

1 FOCUS

Objectives

- 6.3.1 Describe** trends among the elements for atomic size.
- 6.3.2 Explain** how ions form.
- 6.3.3 Describe** periodic trends for first ionization energy, ionic size, and electronegativity.

Guide for Reading

Build Vocabulary L2

Graphic Organizer Have students make a compare/contrast table for *cation* and *anion*, and use the table to decide which type of ion an element is likely to form.

Reading Strategy L2

Relate Text and Visuals Tell students that they will learn about trends related to the location of elements in the periodic table. They should look for the visuals throughout the section that summarize the trends.

2 INSTRUCT

Connecting to Your World

Have students read the opening paragraph. Ask, **What does the geometric pattern indicate about the arrangement of particles in salt?** (*It is orderly.*)

Trends in Atomic Size

Use Visuals L1

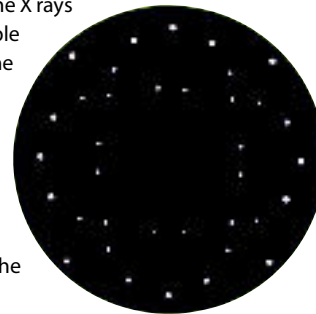
Figure 6.13 Ask, **Why can't a scientist measure the diameter of a single atom?** (*because an atom does not have a sharply defined border*) Discuss how measuring the distance between nuclei solves this problem.

FYI

In Chapter 8 there are formal definitions of molecule and diatomic molecule. The operational definition of a molecule should be sufficient for a discussion of atomic radii.

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.



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

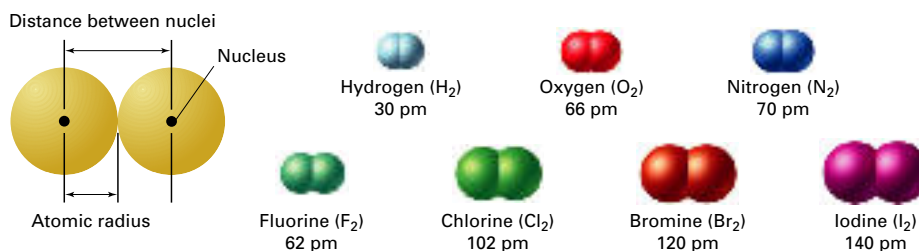
Building Vocabulary After you read this section, explain the difference between a cation and an anion.

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.



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Section Resources

Print

- **Guided Reading and Study Workbook**, Section 6.3
- **Core Teaching Resources**, Section 6.3 Review
- **Transparencies**, T70–T74
- **Small-Scale Chemistry Laboratory Manual**, Lab 9
- **Laboratory Practicals**, 6-1, 6-2

Technology

- **Interactive Textbook with ChemASAP**, Animation 7, Assessment 6.3
- **Go Online**, Section 6.3

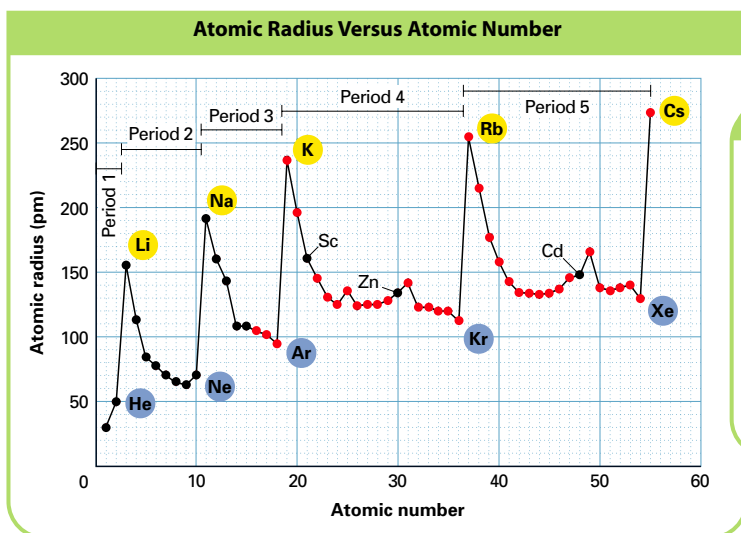


Figure 6.14 This graph plots atomic radius versus atomic number for 55 elements.

