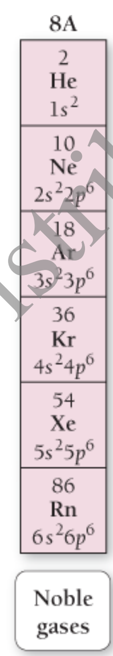
Pearson

3.5: Electron Configurations and Elemental Properties

As we discussed in Section 3.4, *the chemical properties of elements are largely determined by the number of valence electrons the elements contain*. The properties of elements are periodic because the number of valence electrons is periodic. Mendeleev grouped elements into families (or columns) based on observations about their properties. We now know that elements in a family have the same number of valence electrons. In other words, elements in a family have similar properties because they have the same number of valence electrons.

Perhaps the most striking family in the periodic table is the column labeled 8A, known as the **noble gases**. The noble gases are generally inert—they are the most unreactive elements in the entire periodic table. Why? Notice that each noble gas has eight valence electrons (or two in the case of helium), and they all have full outer quantum levels. We do not cover the quantitative (or numerical) aspects of the quantum-mechanical model in this book, but calculations of the overall energy of the electrons within atoms with eight valence electrons (or two for helium) show that these atoms are particularly stable. *In other words, when a quantum level is completely full, the overall potential energy of the electrons that occupy that level is particularly low*.



The noble gases each have eight valence electrons except for helium, which has two. They have full outer quantum levels and are particularly stable and unreactive.

Recall from Section E.6 that, on the one hand, systems with high potential energy tend to change in ways that lower their potential energy. Systems with low potential energy, on the other hand, tend not to change—they are stable. Because atoms with eight electrons (or two for helium) have particularly low potential energy, the noble gases are stable—they *cannot* lower their energy by reacting with other atoms or molecules.

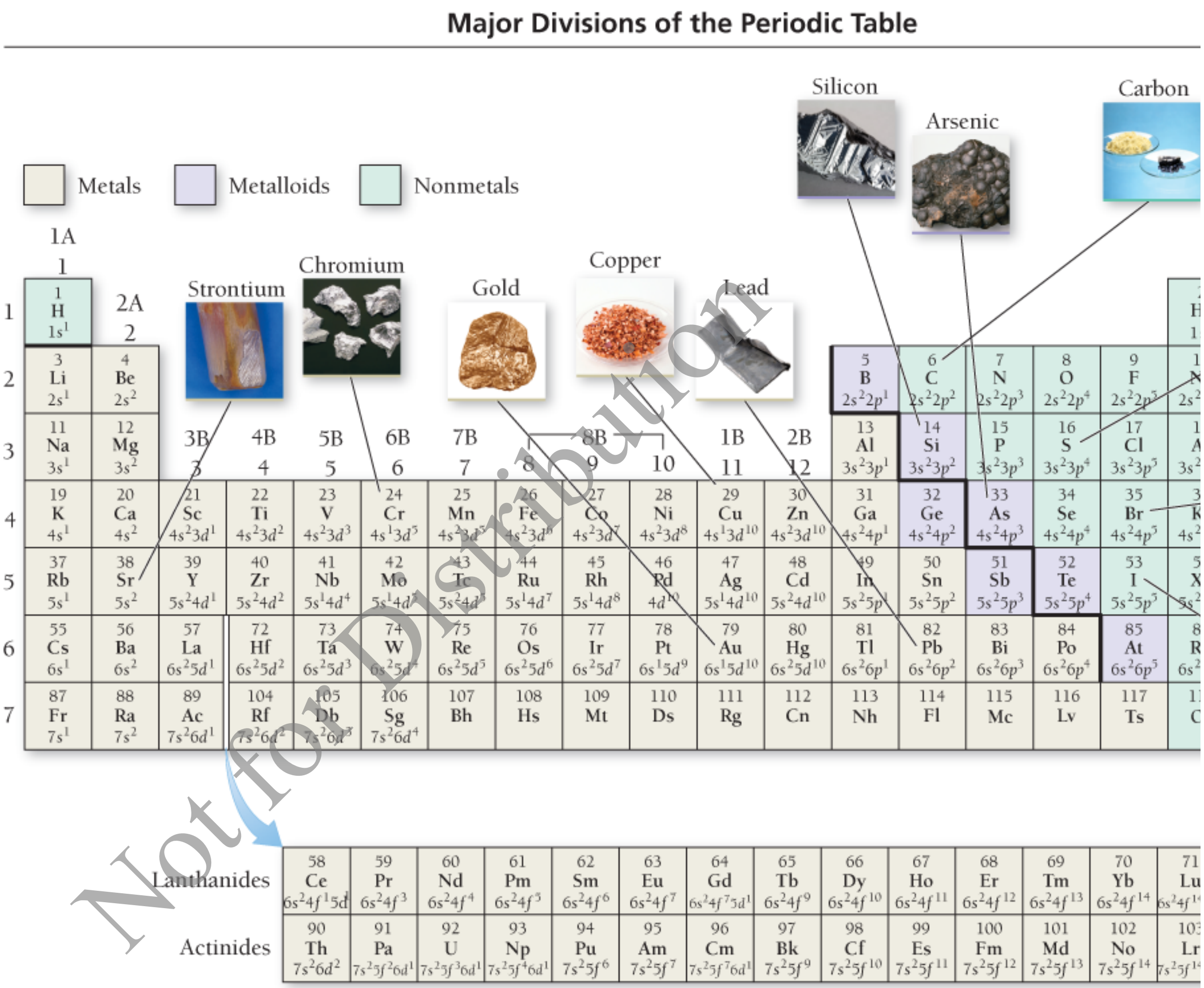
We can explain a great deal of chemical behavior with the simple idea that *elements without a noble gas electron configuration react to attain a noble gas configuration*. This idea applies particularly well to main-group elements. In this section, we first apply this idea to help differentiate between metals and nonmetals. We then apply the idea to understand the properties of several individual families of elements. Lastly, we apply the idea to the formation of ions.

