6.3 Periodic Trends

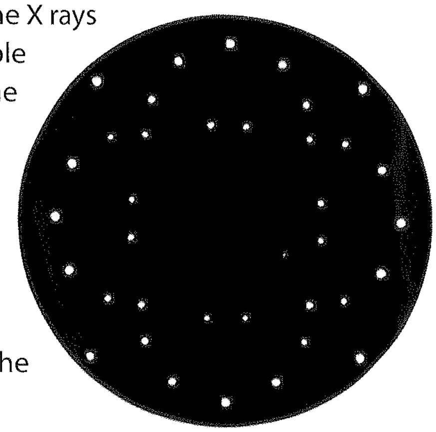
Guide for Reading

Key Concepts
- What are the trends among the elements for atomic size?
- How do ions form?
- What are the trends among the elements for first ionization energy, ionic size, and electronegativity?
- What is the underlying cause of periodic trends?

Vocabulary
atomic radius
ion
cation
anion
ionization energy
electronegativity

Reading Strategy
Building Vocabulary After you read this section, explain the difference between a cation and an anion.

Connecting to Your World An atom doesn't have a sharply defined boundary. So the radius of an atom cannot be measured directly. There are ways to estimate the sizes of atoms. In one method, a solid is bombarded with X rays, and the paths of the X rays are recorded on film. Sodium chloride (table salt) produced the geometric pattern in the photograph. Such a pattern can be used to calculate the position of nuclei in a solid. The distances between nuclei in a solid are an indication of the size of the particles in the solid. In this section, you will learn how properties such as atomic size are related to the location of elements in the periodic table.

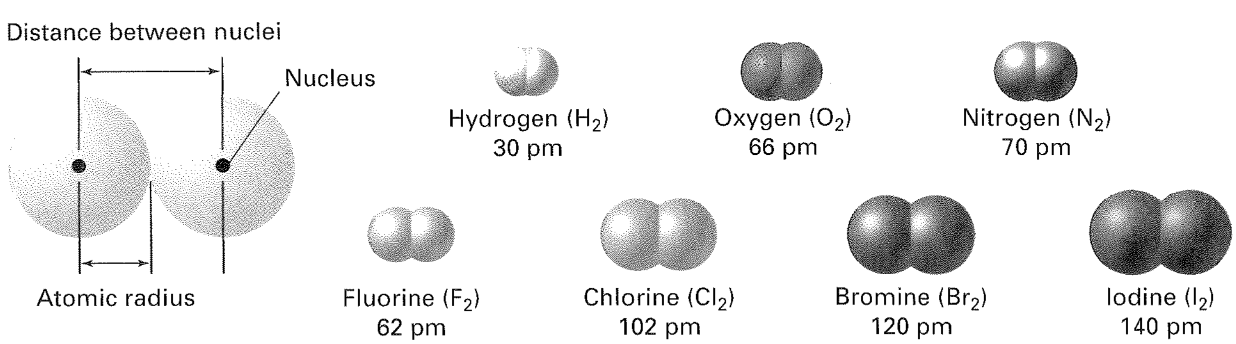


Trends in Atomic Size

Another way to think about atomic size is to look at the units that form when atoms of the same element are joined to one another. These units are called molecules. Figure 6.13 shows models of molecules (molecular models) for seven nonmetals. Because the atoms in each molecule are identical, the distance between the nuclei of these atoms can be used to estimate the size of the atoms. This size is expressed as an atomic radius. The **atomic radius** is one half of the distance between the nuclei of two atoms of the same element when the atoms are joined.

The distances between atoms in a molecule are extremely small. So the atomic radius is often measured in picometers. Recall that there are one trillion, or 10^{12}, picometers in a meter. The molecular model of iodine in Figure 6.13 is the largest. The distance between the nuclei in an iodine molecule is 280 pm. Because the atomic radius is one half the distance between the nuclei, a value of 140 pm (280/2) is assigned as the radius of the iodine atom. **In general, atomic size increases from top to bottom within a group and decreases from left to right across a period.**

Figure 6.13 This diagram lists the atomic radii of seven nonmetals. An atomic radius is half the distance between the nuclei of two atoms of the same element when the atoms are joined.



