2

CHAPTER OUTLINE

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CHAPTER GOALS

In this chapter you will learn how to:

① Identify an element by its symbol and classify it as a metal, nonmetal, or metalloid
② Describe the basic parts of an atom
③ Distinguish isotopes and calculate atomic weight
④ Describe the basic features of the periodic table
⑤ Understand the electronic structure of an atom
⑥ Write an electronic configuration for an element
⑦ Relate the location of an element in the periodic table to its electronic configuration
⑧ Draw an electron-dot symbol for an atom
⑨ Use the periodic table to predict the relative size and ionization energy of atoms

Both the naturally occurring diamond used in jewelry and the synthetic carbon fibers used in high-end, lightweight bicycles are composed of the element **carbon.**

ATOMS AND THE PERIODIC TABLE

EXAMINE the ingredients listed on a box of crackers. They may include flour, added vitamins, sugar for sweetness, a natural or synthetic coloring agent, baking soda, salt for flavor, and BHT as a preservative. No matter how simple or complex each of these substances is, it is composed of the basic building block, the **atom.** The word *atom* comes from the Greek word *atomos* meaning *unable to cut.* In Chapter 2, we examine the structure and properties of atoms, the building blocks that comprise all forms of matter.

2.1 ELEMENTS

You were first introduced to elements in Section 1.3.

- **An *element* is a pure substance that cannot be broken down into simpler substances by a chemical reaction.**

Of the 114 elements currently known, 90 are naturally occurring and the remaining 24 have been prepared by scientists in the laboratory. Some elements, like oxygen in the air we breathe and aluminum in a soft drink can, are familiar to you, while others, like samarium and seaborgium, are probably not. An alphabetical list of all elements appears on the inside front cover.

Each element is identified by a one- or two-letter symbol. The element carbon is symbolized by the single letter **C,** while the element chlorine is symbolized by **Cl.** When two letters are used in the element symbol, the first is upper case while the second is lower case. Thus, **Co** refers to the element cobalt, but CO is carbon monoxide, which is composed of the elements carbon (C) and oxygen (O). Table 2.1 lists common elements and their symbols.

While most element symbols are derived from the first one or two letters of the element name, 11 elements have symbols derived from the Latin names for them. Table 2.2 lists these elements and their symbols.

Elements are named for people, places, and things. For example, *carbon* (C) comes from the Latin word *carbo,* meaning *coal* or *charcoal; neptunium* (Np) was named for the planet Neptune; *einsteinium* (Es) was named for scientist Albert Einstein; and *californium* (Cf) was named for the state of California.

ENVIRONMENTAL NOTE

Carbon monoxide (CO), formed in small amounts during the combustion of fossil fuels like gasoline, is a toxic component of the smoggy air in many large cities. We will learn about carbon monoxide in Section 12.8.

PROBLEM 2.1

Give the symbol for each element.

a. calcium, a nutrient needed for strong teeth and bones
b. radon, a radioactive gas produced in the soil
c. nitrogen, the main component of the earth’s atmosphere
d. gold, a precious metal used in coins and jewelry

PROBLEM 2.2

An alloy is a mixture of two or more elements that has metallic properties. Give the element symbol for the components of each alloy: (a) brass (copper and zinc); (b) bronze (copper and tin); (c) pewter (tin, antimony, and lead).

PROBLEM 2.3

Give the name corresponding to each element symbol: (a) Ne; (b) S; (c) I; (d) Si; (e) B; (f) Hg.

TABLE 2.1 Common Elements and Their Symbols

Element	Symbol	Element	Symbol
Bromine	Br	Magnesium	Mg
Calcium	Ca	Manganese	Mn
Carbon	C	Molybdenum	Mo
Chlorine	Cl	Nitrogen	N
Chromium	Cr	Oxygen	O
Cobalt	Co	Phosphorus	P
Copper	Cu	Potassium	K
Fluorine	F	Sodium	Na
Hydrogen	H	Sulfur	S
Iodine	I	Zinc	Zn
Lead	Pb		

TABLE 2.2 Element Symbols with Latin Origins

Element	Symbol
Antimony	Sb (stibium)
Copper	Cu (cuprum)
Gold	Au (aurum)
Iron	Fe (ferrum)
Lead	Pb (plumbum)
Mercury	Hg (hydrargyrum)
Potassium	K (kalium)
Silver	Ag (argentum)
Sodium	Na (natrium)
Tin	Sn (stannum)
Tungsten	W (wolfram)

2.1A ELEMENTS AND THE PERIODIC TABLE

How the periodic table is organized is discussed in Section 2.7. A periodic table appears on the inside front cover for easy reference.

Long ago it was realized that groups of elements have similar properties, and that these elements could be arranged in a schematic way called the **periodic table** (Figure 2.1). The position of an element in the periodic table tells us much about its chemical properties.

The elements in the periodic table are divided into three groups—**metals, nonmetals,** and **metalloids.** The solid line that begins with boron (B) and angles in steps down to astatine (At) marks the three regions corresponding to these groups. All metals are located to the *left* of the line. All nonmetals except hydrogen are located to the *right.* Metalloids are located along the steps.

- ***Metals* are shiny solids that are good conductors of heat and electricity. All metals are solids at room temperature except for mercury, which is a liquid.**
- ***Nonmetals* do not have a shiny appearance, and they are generally poor conductors of heat and electricity. Nonmetals like sulfur and carbon are solids at room temperature; bromine is a liquid; and nitrogen, oxygen, and nine other elements are gases.**
- ***Metalloids* have properties intermediate between metals and nonmetals. Only seven elements are categorized as metalloids: boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb), tellurium (Te), and astatine (At).**

PROBLEM 2.4

Locate each element in the periodic table and classify it as a metal, nonmetal, or metalloid.

a. titanium
b. chlorine
c. krypton
d. palladium
e. arsenic
f. cesium
g. selenium
h. osmium

2.1B FOCUS ON THE HUMAN BODY
THE ELEMENTS OF LIFE

Because living organisms selectively take up elements from their surroundings, the abundance of elements in the human body is very different from the distribution of elements in the earth's crust. **Four nonmetals—oxygen, carbon, hydrogen, and nitrogen—comprise 96% of the mass of the human body, and are called the *building-block elements*** (Figure 2.2). Hydrogen and oxygen are the elements that form water, the most prevalent substance in the body. Carbon, hydrogen, and oxygen are found in the four main types of biological molecules—proteins, carbohydrates, lipids, and nucleic acids. Proteins and nucleic acids contain the element nitrogen as well. These biological molecules are discussed in Chapters 19–22.

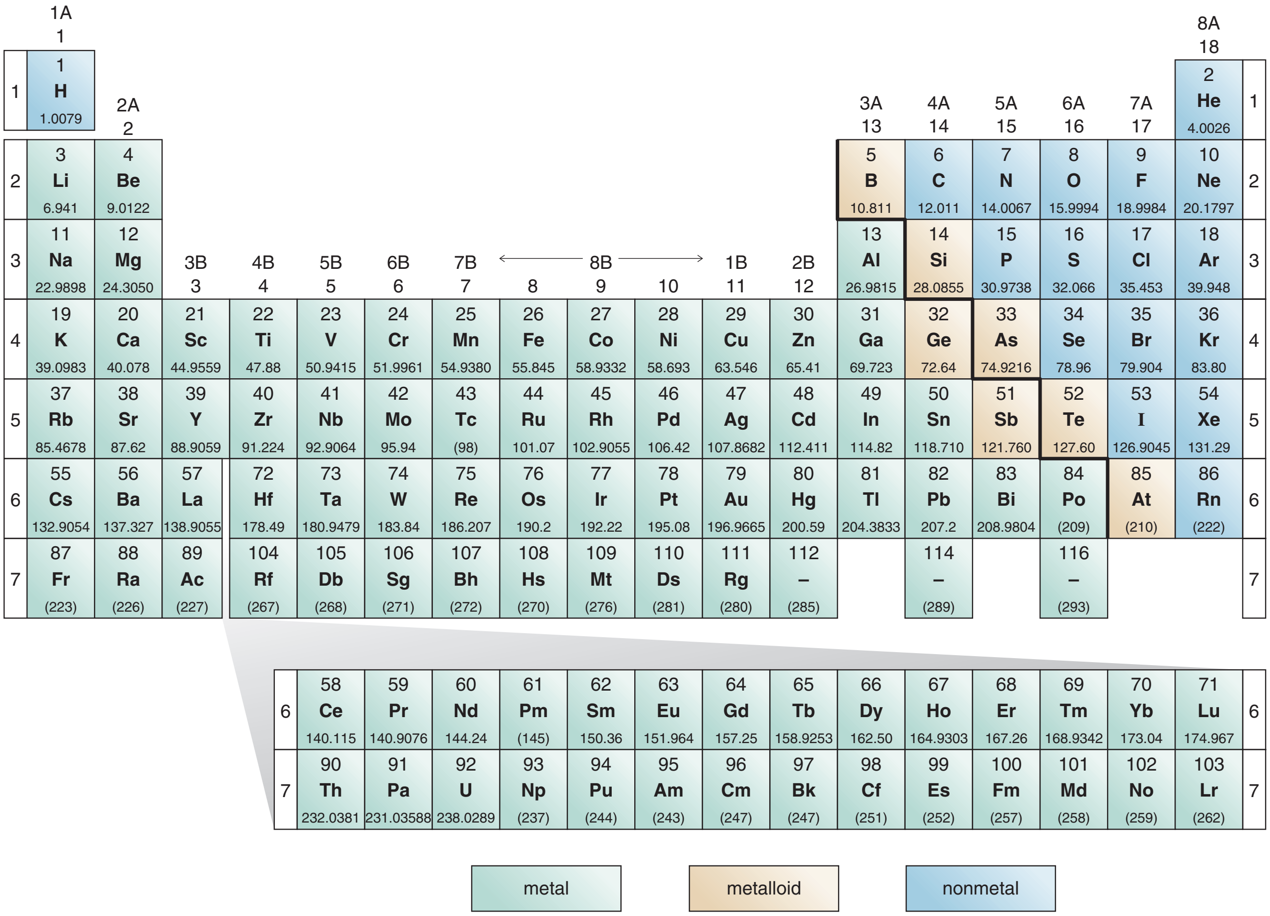
Seven other elements, called the **major minerals** or **macronutrients,** are also present in the body in much smaller amounts (0.1–2% by mass). Sodium, potassium, and chlorine are present in body fluids. Magnesium and sulfur occur in proteins, and calcium and phosphorus are present in teeth and bones. Phosphorus is also contained in all nucleic acids, such as the DNA that transfers genetic information from one generation to another. At least 100 mg of each macronutrient is needed in the daily diet.

Many other elements occur in very small amounts in the body, but are essential to good health. These **trace elements** or **micronutrients** are required in the daily diet in small quantities, usually less than 15 mg. Each trace element has a specialized function that is important for proper cellular function. For example, iron is needed for hemoglobin, the protein that carries oxygen in red blood cells, and myoglobin, the protein that stores oxygen in muscle. Zinc is needed for the proper functioning of many enzymes in the liver and kidneys, and iodine is needed for proper thyroid function. Although most of the trace elements are metals, nonmetals like fluorine and selenium are micronutrients as well.

PROBLEM 2.5

Classify each micronutrient in Figure 2.2 as a metal, nonmetal, or metalloid.

FIGURE 2.1 The Periodic Table of the Elements

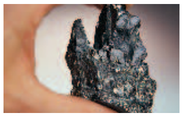


	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8	← 8B → 9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	
1	1 **H** 1.0079																	2 **He** 4.0026	1
2	3 **Li** 6.941	4 **Be** 9.0122											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984	10 **Ne** 20.1797	2
3	11 **Na** 22.9898	12 **Mg** 24.3050											13 **Al** 26.9815	14 **Si** 28.0855	15 **P** 30.9738	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.948	3
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.9380	26 **Fe** 55.845	27 **Co** 58.9332	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.41	31 **Ga** 69.723	32 **Ge** 72.64	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80	4
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.9055	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.411	49 **In** 114.82	50 **Sn** 118.710	51 **Sb** 121.760	52 **Te** 127.60	53 **I** 126.9045	54 **Xe** 131.29	5
6	55 **Cs** 132.9054	56 **Ba** 137.327	57 **La** 138.9055	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.2	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.9665	80 **Hg** 200.59	81 **Tl** 204.3833	82 **Pb** 207.2	83 **Bi** 208.9804	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)	6
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **–** (285)		114 **–** (289)		116 **–** (293)			7

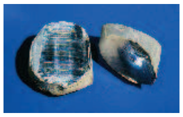
6	58 **Ce** 140.115	59 **Pr** 140.9076	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.964	64 **Gd** 157.25	65 **Tb** 158.9253	66 **Dy** 162.50	67 **Ho** 164.9303	68 **Er** 167.26	69 **Tm** 168.9342	70 **Yb** 173.04	71 **Lu** 174.967	6
7	90 **Th** 232.0381	91 **Pa** 231.03588	92 **U** 238.0289	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)	7

silver (Ag)

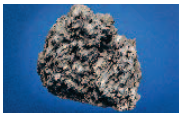
silicon (Si)



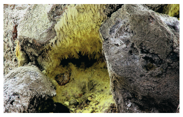
carbon (C)



sodium (Na)



arsenic (As)



sulfur (S)

mercury (Hg)

boron (B)

bromine (Br)

- **Metals** like silver, sodium, and mercury are shiny substances that conduct heat and electricity.
- **Metalloids** like silicon, arsenic, and boron have properties intermediate between metals and nonmetals.
- **Nonmetals** like carbon, sulfur, and bromine are poor conductors of heat and electricity.

