Chapter 5

Atoms and Bonding

CALIFORNIA Standards Preview

S 8.3 Each of the more than 100 elements of matter has distinct properties and a distinct atomic structure. All forms of matter are composed of one or more of the elements. As a basis for understanding this concept:

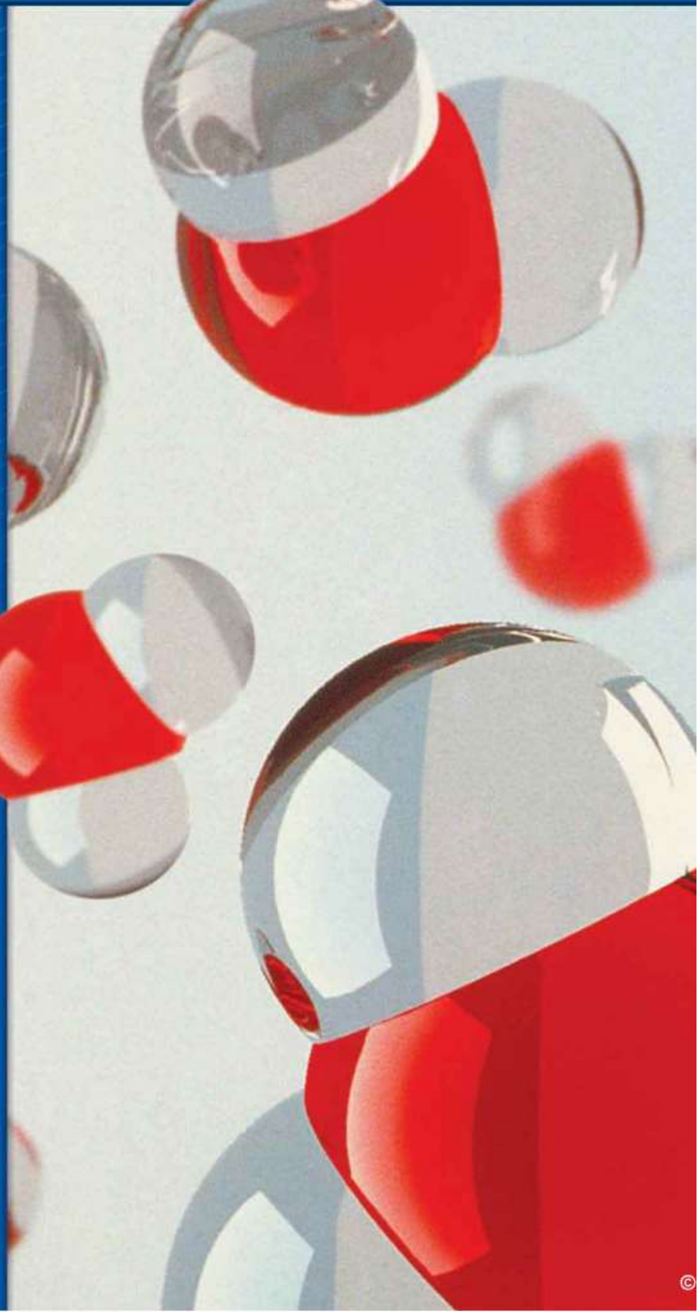
- **b.** Students know that compounds are formed by combining two or more different elements and that compounds have properties that are different from their constituent elements.
- **c.** Students know that atoms and molecules form solids by building up repeating patterns, such as the crystal structure of NaCl or long-chain polymers.
- **f.** Students know how to use the periodic table to identify elements in simple compounds.

S 8.7 The organization of the periodic table is based on the properties of elements and reflects the structure of atoms. As a basis for understanding this concept:

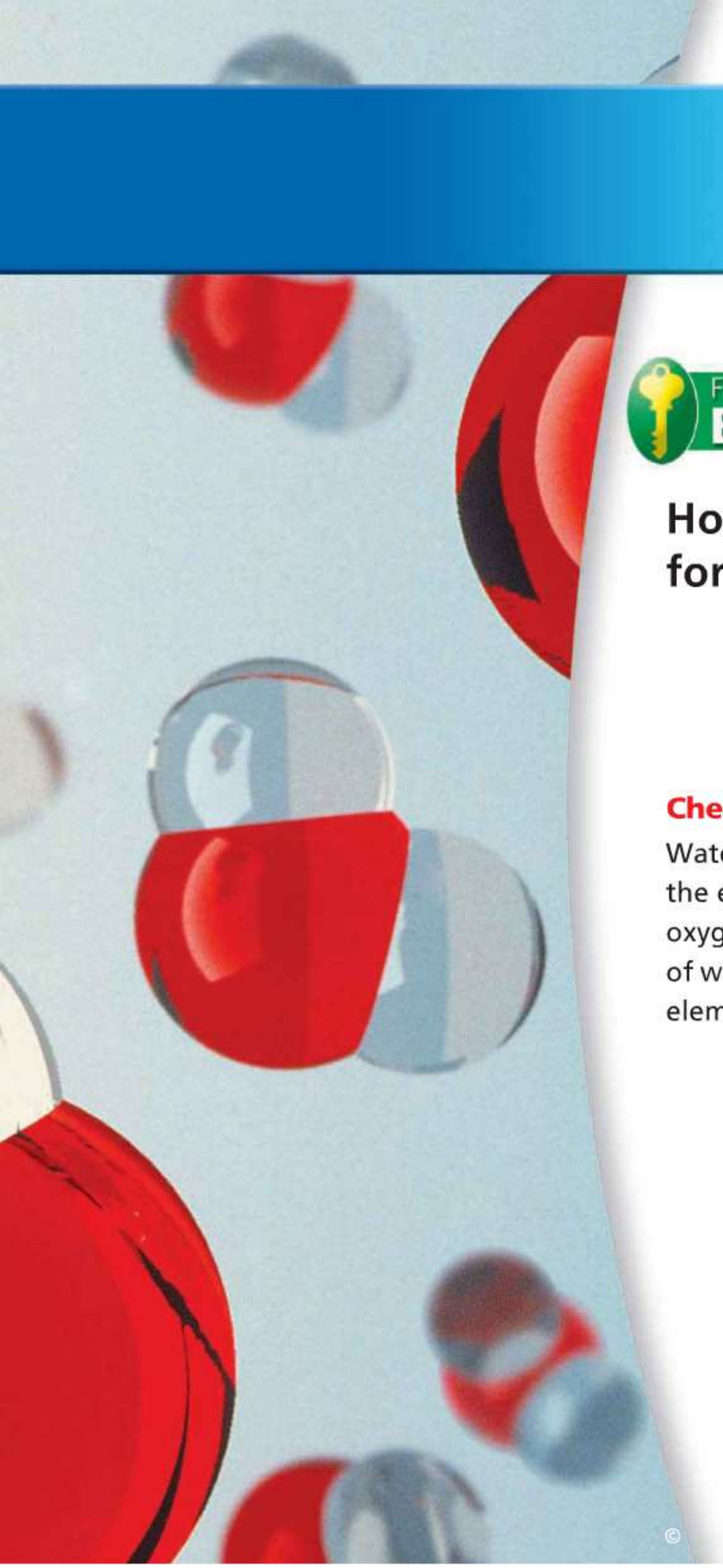
- **c.** Students know substances can be classified by their properties, including their melting temperature, density, hardness, and thermal and electrical conductivity.

S 8.9 Scientific progress is made by asking meaningful questions and conducting careful investigations. As a basis for understanding this concept and addressing the content in the other three strands, students should develop their own questions and perform investigations. Students will:

- **c.** Distinguish between variable and controlled parameters in a test.
- **g.** Distinguish between linear and nonlinear relationships on a graph of data.



These models represent molecules of water (H_2O).



Video Preview
Discovery Channel School
Atoms and Bonding

S 8.3.b

Focus on the **BIG Idea**

How do compounds form?

Check What You Know

Water is a compound made from the elements hydrogen and oxygen. How do the properties of water differ from those of the elements that it is made up of?

Build Science Vocabulary

The images shown here represent some of the key terms in this chapter. You can use this vocabulary skill to help you understand the meaning of some key terms in this chapter.

Vocabulary Skill

High-Use Academic Words

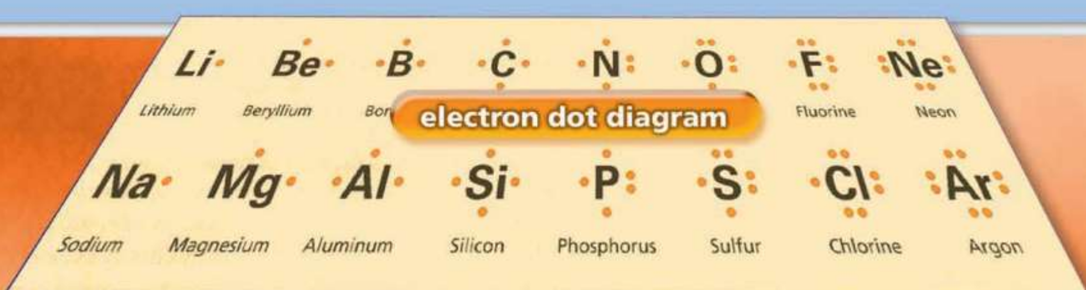
High-use academic words are words you are likely to meet while reading textbooks. Look for the following words in context as you read this chapter.

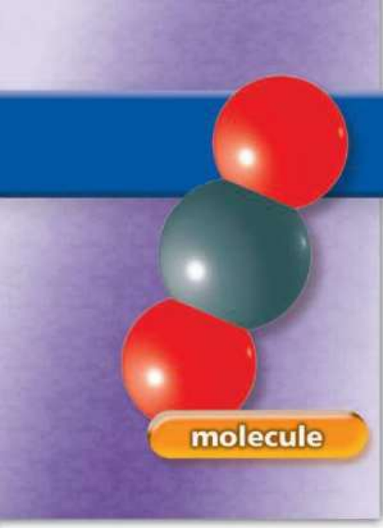
Word	Definition	Example Sentence
conduct (kahn **DUKT**) p. 199	*v.* To allow something to travel along or through it	Metal strips on a circuit board conduct electric current.
stable (**STAY** bul) p. 177	*adj.* Not easily or quickly changed from one state to another	Gold is a stable metal that does not rust or tarnish.
structure (**STRUK** chur) p. 178	*n.* The way in which parts of something are put together	The outside structure of the building is made of brick and concrete.
symbol (**SIM** bul) p. 187	*n.* A written sign that stands for something else	The symbol for the element oxygen is O.

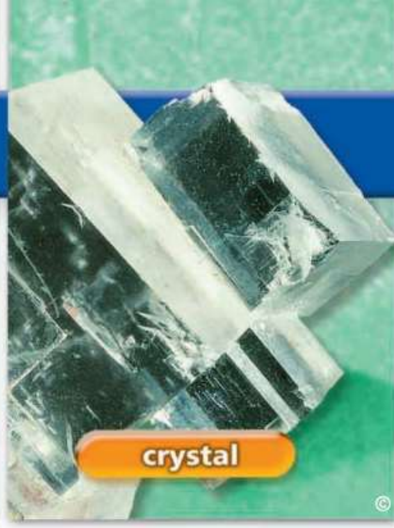
Apply It!

Choose the word that best completes the sentence.

1. "H" is the __________ for hydrogen.
2. The __________ of an atom consists of a nucleus of protons and neutrons, surrounded by a cloud of moving electrons.
3. Platinum jewelry lasts a long time because the metal is very __________.







Chapter 5 Vocabulary

Section 1 (page 176)
valence electron
electron dot diagram
chemical bond

Section 2 (page 184)
ion
polyatomic ion
ionic bond
ionic compound
chemical formula
subscript
crystal

Section 3 (page 192)
covalent bond
molecule
double bond
triple bond
molecular compound
polar bond
nonpolar bond

Section 4 (page 198)
alloy
metallic bond

Interactive Textbook
Build Science Vocabulary
Online
Visit: PHSchool.com
Web Code: cxj-2050

How to Read Science

Reading Skill

Compare and Contrast

When you compare and contrast, you examine the similarities and differences among things. You can compare and contrast by using a table.

Follow these steps to set up a compare–contrast table:

- List the items to be compared in the first column.
- List the characteristics to be compared across the top of the table.
- Complete the table by filling in information about each characteristic.