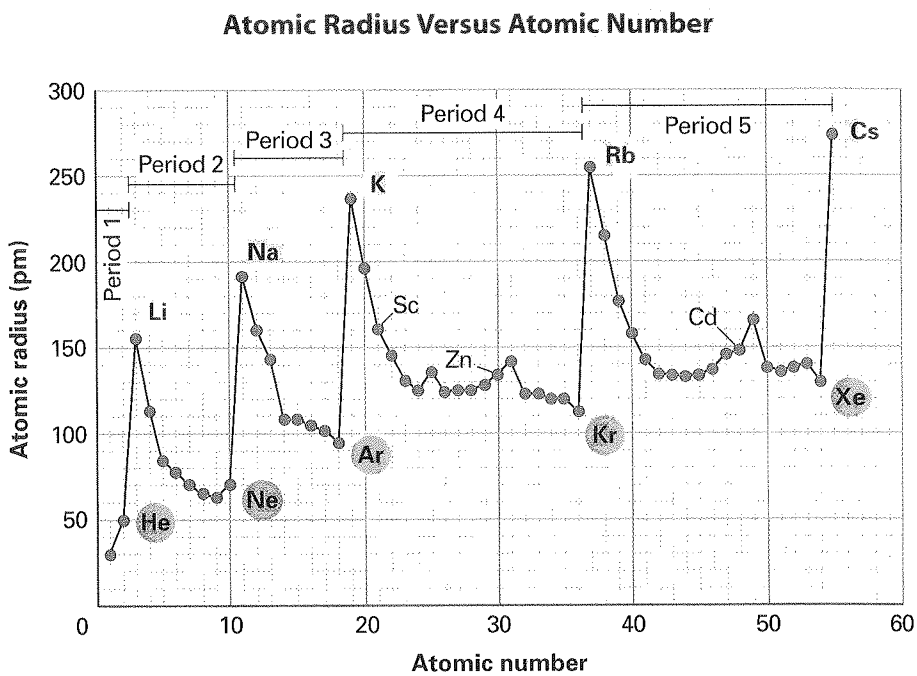


Figure 6.14 This graph plots atomic radius versus atomic number for 55 elements.

INTERPRETING GRAPHS

a. Analyzing Data Which alkali metal has an atomic radius of 238 pm?

b. Drawing Conclusions Based on the data for alkali metals and noble gases, how does atomic size change within a group?

c. Predicting Is an atom of barium, atomic number 56, smaller or larger than an atom of cesium (Cs)?

Group Trends in Atomic Size In the Figure 6.14 graph, atomic radius is plotted versus atomic number. Look at the data for the alkali metals and noble gases. The atomic radius within these groups increases as the atomic number increases. This increase is an example of a trend.

As the atomic number increases within a group, the charge on the nucleus increases and the number of occupied energy levels increases. These variables affect atomic size in opposite ways. The increase in positive charge draws electrons closer to the nucleus. The increase in the number of occupied orbitals shields electrons in the highest occupied energy level from the attraction of protons in the nucleus. The shielding effect is greater than the effect of the increase in nuclear charge. So the atomic size increases.

Periodic Trends in Atomic Size Look again at Figure 6.14. In general, atomic size decreases across a period from left to right. Each element has one more proton and one more electron than the preceding element. Across a period, the electrons are added to the same principal energy level. The shielding effect is constant for all the elements in a period. The increasing nuclear charge pulls the electrons in the highest occupied energy level closer to the nucleus and the atomic size decreases. Figure 6.15 summarizes the group and period trends in atomic size.

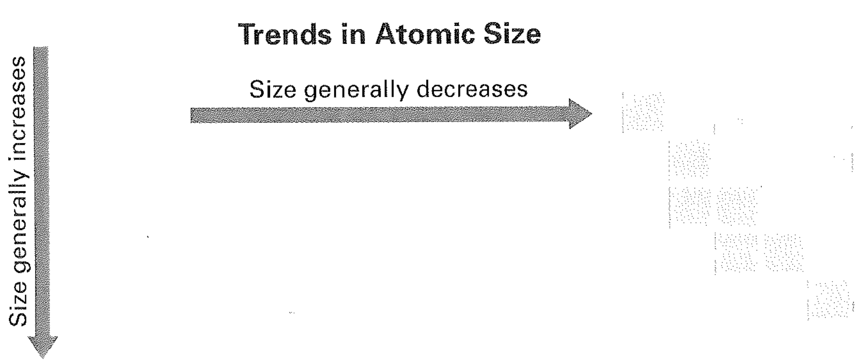


Figure 6.15 The size of atoms tends to decrease from left to right across a period and increase from top to bottom within a group. **Predicting** ***If a halogen and an alkali metal are in the same period, which one will have the larger radius?***