FIGURE 2.2 The Elements of Life

Building-Block Elements

Oxygen (O)
Carbon (C)
Hydrogen (H)
Nitrogen (N)

These four elements compose almost 96% of the mass of the human body.

Major Minerals

Potassium (K), sodium (Na), and chlorine (Cl) are present in body fluids.

Magnesium (Mg) and sulfur (S) are found in proteins.

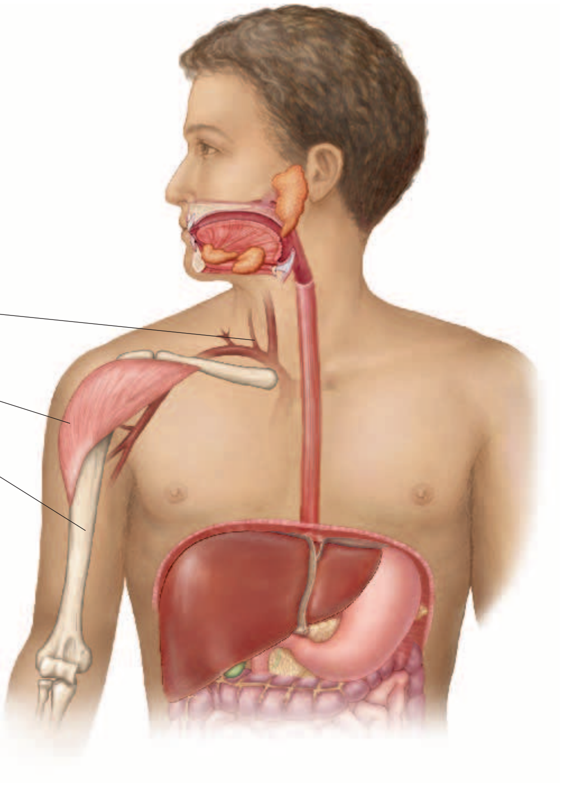
Calcium (Ca) and phosphorus (P) are present in teeth and bones.

Each major mineral is present in 0.1–2% by mass. At least 100 mg of each mineral is needed in the daily diet.

Trace Elements

Arsenic (As)
Boron (B)
Chromium (Cr)
Cobalt (Co)
Copper (Cu)
Fluorine (F)
Iodine (I)
Iron (Fe)
Manganese (Mn)
Molybdenum (Mo)
Nickel (Ni)
Selenium (Se)
Silicon (Si)
Zinc (Zn)

Each trace element is present in less than 0.1% by mass. A small quantity (15 mg or less) of each element is needed in the daily diet.



2.1C COMPOUNDS

In Section 1.3 we learned that a ***compound*** **is a pure substance formed by chemically combining two or more elements together.** Element symbols are used to write chemical formulas for compounds.

- **A *chemical formula* uses element symbols to show the identity of the elements forming a compound and subscripts to show the ratio of atoms (the building blocks of matter) contained in the compound.**

For example, table salt is formed from sodium (Na) and chlorine (Cl) in a ratio of 1:1, so its formula is NaCl. Water, on the other hand, is formed from two hydrogen atoms for each oxygen atom, so its formula is H_2O. The subscript "1" is understood when no subscript is written. Other examples of chemical formulas are shown below.

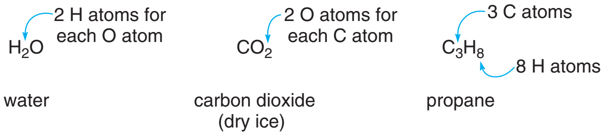
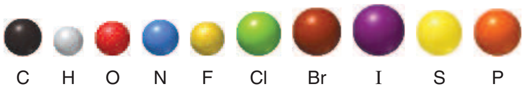
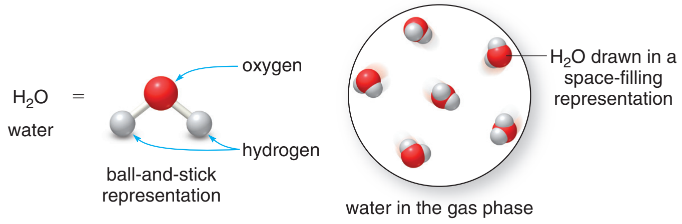


FIGURE 2.3 Common Element Colors Used in Molecular Art



As we learned in Section 1.2, molecular art will often be used to illustrate the composition and state of elements and compounds. Color-coded spheres, shown in Figure 2.3, are used to identify the common elements that form compounds.

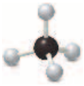
For example, a red sphere is used for the element oxygen and gray is used for the element hydrogen, so H_2O is represented as a red sphere joined to two gray spheres. Sometimes the spheres will be connected by "sticks" to generate a **ball-and-stick** representation for a compound. At other times, the spheres will be drawn close together to form a **space-filling** representation. No matter how the spheres are depicted, H_2O always consists of one red sphere for the oxygen atom and two gray spheres for the two hydrogen atoms.



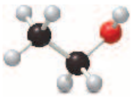
PROBLEM 2.6 Identify the elements in each chemical formula, and give the number of atoms of each element.

a. NaCN (sodium cyanide)
b. H_2S (hydrogen sulfide)
c. C_2H_6 (ethane)
d. SnF_2 (stannous fluoride)
e. CO (carbon monoxide)
f. $C_3H_8O_3$ (glycerol)

PROBLEM 2.7 Identify the elements used in each example of molecular art.

a. 

b. 

c. 

2.2 STRUCTURE OF THE ATOM

All matter is composed of the same basic building blocks called *atoms*. An atom is much too small to be seen even by the most powerful light microscopes. The period at the end of this sentence holds about 1 × 10^8 atoms, and a human cheek cell contains about 1 × 10^{16} atoms. An atom is composed of three subatomic particles.

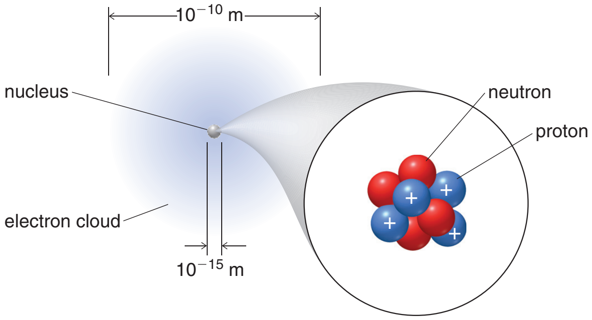
- **A proton, symbolized by p, has a positive (+) charge.**
- **An electron, symbolized by e^-, has a negative (–) charge.**
- **A neutron, symbolized by n, has no charge.**

Protons and neutrons have approximately the same, exceedingly small mass, as shown in Table 2.3. The mass of an electron is much less, 1/1,836 the mass of a proton. These subatomic particles are not evenly distributed in the volume of an atom. There are two main components of an atom.

TABLE 2.3 Summary: The Properties of the Three Subatomic Particles

Subatomic Particle	Charge	Mass (g)	Mass (amu)
Proton	+1	1.6726×10^{-24}	1
Neutron	0	1.6749×10^{-24}	1
Electron	–1	9.1093×10^{-28}	Negligible

- **The *nucleus* is a dense core that contains the protons and neutrons. Most of the mass of an atom resides in the nucleus.**
- **The *electron cloud* is composed of electrons that move rapidly in the almost empty space surrounding the nucleus. The electron cloud comprises most of the volume of an atom.**



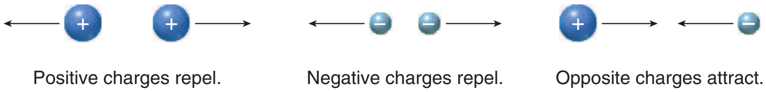
main components of an atom

While the diameter of an atom is about 10^{-10} m, the diameter of a nucleus is only about 10^{-15} m. For a macroscopic analogy, if the nucleus were the size of a baseball, an atom would be the size of Yankee Stadium!

The charged particles of an atom can either attract or repel each other.

- **Opposite charges attract while like charges repel each other.**

Thus, two electrons or two protons repel each other, while a proton and an electron attract each other.



Since the mass of an individual atom is so small (on the order of 10^{-24} g), chemists use a standard mass unit, the **atomic mass unit,** which defines the mass of individual atoms relative to a standard mass.

- **One atomic mass unit (amu) equals one-twelfth the mass of a carbon atom that has six protons and six neutrons; 1 amu = 1.661×10^{-24} g.**

Using this scale, one proton has a mass of 1.0073 amu, a value typically rounded to 1 amu. One neutron has a mass of 1.0087 amu, a value also typically rounded to 1 amu. The mass of an electron is so small that it is ignored.

Every atom of a given type of element always has the *same* number of protons in the nucleus, a value called the *atomic number,* symbolized by ***Z.*** Conversely, two *different* elements have *different* atomic numbers.

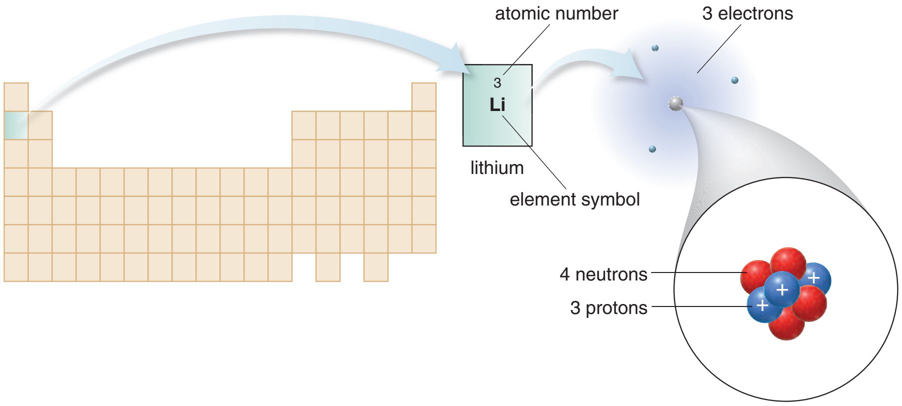
- The *atomic number* (*Z*) = the number of protons in the nucleus of an atom.

Thus, the element hydrogen has one proton in its nucleus, so its atomic number is one. Lithium has three protons in its nucleus, so its atomic number is three. The periodic table is arranged in order of increasing atomic number beginning at the upper left-hand corner. The atomic number appears just above the element symbol for each entry in the table.

Since a neutral atom has no overall charge:

- *Z* = the number of protons in the nucleus = the number of electrons.

Thus, the atomic number tells us *both* the number of protons in the nucleus and the number of electrons in the electron cloud of a neutral atom.



SAMPLE PROBLEM 2.1

Identify the element that has an atomic number of 19, and give the number of protons and electrons in the neutral atom.

ANALYSIS The atomic number is unique to an element and tells the number of protons in the nucleus and the number of electrons in the electron cloud of a neutral atom.

SOLUTION According to the periodic table, the element potassium has atomic number 19. A neutral potassium atom has 19 protons and 19 electrons.

PROBLEM 2.8

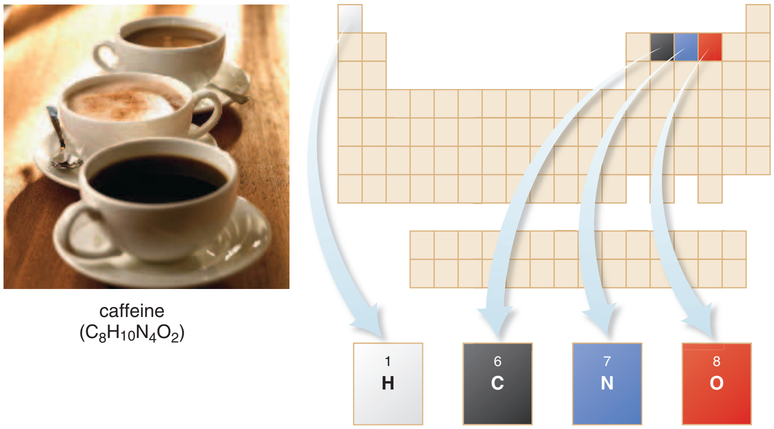
Identify the element with each atomic number, and give the number of protons and electrons in the neutral atom: (a) 2; (b) 11; (c) 20; (d) 47; (e) 78.

Both protons and neutrons contribute to the mass of an atom. The **mass number,** symbolized by ***A,*** is the sum of the number of protons and neutrons.

- Mass number (*A*) = the number of protons (*Z*) + the number of neutrons.

For example, a fluorine atom with nine protons and 10 neutrons in the nucleus has a mass number of 19. Figure 2.4 lists the atomic number, mass number, and number of subatomic particles in the four building-block elements—hydrogen, carbon, nitrogen, and oxygen—found in a wide variety of compounds including caffeine (chemical formula $C_8H_{10}N_4O_2$), the bitter-tasting mild stimulant in coffee, tea, and cola beverages.

FIGURE 2.4 Atomic Composition of the Four Building-Block Elements



Element	Hydrogen	Carbon	Nitrogen	Oxygen
Atomic number	1	6	7	8
Mass number	1	12	14	16
Number of protons	1	6	7	8
Number of electrons	1	6	7	8
Number of neutrons	0	6	7	8

SAMPLE PROBLEM 2.2

How many protons, neutrons, and electrons are contained in an atom of argon, which has an atomic number of 18 and a mass number of 40?

ANALYSIS

- In a neutral atom, the atomic number (Z) = the number of protons = the number of electrons.
- The mass number (A) = the number of protons + the number of neutrons.

