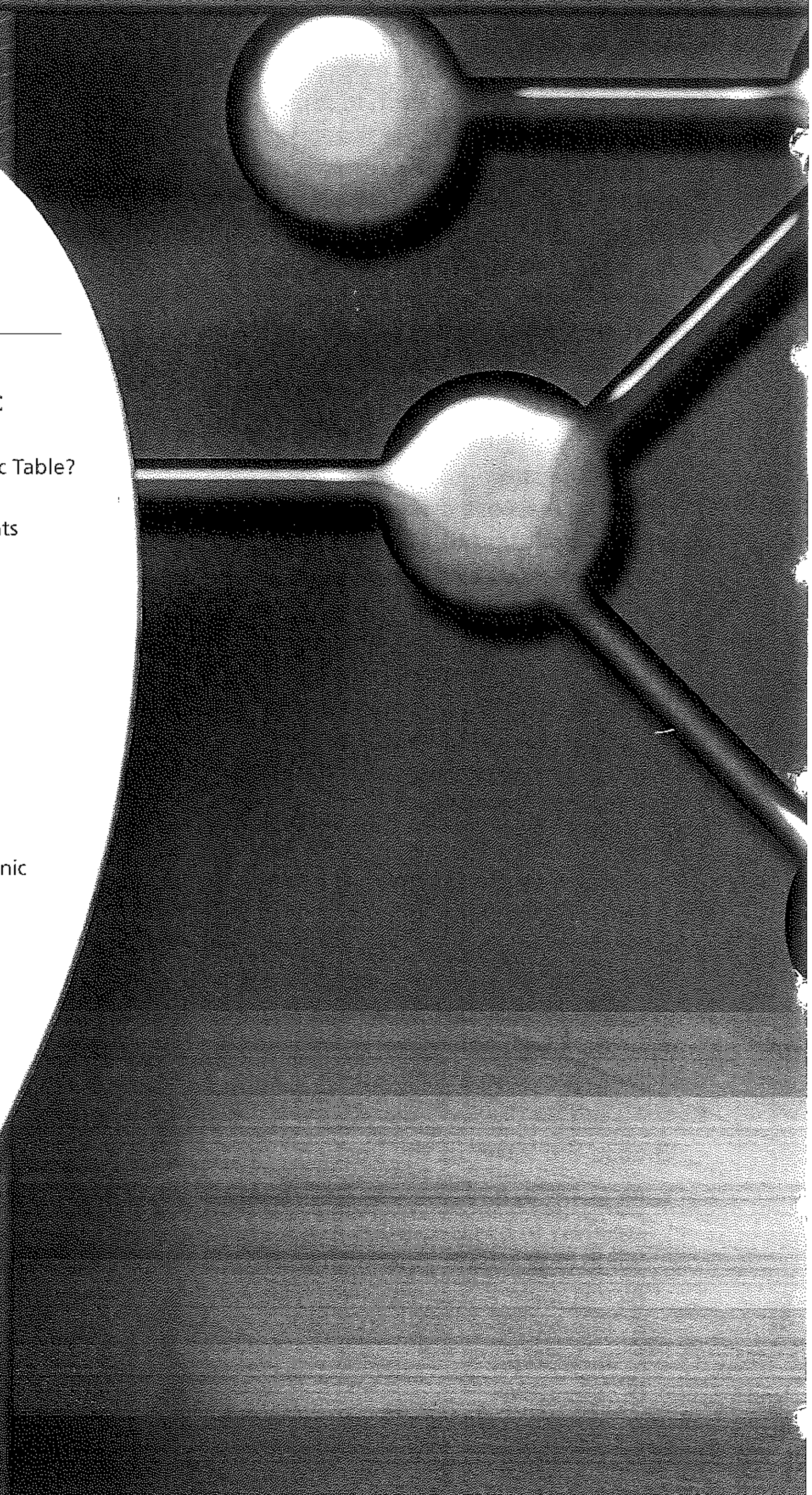
Chapter

5 Atoms and Bonding



The BIG Idea
Properties of Matter

How do compounds form?

