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 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*

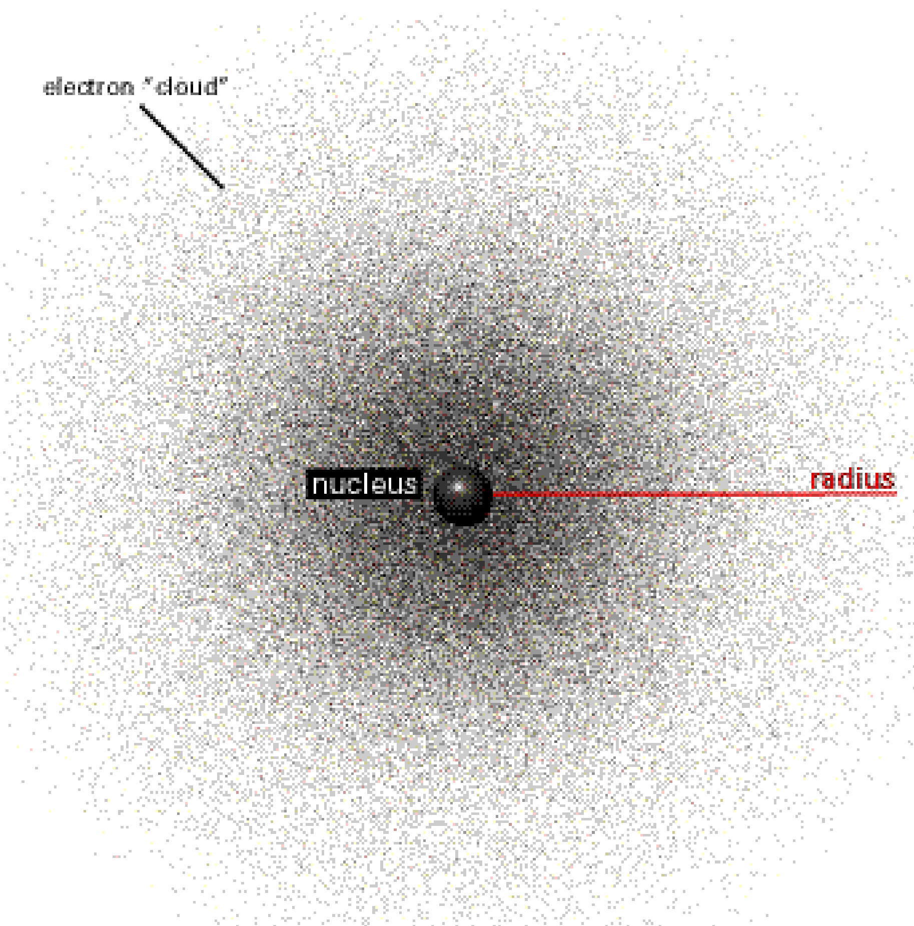


Figure 2.12 **A representation of the radius of an atom**

SKILL FOCUS

- Predicting
- Performing and recording
- Analyzing and interpreting

Analyzing Atomic Radius Data

Examine the main-group elements in the periodic table. Imagine how their size might change as you move down a group or across a period. What knowledge and reasoning can you use to infer the sizes of the atoms?

Question

How do the sizes of main-group atoms compare within a group and across a period?

Prediction

Predict a trend (pattern) that describes how the sizes of main-group atoms change down a group and across a period. Include a brief explanation to justify your prediction.

Safety Precautions

Be careful when handling any sharp instruments or materials that you choose to use.

Materials

to be decided in class

Procedure

1. The table below lists the atomic *radii* (plural of *radius*) for the main-group elements. Design different scale models that could help you visualize and compare the sizes of the atoms. Your models can be two-dimensional or three-dimensional, large or small.
2. Discuss your designs as a group. Choose a design that you think will best show the information you require.
3. Build your models. Arrange them according to their positions in the periodic table.

