Elements 2

> *"Nothing exists except atoms and empty space; everything else is opinion."* —
> Democritus (460–370 B.C.)

MODULE OUTLINE

2.1 The Nuclear Theory of the Atom 19
2.2 Elements: Defined by Their Numbers of Protons 21
2.3 Isotopes: When the Number of Neutrons Varies 24
2.4 Atomic Mass: The Average Mass of an Element's Atoms 26
2.5 Ions: Losing and Gaining Electrons 29
2.6 Arrangements of Electrons Around the Nucleus 32
2.7 Looking for Patterns: The Periodic Law and the Periodic Table 38
2.8 Electron Configurations and the Periodic Table 43
2.9 The Explanatory Power of the Quantum-Mechanical Model 46
2.10 Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character 48

2.1 The Nuclear Theory of the Atom

LO: Describe the respective properties and charges of electrons, neutrons, and protons.

The **nuclear theory of the atom** has three basic parts:

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

In addition to the points above that are the key components of the nuclear theory of the atom, note the following:

The nucleus may also contain uncharged particles called *neutrons*.

Atoms combine in simple, whole-number ratios to form compounds.

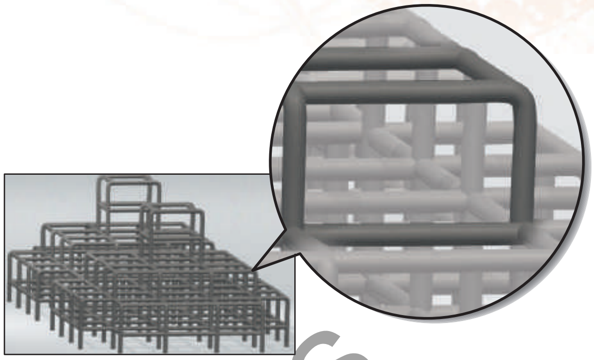
If the nucleus of the atom were the size of this dot ·, the average electron would be about 10 m away. Yet the dot would contain almost the entire mass of the atom. What if matter were composed of atomic nuclei piled on top of each other like marbles? Such matter would be incredibly dense; a single grain of sand composed of solid atomic nuclei would have a mass of 5 million kg. Astronomers believe that black holes and neutron stars are composed of this kind of incredibly dense matter.

Protons and neutrons have very similar masses. In SI units, the mass of the proton is 1.67262×10^{-27} kg, and the mass of the neutron is a close 1.67493×10^{-27} kg. A more common unit to express these masses, however, is the **atomic mass unit (amu)**, defined as one-twelfth of the mass of a carbon atom containing 6 protons and 6 neutrons. A proton has a mass of 1.0073 amu, and a neutron has a mass of 1.0087 amu. Electrons, by contrast, have an almost negligible mass of 0.00091×10^{-27} kg, or approximately 0.00055 amu.

EVERYDAY CHEMISTRY

Solid Matter?

If matter really is mostly empty space, then why does it appear so solid? Why can you tap your knuckles on the table and feel a solid thump? Matter appears solid because the variation in the density is on such a small scale that our eyes can't see it. Imagine a jungle gym 100 stories high and the size of a football field. It is mostly empty space. Yet if you were to view it from an airplane, it would appear as a solid mass. Matter is similar. When you tap your knuckles on the table, it is much like one giant jungle gym (your finger) crashing into another (the table). Even though they are both primarily empty space, one does not fall into the other.



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

◀ If a proton had the mass of a baseball, an electron would have the mass of a rice grain. The proton is nearly 2000 times as massive as an electron. © Maxwell Art And Photo/Pearson.

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

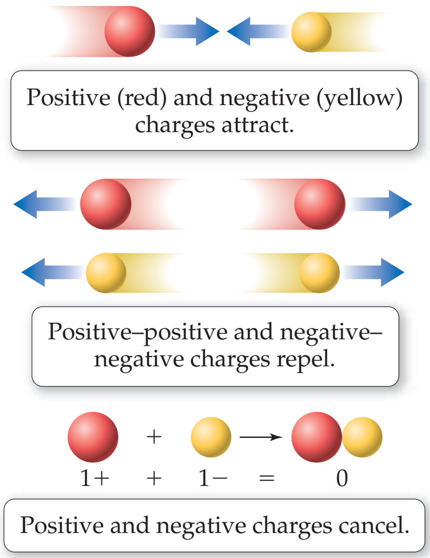
What is electrical charge? Electrical charge is a fundamental property of protons and electrons, just as mass is a fundamental property of matter. Most matter is charge-neutral because protons and electrons occur together and their charges cancel. However, you may have experienced excess electrical charge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

We can summarize the nature of electrical charge as follows (◀ Figure 2.1).

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

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

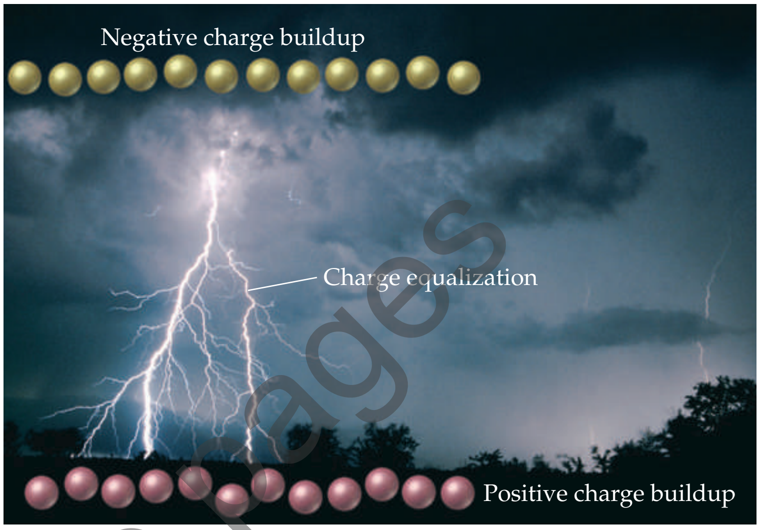
If you had a sample of matter—even a tiny sample, such as a sand grain—that was composed of only protons or only electrons, the forces around that matter would be extraordinary, and the matter would be unstable. Fortunately, matter is not that way—protons and electrons exist together, canceling



◀ **FIGURE 2.1 The properties of electrical charge** each other's charge and making matter charge-neutral. Table 2.1 summarizes the properties of protons, neutrons, and electrons (subatomic particles).

TABLE 2.1 Subatomic Particles

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



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

CONCEPTUAL CHECKPOINT 2.1

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

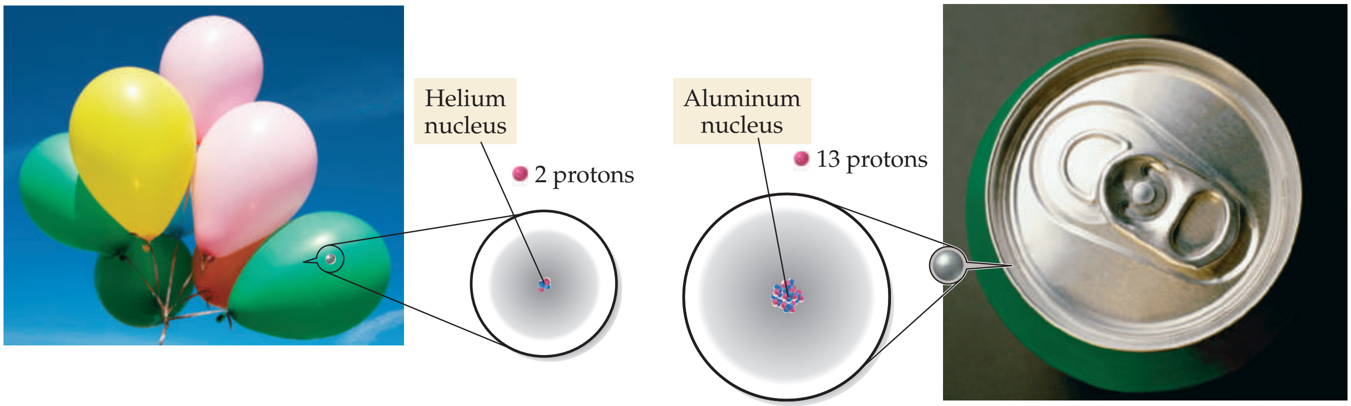
(a) 6 protons and 6 electrons
(b) 3 protons, 3 neutrons, and 6 electrons
(c) 6 protons, 6 neutrons, and 6 electrons
(d) 12 neutrons and 12 electrons

2.2 Elements: Defined by Their Numbers of Protons

LO: Determine the atomic symbol and atomic number for an element using the periodic table.

We have seen that atoms are composed of protons, neutrons, and electrons. However, it is the number of protons in the nucleus of an atom that identifies it as a particular element. For example, atoms with 2 protons in their nucleus are helium atoms, atoms with 13 protons in their nucleus are aluminum atoms, and atoms with 92 protons in their nucleus are uranium atoms. The number of protons in an atom's nucleus defines the element (▶ Figure 2.2). Every aluminum atom has 13 protons in its nucleus; if it had a different number of protons, it would be a different element. The number of protons in the nucleus of an atom is its **atomic number** and is represented with the symbol Z.

The periodic table of the elements (▶ Figure 2.3) lists all known elements according to their atomic numbers. Each element is represented by a unique **chemical symbol**, a one- or two-letter abbreviation for the element that appears directly below its atomic number on the periodic table. The chemical symbol for helium is He; for aluminum, Al; and for uranium, U. The chemical symbol and the atomic number always go together. If the atomic number is 13, the chemical symbol must

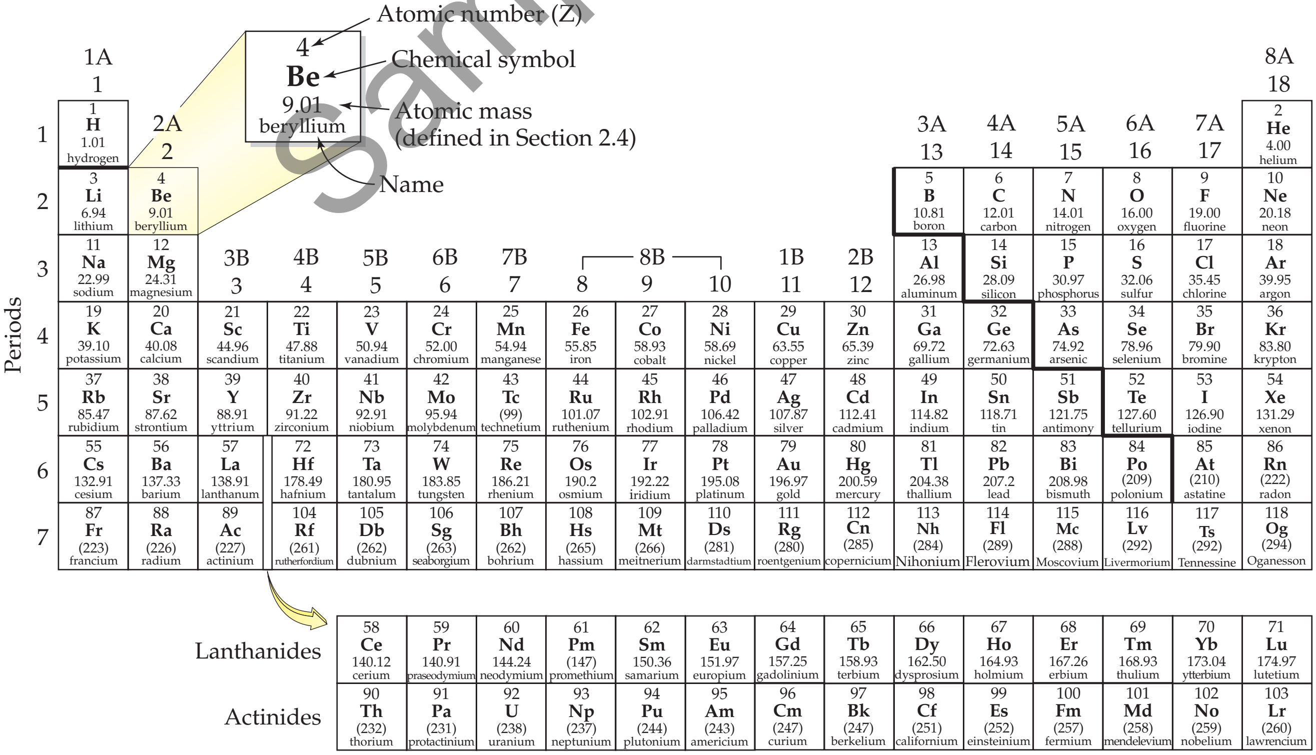


▲ **FIGURE 2.2 The number of protons in the nucleus defines the element** © (left) www.photos.com/Jupiter Images/Getty Images; (right) John A. Rizzo/Digital Vision/Getty Images.

be Al. If the atomic number is 92, the chemical symbol must be U. This is just another way of saying that the number of protons defines the element.

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

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



Periods	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H** 1.01 hydrogen																	2 **He** 4.00 helium
2	3 **Li** 6.94 lithium	4 **Be** 9.01 beryllium											5 **B** 10.81 boron	6 **C** 12.01 carbon	7 **N** 14.01 nitrogen	8 **O** 16.00 oxygen	9 **F** 19.00 fluorine	10 **Ne** 20.18 neon
3	11 **Na** 22.99 sodium	12 **Mg** 24.31 magnesium											13 **Al** 26.98 aluminum	14 **Si** 28.09 silicon	15 **P** 30.97 phosphorus	16 **S** 32.06 sulfur	17 **Cl** 35.45 chlorine	18 **Ar** 39.95 argon
4	19 **K** 39.10 potassium	20 **Ca** 40.08 calcium	21 **Sc** 44.96 scandium	22 **Ti** 47.88 titanium	23 **V** 50.94 vanadium	24 **Cr** 52.00 chromium	25 **Mn** 54.94 manganese	26 **Fe** 55.85 iron	27 **Co** 58.93 cobalt	28 **Ni** 58.69 nickel	29 **Cu** 63.55 copper	30 **Zn** 65.39 zinc	31 **Ga** 69.72 gallium	32 **Ge** 72.63 germanium	33 **As** 74.92 arsenic	34 **Se** 78.96 selenium	35 **Br** 79.90 bromine	36 **Kr** 83.80 krypton
5	37 **Rb** 85.47 rubidium	38 **Sr** 87.62 strontium	39 **Y** 88.91 yttrium	40 **Zr** 91.22 zirconium	41 **Nb** 92.91 niobium	42 **Mo** 95.94 molybdenum	43 **Tc** (99) technetium	44 **Ru** 101.07 ruthenium	45 **Rh** 102.91 rhodium	46 **Pd** 106.42 palladium	47 **Ag** 107.87 silver	48 **Cd** 112.41 cadmium	49 **In** 114.82 indium	50 **Sn** 118.71 tin	51 **Sb** 121.75 antimony	52 **Te** 127.60 tellurium	53 **I** 126.90 iodine	54 **Xe** 131.29 xenon
6	55 **Cs** 132.91 cesium	56 **Ba** 137.33 barium	57 **La** 138.91 lanthanum	72 **Hf** 178.49 hafnium	73 **Ta** 180.95 tantalum	74 **W** 183.85 tungsten	75 **Re** 186.21 rhenium	76 **Os** 190.2 osmium	77 **Ir** 192.22 iridium	78 **Pt** 195.08 platinum	79 **Au** 196.97 gold	80 **Hg** 200.59 mercury	81 **Tl** 204.38 thallium	82 **Pb** 207.2 lead	83 **Bi** 208.98 bismuth	84 **Po** (209) polonium	85 **At** (210) astatine	86 **Rn** (222) radon
7	87 **Fr** (223) francium	88 **Ra** (226) radium	89 **Ac** (227) actinium	104 **Rf** (261) rutherfordium	105 **Db** (262) dubnium	106 **Sg** (263) seaborgium	107 **Bh** (262) bohrium	108 **Hs** (265) hassium	109 **Mt** (266) meitnerium	110 **Ds** (281) darmstadtium	111 **Rg** (280) roentgenium	112 **Cn** (285) copernicium	113 **Nh** (284) Nihonium	114 **Fl** (289) Flerovium	115 **Mc** (288) Moscovium	116 **Lv** (292) Livermorium	117 **Ts** (292) Tennessine	118 **Og** (294) Oganesson

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

▲ **FIGURE 2.3 The periodic table of the elements**

▲ The name *bromine* originates from the Greek word *bromos*, meaning "stench." Bromine vapor, seen as the red-brown gas in this photograph, has a strong odor. © Charles D. Winters/Science Source.