❶ Atoms, Bonding, and the Periodic Table
Discover What Are the Trends in the Periodic Table?
Active Art Periodic Table
Science and History Discovery of the Elements
At-Home Activity Looking for Elements
Skills Lab Comparing Atom Sizes

❷ Ionic Bonds
Discover How Do Ions Form?
Skills Activity Interpreting Data
Try This Crystal Clear
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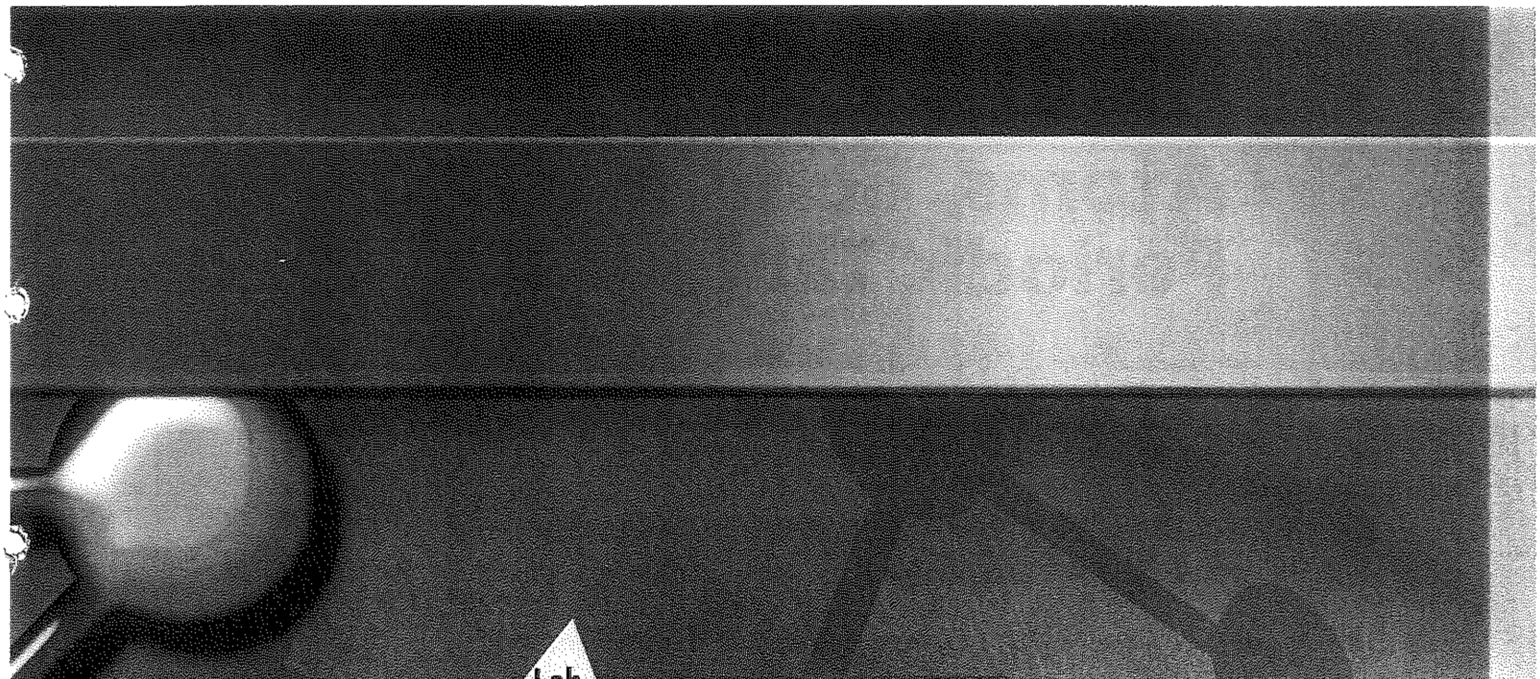
❸ Covalent Bonds
Discover Can Water and Oil Mix?
Analyzing Data Comparing Molecular and Ionic Compounds
At-Home Activity Laundry Chemistry

❹ Bonding in Metals
Discover Are They "Steel" the Same?
Try This What Do Metals Do?