SOLUTION

The atomic number of 18 means that argon has 18 protons and 18 electrons. To find the number of neutrons, subtract the atomic number (Z) from the mass number (A).

$$\begin{aligned}\text{number of neutrons} &= \text{mass number} - \text{atomic number}\\ &= 40 - 18\\ &= 22 \text{ neutrons}\end{aligned}$$

PROBLEM 2.9

How many protons, neutrons, and electrons are contained in each atom with the given atomic number and mass number?

a. $Z = 17, A = 35$
b. $Z = 14, A = 28$
c. $Z = 92, A = 238$

PROBLEM 2.10

What element has an atomic number of 53 and contains 74 neutrons? How many electrons does this atom contain? What is its mass number?

PROBLEM 2.11

What is the mass number of an atom that contains

a. 42 protons, 42 electrons, and 53 neutrons?
b. 24 protons, 24 electrons, and 28 neutrons?

2.3 ISOTOPES

Two atoms of the same element always have the same number of protons, but the number of neutrons can vary.

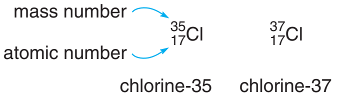
- ***Isotopes*** **are atoms of the same element having a different number of neutrons.**

2.3A ISOTOPES, ATOMIC NUMBER, AND MASS NUMBER

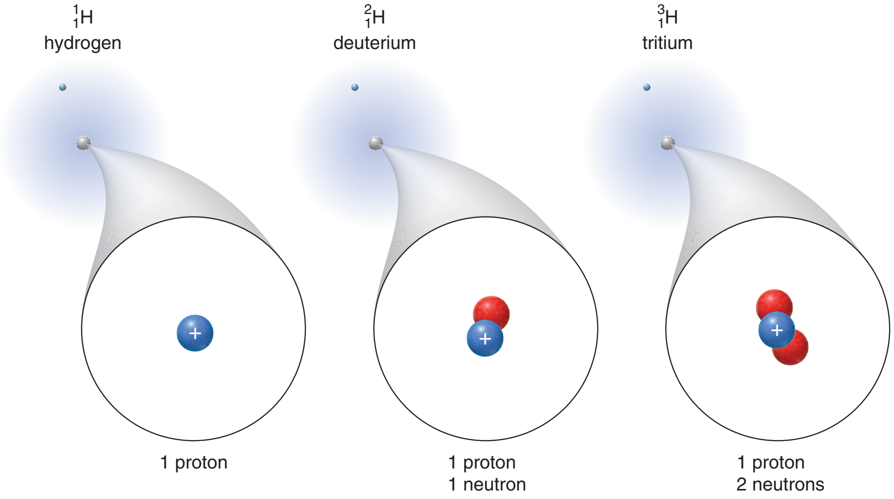
Most elements in nature exist as a mixture of isotopes. For example, all atoms of the element chlorine contain 17 protons in the nucleus, but some of these atoms have 18 neutrons in the nucleus and some have 20 neutrons. Thus, chlorine has two isotopes with different mass numbers, 35 and 37. These isotopes are often referred to as chlorine-35 (or Cl-35) and chlorine-37 (or Cl-37).

Isotopes are also written using the element symbol with the atomic number written as a subscript and the mass number written as a superscript, both to the left.

Two isotopes of the element chlorine



The element hydrogen has three isotopes. Most hydrogen atoms have one proton and no neutrons, giving them a mass number of one. About 1% of hydrogen atoms have one proton and one neutron, giving them a mass number of two. This isotope is called **deuterium,** and it is often symbolized as **D.** An even smaller number of hydrogen atoms contain one proton and two neutrons, giving them a mass number of three. This isotope is called **tritium,** symbolized as **T.**



SAMPLE PROBLEM 2.3

For each atom give the following information: [1] the atomic number; [2] the mass number; [3] the number of protons; [4] the number of neutrons; [5] the number of electrons.

a. $^{118}_{50}Sn$ b. $^{195}_{78}Pt$

ANALYSIS

- The superscript gives the mass number and the subscript gives the atomic number for each element.
- The atomic number = the number of protons = the number of electrons.
- The mass number = the number of protons + the number of neutrons.

SOLUTION

	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
a. $^{118}_{50}Sn$	50	118	50	118 – 50 = 68	50
b. $^{195}_{78}Pt$	78	195	78	195 – 78 = 117	78

PROBLEM 2.12

For each atom give the following information: [1] the atomic number; [2] the mass number; [3] the number of protons; [4] the number of neutrons; [5] the number of electrons.

a. ${}^{13}_{6}\text{C}$ b. ${}^{121}_{51}\text{Sb}$

SAMPLE PROBLEM 2.4

Determine the number of neutrons in each isotope: (a) carbon-14; (b) ${}^{81}\text{Br}$.

ANALYSIS

- The identity of the element tells us the atomic number.
- The number of neutrons = mass number (*A*) – atomic number (*Z*).

SOLUTION

a. Carbon's atomic number (*Z*) is 6. Carbon-14 has a mass number (*A*) of 14.

$$\begin{aligned}\text{number of neutrons} &= A - Z\\ &= 14 - 6 = 8 \text{ neutrons}\end{aligned}$$

b. Bromine's atomic number is 35 and the mass number of the given isotope is 81.

$$\begin{aligned}\text{number of neutrons} &= A - Z\\ &= 81 - 35 = 46 \text{ neutrons}\end{aligned}$$

PROBLEM 2.13

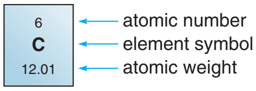
Magnesium has three isotopes that contain 12, 13, and 14 neutrons. For each isotope give the following information: (a) the number of protons; (b) the number of electrons; (c) the atomic number; (d) the mass number. Write the element symbol of each isotope using a superscript and subscript for mass number and atomic number, respectively.

2.3B ATOMIC WEIGHT

Some elements like fluorine occur naturally as a single isotope. More commonly, an element is a mixture of isotopes, and it is useful to know the average mass, called the **atomic weight** (or **atomic mass**), of the atoms in a sample.

- **The *atomic weight* is the weighted average of the mass of the naturally occurring isotopes of a particular element reported in atomic mass units.**

The atomic weights of the elements appear in the alphabetical list of elements on the inside front cover. The atomic weight is also given under the element symbol in the periodic table on the inside front cover.



carbon

To determine the atomic weight of an element, two quantities must be known: the mass of each isotope in atomic mass units, and the frequency with which each isotope occurs.

HOW TO Determine the Atomic Weight of an Element

EXAMPLE **What is the atomic weight of the element chlorine?**

Step [1] **List each isotope, along with its mass in atomic mass units (amu) and the percentage that each isotope occurs in nature.**

- Chlorine has two isotopes—Cl-35 and Cl-37.
- To solve the problem, the masses and abundances of the isotopes must be known.

	Mass (amu)	Isotopic Abundance
Cl-35	34.97	75.78% = 0.7578
Cl-37	36.97	24.22% = 0.2422

- The mass of any isotope is very close to the mass number of the isotope.
- To convert a percent to a decimal, divide by 100%, which is the same as moving the decimal point two places to the left; thus,

75.78% = 0.7578

Step [2] **Multiply the isotopic abundance by the mass of each isotope, and add up the products. The sum is the atomic weight for the element.**

Mass due to Cl-35: 0.7578×34.97 amu = 26.5003 amu
Mass due to Cl-37: 0.2422×36.97 amu = 8.9541 amu
Atomic weight = 35.4544 amu rounded to 35.45 amu
Answer

SAMPLE PROBLEM 2.5

Calculate the atomic weight of copper, which has two isotopes with the following properties: Cu-63 (62.93 amu, 69.17% natural occurrence) and Cu-65 (64.93 amu, 30.83% natural occurrence).

ANALYSIS Multiply the isotopic abundance by the mass of each isotope, and add up the products to give the atomic weight for the element.

SOLUTION

Mass due to Cu-63: 0.6917×62.93 amu = 43.5287 amu
Mass due to Cu-65: 0.3083×64.93 amu = 20.0179 amu
Atomic weight = 63.5466 amu rounded to 63.55 amu
Answer

PROBLEM 2.14

Calculate the atomic weight of each element given the mass and natural occurrence of each isotope.

a. Magnesium	Mass (amu)	Isotopic Abundance
Mg-24	23.99	78.99%
Mg-25	24.99	10.00%
Mg-26	25.98	11.01%

b. Vanadium	Mass (amu)	Isotopic Abundance
V-50	49.95	0.250%
V-51	50.94	99.750%

2.3C FOCUS ON HEALTH & MEDICINE ISOTOPES IN MEDICINE

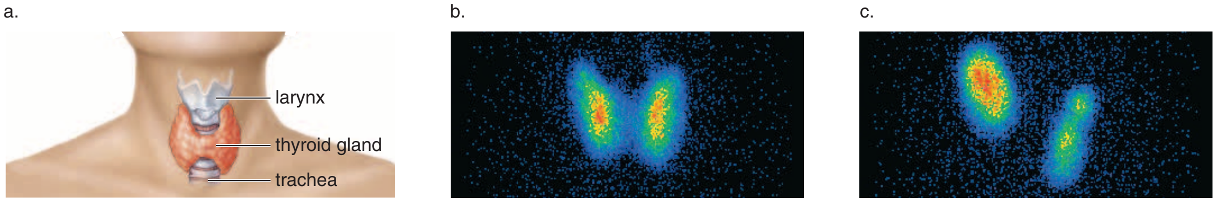
Generally the chemical properties of isotopes are identical. Sometimes, however, one isotope of an element is radioactive—that is, it emits particles or energy as some form of radiation. Radioactive isotopes have both diagnostic and therapeutic uses in medicine.

As an example, iodine-131 is used in at least two different ways for thyroid disease. Iodine is a micronutrient needed by the body to synthesize the thyroid hormone thyroxine, which contains four iodine atoms. To evaluate the thyroid gland, a patient can be given sodium iodide (NaI) that contains radioactive iodine-131. Iodine-131 is taken up in the thyroid gland and as it emits radiation, it produces an image in a thyroid scan, which is then used to determine the condition of the thyroid gland, as shown in Figure 2.5.

Higher doses of iodine-131 can also be used to treat thyroid disease. Since the radioactive isotope is taken up by the thyroid gland, the radiation it emits can kill overactive or cancerous cells in the thyroid.

Other applications of radioactive isotopes in medicine are discussed in Chapter 10.

FIGURE 2.5 Iodine-131 in Medicine



The thyroid gland is a butterfly-shaped gland in the neck, shown in (a). Uptake of radioactive iodine-131 can reveal the presence of a healthy thyroid as in (b), or an unsymmetrical thyroid gland with dense areas of iodine uptake as in (c), which may be indicative of cancer or other thyroid disease.

2.4 THE PERIODIC TABLE

Every beginning chemistry text has a periodic table in a prominent location—often the inside front cover—because it is a valuable list of all known elements organized so that groups of elements with similar characteristics are arranged together. The periodic table evolved over many years, and it resulted from the careful observations and experiments of many brilliant scientists in the nineteenth century. Most prominent was Russian chemist Dmitri Mendeleev, whose arrangement in 1869 of the 60 known elements into groups having similar properties in order of increasing atomic number became the precursor of the modern periodic table (inside front cover and Figure 2.6).

2.4A BASIC FEATURES OF THE PERIODIC TABLE

The periodic table is arranged into seven horizontal rows and 18 vertical columns. The particular row and column tell us much about the properties of an element.

- **A row in the periodic table is called a *period.* Elements in the same row are similar in size.**
- **A column in the periodic table is called a *group.* Elements in the same group have similar electronic and chemical properties.**