In this chapter, you will learn about chemical compounds, such as table salt (NaCl). Look at the compare–contrast table below.

Compounds and Their Component Elements

Substance	Color	State at Room Temperature
Table Salt (NaCl)	White	
Sodium (Na)	Silvery white	Solid
Chlorine (Cl)	Greenish yellow	Gas

Apply It!

After reading Section 2, copy and complete the compare–contrast table above to compare the properties of sodium chloride with those of sodium and chlorine. After reading Section 3, compare the properties of molecular compounds with those of ionic compounds.

Lab zone

Standards Investigation

S 8.3.c

Models of Compounds

In this chapter, you will learn how atoms of elements react with one another to form compounds. When they form compounds, the atoms become chemically bonded to each other. In this investigation, you will create models of chemical compounds.

Your Goal

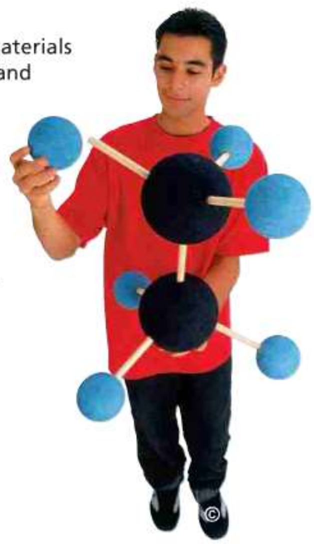
To make models demonstrating how atoms bond in ionic compounds and in molecular compounds

To complete the investigation, you must

- select appropriate materials to make models of atoms
- indicate the number of bonds each atom forms
- use your model atoms to compare compounds that contain ionic bonds with compounds that contain covalent bonds
- follow the safety guidelines in Appendix A

Plan It!

Brainstorm with some classmates about materials you can use to represent different atoms and chemical bonds. Look ahead in the chapter to preview ionic and covalent bonding. Think about how you will show that ionic and covalent bonding are different. You may need to find some small, but highly visible, objects to represent electrons. Be ready to display your models and explain what they show.



Section 1

Atoms, Bonding, and the Periodic Table

CALIFORNIA Standards Focus

S 8.3.f Students know how to use the periodic table to identify elements in simple compounds.

- How is the reactivity of elements related to valence electrons in atoms?
- What does the periodic table tell you about the atoms of elements?

Key Terms

- valence electron
- electron dot diagram
- chemical bond

Lab zone **Standards Warm-Up**

What Are the Trends in the Periodic Table?

1. Examine the periodic table of the elements that your teacher provides. Look in each square for the whole number located above the symbol of the element. As you read across a row from left to right, what trend do you see?
2. Now look at a column from top to bottom. What trend do you see in these numbers?

Think It Over

Interpreting Data Can you explain why one row ends and a new row starts? Why are certain elements in the same column?

Why isn't the world made only of elements? How do the atoms of different elements combine to form compounds? The answers to these questions are related to electrons and their energy levels. And the roadmap to understanding how electrons determine the properties of elements is the periodic table.

Valence Electrons and Bonding

You learned earlier about electrons and energy levels. An atom's **valence electrons** (VAY luns) are those electrons that have the highest energy level and are held most loosely. **The number of valence electrons in an atom of an element determines many properties of that element, including the ways in which the atom can bond with other atoms.**

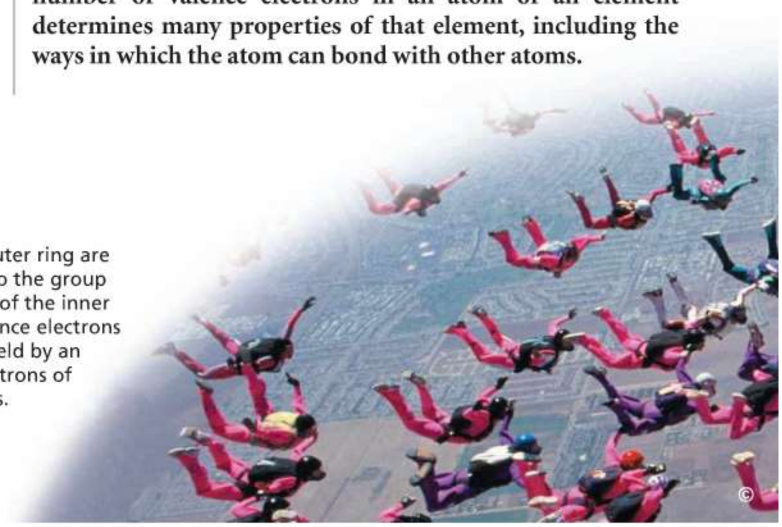


FIGURE 1
Valence Electrons
Sky divers in the outer ring are less securely held to the group than are members of the inner ring. Similarly, valence electrons are more loosely held by an atom than are electrons of lower energy levels.

Electron Dot Diagrams Each element has a specific number of valence electrons, ranging from 1 to 8. Figure 2 shows one way to depict the number of valence electrons in an element. An **electron dot diagram** includes the symbol for the element surrounded by dots. Each dot stands for one valence electron.

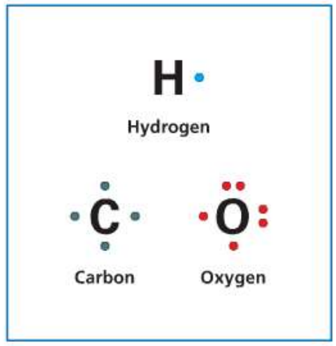
Chemical Bonds and Stability Atoms of most elements are more stable—that is, less likely to react—when they have eight valence electrons. For example, atoms of neon, argon, krypton, and xenon all have eight valence electrons and are very unreactive. These elements do not easily form compounds. Some small atoms, such as helium, are stable with just two valence electrons.

Atoms usually react in a way that makes each atom more stable. One of two things can happen: Either the number of valence electrons increases to eight (or two, in the case of hydrogen). Or, the atom gives up loosely held valence electrons. Atoms that react this way can become chemically combined, that is, bonded to other atoms. A **chemical bond** is the force of attraction that holds two atoms together as a result of the rearrangement of electrons between them.

Chemical Bonds and Chemical Reactions When atoms bond, electrons may be transferred from one atom to another, or they may be shared between the atoms. In either case, the change results in a chemical reaction—that is, new substances form. Later in this chapter, you will learn which elements are likely to gain electrons, which are likely to give up electrons, and which are likely to share electrons. You will also learn how the periodic table of the elements can help you predict how atoms of different elements react.

Reading Checkpoint **What information does an electron dot diagram show?**

FIGURE 2
Electron Dot Diagrams
An atom's valence electrons are shown as dots around the symbol of the element. Notice that oxygen atoms have six valence electrons. **Predicting** *How many more electrons are needed to make an oxygen atom stable?*



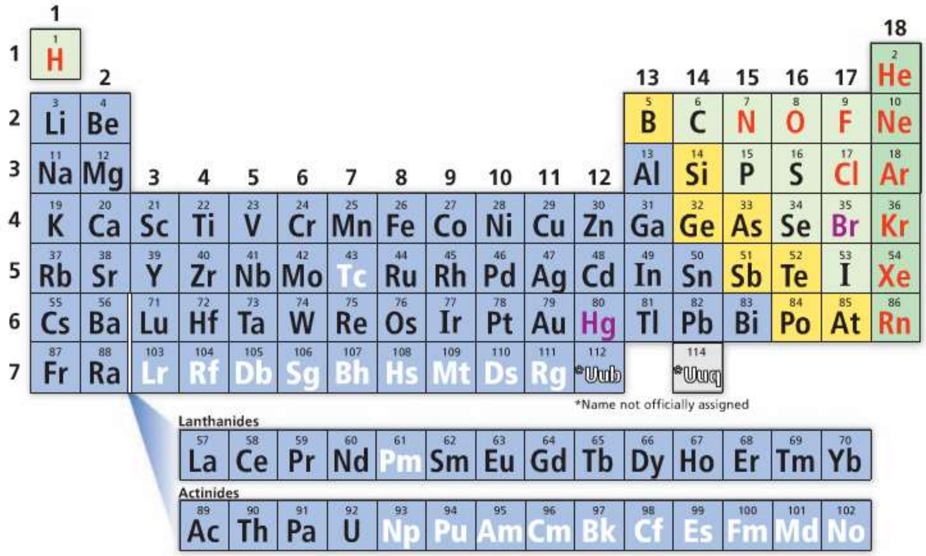


FIGURE 3
Periodic Table of the Elements
The periodic table of elements is arranged horizontally in order of increasing atomic number (number of protons) and vertically in columns of elements with similar chemical properties.

How the Periodic Table Works

The periodic table is a useful tool for identifying the elements in simple compounds. The table can also help you understand how atoms form bonds. **The periodic table reveals the underlying atomic structure of atoms, including the arrangement of the electrons.** If you know the number of valence electrons that atoms of different elements have, you have a clue as to which elements combine and how.

Relating Periods and Groups Look at Figure 3. Think about how the atoms change from left to right across a period. As the atomic number increases, the number of electrons also increases. Except for Period 1, a given period ends when the number of valence electrons reaches eight. The next period begins with atoms having valence electrons with higher energy. This repeating pattern means that the elements within a group always have the same number of valence electrons. As a result, they have similar properties.

Figure 4 compares the electron dot diagrams of elements in Periods 2 and 3. Notice that each element has one more valence electron than the element to its left. For example, Group 1 elements have one valence electron. The elements in Group 2 have two. Elements in Group 13 have three valence electrons, elements in Group 14 have four, and so on. (Elements in Groups 3 to 12 follow a slightly different pattern.)

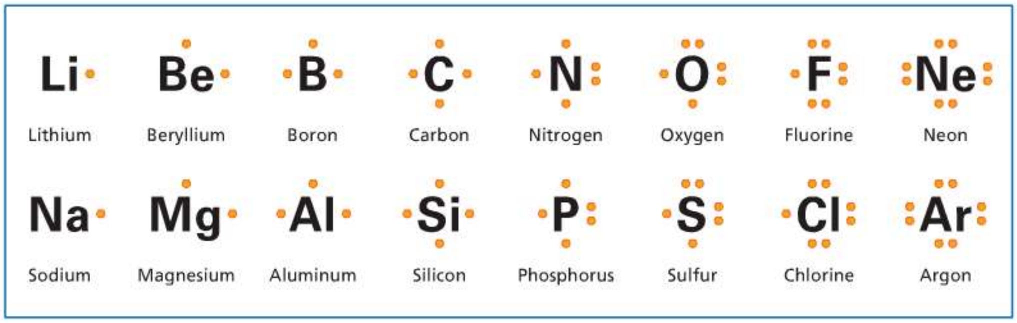


FIGURE 4
Patterns of Valence Electrons
After the number of valence electrons reaches 8, a new period begins.

Inert Gases The Group 18 elements are the inert gases. Atoms of these elements have eight valence electrons, except for helium, which has two. As you have read, atoms with eight valence electrons (or two, in the case of helium) are stable. Such atoms are unlikely to transfer electrons to other atoms or to share electrons with other atoms. As a result, inert gases do not react easily with other elements. Even so, chemists have been able to make inert gases form compounds with a few other elements.

Reactive Nonmetals and Metals Now look at the elements in the column just to the left of the inert gases. The elements in Group 17, the halogens, have atoms with seven valence electrons. A gain of just one more electron gives these atoms the stable number of eight electrons, as in the inert gases. As a result, the halogens react easily with other elements whose atoms can give up or share electrons.

At the far left side of the periodic table is Group 1, the alkali metal family. Atoms of the alkali metals have only one valence electron. Except for lithium, the next lowest energy level has a stable set of eight electrons. (Lithium atoms have a stable set of two electrons at the next lowest energy level.) Therefore, alkali metal atoms can become chemically more stable by losing their one valence electron. This property makes the alkali metals very reactive.

Reading Checkpoint **How are atoms of the elements in Group 1 similar?**

FIGURE 5
Reactivity of Chlorine
Chlorine is so reactive that steel wool burns when exposed to the chlorine gas in this jar.
Relating Cause and Effect *Why is chlorine so reactive?*

Video Field Trip
Discovery Channel School
Atoms and Bonding

Other Metals Look at the elements in Groups 2 through 12 of the periodic table. Like the Group 1 elements, these elements are metals. Most have one, two, or three valence electrons. They react by losing these electrons, especially when they combine with oxygen or one of the halogens.