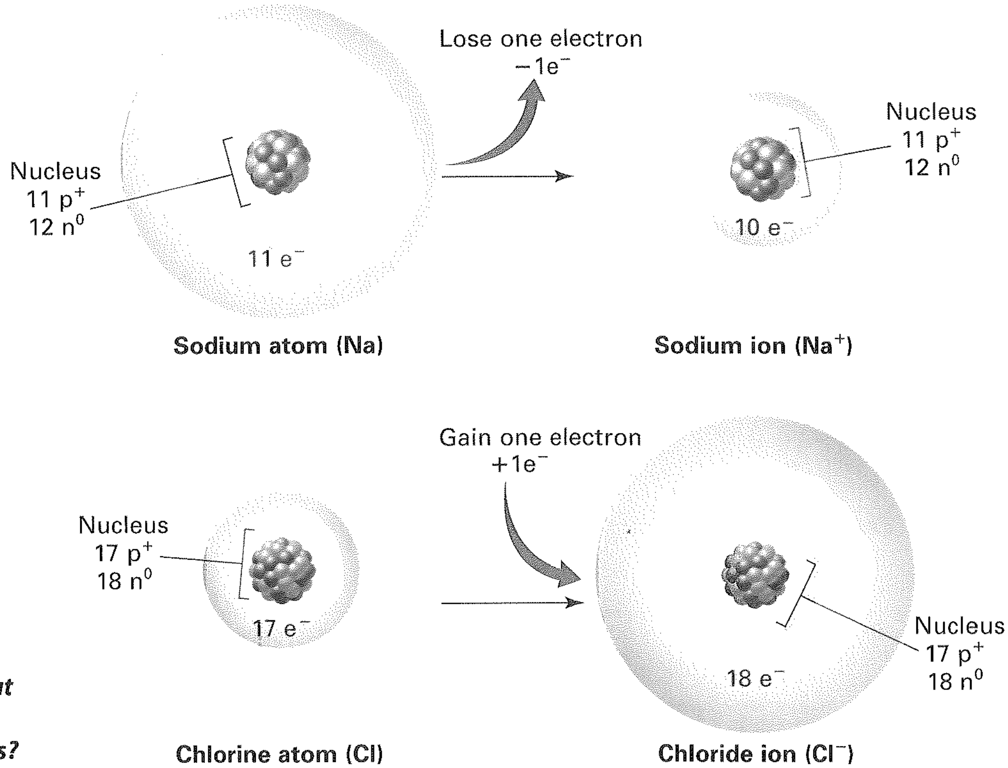


Figure 6.16 When a sodium atom loses an electron, it becomes a positively charged ion. When a chlorine atom gains an electron, it becomes a negatively charged ion. **Interpreting Diagrams** ***What happens to the protons and neutrons during these changes?***

Ions

Some compounds are composed of particles called ions. An **ion** is an atom or group of atoms that has a positive or negative charge. An atom is electrically neutral because it has equal numbers of protons and electrons. For example, an atom of sodium (Na) has 11 positively charged protons and 11 negatively charged electrons. The net charge on a sodium atom is zero $[(11+) + (-11) = 0]$.

Positive and negative ions form when electrons are transferred between atoms. Atoms of metals, such as sodium, tend to form ions by losing one or more electrons from their highest occupied energy levels. A sodium atom tends to lose one electron. Figure 6.16 compares the atomic structure of a sodium atom and a sodium ion. In the sodium ion, the number of electrons (10) is no longer equal to the number of protons (11). Because there are more positively charged protons than negatively charged electrons, the sodium ion has a net positive charge. An ion with a positive charge is called a **cation.** The charge for a cation is written as a number followed by a plus sign. If the charge is 1+, the number 1 is usually omitted from the symbol for the ion. So Na^+ is equivalent to Na^{1+}.

Atoms of nonmetals, such as chlorine, tend to form ions by gaining one or more electrons. A chlorine atom tends to gain one electron. Figure 6.16 compares the atomic structure of a chlorine atom and a chloride ion. In a chloride ion, the number of electrons (18) is no longer equal to the number of protons (17). Because there are more negatively charged electrons than positively charged protons, the chloride ion has a net negative charge. An ion with a negative charge is called an **anion.** The charge for an anion is written as a number followed by a minus sign.