Lab zone™ Chapter **Project**

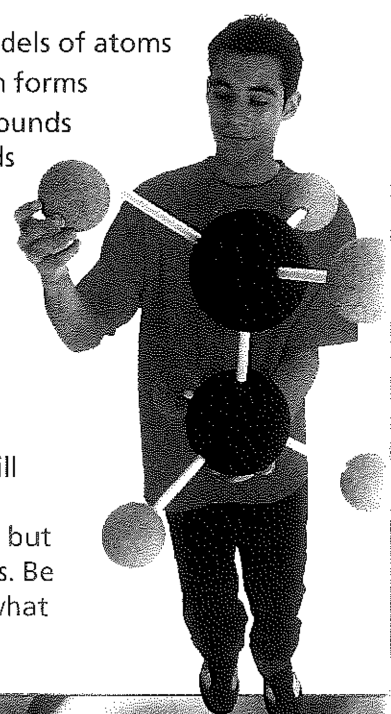
Models of Compounds

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 project, 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 project, 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

Reading Preview

Key Concepts

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

Key Terms

- valence electron
- electron dot diagram
- chemical bond

Target Reading Skill

Identifying Main Ideas As you read *How the Periodic Table Works*, write the main idea in a graphic organizer like the one below. Then write three supporting details that give examples of the main idea.

Main Idea

The number of valence electrons in atoms affects the reactivity of elements.

Detail	Detail	Detail

Lab zone Discover Activity

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
Skydivers 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.

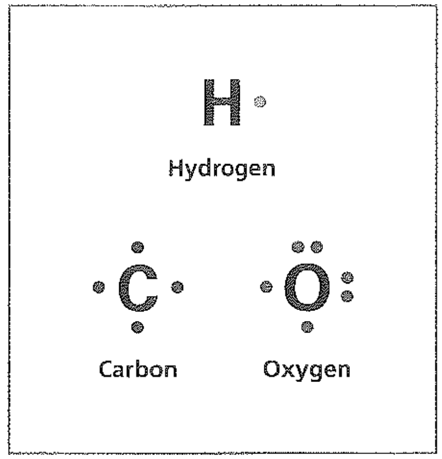
Chemical Bonds and Stability Atoms of most elements are more stable—that is, less likely to react—when they have eight valence electrons. For example, atoms of neon, argon, krypton, and xenon all have eight valence electrons and are very unreactive. These elements do not easily form compounds. Some small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



Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	103	104	105	106	107	108	109	110	111	112 *Uub	113 *Uut	114 *Uuq	115 *Uup	116 *Uuh		118 *Uuo

*Discovery not yet confirmed

Lanthanides

57 La	58 Ce	59 Pr	60 Nd	61	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb

Actinides

89 Ac	90 Th	91 Pa	92 U	93	94	95	96	97	98	99	100	101	102

FIGURE 3
Periodic Table of the Elements
The periodic table is a system used worldwide for organizing the elements. Clues to an element's properties relate to its position in the table.

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How the Periodic Table Works

Recall that the periodic table is organized by atomic number—the number of protons in the nucleus of an atom. **The periodic table gives you information about the arrangement of electrons in atoms.** 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 and think about how atoms change from left to right across a period, or row. As the atomic number increases, the number of electrons also increases. Except for Period 1, a 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, or column, always have the same number of valence electrons.

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.) **The elements within a group have similar properties because they all have the same number of valence electrons in their atoms.**