How reactive a metal is depends on how easily its atoms lose valence electrons. Some metals, such as those in Group 2 (the alkaline earth metals), lose electrons easily and are almost as reactive as the alkali metals of Group 1. Other metals, such as platinum (Pt) in Group 10 and gold (Au) in Group 11, are unreactive. In general, the reactivity of metals decreases from left to right across the periodic table. Among Groups 1 and 2, reactivity increases from top to bottom.

Science and **History**

1875
Gallium
The French chemist Paul-Émile Lecoq de Boisbaudran discovered an element that he called gallium. It had properties predicted by Mendeleev for an unknown element that would fit directly below aluminum in the periodic table.

1894
Argon, Neon, Krypton, and Xenon
British chemist William Ramsay discovered an element he named argon, after the Greek word for "lazy." The name fits because argon does not react with other elements. Ramsay looked for other nonreactive gases and discovered neon, krypton, and xenon.

1898
Polonium and Radium
Chemists Marie Curie and her husband Pierre had to start with three tons of uranium ore to isolate a few grams of two new elements, which were named polonium and radium.

1830 1865 1900

Other Nonmetals Elements in the green color-coded region of the periodic table are the nonmetals. Five nonmetals are solids, four are gases, and one is a liquid. All of the nonmetals have four or more valence electrons. Like the halogens, other nonmetals become stable when they gain or share enough electrons to have a set of eight valence electrons.

The nonmetals combine with metals usually by gaining electrons. But nonmetals can also combine with other nonmetals by sharing electrons. Of the nonmetals, oxygen and the halogens are highly reactive. In fact, fluorine is the most reactive element known. It even forms compounds with some of the inert gases.

Writing in Science

Research and Write Various forms of the periodic table exist. The periodic tables used by physicists and engineers may include different information than the tables used by chemists. Use the library to find two different forms of the periodic table. Write a paragraph comparing the two tables and how scientists might use them.

1939
Francium
Although Mendeleev predicted the properties of an element he called "eka-cesium," the element was not discovered until 1939. French chemist Marguerite Perey named her discovery francium, after the country France.

1941
Plutonium
American chemist Glenn Seaborg was the first to isolate plutonium, which is found in small amounts in uranium ores. Plutonium is used as fuel in certain nuclear reactors. It has also been used to power equipment used in space exploration.

1997
Elements 101 to 109
The International Union of Pure and Applied Chemists (IUPAC) agreed on names for elements 101 to 109. Many of the names honor scientists. Seaborgium is named after Glenn Seaborg. Meitnerium is named after Lise Meitner, shown here in 1946. All the new elements are synthetic, and none is stable enough to exist in nature.

2003 to Present
Darmstadtium
Element 110, first created in the mid-1990s, is named darmstadtium. Research to produce and study new synthetic elements continues.

1935 | 1970 | 2005

The quartz movement of the watch

FIGURE 6
A Semimetal at Work
This quartz-movement watch keeps time with a small quartz crystal, a compound made of the semimetal silicon and the nonmetal oxygen. The crystal vibrates at about 32,000 vibrations per second when a voltage is applied.

Semimetals Several elements known as semimetals lie along a zigzag line between the metals and nonmetals. The semimetals have from three to six valence electrons. They can either lose or share electrons when they combine with other elements. So, depending on the conditions, these elements can behave as either metals or nonmetals.

Hydrogen Notice that hydrogen is located above Group 1 in the periodic table. It is placed there because it has only one valence electron. However, hydrogen is not considered a metal. It is a reactive element, but its properties differ greatly from those of the alkali metals.

Reading Checkpoint **Why is hydrogen grouped above the Group 1 elements even though it is not a metal?**

Section 1 Assessment

S 8.3.f, E-LA: Reading 8.1.0

Vocabulary Skill **High-Use Academic Words** Use the word *stable* to explain why the halogens tend to combine easily with other elements.

Reviewing Key Concepts

HINT 1. a. **Defining** What are valence electrons?

HINT b. **Reviewing** What role do valence electons play in the formation of compounds from elements?

HINT c. **Comparing and Contrasting** Do oxygen atoms become more stable or less stable when oxygen forms compounds? Explain.

HINT 2. a. **Summarizing** Summarize how the periodic table is organized, and tell why this organization is useful.

HINT b. **Explaining** Why do the properties of elements change in a regular way across a period?

HINT c. **Relating Cause and Effect** Explain the reactivity of the inert gases in terms of valence electrons.

Lab zone At-Home Activity

Looking for Elements Find some examples of elements at home. Then locate the elements on the periodic table. Show your examples and the periodic table to your family. Point out the positions of the elements on the table and explain what the periodic table tells you about the elements. Include at least two nonmetals in your discussion. (*Hint:* The nonmetals may be invisible.)

Lab zone **Skills Lab**

Comparing Atom Sizes

S 8.3.f, 8.9.g

Problem

How is the radius of an atom related to its atomic number?

Skills Focus

making models, graphing, interpreting data

Materials

- drawing compass
- metric ruler
- calculator
- periodic table of the elements

Procedure

1. Using the periodic table as a reference, predict whether the size (radius) of atoms will increase, remain the same, or decrease as you go from the top to the bottom of a group, or family, of elements.
2. The data table lists the elements in Group 2 in the periodic table. The atomic radius of each element is given in picometers (pm). Copy the data table into your notebook.
3. Use the periodic table to look up the atomic numbers of the Group 2 elements. Record the values in your data table.
4. Calculate the relative radius of each atom compared to beryllium, the smallest atom listed. Do this by dividing each radius by the radius of beryllium. (*Hint:* The relative radius of magnesium would be 160 pm divided by 112 pm, or 1.4.) Record these values, rounded to the nearest tenth, in your data table.
5. Using a compass, draw a circle for each element with a radius that corresponds to the relative radius you calculated in Step 3. Use centimeters as your unit for the radius of each circle. **CAUTION:** *Do not push the sharp point of the compass against your skin.*
6. Label each model with the symbol of the element it represents.

Data Table

Element	Atomic Number	Radius (pm)*	Relative Radius
Be		112	1
Mg		160	
Ca		197	
Sr		215	
Ba		222	

**A picometer (pm) is one billionth of a millimeter.*

Analyze and Conclude

1. **Making Models** Based on your models, was your prediction in Step 1 correct? Explain.
2. **Graphing** Make a graph of the data in the second and third columns of the data table. Label the horizontal axis *Atomic Number.* Mark the divisions from 0 to 60. Then label the vertical axis *Radius* and mark its divisions from 0 to 300 picometers.
3. **Interpreting Data** What trend does the graph show? Is the relationship between the variables linear or nonlinear?
4. **Predicting** Predict where you would find the largest atom in any group, or family, of elements. What evidence would you need to tell if your prediction is correct?
5. **Communicating** Write a paragraph explaining why it is useful to draw a one- to two-centimeter model of an atom that has an actual radius of 100 to 200 picometers.

More to Explore

Look up the atomic masses for the Group 2 elements. Devise a plan to model their relative atomic masses using real-world objects.

Section 2 Ionic Bonds

CALIFORNIA Standards Focus

S 8.3.b Students know that compounds are formed by combining two or more different elements and that compounds have properties that are different from their constituent elements.

S 8.3.c Students know that atoms and molecules form solids by building up repeating patterns, such as the crystal structure of NaCl or long-chain polymers.

- How do ions form bonds?
- How are the formulas and names of ionic compounds written?
- What are the properties of ionic compounds?

Key Terms
- ion
- polyatomic ion
- ionic bond
- ionic compound
- chemical formula
- subscript
- crystal

Lab zone Standards Warm-Up

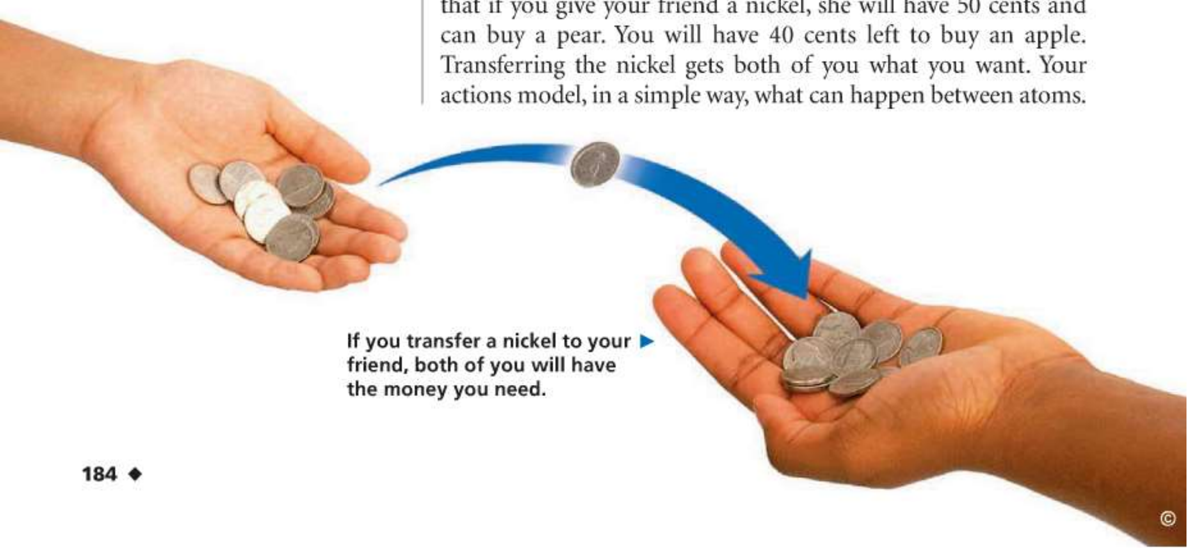
How Do Ions Form?

1. Place three pairs of checkers (three red and three black) on your desk. The red represent electrons and the black represent protons.
2. Place nine pairs of checkers (nine red and nine black) in a separate group on your desk.
3. Move a red checker from the smaller group to the larger group.
4. Count the number of positive charges (protons) and negative charges (electrons) in each group.
5. Now sort the checkers into a group of four pairs and a group of eight pairs. Repeat Steps 3 and 4, this time moving two red checkers from the smaller group to the larger group.

Think It Over

Inferring What was the total charge on each group before you moved the red checkers (electrons)? What was the charge on each group after you moved the checkers? Based on this activity, what do you think happens to the charge on an atom when it loses electrons? When it gains electrons?

You and a friend walk past a market that sells apples for 40 cents each and pears for 50 cents each. You have 45 cents and want an apple. Your friend also has 45 cents but wants a pear. You realize that if you give your friend a nickel, she will have 50 cents and can buy a pear. You will have 40 cents left to buy an apple. Transferring the nickel gets both of you what you want. Your actions model, in a simple way, what can happen between atoms.



If you transfer a nickel to your friend, both of you will have the money you need.

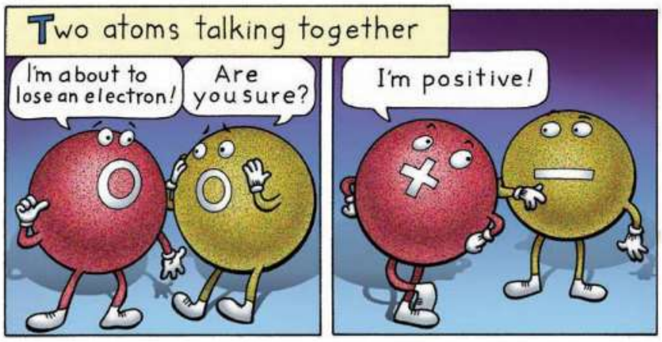


FIGURE 7
How Ions Form
When an atom loses one of its electrons, it becomes a positively charged ion. The atom that gains the electron becomes a negatively charged ion.

Ions

Atoms with five, six, or seven valence electrons usually become more stable when this number increases to eight. Likewise, most atoms with one, two, or three valence electrons can lose electrons and become more stable. When these two types of atoms combine, or bond, electrons are transferred from one type of atom to the other. The transfer makes both types of atoms more stable.