Atomic Radii of Main-Group Elements

Name of element	Atomic radius in picometres (pm)	Name of element	Atomic radius in picometres (pm)	Name of element	Atomic radius in picometres (pm)
aluminum	143	gallium	141	polonium	167
antimony	159	germanium	137	potassium	235
argon	88	helium	49	radon	134
astatine	145	hydrogen*	79	rubidium	248
barium	222	indium	166	selenium	140
beryllium	112	iodine	132	silicon	132
bismuth	170	krypton	103	sodium	190
boron	98	lead	175	strontium	215
bromine	112	lithium	155	sulfur	127
calcium	197	magnesium	160	tellurium	142
carbon	91	neon	51	thallium	171
cesium	267	nitrogen	92	tin	162
chlorine	97	oxygen	65	xenon	124
fluorine	57	phosphorus	128		

*Quantum mechanical value for a free hydrogen atom

Analysis

1. How do atomic radii change as you look from top to bottom within a group?
2. How do atomic radii change as you look from left to right across a period?
3. Compare your observations with your prediction. Explain why your results did, or did not, agree with your prediction.

Conclusion

4. State whether or not atomic radius is a periodic property of atoms. Give evidence to support your answer.

Application

5. Would you expect atoms of the transition elements to follow the same trend you observed for the main-group elements? Locate atomic radius data for the transition elements (not including the inner transition elements). Make additional models, or draw line or bar graphs, to verify your expectations.

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.

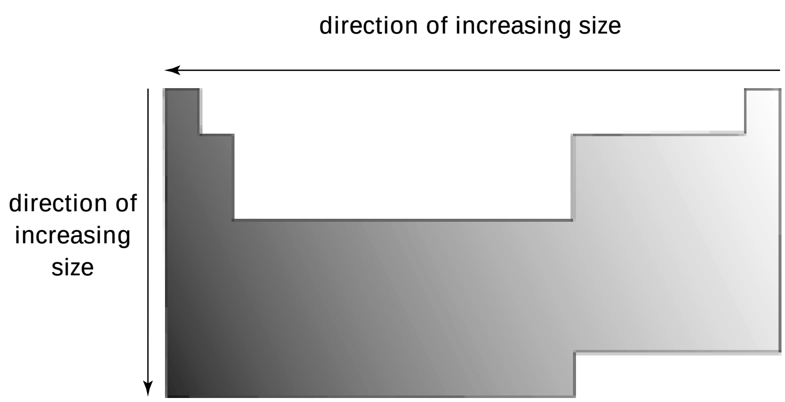


Figure 2.13 Atomic size increases down a group and decreases across a row in the periodic table.

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 ion that gains electrons becomes a negatively charged* **anion**. *An ion 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

