238

2.81 Do isotopes of the heavy elements (for example, those from atomic number 37 to 53) contain more, the same, or fewer neutrons than protons?

2.82 What is the symbol for each of the following elements? (Try not to look at a Periodic Table.)
(a) Phosphorus (b) Potassium
(c) Sodium (d) Nitrogen
(e) Bromine (f) Silver
(g) Calcium (h) Carbon
(i) Tin (j) Zinc

2.83 The natural abundance of boron isotopes is as follows: 19.9% boron-10 (10.013 amu) and 80.1% boron-11 (11.009 amu). Calculate the atomic weight of boron (watch the significant figures) and compare your calculated value with that given in the Periodic Table.

2.84 How many electrons are in the outer shell of each of the following elements?
(a) Si (b) Br
(c) P (d) K
(e) He (f) Ca
(g) Kr (h) Pb
(i) Se (j) O

2.85 The mass of a proton is 1.67×10^{-24} g. The mass of a grain of salt is 1.0×10^{-2} g. How many protons would it take to have the same mass as a grain of salt?

2.86 (a) What are the charges of an electron, a proton, and a neutron?
(b) What are the masses (in amu, to one significant figure) of an electron, a proton, and a neutron?

2.87 What is the name of this element, and how many protons and neutrons does this isotope have in its nucleus: $^{131}_{54}X$?

2.88 Based on the data presented in Figure 2.16, which atom would have the highest ionization energy: I, Cs, Sn, or Xe?

2.89 Assume that a new element has been discovered with atomic number 117. Its chemical properties should be similar to those of astatine (At). Predict whether the new element's ionization energy will be greater than, the same as, or smaller than that of:
(a) At (b)Ra

2.90 Explain why the sizes of atoms change when proceeding across a period of the Periodic Table.

2.91 These are the first two ionization energy for lithium:

$$Li(g) \longrightarrow Li^{+}(g) + e^{-}(g) \quad \text{Ionization energy} = 523 \text{ kJ/mol}$$

$$Li^{+}(g) \longrightarrow Li^{2+}(g) + e^{-}(g) \quad \text{Ionization energy} = 7298 \text{ kJ/mol}$$

(a) Explain the large increase in ionization energy that occurs for the removal of the second electron.
(b) The radius of Li^{+} is 78 pm (1 pm $=10^{-12}$ m) while that of a lithium atom, Li, is 152 pm. Explain why the radius of Li^{+} is so much smaller than the radius of Li.

2.92 Which has the largest radius: O^{2-}, F^{-} or F? Explain your reasoning.

2.93 Arrange the following elements in order of increasing size: Al, B, C, and Na. Try doing it without looking at Figure 2.16 and then check yourself by looking at the figure.

2.94 Using your knowledge of trends in element sizes in going across a period of the Periodic Table, explain why the density of the elements increases from potassium through vanadium. (Recall from Section 1.7

that specific gravity is numerically the same as density but has no units.)

Element	Specific Gravity
K	0.862
Ca	1.55
Se	2.99
Ti	4.54
V	6.11

2.95 Name the elements in Group 3A. What does the group designation tell you about the electron configuration of these elements?

2.96 Using the orbital box diagrams and the noble gas notation, write the electron configuration for each atom and ion.

(a) Ti (b) Ti^{2+} (c) Ti^{4+}

2.97 Explain why the Ca^{3+} ion is not found in chemical compounds.

2.98 Explain how the ionization energy of atoms changes when proceeding down a group of the Periodic Table and explain why this change occurs.

Looking Ahead

2.99 Suppose that you face a problem similar to Mendeleyev: You must predict the properties of an element not yet discovered. What will element 118 be like if and when enough of it is made for chemists to study its physical and chemical properties?

2.100 Compare the neutron to proton ratio for the heavier and lighter elements. Does the value of this ratio generally increase, decrease, or remain the same as atomic number increases?