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.

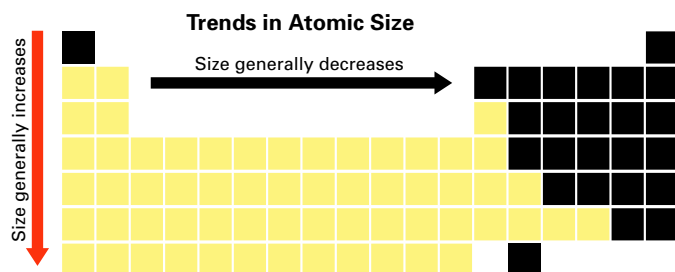


Figure 6.15 The size of atoms tends to decrease from left to right across a period and increase from top to bottom within a group. **Predicting** If a halogen and an alkali metal are in the same period, which one will have the larger radius?

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Interpreting Graphs

L2

- a.** potassium
b. It increases.
c. smaller

Because of the amount of data in this graph, you may need to help students get oriented before they begin to interpret the graph.

Enrichment Question

Emphasize the key roles electrical attraction and repulsion play within atoms and ions. Review the effects of increasing nuclear charge and changes in the shielding effect of electrons on the size of an atom. (Nuclear charge increases within groups and across periods; the shielding effect increases within groups, but is constant across periods.) Have students use these effects to describe the trends for atomic size within a period and within groups.

Discuss

L2

As an analogy to positions and trends in properties of elements in the periodic table, use seating charts and pricing data from local theaters or sports venues to discover trends. Ask students to determine patterns that relate the position of a seat to its price. Students should discover that variables such as distance from the stage or field, location relative to the center of the action, and whether the view will be obstructed, all affect price.

Facts and Figures

Elements and the Big Bang

At the time of the Big Bang, the temperature was many billions of degrees. Neutrons, protons, and electrons may have formed within 10^{-4} second after the Big Bang, and the lightest nuclei formed within 3 minutes. Matter was in the form of plasma, a sea of positive

nuclei and negative electrons. It took an estimated 500,000 years for electrons and nuclei to cool enough to form atoms. According to the Big Bang theory, Earth, with its wealth of chemical elements, formed from the debris of supernova explosions.

Answers to...

Figure 6.15 the alkali metal

Ions

CLASS Activity

Listing Elements

L2

Give students a list of elements. Ask them to locate each element in the periodic table, and decide whether its atoms are likely to form positive or negative ions. Have students make a list of elements that are likely to form positive ions and another list of elements that are likely to form negative ions.

FYI

The major discussion of ions and ionic bonding is in Chapter 7. This section presents the limited information that you will need to discuss the trends in ionization energy.

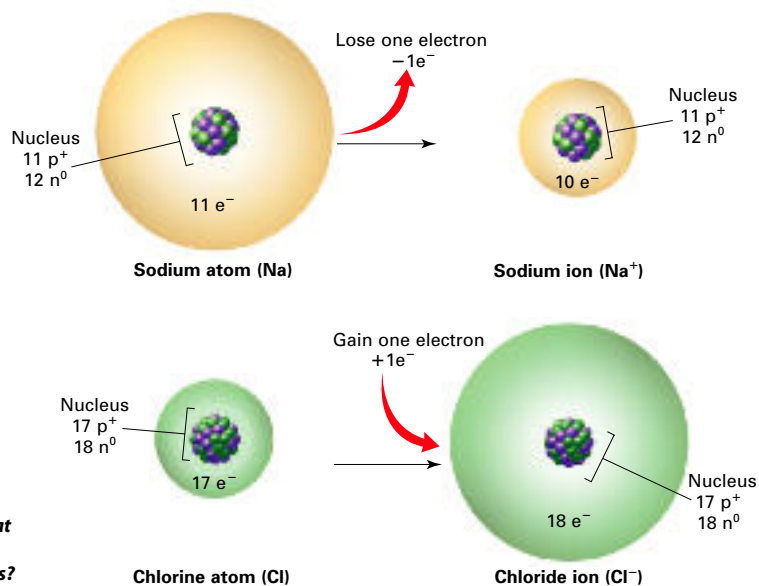


Figure 6.16 When a sodium atom loses an electron, it becomes a positively charged ion. When a chlorine atom gains an electron, it becomes a negatively charged ion.