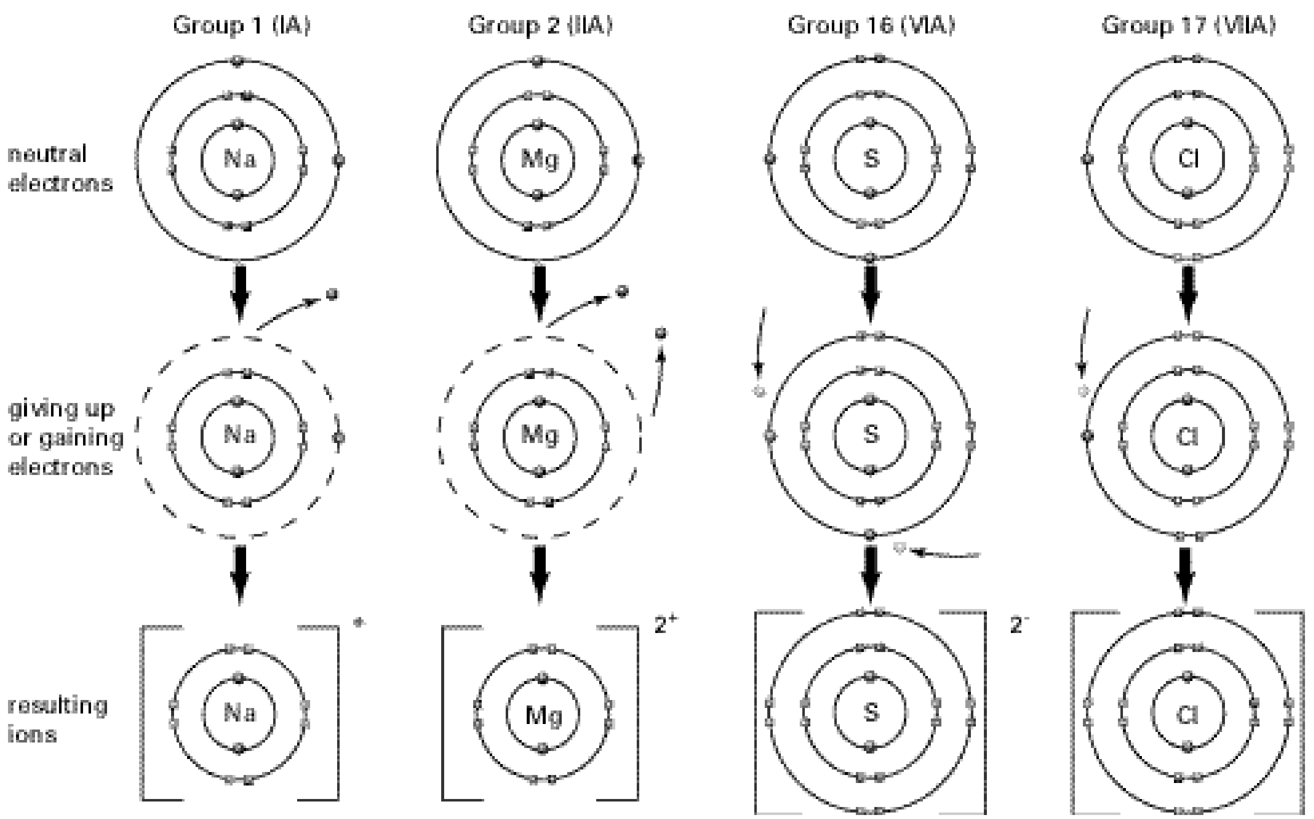


Figure 2.14 **These diagrams show the ions that are formed from neutral atoms of sodium, magnesium, sulfur, and chlorine. What other element has the same electron arrangement that sodium, magnesium, sulfur, and chlorine ions have?**

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? Examin 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.

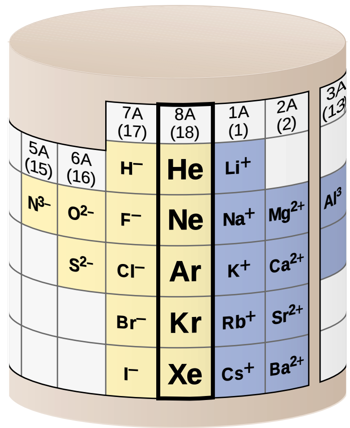


Figure 2.15 **Examine the relationship between ion charge and noble gas electron arrangement.**

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 sulfur 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 (1A) 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?