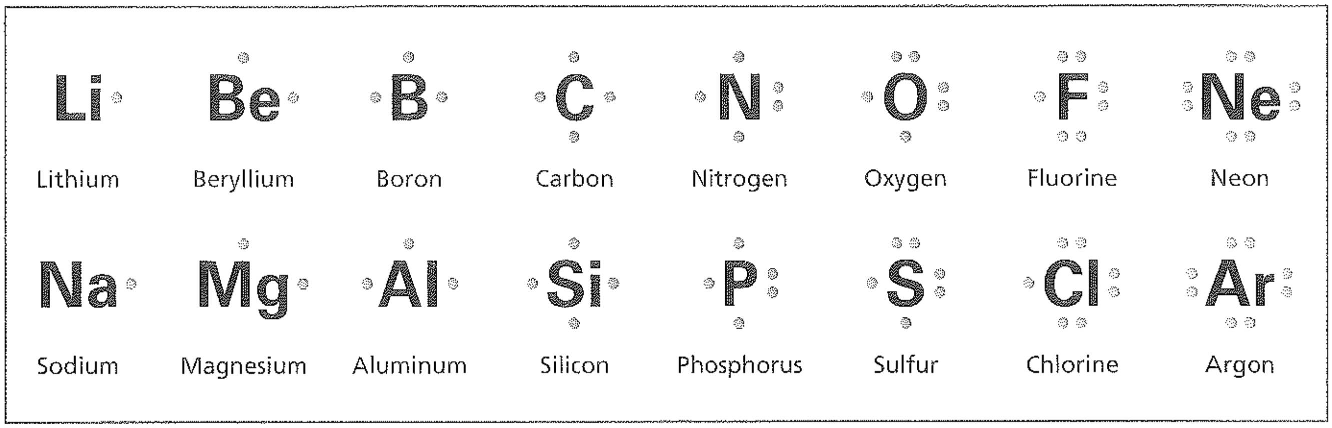


FIGURE 4
Patterns of Valence Electrons
After the number of valence electrons reaches 8, a new period begins.

Noble Gases The Group 18 elements are the noble 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, noble gases do not react easily with other elements. Even so, chemists have been able to make noble 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 noble 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 noble 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?*

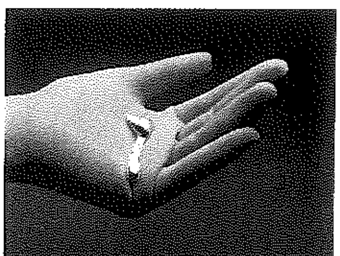
Other Metals Look at the elements in Groups 2 through 12 of the periodic table. Like the Group 1 elements, these elements are metals. Most have one, two, or three valence electrons. They react by losing these electrons, especially when they combine with oxygen or one of the halogens.

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**

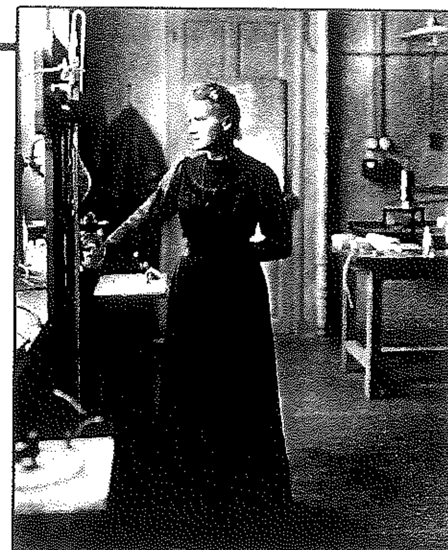
Discovery of the Elements

In 1869, Dmitri Mendeleev published the first periodic table. At that time, 63 elements were known. Since then, scientists have discovered or created about 50 new elements.



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
Polish chemist Marie Curie started with three tons of uranium ore before she eventually isolated a few grams of two new elements. She named them polonium and radium.

1830 1865 1900

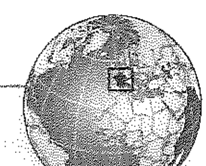
Other Nonmetals Elements in the green section of the periodic table are the nonmetals. Notice that, unlike the metals, most nonmetals are gases at room temperature. Five nonmetals are solids, and one is a liquid. All of the nonmetals have four or more valence electrons. Like the halogens, other nonmetals 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 noble gases.

Writing in Science

Research and Write Select three elements that interest you and find out more about them. Who identified or discovered the elements? How did the elements get their names? How are the elements used? To answer these questions, look up the elements in reference books.

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.