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

EXAMPLE 2.1 ATOMIC NUMBER, ATOMIC SYMBOL, AND ELEMENT NAME

List the atomic symbol and atomic number for each element.

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

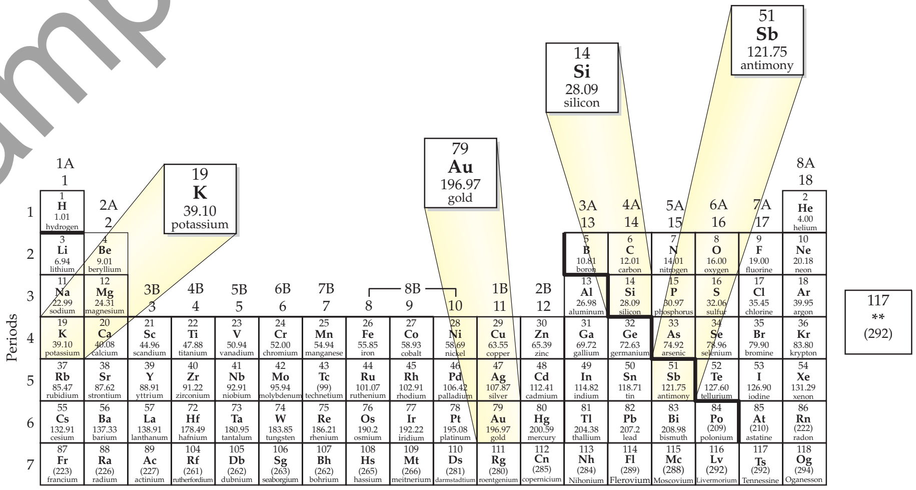
SOLUTION

As you become familiar with the periodic table, you will be able to quickly locate elements on it. At first you may find it easier to locate them in the alphabetical listing on the inside back cover of this book, but you should become familiar with their positions in the periodic table.

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

Curium
96
Cm
(247)

▲ Curium is named after Marie Curie (1867–1934), a chemist who helped discover radioactivity and also discovered two new elements. Curie won two Nobel Prizes for her work. © Library of Congress Prints and Photographs Division.



▶SKILLBUILDER 2.1 | Atomic Number, Atomic Symbol, and Element Name

Find the name and atomic number for each element.

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

▶FOR MORE PRACTICE Problems 42, 43, 44, 45, 46, 47, 48, 49.

2.3 Isotopes: When the Number of Neutrons Varies

LO: Determine atomic numbers, mass numbers, and isotope symbols for an isotope.

LO: Determine number of protons and neutrons from isotope symbols.

All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons. Because neutrons and protons have nearly the same mass (approximately 1 amu), and the number of neutrons in the atoms of a given element can vary, all atoms of a given element *do not* have the same mass. For example, all neon atoms in nature contain 10 protons, but they may have 10, 11, or 12 neutrons (▼ Figure 2.4). All three types of neon atoms exist, and each has a slightly different mass. Atoms with the same number of protons but different numbers of neutrons are **isotopes**. Some elements, such as beryllium (Be) and aluminum (Al), have only one naturally occurring isotope, while other elements, such as neon (Ne) and chlorine (Cl), have two or more.

For a given element, the relative amounts of each different isotope in a naturally occurring sample of that element are always the same. For example, in any natural sample of neon atoms, 90.48 % of the atoms are the isotope with 10 neutrons, 0.27 % are the isotope with 11 neutrons, and 9.25 % are the isotope with 12 neutrons as summarized in Table 2.2. This means that in a sample of 10,000 neon atoms, 9048 have 10 neutrons, 27 have 11 neutrons, and 925 have 12 neutrons. These percentages are the **percent natural abundance** of the isotopes. The preceding numbers are for neon only; each element has its own unique percent natural abundance of isotopes.

Recent studies have shown that for some elements, the relative amounts of each different isotope vary depending on the history of the sample. However, these variations are usually small and beyond the scope of this book.

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

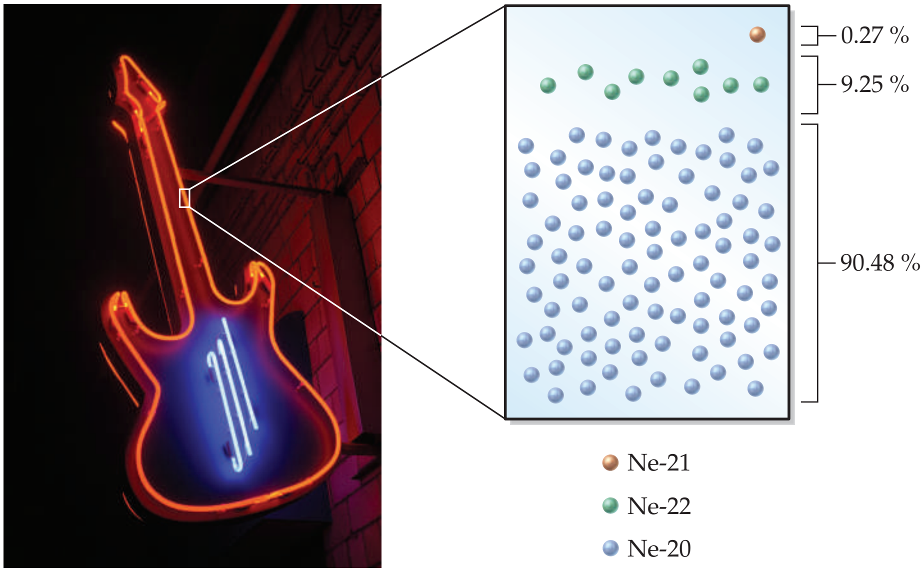
TABLE 2.2 Neon Isotopes

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

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

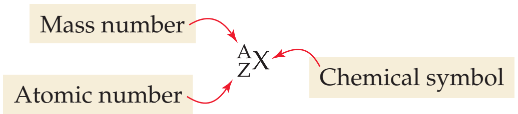
$$A = \text{Number of protons} + \text{Number of neutrons}$$

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



▶ **FIGURE 2.4 Isotopes of neon** Naturally occurring neon contains three different isotopes: Ne-20 (with 10 neutrons), Ne-21 (with 11 neutrons), and Ne-22 (with 12 neutrons). © U.S. Department of Energy

We often symbolize isotopes in the following way:



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

For example, the symbols for the neon isotopes are:

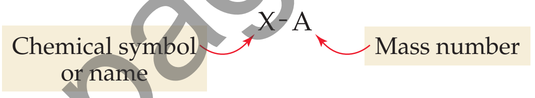
$$ {}^{20}_{10}\mathrm{Ne} \quad {}^{21}_{10}\mathrm{Ne} \quad {}^{22}_{10}\mathrm{Ne} $$

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

CONCEPTUAL CHECKPOINT 2.2

Carbon has two naturally occurring isotopes: ${}^{12}_{6}\mathrm{C}$ and ${}^{13}_{6}\mathrm{C}$. Using circles to represent protons and squares to represent neutrons, draw the nucleus of each isotope.

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



In this notation, the neon isotopes are:

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

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

In general, mass number increases with increasing atomic number.

EXAMPLE 2.2 ATOMIC NUMBERS, MASS NUMBERS, AND ISOTOPE SYMBOLS

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

SOLUTION

You can determine that the atomic number (Z) of carbon is 6 (from the periodic table). This means that carbon atoms have 6 protons. The mass number (A) for the isotope with 7 neutrons is the sum of the number of protons and the number of neutrons.

$$A = 6 + 7 = 13$$

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

▶SKILLBUILDER 2.2 | Atomic Numbers, Mass Numbers, and Isotope Symbols

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

▶FOR MORE PRACTICE Example 2.17; Problems 52, 54, 56, 57.

EXAMPLE 2.3 NUMBERS OF PROTONS AND NEUTRONS FROM ISOTOPE SYMBOLS

How many protons and neutrons are in the chromium isotope $^{52}_{24}\text{Cr}$?

	SOLUTION
The number of protons is equal to Z (lower left number).	$\#p^+ = Z = 24$
The number of neutrons is equal to A (upper left number) –Z (lower left number).	$\#n = A - Z$ $= 52 - 24$ $= 28$

▶SKILLBUILDER 2.3 | Numbers of Protons and Neutrons from Isotope Symbols

How many protons and neutrons are in the potassium isotope $^{39}_{19}\text{K}$?

▶FOR MORE PRACTICE Example 2.18; Problems 58, 59.

✓ CONCEPTUAL CHECKPOINT 2.3

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

(a) silicon

(b) aluminum

(c) cobalt

(d) niobium

✓ CONCEPTUAL CHECKPOINT 2.4

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

2.4 Atomic Mass: The Average Mass of an Element's Atoms

LO: Calculate atomic mass from percent natural abundances and isotopic masses.

As we just learned, the atoms of a given element may have different masses (because of isotopes). Therefore, mass number is not always a useful quantity describing a significant quantity of an element. We can, however, calculate an average mass—called the **atomic mass**—for each element. You can find the atomic mass of each element in the periodic table directly beneath the element's symbol; it represents the average mass of the atoms that compose that element. For example, the periodic table lists the atomic mass of chlorine as 35.45 amu. Naturally occurring chlorine consists of 75.77 % chlorine-35 (mass 34.97 amu) and 24.23 % chlorine-37 (mass 36.97 amu). Its atomic mass is:

Some books use the terms *average atomic mass* or atomic weight instead of simply *atomic mass*.

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

CHEMISTRY **IN THE ENVIRONMENT**

▶ Radioactive Isotopes at Hanford, Washington

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

The nuclear power and nuclear weapons industries produce by-products containing unstable isotopes of several different elements. Many of these isotopes emit nuclear radiation for a long time, and their disposal is an environmental problem. For example, in Hanford, Washington, which for 50 years produced fuel for nuclear weapons, 177 underground storage tanks contain 200 million litres of highly radioactive nuclear waste. Certain radioactive isotopes within that waste will produce nuclear radiation for the foreseeable future. Unfortunately, some of the underground storage tanks in Hanford are aging, and leaks have allowed some of the waste to seep into the environment. While the danger from short-term external exposure to this waste is minimal, ingestion of the waste through contamination of drinking water or food supplies would pose significant health risks. Consequently, Hanford is now the site of the largest environmental cleanup project in U.S. history, involving 11,000 workers. The U.S. government expects the project to last for decades, and current costs are about $2 billion per year.

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

B2.1 CAN YOU ANSWER THIS? *Give the number of neutrons in each of the following isotopes: Pb-206, Pb-207, Pb-208, Pb-185, Tc-99.*

◀ Storage tanks at Hanford, Washington, contain 200 million litres of high-level nuclear waste. Each tank pictured here holds 4 million litres.

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

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

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

In general, we calculate atomic mass according to the following equation:

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

where the fractions of each isotope are the percent natural abundances converted to their decimal values. Atomic mass is useful because it allows us to assign a characteristic mass to each element and, as we will see in Module 5, it allows us to quantify the number of atoms in a sample of that element.

EXAMPLE 2.4 CALCULATING ATOMIC MASS

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

Remember to convert the percent natural abundances into decimal form by dividing by 100.	**SOLUTION** $\text{Fraction Ga-69} = \frac{60.11}{100} = 0.6011$ $\text{Fraction Ga-71} = \frac{39.89}{100} = 0.3989$
Use the fractional abundances and the atomic masses of the isotopes to calculate the atomic mass according to the atomic mass definition.	$\text{Atomic mass} = (0.6011 \times 68.9256\ \text{amu}) + (0.3989 \times 70.9247\ \text{amu})$ $= 41.4312\ \text{amu} + 28.2919\ \text{amu}$ $= 69.7231 = 69.72\ \text{amu}$

▶SKILLBUILDER 2.4 | Calculating Atomic Mass

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

▶FOR MORE PRACTICE Example 2.19; Problems 62, 63.

✓ CONCEPTUAL CHECKPOINT 2.5

A fictitious element is composed of isotopes A and B with masses of 61.9887 and 64.9846 amu, respectively. The atomic mass of the element is 64.52. What can you conclude about the natural abundances of the two isotopes?

(a) The natural abundance of isotope A must be greater than the natural abundance of isotope B.

(b) The natural abundance of isotope B must be greater than the natural abundance of isotope A.

(c) The natural abundances of both isotopes must be about equal.

(d) Nothing can be concluded about the natural abundances of the two isotopes from the given information.

2.5 Ions: Losing and Gaining Electrons

LO: **Determine ion charge from numbers of protons and electrons.**

LO: **Determine the number of protons and electrons in an ion.**

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

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

$$Li \rightarrow Li^+ + e^-$$

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

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

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

For the Li^+ ion with 3 protons and 2 electrons the charge is:

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

A neutral fluorine (F) atom contains 9 protons and 9 electrons; however, in chemical reactions a fluorine atom gains 1 electron to form F^- ions:

$$F + e^- \rightarrow F^-$$

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

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

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

EXAMPLE 2.5 DETERMINING ION CHARGE FROM NUMBERS OF PROTONS AND ELECTRONS

Determine the charge of each ion.

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

SOLUTION

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

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

You are given the number of electrons in the problem. You can obtain the number of protons from the element's atomic number in the periodic table.

(a) Magnesium's atomic number is 12.

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

(b) Sulfur's atomic number is 16.

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

(c) Iron's atomic number is 26.

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

▶SKILLBUILDER 2.5 | Determining Ion Charge from Numbers of Protons and Electrons

Determine the charge of each ion.

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

▶FOR MORE PRACTICE Example 2.20; Problems 66, 67.