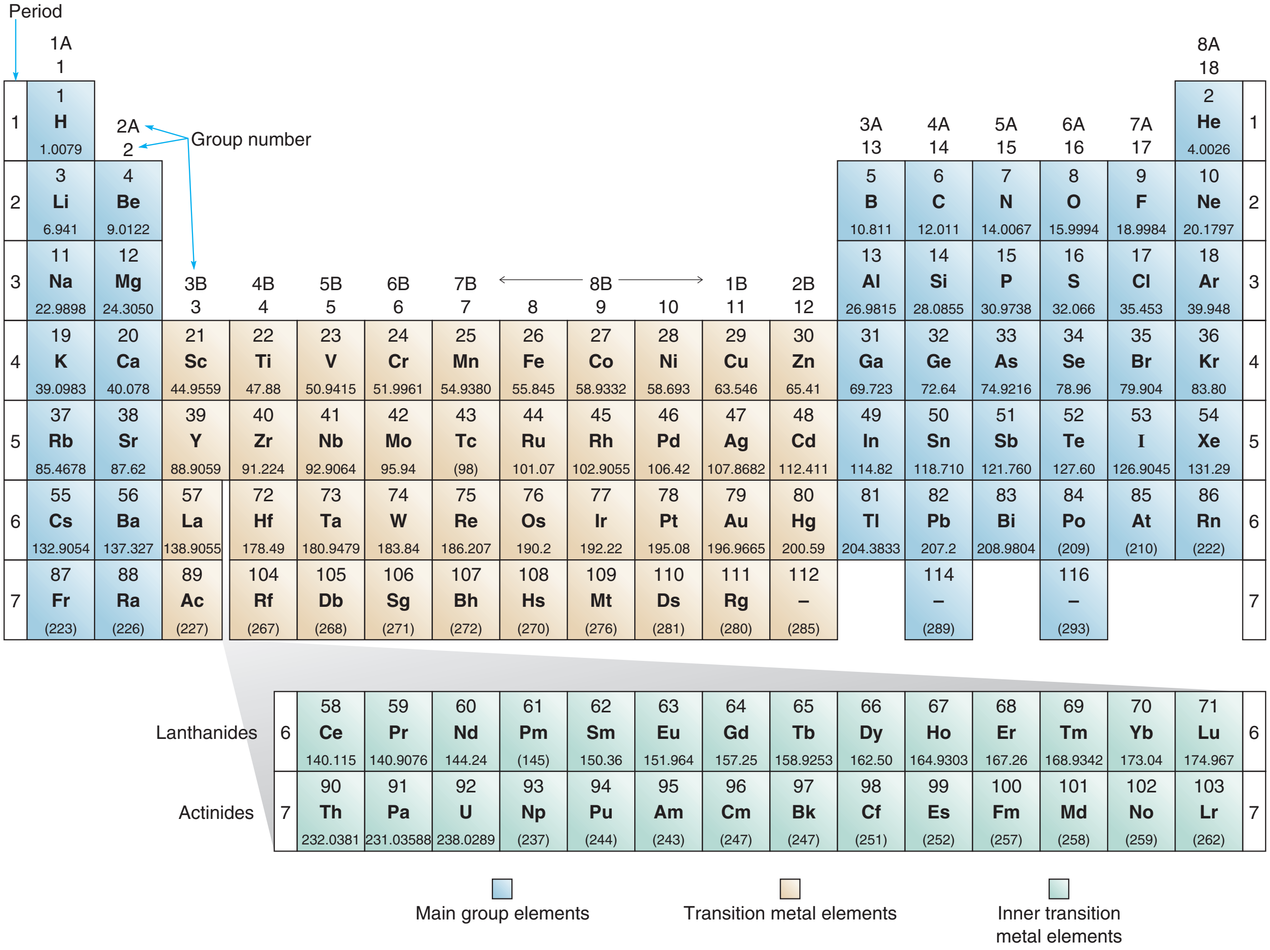
The rows in the periodic table are numbered 1–7. The number of elements in each row varies. The first period has just two elements, hydrogen and helium. The second and third rows have eight elements each, and the fourth and fifth rows have 18 elements. Also note that two groups of fourteen elements appear at the bottom of the periodic table. The **lanthanides,** beginning with the element cerium ($Z = 58$), immediately follow the element lanthanum (La). The **actinides,** beginning with thorium ($Z = 90$), immediately follow the element actinium (Ac).

Each column in the periodic table is assigned a **group number.** Groups are numbered in two ways. In one system, the 18 columns of the periodic table are assigned the numbers 1–18, beginning with the column farthest to the left. An older but still widely used system numbers the groups 1–8, followed by the letter A or B.

- **The *main group elements* consist of the two columns on the far left and the six columns on the far right of the table. These groups are numbered 1A–8A.**
- **The *transition metal elements* are contained in the 10 short columns in the middle of the table, numbered 1B–8B.**
- **The *inner transition elements* consist of the lanthanides and actinides, and they are not assigned group numbers.**

The periodic table in Figure 2.6 has both systems of numbering groups. For example, the element carbon (C) is located in the second row (period 2) of the periodic table. Its group number is 4A (or 14).

FIGURE 2.6 Basic Features of the Periodic Table



Period	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8	8B 9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	
1	1 H 1.0079																	2 He 4.0026	1
2	3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797	2
3	11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.453	18 Ar 39.948	3
4	19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9380	26 Fe 55.845	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.41	31 Ga 69.723	32 Ge 72.64	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80	4
5	37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.82	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.9045	54 Xe 131.29	5
6	55 Cs 132.9054	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.9665	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.9804	84 Po (209)	85 At (210)	86 Rn (222)	6
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (267)	105 Db (268)	106 Sg (271)	107 Bh (272)	108 Hs (270)	109 Mt (276)	110 Ds (281)	111 Rg (280)	112 – (285)		114 – (289)		116 – (293)			7

Lanthanides	6	58 Ce 140.115	59 Pr 140.9076	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.9253	66 Dy 162.50	67 Ho 164.9303	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.967	6
Actinides	7	90 Th 232.0381	91 Pa 231.03588	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)	7

- Each element of the periodic table is part of a horizontal row and a vertical column.
- The periodic table consists of seven rows, labeled periods 1–7, and 18 columns that are assigned a group number. Two different numbering systems are indicated.
- Elements are divided into three categories: main group elements (groups 1A–8A, shown in light blue), transition metals (groups 1B–8B, shown in tan), and inner transition metals (shown in light green).

SAMPLE PROBLEM 2.6 Give the period and group number for each element: (a) magnesium; (b) manganese.

ANALYSIS Use the element symbol to locate an element in the periodic table. Count down the rows of elements to determine the period. The group number is located at the top of each column.

SOLUTION

a. Magnesium (Mg) is located in the third row (period 3), and has group number 2A (or 2).

b. Manganese (Mn) is located in the fourth row (period 4), and has group number 7B (or 7).

PROBLEM 2.15 Give the period and group number for each element: (a) oxygen; (b) calcium; (c) phosphorus; (d) platinum; (e) iodine.

2.4B CHARACTERISTICS OF GROUPS 1A, 2A, 7A, AND 8A

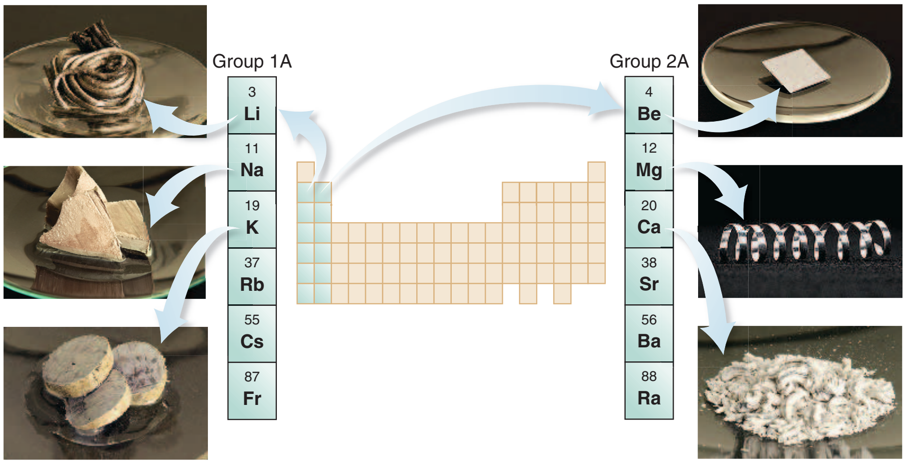
Four columns of main group elements illustrate an important fact about the periodic table.

- **Elements that comprise a particular group have similar chemical properties.**

Alkali Metals (Group 1A) and Alkaline Earth Elements (Group 2A)

The alkali metals and the alkaline earth elements are located on the far left side of the periodic table.

Although hydrogen is also located in group 1A, it is *not* an alkali metal.



The **alkali metals,** located in group 1A (group 1), include lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr). Alkali metals share the following characteristics:

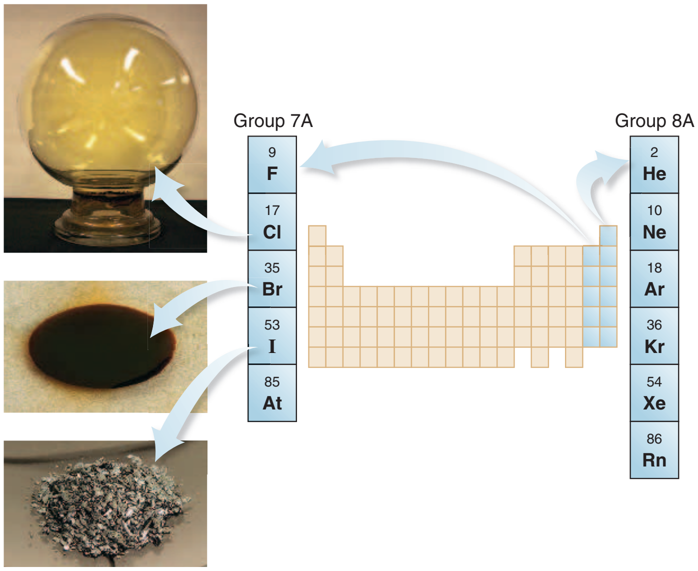
- They are soft and shiny and have low melting points.
- They are good conductors of heat and electricity.
- They react readily with water to form basic solutions.

The **alkaline earth elements,** located in group 2A (group 2), include beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). Alkaline earth metals are also shiny solids but less reactive than the alkali metals.

None of the metals in groups 1A or 2A exist in nature as pure elements; rather, they are always combined with other elements to form compounds. Examples of compounds from group 1A elements include sodium chloride (NaCl), table salt; potassium iodide (KI), an essential nutrient added to make iodized salt; and lithium carbonate (Li_2CO_3), a drug used to treat bipolar disorder. Examples of compounds from group 2A elements include magnesium sulfate ($MgSO_4$), an anticonvulsant used to prevent seizures in pregnant women; and barium sulfate ($BaSO_4$), which is used to improve the quality of X-ray images of the gastrointestinal tract.

Halogens (Group 7A) and Noble Gases (Group 8A)

The halogens and noble gases are located on the far right side of the periodic table.



HEALTH NOTE

Radon detectors are used to measure high levels of radon, a radioactive noble gas linked to an increased incidence of lung cancer.

The **halogens,** located in group 7A (group 17), include fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and the rare radioactive element astatine (At). In their elemental form, halogens contain two atoms joined together—F_2, Cl_2, Br_2, and I_2. Fluorine and chlorine are gases at room temperature, bromine is a liquid, and iodine is a solid. Halogens are very reactive and combine with many other elements to form compounds. In Chapter 14, we will learn about carbon compounds that contain halogen atoms.

The **noble gases,** located in group 8A (group 18), include helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). Unlike other elements, the noble gases are especially stable as atoms, and so they rarely combine with other elements to form compounds.

The noble gas **radon** has received attention in recent years. Radon is a radioactive gas, and generally its concentration in the air is low and therefore its presence harmless. In some types of soil, however, radon levels can be high and radon detectors are recommended for the basement of homes to monitor radon levels. High radon levels are linked to an increased risk of lung cancer.

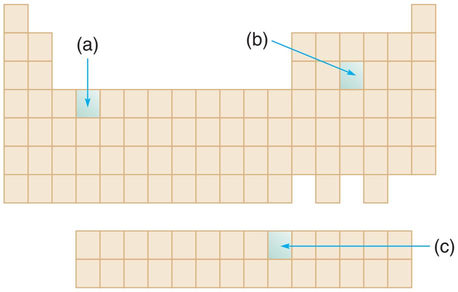
PROBLEM 2.16

Identify the element fitting each description.

a. an alkali metal in period 4
b. a second-row element in group 7A
c. a noble gas in the third period
d. a main group element in period 5 and group 2A
e. a transition metal in group 12, period 4
f. a transition metal in group 5, period 5

PROBLEM 2.17

Identify each highlighted element in the periodic table and give its [1] element name and symbol; [2] group number; [3] period; [4] classification (main group element, transition metal, or inner transition metal).



2.4C THE UNUSUAL NATURE OF CARBON

Carbon, a second-row element in group 4A of the periodic table, is different from most other elements in that it has three elemental forms (Figure 2.7). The two most common forms of carbon are diamond and graphite. **Diamond** is hard because it contains a dense three-dimensional network of carbon atoms in six-membered rings. **Graphite,** on the other hand, is a slippery black substance used as a lubricant. It contains parallel sheets of carbon atoms in flat six-membered rings.

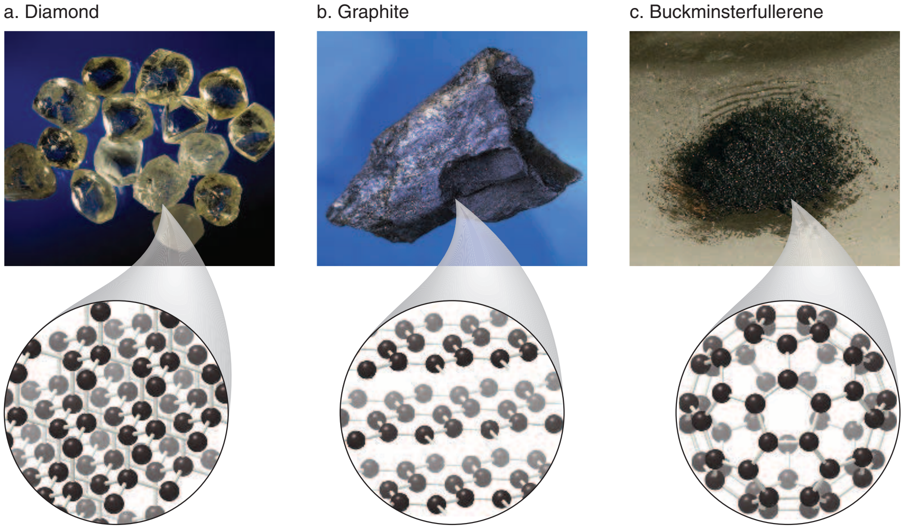
Buckminsterfullerene, also referred to as a bucky ball, is a third form that contains 60 carbon atoms joined together in a sphere of 20 hexagons and 12 pentagons in a pattern that resembles a soccer ball. A component of soot, this form of carbon was not discovered until 1985. Its unusual name stems from its shape, which resembles the geodesic dome invented by R. Buckminster Fuller.

Carbon's ability to join with itself and other elements gives it versatility not seen with any other element in the periodic table. In the unscientific but eloquent description by writer Bill Bryson in *A Short History of Nearly Everything,* carbon is described as "the party animal of the atomic world, latching on to many other atoms (including itself) and holding tight, forming molecular conga lines of hearty robustness—the very trick of nature necessary to build proteins and DNA." As a result, millions of compounds that contain the element carbon are known. The chemistry of these compounds is discussed at length in Chapters 11–24.