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 Na¹⁺.

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.

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



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

with ChemASAP

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.

Trends in Ionization Energy

Discuss

L2

Explain that ionization energy is a measure of the difficulty in removing an electron from the highest occupied energy level. Ask, **Why is the first ionization energy of a nonmetal much higher than that of an alkali metal?** (Because the nuclear charge increases from left to right across a period and the shielding effect stays the same, it is more difficult to remove an electron.) There are ionization energy graphs in the Elements Handbook on R7, R11, R15, R19, R25, R29, and R33.

TEACHER Demo

Effective Nuclear Charge and Electron Shielding

L2

To help students understand the concepts of effective nuclear charge and electron shielding, choose four students to be “protons” and four to be “electrons.” Construct a lithium “nucleus” by having three protons stand together at the front of the room. Note that for purposes of this demo you are ignoring the neutrons. Place two electrons together at a short distance from the nucleus to represent the 1s electrons. Place the third electron a bit farther away to represent the 2s electron. You should be able to draw a line from the nucleus through the 1s electrons to the 2s electron. Point out that there are no other electrons between the 1s electrons and the nucleus. Thus, these electrons experience the full impact of the 3+ charge. Because the third electron’s “view” of the nucleus is partially blocked, it is shielded somewhat from the full force of the 3+ charge.

FYI

Because the unit for ionization energy is kJ/mol, a footnote in Table 6.1 supplies an operational definition of mole, which is introduced in Section 10.1.

Answers to...

Figure 6.16 nothing



Checkpoint A cation has a positive charge. An anion has a negative charge.

Section 6.3 (continued)

Interpreting Graphs

L2

- a. lithium; sodium
 b. First ionization energy decreases as atomic number increases.
 c. hydrogen; it has only one electron.
- Because of the amount of data in this graph, you may need to help students get oriented before they begin to interpret the graph.

Enrichment Question

L3

Challenge students to explain why the portion of the graph for periods 4 and 5 is different from the portion of the graph for periods 2 and 3.

(Periods 4 and 5 include transition metals, whose atoms have electrons in *d* orbitals.)

TEACHER Demo

Predicting Reactivity

L2

Purpose Students observe the relative reactivities of magnesium and calcium and predict relative reactivities for other pairs of elements.

Materials 20 mL 1M HCl, two 50-mL beakers, overhead projector, 20 cm magnesium ribbon, 1 g calcium

Safety Wear goggles for this demo.

Procedure Pour 20 mL HCl into each beaker. Set the beakers on an overhead projector. Coil the magnesium ribbon and drop it into one beaker. Drop 1 g calcium into the other beaker. Compare the reaction rates in the two beakers. Point out the positions of the two elements in the periodic table, and relate the difference in reactivity to their first and second ionization energies. Ask students to predict the relative reactivities of other pairs of elements in Groups 1A and 2A.

Expected Outcome The calcium fizzes in the HCl. The magnesium reacts more slowly with the HCl.

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

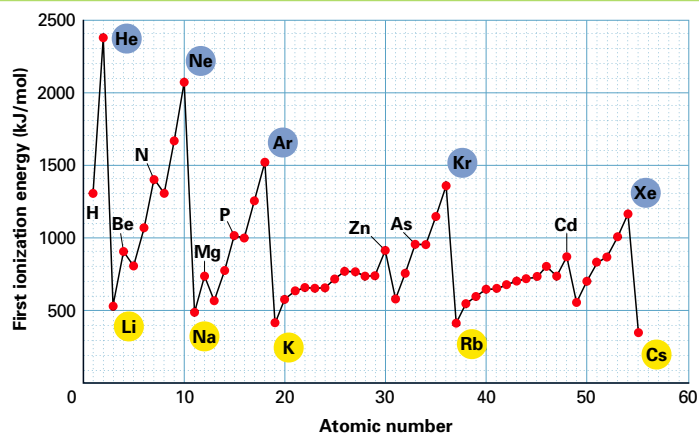
INTERPRETING GRAPHS

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.

First Ionization Energy Versus Atomic Number



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.

