Periodic Trends Involving the Sizes and Energy Levels of Atoms

2.3

In section 2.1, you learned that the size of a typical atom is about 10^{-10} m. You know, however, that the atoms of each element are distinctly different. For example, the atoms of different elements have different numbers of protons. This means, of course, that they also have different numbers of electrons. You might predict that the size of an atom is related to the number of protons and electrons it has. Is there evidence to support this prediction? If so, is there a pattern that can help you predict the relative size of an atom for any element in the periodic table?

In Investigation 2-A, you will look for a pattern involving the size of atoms. Chemists define, and measure, an atom's size in terms of its radius. The radius of an atom is the distance from its nucleus to the *approximate* outer boundary of the cloud-like region of its electrons. This boundary is approximate because atoms are not solid spheres. They do not have a fixed outer boundary.

Figure 2.12 represents how the radius of an atom extends from its nucleus to the approximate outer boundary of its electron cloud. Notice that the radius line in this diagram is just inside the outer boundary of the electron cloud. An electron may also spend time beyond the end of the radius line.

Section Preview/ Specific Expectations

In this section, you will

- **use** your understanding of electron arrangement and forces in atoms to explain the following periodic trends: atomic radius, ionization energy, electron affinity
- **analyze** data involving atomic radius, ionization energy, and electron affinity to identify and describe general periodic trends
- **communicate** your understanding of the following terms: *ion, anion, cation ionization energy, electron affinity*

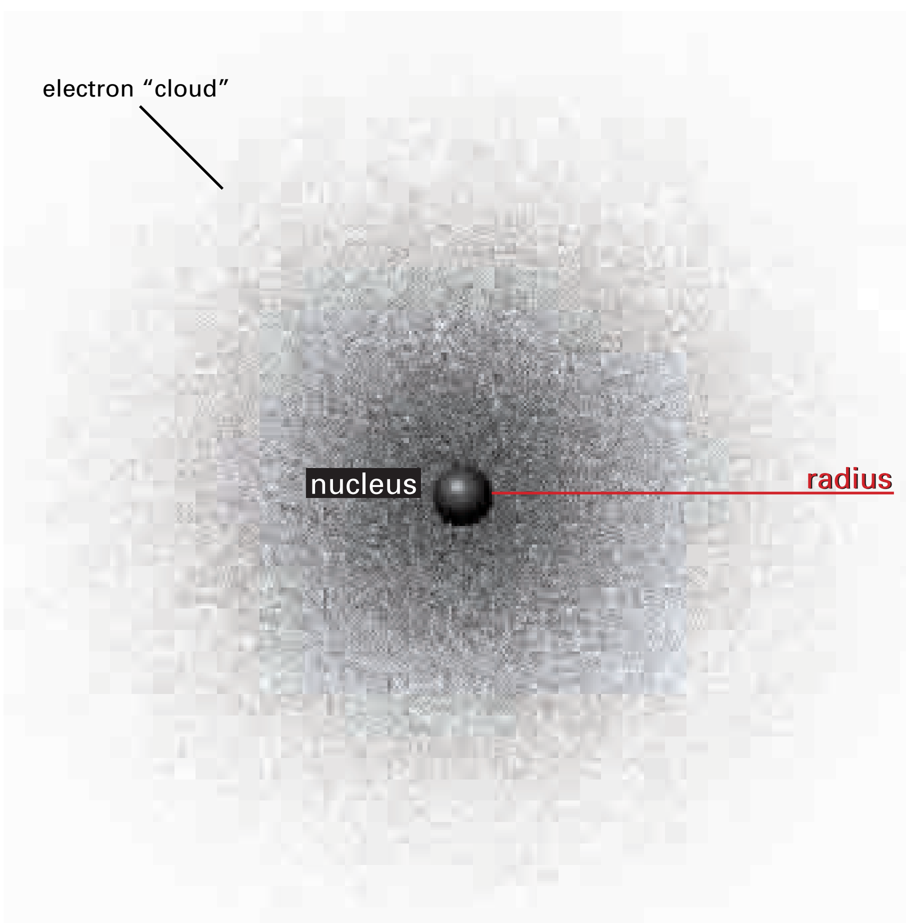


Figure 2.12 A representation of the radius of an atom

Trends for Atomic Size (Radius)