2.5 ELECTRONIC STRUCTURE

Why do elements in a group of the periodic table have similar chemical properties? **The chemistry of an element is determined by the number of *electrons* in an atom.** To understand the properties of an element, therefore, we must learn more about the electrons that surround the nucleus.

FIGURE 2.7 Three Elemental Forms of Carbon



a. Diamond

- Diamond consists of an intricate three-dimensional network of carbon atoms.

b. Graphite

- Graphite contains parallel sheets of carbon atoms.

c. Buckminsterfullerene

- Buckminsterfullerene contains a sphere with 60 carbon atoms.

The modern description of the electronic structure of an atom is based on the following principles.

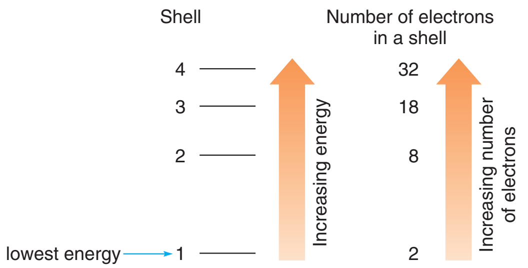
- Electrons do not move freely in space; rather, an electron is confined to a specific region, giving it a particular energy.
- Electrons occupy discrete energy levels. The energy of electrons is *quantized;* that is, the energy is restricted to specific values.

The electrons that surround a nucleus are confined to regions called the **principal energy levels,** or **shells.**

- The shells are numbered, $n = 1, 2, 3, 4$, and so forth, beginning closest to the nucleus.
- Electrons closer to the nucleus are held more tightly and are lower in energy.
- Electrons farther from the nucleus are held less tightly and are higher in energy.

The number of electrons that can occupy a given shell is determined by the value of n. **The farther a shell is from the nucleus, the larger its volume becomes, and the more electrons it can hold.** Thus, the first shell can hold only two electrons, the second holds eight, the third 18, and so forth.

Distribution of electrons in the first four shells

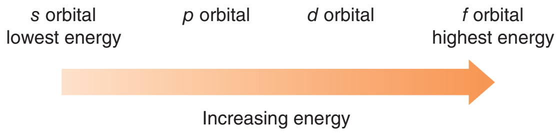


Shells are divided into **subshells,** identified by the letters ***s, p, d,*** and ***f.*** The subshells consist of **orbitals.**

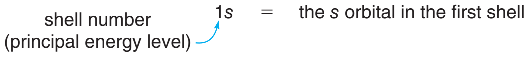
- An *orbital* is a region of space where the probability of finding an electron is high. Each orbital can hold *two* electrons.

A particular type of subshell contains a specific number of orbitals. An *s* subshell contains only **one** *s* orbital. A *p* subshell has **three** *p* orbitals. A *d* subshell has **five** *d* orbitals. An *f* subshell has **seven** *f* orbitals. The number of subshells in a given shell equals the value of n. The energy of orbitals shows the following trend:

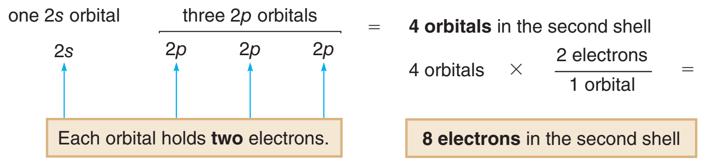
The two electrons in an orbital must have opposite spins. If one electron has a clockwise spin, the second electron in the orbital must have a counterclockwise spin.



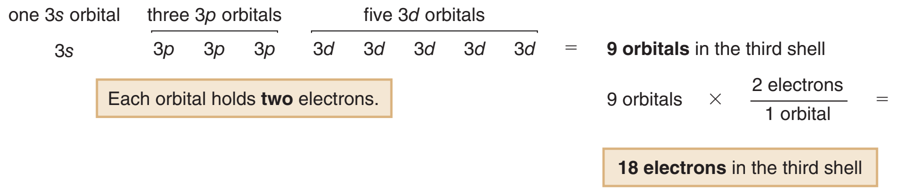
The first shell of electrons around a nucleus ($n = 1$) has only one *s* orbital. This orbital is called the 1*s* orbital since it is the *s* orbital in the first shell. Since each orbital can hold two electrons and the first shell has only one orbital, the **first shell can hold two electrons.**



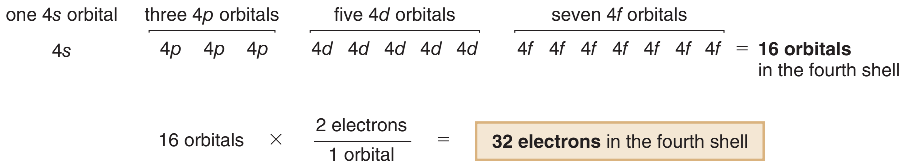
The second shell of electrons ($n = 2$) has two types of orbitals—one *s* and three *p* orbitals. These orbitals are called the 2*s* and 2*p* orbitals since they are located in the second shell. Since each orbital can hold two electrons and there are four orbitals, the **second shell can hold eight electrons.**



The third shell of electrons ($n = 3$) has three types of orbitals—one *s,* three *p,* and five *d* orbitals. These orbitals are called the 3*s,* 3*p,* and 3*d* orbitals since they are located in the third shell. Since each orbital can hold two electrons and the third shell has a total of nine orbitals, the **third shell can hold 18 electrons.**



The fourth shell of electrons ($n = 4$) has four types of orbitals—one *s,* three *p,* five *d,* and seven *f* orbitals. These orbitals are called the 4*s,* 4*p,* 4*d,* and 4*f* orbitals since they are located in the fourth shell. Since each orbital can hold two electrons and the fourth shell has a total of sixteen orbitals, the **fourth shell can hold 32 electrons.**



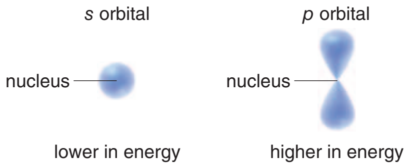
Thus, the maximum number of electrons that can occupy a shell is determined by the number of orbitals in the shell. Table 2.4 summarizes the orbitals and electrons in the first four shells.

TABLE 2.4 Orbitals and Electrons Contained in the Principal Energy Levels (n = 1–4)

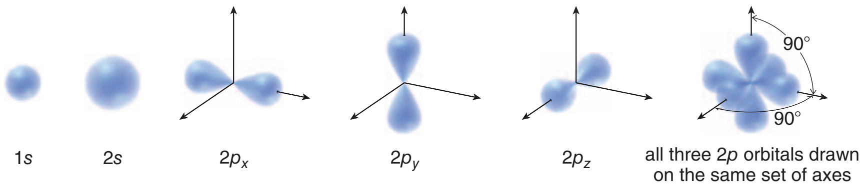
Shell	Orbitals	Electrons in Each Subshell	Maximum Number of Electrons
1	1*s*	2	2
2	2*s* 2*p* 2*p* 2*p*	2 3 × 2 = 6	8
3	3*s* 3*p* 3*p* 3*p* 3*d* 3*d* 3*d* 3*d* 3*d*	2 3 × 2 = 6 5 × 2 = 10	18
4	4*s* 4*p* 4*p* 4*p* 4*d* 4*d* 4*d* 4*d* 4*d* 4*f* 4*f* 4*f* 4*f* 4*f* 4*f* 4*f*	2 3 × 2 = 6 5 × 2 = 10 7 × 2 = 14	32

Each type of orbital has a particular shape.

- **An *s* orbital has a sphere of electron density. It is lower in energy than other orbitals in the same shell because electrons are kept closer to the positively charged nucleus.**
- **A *p* orbital has a dumbbell shape. A *p* orbital is higher in energy than an *s* orbital in the same shell because its electron density is farther from the nucleus.**



All *s* orbitals are spherical, but the orbital gets larger in size as the shell number increases. Thus, both a 1*s* orbital and a 2*s* orbital are spherical, but the 2*s* orbital is larger. The three *p* orbitals in a shell are perpendicular to each other along the *x, y,* and *z* axes.



PROBLEM 2.18

How many electrons are present in each shell, subshell, or orbital?

a. a 2*p* orbital
b. the 3*d* subshell
c. a 3*d* orbital
d. the third shell

PROBLEM 2.19

What element fits each description?

a. the element with electrons that completely fill the first and second shells
b. the element with a completely filled first shell and four electrons in the second shell
c. the element with a completely filled first and second shell, and two electrons in the third shell

2.6 ELECTRONIC CONFIGURATIONS

We can now examine the **electronic configuration** of an individual atom—that is, how the electrons are arranged in an atom's orbitals. **The lowest energy arrangement of electrons is called the *ground state*.** Three rules are followed.

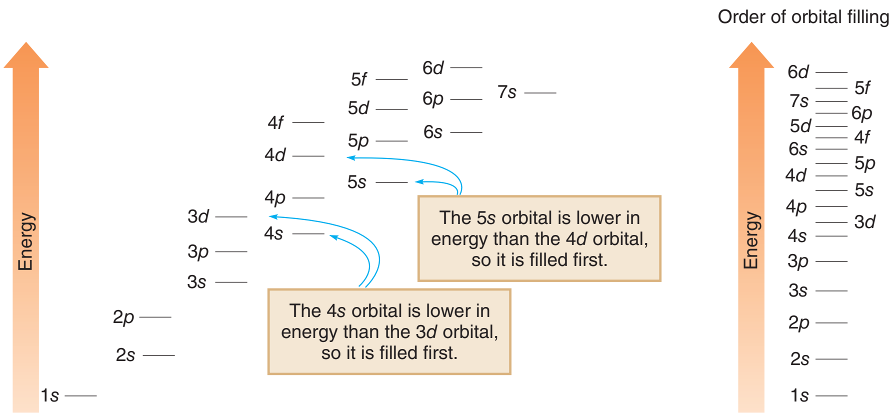
Rules to Determine the Ground State Electronic Configuration of an Atom

Rule [1] Electrons are placed in the lowest energy orbitals beginning with the 1*s* orbital.

- In comparing similar types of orbitals from one shell to another (e.g., 2*s* and 3*s*), an orbital closer to the nucleus is lower in energy. Thus, the energy of a 2*s* orbital is lower than a 3*s* orbital.
- Within a shell, orbital energies increase in the following order: *s, p, d, f.*

These guidelines result in the following order of energies in the first three periods: 1*s,* 2*s,* 2*p,* 3*s,* 3*p.* Above the 3*p* level, however, all orbitals of one shell do *not* have to be filled before any orbital in the next higher shell gets electrons. For example, a 4*s* orbital is lower in energy than a 3*d* orbital, so it is filled first. Figure 2.8 lists the relative energy of the orbitals used by atoms in the periodic table.

FIGURE 2.8 Relative Energies of Orbitals



Order of orbital filling

The 5*s* orbital is lower in energy than the 4*d* orbital, so it is filled first.

The 4*s* orbital is lower in energy than the 3*d* orbital, so it is filled first.

Electrons are added to orbitals in order of increasing energy. The 4*s* orbital is filled with electrons before the 3*d* orbital since it is lower in energy. The same is true for filling the 5*s* orbital with electrons before the 4*d* orbital. Likewise, the 6*s* orbital is filled before both the 4*f* and 5*d* orbitals, and the 7*s* orbital is filled before both the 5*f* and 6*d* orbitals.

Rule [2] Each orbital holds a maximum of two electrons.

Rule [3] When orbitals are equal in energy, one electron is added to each orbital until the orbitals are half-filled, before any orbital is completely filled.

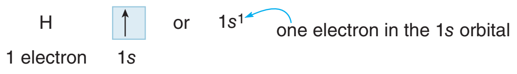
- For example, one electron is added to each of the three *p* orbitals before filling any *p* orbital with two electrons.
- Because like charges repel each other (Section 2.2), adding electrons to *different* *p* orbitals keeps them farther away from each other, which is energetically favorable.

To illustrate how these rules are used, we can write the electronic configuration for several elements using **orbital diagrams.** An orbital diagram uses a box to represent each orbital and arrows to represent electrons. A single electron, called an **unpaired electron,** is shown with a single arrow pointing up (↑). Two electrons in an orbital have **paired spins**—that is, the spins are opposite in direction—so up and down arrows (↑↓) are used.

2.6A FIRST-ROW ELEMENTS (PERIOD 1)

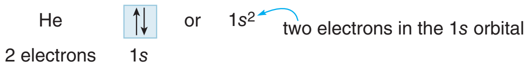
The first row of the periodic table contains only two elements—hydrogen and helium. Since the number of protons in the nucleus equals the number of electrons in a neutral atom, the atomic **number tells us how many electrons must be placed in orbitals.**

Hydrogen (H, $Z = 1$) has one electron. In the ground state, this electron is added to the lowest energy orbital, the 1*s* orbital. To draw an orbital diagram we use one box to represent the 1*s* orbital, and one up arrow to represent the electron. We can also write out the electron configuration without boxes and arrows, using a superscript with each orbital to show how many electrons it contains.



one electron in the 1*s* orbital

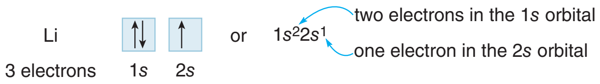
Helium (He, $Z = 2$) has two electrons. In the ground state, both electrons are added to the 1*s* orbital. To draw an orbital diagram we use one box to represent the 1*s* orbital, and a set of up and down arrows to represent the two electrons with paired spins. The electron configuration can also be written as $1s^2$, meaning the 1*s* orbital has two electrons. Helium has a filled first shell of electrons.