Interactive Textbook

Animation 7 Discover the ways that atoms of elements combine to form compounds.

with ChemASAP

Checkpoint ***What is the difference between a cation and an anion?***

Trends in Ionization Energy

Recall that electrons can move to higher energy levels when atoms absorb energy. Sometimes there is enough energy to overcome the attraction of the protons in the nucleus. The energy required to remove an electron from an atom is called **ionization energy.** This energy is measured when an element is in its gaseous state. The energy required to remove the first electron from an atom is called the first ionization energy. The cation produced has a 1+ charge. **First ionization energy tends to decrease from top to bottom within a group and increase from left to right across a period.**

Table 6.1 lists the first, second, and third ionization energies for the first 20 elements. The second ionization energy is the energy required to remove an electron from an ion with a 1+ charge. The ion produced has a 2+ charge. The third ionization energy is the energy required to remove an electron from an ion with a 2+ charge. The ion produced has a 3+ charge.

Ionization energy can help you predict what ions elements will form. Look at the data in Table 6.1 for lithium (Li), sodium (Na), and potassium (K). The increase in energy between the first and second ionization energies is large. It is relatively easy to remove one electron from a Group 1A metal atom, but it is difficult to remove a second electron. So Group 1A metals tend to form ions with a 1+ charge.

Table 6.1

Ionization Energies of First 20 Elements (kJ/mol*)

Symbol	First	Second	Third
H	1312		
He (noble gas)	2372	5247	
Li	520	7297	11,810
Be	899	1757	14,840
B	801	2430	3659
C	1086	2352	4619
N	1402	2857	4577
O	1314	3391	5301
F	1681	3375	6045
Ne (noble gas)	2080	3963	6276
Na	496	4565	6912
Mg	738	1450	7732
Al	578	1816	2744
Si	786	1577	3229
P	1012	1896	2910
S	999	2260	3380
Cl	1256	2297	3850
Ar (noble gas)	1520	2665	3947
K	419	3069	4600
Ca	590	1146	4941

*An amount of matter equal to the atomic mass in grams.

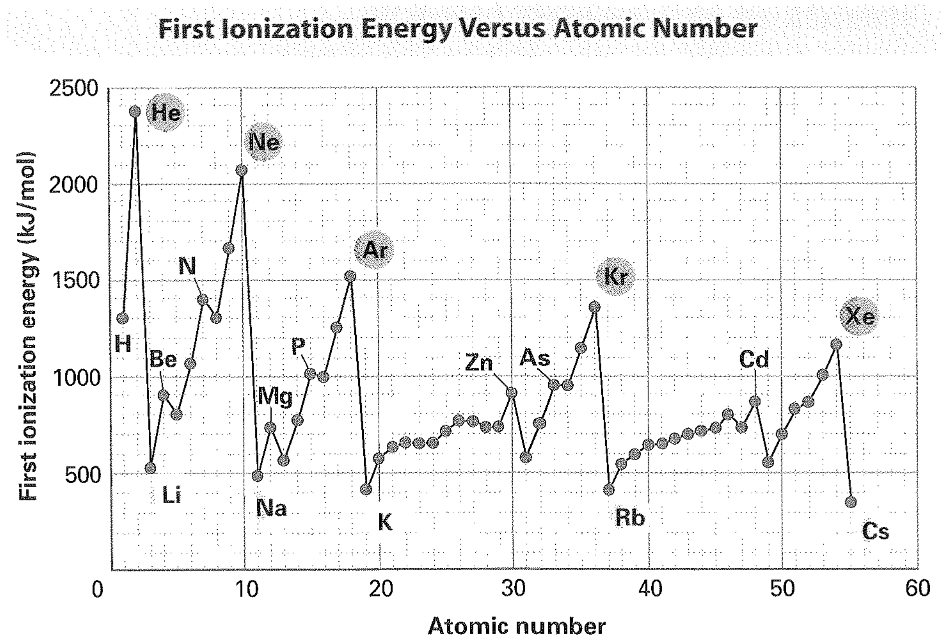
Figure 6.17 This graph reveals group and period trends for ionization energy.