There are two general trends for atomic size:

- *As you go down each group in the periodic table, the size of an atom increases*. This makes sense if you consider energy levels. As you go down a group, the valence electrons occupy an energy level that is farther and farther from the nucleus. Thus, the valence electrons experience less attraction for the nucleus. In addition, electrons in the inner energy levels block, or *shield*, the valence electrons from the attraction of the nucleus. As a result, the total volume of the atom, and thus the size, increases with each additional energy level.
- *As you go across a period, the size of an atom decreases*. This trend might surprise you at first, since the number of electrons increases as you go across a period. You might think that more electrons would occupy more space, making the atom larger. You might also think that repulsion from their like charges would force the electrons farther apart. The size of an atom decreases, however, because the positive charge on the nucleus also increases across a period. As well, without additional energy the electrons are restricted to their outer energy level. For example, the outer energy level for Period 2 elements is the second energy level. Electrons cannot move beyond this energy level. As a result, the positive force exerted by the nucleus pulls the outer electrons closer, reducing the atom's total size.

Figure 2.13 summarizes the trends for atomic size. The Practice Problems that follow give you a chance to apply your understanding of these trends.

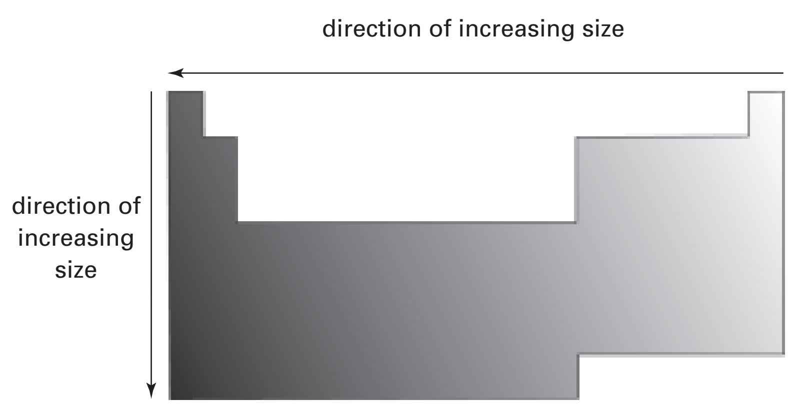


Figure 2.13 **Atomic size increases down a group and decreases across a period in the periodic table.**

Practice Problems

7. Using only their location in the periodic table, rank the atoms in each set by decreasing atomic size. Explain your answers.

(a) Mg, Be, Ba
(b) Ca, Se, Ga
(c) Br, Rb, Kr
(d) Se, Br, Ca
(e) Ba, Sr, Cs
(f) Se, Br, Cl
(g) Mg, Ca, Li
(h) Sr, Te, Se
(i) In, Br, I
(j) S, Se, O

Trends for Ionization Energy

A neutral atom contains equal numbers of positive charges (protons) and negative charges (electrons). The particle that results when a neutral atom gains electrons or gives up electrons is called an **ion**. Thus, an ion is a charged particle. *An atom that gains electrons becomes a negatively charged* **anion**. *An atom that gives up electrons becomes a positively charged* **cation**. Figure 2.14 shows the formation of ions for several elements. As you examine the diagrams, pay special attention to

- the energy level from which electrons are gained or given up
- the charge on the ion that is formed when an atom gains or gives up electrons
- the arrangement of the electrons that remain after electrons are gained or given up