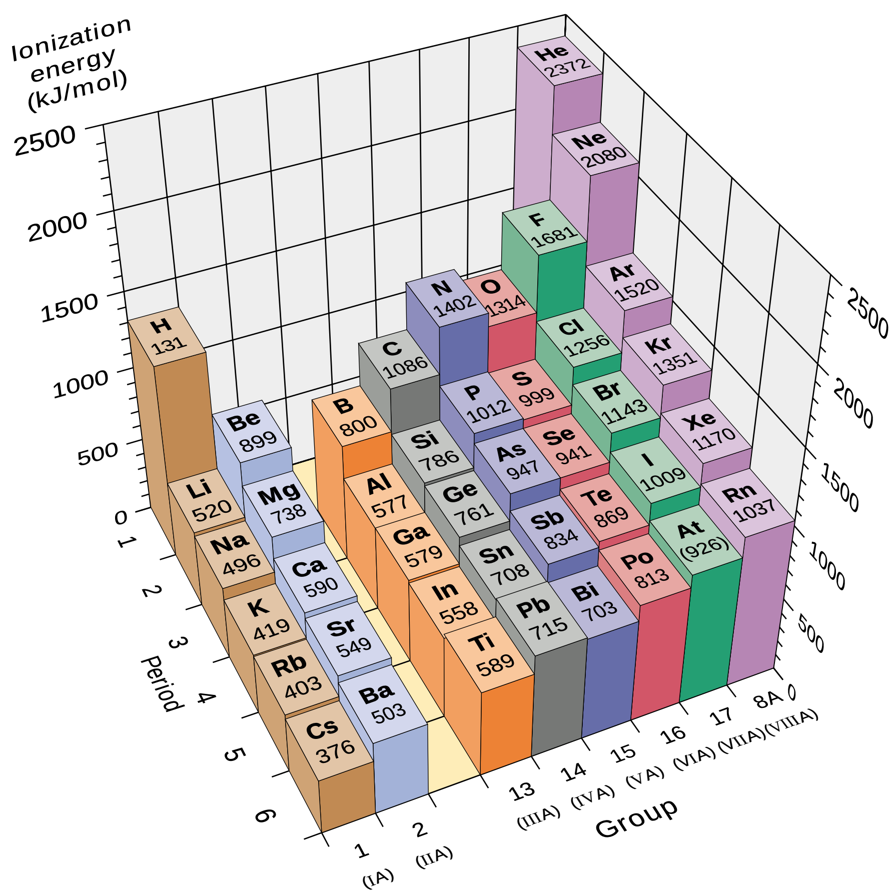


Figure 2.16 **This graph represents the first ionization energy for the main-group elements.**

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 the noble gases to have the highest ionization energies of all the elements. 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.

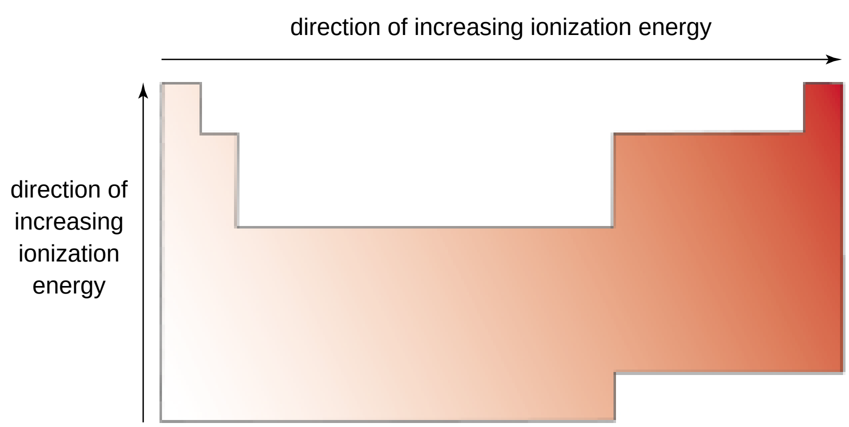


Figure 2.17 **Ionization energy tends to decrease down a group and increase across a period.**

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

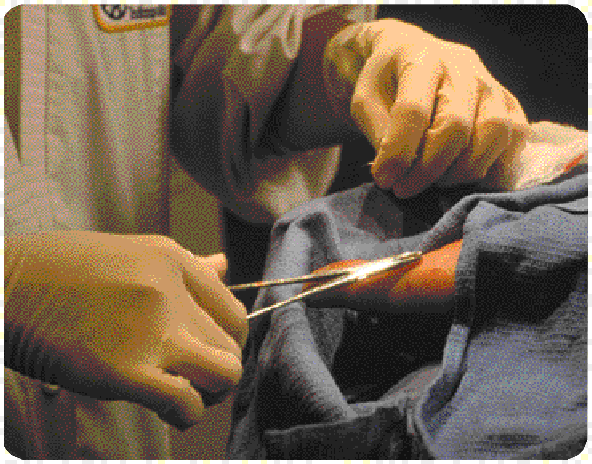
COURSE CHALLENGE

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.