EXAMPLE 2.6 DETERMINING THE NUMBER OF PROTONS AND ELECTRONS IN AN ION

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

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

SOLUTION

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

$2+ = 20 - \#e^{-}$

$\#e^{-} = 20 - 2 = 18$

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

▶SKILLBUILDER 2.6 | Determining the Number of Protons and Electrons in an Ion

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

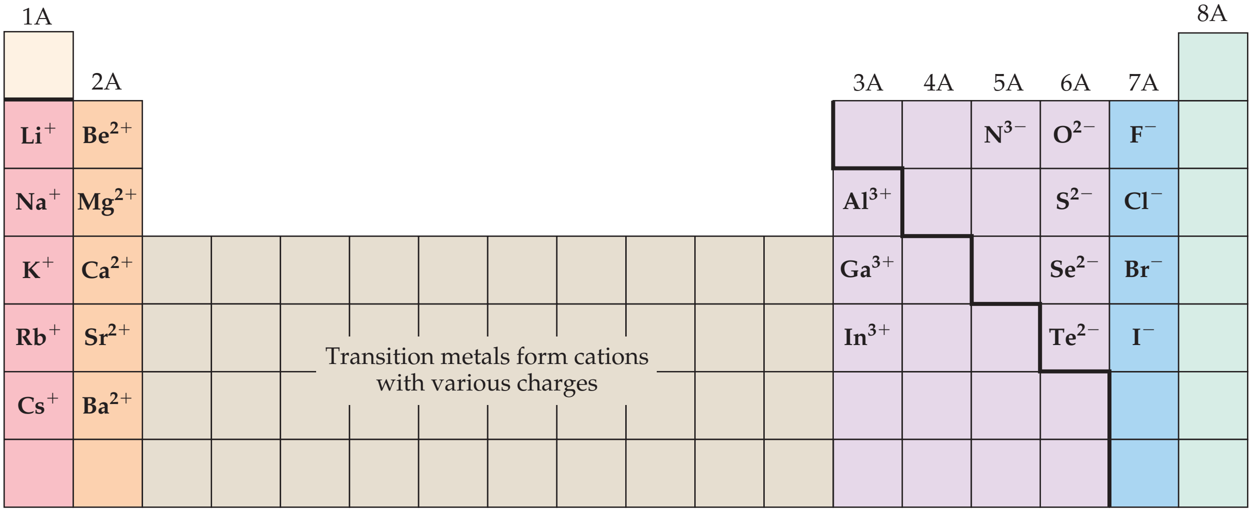
▶FOR MORE PRACTICE Example 2.21; Problems 68, 69.

Ions and the Periodic Table

For many main-group elements (Section 2.7), we can use the periodic table to predict how many electrons tend to be lost or gained when an atom of that particular element ionizes. The number associated with the letter A above each *main-group* column in the periodic table—1 through 8—gives the number of *valence electrons* for the elements in that column. We will discuss the concept of valence electrons more fully in the next section in this module; for now, you can think of valence electrons as the outermost electrons in an atom. Because oxygen is in column 6A, we can deduce that it has 6 valence electrons; because magnesium is in column 2A, it has 2 valence electrons, and so on. An important exception to this rule is helium—it is in column 8A, but has only 2 valence electrons. Valence electrons are particularly important because, as we shall see in Module 3, these electrons are the ones that are most important in chemical bonding.

We can predict the charge acquired by a particular element when it ionizes from its position in the periodic table relative to the noble gases.

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



▲ **FIGURE 2.5 Elements that form predictable ions**

For example, the closest noble gas to oxygen is neon. When oxygen ionizes, it *acquires* two additional electrons for a total of 8 valence electrons—the same number as neon. When determining the closest noble gas, we can move either forward or backward on the periodic table. For example, the closest noble gas to magnesium is also neon, even though neon (atomic number 10) falls before magnesium (atomic number 12) in the periodic table. Magnesium *loses* its 2 valence electrons to attain the same number of valence electrons as neon.

In accordance with this principle, the alkali metals (Group 1A) tend to lose 1 electron and form 1+ ions, while the alkaline earth metals (Group 2A) tend to lose 2 electrons and form 2+ ions. The halogens (Group 7A) tend to gain 1 electron and form 1– ions. The groups in the periodic table that form predictable ions are shown in ▲ Figure 2.5. Become familiar with these groups and the ions they form. Later in this module (Section 2.9), we will examine a theory that more fully explains why these groups form ions as they do.

EXAMPLE 2.7 CHARGE OF IONS FROM POSITION IN PERIODIC TABLE

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

SOLUTION

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

▶SKILLBUILDER 2.7 | Charge of Ions from Position in Periodic Table

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

▶FOR MORE PRACTICE Problems 126, 127.

✓ CONCEPTUAL CHECKPOINT 2.6

Which pair of ions has the same total number of electrons?

(a) Na^+ and Mg^{2+}

(b) F^- and Cl^-

(c) O^- and O^{2-}

(d) Ga^{3+} and Fe^{3+}

2.6 Arrangements of Electrons Around the Nucleus

LO: Write electron configurations and orbital diagrams for atoms.

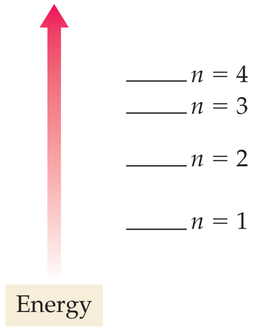
The main points made about electrons so far is that they are negatively charged, have negligible mass, and they are dispersed throughout the volume of an atom outside the nucleus. In this section the way that electrons are arranged around the nucleus of an atom is introduced. The current understanding of these electron arrangements is called the quantum-mechanical model.

The number and arrangement of electrons in atoms dictates the chemical properties of that element. Experimental data has led to the discovery that electrons occupy discreet energy levels around the nucleus. These energy levels are called **orbitals**. When energy is put into an atom, electrons are excited to higher-energy orbitals. When an electron in an atom relaxes from a higher-energy orbital to a lower-energy orbital, the atom emits light. The energy (and therefore the colour) of the emitted light corresponds to the energy difference between the two orbitals in the transition.

In this section, we examine quantum-mechanical orbitals and electron configurations. An electron configuration is a compact way to specify the occupation of quantum-mechanical orbitals by electrons.

Quantum-Mechanical Orbitals

The lowest-energy orbital in the quantum-mechanical model is the *1s orbital*. We specifiy it by the number 1 and the letter *s*. The number is the **principal quantum number** (n) and specifies the **principal shell** of the orbital. The higher the principal quantum number, the higher the energy of the orbital. The possible principal quantum numbers are $n = 1, 2, 3 \ldots$, with energy increasing as n increases (◀ Figure 2.6). Because the 1*s* orbital has the lowest possible principal quantum number, it is in the lowest-energy shell and has the lowest possible energy.

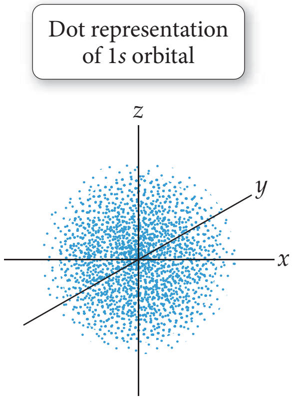


▲ **FIGURE 2.6 Principal quantum numbers** The principal quantum numbers ($n = 1, 2, 3 \ldots$) determine the energy of the hydrogen quantum-mechanical orbitals.

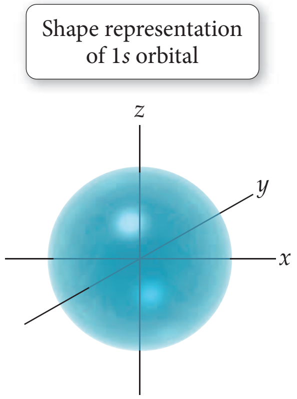
The letter indicates the **subshell** of the orbital and specifies its shape. The possible letters are *s*, *p*, *d*, and *f*, and each letter corresponds to a different shape. For example, orbitals within the *s* subshell have a spherical shape. The 1*s* quantum-mechanical orbital is a three-dimensional probability map. We sometimes represent orbitals with dots (▼ Figure 2.7), where the dot density is proportional to the probability of finding the electron.

This analogy is purely hypothetical. It is impossible to photograph electrons in this way.

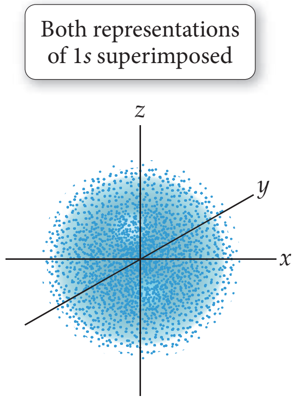
We can understand the dot representation of an orbital better with another analogy. Imagine taking a photograph of an electron in an atom every second for 10 or 15 minutes. One second the electron is very close to the nucleus; the next second it is farther away and so on. Each photo shows a dot representing the electron's position relative to the nucleus at that time. If we took hundreds of photos and superimposed



▲ **FIGURE 2.7 1s orbital** The dot density in this plot is proportional to the probability of finding the electron. The greater dot density near the middle indicates a higher probability of finding the electron near the nucleus.



▲ **FIGURE 2.8 Shape representation of the 1s orbital** Because the distribution of electron density around the nucleus in Figure 2.7 is symmetrical—the same in all directions—we can represent the 1*s* orbital as a sphere.



▲ **FIGURE 2.9 Orbital shape and dot representation for the 1s orbital** The shape representation of the 1*s* orbital superimposed on the dot density representation. We can see that when the electron is in the 1*s* orbital, it is most likely to be found within the sphere.

FIGURE 2.10 Subshells The number of subshells in a given principal shell is equal to the value of n.

Shell	Number of subshells	Letters specifying subshells			
$n = 4$	4	s	p	d	f
$n = 3$	3	s	p	d	
$n = 2$	2	s	p		
$n = 1$	1	s			

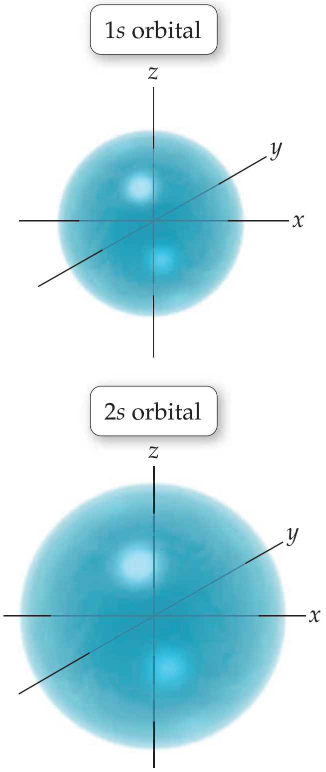


FIGURE 2.11 The 2s orbital The 2s orbital is similar to the 1s orbital, but larger in size.

all of them, we would have an image like ◀ Figure 2.7—a statistical representation of where the electron is found. Notice that the dot density for the 1*s* orbital is greatest near the nucleus and decreases farther away from the nucleus. This means that the electron is more likely to be found close to the nucleus than far away from it.

Orbitals can also be represented as geometric shapes that encompass most of the volume where the electron is likely to be found. For example, the 1*s* orbital can be represented as a sphere (◀ Figure 2.8) that encompasses the volume within which the electron is found 90 % of the time. If we superimpose the dot representation of the 1*s* orbital on the shape representation (◀ Figure 2.9), we can see that most of the dots are within the sphere, meaning that the electron is most likely to be found within the sphere when it is in the 1*s* orbital.

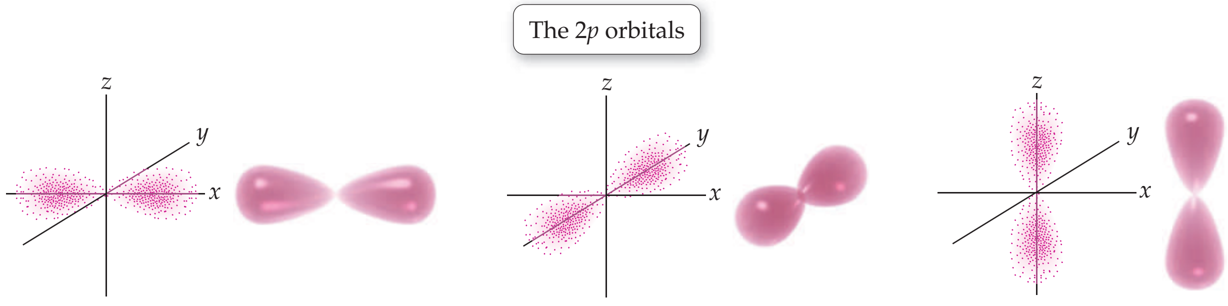
The single electron of an undisturbed hydrogen atom at room temperature is in the 1*s* orbital. This is the **ground state**, or lowest energy state, of the hydrogen atom. However, the quantum-mechanical model allows transitions to higher-energy orbitals upon the absorption of energy. What are these higher-energy orbitals? What do they look like?

The next orbitals in the quantum-mechanical model are those with principal quantum number $n = 2$ Unlike the $n = 1$ principal shell, which contains only one subshell (specified by *s*), the $n = 2$ principal shell contains two subshells, specified by *s* and *p*.

The number of subshells in a given principal shell is equal to the value of n. Therefore the $n = 1$ principal shell has one subshell, the $n = 2$ principal shell has two subshells, and so on (▲ Figure 2.10). The *s* subshell contains the 2*s* orbital, higher in energy than the 1*s* orbital and slightly larger (◀ Figure 2.11), but otherwise similar in shape. The *p* subshell contains three 2*p* orbitals (▼ Figure 2.12), all with the same dumbbell-like shape but with different orientations.