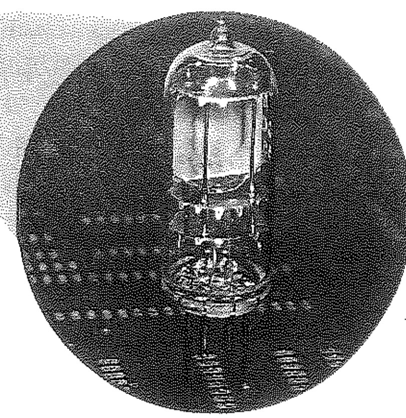


1939
Francium
Although Mendeleev predicted the properties of an element he called "eka-cesium," the element was not discovered until 1939. French chemist Marguerite Perey named her discovery francium, after the country France.

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, such as Lise Meitner, shown here in 1946. All of the new elements were created in laboratories, and none is stable enough to exist in nature.

2003 to Present
More New Elements
Darmstadtium (110) and roentgenium (111) are named. Research to produce and study new synthetic elements continues.

1935 | 1970 | 2005



◀ The quartz movement of the watch

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

Metalloids Several elements known as metalloids lie along a zigzag line between the metals and nonmetals. The metalloids 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 considered to be a nonmetal. 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

Target Reading Skill Identifying Main Ideas Use your graphic organizer to help you answer Question 2 below.

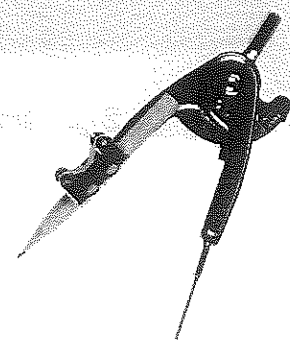
Reviewing Key Concepts

1. a. Defining What are valence electrons?
   b. Reviewing What role do valence electons play in the formation of compounds from elements?
   c. Comparing and Contrasting Do oxygen atoms become more stable or less stable when oxygen forms compounds? Explain.
2. a. Summarizing Summarize how the periodic table is organized, and tell why this organization is useful.
   b. Explaining Why do the properties of elements change in a regular way across a period?
   c. Relating Cause and Effect How reactive are the elements in Group 18? Explain this reactivity in terms of the number 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

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. 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.
4. 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.*
5. Label each model with the symbol of the element it represents.

Data Table			
Atomic Number	Element	Radius (pm)*	Relative Radius
4	Be	112	1
12	Mg	160	
20	Ca	197	
38	Sr	215	
56	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 bar graph of the data given in the first 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 Do the points on your graph fall on a straight line or on a curve? What trend do the data show?
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

Reading Preview

Key Concepts

- What are ions, and how do they 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

Target Reading Skill

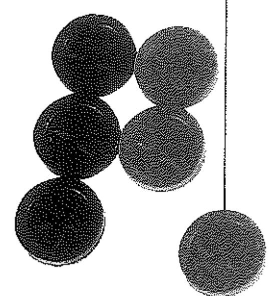
Previewing Visuals Before you read, preview Figure 9. Then write two questions that you have about the diagram in a graphic organizer like the one below. As you read, answer your questions.

Formation of an Ionic Bond

Q. What is an ionic bond?
A.
Q.

Lab zone Discover Activity

How Do Ions Form?