INTERPRETING GRAPHS

a. Analyzing Data Which element in period 2 has the lowest first ionization energy? In period 3?

b. Drawing Conclusions What is the group trend for first ionization energy for noble gases and alkali metals?

c. Predicting If you drew a graph for second ionization energy, which element would you have to omit? Explain.



Group Trends in Ionization Energy Figure 6.17 is a graph of first ionization energy versus atomic number. Each red dot represents the data for one element. Look at the data for the noble gases and the alkali metals. In general, first ionization energy decreases from top to bottom within a group. Recall that the atomic size increases as the atomic number increases within a group. As the size of the atom increases, nuclear charge has a smaller effect on the electrons in the highest occupied energy level. So less energy is required to remove an electron from this energy level and the first ionization energy is lower.

Periodic Trends in Ionization Energy In general, the first ionization energy of representative elements tends to increase from left to right across a period. This trend can be explained by the nuclear charge, which increases, and the shielding effect, which remains constant. The nuclear charge increases across the period, but the shielding effect remains constant. So there is an increase in the attraction of the nucleus for an electron. Thus, it takes more energy to remove an electron from an atom. Figure 6.18 summarizes the group and period trends for first ionization energy.

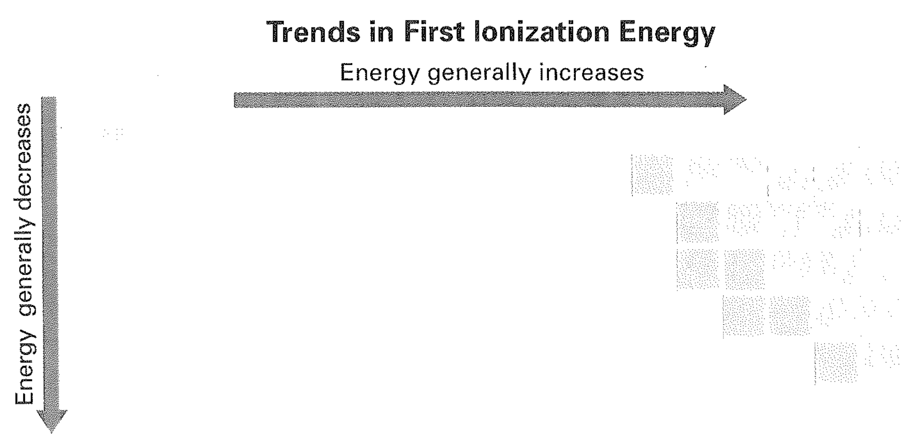


Figure 6.18 First ionization energy tends to increase from left to right across a period and decrease from top to bottom within a group.
Predicting ***Which element would have the larger first ionization energy—an alkali metal in period 2 or an alkali metal in period 4?***

Quick LAB

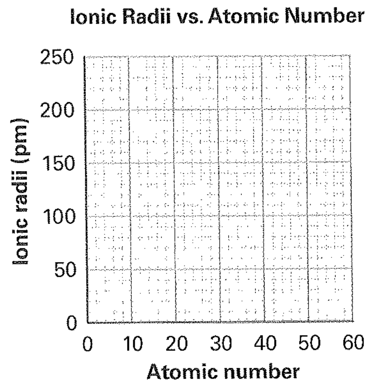
Periodic Trends in Ionic Radii

Purpose
Make a graph of ionic radius versus atomic number and use the graph to identify periodic and group trends.