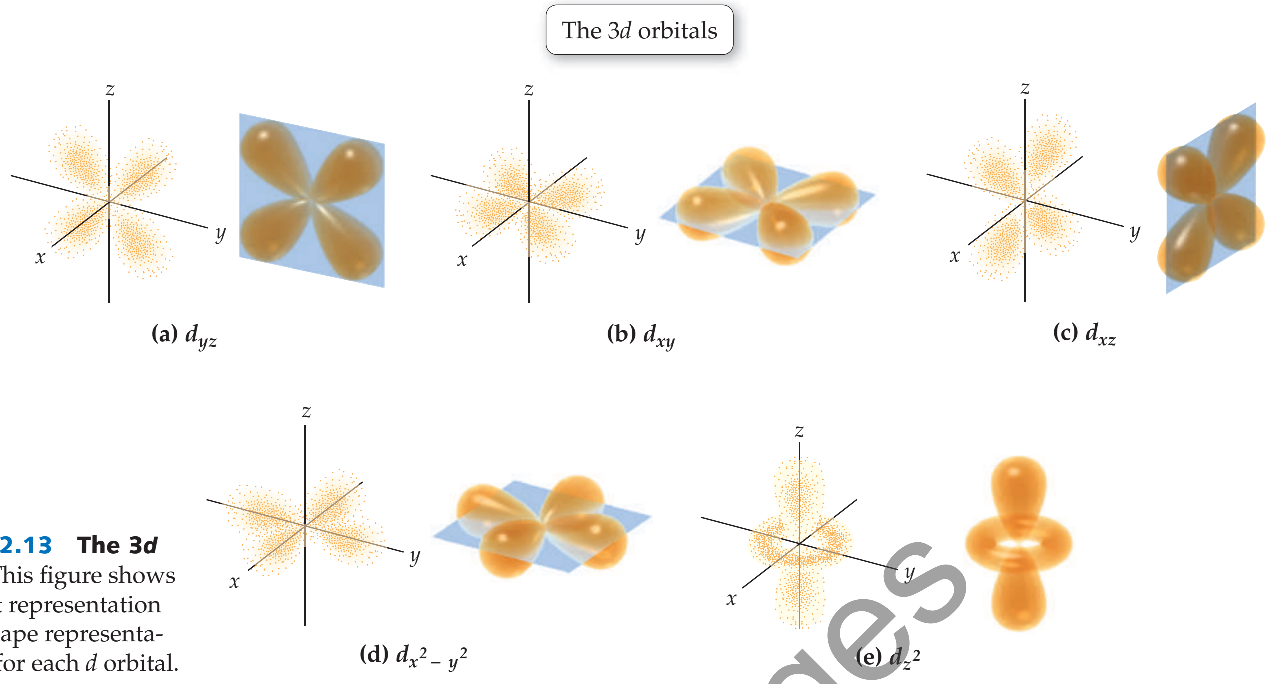
The next principal shell, $n = 3$ contains three subshells specified by *s*, *p*, and *d*. The *s* and *p* subshells contain the 3*s* and 3*p* orbitals, similar in shape to the 2*s* and 2*p* orbitals, but slightly larger and higher in energy. The *d* subshell contains the five *d* orbitals shown in ▶ Figure 2.13. The next principal shell, $n = 4$ contains four subshells specified by *s*, *p*, *d*, and *f*. The *s*, *p*, and *d* subshells are similar to those in $n = 3$. The *f* subshell contains seven orbitals (called the 4*f* orbitals), whose shape we do not consider in this book.

As we have already discussed, hydrogen's single electron is usually in the 1*s* orbital because electrons generally occupy the lowest-energy orbital available. In



(a) p_x (b) p_y (c) p_z

FIGURE 2.12 The 2*p* orbitals This figure shows both the dot representation (left) and shape representation (right) for each *p* orbital.



(a) d_{yz} (b) d_{xy} (c) d_{xz} (d) $d_{x^2-y^2}$ (e) d_{z^2}

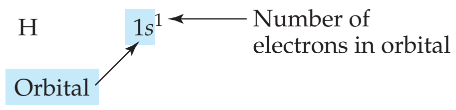
FIGURE 2.13 The 3*d* orbitals This figure shows both the dot representation (left) and shape representation (right) for each *d* orbital.

hydrogen, the rest of the orbitals are normally empty. However, the absorption of energy by a hydrogen atom can cause the electron to jump (or make a transition) from the 1*s* orbital to a higher-energy orbital. When the electron is in a higher-energy orbital, we say that the hydrogen atom is in an **excited state**.

Because of their higher energy, excited states are unstable, and the electron will usually fall (or relax) back to a lower-energy orbital. In the process the electron emits energy, often in the form of light. The quantum-mechanical model predicts the different colours of light emitted in hydrogen, with one electron, as well as other elements with more than one electron.

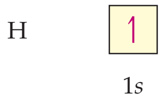
Electron Configurations: How Electrons Occupy Orbitals

An **electron configuration** illustrates the occupation of orbitals by electrons for a particular atom. For example, the electron configuration for a ground-state (or lowest energy) hydrogen atom is:



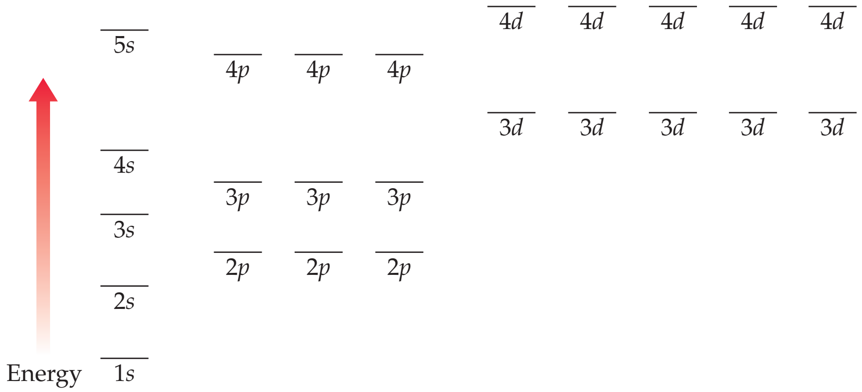
The electron configuration tells us that hydrogen's single electron is in the 1*s* orbital.

Another way to represent this information is with an **orbital diagram**, which gives similar information but shows the electrons as arrows in a box representing the orbital. The orbital diagram for a ground-state hydrogen atom is:



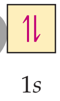
The box represents the 1*s* orbital, and the arrow within the box represents the electron in the 1*s* orbital. In orbital diagrams, the direction of the arrow (pointing up or pointing down) represents **electron spin**, a fundamental property of electrons. All electrons have spin. The **Pauli exclusion principle** states that *orbitals may hold no more than two electrons with opposing spins*. We symbolize this as two arrows pointing in opposite directions:

⥮



▶ **FIGURE 2.14 Energy ordering of orbitals for multi-electron atoms** Different subshells within the same principal shell have different energies.

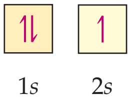
A helium atom, for example, has two electrons. The electron configuration and orbital diagram for helium are:

	Electron configuration	**Orbital diagram**
He	$1s^2$	

Since we know that electrons occupy the lowest-energy orbitals available, and since we know that only two electrons (with opposing spins) are allowed in each orbital, we can continue to build ground-state electron configurations for the rest of the elements as long as we know the energy ordering of the orbitals. ▲ Figure 2.14 shows the energy ordering of a number of orbitals for multi-electron atoms.

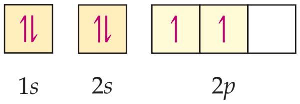
In multi-electron atoms, the subshells within a principal shell do not have the same energy because of electron–electron interactions.

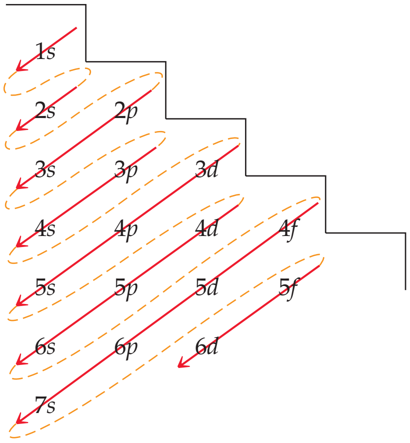
Notice that, for multi-electron atoms (in contrast to hydrogen which has only one electron), the subshells within a principal shell *do not* have the same energy. In elements other than hydrogen, the energy ordering is not determined by the principal quantum number alone. For example, in multi-electron atoms, the 4*s* subshell is lower in energy than the 3*d* subshell, even though its principal quantum number is higher. Using this relative energy ordering, we can write ground-state electron configurations and orbital diagrams for other elements. For lithium, which has three electrons, the electron configuration and orbital diagram are:

	Electron configuration	**Orbital diagram**
Li	$1s^22s^1$	

Remember that the number of electrons in an atom is equal to its atomic number.

For carbon, which has 6 electrons, the electron configuration and orbital diagram are:

	Electron configuration	**Orbital diagram**
C	$1s^22s^22p^2$	



▲ **FIGURE 2.15 Orbital filling order** The arrows indicate the order in which orbitals fill.

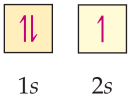
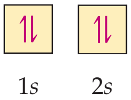
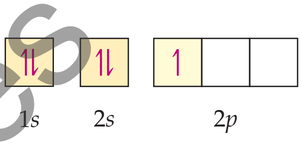
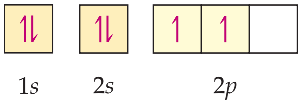
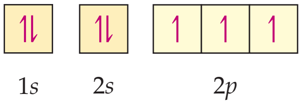
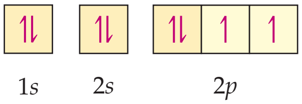
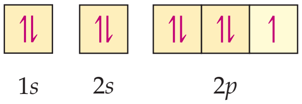
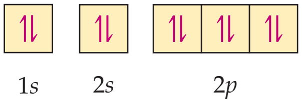
Notice that the 2*p* electrons occupy the *p* orbitals (of equal energy) singly rather than pairing in one orbital. This is the result of **Hund's rule**, which states that *when filling orbitals of equal energy, electrons fill them singly first, with parallel spins.*

Before we write electron configurations for other elements, let us summarize what we have learned so far:

- Electrons occupy orbitals so as to minimize the energy of the atom; therefore, lower-energy orbitals fill before higher-energy orbitals. Orbitals fill in the following order: 1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p* 5*s* 4*d* 5*p* 6*s* (◀ Figure 2.15).

- Orbitals can hold no more than two electrons each. When two electrons occupy the same orbital, they must have opposing spins. This is known as the Pauli exclusion principle.
- When orbitals of identical energy are available, these are first occupied singly with parallel spins rather than in pairs. This is known as Hund's rule.

Consider the electron configurations and orbital diagrams for elements with atomic numbers 3 through 10:

Symbol (#e^-)	**Electron configuration**	**Orbital diagram**
Li (3)	$1s^2 2s^1$	
Be (4)	$1s^2 2s^2$	
B (5)	$1s^2 2s^2 2p^1$	
C (6)	$1s^2 2s^2 2p^2$	
N (7)	$1s^2 2s^2 2p^3$	
O (8)	$1s^2 2s^2 2p^4$	
F (9)	$1s^2 2s^2 2p^5$	
Ne (10)	$1s^2 2s^2 2p^6$	

Notice how the *p* orbitals fill. As a result of Hund's rule, the *p* orbitals fill with single electrons before they fill with paired electrons. The electron configuration of neon represents the complete filling of the $n = 2$ principal shell. When writing electron configurations for elements beyond neon—or beyond any other noble gas—we often abbreviate the electron configuration of the previous noble gas by the symbol for the noble gas in brackets. For example, the electron configuration of sodium is:

Na $1s^2 2s^2 2p^6 3s^1$

We can write this using the noble gas core notation as:

Na $[\text{Ne}]3s^1$

where [Ne] represents $1s^2 2s^2 2p^6$, the electron configuration for neon.

To write an electron configuration for an element, we first find its atomic number from the periodic table—this number equals the number of electrons in the neutral atom. Then we use the order of filling from Figure 2.14 to distribute the electrons in the appropriate orbitals. Remember that each orbital can hold a maximum of 2 electrons. Consequently:

- the *s* subshell has only 1 orbital and therefore can hold only 2 electrons.
- the *p* subshell has 3 orbitals and therefore can hold 6 electrons.
- the *d* subshell has 5 orbitals and therefore can hold 10 electrons.
- the *f* subshell has 7 orbitals and therefore can hold 14 electrons.

EXAMPLE 2.8 ELECTRON CONFIGURATIONS

Write electron configurations for each element.

(a) Mg **(b)** S **(c)** Ga

SOLUTION

(a) Magnesium has 12 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbitals, and two into the 3*s* orbital.

You can also write the electron configuration more compactly using the noble gas core notation. For magnesium, use [Ne] to represent $1s^2 2s^2 2p^6$.

Mg $1s^2 2s^2 2p^6 3s^2$

or

Mg $[\text{Ne}]3s^2$

(b) Sulfur has 16 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbitals, two into the 3*s* orbital, and four into the 3*p* orbitals.

You can write the electron configuration more compactly by using [Ne] to represent $1s^2 2s^2 2p^6$.

S $1s^2 2s^2 2p^6 3s^2 3p^4$

or

S $[\text{Ne}]3s^2 3p^4$

(c) Gallium has 31 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbitals, two into the 3*s* orbital, six into the 3*p* orbitals, two into the 4*s* orbital, ten into the 3*d* orbitals, and one into the 4*p* orbitals. Notice that the *d* subshell has five orbitals and can therefore hold 10 electrons.

You can write the electron configuration more compactly by using [Ar] to represent $1s^2 2s^2 2p^6 3s^2 3p^6$.

Ga $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$

or

Ga $[\text{Ar}]4s^2 3d^{10} 4p^1$

▶ **SKILLBUILDER 2.8 | Electron Configurations**

Write electron configurations for each element.

(a) Al **(b)** Br **(c)** Sr

▶ **SKILLBUILDER PLUS |** Write electron configurations for each ion. (*Hint:* To determine the number of electrons to include in the electron configuration of an ion, add or subtract electrons as needed to account for the charge of the ion.)

(a) Al^{3+} **(b)** Cl^- **(c)** O^{2-}

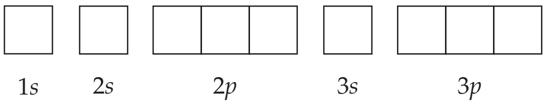
▶ **FOR MORE PRACTICE** Problems 76, 77, 80, 81.

EXAMPLE 2.9 WRITING ORBITAL DIAGRAMS

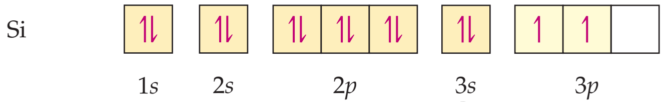
Write an orbital diagram for silicon.

SOLUTION

Since silicon is atomic number 14, it has 14 electrons. Draw a box for each orbital, putting the lowest-energy orbital (1*s*) on the far left and proceeding to orbitals of higher energy to the right.



Distribute the 14 electrons into the orbitals, allowing a maximum of 2 electrons per orbital and remembering Hund's rule. The complete orbital diagram is:



▶ SKILLBUILDER 2.9 | Writing Orbital Diagrams

Write an orbital diagram for argon.

▶ FOR MORE PRACTICE Example 2.22; Problems 78, 79.

✓ CONCEPTUAL CHECKPOINT 2.7

Which pair of elements has the same *total* number of electrons in *p* orbitals?

(a) Na and K

(b) K and Kr

(c) P and N

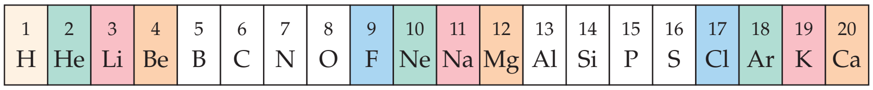
(d) Ar and Ca

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

LO: Use the periodic table to classify elements by group.

The organization of the periodic table has its origins in the work of Dmitri Mendeleev (1834–1907), a nineteenth-century Russian chemistry professor. In his time, about 65 different elements had been discovered. Thanks to the work of a number of chemists, much was known about each of these elements, including their relative masses, chemical activity, and some of their physical properties. However, there was no systematic way of organizing them.

▲ Dmitri Mendeleev, a Russian chemistry professor who arranged early versions of the periodic table. © Popova Olga/Fotolia.