Trends in First Ionization Energy

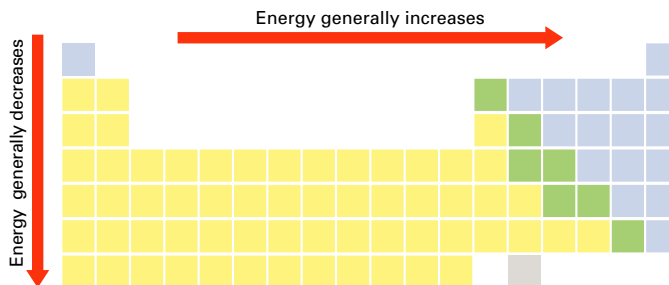


Figure 6.18 First ionization energy tends to increase from left to right across a period and decrease from top to bottom within a group.

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

- Describe how the size changes when an atom forms a cation and when an atom forms an anion.
- How do the ionic radii vary within a group of metals? How do they vary within a group of nonmetals?
- Describe the shape of a portion of the graph that corresponds to one period.
- Is the trend across a period similar or different for periods 2, 3, 4, and 5?
- Propose explanations for the trends you have described for ionic radii within groups and across periods.

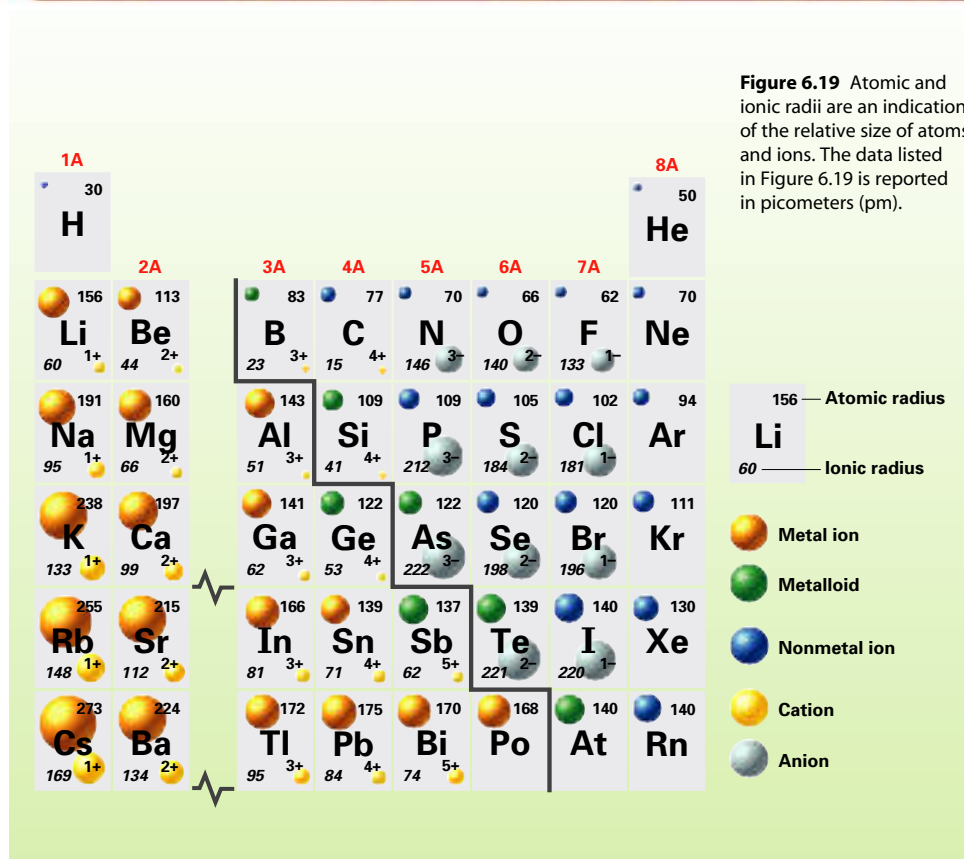
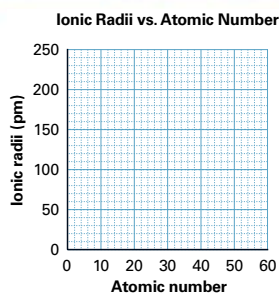


Figure 6.19 Atomic and ionic radii are an indication of the relative size of atoms and ions. The data listed in Figure 6.19 is reported in picometers (pm).