Chemistry Bulletin

Science | Technology | Society | Environment

Manitoba Mine Specializes in Rare Metals



At TANCO in Bernic Lake, Manitoba, miners are busy finding and processing two rare and very different metallic elements: tantalum and cesium. Both of these metals are important parts of "high-tech" applications around the world. They are used in nuclear reactors and as parts of aircraft, missiles, camera lenses, and surgical instruments like the one shown above.

Tantalum is found only in Canada, Australia, Brazil, Zaire, and China. TANCO (the Tantalum Mining Corporation of Canada) is the only mine in North America that produces tantalum. TANCO is also the world's main producer of cesium. Other than the fact that tantalum and cesium are both found at TANCO and both are used in high-tech applications, they share little in common.

Tantalum is a heavy, hard, and brittle grey metal. In its pure form, it is extremely ductile and can be made into a fine wire. This has proved useful for making surgical sutures. Another property that makes tantalum useful is its resistance to corrosion by most acids, due to its very limited reactivity. At normal temperatures, tantalum is virtually non-reactive. In fact, tantalum has about the same resistance to corrosion as glass. Tantalum can withstand higher temperatures than glass, however. It has a melting point of 3290 K—higher than the melting points of all other elements, except tungsten and rhenium. Tantalum's resistance to corrosion and high melting point make it suitable for use in surgical equipment and implants. For example, some of the pins that are used by surgeons to hold a patient's broken bones together are made of tantalum.

Tantalum is resistant to corrosion because a thin film of tantalum oxide forms when tantalum is exposed to oxygen. The metal oxide acts as a protective layer. The oxide also has special refractive properties that make it ideal for use in camera lenses.

Cesium is quite different from tantalum, but it, too, has many high-tech applications. Cesium is a silvery-white metal. It is found in a mineral called pollucite. Cesium is the softest of all the metals and is a liquid at room temperature. It is also the most reactive metal on Earth.

Cesium has a high ionization energy. It readily gives up its single valence electron to form crystalline compounds with all the halogen non-metals. Cesium is also very photoelectric. This means that it easily gives up its lone outer electron when it is exposed to light. Thus, cesium is used in television cameras and traffic signals. As well, it has the potential to be used in ion propulsion engines for travel into deep space.

Making Connections

1. Make a table to show the differences and similarities between tantalum and cesium. For each metal, add a column to describe how its different properties make it useful for specific applications.

2. Bernic Lake is one of the few locations where tantulum can be found. As well, it is the most important cesium source in the world. Research and describe what geographical conditions led to the presence of two such rare metals in one location.

Trends for Electron Affinity

In everyday conversation, if you like something, you may say that you have an affinity for it. For example, what if you enjoy pizza and detest asparagus? You may say that you have a high affinity for pizza and a low affinity for asparagus. If you prefer asparagus to pizza, your affinities are reversed.

Atoms are not living things, so they do not like or dislike anything. You know, however, that some atoms have a low attraction for electrons. Other atoms have a greater attraction for electrons. **Electron affinity** is a measure of the change in energy that occurs when an electron is added to the outer energy level of an atom to form a negative ion.

Figure 2.18 identifies the electron affinities of the main-group elements. If the ion that is formed by gaining an electron is stable, the electron affinity is expressed as a negative integer. The more unstable the ion, the higher is the negative integer for the electron affinity. Notice, for example, that fluorine has the highest electron affinity. This indicates that fluorine is very likely to be involved in chemical reactions. In fact, fluorine is the most reactive of all the elements.