Metals and Nonmetals

We can understand the broad chemical behavior of the elements by superimposing one of the most general properties of an element—whether it is a metal or nonmetal—with its outer electron configuration in the form of a periodic table (Figure 3.11). **Metals** lie on the lower left side and middle of the periodic table and share some common 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 most importantly, *they tend to lose electrons when they undergo chemical changes*.

Figure 3.11 Metallic Behavior and Electron Configuration

The elements in the periodic table fall into three broad classes: metals, metalloids, and nonmetals. Notice the correlations between elemental properties and electron configurations.



For example, sodium is among the most reactive metals. Its electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^1$. Notice that its electron configuration is one electron beyond the configuration of neon, a noble gas. Sodium can attain a noble gas electron configuration by losing that one valence electron—and that is exactly what it does. When we find sodium in nature, we most often find it as Na^+, which has the electron configuration of neon $(1s^2\ 2s^2\ 2p^6)$. The other main-group metals in the periodic table behave similarly: They tend to lose their valence electrons in chemical changes to attain noble gas electron configurations. The transition metals also tend to lose electrons in their chemical changes, but they do not generally attain noble gas electron configurations.

Nonmetals lie on the upper right side of the periodic table. The division between metals and nonmetals is the zigzag diagonal line running from boron to astatine. Nonmetals have varied properties—some are solids at room temperature, others are liquids or gases—but as a whole they tend to be poor conductors of heat and electricity, and most importantly *they all tend to gain electrons when they undergo chemical changes*.

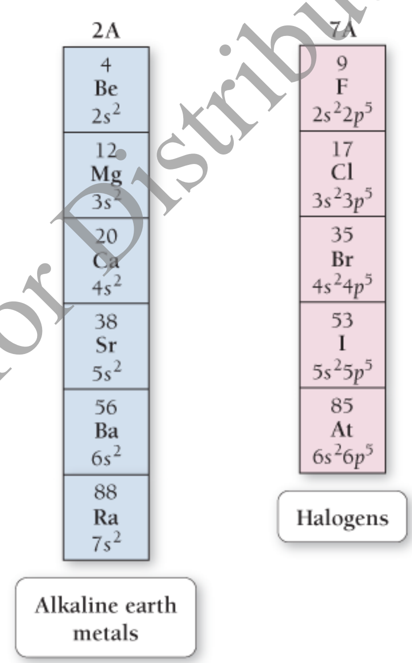
Chlorine is among the most reactive nonmetals. Its electron configuration is $1s^2\ 2s^2\ 2p^6 3s^2 3p^5$. Notice that its electron configuration is one electron short of the configuration of argon, a noble gas. Chlorine can attain a noble gas electron configuration by gaining one electron—and that is exactly what it does. When we find chlorine in nature, we often find it as Cl^-, which has the electron configuration of argon $(1s^2\ 2s^2\ 2p^6 3s^2 3p^6)$. The other nonmetals in the periodic table behave similarly: They tend to gain electrons in chemical changes to attain noble gas electron configurations.

Many of the elements that lie along the zigzag diagonal line that divides metals and nonmetals are **metalloids** and exhibit mixed properties. Several metalloids are classified as **semiconductors** because of their intermediate (and highly temperature-dependent) electrical conductivity. Our ability to change and control the conductivity of semiconductors makes them useful in the manufacture of the electronic chips and circuits central to computers, cellular telephones, and many other devices. Examples of metalloids include silicon, arsenic, and antimony.

> Metalloids are sometimes called semimetals.

Families of Elements

We can also understand the properties of families of elements (those in the same column in the periodic table) based on their electron configurations. We have already seen that the group 8A elements, called the *noble gases*, have eight valence electrons and are mostly unreactive. The most familiar noble gas is probably helium, used to fill buoyant balloons. Helium is chemically stable—it does not combine with other elements to form compounds—and is therefore safe to put into balloons. Other noble gases are neon (often used in electronic signs), argon (a small component of our atmosphere), krypton, and xenon.



The group 1A elements, called the **alkali metals**, all have an outer electron configuration of ns^1. Like sodium, a member of this family, the alkali metals have electron configurations that are one electron beyond a noble gas electron configuration. In their reactions, alkali metals readily, and sometimes violently, lose the ns^1 electron to form ions with a 1+ charge. A marble-sized piece of sodium, for example, explodes violently when dropped into water. Lithium, potassium, and rubidium are also alkali metals.