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Quick LAB

Periodic Trends in Ionic Radii **L2**

Objective After completing this activity, students will be able to

- identify periodic trends in ionic size.

Skills Focus Using tables and graphs, predicting, drawing conclusions



Prep Time none

Class Time 40 minutes

Teaching Tips

- If time is too limited for students to make the graph, use Figure 6.19 to answer Questions 1, 2, 4, and 5.
- You may want to reference the radii diagrams in the Elements Handbook on R7, R11, R15, R19, R25, R29, and R33.

Expected Outcome

Ionic radii increase from top to bottom within a group. The radii of cations and anions decrease from left to right across a period.

Analyze and Conclude

- Cations are smaller than their atoms; anions are larger than their atoms.
- Ionic radii increase from top to bottom within a group of metals or within a group of nonmetals.
- Two portions of the curve slope down from left to right.
- The trend is similar for the periods.
- The radii increase within a group because the number of occupied energy levels increases. The radii of cations decrease across a period because the nuclear charge increases, the shielding effect is constant, and the number of electrons decreases. (The effect is smaller with anions because the number of electrons increases.)

For Enrichment **L3**

Have students use the graph on page R41 to describe the periodic trend in atomic size for transition metals. Ask how the trend for transition metals compares to the trend for representative elements.

Answers to...

Figure 6.18 an alkali metal in period 2

Trends in Ionic Size

Discuss

L2

Relate the periodic trends in ionic size to those discussed earlier for atomic size. Explain that the effective nuclear charge experienced by an electron in the highest occupied orbital of an atom or ion is equal to the total nuclear charge (the number of protons) minus the shielding effect due to electrons in lower energy levels. The effective nuclear charge determines the atomic and ionic radii. As you proceed from left to right in any given period, the principal quantum number, n , of the highest occupied energy level remains constant, but the effective nuclear charge increases. Therefore, atomic and ionic radii decrease as you move to the right in a period. In contrast, within any group, as you proceed from top to bottom, the effective nuclear charge remains nearly constant, but the principal quantum number, n , increases. Consequently, atomic and ionic radii increase from top to bottom within a group.

TEACHER Demo

Trends in Ionic Size

L1

Purpose Students observe an analogy for the effect of adding or removing electrons from an atom.

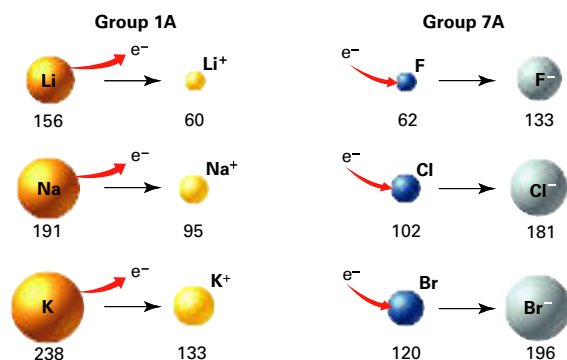
Materials washers or other small circular items, smaller item to represent the nucleus, overhead projector

Procedure On the overhead projector, make a circle of washers to represent an electron cloud in a neutral atom. The washers should be touching. Place the “nucleus” in the center of the circle. Add or subtract washers to mimic ion formation. With each change, adjust the circle so that the washers are still touching. Explain that the change in the diameter of the circle is analogous to the change in the effective attraction of the nuclear charge for electrons.

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.

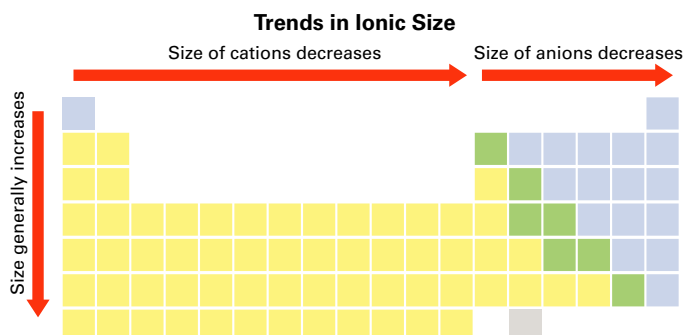


Figure 6.21 The ionic radii for cations and anions decrease from left to right across periods and increase from top to bottom within groups.

