

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.

8A
2 He $1s^2$
10 Ne $2s^2 2p^6$
18 Ar $3s^2 3p^6$
36 Kr $4s^2 4p^6$
54 Xe $5s^2 5p^6$
86 Rn $6s^2 6p^6$
Noble gases

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.

Major Divisions of the Periodic Table

Metals Metalloids Nonmetals

1A 1 2A 2

1 1 H 1s¹ 2 2 Li 2s¹ 3 3 Na 3s¹ 4 4 Be 2s² 5 5 B 2s²2p¹ 6 6 C 2s²2p² 7 7 N 2s²2p³ 8 8 O 2s²2p⁴ 9 9 F 2s²2p⁵ 10 10 Ne 2s²2p⁶

3 3B 4 4B 5 5B 6 6B 7 7B 8 8B 9 9B 10 10B 11 1B 12 2B

13 13 Al 3s²3p¹ 14 14 Si 3s²3p² 15 15 P 3s²3p³ 16 16 S 3s²3p⁴ 17 17 Cl 3s²3p⁵ 18 18 Ar 3s²3p⁶

19 19 K 4s¹ 20 20 Ca 4s² 21 21 Sc 4s²3d¹ 22 22 Ti 4s²3d² 23 23 V 4s²3d³ 24 24 Cr 4s¹3d⁵ 25 25 Mn 4s²3d⁵ 26 26 Fe 4s²3d⁶ 27 27 Co 4s²3d⁷ 28 28 Ni 4s²3d⁸ 29 29 Cu 4s¹3d¹⁰ 30 30 Zn 4s²3d¹⁰ 31 31 Ga 4s²4p¹ 32 32 Ge 4s²4p² 33 33 As 4s²4p³ 34 34 Se 4s²4p⁴ 35 35 Br 4s²4p⁵ 36 36 Kr 4s²4p⁶

37 37 Rb 5s¹ 38 38 Sr 5s² 39 39 Y 5s²4d¹ 40 40 Zr 5s²4d² 41 41 Nb 5s¹4d⁴ 42 42 Mo 5s¹4d⁵ 43 43 Tc 5s²4d⁵ 44 44 Ru 5s¹4d⁷ 45 45 Rh 5s¹4d⁸ 46 46 Pd 4d¹⁰ 47 47 Ag 5s¹4d¹⁰ 48 48 Cd 5s²4d¹⁰ 49 49 In 5s²5p¹ 50 50 Sn 5s²5p² 51 51 Sb 5s²5p³ 52 52 Te 5s²5p⁴ 53 53 I 5s²5p⁵ 54 54 Xe 5s²5p⁶

55 55 Cs 6s¹ 56 56 Ba 6s² 57 57 La 6s²5d¹ 72 72 Hf 6s²5d² 73 73 Ta 6s²5d³ 74 74 W 6s²5d⁴ 75 75 Re 6s²5d⁵ 76 76 Os 6s²5d⁶ 77 77 Ir 6s²5d⁷ 78 78 Pt 6s¹5d⁹ 79 79 Au 6s¹5d¹⁰ 80 80 Hg 6s²5d¹⁰ 81 81 Tl 6s²6p¹ 82 82 Pb 6s²6p² 83 83 Bi 6s²6p³ 84 84 Po 6s²6p⁴ 85 85 At 6s²6p⁵ 86 86 Rn 6s²6p⁶

87 87 Fr 7s¹ 88 88 Ra 7s² 89 89 Ac 7s²6d¹ 104 104 Rf 7s²6d² 105 105 Db 7s²6d³ 106 106 Sg 7s²6d⁴ 107 107 Bh 7s²6d⁵ 108 108 Hs 7s²6d⁶ 109 109 Mt 7s²6d⁷ 110 110 Ds 7s²6d⁸ 111 111 Rg 7s²6d⁹ 112 112 Cn 7s²6d¹⁰ 113 113 Nh 7s²7p¹ 114 114 Fl 7s²7p² 115 115 Mc 7s²7p³ 116 116 Lv 7s²7p⁴ 117 117 Ts 7s²7p⁵ 118 118 Og 7s²7p⁶

58 58 Ce 6s²4f¹5d¹ 59 59 Pr 6s²4f³ 60 60 Nd 6s²4f⁴ 61 61 Pm 6s²4f⁵ 62 62 Sm 6s²4f⁶ 63 63 Eu 6s²4f⁷ 64 64 Gd 6s²4f⁷5d¹ 65 65 Tb 6s²4f⁹ 66 66 Dy 6s²4f¹⁰ 67 67 Ho 6s²4f¹¹ 68 68 Er 6s²4f¹² 69 69 Tm 6s²4f¹³ 70 70 Yb 6s²4f¹⁴ 71 71 Lu 6s²4f¹⁴

90 90 Th 7s²6d² 91 91 Pa 7s²5f²6d¹ 92 92 U 7s²5f³6d¹ 93 93 Np 7s²5f⁴6d¹ 94 94 Pu 7s²5f⁶ 95 95 Am 7s²5f⁷ 96 96 Cm 7s²5f⁷6d¹ 97 97 Bk 7s²5f⁹ 98 98 Cf 7s²5f¹⁰ 99 99 Es 7s²5f¹¹ 100 100 Fm 7s²5f¹² 101 101 Md 7s²5f¹³ 102 102 No 7s²5f¹⁴ 103 103 Lr 7s²5f¹⁴

Lanthanides Actinides

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.

2A	7A
4 Be $2s^2$	9 F $2s^2 2p^5$
12 Mg $3s^2$	17 Cl $3s^2 3p^5$
20 Ca $4s^2$	35 Br $4s^2 4p^5$
38 Sr $5s^2$	53 I $5s^2 5p^5$
56 Ba $6s^2$	85 At $6s^2 6p^5$
88 Ra $7s^2$	
Alkaline earth metals	Halogens

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

Elements That Form Ions with Predictable Charges																	
	1A	2A									3A	4A	5A	6A	7A	8A	
1	Li ⁺												N ³⁻	O ²⁻	F ⁻		
2	Na ⁺	Mg ⁺¹	3B	4B	5B	6B	7B	8B		1B	2B	Al ³⁺			S ²⁻	Cl ⁻	
3	K ⁺	Ca ⁺¹													Se ²⁻	Br ⁻	
4	Rb ⁺	Sr ⁺¹													Te ²⁻	I ⁻	
5	Cs ⁺	Ba ⁺¹															

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.

- Al
- 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

Not for Distribution

Not for Distribution