The group 2A elements, called the **alkaline earth metals**, all have an outer electron configuration of ns^2. They have electron configurations that are two electrons beyond a noble gas configuration. In their reactions, they tend to lose the two ns^2 electrons—though not quite as violently as the alkali metals—to form ions with a 2+

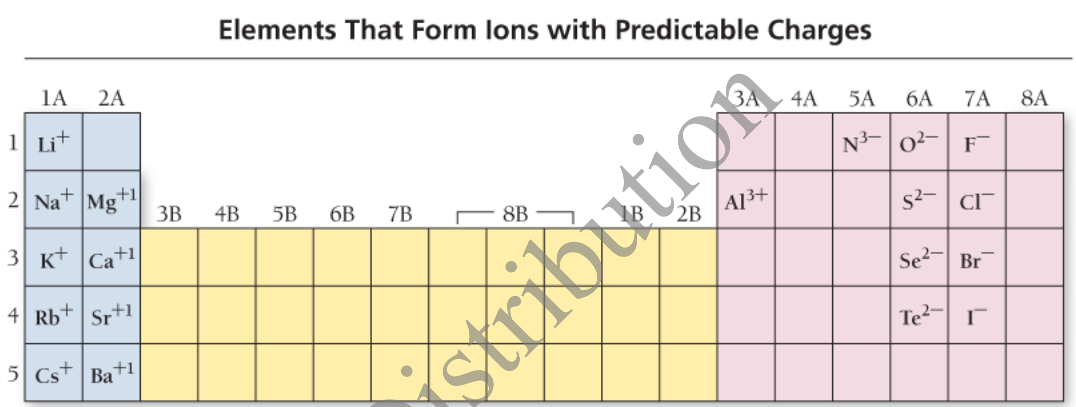
charge. Calcium, for example, reacts fairly vigorously when dropped into water but does not explode as dramatically as sodium. Magnesium (a common low-density structural metal), strontium, and barium are other alkaline earth metals.

The group 7A elements, the **halogens**, all have an outer electron configuration of $ns^2\ np^5$. Like chlorine, a member of this family, their electron configurations are one electron short of a noble gas configuration. Consequently, in their reactions with metals, halogens tend to gain one electron to form ions with a 1– charge. Chlorine, a greenish-yellow gas with a pungent odor, is one of the most familiar halogens. Because of its reactivity, chlorine is used as a sterilizing and disinfecting agent. Other halogens are bromine, a red-brown liquid that easily evaporates into a gas; iodine, a purple solid; and fluorine, a pale-yellow gas.

The Formation of Ions

In Section 1.8, we learned that atoms can lose or gain electrons to form ions. We have just seen that metals tend to form positively charged ions (cations) and nonmetals tend to form negatively charged ions (anions). A number of main-group elements in the periodic table always form ions with a noble gas electron configuration. Consequently, we can reliably predict their charges (Figure 3.12).

Figure 3.12 Elements That Form Ions with Predictable Charges



As we have already seen, the alkali metals tend to form cations with a 1+ charge, the alkaline earth metals tend to form ions with a 2+ charge, and the halogens tend to form ions with a 1– charge. In each of these cases, the ions have noble gas electron configurations. This is true of the rest of the ions in Figure 3.12. Nitrogen, for example, has an electron configuration of $1s^2\ 2s^2\ 2p^3$. The N^{3-} ion has three additional electrons and an electron configuration of $1s^2\ 2s^2\ 2p^6$, which is the same as the configuration of neon, the nearest noble gas.

Notice that, for the main-group elements that form cations with predictable charge, the charge is equal to the group number. For main-group elements that form anions with predictable charge, the charge is equal to the group number minus eight. Transition elements may form various ions with different charges.

The tendency for many main-group elements to form ions with noble gas electron configurations *does not* mean that the process is in itself energetically favorable. In fact, forming cations always requires energy, and forming anions sometimes requires energy as well. However, the energy cost of forming a cation or anion with *a noble gas configuration* is often less than the energy payback that occurs when that cation or anion forms chemical bonds, as we shall see in Chapter 4.

Example 3.5 Predicting the Charge of Ions

Predict the charges of the monoatomic (single atom) ions formed by each main-group element.

a. Al
b. S

SOLUTION

a. Aluminum is a main-group metal and tends to lose electrons to form a cation with the same electron configuration as the nearest noble gas. The electron configuration of aluminum is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$. The nearest noble gas is neon, which has an electron configuration of $1s^2\ 2s^2\ 2p^6$. Therefore, aluminum loses three electrons to form the cation Al^{3+}.

b. Sulfur is a nonmetal and tends to gain electrons to form an anion with the same electron configuration as the nearest noble gas. The electron configuration of sulfur is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. The nearest noble gas is argon, which has an electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. Therefore, sulfur gains two electrons to form the anion S^{2-}.

FOR PRACTICE 3.5 Predict the charges of the monoatomic ions formed by each main-group element.

a. N

b. Rb