Differentiated Instruction

Gifted and Talented


L3

Have students research and describe the phenomenon of the lanthanide contraction. Ask them to discuss how the lanthanide contraction accounts for the fact that zirconium and hafnium have virtually the same atomic radius even though hafnium is below zirconium in Group 4B of the periodic table.

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?

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		


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

Trends in Electronegativity Discuss

Lead a class discussion on periodic and group trends in electronegativities. Point out that electronegativity values help chemists predict the type of bonding that exists between atoms in compounds. Ask, **Why aren't the noble gases included in a discussion on electronegativity?** (*because they form very few compounds*) **Which element is the most electronegative and which is the least electronegative?** (*fluorine; cesium*) You may want to reference the electronegativity graphs in the Elements Handbook on R7, R11, R15, R19, R25, R29, and R33.

FYI

The values for electronegativity are often based on values for ionization energy and electron affinity. Ionization energy is a measure of an atom's ability to lose electrons. Electron affinity is a measure of an atom's ability to gain electrons.



Download a worksheet on **Electronegativity** for students to complete, and find additional teacher support from NSTA SciLinks.

Answers to...

Figure 6.20 The radius decreases for cations and increases for anions.

 **Checkpoint** They rarely form compounds.

Summary of Trends

Use Visuals

L1

Figure 6.22 Point out that this diagram incorporates information from several diagrams earlier in the chapter. Ask students whether they find this or earlier diagrams more helpful, and why.

ASSESS

Evaluate Understanding

L2

Have students compare two elements in the same group in terms of atomic radius, ionic radius, ionization energy, and electronegativity. Repeat the exercise with a metal and nonmetal from the same period. Have students write general statements to summarize the trends revealed by these comparisons.

Reteach

L1

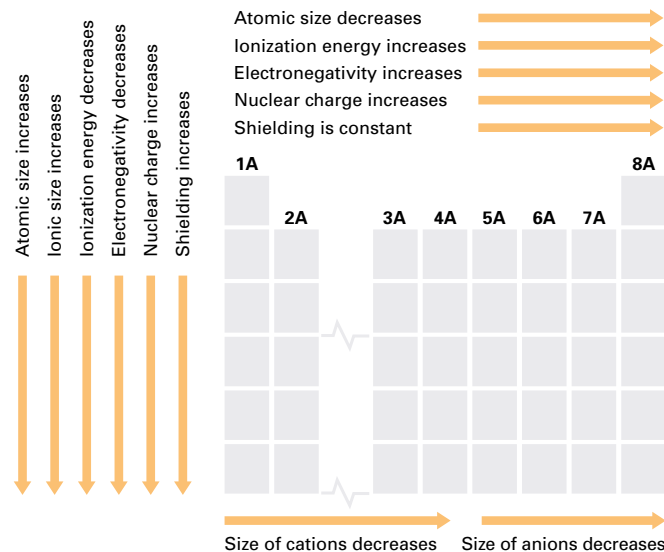
Review the terms used in Figure 6.22. Then, use the periodic table and the terms to play a version of "I'm Thinking of..." For example, choose fluorine and say you are thinking of an element that has a very small atomic size and a very high electronegativity. Let students guess, then discuss the correct answer. Have students continue the game in small groups.

Writing Activity

From top to bottom within a group, the effective nuclear charge remains nearly constant and the principal quantum number increases. From left to right in a period, the principal quantum number remains constant, but the effective nuclear charge increases.

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

- Key Concept** How does atomic size change within groups and across periods?
- Key Concept** When do ions form?
- Key Concept** What happens to first ionization energy within groups and across periods?
- Key Concept** Compare the size of ions to the size of the atoms from which they form.
- Key Concept** How does electronegativity vary within groups and across periods?
- Key Concept** In general, how can the periodic trends displayed by elements be explained?
- 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?
- Which element in each pair has the larger first ionization energy?
 - sodium, potassium
 - 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

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Section 6.3 Assessment

- Atomic size generally increases within a group and decreases from left to right across a period.
- Ions form when electrons are transferred between atoms.
- decreases within a group and increases from left to right across a period
- Anions are larger and cations are smaller than the atoms from which they form.
- decreases from top to bottom within a group and increases from left to right across a period
- The trends can be explained by variations in atomic structure.
- sodium, aluminum, sulfur, chlorine; periodic trend
- a. sodium b. phosphorus

Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 6.3.

with ChemASAP

Small-Scale LAB

Periodicity in Three Dimensions

Purpose

To build three-dimensional models for periodic trends.