Metals have very low electron affinities. This is especially true for the Group 1 (1A) and 2 (2A) elements. Atoms of these elements form stable positive ions. A negative ion that is formed by the elements of these groups is unstable. It breaks apart into a neutral atom and a free electron.

Examine Figure 2.18. What trends can you observe? How regular are these trends?

1 (IA)	2 (IIA)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
H −72.8							He (+21)
Li −59.6	Be (+241)	B −26.7	C −122	N 0	O −141	F −328	Ne (+29)
Na −52.9	Mg (+230)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+34)
K −48.4	Ca (+156)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+167)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+40)
Cs −45.5	Ba (+52)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

Figure 2.18 **The units for electron affinity are the same as the units for ionization energy: kJ/mol. High negative numbers mean a high electron affinity. Low negative numbers and any positive numbers mean a low electron affinity.**

Tools & Techniques

Analyzing the Ice Man's Axe

In September 1991, hikers in the Alps Mountains near the Austrian-Italian border discovered the body of a man who had been trapped in a glacier. He was almost perfectly preserved. With him was an assortment of tools, including an axe with a metal blade.

Scientists were particularly interested in the axe. At first, they believed that it was bronze, which is an alloy of copper and tin. There was a complication, however. Dating techniques that were used for the clothing and body suggested that the "Ice Man" was about 5300 years old. Bronze implements do not appear in Europe's fossil record until about 4000 years ago. Either Europeans were using bronze earlier than originally thought, or the axe was made of a different material. Copper was consistent with the Ice Man's age, since it has been used for at least the past 6000 years.

One technique to determine a metal's identity is to dissolve it in acid. The resulting solution is examined for evidence of ions. Scientists did not want to damage the precious artifact in any way, though.

The solution was an analytical technique called *X-ray fluorescence*. The object is irradiated with high-energy X-ray radiation. Its atoms absorb the radiation, causing electrons from a lower energy level to be ejected from the atom. This causes electrons from an outer energy level to "move in," to occupy the vacated space. As the electrons fall to a less energetic state, they emit X-rays. The electrons of each atom emit X-rays of a particular wavelength. Scientists use this energy "signature" to identify the atom.

Analysis by X-ray fluorescence revealed that the metal in the blade of the axe was almost pure copper.

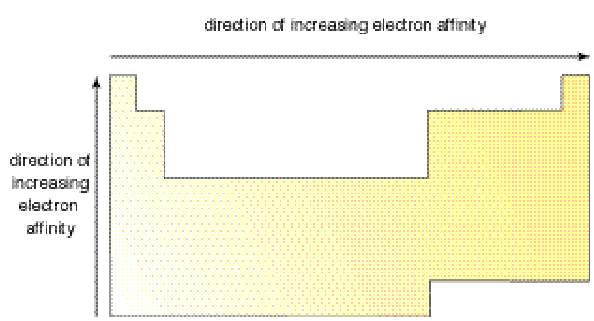


Figure 2.19 **Electron affinity tends to decrease down a group and increase across a period.**

The trends for electron affinity, shown in Figure 2.19, are more irregular than the trends for ionization energy and atomic radius. Nevertheless, the following *general* trends can be observed:

- *Electron affinity tends to decrease down a group*. For example, fluorine has a higher electron affinity than iodine.
- *Electron affinity tends to increase across a period*. For example, calcium has a lower electron affinity than sulfur.

ThoughtLab Design an Annotated Periodic Table

You have learned a great deal about the properties of the elements. In the following chapters, you will learn more. With your classmates, develop your own large-scale periodic table to record the properties and common uses of the elements.

Procedure

1. Use print and electronic resources (including this textbook) to find information about one element. Consult with your classmates to make sure that everyone chooses a different element.
2. Find the following information about your element:
   - atomic number
   - atomic mass
   - atomic symbol
   - melting point
   - boiling point
   - density
   - atomic radius
   - ionization energy
   - electron affinity
   - place and date discovered, and the name of the scientist who discovered it
   - uses, both common and unusual
   - hazards and methods for safe handling

   If possible, find a photograph of the element in its natural form. If this is not possible, find a photograph that shows one or more compounds in which the element is commonly found.
3. Record your findings on a sheet of notepaper or blank paper. Arrange all the sheets of paper, for all the elements, in the form of a periodic table on a wall in the classroom. Make sure that you leave space to insert additional properties and uses of your element as you learn about them during this course.