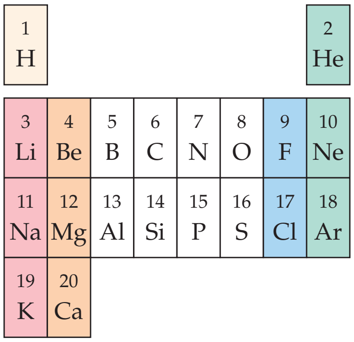
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
H	He	Li	Be	B	C	N	O	F	Ne	Na	Mg	Al	Si	P	S	Cl	Ar	K	Ca

▲ **FIGURE 2.16 Recurring properties** These elements are listed in order of increasing atomic number (Mendeleev used relative mass, which is similar). The color of each element represents its properties. Notice that the properties (colors) of these elements form a repeating pattern.

In 1869, Mendeleev noticed that certain groups of elements had similar properties. He found that if he listed the elements in order of increasing relative mass, those similar properties recurred in a regular pattern (▲ Figure 2.16). Mendeleev summarized these observations in the **periodic law**:

Periodic means "recurring regularly."

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



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

▲ **FIGURE 2.17 Making a periodic table** If we place the elements from Figure 2.16 in a table, we can arrange them in rows so that similar properties align in the same vertical columns. This is similar to Mendeleev's first periodic table.

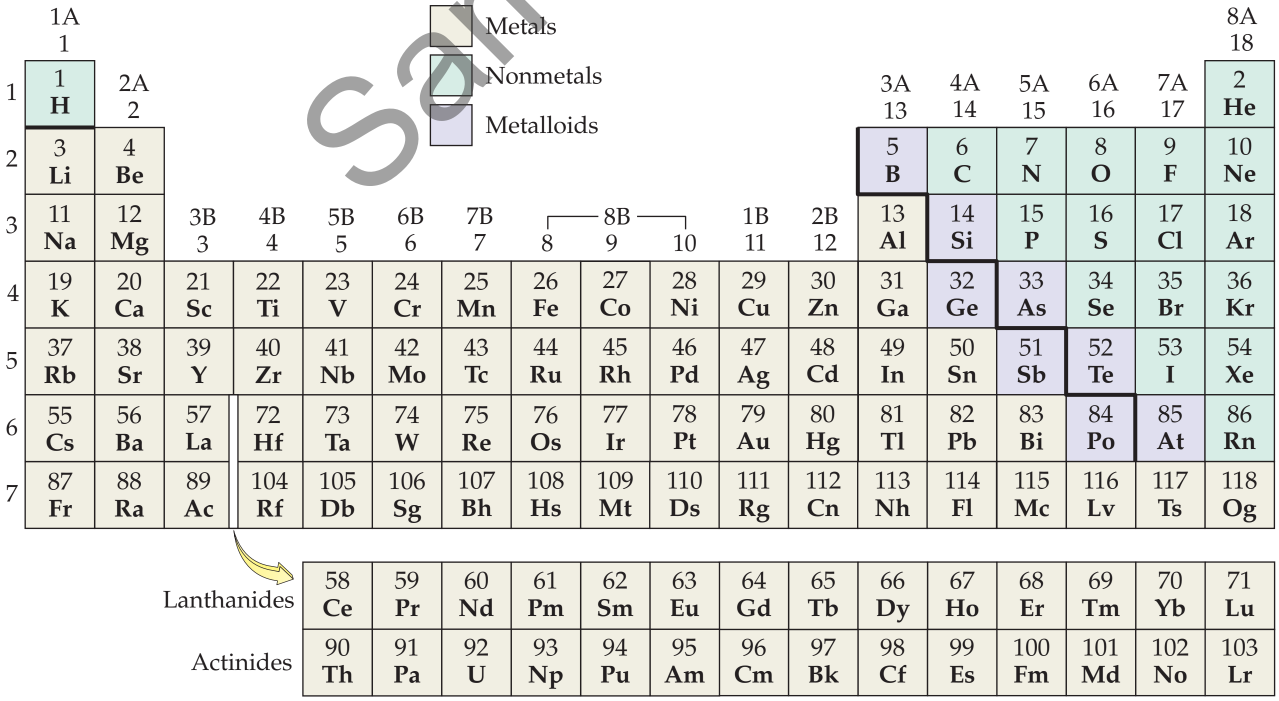
Mendeleev organized all the known elements in a table in which relative mass increased from left to right and elements with similar properties were aligned in the same vertical columns (◀ Figure 2.17). Because many elements had not yet been discovered, Mendeleev's table contained some gaps, which allowed him to predict the existence of yet-undiscovered elements. For example, Mendeleev predicted the existence of an element he called *eka-silicon*, which fell below silicon on the table and between gallium and arsenic. In 1886, eka-silicon was discovered by German chemist Clemens Winkler (1838–1904) and was found to have almost exactly the properties that Mendeleev had anticipated. Winkler named the element germanium, after his home country.

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

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

We can broadly classify the elements in the periodic table as metals, nonmetals, and metalloids (▼ Figure 2.18). **Metals** occupy the left side of the periodic table and have similar properties: They are good conductors of heat and electricity; they can be pounded into flat sheets (malleability); they can be drawn into wires (ductility); they are often shiny; and they tend to lose electrons when they undergo chemical changes. Examples of metals are iron, magnesium, chromium, and sodium.

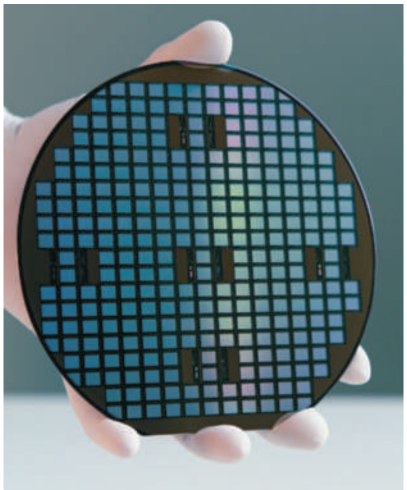
Nonmetals occupy the upper right side of the periodic table. The dividing line between metals and nonmetals is the zigzag diagonal line running from boron to astatine in Figure 2.17. Nonmetals have more varied properties—some are solids at room temperature, others are gases—but as a whole they tend to be poor



	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

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

▲ **FIGURE 2.18 Metals, nonmetals, and metalloids** The elements in the periodic table can be broadly classified as metals, nonmetals, or metalloids.



▲ Silicon is a metalloid used extensively in the computer and electronics industries. © Tom Grill/Corbis.

conductors of heat and electricity, and they all tend to gain electrons when they undergo chemical changes. Examples of nonmetals are oxygen, nitrogen, chlorine, and iodine.

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

EXAMPLE 2.10 CLASSIFYING ELEMENTS AS METALS, NONMETALS, OR METALLOIDS

Classify each element as a metal, nonmetal, or metalloid.

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

SOLUTION

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

▶SKILLBUILDER 2.10 | Classifying Elements as Metals, Nonmetals, or Metalloids

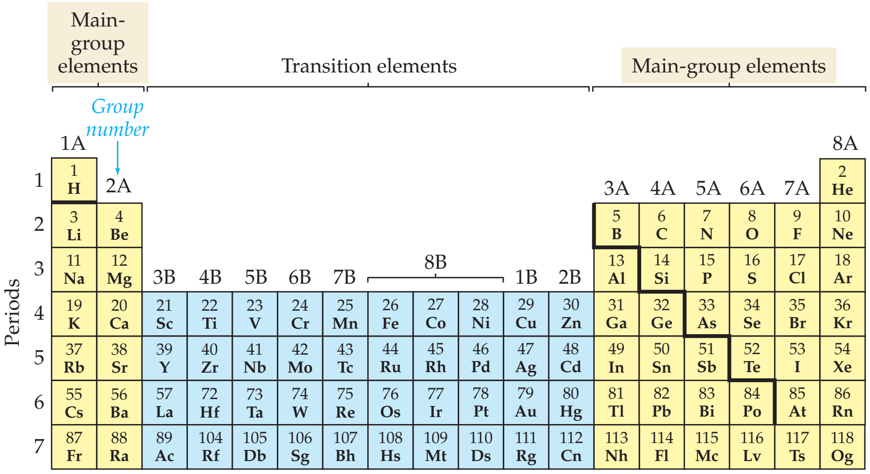
Classify each element as a metal, nonmetal, or metalloid.

(a) S **(b)** Cl **(c)** Ti **(d)** Sb

▶FOR MORE PRACTICE Problems 82, 83.

We can also broadly divide the periodic table into **main-group elements**, whose properties tend to be more predictable based on their position in the periodic table, and **transition elements** or **transition metals**, whose properties are less easily predictable based simply on their position in the periodic table (▼ Figure 2.19). Main-group elements are in columns labeled with a number and the letter A. Transition elements are in columns labeled with a number and the letter B. A competing numbering system does not use letters, but only the numbers 1–18. We show both numbering systems in the periodic table in the inside front cover of this book.

▶ FIGURE 2.19 Main-group and transition elements We can broadly divide the periodic table into main-group elements, whose properties we can generally predict based on their position, and transition elements, whose properties tend to be less predictable based on their position.



	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

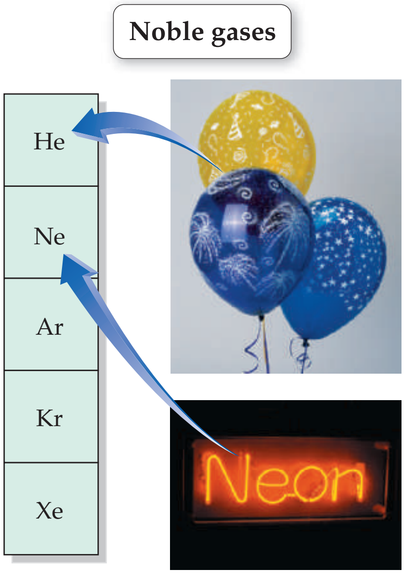
CONCEPTUAL CHECKPOINT 2.8

Which element is a main-group metal?

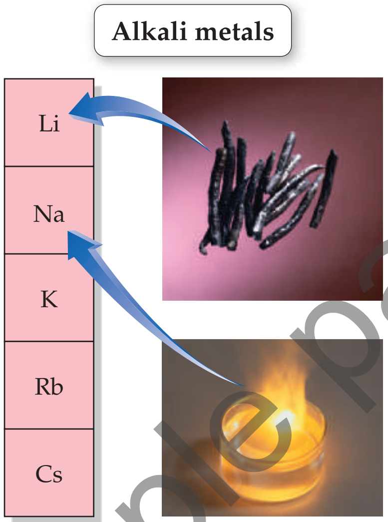
(a) O **(b)** Ag **(c)** P **(d)** Pb

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

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

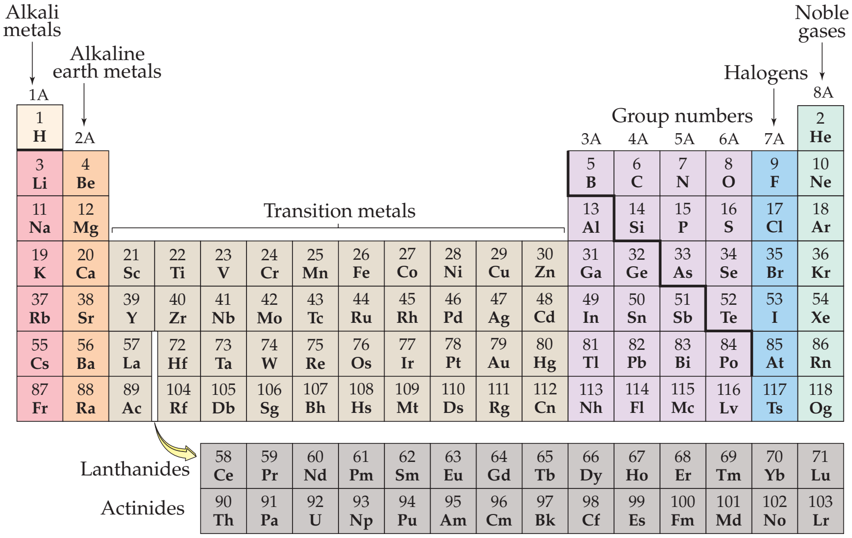


▲ The noble gases include helium (used in balloons), neon (used in neon signs), argon, krypton, and xenon. (top left) PhotoDisc/Getty Images. (bottom left) Graeme Dawes/Shutterstock. (top right) Richard Megna/Fundamental Photographs. (bottom right) Sciencephotos/Alamy.



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

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



EXAMPLE 2.11 GROUPS OF ELEMENTS

To which group of elements does each element belong?

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

SOLUTION

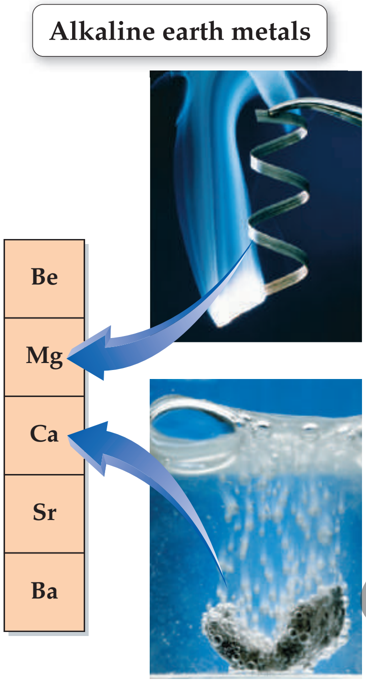
(a) Mg is in Group 2A; it is an alkaline earth metal.
(b) N is in Group 5A.
(c) K is in Group 1A; it is an alkali metal.
(d) Br is in Group 7A; it is a halogen.

▶SKILLBUILDER 2.11 | Groups and Families of Elements

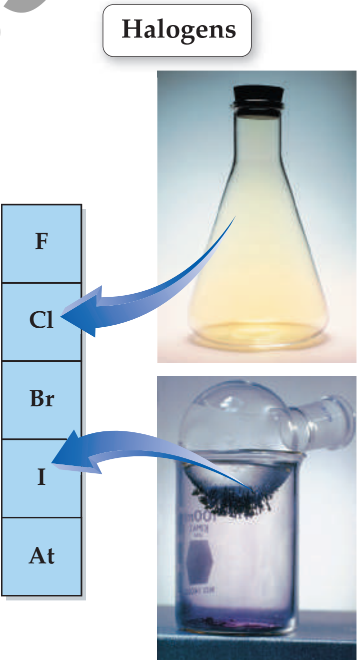
To which group of elements does each element belong?

(a) Li **(b)** B **(c)** I **(d)** Ar

▶FOR MORE PRACTICE Problems 88, 89, 90, 91, 92, 93, 94, 95.



◀ The alkaline earth metals include beryllium, magnesium (shown burning in the first photo), calcium (shown reacting with water in the second photo), strontium, and barium. (top) Richard Megna/Fundamental Photographs. (bottom) Richard Megna/Fundamental Photographs.



▶ The halogens include fluorine, chlorine, bromine, iodine, and astatine. © (top) Charles D. Winters/Science Source. (bottom) Tom Bochsler/Pearson.

✓ CONCEPTUAL CHECKPOINT 2.9

Which statement is NEVER true?