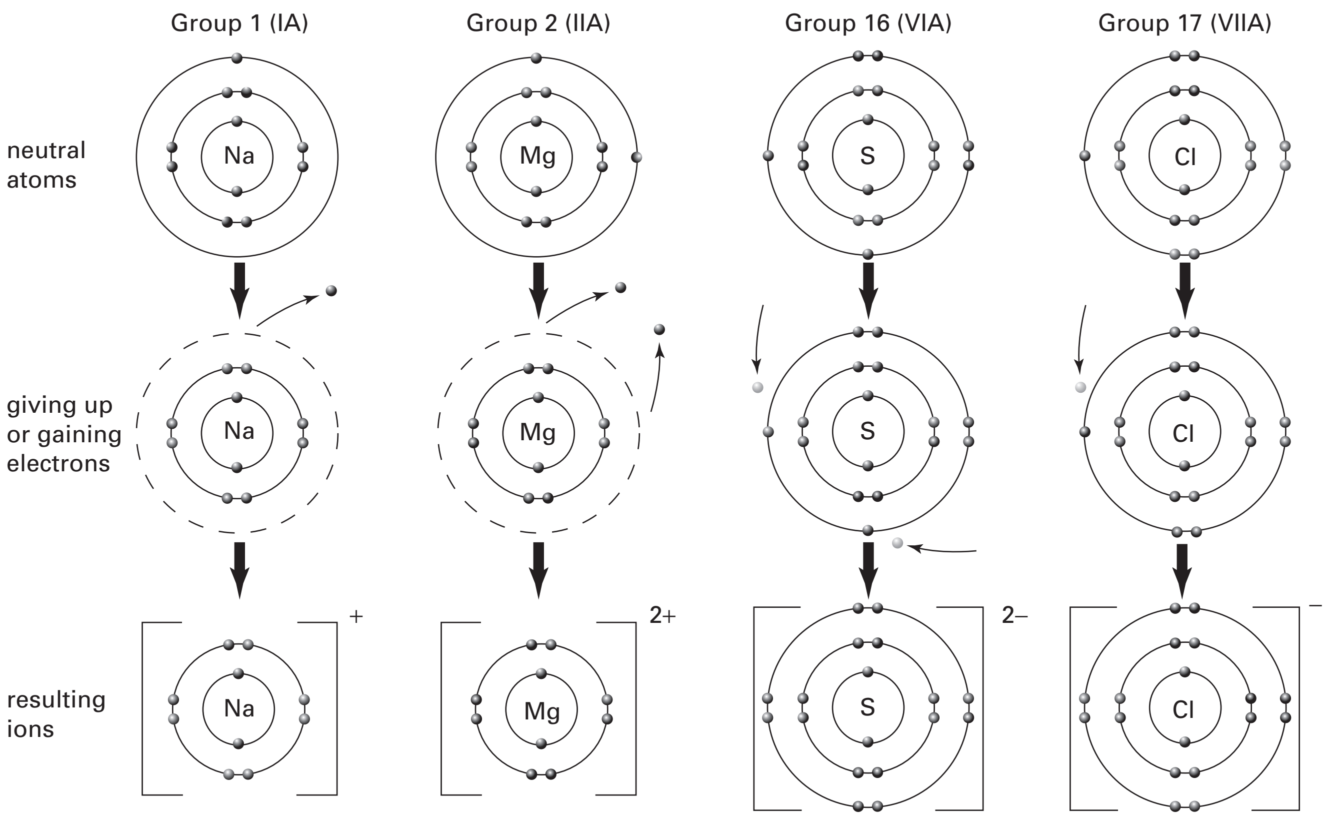


Figure 2.14 **These diagrams show the ions that are formed from neutral atoms of sodium, magnesium, sulfur, and chlorine. What other element has the same electron arrangement that sodium, magnesium, sulfide, and chloride ions have?**

Try to visualize the periodic table as a cylinder, rather than a flat plane. Can you see a relationship between ion formation and the electron arrangement of noble gases? Examine Figure 2.14 as well as 2.15 on the next page. The metals that are main-group elements tend to *give up* electrons and form ions that have the same number of electrons as the nearest noble gases. Non-metals tend to *gain* electrons and form ions that have the same number of electrons as the nearest noble gases. For example, when a sodium atom gives up its single valence electron, it becomes a positively charged sodium ion. Its outer electron arrangement is like neon's outer electron arrangement. When a fluorine atom gains an electron, it becomes a negatively charged ion with an outer electron arrangement like that of neon.