How Ions Form An **ion** (EYE ahn) is an atom or group of atoms that has an electric charge. When an atom loses an electron, it loses a negative charge and becomes a positive ion. When an atom gains an electron, it gains a negative charge and becomes a negative ion. Figure 8 lists some ions you will often see in this book. Use this table as a reference while you read this section and other chapters.

Polyatomic Ions Notice in Figure 8 that some ions are made of several atoms. For example, the ammonium ion is made of nitrogen and hydrogen atoms. Ions that are made of more than one atom are called **polyatomic ions** (pahl ee uh TAHM ik). The prefix *poly* means "many," so *polyatomic* means "many atoms." You can think of a polyatomic ion as a group of atoms that reacts as a unit. Like other ions, polyatomic ions have an overall positive or negative charge.

Reading Checkpoint **How does an ion with a charge of 2+ form?**

Ions and Their Charges

Name	Charge	Symbol or Formula
Lithium	1+	Li^{+}
Sodium	1+	Na^{+}
Potassium	1+	K^{+}
Ammonium	1+	NH_4^{+}
Calcium	2+	Ca^{2+}
Magnesium	2+	Mg^{2+}
Aluminum	3+	Al^{3+}
Fluoride	1–	F^{-}
Chloride	1–	Cl^{-}
Iodide	1–	I^{-}
Bicarbonate	1–	HCO_3^{-}
Nitrate	1–	NO_3^{-}
Oxide	2–	O^{2-}
Sulfide	2–	S^{2-}
Carbonate	2–	CO_3^{2-}
Sulfate	2–	SO_4^{2-}
Phosphate	3–	PO_4^{3-}

FIGURE 8
Ions are atoms that have lost or gained electrons. **Interpreting Tables** *How many electrons does a sulfur atom gain when it becomes a sulfide ion?*

Ionic Bonds Look at Figure 9 to see how sodium atoms and chlorine atoms combine to form sodium chloride (table salt). Notice that sodium has one valence electron and chlorine has seven valence electrons. When sodium's valence electron is transferred to chlorine, both atoms become ions. The sodium atom becomes a positive ion (Na^+). The chlorine atom becomes a negative ion (Cl^-).

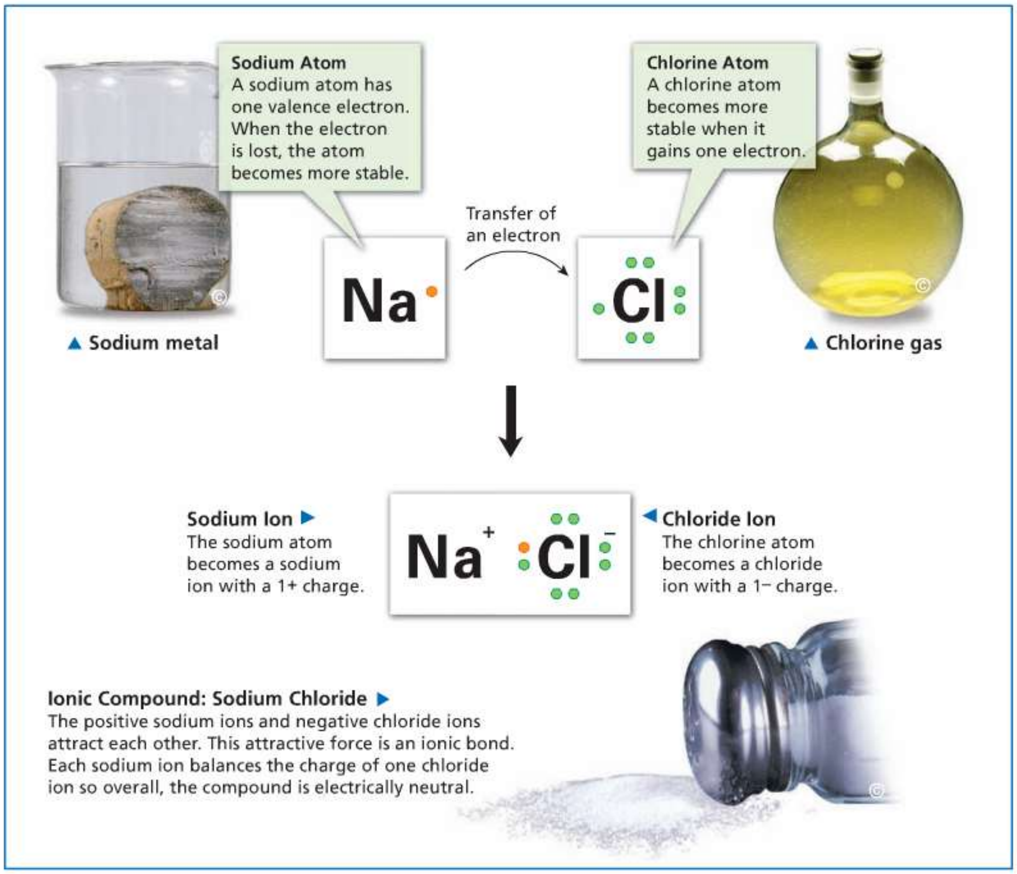
Because oppositely charged particles attract, the positive Na^+ ion and the negative Cl^- ion attract each other. An **ionic bond** is the attraction between two oppositely charged ions. **Ionic bonds form as a result of the attraction between positive and negative ions.** A compound that consists of positive and negative ions, such as sodium chloride, is called an **ionic compound.**

FIGURE 9

Formation of an Ionic Bond

Reactions occur easily between metals in Group 1 and nonmetals in Group 17. Follow the process below to see how an ionic bond forms between a sodium atom and a chlorine atom.

Relating Cause and Effect *Why is sodium chloride electrically neutral?*



Chemical Formulas and Names

Compounds can be represented by chemical formulas. A **chemical formula** is a combination of symbols that shows the ratio of elements in a compound. For example, the formula for magnesium chloride is $MgCl_2$. What does the formula tell you?

Formulas of Ionic Compounds From Figure 8 you know that the charge on the magnesium ion is 2+. **When ionic compounds form, the ions come together in a way that balances out the charges on the ions. The chemical formula for the compound reflects this balance.** Two chloride ions, each with a charge of 1– will balance the charge on the magnesium ion. That's why the formula of magnesium chloride is $MgCl_2$. The number "2" is a subscript. A **subscript** tells you the ratio of elements in the compound. For $MgCl_2$, the ratio of magnesium ions to chloride ions is 1 to 2.

If no subscript is written, the number 1 is understood. For example, the formula NaCl tells you that there is a 1 to 1 ratio of sodium ions to chloride ions. Formulas for compounds of polyatomic ions are written in a similar way. For example, calcium carbonate has the formula $CaCO_3$.

Naming Ionic Compounds Magnesium chloride, sodium bicarbonate, sodium oxide—where do these names come from? **For an ionic compound, the name of the positive ion comes first, followed by the name of the negative ion.** The name of the positive ion is usually the name of a metal. But, a few positive polyatomic ions exist, such as the ammonium ion (NH_4^+). If the negative ion is a single element, as you've already seen with sodium chloride, the end of its name changes to *-ide*. For example, MgO is named magnesium oxide. If the negative ion is polyatomic, its name usually ends in *-ate* or *-ite*, as in Figure 8. The compound NH_4NO_3, named ammonium nitrate, is a common fertilizer for gardens and crop plants.

Reading Checkpoint **What is the name of the ionic compound with the formula K_2S?**

Lab zone Skills Activity

Interpreting Data

Use the periodic table and Figure 8 to identify the charges of the ions in each ionic compound listed below. Then write the formula for each compound.

- sodium bromide
- lithium oxide
- magnesium sulfide
- aluminum fluoride
- potassium nitrate
- ammonium chloride

How did you know how many of each ion to write in the formula?

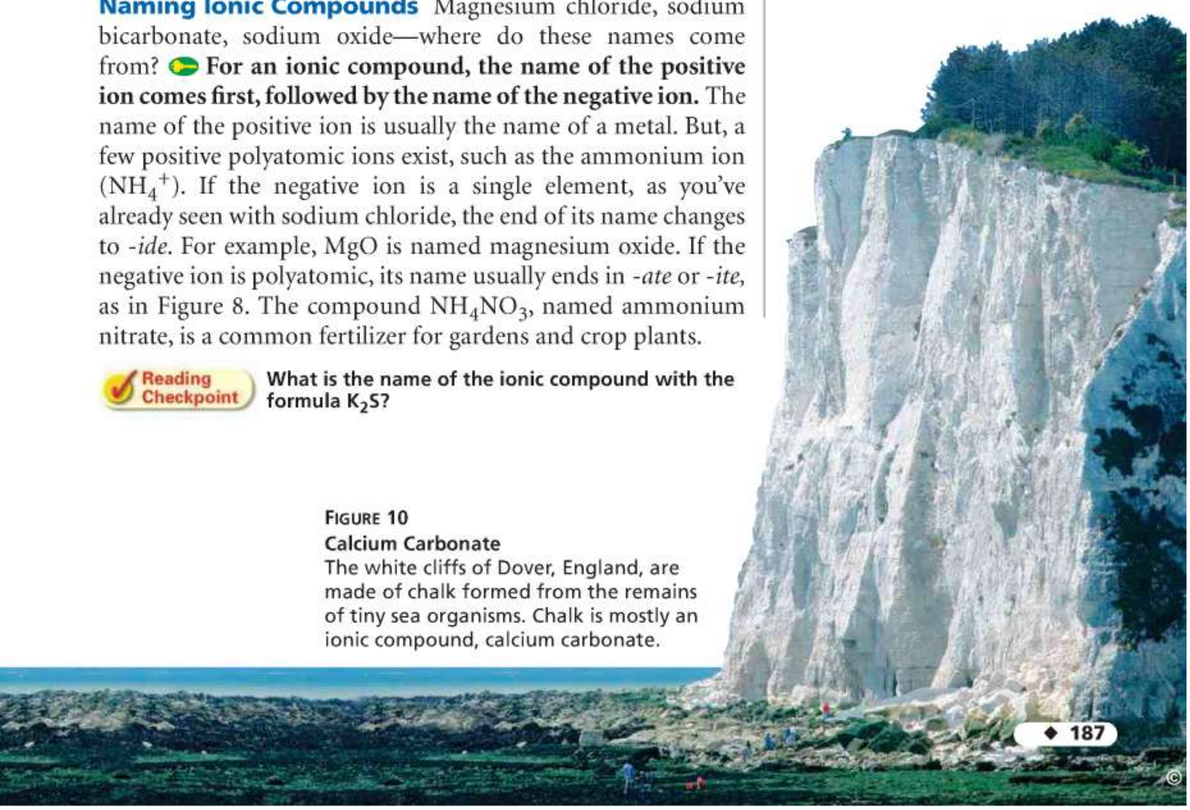


FIGURE 10
Calcium Carbonate
The white cliffs of Dover, England, are made of chalk formed from the remains of tiny sea organisms. Chalk is mostly an ionic compound, calcium carbonate.

Lab zone **Try This Activity**

Crystal Clear

Can you grow a salt crystal?

1. Add table salt to a jar containing about 200 mL of hot tap water and stir. Keep adding salt until no more dissolves and it settles out when you stop stirring.
2. Tie a large crystal of coarse salt into the middle of a piece of thread.
3. Tie one end of the thread to the middle of a pencil.
4. Suspend the other end of the thread in the solution by laying the pencil across the mouth of the jar. Do not allow the crystal to touch the solution.
5. Place the jar in a quiet, undisturbed area. Check the size of the crystal over the next few days.

Observing Does the salt crystal change size over time? What is its shape? What do you think is happening to the ions in the solution?