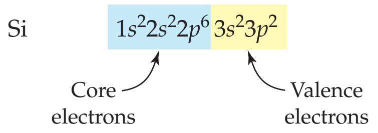
(a) An element can be both a transition element and a metal.
(b) An element can be both a transition element and a metalloid.
(c) An element can be both a metalloid and a halogen.
(d) An element can be both a main-group element and a halogen.

2.8 Electron Configurations and the Periodic Table

LO: Identify valence electrons and core electrons.

LO: Write electron configurations for elements based on their positions in the periodic table.

Valence electrons are the electrons in the outermost principal shell (the principal shell with the highest principal quantum number, n). These electrons are important because, as we will see in the next module, they are held most loosely and are most easily lost or shared; therefore they are involved in chemical bonding. Electrons that are *not* in the outermost principal shell are **core electrons**. For example, silicon, with the electron configuration of $1s^22s^22p^63s^23p^2$, has 4 valence electrons (those in the $n = 3$ principal shell) and 10 core electrons.



EXAMPLE 2.12 VALENCE ELECTRONS AND CORE ELECTRONS

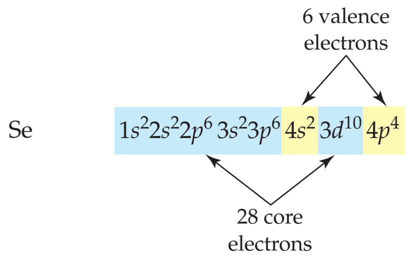
Write an electron configuration for selenium and identify the valence electrons and the core electrons.

SOLUTION

Write the electron configuration for selenium by determining the total number of electrons from selenium's atomic number (34) and distributing them into the appropriate orbitals.

$$\text{Se} \quad 1s^22s^22p^63s^23p^64s^23d^{10}4p^4$$

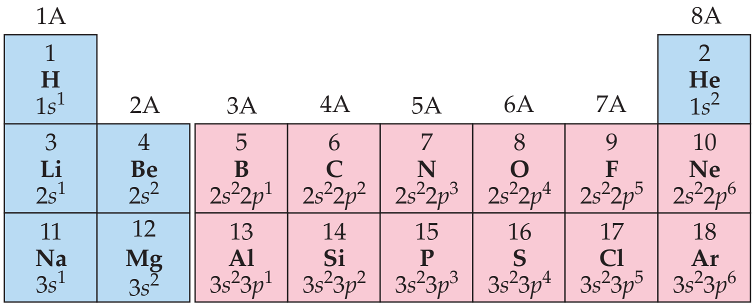
The valence electrons are those in the outermost principal shell. For selenium, the outermost principal shell is the $n = 4$ shell, which contains 6 electrons (2 in the $4s$ orbital and 4 in the three $4p$ orbitals). All other electrons, including those in the $3d$ orbitals, are core electrons.



▶ SKILLBUILDER 2.12 | Valence Electrons and Core Electrons

Write an electron configuration for chlorine and identify the valence electrons and core electrons.

▶ FOR MORE PRACTICE Example 2.23; Problems 104, 105.



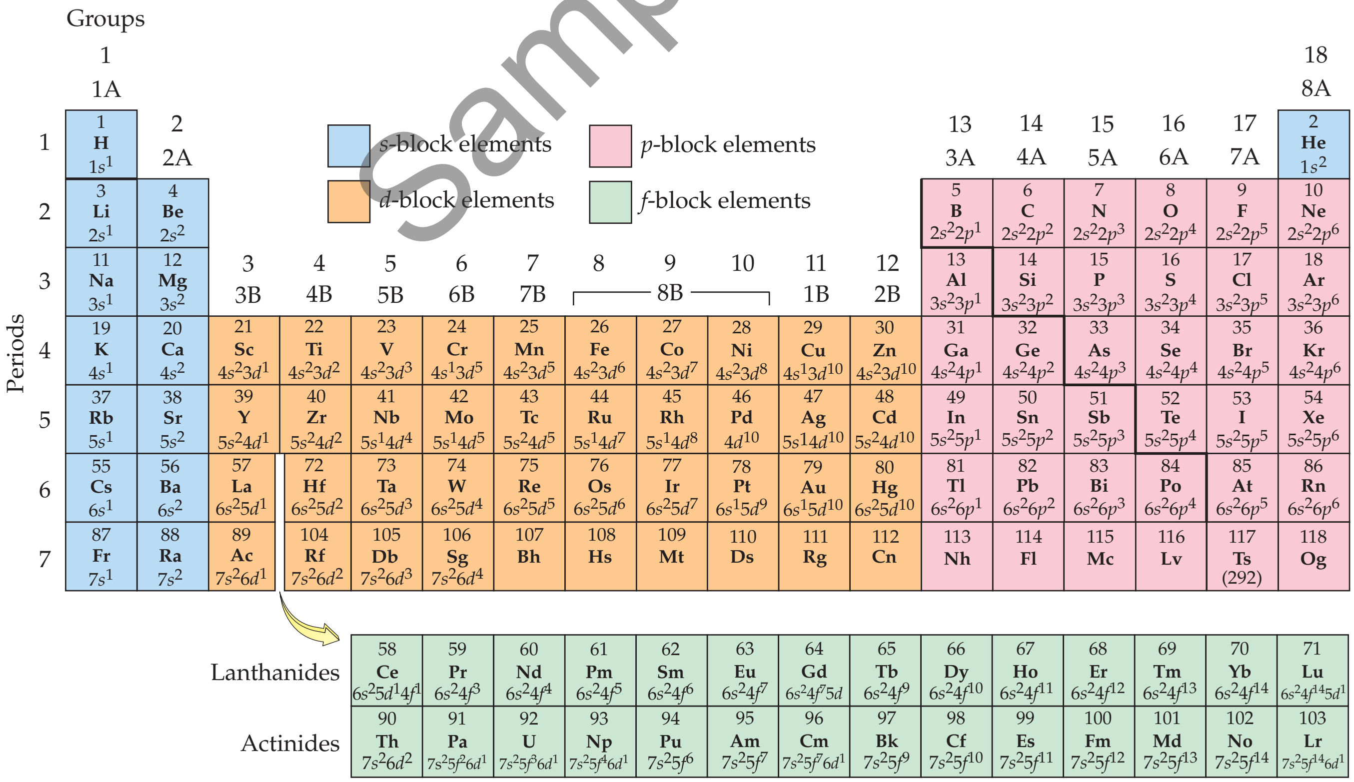
1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$	5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$	13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$

▶ **FIGURE 2.20 Outer electron configurations of the first 18 elements**

▲ Figure 2.20 shows the first 18 elements in the periodic table with an outer electron configuration listed below each one. As we move across a row, the orbitals are simply filling in the correct order. As we move down a column, the highest principal quantum number increases, but the number of electrons in each subshell remains the same. Consequently, the elements within a column (or group) all have the same number of valence electrons and similar outer electron configurations.

A similar pattern exists for the entire periodic table (▼ Figure 2.21). Notice that, because of the filling order of orbitals, we can divide the periodic table into blocks representing the filling of particular subshells.

- The first two columns on the left side of the periodic table are the *s* block with outer electron configurations of ns^1 (first column) and ns^2 (second column).
- The six columns on the right side of the periodic table are the *p* block with outer electron configurations of: ns^2np^1, ns^2np^2, ns^2np^3, ns^2np^4, ns^2np^5 (halogens), and ns^2np^6 (noble gases).
- The transition metals are the *d* block.
- The lanthanides and actinides (also called the inner transition metals) are the *f* block.



Groups

Periods	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cn**	113 **Nh**	114 **Fl**	115 **Mc**	116 **Lv**	117 **Ts** (292)	118 **Og**

🟦 *s*-block elements
🟧 *d*-block elements
🟥 *p*-block elements
🟩 *f*-block elements

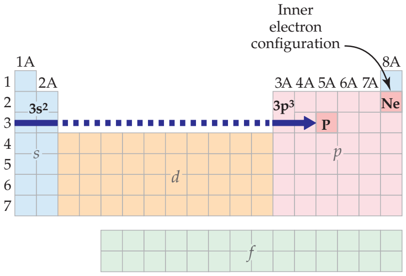
Lanthanides	58 **Ce** $6s^25d^14f^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
Actinides	90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

▲ **FIGURE 2.21 Outer electron configurations of the elements**

Remember that main-group elements are those in the two far–left columns (1A, 2A) and the six far–right columns (3A–8A) of the periodic table (see Section 2.5).

Notice that, except for helium, the number of valence electrons for any main-group element is equal to the group number of its column. For example, we can tell that chlorine has 7 valence electrons because it is in the column with group number 7A. The row number in the periodic table is equal to the number of the highest principal shell (n value). For example, since chlorine is in row 3, its highest principal shell is the $n = 3$ shell.

The transition metals have electron configurations with trends that differ somewhat from main-group elements. As we move across a row in the d block, the d orbitals are filling (▲ Figure 2.21). However, the principal quantum number of the d orbital being filled across each row in the transition series is equal to the row number minus one (in the fourth row, the $3d$ orbitals fill; in the fifth row, the $4d$ orbitals fill; and so on). For the first transition series, the outer configuration is $4s^2 3d^x$ (x = number of d electrons) with two exceptions: Cr is $4s^1 3d^5$ and Cu is $4s^1 3d^{10}$. These exceptions occur because a half-filled d subshell and a completely filled d subshell are particularly stable. Otherwise, the number of outershell electrons in a transition series does not change as we move across a period. In other words, *the transition series represents the filling of core orbitals, and the number of outershell electrons is mostly constant*.



▲ **FIGURE 2.22 Electron configuration of phosphorus** Determining the electron configuration for P from its position in the periodic table.

We can now see that the organization of the periodic table allows us to write the electron configuration for any element based simply on its position in the periodic table. For example, suppose we want to write an electron configuration for P. The inner electrons of P are those of the noble gas that precedes P in the periodic table, Ne. So we can represent the inner electrons with [Ne]. We obtain the outer electron configuration by tracing the elements between Ne and P and assigning electrons to the appropriate orbitals (◀ Figure 2.22). Remember that the highest n value is given by the row number (3 for phosphorus). So we begin with [Ne], then add in the two $3s$ electrons as we trace across the s block, followed by three $3p$ electrons as we trace across the p block to P, which is in the third column of the p block. The electron configuration is:

$$\text{P} \qquad [\text{Ne}]3s^2 3p^3$$

Notice that P is in column 5A and therefore has 5 valence electrons and an outer electron configuration of ns^2np^3.

To summarize writing an electron configuration for an element based on its position in the periodic table:

- The inner electron configuration for any element is the electron configuration of the noble gas that immediately precedes that element in the periodic table. We represent the inner configuration with the symbol for the noble gas in brackets.
- We can determine the outer electrons from the element's position within a particular block (s, p, d, or f) in the periodic table. We trace the elements between the preceding noble gas and the element of interest, and assign electrons to the appropriate orbitals.
- The highest principal quantum number (highest n value) is equal to the row number of the element in the periodic table.
- For any element containing d electrons, the principal quantum number (n value) of the outermost d electrons is equal to the row number of the element minus 1.

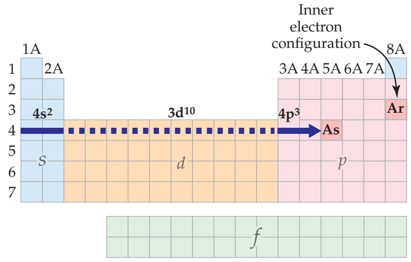
EXAMPLE 2.13 WRITING ELECTRON CONFIGURATIONS FROM THE PERIODIC TABLE

Write an electron configuration for arsenic based on its position in the periodic table.

SOLUTION

The noble gas that precedes arsenic in the periodic table is argon, so the inner electron configuration is [Ar]. Obtain the outer electron configuration by tracing the elements between Ar and As and assigning electrons to the appropriate orbitals.

Remember that the highest *n* value is given by the row number (4 for arsenic). So, begin with [Ar], then add in the two 4*s* electrons as you trace across the *s* block, followed by ten 3*d* electrons as you trace across the *d* block (the *n* value for *d* subshells is equal to the row number minus one), and finally the three 4*p* electrons as you trace across the *p* block to As, which is in the third column of the *p* block:



The electron configuration is:

$$\text{As} \qquad [\text{Ar}]4s^2 3d^{10} 4p^3$$

▶ SKILLBUILDER 2.13 | Writing Electron Configurations from the Periodic Table

Use the periodic table to determine the electron configuration for tin.

▶ FOR MORE PRACTICE Example 2.24; Problems 112, 113, 114, 115.

✓ CONCEPTUAL CHECKPOINT 2.10

Which element has the *fewest* valence electrons?

(a) B **(b)** Ca **(c)** O
(d) K **(e)** Ga

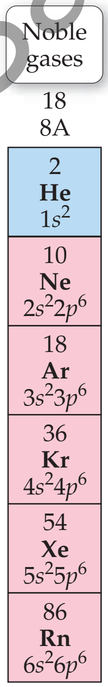
2.9 The Explanatory Power of the Quantum-Mechanical Model

LO: Recognize that the chemical properties of elements are largely determined by the number of valence electrons they contain.

We can now see how *the chemical properties of elements are largely determined by the number of valence electrons they contain*. The properties of elements vary in a periodic fashion because the number of valence electrons is periodic.

Because elements within a group in the periodic table have the same number of valence electrons, they also have similar chemical properties. The noble gases, for example, all have 8 valence electrons, except for helium, which has 2 (◀ Figure 2.23). Although we don't get into the quantitative (or numerical) aspects of the quantum-mechanical model in this book, calculations show that atoms with 8 valence electrons (or 2 for helium) are particularly low in energy, and therefore stable. The noble gases are indeed chemically stable and thus relatively inert or nonreactive as accounted for by the quantum model.

Elements with electron configurations close to the noble gases are the most reactive because they can attain noble gas electron configurations by losing or gaining a small number of electrons. Alkali metals (Group 1) are among the most reactive metals since their outer electron configuration (*ns*¹) is 1 electron beyond a noble gas



▶ FIGURE 2.23 Electron configurations of the noble gases The noble gases (except for helium) all have 8 valence electrons and completely full outer principal shells.

configuration (◀ Figure 2.24). If they can react to lose the ns^1 electron, they attain a noble gas configuration. This explains why—as we learned earlier in this module—the Group 1A metals tend to form 1+ cations. As an example, consider the electron configuration of sodium:

$$\text{Na} \qquad 1s^22s^22p^63s^1$$

In reactions, sodium loses its 3*s* electron, forming a 1+ ion with the electron configuration of neon.

$$\text{Na}^+ \qquad 1s^22s^22p^6$$

$$\text{Ne} \qquad 1s^22s^22p^6$$

Similarly, alkaline earth metals, with an outer electron configuration of ns^2 also tend to be reactive metals, losing their ns^2 electrons to form 2+ cations (▼ Figure 2.25). For example, consider magnesium:

$$\text{Mg} \qquad 1s^22s^22p^63s^2$$

In reactions, magnesium loses its two 3*s* electrons, forming a 2+ ion with the electron configuration of neon.

$$\text{Mg}^{2+} \qquad 1s^22s^22p^6$$

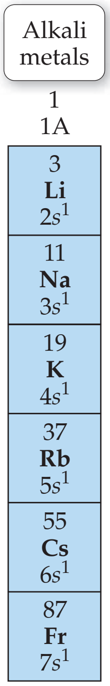
On the other side of the periodic table, halogens are among the most reactive nonmetals because of their ns^2np^5 electron configurations (▼ Figure 2.26). They are only one electron away from a noble gas configuration and tend to react to gain that one electron, forming 1– ions. For example, consider fluorine:

$$\text{F} \qquad 1s^22s^22p^5$$

In reactions, fluorine gains an additional 2*p* electron, forming a 1– ion with the electron configuration of neon.