Materials
- graph paper

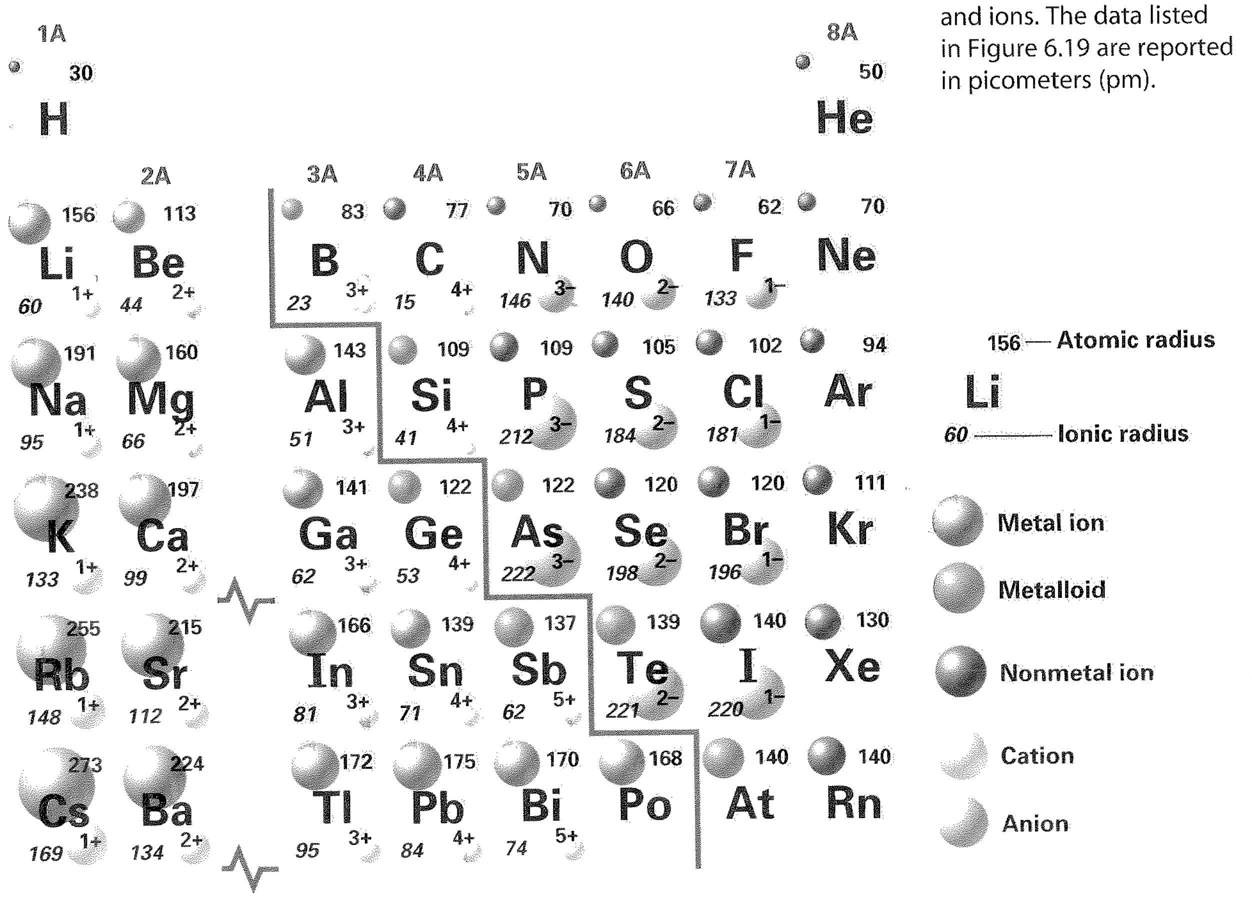
Procedure
Use the data presented in Figure 6.19 to plot ionic radius versus atomic number.

Analyze and Conclude
1. Describe how the size changes when an atom forms a cation and when an atom forms an anion.
2. How do the ionic radii vary within a group of metals? How do they vary within a group of nonmetals?
3. Describe the shape of a portion of the graph that corresponds to one period.



4. Is the trend across a period similar or different for periods 2, 3, 4, and 5?
5. Propose explanations for the trends you have described for ionic radii within groups and across periods.

Figure 6.19 Atomic and ionic radii are an indication of the relative size of atoms and ions. The data listed in Figure 6.19 are reported in picometers (pm).



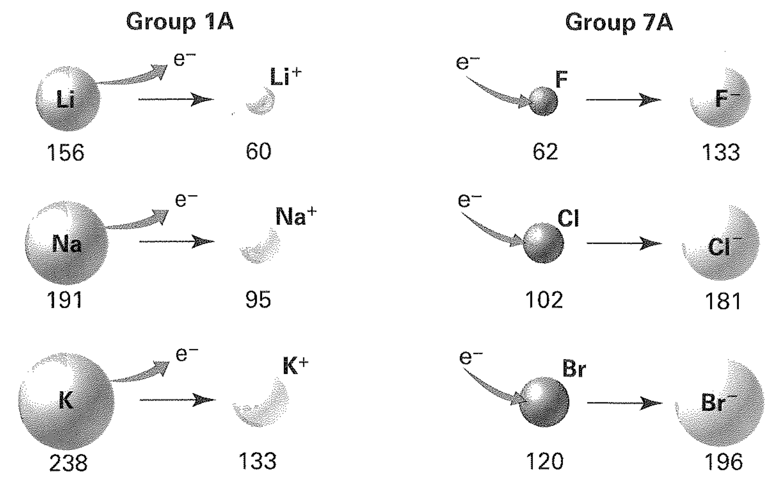


Figure 6.20 This diagram compares the relative sizes of atoms and ions for selected alkali metals and halogens. The data are given in picometers. **Comparing and Contrasting** ***What happens to the radius when an atom forms a cation? When an atom forms an anion?***