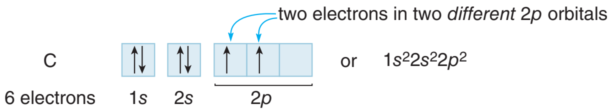
2.6B SECOND-ROW ELEMENTS (PERIOD 2)

To write orbital diagrams for the second-row elements, we must now use the four orbitals in the second shell—the 2*s* orbital and the three 2*p* orbitals. Since electrons are always added to the lowest energy orbitals first, all second-row elements have the 1*s* orbital filled with electrons, and then the remaining electrons are added to the orbitals in the second shell. Since the 2*s* orbital is lower in energy than the 2*p* orbitals, it is completely filled before adding electrons to the 2*p* orbitals.

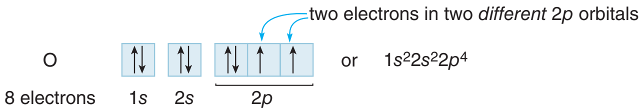
Lithium (Li, $Z = 3$) has three electrons. In the ground state, two electrons are added to the 1*s* orbital and the remaining electron is an unpaired electron in the 2*s* orbital. Lithium's electronic configuration can also be written as $1s^22s^1$ to show the placement of its three electrons.



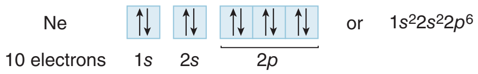
Carbon (C, $Z = 6$) has six electrons. In the ground state, two electrons are added to both the 1*s* and 2*s* orbitals. The two remaining electrons are added to two different 2*p* orbitals, giving carbon **two unpaired electrons.** These electrons spin in the same direction, so the arrows used to represent them are drawn in the same direction as well (both ↑ in this case). Carbon's electronic configuration is also written as $1s^22s^22p^2$. This method of writing an electronic configuration indicates that carbon has two electrons in its 2*p* orbitals, but it does not explicitly show that the two 2*p* electrons occupy *different* 2*p* orbitals.



Oxygen (O, $Z = 8$) has eight electrons. In the ground state, two electrons are added to both the 1*s* and 2*s* orbitals. The remaining four electrons must be distributed among the three 2*p* orbitals to give the lowest energy arrangement. This is done by pairing two electrons in one 2*p* orbital, and giving the remaining 2*p* orbitals one electron each. Oxygen has two unpaired electrons.



Neon (Ne, $Z = 10$) has 10 electrons. In the ground state, two electrons are added to the 1*s*, 2*s*, and each of the three 2*p* orbitals, so that the second shell of orbitals is now completely filled with electrons.



Sometimes the electronic configuration of an element is shortened by using the name of the noble gas that has a filled shell of electrons from the preceding row, and then adding the electronic configuration of all remaining electrons using orbitals and superscripts. For example, each

TABLE 2.5 Electronic Configurations of the First- and Second-Row Elements

		Orbital Diagram						
Atomic Number	Element	1*s*	2*s*	2*p*			Electronic Configuration	Noble Gas Notation
1	H	↑					$1s^1$	
2	He	↑↓					$1s^2$	
3	Li	↑↓	↑				$1s^22s^1$	$[\text{He}]\ 2s^1$
4	Be	↑↓	↑↓				$1s^22s^2$	$[\text{He}]\ 2s^2$
5	B	↑↓	↑↓	↑			$1s^22s^22p^1$	$[\text{He}]\ 2s^22p^1$
6	C	↑↓	↑↓	↑	↑		$1s^22s^22p^2$	$[\text{He}]\ 2s^22p^2$
7	N	↑↓	↑↓	↑	↑	↑	$1s^22s^22p^3$	$[\text{He}]\ 2s^22p^3$
8	O	↑↓	↑↓	↑↓	↑	↑	$1s^22s^22p^4$	$[\text{He}]\ 2s^22p^4$
9	F	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^5$	$[\text{He}]\ 2s^22p^5$
10	Ne	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^22s^22p^6$	$[\text{He}]\ 2s^22p^6$

second-row element has a $1s^2$ configuration like the noble gas helium in the preceding row, so the electronic configuration for carbon can be shortened to $[\text{He}]2s^22p^2$.

Electronic configuration for helium: [He] = $1s^2$

For carbon: $1s^22s^22p^2$ $[\text{He}]2s^22p^2$

replace

carbon's electronic configuration using noble gas notation

The electronic configurations of all the first- and second-row elements are listed in Table 2.5.

PROBLEM 2.20

What element has each electronic configuration?

a. $1s^22s^22p^63s^23p^2$

b. $[\text{Ne}]3s^23p^4$

c. $1s^22s^22p^63s^23p^64s^23d^1$

d. $[\text{Ar}]4s^23d^{10}$

2.6C OTHER ELEMENTS

Orbital diagrams and electronic configurations can be written in much the same way for every element in the periodic table. Sample Problems 2.7 and 2.8 illustrate two examples.

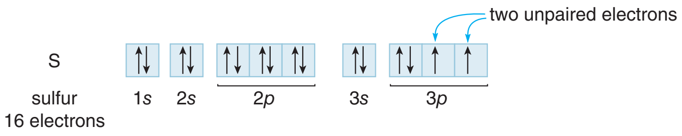
SAMPLE PROBLEM 2.7

Give the orbital diagram for the ground state electronic configuration of the element sulfur. Then, convert this orbital diagram to noble gas notation.

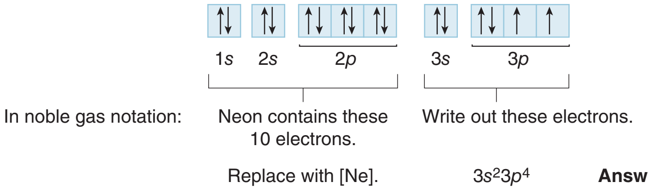
ANALYSIS

- Use the atomic number to determine the number of electrons.
- Place electrons two at a time into the lowest energy orbitals, following the order of orbital filling in Figure 2.8. When orbitals have the same energy, place electrons one at a time in the orbitals until they are half-filled.
- To convert an orbital diagram to noble gas notation, replace the electronic configuration corresponding to the noble gas in the preceding row by the symbol for the noble gas in brackets.

SOLUTION The atomic number of sulfur is 16, so 16 electrons must be placed in orbitals. Twelve electrons are added in pairs to the 1*s*, 2*s*, three 2*p*, and 3*s* orbitals. The remaining four electrons are then added to the three 3*p* orbitals to give two paired electrons and two unpaired electrons.



Since sulfur is in the third period, use the noble gas neon in the preceding row to write the electronic configuration in noble gas notation. **Substitute [Ne] for all of the electrons in the first and second shells.**



Answer: [Ne]$3s^23p^4$

ENVIRONMENTAL NOTE

Coal that is high in **sulfur** content burns to form sulfur oxides, which in turn react with water to form sulfurous and sulfuric acids. Rain that contains these acids has destroyed acres of forests worldwide.

SAMPLE PROBLEM 2.8

Give the ground state electronic configuration of the element calcium. Convert the electronic configuration to noble gas notation.

ANALYSIS

- Use the atomic number to determine the number of electrons.
- Place electrons two at a time into the lowest energy orbitals, following the order of orbital filling in Figure 2.8.
- To convert the electronic configuration to noble gas notation, replace the electronic configuration corresponding to the noble gas in the preceding row by the symbol for the noble gas in brackets.

SOLUTION The atomic number of calcium is 20, so 20 electrons must be placed in orbitals. Eighteen electrons are added in pairs to the 1*s*, 2*s*, three 2*p*, 3*s*, and three 3*p* orbitals. Figure 2.8 shows that the 4*s* orbital is next highest in energy, not the 3*d* orbitals, so the remaining two electrons are added to the 4*s* orbital. Since calcium is an element in period 4, use the noble gas argon in period 3 to write the noble gas configuration.

$$\underset{\text{(20 electrons)}}{\text{Electronic configuration for Ca}} = \underbrace{1s^22s^22p^63s^23p^6}_{\substack{\text{The noble gas argon contains}\\ \text{these 18 electrons.}\\ \text{Replace with [Ar].}}}4s^2 = \underset{\text{noble gas notation}}{[\text{Ar}]4s^2}$$

PROBLEM 2.21

Draw an orbital diagram for each element: (a) magnesium; (b) aluminum; (c) bromine.

PROBLEM 2.22

Give the electronic configuration for each element and then convert it to noble gas notation: (a) sodium; (b) silicon; (c) iodine.

2.7 ELECTRONIC CONFIGURATIONS AND THE PERIODIC TABLE

Having learned how electrons are arranged in the orbitals of an atom, we can now understand more about the structure of the periodic table. Considering electronic configuration, the periodic table can be divided into four regions, called **blocks,** labeled ***s*, *p*, *d*,** and ***f*,** and illustrated in

FIGURE 2.9 The Blocks of Elements in the Periodic Table

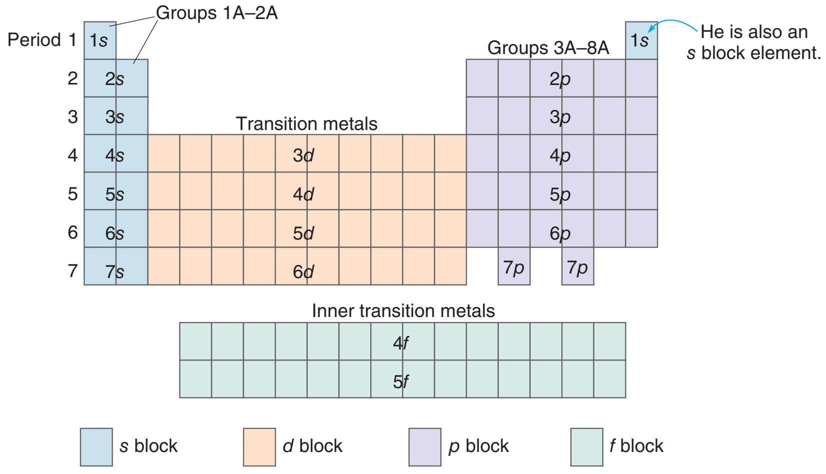


Figure 2.9. **The blocks are labeled according to the subshells that are filled with electrons *last*.**

- The *s block* consists of groups 1A and 2A and the element helium. The *s* subshell is filled last in these elements.
- The *p block* consists of groups 3A–8A (except helium). The *p* subshell is filled last in these elements.
- The *d block* consists of the 10 columns of transition metals. The *d* subshell is filled last in these elements.
- The *f block* consists of the two groups of 14 inner transition metals. The *f* subshell is filled last in these elements.

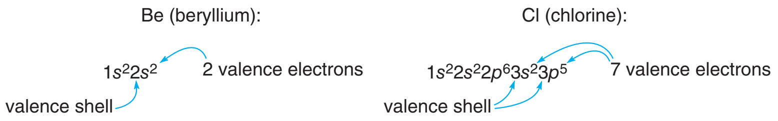
The number of electrons that can fill a given subshell determines the number of columns in a block. Since each shell contains only one *s* orbital, which can hold two electrons, the *s* block is composed of two columns, one that results from adding one electron to an *s* orbital, and one that results from adding two. Similarly, because a shell has three *p* orbitals that can hold two electrons each, there are six columns in the *p* block. The 10 columns of the *d* block result from adding up to 10 electrons to five *d* orbitals. The 14 columns of the *f* block result from adding up to 14 electrons to the seven *f* orbitals.

2.7A VALENCE ELECTRONS

The chemical properties of an element depend on the most loosely held electrons—that is, those electrons in the outermost shell, called the **valence shell.** The period number tells the number of the valence shell.

- The electrons in the outermost shell are called the *valence electrons.*

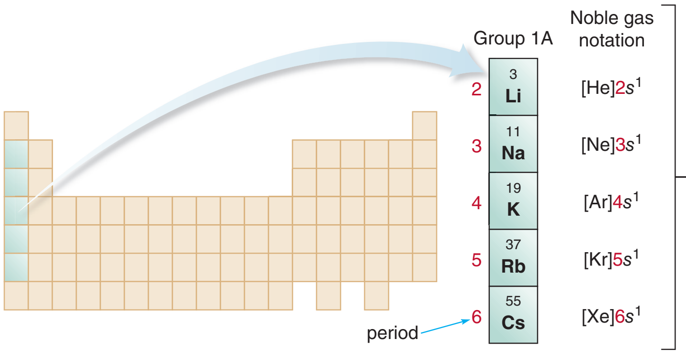
To identify the electrons in the valence shell, always look for the shell with the *highest* number. Thus, beryllium has two valence electrons that occupy the 2*s* orbital. Chlorine has seven valence electrons since it has a total of seven electrons in the third shell, two in the 3*s* orbital and five in the 3*p* orbitals.



If we examine the electronic configuration of a group in the periodic table, two facts become apparent.

- **Elements in the same group have the same number of valence electrons and similar electronic configurations.**
- **The group number (using the 1A–8A system) equals the number of valence electrons for main group elements (except helium).**

As an example, the alkali metals in group 1A all have one valence electron that occupies an *s* orbital. Thus, a general electronic configuration for the valence electrons of an alkali metal is ns^1, where n = the period in which the element is located.



- Each element in group 1A has one electron in an *s* orbital.
- The period number indicates the valence shell.

Thus, the periodic table is organized into groups of elements with similar valence electronic configurations in the same column. The valence electronic configurations of the main group elements in the first three rows of the periodic table are given in Table 2.6.