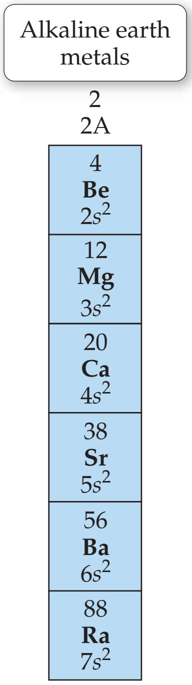
$$\text{F}^- \qquad 1s^22s^22p^6$$

The elements that form predictable ions are shown in ▶ Figure 2.27 (first introduced earlier in this module). Notice how the charge of these ions reflects their electron configurations—these elements form ions with noble gas electron configurations.

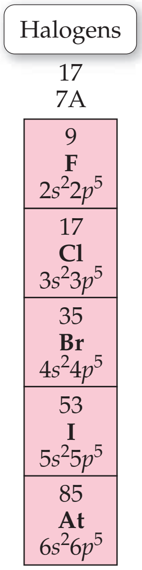


▲ **FIGURE 2.24 Electron configurations of the alkali metals** The alkali metals all have ns^1 electron configurations and are therefore 1 electron beyond a noble gas configuration. In their reactions, they tend to lose that electron, forming 1+ ions and attaining a noble gas configuration.

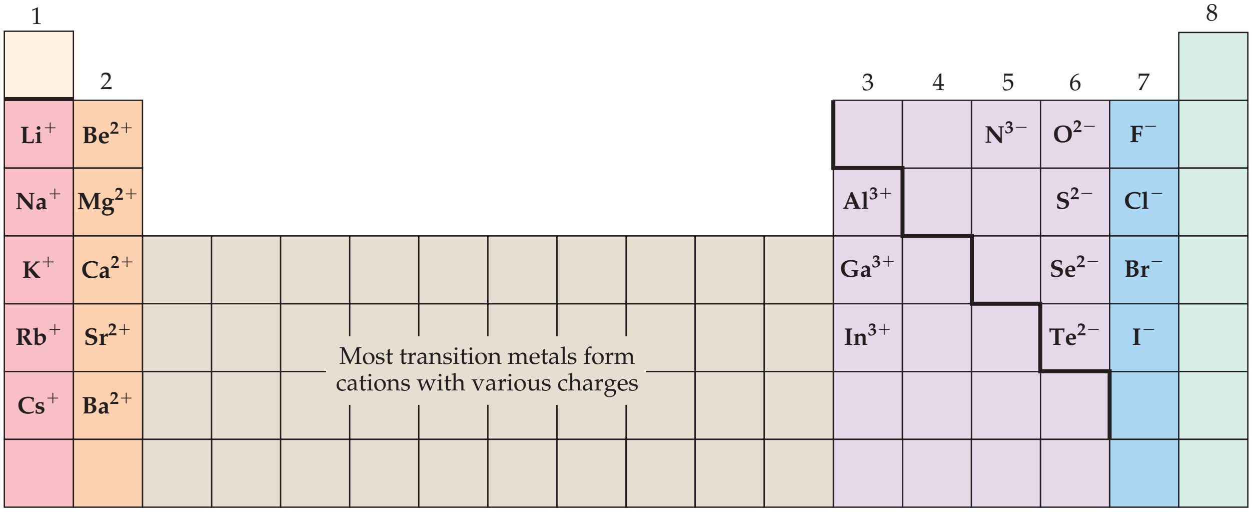
Atoms and/or ions that share the same electron configuration are termed isoelectronic.



◀ **FIGURE 2.25 Electron configurations of the alkaline earth metals** The alkaline earth metals all have ns^2 electron configurations and are therefore 2 electrons beyond a noble gas configuration. In their reactions, they tend to lose 2 electrons, forming 2+ ions and attaining a noble gas configuration.



▶ **FIGURE 2.26 Electron configurations of the halogens** The halogens all have ns^2np^5 electron configurations and are therefore 1 electron short of a noble gas configuration. In their reactions, they tend to gain 1 electron, forming 1– ions and attaining a noble gas configuration.



▲ **FIGURE 2.27** **Elements that form predictable ions**

CONCEPTUAL CHECKPOINT 2.11

Below is the electron configuration of calcium:

$$\text{Ca} \quad 1s^22s^22p^63s^23p^64s^2$$

In its reactions, calcium tends to form the Ca^{2+} ion. Which electrons are lost upon ionization?

(a) the 4*s* electrons

(b) two of the 3*p* electrons

(c) the 3*s* electrons

(d) the 1*s* electrons

2.10 Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character

LO: **Identify and understand periodic trends such as atomic size, ionization energy, and metallic character.**

The quantum-mechanical model also explains other periodic trends such as atomic size, ionization energy, and metallic character. We examine these trends individually in this section of the module.

Atomic Size

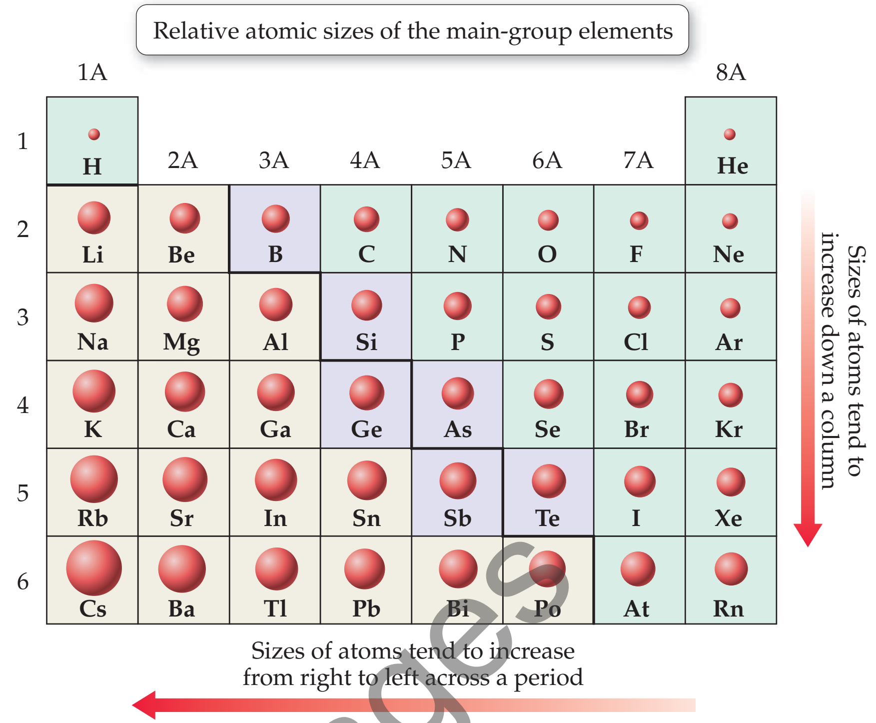
The **atomic size** of an atom is determined by the distance between the outermost electrons and the nucleus. As we move across a period in the periodic table, we know that electrons occupy orbitals with the same principal quantum number, *n*. Since the principal quantum number largely determines the size of an orbital, electrons are therefore filling orbitals of approximately the same size, and we might expect atomic size to remain constant across a period. However, with each step across a period, the number of protons in the nucleus also increases. This increase in the number of protons results in a greater pull on the electrons from the nucleus, causing atomic size to actually decrease. Therefore:

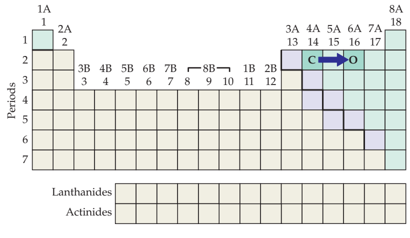
> As we move to the left across a period, or row, in the periodic table, atomic size increases, as shown in ▶Figure 2.28.

As we move down a column in the periodic table, the highest principal quantum number, *n*, increases. Because the size of an orbital increases with increasing principal quantum number, the electrons that occupy the outermost orbitals are farther from the nucleus as we move down a column. Therefore:

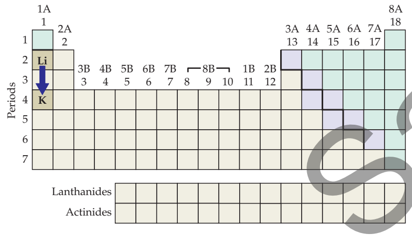
> As we move down a column, or group in the periodic table, atomic size increases, as shown in ▶ Figure 2.28.

▶ **FIGURE 2.28 Periodic properties: atomic size** Atomic size increases as we move to the left across a period and increases as we move down a column in the periodic table.

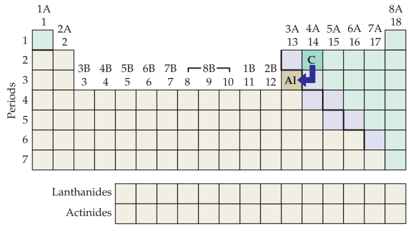




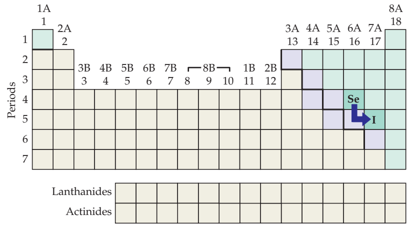
(a)



(b)



(c)



(d)

EXAMPLE 2.14 ATOMIC SIZE

Choose the larger atom in each pair.

(a) C or O **(b)** Li or K **(c)** C or Al **(d)** Se or I

SOLUTION

(a) C or O

Carbon atoms are larger than O atoms because, as you trace the path between C and O on the periodic table, you move to the right within the same period. Atomic size decreases as you go to the right.

(b) Li or K

Potassium atoms are larger than Li atoms because, as you trace the path between Li and K on the periodic table, you move down a column. Atomic size increases as you go down a column.

(c) C or Al

Aluminum atoms are larger than C atoms because, as you trace the path between C and Al on the periodic table you move down a column (atomic size increases) and then to the left across a period (atomic size increases). These effects add together for an overall increase.

(d) Se or I

Based on periodic properties alone, you cannot tell which atom is larger because as you trace the path between Se and I, you go down a column (atomic size increases) and then to the right across a period (atomic size decreases). These effects tend to cancel one another.

▶ SKILLBUILDER 2.14 | Atomic Size

Choose the larger atom in each pair.

(a) Pb or Po **(b)** Rb or Na **(c)** Sn or Bi **(d)** F or Se

▶ **FOR MORE PRACTICE** Example 2.25a; Problems 136, 137, 138, 139.

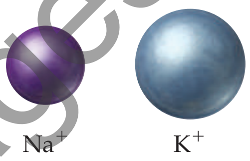
CHEMISTRY **AND HEALTH**

▶ Pumping Ions: Atomic Size and Nerve Impulses

No matter what you are doing at this moment, tiny pumps in each of the trillions of cells that make up your body are hard at work. These pumps, located in the cell membrane, move a number of different ions into and out of the cell. The most important of these ions are sodium (Na^+) and potassium (K^+) which happen to be pumped in opposite directions. Sodium ions are pumped *out of cells*, while potassium ions are pumped *into cells*. The result is a *chemical gradient* for each ion: The concentration of sodium is higher outside the cell than within, while exactly the opposite is true for potassium.

The ion pumps within the cell membrane are analogous to water pumps in a high-rise building that pump water against the force of gravity to a tank on the roof. Other structures within the membrane, called ion channels, are like the building's faucets. When they open momentarily, bursts of sodium and potassium ions, driven by their concentration gradients, flow back across the membrane—sodium flowing in and potassium flowing out. These ion pulses are the basis for the transmission of nerve signals in the brain, heart, and throughout the body. Consequently, every move you make or every thought you have is mediated by the flow of these ions.

How do the pumps and channels differentiate between sodium and potassium ions? How do the ion pumps selectively move sodium out of the cell and potassium into the cell? To answer this question, we must examine the sodium and potassium ions more closely. In what ways do they differ? Both are cations of Group I metals. All Group I metals tend to lose one electron to form cations with 1+ charge, so the magnitude of the charge cannot be the decisive factor. But potassium (atomic number 19) lies directly below sodium in the periodic table (atomic number 11) and based on periodic properties is therefore larger than sodium. The potassium ion has a radius of 133 pm, while the sodium ion has a radius of 95 pm (1 pm = 10^{-12} m). The pumps and channels within cell membranes are so sensitive that they distinguish between the sizes of these two ions and selectively allow only one or the other to pass. The result is the transmission of nerve signals that allows you to read this page.



B2.2 CAN YOU ANSWER THIS? *Other ions, including calcium and magnesium, are also important to nerve signal transmission. Arrange these four ions in order of increasing size:* K^+, Na^+, Mg^{2+}, *and* Ca^{2+}.

Ionization Energy

The **ionization energy** of an atom is the energy required to remove n electron from the atom in the gaseous state. For example, the ionization of sodium can be represented with the equation:

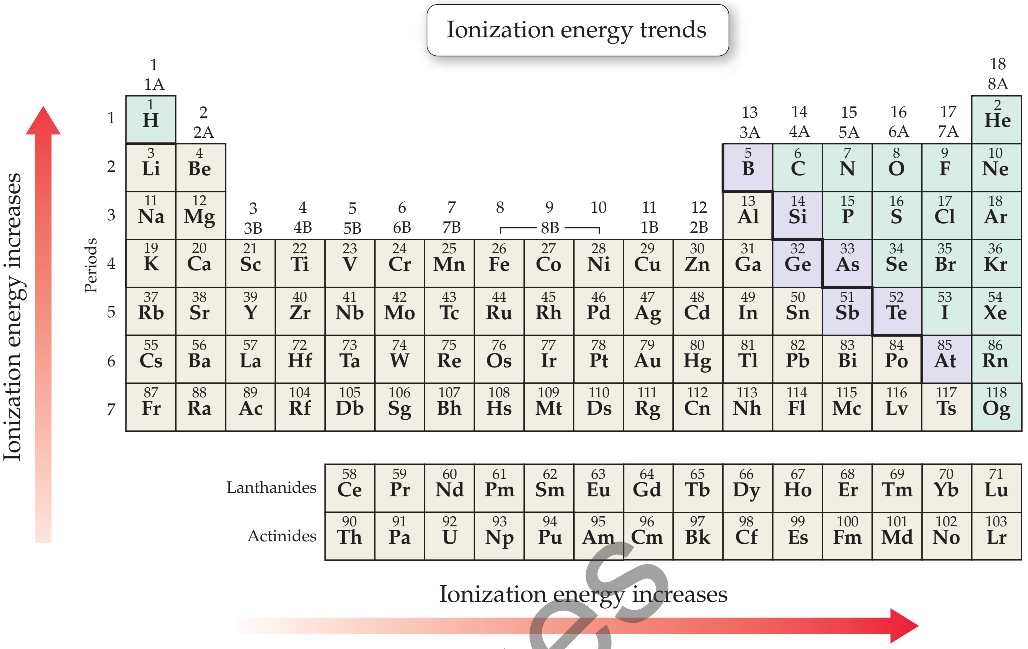
$$\text{Na} + \text{Ionization energy} \rightarrow \text{Na}^+ + 1\text{e}^-$$

Based on what we know about electron configurations, what can we predict about ionization energy trends? Would it take more or less energy to remove an electron from Na than from Cl? The explanation is exactly in line with that for the trend in atomic radius. With each step across a period, the number of protons increases. Since this increase in the number of protons results in a greater pull on the electrons from the nucleus, the energy required to remove the outermost electron increases. It is easier to remove an electron from sodium than it is from chlorine. We can generalize this idea in this statement:

> As we move across a period, or row, to the right in the periodic table, ionization energy increases (▶Figure 2.29).