Analysis

1. What uses of your element did you know about? Which uses surprised you? Why?
2. Examine the dates on which the elements were discovered. What pattern do you notice? How can you explain this pattern?
3. Do you think that scientists have discovered all the naturally occurring elements? Do you think they have discovered all the synthetically produced elements? Give reasons to justify your opinions.

Section Wrap-up

Despite some irregularities and exceptions, the following periodic trends summarize the relationships among atomic size, ionization energy, and electron affinity:

- Trends for atomic size are the reverse of trends for ionization energy and electron affinity. Larger atoms tend to have lower ionization energies and lower electron affinities.
- Group 16 (VI A) and 17 (VII A) elements attract electrons strongly. They do not give up electrons readily. In other words, they have a strong tendency to form negative ions. Thus, they have high ionization energies and high electron affinities.
- Group 1 (I A) and 2 (II A) elements give up electrons readily. They have low or no attraction for electrons. In other words, they have a strong tendency to form positive ions. Thus, they have low ionization energies and low electron affinities.
- Group 18 (VIII A) elements do not attract electrons and do not give up electrons. In other words, they do not naturally form ions. (They are very stable.) Thus, they have very high ionization energies and very low electron affinities.

The trends you have examined in this chapter have an enormous influence on the ability of atoms to combine and form compounds. In the next chapter, you will use these trends to help you understand and predict the kinds of compounds that atoms form. As well, you will learn about another periodic trend. This trend called electronegativity, is related to the formation of some of the most common compounds in your life, such as water, carbon dioxide, and sugar.

Section Review

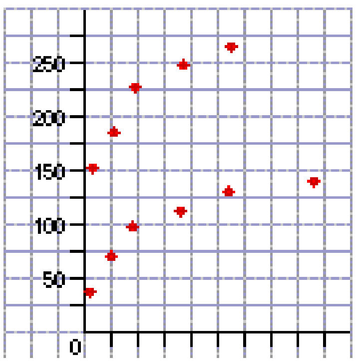
1 K/U How does your understanding of electron arrangement and forces in atoms help you explain the following periodic trends?
(a) atomic radius
(b) ionization energy
(c) electron affinity

2 K/U Using only their location in a periodic table, rank each of the following sets of elements in order of increasing atomic size. Explain your answer in each case.
(a) Mg, S, Cl
(b) Al, B, In
(c) Ne, Ar, Xe
(d) Rb, Xe, Te
(e) P, Na, F
(f) O, S, N

3 K/U Using only their location in a periodic table, rank each of the following sets of elements in order of decreasing ionization energy. Explain your answer in each case.
(a) Cl, Br, I
(b) Ga, Ge, Se
(c) K, Ca, Kr
(d) Na, Li, Cs
(e) S, Cl, Br
(f) Cl, Ar, K

4 K/U Which element in each of the following pairs will have the lower electron affinity? Explain your answer in each case.
(a) K or Ca
(b) O or Li
(c) S or Se
(d) Cs or F

5 C The graph shows a periodic trend, but is only partially complete. Copy it into your notebook and fill in all the data and labels that will make it complete. Title the graph with the trend it shows.



What data does this graph need to be complete?

6 I Use your understanding of periodic trends to sketch the shape of a graph that shows a trend that is opposite to that shown in question 5. Label the x- and y-axes, and add any other labels that you think are necessary to represent the trend you are showing.

Unit Project Prep

Look ahead to the project at the end of this unit. Many common chemical products contain elements (as components of compounds) from some groups of the periodic table more than others. Examine ingredient labels from different chemical products. Which element groups are represented most frequently? Why might that be?