Properties of Ionic Compounds

Compounds have properties that are different from their constituent elements. You have already read about the properties of metals and nonmetals. But what about the properties of ionic compounds formed when metals and nonmetals react? **In general, ionic compounds are hard, brittle solids with high melting points. When melted or dissolved in water, they conduct electric current.**

Ionic Crystals Ionic compounds form solids by building up repeating patterns of ions. Figure 11 shows a chunk of a halite, or rock salt, which is how sodium chloride occurs naturally. Pieces of halite have sharp edges, corners, flat surfaces, and a cubic shape. Equal numbers of Na^+ and Cl^- ions in solid sodium chloride are attracted in an alternating pattern, as shown in the diagram. The ions form an orderly, three-dimensional arrangement called a **crystal.**

In an ionic compound, every ion is attracted to ions of opposite charge that surround it. The pattern formed by the ions remains the same no matter what the size of the crystal. In a single grain of salt, the crystal pattern extends for millions of ions in every direction. Many crystals of ionic compounds are hard and brittle, due to the strength of their ionic bonds and the attractions among all the ions.

High Melting Points When you heat a substance such as table salt, its energy increases. When the ions have enough energy to overcome the attractive forces between them, they break away from each other. The ionic crystal melts into a liquid. Because ionic bonds are strong, a lot of energy is needed to break them. As a result, ionic compounds have high melting points. For example, the melting point of table salt is 801°C.

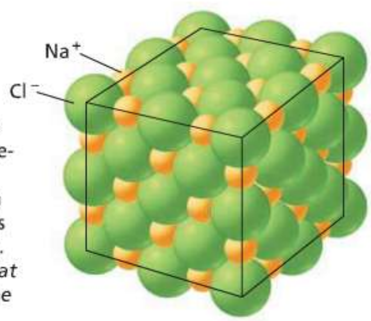


FIGURE 11
Ionic Crystals
The ions in ionic compounds are arranged in specific three-dimensional shapes called crystals. Some crystals have a cube shape like these crystals of halite, or sodium chloride.
Making Generalizations *What holds the ions together in the crystal?*

FIGURE 12
Ions in Solution
A solution of sodium chloride conducts electric current across the gap between the two black rods of a conductivity tester. As a result, the bulb lights up.

Electrical Conductivity Electric current is the flow of charged particles. When ionic crystals dissolve in water, the bonds between ions are broken. As a result, the ions are free to move about, and the solution conducts current. Likewise, after an ionic compound melts, the ions are able to move freely, and the liquid conducts current. In contrast, ionic compounds in solid form do not conduct current well. The ions in the solid crystal are tightly bound to each other and cannot move from place to place. If charged particles cannot move, there is no current.

Go Online
SCILINKS NSTA
For: Links on ionic compounds
Visit: www.SciLinks.org
Web Code: scn-1213

Reading Checkpoint **What is a crystal?**

Section 2 Assessment

S 8.3.b, 8.3.c, E-LA: Reading 8.2.0, Writing 8.2.1

Target Reading Skill Compare and Contrast Complete your table comparing the properties of sodium, chlorine, and sodium chloride. Expand the table by adding columns for melting point and conductivity. Then compare these properties.

Reviewing Key Concepts

HINT
1. a. Reviewing What is an ion?

HINT
b. Comparing and Contrasting Contrast sodium and chloride ions, including how they form. Write the symbol for each ion.

HINT
c. Relating Cause and Effect What holds the ions together in sodium chloride? Indicate the specific charges that are involved.

HINT
2. a. Identifying What information is given by the formula of an ionic compound?

HINT
b. Explaining The formula for sodium sulfide is Na_2S. Explain what this formula means.

c. Applying Concepts Write the formula for calcium chloride. Explain how you determined this formula. HINT

3. a. Listing List three properties of ionic compounds. HINT

b. Making Generalizations Relate each property that you listed to the characteristics of ionic bonds. HINT

Writing in Science

Firsthand Account Pretend that you are the size of an atom, and you are observing a reaction between a potassium atom and a fluorine atom. Write an account of the formation of an ionic bond as the atoms react. Tell what happens to the valence electrons on each atom and how each atom is changed by losing or gaining electrons.

Lab zone **Skills Lab**

Shedding Light on Ions

S 8.7.c, 8.9.c

Problem

What kinds of compounds produce ions in solution?

Skills Focus

controlling variables, interpreting data, inferring

Materials

- 2 dry cells, 1.5 V
- small light bulb and socket
- 4 lengths of wire with alligator clips on both ends
- 2 copper strips

(or conductivity probe)

- distilled water
- small beaker
- small plastic spoon
- sodium chloride
- graduated cylinder, 100-mL
- sucrose
- additional materials supplied by your teacher

Data Table	
Sample	Observations
Tap water	
Distilled water	
Sodium chloride	
Sodium chloride in water	

Procedure

1. Make a conductivity tester as described below or, if you are using a conductivity probe, see your teacher for instructions. Then make a data table in your notebook similar to the one above.
2. Pour about 50 mL of tap water into a small beaker. Place the copper strips in the beaker. Be sure the strips are not touching each other. Attach the alligator clip of the free end of one wire to a copper strip. Do the same with the other wire and the other copper strip. Record your observations.
3. Disconnect the wires from the copper strips. Take the strips out of the beaker, and pour out the tap water. Dry the inside of the beaker and the copper strips with a paper towel.
4. Pour 50 mL of distilled water into the beaker. Reconnect the conductivity tester and test the water as in Step 2. Keep the copper strips about the same distance apart as in Step 2. Record your observations.

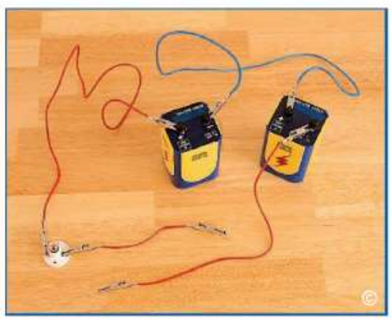
Making a Conductivity Tester

A. Use wire with alligator clips to connect the positive terminal of a dry cell to a lamp socket. **CAUTION:** *The bulb is fragile and can break.*

B. Similarly connect another wire between the negative terminal of the cell and the positive terminal of the second cell.

C. Connect one end of a third wire to the negative terminal of the second dry cell.

D. Connect one end of a fourth wire to the other terminal of the lamp socket.



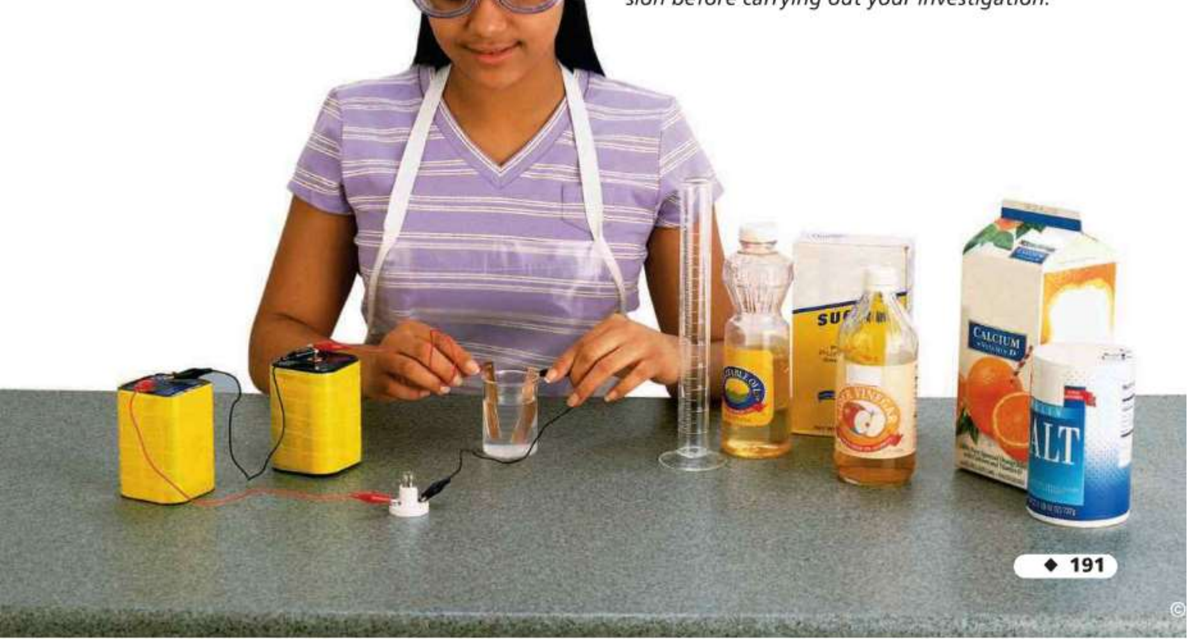
5. Use 3 spoonfuls of sodium chloride to make a small pile on a clean piece of paper. Dry off the copper strips of the conductivity tester and use it to test the conductivity of the sodium chloride. Record your observations.
6. Add 1 spoonful of sodium chloride to the distilled water in the beaker. Stir with the spoon until the salt dissolves. Repeat the conductivity test and record your observations.
7. Disconnect the conductivity tester and rinse the beaker, spoon, and copper strips with distilled water. Dry the beaker as in Step 3.
8. Test sucrose (table sugar) in the same ways that you tested sodium chloride in Steps 4 through 7. Test additional materials supplied by your teacher.
   - If the material is a solid, mix 1 spoonful of it with about 50 mL of distilled water and stir until the material dissolves. Test the resulting mixture.
   - If the substance is a liquid, simply pour about 50 mL into the beaker. Test it as you did the other mixtures.

Analyze and Conclude

1. **Designing Experiments** What were the variable parameters in your experiment? What were the controlled parameters?
2. **Controlling Variables** Why did you test both tap water and distilled water before testing the sodium chloride solution?
3. **Inferring** Could you have used tap water in your tests instead of distilled water? Explain.
4. **Drawing Conclusions** Based on your observations, add a column to your data table indicating whether each substance produced ions in solution.
5. **Inferring** How can you account for any observed differences in conductivity between dry and dissolved sodium chloride?
6. **Communicating** Based on your observations, decide whether or not you think sucrose is made up of ions. Explain your answer.

Design an Experiment

Design an experiment to test the effects of varying the spacing between the copper strips of the conductivity tester. *Obtain your teacher's permission before carrying out your investigation.*



Section 3

Covalent Bonds

CALIFORNIA Standards Focus

S 8.3.b Students know that compounds are formed by combining two or more different elements and that compounds have properties that are different from their constituent elements.

S 8.7.c Students know substances can be classified by their properties, including their melting temperature, density, hardness, and thermal and electrical conductivity.

- What holds covalently bonded atoms together?
- What are the properties of molecular compounds?
- How does unequal sharing of electrons affect molecules?

Key Terms

- covalent bond
- molecule
- double bond
- triple bond
- molecular compound
- polar bond
- nonpolar bond

Lab zone Standards Warm-Up

Can Water and Oil Mix?

1. Pour water into a small jar that has a tight-fitting lid until the jar is about a third full.
2. Add an equal amount of vegetable oil to the jar. Cover the jar tightly.
3. Shake the jar vigorously for 20 seconds. Observe the contents.
4. Allow the jar to sit undisturbed for 1 minute. Observe again.
5. Remove the top and add 3 drops of liquid detergent. Cover the jar and repeat Steps 3 and 4.

Think It Over

Forming Operational Definitions Based on your observations, write an operational definition of *detergent.* How might your observations relate to chemical bonds in the detergent, oil, and water molecules?

Uh oh, you have a big project due in English class next week! You need to write a story and illustrate it with colorful posters. Art has always been your best subject, but writing takes more effort. Luckily, you're working with a partner who writes well but doesn't feel confident in art. If you each contribute your skills, together you can produce a high-quality finished project.

FIGURE 13
Sharing Skills
One student is a skilled artist, while the other is a skilled writer. By pooling their skills, the students can complete their project.

How Covalent Bonds Form

Just as you and your friend can work together by sharing your talents, atoms can become more stable by sharing electrons. The chemical bond formed when two atoms share electrons is called a **covalent bond.** Covalent bonds usually form between atoms of nonmetals. In contrast, ionic bonds usually form when a metal combines with a nonmetal.

Electron Sharing Nonmetals can bond to other nonmetals by sharing electrons. So can hydrogen. Most nonmetals can even bond with another atom of the same element. Figure 14 illustrates how two fluorine atoms can react by sharing a pair of electrons. By sharing electrons, each fluorine atom has a stable set of eight. **The force that holds atoms together in a covalent bond is the attraction of each atom's nucleus for the shared pair of electrons.** The two bonded fluorine atoms form a molecule. A **molecule** is a neutral group of atoms joined by covalent bonds.