- **The chemical properties of a group are similar because these elements contain the same electronic configuration of valence electrons.**

Take particular note of the electronic configuration of the noble gases in group 8A. **All of these elements have a completely filled outer shell of valence electrons.** Helium has a filled first

TABLE 2.6 Valence Electronic Configurations for the Main Group Elements in Periods 1–3

Group Number	1A	2A	3A	4A	5A	6A	7A	8A[a]
Period 1	H $1s^1$							He $1s^2$
Period 2	Li $2s^1$	Be $2s^2$	B $2s^22p^1$	C $2s^22p^2$	N $2s^22p^3$	O $2s^22p^4$	F $2s^22p^5$	Ne $2s^22p^6$
Period 3	Na $3s^1$	Mg $3s^2$	Al $3s^23p^1$	Si $3s^23p^2$	P $3s^23p^3$	S $3s^23p^4$	Cl $3s^23p^5$	Ar $3s^23p^6$
General configuration	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6

[a] The general electronic configuration in group 8A applies to all of the noble gases except helium. Since helium is a first-row element, it has only two electrons, and these occupy the only available orbital in the first shell, the 1*s* orbital.

shell ($1s^2$ configuration). The remaining elements have a completely filled valence shell of s and p orbitals (s^2p^6). This electronic arrangement is especially stable, and as a result, these elements exist in nature as single atoms. We will learn about the consequences of having a completely filled valence shell in Chapter 3.

SAMPLE PROBLEM 2.9 Identify the total number of electrons, the number of valence electrons, and the name of the element with each electronic configuration.

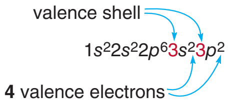
a. $1s^22s^22p^63s^23p^2$

b. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}$

ANALYSIS To obtain the total number of electrons, add up the superscripts. This gives the atomic number and identifies the element. To determine the number of valence electrons, add up the number of electrons in the shell with the highest number.

SOLUTION

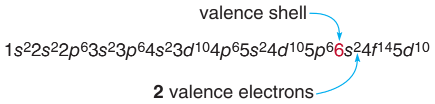
a.



Total number of electrons = 2 + 2 + 6 + 2 + 2 = **14**

Answer: Silicon (Si), 14 total electrons and 4 valence electrons

b.



Total number of electrons = **80**

Answer: Mercury (Hg), 80 total electrons and 2 valence electrons

PROBLEM 2.23 Identify the total number of electrons, the number of valence electrons, and the name of the element with each electronic configuration.

a. $1s^22s^22p^63s^2$

b. $1s^22s^22p^63s^23p^3$

c. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^2$

d. $[Ar]4s^23d^6$

SAMPLE PROBLEM 2.10 Determine the number of valence electrons and give the electronic configuration of the valence electrons of each element: (a) nitrogen; (b) potassium.

ANALYSIS The group number of a main group element = the number of valence electrons. Use the general electronic configurations in Table 2.6 to write the configuration of the valence electrons.

SOLUTION

a. Nitrogen is located in group 5A so it has five valence electrons. Since nitrogen is a second-period element, its valence electronic configuration is $2s^22p^3$.

b. Potassium is located in group 1A so it has one valence electron. Since potassium is a fourth-period element, its valence electronic configuration is $4s^1$.

PROBLEM 2.24 Determine the number of valence electrons and give the electronic configuration of the valence electrons of each element: (a) fluorine; (b) krypton; (c) magnesium; (d) germanium.

PROBLEM 2.25 Write the valence shell electronic configuration for the elements in periods 4, 5, and 6 of group 6A.

2.7B ELECTRON-DOT SYMBOLS

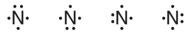
The number of valence electrons around an atom is often represented by an **electron-dot symbol.** Representative examples are shown.

	H	C	O	Cl
Number of valence electrons:	1	4	6	7

Electron-dot symbol:

- **Each dot represents one electron.**
- **The dots are placed on the four sides of an element symbol.**
- **For one to four valence electrons, single dots are used. With more than four electrons, the dots are paired.**

The location of the dots around the symbol—side, top, or bottom—does not matter. Each of the following representations for the five valence electrons of nitrogen is equivalent.



SAMPLE PROBLEM 2.11 Write an electron-dot symbol for each element: (a) sodium; (b) phosphorus.

ANALYSIS Write the symbol for each element and use the group number to determine the number of valence electrons for a main group element. Represent each valence electron with a dot.

SOLUTION

a. The symbol for sodium is Na. Na is in group 1A and has one valence electron. Electron-dot symbol:

Na·

b. The symbol for phosphorus is P. P is in group 5A and has five valence electrons. Electron-dot symbol:

PROBLEM 2.26 Give the electron-dot symbol for each element: (a) bromine; (b) lithium; (c) aluminum; (d) sulfur; (e) neon.

2.8 PERIODIC TRENDS

Many properties of atoms exhibit **periodic trends;** that is, they change in a regular way across a row or down a column of the periodic table. Two properties that illustrate this phenomenon are **atomic size** and **ionization energy.**

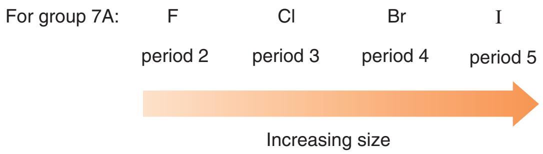
HEALTH NOTE

Mercury (Sample Problem 2.9) is safely used in dental amalgam to fill cavities in teeth. Mercury released into the environment, however, is converted to toxic methylmercury by microorganisms in water, so hazardous levels of this soluble mercury compound can accumulate in fish at the top of the food chain, such as sharks and swordfish.

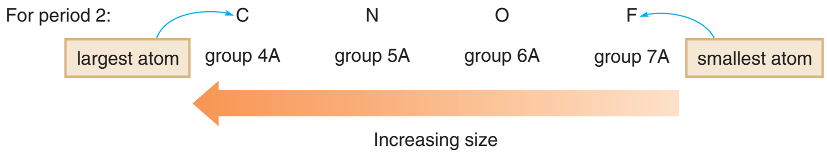
2.8A ATOMIC SIZE

The size of an atom is measured by its atomic radius—that is, the distance from the nucleus to the outer edge of the valence shell. Two periodic trends characterize the size of atoms.

- **The size of atoms increases down a column of the periodic table, as the valence electrons are farther from the nucleus.**



- **The size of atoms decreases across a row of the periodic table as the number of protons in the nucleus increases. An increasing number of protons pulls the electrons closer to the nucleus, so the atom gets smaller.**



PROBLEM 2.27 Rank the atoms in each group in order of increasing size.

a. boron, carbon, neon
b. calcium, magnesium, beryllium
c. silicon, sulfur, magnesium
d. krypton, neon, xenon
e. sulfur, oxygen, silicon
f. fluorine, sulfur, aluminum

2.8B IONIZATION ENERGY

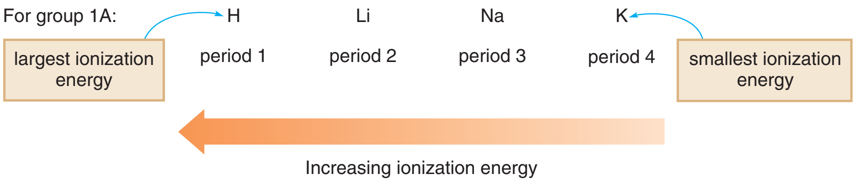
Since a negatively charged electron is attracted to a positively charged nucleus, energy is required to remove an electron from a neutral atom. The more tightly the electron is held, the greater the energy required to remove it. Removing an electron from a neutral atom forms a **cation.**

$$\underset{\text{neutral atom}}{\text{Na}} + \underset{\boxed{\text{ionization energy}}}{\text{energy}} \longrightarrow \underset{\boxed{\text{cation}}}{\text{Na}^+} + e^-$$

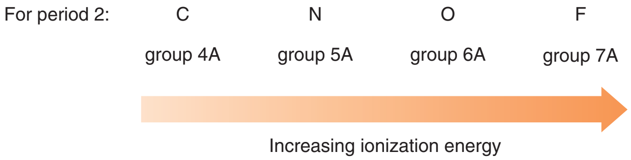
- **The *ionization energy* is the energy needed to remove an electron from a neutral atom.**
- **A *cation* is positively charged, and has fewer electrons than the neutral atom.**

Two periodic trends characterize ionization energy.

- **Ionization energies decrease down a column of the periodic table as the valence electrons get farther from the positively charged nucleus.**



- **Ionization energies generally increase across a row of the periodic table as the number of protons in the nucleus increases.**



PROBLEM 2.28 Arrange the elements in each group in order of increasing ionization energy.

a. phosphorus, silicon, sulfur
b. magnesium, calcium, beryllium
c. carbon, fluorine, beryllium
d. neon, krypton, argon
e. tin, silicon, sulfur
f. calcium, aluminum, nitrogen

CHAPTER HIGHLIGHTS

KEY TERMS

Actinide (2.4)
Alkali metal (2.4)
Alkaline earth element (2.4)
Atom (2.1)
Atomic mass unit (2.2)
Atomic number (2.2)
Atomic weight (2.3)
Building-block element (2.1)
Cation (2.8)
Chemical formula (2.1)
Compound (2.1)
d Block (2.7)
Deuterium (2.3)
Electron (2.2)
Electron cloud (2.2)
Electron-dot symbol (2.7)
Electronic configuration (2.6)
Element (2.1)
f Block (2.7)
Ground state (2.6)
Group (2.4)
Group number (2.4)
Halogen (2.4)
Inner transition metal element (2.4)
Ionization energy (2.8)
Isotope (2.3)
Lanthanide (2.4)
Main group element (2.4)
Major mineral (Macronutrient, 2.1)
Mass number (2.2)
Metal (2.1)
Metalloid (2.1)
Neutron (2.2)
Noble gas (2.4)
Nonmetal (2.1)
Nucleus (2.2)
Orbital (2.5)
p Block (2.7)
Period (2.4)
Periodic table (2.1)
p Orbital (2.5)
Proton (2.2)
s Block (2.7)
Shell (2.5)
s Orbital (2.5)
Subshell (2.5)
Trace element (Micronutrient, 2.1)
Transition metal element (2.4)
Tritium (2.3)
Unpaired electron (2.6)
Valence electron (2.7)

KEY CONCEPTS

❶ **How is the name of an element abbreviated and how does the periodic table help to classify it as a metal, nonmetal, or metalloid? (2.1)**

- An element is abbreviated by a one- or two-letter symbol. The periodic table contains a stepped line from boron to astatine. All metals are located to the left of the line. All nonmetals except hydrogen are located to the right of the line. The seven elements located along the line are metalloids.

❷ **What are the basic components of an atom? (2.2)**

- An atom is composed of two parts: a dense nucleus containing positively charged protons and neutral neutrons, and an electron cloud containing negatively charged electrons. Most of the mass of an atom resides in the nucleus, while the electron cloud contains most of its volume.
- The atomic number (Z) of a neutral atom tells the number of protons and the number of electrons. The mass number (A) is the sum of the number of protons (Z) and the number of neutrons.

❸ **What are isotopes and how are they related to the atomic weight? (2.3)**

- Isotopes are atoms that have the same number of protons but a different number of neutrons. The atomic weight is the weighted average of the mass of the naturally occurring isotopes of a particular element.

❹ **What are the basic features of the periodic table? (2.4)**

- The periodic table is a schematic of all known elements, arranged in rows (periods) and columns (groups), organized so that elements with similar properties are grouped together.
- The vertical columns are assigned group numbers using two different numbering schemes—1–8 plus the letters A or B; or 1–18.
- The periodic table is divided into the main group elements (groups 1A–8A), the transition metals (groups 1B–8B), and the inner transition metals located at the bottom.

❺ **How are electrons arranged around an atom? (2.5)**

- Electrons occupy discrete energy levels, organized into shells (numbered 1, 2, 3, and so on), subshells (*s, p, d,* and *f*), and orbitals.
- Each orbital can hold two electrons.

❻ **What rules determine the electronic configuration of an atom? (2.6)**

- To write the ground state electronic configuration of an atom, electrons are added to the lowest energy orbitals, giving each orbital two electrons. When two orbitals are equal in energy, one electron is added to each orbital until the orbitals are half-filled.
- Orbital diagrams that use boxes for orbitals and arrows for electrons indicate electronic configuration. Electron configuration can also be shown using superscripts to show how many electrons an orbital contains. For example, the electron configuration of the six electrons in a carbon atom is $1s^22s^22p^2$.

❼ How is the location of an element in the periodic table related to its electronic configuration? (2.7)

- The periodic table is divided into four regions—the *s* block, *p* block, *d* block, and *f* block—based on the subshells that are filled with electrons last.
- Elements in the same group have the same number of valence electrons and similar electronic configurations.

❽ What is an electron-dot symbol? (2.7)

- An electron-dot symbol uses a dot to represent each valence electron around the symbol for an element.

❾ How are atomic size and ionization energy related to location in the periodic table? (2.8)

- The size of an atom decreases across a row and increases down a column.
- Ionization energy—the energy needed to remove an electron from an atom—increases across a row and decreases down a column.

PROBLEMS

Selected in-chapter and end-of-chapter problems have brief answers provided in Appendix B.