Trends in Ionic Size

During reactions between metals and nonmetals, metal atoms tend to lose electrons and nonmetal atoms tend to gain electrons. The transfer has a predictable affect on the size of the ions that form. **Cations are always smaller than the atoms from which they form. Anions are always larger than the atoms from which they form.**

Figure 6.20 compares the relative sizes of the atoms and ions for three metals in Group 1A. For each of these elements, the ion is much smaller than the atom. For example, the radius of a sodium ion (95 pm) is about half the radius of a sodium atom (191 pm). When a sodium atom loses an electron, the attraction between the remaining electrons and the nucleus is increased. The electrons are drawn closer to the nucleus. Also, metals that are representative elements tend to lose all their outermost electrons during ionization. So the ion has one fewer occupied energy level.

The trend is the opposite for nonmetals like the halogens in Group 7A. For each of these elements, the ion is much larger than the atom. For example, the radius of a fluoride ion (133 pm) is more than twice the radius of a fluorine atom (62 pm). As the number of electrons increases, the attraction of the nucleus for any one electron decreases.

Look back at Figure 6.19. From left to right across a period, two trends are visible—a gradual decrease in the size of the positive ions followed by a gradual decrease in the size of the negative ions. Figure 6.21 summarizes the group and periodic trends in ionic size.

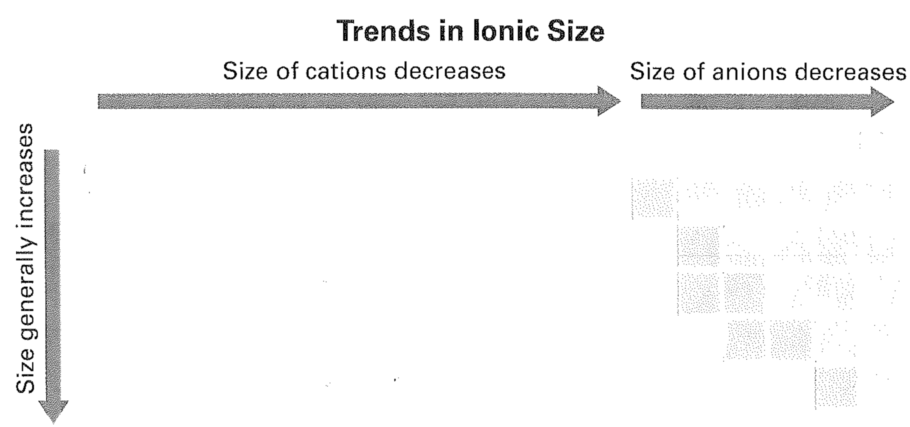


Figure 6.21 The ionic radii for cations and anions decrease from left to right across periods and increase from top to bottom within groups.

Trends in Electronegativity

In Chapters 7 and 8, you will study two types of bonds that can exist in compounds. Electrons are involved in both types of bonds. There is a property that can be used to predict the type of bond that will form during a reaction. This property is called electronegativity. **Electronegativity** is the ability of an atom of an element to attract electrons when the atom is in a compound. Scientists use factors such as ionization energy to calculate values for electronegativity.

Table 6.2 lists electronegativity values for representative elements in Groups 1A through 7A. The elements are arranged in the same order as in a periodic table. The noble gases are omitted because they do not form many compounds. The data in Table 6.2 is expressed in units called Paulings. Linus Pauling won a Nobel Prize in Chemistry for his work on chemical bonds. He was the first to define electronegativity.

In general, electronegativity values decrease from top to bottom within a group. For representative elements, the values tend to increase from left to right across a period. Metals at the far left of the periodic table have low values. By contrast, nonmetals at the far right (excluding noble gases) have high values. The electronegativity values among the transition metals are not as regular.