Materials

- 96-well spot plate
- straws
- scissors
- metric ruler
- permanent fine-line marker

Procedure



1. Measure the depth of a well in the spot plate by inserting a straw into a well and holding the straw upright as shown in the photograph. Make a mark on the straw at the point where the straw meets the surface of the plate. Measure the distance from the end of the straw to the mark in centimeters. Record this distance as well depth.
2. Cut the straw to a length that is 4.0 cm plus well depth. The straw will extend exactly 4.0 cm above the surface of the plate.
3. Fluorine has an electronegativity value of 4.0. On a scale of one cm equals one unit of electronegativity, the portion of the straw that extends above the surface of the plate represents the electronegativity value for fluorine. Using the same scale, cut straws to represent the electronegativity values for all the elements listed in Table 6.2. Remember to add the well depth to the electronegativity value before cutting a straw. As you cut the straws, mark each straw with the chemical symbol of the element that the straw represents.
4. Arrange the straws in the spot plate in rows and columns to match the locations of the elements in the periodic table.
5. Make a rough sketch of your completed model.

Analyze

Observe your model and record the answers to the following questions below your sketch.

1. Which element represented in your model is the most electronegative?
2. Based on your model, what is the general trend in electronegativity from left to right across a period?



3. Relate the trend in electronegativity across a period to the location of metals and nonmetals in the periodic table.
4. What is the general trend in electronegativity within a group? Are there any notable exceptions?
5. Why do you think that the electronegativity value for hydrogen is so high given its location in the table?

You're the Chemist

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Design It!** Construct a similar 3-D model for first ionization energies. Use the data in Table 6.1. Use a scale of one cm equals 300 kJ/mol.
2. **Design It!** Design and construct a 3-D model that shows trends in atomic and ionic radii for the elements in Groups 1A and 7A. Devise a way to display both ionic and atomic radii in the same model.
3. **Analyze it!** Xenon has an electronegativity value of 2.6. Cut and place a straw in your first model to represent xenon. Does xenon support the trend for electronegativity across a period? Is xenon likely to form compounds? Explain your answers.

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Small-Scale LAB

Periodicity in Three Dimensions

L2

Objective

After completing this activity, students should be able to:

- build concrete models to reinforce periodic trends.
- apply a procedure to a new variable and design a model on their own.



Prep Time 10 minutes

Class Time 40 minutes

Materials 96-well spot plates, straws, scissors, metric rulers, permanent fine-line markers

Advance Prep Straws with a 1/4-inch diameter fit snugly in the wells.

Teaching Tips

- Students can use colored straws to color code groups or periods.
- If you don't have spot plates, press a lump of clay the size of a golf ball flat on a table with a block of wood. Students mark out a 1-cm square grid and insert the straws in the clay.

Expected Outcome Students produce 3-D models for periodic trends.

Analyze

1. fluorine
2. Electronegativity generally increases from left to right along a period.
3. Metals, which are on the left side of the table, have lower electronegativity values than nonmetals, which are on the right.
4. Electronegativity generally increases from bottom to top within a group. Except for boron, the rest of Group 3A shows a reverse in this trend.
5. Hydrogen is placed in Group 1A based on its electron configuration, but is classified as a nonmetal.

For Enrichment

L3

Have students use the data on page R41 to make a 3-D model of trends in atomic size for transition metals.

Answers to...

Figure 6.22 size of atoms and ions

You're the Chemist

1. Students divide the values of first ionization energies by 300 and measure the appropriate length of straws.
2. Students must determine their own scale before they begin. Students often use two wells to represent both ionic and atomic radii. Other students cut a straw to a length that represents the larger radius of an atom and mark the straw to show the smaller radius of the corresponding cation.

3. The value for xenon is similar to iodine, which is consistent with the general trend. Based on this value, xenon appears to have the ability to attract electrons and form compounds.