How Many Bonds? Look at the electron dot diagrams in Figure 15. Count the valence electrons around each atom that reacts. Hydrogen has one valence electron. Oxygen has six. Nitrogen has five. The number of covalent bonds that a nonmetal atom can form equals the number of electrons needed to make a total of eight. For example, oxygen has six valence electrons, so it can form two covalent bonds. In a water molecule, oxygen forms one covalent bond with each of two hydrogen atoms. As a result, the oxygen atom has a stable set of eight valence electrons. Each hydrogen atom can form one bond because it needs only a total of two electrons to be stable.

FIGURE 14
Sharing Electrons
By sharing electrons in a covalent bond, each fluorine atom has a stable set of eight valence electrons.

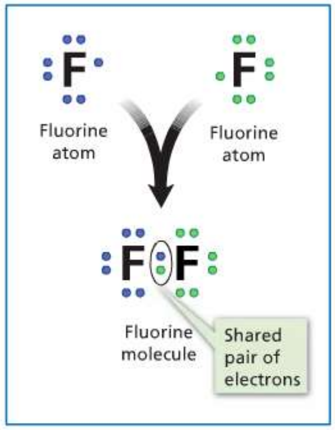
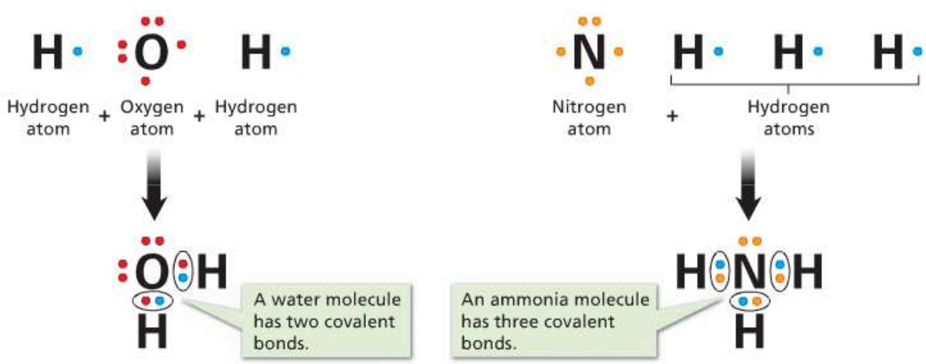


FIGURE 15
Covalent Bonds
The oxygen atom in water and the nitrogen atom in ammonia each have eight valence electrons as a result of forming covalent bonds with hydrogen atoms.
Interpreting Diagrams *How many covalent bonds can a nitrogen atom form?*



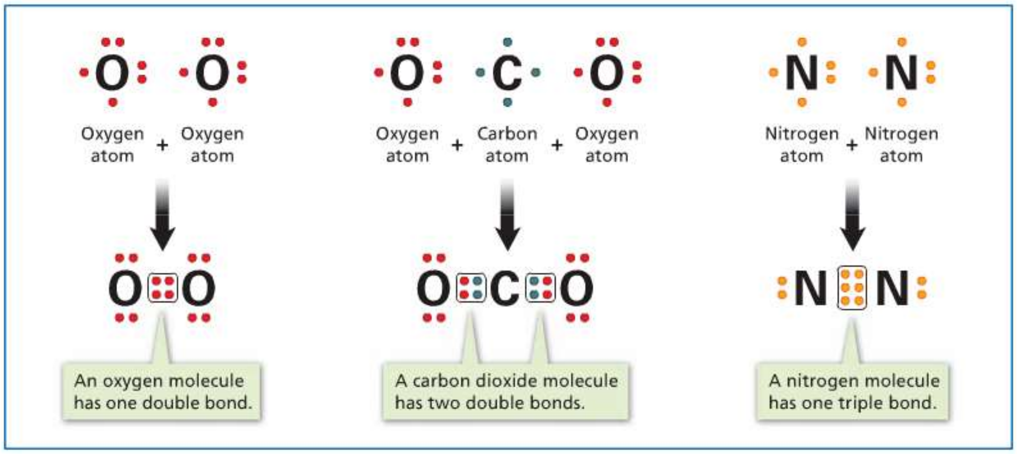


FIGURE 16
Double and Triple Bonds
Double and triple bonds can form when atoms share more than one pair of electrons.
Interpreting Diagrams *In a nitrogen molecule, how many electrons does each nitrogen atom share with the other?*

Double Bonds and Triple Bonds Look at the diagram of the oxygen molecule (O_2) in Figure 16. What do you see that's different? This time the two atoms share two pairs of electrons, forming a **double bond.** In a carbon dioxide molecule (CO_2), carbon forms a double bond with each of two oxygen atoms. Elements such as nitrogen and carbon can form **triple bonds** in which their atoms share three pairs of electrons.

Reading Checkpoint **What is the difference between a double bond and a triple bond?**

Molecular Compounds

A **molecular compound** is a compound that is composed of molecules. The molecules of a molecular compound contain atoms that are covalently bonded. **Compared to ionic compounds, molecular compounds generally have lower melting points and boiling points. And, unlike ionic compounds, molecular compounds do not conduct electric current when melted or dissolved in water.**

Low Melting Points and Boiling Points Study the table in the Analyzing Data box on the next page. It lists the melting points and boiling points for a few molecular compounds and ionic compounds. In molecular solids, forces hold the molecules close to one another. But, the forces between molecules are much weaker than the forces between ions in an ionic solid. Compared with ionic solids, less heat must be added to molecular solids to separate the molecules and change the solid to a liquid. That is why most familiar compounds that are liquids or gases at room temperature are molecular compounds.

Reviewing Math: Statistics, Data Analysis, and Probability 7.1.1

Math Analyzing Data

Comparing Molecular and Ionic Compounds

The table compares the melting points and boiling points of a few molecular compounds and ionic compounds. Use the table to answer the following questions.

1. **Graphing** Create a bar graph of just the melting points of these compounds. Arrange the bars in order of increasing melting point. The *y*-axis should start at −200°C and go to 900°C.
2. **Interpreting Data** Describe what your graph reveals about the melting points of molecular compounds compared to those of ionic compounds.
3. **Inferring** How can you account for the differences in melting points between molecular compounds and ionic compounds?
4. **Interpreting Data** How do the boiling points of the molecular and ionic compounds compare?

Melting Points and Boiling Points of Molecular and Ionic Compounds

Substance	Formula	Melting Point (°C)	Boiling Point (°C)
Methane	CH_4	–182.4	–161.5
Rubbing alcohol	C_3H_8O	–89.5	82.4
Water	H_2O	0	100
Zinc chloride	$ZnCl_2$	290	732
Magnesium chloride	$MgCl_2$	714	1,412
Sodium chloride	NaCl	800.7	1,465

Molecular compound   Ionic compound

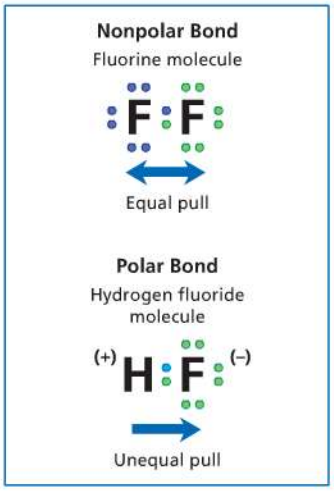
5. **Predicting** Ammonia's melting point is −78°C and its boiling point is −34°C. Is ammonia a molecular compound or an ionic compound? Explain.

Poor Conductivity Most molecular compounds do not conduct electric current. No charged particles are available to move, so there is no current. Materials such as plastic and rubber are used to insulate wires because these materials are composed of molecular substances. Even as liquids, molecular compounds are poor conductors. Pure water, for example, does not conduct electric current. Neither does table sugar or alcohol when they are dissolved in pure water.

Unequal Sharing of Electrons

Have you ever played tug of war? If you have, you know that if both teams pull with equal force, the contest is a tie. But what if the teams pull with unequal forces? Then the rope moves toward the side of the stronger team. The same is true of electrons in a covalent bond. Atoms of some elements pull more strongly on shared electrons than do atoms of other elements. As a result, the electrons are shared unequally. **Unequal sharing of electrons causes the bonded atoms to have slight electrical charges.**

Figure 17
Nonpolar and Polar Bonds
Fluorine forms a nonpolar bond with another fluorine atom. In hydrogen fluoride, fluorine attracts electrons more strongly than hydrogen does, so the bond formed is polar.



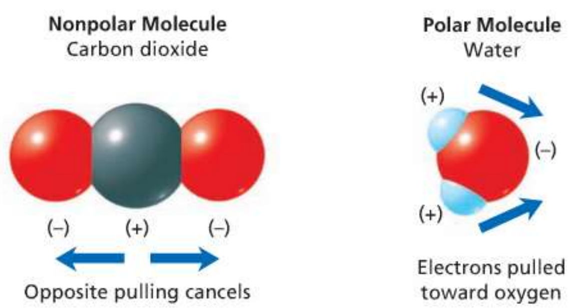
Polar Bonds and Nonpolar Bonds The unequal sharing of electrons is enough to make the atom with the stronger pull slightly negative and the atom with the weaker pull slightly positive. A covalent bond in which electrons are shared unequally is called a **polar bond.** Of course, if two atoms pull equally on the electrons, neither atom becomes charged. A covalent bond in which electrons are shared equally is a **nonpolar bond.** Compare the bond in fluorine (F_2) with the bond in hydrogen fluoride (HF) in Figure 17.

Polar Bonds in Molecules A molecule is considered polar if it has a positively charged end opposite a negatively charged end. For example, hydrogen fluoride is a polar molecule. However, not all molecules containing polar bonds are polar. For example, in carbon dioxide, the oxygen atoms attract electrons much more strongly than carbon does. So, the bonds between the oxygen and carbon atoms are polar. But, as you can see in Figure 18, a carbon dioxide molecule has a straight-line shape. As a result, the two oxygen atoms pull with equal strength in opposite directions. The attractions cancel out, making the molecule nonpolar.

In contrast, a water molecule, with its two polar bonds, is itself polar. A water molecule has two hydrogen atoms at one end and an oxygen atom at the other end. The oxygen atom attracts electrons more strongly than do the hydrogen atoms. As a result, the oxygen end has a slight negative charge and the hydrogen end has a slight positive charge.

Attractions Among Molecules If you could shrink small enough to move among a bunch of water molecules, what would you find? The negatively charged oxygen ends of the polar water molecules attract the positively charged hydrogen ends of nearby water molecules. These attractions pull water molecules toward each other. In contrast, there is little attraction among nonpolar molecules, such as carbon dioxide molecules.

Figure 18
Nonpolar and Polar Molecules
A carbon dioxide molecule is a nonpolar molecule because of its straight-line shape. In contrast, a water molecule is a polar molecule because of its bent shape.
Interpreting Diagrams *What do the arrows in the diagram show?*



The properties of polar and nonpolar compounds differ because of differences in attractions between their molecules. For example, water and vegetable oil don't mix. The molecules in vegetable oil are nonpolar, and nonpolar molecules have little attraction for polar water molecules. On the other hand, the water molecules are attracted more strongly to one another than to the molecules of oil. Thus, water stays with water, and oil stays with oil.

If you did the Standards Warm-Up activity, you found that adding detergent helped oil and water to mix. This is because one end of a detergent molecule has nonpolar covalent bonds. The other end includes an ionic bond. The detergent's nonpolar end mixes easily with the oil. Meanwhile, the charged ionic end is attracted to polar water molecules, so the detergent dissolves in water.

Reading Checkpoint **Why is water (H_2O) a polar molecule but a fluorine molecule (F_2) is not?**

FIGURE 19
Getting Out the Dirt
Most laundry dirt is oily or greasy. Detergents can mix with both oil and water, so when the wash water goes down the drain, the soap and dirt go with it.

Section 3 Assessment

S 8.3.b, 8.7.c, E-LA: Reading 8.2.0, Math: 7PS1.1

Target Reading Skill **Compare and Contrast** Create a table comparing the properties of molecular compounds and ionic compounds.