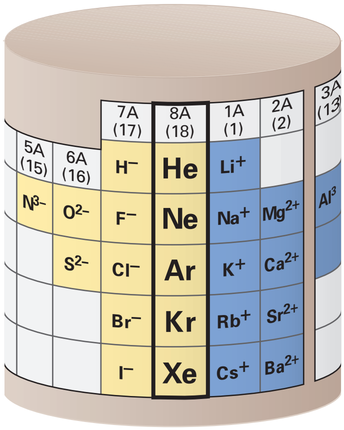


Figure 2.15 **Examine the relationship between ion charge and noble gas electron arrangement.**

Figure 2.15 can help you determine the charge on an ion. Count the number of groups an ion is from the nearest noble gas. That number is the charge on the ion. For example, aluminum is three groups away from neon. Thus, an aluminum ion has a charge of 3+. Sulfur is two groups away from argon. Thus, a sulfide ion has a charge of 2–. **Remember:** Metals form positive ions (cations) and non-metals form negative ions (anions).

It takes energy to overcome the attractive force of a nucleus and pull an electron away from a neutral atom. The energy that is required to remove an electron from an atom is called **ionization energy**. The bar graph in Figure 2.16 shows the ionization energy that is needed to remove one electron from the outer energy level of the atoms of the main-group elements. This energy is called the *first ionization energy.* It is measured in units of kJ/mol. A kilojoule (kJ) is a unit of energy. A mole (mol) is an amount of a substance. (You will learn about the mole in Unit 2.)

As you can see, atoms that give up electrons easily have low ionization energies. You would probably predict that the alkali metals of Group 1 (IA) would have low ionization energies. These elements are, in fact, extremely reactive because it takes so little energy to remove their single valence electron.

CHEM FACT

All elements, except hydrogen, have more than one electron that can be removed. Therefore, they have more than one ionization energy. The energy that is needed to remove a second electron is called the *second ionization energy.* The energy that is needed to remove a third electron is the *third ionization energy*, and so on. What trend would you expect to see in the values of the first, second, and third ionization energies for main-group elements? What is your reasoning?