What happens to ionization energy as we move down a column? As we have learned, the principal quantum number, n, increases as we move down a column. Within a given subshell, orbitals with higher principal quantum numbers are larger than orbitals with smaller principal quantum numbers. Consequently, electrons in the outermost principal shell are farther away from the positively charged nucleus—and therefore are held less tightly—as we move down a column. This

▶ **FIGURE 2.29 Periodic properties: ionization energy** Ionization energy increases as we move to the right across a period and increases as we move up a column in the periodic table.



results in a lower ionization energy (if the electron is held less tightly, it is easier to pull away) as we move down a column. Therefore:

> As we move up a column (or group) in the periodic table, ionization energy increases (▲ Figure 2.29).

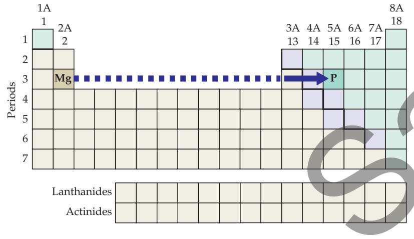
Notice that the trends in ionization energy are consistent with the trends in atomic size. Smaller atoms are more difficult to ionize because their electrons are held more tightly. Therefore, as we move across a period, atomic size decreases and ionization energy increases. Similarly, as we move down a column, atomic size increases and ionization energy decreases since electrons are farther from the nucleus and are therefore less tightly held.

EXAMPLE 2.15 IONIZATION ENERGY

Choose the element with the higher ionization energy from each pair.

(a) Mg or P
(b) As or Sb
(c) N or Si
(d) O or Cl

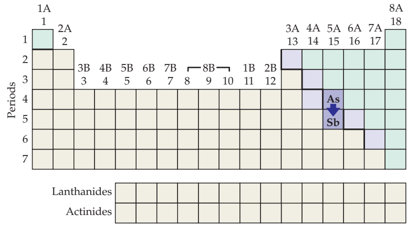
SOLUTION



(a)

(a) Mg or P

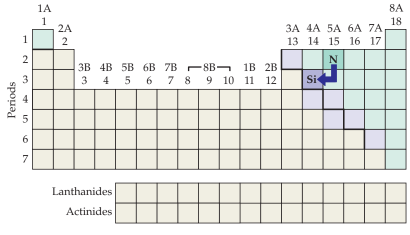
P has a higher ionization than Mg because, as you trace the path between Mg and P on the periodic table, you move to the right within the same period. Ionization energy increases as you go to the right.



(b)

(b) As or Sb

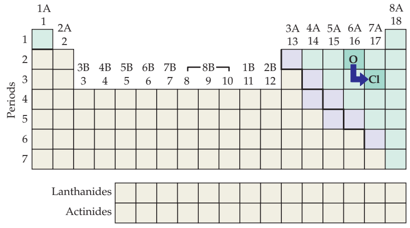
As has a higher ionization energy than Sb because, as you trace the path between As and Sb on the periodic table, you move down a column. Ionization energy decreases as you go down a column.



(c)

(c) N or Si

N has a higher ionization energy than Si because, as you trace the path between N and Si on the periodic table, you move down a column (ionization energy decreases) and then to the left across a period (ionization energy decreases). These effects sum together for an overall decrease.



(d)

(d) O or Cl
Based on periodic properties alone, you cannot tell which has a higher ionization energy because, as you trace the path between O and Cl you move down a column (ionization energy decreases) and then to the right across a period (ionization energy increases). These effects tend to cancel.

▶ SKILLBUILDER 2.15 | Ionization Energy

Choose the element with the higher ionization energy from each pair.

(a) Mg or Sr
(b) In or Te
(c) C or P
(d) F or S

▶ FOR MORE PRACTICE Example 2.25b; Problems 132, 133, 134, 135.

Metallic Character

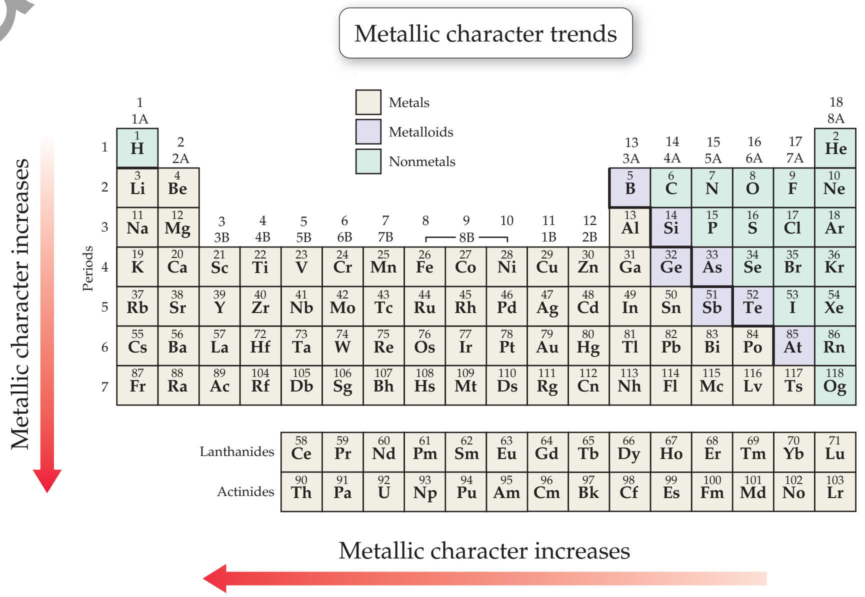
As we learned earlier in this module, metals tend to lose electrons in their chemical reactions, while nonmetals tend to gain electrons. As we move left across a period in the periodic table, ionization energy decreases, which means that electrons are more likely to be lost in chemical reactions. Consequently:

> As we move across a period, or row, to the left in the periodic table, **metallic character** increases (▼ Figure 2.30).

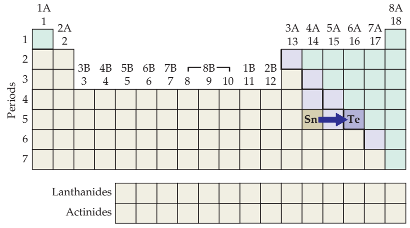
As we move down a column in the periodic table, ionization energy decreases, making electrons more likely to be lost in chemical reactions. Consequently:

> As we move down a column, or group in the periodic table, metallic character increases (▼ Figure 2.30).

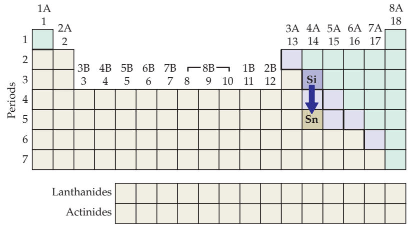
These trends, based on the quantum-mechanical model, explain the distribution of metals and nonmetals that we were introduced to earlier in this module. Metals are found toward the left side of the periodic table and nonmetals (with the exception of hydrogen) toward the upper right.



▲ FIGURE 2.30 Periodic properties: metallic character Metallic character decreases as we move to the right across a period and increases as we move down a column in the periodic table.



(a)



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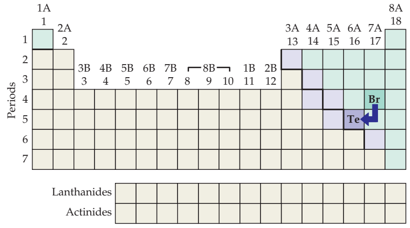

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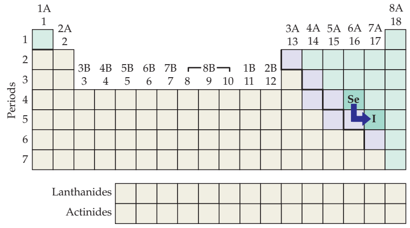


(d)

EXAMPLE 2.16 METALLIC CHARACTER

Choose the more metallic element from each pair.

(a) Sn or Te
(b) Si or Sn
(c) Br or Te
(d) Se or I

SOLUTION

(a) Sn or Te

Sn is more metallic than Te because, as you trace the path between Sn and Te on the periodic table, you move to the right within the same period. Metallic character decreases as you go to the right.

(b) Si or Sn

Sn is more metallic than Si because, as you trace the path between Si and Sn on the periodic table, you move down a column. Metallic character increases as you go down a column.

(c) Br or Te

Te is more metallic than Br because, as you trace the path between Br and Te on the periodic table, you move down a column (metallic character increases) and then to the left across a period (metallic character increases). These effects add together for an overall increase.

(d) Se or I

Based on periodic properties alone, you cannot tell which is more metallic because, as you trace the path between Se and I, you go down a column (metallic character increases) and then to the right across a period (metallic character decreases). These effects tend to cancel.

▶ SKILLBUILDER 2.16 | Metallic Character

Choose the more metallic element from each pair.

(a) Ge or In
(b) Ga or Sn
(c) P or Bi
(d) B or N

▶ **FOR MORE PRACTICE** Example 2.25c; Problems 140, 141, 142, 143.

CONCEPTUAL CHECKPOINT 2.12

Which property *increases* as you move from left to right across a row in the periodic table?

(a) Atomic size
(b) Ionization energy
(c) Metallic character

MODULE IN REVIEW

Self-Assessment Quiz

Q1. Which statement best summarizes the nuclear model of the atom?

(a) The atom is composed of a dense core that contains most of its mass and all of its positive charge, while low-mass negatively charged particles compose most of its volume.
(b) The atom is composed of a sphere of positive charge with many negatively charged particles within the sphere.
(c) Most of the mass of the atom is evenly distributed throughout its volume.
(d) All of the particles that compose an atom have exactly the same mass.

Q2. How many neutrons does the Fe-56 isotope contain?

(a) 26 **(b)** 30 **(c)** 56 **(d)** 112

Q3. Determine the number of protons, neutrons, and electrons in $^{32}_{16}S^{2-}$.

(a) 16 protons; 32 neutrons; and 18 electrons
(b) 16 protons; 16 neutrons; and 18 electrons
(c) 32 protons; 16 neutrons; and 2 electrons
(d) 16 protons; 48 neutrons; and 16 electrons

Q4. An element has four naturally occurring isotopes; the table lists the mass and natural abundance of each isotope. Find the atomic mass of the element.

Isotope	Mass (amu)	Natural abundance
A	203.9730	1.4 %
B	205.9744	24.1 %
C	206.9758	22.1 %
D	207.9766	52.4 %

(a) 207.2 amu
(b) 2.072×10^4 amu
(c) 206.2 amu
(d) 206.5 amu

Q5. An ion composed of which of these particles would have a mass of approximately 16 amu and a charge of 2–?

(a) 8 protons and 8 electrons
(b) 8 protons, 8 neutrons, and 10 electrons
(c) 8 protons, 8 neutrons, and 8 electrons
(d) 8 protons, 8 neutrons, and 6 electrons

Q6. What is the charge of the Cr ion that contains 21 electrons?

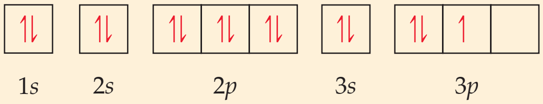
(a) 2– **(b)** 3– **(c)** 2+ **(d)** 3+

Q7. What is the electron configuration of arsenic (As)?

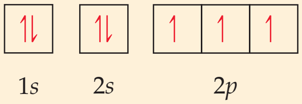
(a) $[Ar]4s^2 4p^3$ **(c)** $[Ar]4s^2 3d^6 4p^3$
(b) $[Ar]4s^2 4d^{10} 4p^3$ **(d)** $[Ar]4s^2 3d^{10} 4p^3$

Q8. Which orbital diagram corresponds to phosphorus (P)?

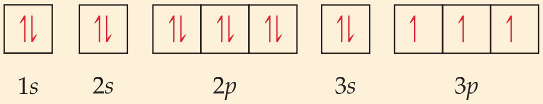
(a)



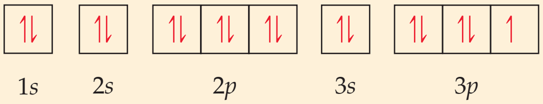
(b)



(c)



(d)



Q9. How many valence electrons does tellurium (Te) have?

(a) 5 **(b)** 6 **(c)** 16 **(d)** 52

Q10. Which element is a main-group metal with an even atomic number?

(a) K **(b)** Ca **(c)** Cr **(d)** Se

Q11. Which element is a halogen?

(a) Ne **(b)** O **(c)** Ca **(d)** I

Q12. Which pair of elements has the most similar properties?

(a) Sr and Ba **(b)** S and Ar
(c) H and He **(d)** K and Se

Q13. Which element is a row 4 noble gas?

(a) Ne **(b)** Br **(c)** Zr **(d)** Kr

Q14. How many electrons does the predictable (most common) ion of fluorine contain?

(a) 1 **(b)** 4 **(c)** 9 **(d)** 10

Q15. The element sulfur forms an ion with what charge?

(a) 2– **(b)** 1– **(c)** 1+ **(d)** 2+

Q16. When aluminum forms an ion, it loses electrons. How many electrons does it lose, and which orbitals do the electrons come from?

(a) 1 electron from the 3*s* orbital
(b) 2 electrons; one from the 3*s* orbital and one from the 2*s* orbital
(c) 3 electrons; two from the 3*s* orbital and one from the 3*p* orbital
(d) 5 electrons from the 3*p* orbital

Q17. Order the elements Sr, Ca, and Se in order of decreasing atomic size.

(a) Se > Sr > Ca **(c)** Sr > Ca > Se
(b) Ca > Se > Sr **(d)** Se > Ca > Sr

Q18. Which element has the highest ionization energy?

(a) Sn **(b)** S **(c)** Si **(d)** F

Q19. Which element is most metallic?

(a) Al **(b)** N **(c)** P **(d)** O

Q20. Which property decreases as you move down a column in the periodic table?

(a) Atomic size
(b) Ionization energy
(c) Metallic character
(d) None of the above (all increase as you move down a column).

Answers: 1:a; 2:b; 3:b; 4:a; 5:b; 6:d; 7:d; 8:c; 9:b; 10:b; 11:d; 12:a; 13:d; 14:d; 15:a; 16:c; 17:c; 18:d; 19:a; 20:b