Reviewing Key Concepts

1. a. **Identifying** What is the attraction that holds two covalently bonded atoms together? HINT
   b. **Inferring** A carbon atom can form four covalent bonds. How many valence electrons does it have? HINT
   c. **Interpreting Diagrams** What is a double bond? Use Figure 16 to explain how a carbon dioxide molecule has a stable set of eight valence electrons for each atom. HINT
2. a. **Reviewing** How are the properties of molecular compounds different from those of ionic compounds? HINT
   b. **Relating Cause and Effect** Why are most molecular compounds poor conductors? HINT
3. a. **Reviewing** How do some atoms in covalent bonds become slightly negative or slightly positive? What type of covalent bonds do these atoms form? HINT
   b. **Comparing and Contrasting** Both carbon dioxide molecules and water molecules have polar bonds. Why then is carbon dioxide a nonpolar molecule while water is a polar molecule? HINT
   c. **Predicting** Predict whether carbon dioxide or water would have a higher boiling point. Explain your prediction in terms of the attractions between molecules. HINT

Lab zone At-Home Activity

Laundry Chemistry Demonstrate the action of soaps and detergents to your family. Pour some vegetable oil on a clean cloth and show how a detergent solution can wash the oil away better than water alone can. Explain to your family the features of soap and detergent molecules in terms of their chemical bonds.

Bonding in Metals

CALIFORNIA Standards Focus

S 8.7.c Students know substances can be classified by their properties, including their melting temperature, density, hardness, and thermal and electrical conductivity.

- How do the properties of metals and alloys compare?
- How do metal atoms combine?
- How does metallic bonding result in useful properties of metals?

Key Terms

- alloy
- metallic bond

Lab zone Standards Warm-Up

Are They "Steel" the Same?

1. Wrap a stainless steel bolt, a wire nail (high-carbon steel), and a cut nail (low-carbon steel) together in a paper towel.
2. Place the towel in a plastic bag. Add about 250 mL of salt water and seal the bag.
3. After one or two days, remove the nails and bolt. Note any changes in the metals.

Think It Over

Developing Hypotheses What happened to the three types of steel? Which one changed the most, and which one changed the least? What do you think accounts for the difference?

Why would you choose metal to cover the complex shape of the building in Figure 20? You couldn't cover the building with brittle, crumbly nonmetals such as sulfur or silicon. What physical properties make metals ideal materials for making furniture, musical instruments, electrical wire, pots and pans, eating utensils, and strong beams for buildings? Why do metals have these physical properties?

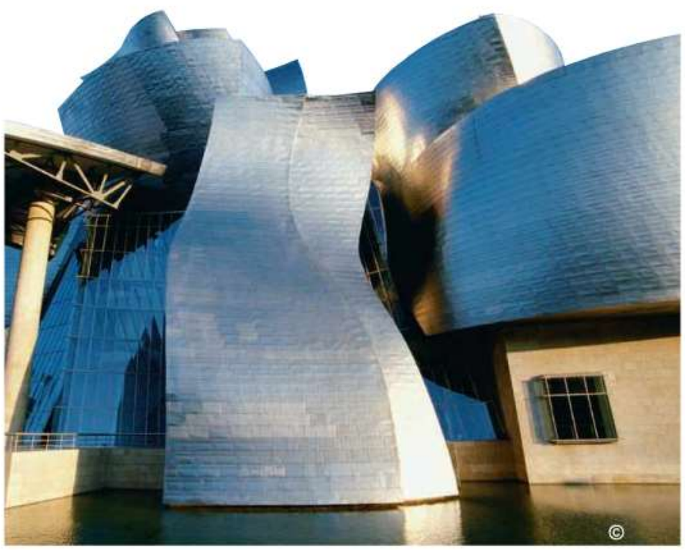


FIGURE 20
Metal in Architecture
The Guggenheim Museum in Bilbao, Spain, makes dramatic use of some properties of metals. The museum's shiny outer "skin" is made of the lightweight metal titanium, which can be pressed into large, thin, flexible sheets.

FIGURE 21
Brass Trumpet
Brass is an alloy of the elements copper and zinc. **Observing** *What are some metallic properties of brass that you can see here?*

Metals and Alloys

You know a piece of metal when you see it. It's usually hard, dense, and shiny. At room temperature, most metals are solids. They can be hammered or drawn out into thin wire. Electronics such as stereos, computers, and MP3 players have metal parts because metals conduct electric current.

Yet, very few of the "metals" you use every day consist of just one element. Instead, most of the metallic objects you see and use are made of alloys. An **alloy** is a mixture made of two or more elements, at least one of which is a metal. **Alloys are generally stronger and less reactive than the pure metals from which they are made.**

Physical Properties The properties of an alloy can differ greatly from those of its individual elements. But depending on how they are mixed, alloys also retain many of the physical properties of metals. For example, pure gold is shiny, but it is soft and easily bent. For that reason, gold jewelry and coins are made of an alloy of gold mixed with a harder element, such as copper or silver. These gold alloys are much harder than pure gold but still retain their beauty and shine. Even after thousands of years, objects made of gold alloys still look exactly the same as when they were first made.

Chemical Properties Iron is strong metal that you might think would be good for making tools. However, iron objects rust when they are exposed to air and water. For this reason, iron is often alloyed with one or more other elements to make steel. Tools made of steel are much stronger than iron and resist rust much better. For example, forks and spoons made of stainless steel can be washed over and over again without rusting. That's because stainless steel—an alloy of iron, carbon, nickel, and chromium—does not react with air and water as iron does.

Reading Checkpoint **Why is most jewelry made of gold alloys rather than pure gold?**

FIGURE 22
Gold and Steel
This pipe wrench is made of steel. The necklace is made of gold alloys.

Lab zone Try This **Activity**

What Do Metals Do?

1. Your teacher will give you pieces of different metals. Examine each metal and try changing its shape by bending, stretching, folding, or any other action you can think of.
2. Write down the properties that are common to these metals. Write down the properties that are different.
3. What properties make each metal suitable for its intended use?

Inferring What properties must aluminum have in order to be made into foil?

Metallic Bonding

The properties of solid metals and their alloys can be explained by the structure of metal atoms and the bonding between those atoms. Recall that most metals have 1, 2, or 3 valence electrons. When metal atoms combine chemically with atoms of other elements, they usually lose valence electrons, becoming positively charged metal ions. Metals lose electrons easily because their valence electrons are not strongly held.

Metal atoms combine in regular patterns in which the valence electrons are free to move from atom to atom. Most metals are crystalline solids. Within each crystal, the metal atoms exist as closely packed, positively charged ions. The valence electrons drift among the ions. Each metal ion is held in place by a **metallic bond**—an attraction between a positive metal ion and the many electrons surrounding it. Figure 23 illustrates the metallic bonds that hold together aluminum foil. The positively charged metal ions are embedded in a "sea" of valence electrons. The more valence electrons an atom can add to the "sea," the stronger the metallic bonds will be.

Reading Checkpoint **What is a metallic bond?**

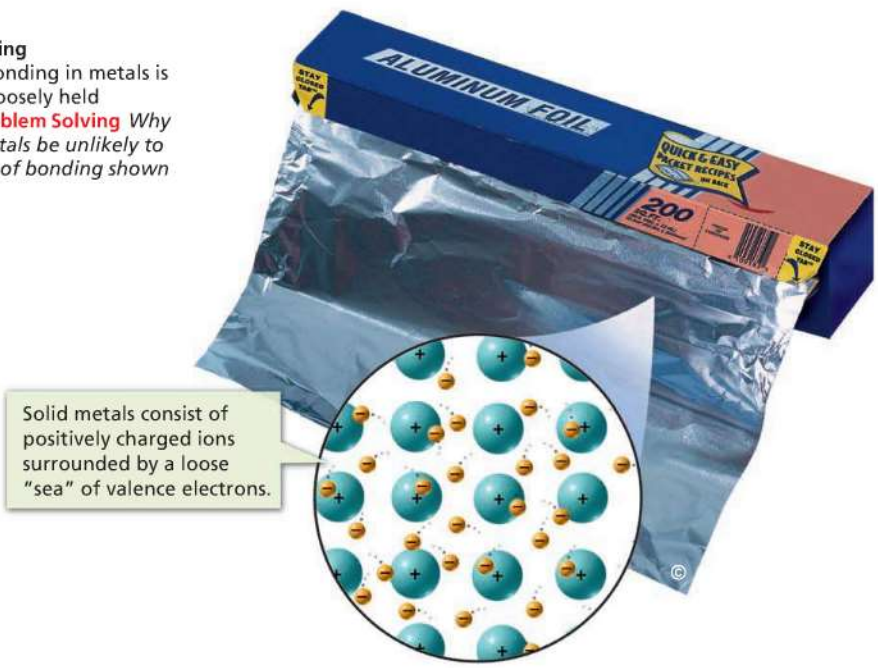


FIGURE 23
Metallic Bonding
The type of bonding in metals is the result of loosely held electrons. **Problem Solving** *Why would nonmetals be unlikely to have the type of bonding shown here?*

Metallic Properties

Suppose that you placed one hand on an unheated aluminum pan and the other hand on a wooden tabletop. The aluminum pan would feel cooler than the tabletop even though both are at the same temperature. You feel the difference because aluminum conducts heat away from your hand much faster than wood does. Metal fins called a "heat sink" are used inside many electronics to cool their insides. However, a metal's ability to conduct heat is even more useful if the metal can be bent or hammered into a useful shape.

The "sea of electrons" model of metallic bonding helps explain the malleability, ductility, luster, high electrical conductivity, and high thermal conductivity of solid metals. Each of these properties is related to the behavior of valence electrons in metal atoms.

Malleability and Ductility Most metals are flexible and can be reshaped easily. They can be stretched, pushed, or compressed into different shapes without breaking. Metals act this way because the positive ions are attracted to the loose electrons all around them rather than to other metal ions. These ions can be made to change position, as shown in Figure 24. However, the metallic bonds between the ion and the surrounding electrons keep the metal from breaking.

Because the metal ions move easily, metals are ductile, which means that they can be bent easily and pulled into thin strands or wires. Metals are also malleable—able to be rolled into thin sheets, as in aluminum foil, or beaten into complex shapes.

Go Online SCiLINKS NSTA
For: Links on metallic bonding
Visit: www.SciLinks.org
Web Code: scn-1215

FIGURE 24
Flexibility of Metals
Most metals can be reshaped because their ions can be pushed easily out of position.

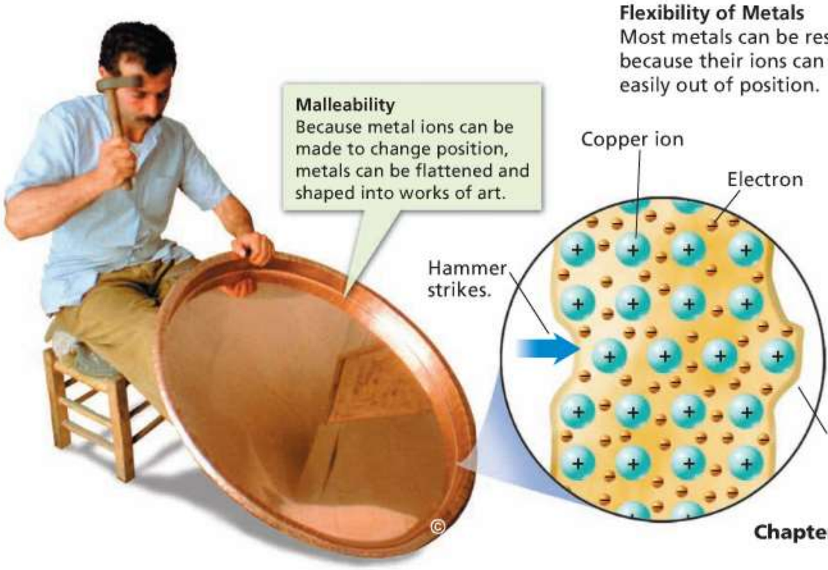
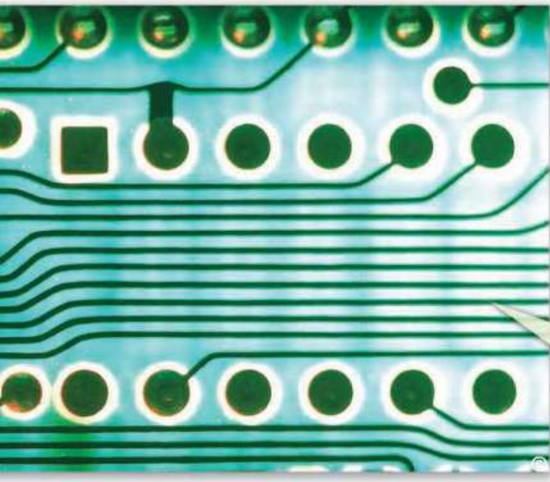


FIGURE 25
Luster and Conductivity of Metals

The unique properties of metals result from the ability of their electrons to move about freely.