Elements

2.29 Give the name of the elements in each group of three element symbols.

a. Au, At, Ag
b. N, Na, Ni
c. S, Si, Sn
d. Ca, Cr, Cl
e. P, Pb, Pt
f. Ti, Ta, Tl

2.30 What element(s) are designated by each symbol or group of symbols?

a. CU and Cu
b. Os and OS
c. Ni and NI
d. BIN, BiN, and BIn

2.31 Does each chemical formula represent an element or a compound?

a. H_2 b. H_2O_2 c. S_8 d. Na_2CO_3 e. C_{60}

2.32 Identify the elements in each chemical formula and tell how many atoms of each are present.

a. $K_2Cr_2O_7$
b. $C_5H_8NNaO_4$ (MSG, flavor enhancer)
c. $C_{10}H_{16}N_2O_3S$ (vitamin B_7)

2.33 Identify the element that fits each description.

a. an alkali metal in period 6
b. a transition metal in period 5, group 8
c. a main group element in period 3, group 7A
d. a main group element in period 2, group 2A
e. a halogen in period 2
f. an inner transition metal with one 4*f* electron

2.34 Identify the element that fits each description.

a. an alkaline earth element in period 3
b. a noble gas in period 6
c. a main group element in period 3 that has *p* orbitals half-filled with electrons
d. a transition metal in period 4, group 11
e. an inner transition metal with its 5*f* orbitals completely filled with electrons
f. a transition metal in period 6, group 10

2.35 Give all of the terms that apply to each element: [1] metal; [2] nonmetal; [3] metalloid; [4] alkali metal; [5] alkaline earth element; [6] halogen; [7] noble gas; [8] main group element; [9] transition metal; [10] inner transition metal.

a. sodium
b. silver
c. xenon
d. platinum
e. uranium
f. tellurium

2.36 Give all of the terms that apply to each element: [1] metal; [2] nonmetal; [3] metalloid; [4] alkali metal; [5] alkaline earth element; [6] halogen; [7] noble gas; [8] main group element; [9] transition metal; [10] inner transition metal.

a. bromine
b. silicon
c. cesium
d. gold
e. calcium
f. chromium

Atomic Structure

2.37 Complete the following table for neutral elements.

	Element Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
a.	C		12			
b.			31			15
c.					35	30
d.	Mg		24			
e.				53	74	
f.		4			5	
g.		40	91			
h.					16	16

2.38 For the given atomic number (*Z*) and mass number (*A*): [1] identify the element; [2] give the element symbol; [3] give the number of protons, neutrons, and electrons.

a. $Z = 10, A = 20$
b. $Z = 13, A = 27$
c. $Z = 38, A = 88$
d. $Z = 55, A = 133$
e. $Z = 28, A = 59$
f. $Z = 79, A = 197$

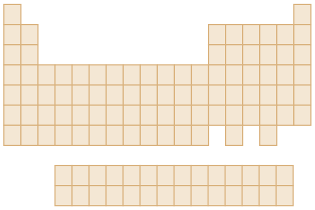
2.39 Convert the mass of a proton (1.6726×10^{-24} g) to a standard number.

2.40 Convert the mass of an electron (9.1093×10^{-28} g) to a standard number.

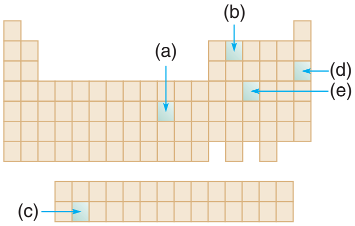
Periodic Table

2.41 Label each region on the periodic table.

a. noble gases	e. alkaline earth elements
b. period 3	f. *f* block elements
c. group 4A	g. transition metals
d. *s* block elements	h. group 10



2.42 Identify each highlighted element in the periodic table and give its [1] element name and symbol; [2] group number; [3] period; [4] classification (i.e., main group element, transition metal, or inner transition metal).



2.43 What element is located in group 1A but is not an alkali metal?

2.44 What *s* block element is not located in either group 1A or group 2A in the periodic table?

2.45 Name two elements in the periodic table that have chemical properties similar to carbon.

2.46 Name two elements in the periodic table that have chemical properties similar to calcium.

2.47 Classify each element in the fourth row of the periodic table as a metal, nonmetal, or metalloid.

2.48 To which blocks do the elements in the fifth period belong?

2.49 Which group(s) in the periodic table contain only nonmetals?

2.50 Which groups in the periodic table contain metals, nonmetals, and metalloids?

Isotopes and Atomic Weight

2.51 The most common isotope of oxygen has a mass number of 16, but two other isotopes having mass numbers of 17 and 18 are also known. For each isotope, give the following information: (a) the number of protons; (b) the number of neutrons; (c) the number of electrons in the neutral atom; (d) the group number; (e) the element symbols using superscripts and subscripts.

2.52 The three most common isotopes of tin have mass numbers 116, 118, and 120. For each isotope, give the following information: (a) the number of protons; (b) the number of neutrons; (c) the number of electrons in the neutral atom; (d) the group number; (e) the element symbols using superscripts and subscripts.

2.53 How many protons, neutrons, and electrons are contained in each element?

a. $^{27}_{13}\text{Al}$ b. $^{35}_{17}\text{Cl}$ c. $^{34}_{16}\text{S}$

2.54 Give the number of protons, neutrons, and electrons in each element: (a) silver-115; (b) Au-197; (c) Rn-222; (d) osmium-192.

2.55 Write the element symbol that fits each description, using a superscript for the mass number and a subscript for the atomic number.

a. an element that contains 53 protons and 74 neutrons
b. an element with 35 electrons and a mass number of 79
c. an element with 47 protons and 60 neutrons

2.56 Write the element symbol that fits each description. Use a superscript for the mass number and a subscript for the atomic number.

a. an element that contains 10 protons and 12 neutrons
b. an element with atomic number 24 and mass number 52
c. an element with 10 electrons and 10 neutrons

2.57 Calculate the atomic weight of silver, which has two isotopes with the following properties: Ag-107 (106.91 amu, 51.84% natural occurrence) and Ag-109 (108.90 amu, 48.16% natural occurrence).

2.58 Calculate the atomic weight of antimony, which has two isotopes with the following properties: Sb-121 (120.90 amu, 57.21% natural occurrence) and Sb-123 (122.90 amu, 42.79% natural occurrence).

2.59 Can the neutral atoms of two different elements have the same number of electrons? Explain.

2.60 Can the neutral atoms of two different elements have the same number of neutrons? Explain.

Electronic Configuration

2.61 What is the difference between a shell and a subshell of electrons?

2.62 What is the difference between a subshell and an orbital of electrons?

2.63 What is the difference between a 1*s* and 2*s* orbital?

2.64 What is the difference between a 2*s* and 2*p* orbital?

2.65 How many orbitals are contained in each of the following shells of electrons: (a) first shell ($n = 1$); (b) second shell ($n = 2$); (c) third shell ($n = 3$); (d) fourth shell ($n = 4$)?

2.66 What is the maximum number of electrons that can be contained in each shell, subshell, or orbital?

a. second shell	c. 3*p* subshell	e. fourth shell
b. 3*s* orbital	d. 4*f* orbital	f. 5*p* orbital

2.67 Why are there 10 columns of transition metal elements in the periodic table?

2.68 Why are there six columns of *p* block elements in the periodic table?

2.69 Write the electronic configuration of each element using an orbital diagram: (a) B; (b) K; (c) Se; (d) Ar; (e) Zn.

2.70 Write the electronic configuration of each element using an orbital diagram: (a) N; (b) I; (c) Ga; (d) Ti; (e) Mn.

2.71 For each element in Problem 2.69: (a) Write out the electronic configuration using a superscript with each orbital; (b) write out the electronic configuration using noble gas notation.

2.72 For each element in Problem 2.70: (a) Write out the electronic configuration using a superscript with each orbital; (b) write out the electronic configuration using noble gas notation.

2.73 How many unpaired electrons are contained in each element: (a) Al; (b) P; (c) Na?

2.74 How many unpaired electrons are contained in each element: (a) Cl; (b) Se; (c) Cs?

2.75 Give the total number of electrons, the number of valence electrons, and the identity of the element with each electronic configuration.

a. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$
b. $1s^22s^22p^63s^23p^4$
c. $1s^22s^22p^63s^1$
d. $[\text{Ne}]3s^23p^5$

2.76 Give the total number of electrons, the number of valence electrons, and the identity of the element with each electronic configuration.

a. $1s^22s^22p^63s^23p^6$
b. $1s^22s^22p^63s^23p^64s^23d^7$
c. $1s^22s^22p^3$
d. $[\text{Kr}]5s^24d^{10}5p^2$

2.77 How do an alkali metal and an alkaline earth element in the same row differ in the electronic configuration of the valence shell electrons?

2.78 How do a halogen and a noble gas in the same row differ in the electronic configuration of the valence shell electrons?

2.79 For each element, give the following information: [1] total number of electrons; [2] group number; [3] number of valence electrons; [4] period; [5] number of the valence shell.

a. carbon b. calcium c. krypton

2.80 For each element, give the following information: [1] total number of electrons; [2] group number; [3] number of valence electrons; [4] period; [5] number of the valence shell.

a. oxygen b. sodium c. phosphorus

2.81 For each element in Problem 2.79, first write the electronic configuration of all the electrons. Then give the electronic configuration of the valence electrons only.

2.82 For each element in Problem 2.80, first write the electronic configuration of all the electrons. Then give the electronic configuration of the valence electrons only.

2.83 Give the number of valence electrons for each element in Problem 2.37.

2.84 Are the valence electrons always written last when an electronic configuration is written? Explain.

2.85 How many valence electrons does an element in each group contain: (a) 2A; (b) 4A; (c) 7A?

2.86 In what shell do the valence electrons reside for an element in period: (a) 2; (b) 3; (c) 4; (d) 5?

2.87 Can a *d* block element have valence electrons that occupy an *s* orbital? Explain.

2.88 Can an *f* block element have valence electrons that occupy an *s* orbital? Explain.

2.89 Give the number of valence electrons in each element. Write out the electronic configuration for the valence electrons.

a. sulfur b. chlorine c. barium d. titanium e. tin

2.90 Give the number of valence electrons in each element. Write out the electronic configuration for the valence electrons.

a. neon
b. rubidium
c. aluminum
d. manganese
e. zirconium

2.91 Write an electron-dot symbol for each element: (a) beryllium; (b) silicon; (c) iodine; (d) magnesium; (e) argon.

2.92 Write an electron-dot symbol for each element: (a) K; (b) B; (c) F; (d) Ca; (e) Se.

Periodic Trends

2.93 Which element in each pair is larger?

a. bromine and iodine
b. carbon and nitrogen
c. silicon and potassium
d. chlorine and selenium

2.94 Which element in each pair has its valence electrons farther from the nucleus?

a. sodium and magnesium
b. carbon and fluorine
c. neon and krypton
d. argon and bromine

2.95 For each pair of elements in Problem 2.93, label the element with the higher ionization energy.

2.96 For each pair of elements in Problem 2.94, label the element from which it is easier to remove an electron.

2.97 Rank the following elements in order of increasing size: sulfur, silicon, oxygen, magnesium, and fluorine.

2.98 Rank the following elements in order of increasing size: aluminum, nitrogen, potassium, oxygen, and phosphorus.

2.99 Rank the following elements in order of increasing ionization energy: nitrogen, fluorine, magnesium, sodium, and phosphorus.

2.100 Rank the following elements in order of decreasing ionization energy: calcium, silicon, oxygen, magnesium, and carbon.

Applications

2.101 Sesame seeds, sunflower seeds, and peanuts are good dietary sources of the trace element copper. Copper is needed for the synthesis of neurotransmitters, compounds that transmit nerve signals from one nerve cell to another. Copper is also needed for the synthesis of collagen, a protein found in bone, tendons, teeth, and blood vessels. What is unusual about the electronic configuration of the trace element copper: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$?

2.102 Platinum is a precious metal used in a wide variety of products. Besides fine jewelry, platinum is also the catalyst found in the catalytic converters of automobile exhaust systems, and platinum-containing drugs like cisplatin are used to treat some lung and ovarian cancers. Answer the following questions about the element platinum.
  a. What is its element symbol?
  b. What group number and period are assigned to platinum?
  c. What is its atomic number?
  d. Is platinum classified as a main group element, transition metal, or inner transition metal?
  e. In what block does platinum reside?
  f. What is unusual about the electronic configuration for platinum: $[Xe]6s^1 4f^{14} 5d^9$?

2.103 Answer the following questions about the macronutrients sodium, potassium, and chlorine.
  a. Is each element classified as a metal, nonmetal, or metalloid?
  b. In which block does each element reside?
  c. Which element has the smallest atomic radius?
  d. Which element has the largest atomic radius?
  e. Which element has the largest ionization energy?
  f. Which element has the smallest ionization energy?
  g. How many valence electrons does each element possess?

2.104 Answer the following questions about the macronutrients calcium, magnesium, and sulfur.
  a. Is each element classified as a metal, nonmetal, or metalloid?
  b. In which block does each element reside?
  c. Which element has the smallest atomic radius?
  d. Which element has the largest atomic radius?
  e. Which element has the largest ionization energy?
  f. Which element has the smallest ionization energy?
  g. How many valence electrons does each element possess?

2.105 Carbon-11 is an unnatural isotope used in positron emission tomography (PET) scans. PET scans are used to monitor brain activity and diagnose dementia. How does carbon-11 compare to carbon-12 in terms of the number of protons, neutrons, and electrons? Write the element symbol of carbon-11 using superscripts and subscripts.

2.106 Strontium-90 is a radioactive element formed in nuclear reactors. When an unusually high level of strontium is released into the air, such as occurred during the Chernobyl nuclear disaster in 1986, the strontium can be incorporated into the bones of exposed individuals. High levels of strontium can cause bone cancer and leukemia. Why does Sr-90 cause this particular health problem? (Hint: What macronutrient has similar chemical properties to strontium?)