The least electronegative element is cesium, with an electronegativity value of 0.7. It has the least tendency to attract electrons. When it reacts, it tends to lose electrons and form positive ions. The most electronegative element is fluorine, with a value of 4.0. Because fluorine has such a strong tendency to attract electrons, when it is bonded to any other element it either attracts the shared electrons or forms a negative ion.

Checkpoint ***Why are values for noble gases omitted from Table 6.2?***

Go Online SCILINKS

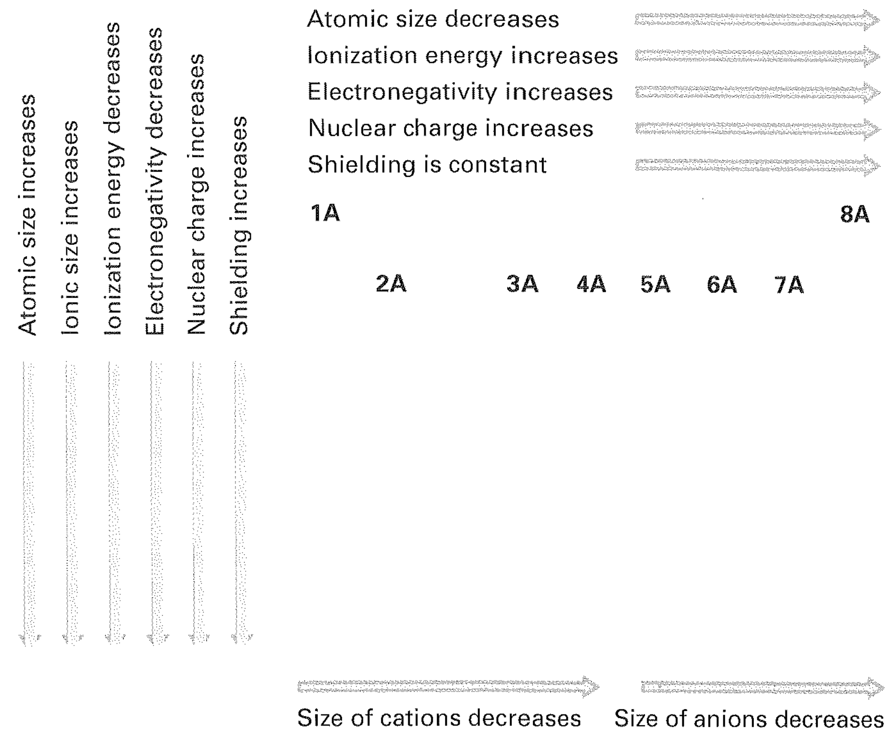
For: Links on Electronegativity
Visit: www.SciLinks.org
Web Code: cdn-1063

Table 6.2

Electronegativity Values for Selected Elements

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9		

Figure 6.22 Properties that vary within groups and across periods include atomic size, ionic size, ionization energy, electronegativity, nuclear charge, and shielding effect. **Interpreting Diagrams** ***Which properties tend to decrease across a period?***



Summary of Trends

Figure 6.22 shows the trends for atomic size, ionization energy, ionic size, and electronegativity in Groups 1A through 8A. These properties vary within groups and across periods. **The trends that exist among these properties can be explained by variations in atomic structure.** The increase in nuclear charge within groups and across periods explains many trends. Within groups an increase in shielding has a significant effect.

6.3 Section Assessment

16. **Key Concept** How does atomic size change within groups and across periods?

17. **Key Concept** When do ions form?

18. **Key Concept** What happens to first ionization energy within groups and across periods?

19. **Key Concept** Compare the size of ions to the size of the atoms from which they form.

20. **Key Concept** How does electronegativity vary within groups and across periods?

21. **Key Concept** In general, how can the periodic trends displayed by elements be explained?

22. Arrange these elements in order of decreasing atomic size: sulfur, chlorine, aluminum, and sodium. Does your arrangement demonstrate a periodic trend or a group trend?

23. Which element in each pair has the larger first ionization energy?

a. sodium, potassium

b. magnesium, phosphorus

Writing Activity

Explaining Trends in Atomic Size Explain why the size of an atom tends to increase from top to bottom within a group. Explain why the size of an atom tends to decrease from left to right across a period.

Interactive Textbook

Assessment 6.3 Test yourself on the concepts in Section 6.3.

with ChemASAP