Luster
Gold in an astronaut's face shield reflects sunlight, protecting the wearer's eyes.



Electrical Conductivity
Metal strips on a circuit board conduct electric current throughout the circuit.

Luster Polished metals have a high luster—that is, they are shiny and reflective. A metal's luster is due to its valence electrons. When light strikes these electrons, they absorb the light and then give it off again. This property makes metals useful for making products as varied as mirrors, buildings, jewelry, and astronaut helmets.

Electrical Conductivity You may recall that when charged particles are free to move an electric current is possible. Metals conduct current easily because the electrons in a metal can move freely among the atoms. When connected to a device such as a battery, electric current flows into the metal at one point and out at another point.

Thermal Conductivity Recall that thermal energy flows from warmer matter to cooler matter. When this happens, the greater motion of the particles in the warmer parts of the material is passed along to the particles in the cooler parts. This transfer of thermal energy is known as heat. Metals conduct heat easily because of the valence electrons' freedom of motion within a metal or metal alloy.

Reading Checkpoint **Why do metals conduct electric current easily?**

Section 4 Assessment

S 8.7.c, E-LA: Reading 8.1.0, Writing 8.2.6

Vocabulary Skill **High-Use Academic Words** Use the word *conduct* in a sentence that explains one of the physical properties of metals.

Reviewing Key Concepts

1. a. **Defining** What is an alloy?
   b. **Reviewing** From what pure metals is stainless steel made?
   c. **Comparing and Contrasting** Compare and contrast the general properties of alloys and pure metals.
2. a. **Describing** What is a metallic bond?
   b. **Relating Cause and Effect** Explain how metal atoms form metallic bonds. What role do the valence electrons play?
   c. **Comparing and Contrasting** Review what you learned earlier about ionic bonds. How does a metallic bond differ from an ionic bond?
3. a. **Listing** Name four properties of metals. What accounts for these properties?
   b. **Describing** In a light bulb, a thin tungsten wire filament that is wound in a coil conducts electric current. Describe two properties of the metal tungsten that make it good material for the filament.
   c. **Applying Concepts** Why is it safer to use a nonmetal mixing spoon when cooking something on a stove?

Writing in Science

Product Label Choose a familiar metal object and create a "product label" for it. Your label should describe at least two of the metal's properties and explain why it exhibits those properties. You should include illustrations on your label as well.

Chapter 5

Study Guide

The BIG Idea Atoms of different elements combine to form compounds by gaining, losing, or sharing electrons.

1 Atoms, Bonding, and the Periodic Table

Key Concepts

S 8.3.f

- The number of valence electrons in an atom of an element determines many properties of that element, including the ways in which the atom can bond with other atoms.
- The periodic table reveals the underlying atomic structure of atoms, including the arrangement of the electrons.

Key Terms

valence electron
electron dot diagram
chemical bond

2 Ionic Bonds

Key Concepts

S 8.3.b, 8.3.c

- Ionic bonds form as a result of the attraction between positive and negative ions.
- When ionic compounds form, the charges on the ions balance out.
- For an ionic compound, the name of the positive ion comes first, followed by the name of the negative ion.
- In general, ionic compounds are hard, brittle crystals that have high melting points and conduct electricity when dissolved in water.

Key Terms

ion
polyatomic ion
ionic bond
ionic compound
chemical formula
subscript
crystal

3 Covalent Bonds

Key Concepts

S 8.3.b, 8.7.c

- The force that holds atoms together in a covalent bond is the attraction of each atom's nucleus for the shared pair of electrons.
- Molecular compounds have low melting and boiling points and do not conduct electric current.
- Unequal sharing of electrons causes bonded atoms to have slight electrical charges.

Key Terms

covalent bond
molecule
double bond
triple bond
molecular compound
polar bond
nonpolar bond

4 Bonding in Metals

Key Concepts

S 8.7.c

- Alloys are generally stronger and less reactive than the pure metals from which they are made.
- Metal atoms combine in regular patterns in which the valence electrons are free to move from atom to atom.
- The "sea of electrons" model of metallic bonding helps explain the malleability, ductility, luster, high electrical conductivity, and high thermal conductivity of metals.

Key Terms

alloy
metallic bond

Review and Assessment

Go Online
PHSchool.com
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Web Code: cxa-2050

Target Reading Skill

Compare and Contrast Complete the compare–contrast table on Types of Chemical Bonds.

Type of Chemical Bond	How Bonds Forms	Charge on Bonded Atoms	Example
Ionic bond	a. ?	Yes; positive and negative	b. ?
Polar bond	Unequal sharing of electrons	c. ?	d. ?
Nonpolar bond	e. ?	f. ?	O_2 molecule
Metallic bond	g. ?	yes; positive	h. ?

Reviewing Key Terms

Choose the letter of the best answer.

HINT 1. Valence electrons in an atom are those that are
- a. held most loosely.
- b. of the lowest energy level.
- c. always easily lost.
- d. never easily lost.

HINT 2. An electron dot diagram shows an atom's number of
- a. protons.
- b. electrons.
- c. valence electrons.
- d. chemical bonds.

HINT 3. When an atom loses or gains electrons, it becomes a(n)
- a. ion.
- b. formula.
- c. crystal.
- d. subscript.

HINT 4. A covalent bond in which electrons are shared unequally is a
- a. double bond.
- b. triple bond.
- c. polar bond.
- d. nonpolar bond.

HINT 5. The metal atoms in stainless steel are held together by
- a. ionic bonds.
- b. polar bonds.
- c. covalent bonds.
- d. metallic bonds.

Complete the following sentences so that your answers clearly explain the key terms.

6. When atoms react, they form a **chemical bond,** which is ________ . HINT
7. **Polyatomic ions** such as ammonium ions (NH_4^-) and nitrate ions (NO_3^-) are ions that consist of ________ . HINT
8. Magnesium chloride is an example of an **ionic compound,** which means a compound composed of ________ . HINT
9. The formulas N_2, H_2O, and CO_2 all represent **molecules,** which are defined as ________ . HINT
10. Pure metals tend to be weaker and more reactive than an **alloy,** which is a ________ . HINT

Writing in Science

Comparing and Contrasting Go to your local grocery store and observe how the products on the shelves are organized. Write a paragraph comparing how foods are organized in a grocery store and how elements are organized in the periodic table.

Video Assessment
Discovery Channel School
Atoms and Bonding

Review and Assessment

Checking Concepts

11. Which element is less reactive, an element whose atoms have seven valence electrons or an element whose atoms have eight valence electrons? Explain.

12. Why do ionic compounds generally have high melting points?

13. The formula of sulfuric acid is H_2SO_4. How many atoms of hydrogen, sulfur, and oxygen are in one molecule of sulfuric acid?

14. How is the formation of an ionic bond different from the formation of a covalent bond?

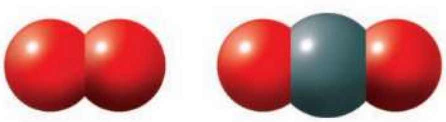
15. Why is the covalent bond between two atoms of the same element a nonpolar bond?

16. Explain how metallic bonding causes metals to conduct electric current.

Thinking Critically

17. Making Generalizations What information does the organization of the periodic table tell you about how reactive an element may be?

18. Classifying Classify each molecule below as either a polar molecule or a nonpolar molecule. Explain your reasoning.



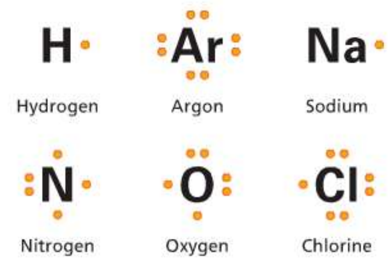
Oxygen **Carbon dioxide**

19. Relating Cause and Effect Many molecular compounds with small molecules are gases at room temperature. Water, however, is a liquid. Use what you know about polar and nonpolar molecules to explain this difference. (*Hint:* Molecules of a gas are much farther apart than molecules of a liquid.)

20. Applying Concepts Why does a metal horseshoe bend but not break when a blacksmith pounds it into shape?

Applying Skills

Use the electron dot diagrams below to answer Questions 21–25.



21. Predicting When nitrogen and hydrogen combine, what will be the ratio of hydrogen atoms to nitrogen atoms in a molecule of the resulting compound? Explain.

22. Inferring Which of these elements can become stable by losing one electron? Explain.

23. Drawing Conclusions Which of these elements is least likely to react with other elements? Explain.

24. Interpreting Diagrams Which of these elements would react with two atoms of sodium to form an ionic compound? Explain.

25. Classifying What type of bond forms when two atoms of nitrogen join to form a nitrogen molecule? When two atoms of oxygen join to form an oxygen molecule?

Lab zone Standards Investigation

Performance Assessment Present your models to the class, telling what the parts of each model represent. Explain why you chose particular items to model the atoms and chemical bonds. Which kind of bonds were easier to show? Why? What more would you like to know about bonding that could help improve your models?

CALIFORNIA Standards Practice

Success Tracker™
Online at PHSchool.com

Choose the letter of the best answer.

1. What is the atomic number of calcium?

A 6 **B** 20
C 40.08 **D** 48 **S 8.3.f**

Use the electron dot diagrams below to answer Questions 2–5.

Electron Dot Diagrams

2. Which element is the most likely to lose two electrons and form an ion with a charge of 2+?

A potassium (K)
B oxygen (O)
C magnesium (Mg)
D aluminum (Al) **S 8.3.b**

3. Oxygen has 6 valence electrons, as indicated by the 6 dots around the letter symbol "O." Based on this information, how many covalent bonds could an oxygen atom form?

A six **B** three
C two **D** none **S 8.3.b**

4. If a reaction occurs between potassium (K) and oxygen (O), what will be the ratio of potassium ions to oxide ions in the resulting compound, potassium oxide?

A 1 : 1 **B** 1 : 2
C 2 : 1 **D** 2 : 2 **S 8.3.b**

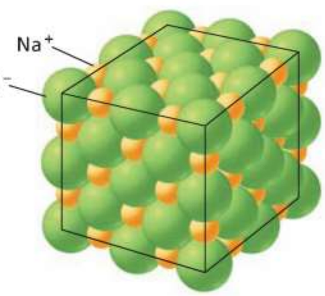
5. The element boron (B) is directly above aluminum (Al) on the periodic table. Which statement about boron is true?

A Boron is in the same period as aluminum and has two valence electrons.
B Boron is in the same group as aluminum and has two valence electrons.
C Boron is in the same period as aluminum and has three valence electrons.
D Boron is in the same group as aluminum and has three valence electrons. **S 8.3.f**

6. An ice cube (solid H_2O) and a scoop of table salt (NaCl) are left outside on a warm, sunny day. Which best explains why the ice cube melts and the salt does not?

A The attractive forces between molecules of H_2O are much weaker than those between ions in NaCl.
B NaCl can dissolve in H_2O.
C The mass of the H_2O was less than the mass of the NaCl.
D NaCl is white and H_2O is colorless. **S 8.3.c**

The diagram below shows the crystal structure of sodium chloride. Use the diagram to answer Question 7.



7. In a crystal of sodium chloride, each sodium ion is attracted to the

A other sodium ions surrounding it.
B chloride ions surrounding it.
C neutral sodium atoms surrounding it.
D neutral chlorine atoms surrounding it. **S 8.3.c**

Apply the BIG Idea

8. Use the periodic table to find the number of valence electrons for potassium (K), calcium (Ca), aluminum (Al), oxygen (O), and iodine (I). Then write the formulas for the following compounds: potassium iodide, calcium oxide, aluminum iodide, and potassium oxide. **S 8.3.b, 8.3.f**