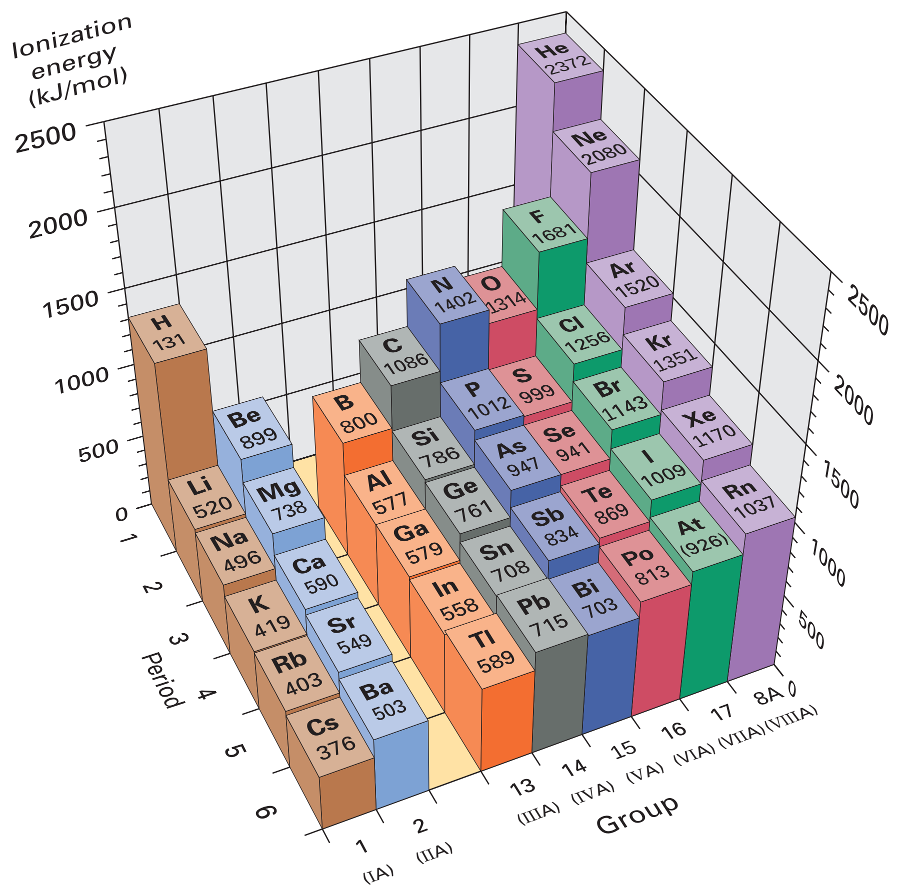


Figure 2.16 **This graph represents the first ionization energy for the main-group elements.**

Summarizing Trends for Ionization Energy

Although you can see a few exceptions in Figure 2.16, there are two general trends for ionization energy:

- *Ionization energy tends to decrease down a group.* This makes sense in terms of the energy level that the valence electrons occupy. Electrons in the outer energy level are farther from the positive force of the nucleus. Thus, they are easier to remove than electrons in lower energy levels.
- *Ionization energy tends to increase across a period.* As you go across a period, the attraction between the nucleus and the electrons in the outer energy level increases. Thus, more energy is needed to pull an electron away from its atom. For this trend to be true, you would expect a noble gas to have the highest ionization energy of all the elements in the same period. As you can see in Figure 2.16, they do.

Figure 2.17 summarizes these general trends for ionization energy. The Practice Problems below give you a chance to apply your understanding of these trends.

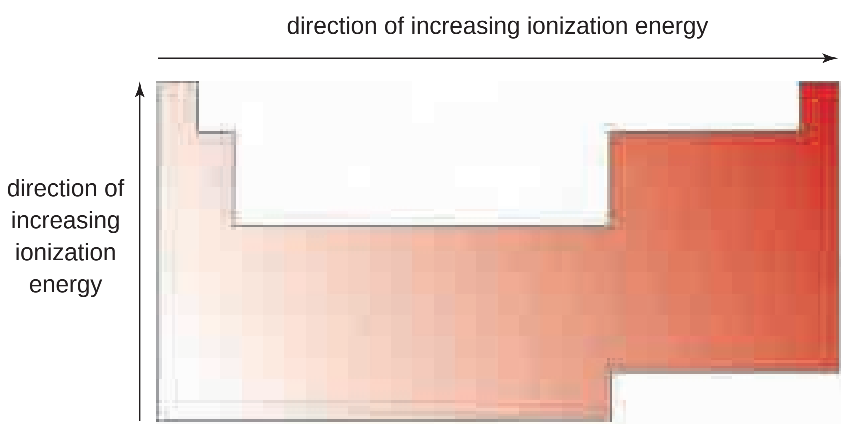


Figure 2.17 **Ionization energy tends to decrease down a group and increase across a period.**

Practice Problems

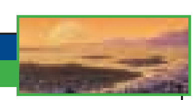
8. Using only a periodic table, rank the elements in each set by increasing ionization energy. Explain your answers.

(a) Xe, He, Ar
(b) Sn, In, Sb
(c) Sr, Ca, Ba
(d) Kr, Br, K
(e) K, Ca, Rb
(f) Kr, Br, Rb

9. Using only a periodic table, identify the atom in each of the following pairs with the *lower* first ionization energy.

(a) B, O
(b) B, In
(c) I, F
(d) F, N
(e) Ca, K
(f) B, Tl

COURSE CHALLENGE



Your understanding of periodic trends such as atomic radius and ionization energy will help you identify some unknown elements in the Chemistry Course Challenge at the end of this book.