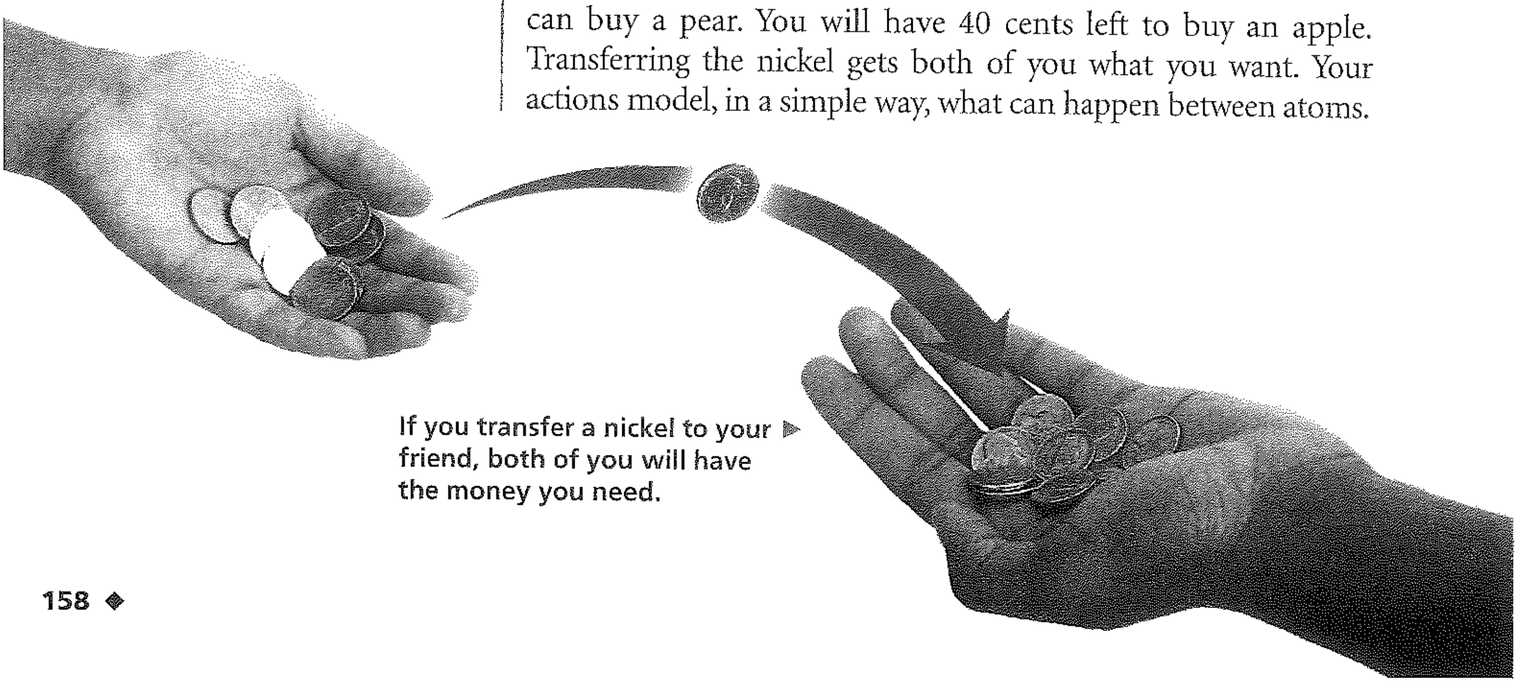


1. Place three pairs of checkers (three red and three black) on your desk. The red represent electrons and the black represent protons.
2. Place nine pairs of checkers (nine red and nine black) in a separate group on your desk.
3. Move a red checker from the smaller group to the larger group.
4. Count 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. ▶

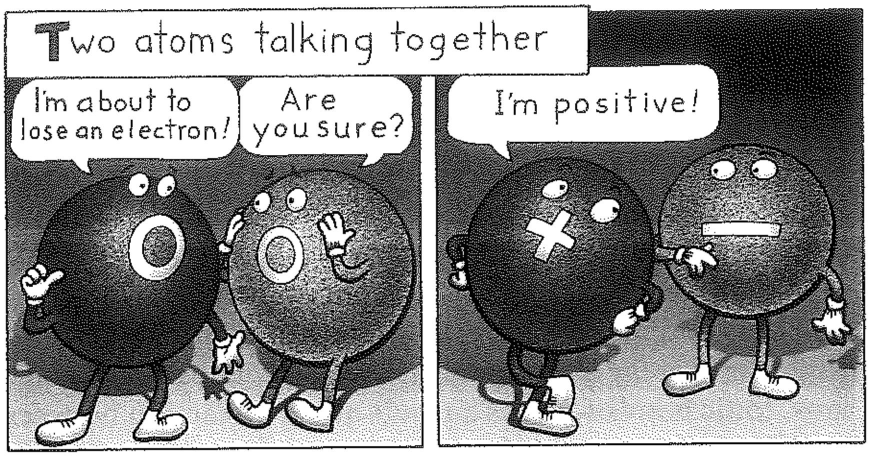


FIGURE 7
How Ions Form
When an atom loses one of its electrons, it becomes a positively charged ion. The atom that gains the electron becomes a negatively charged ion.

Ions and Ionic Bonds

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, 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?*

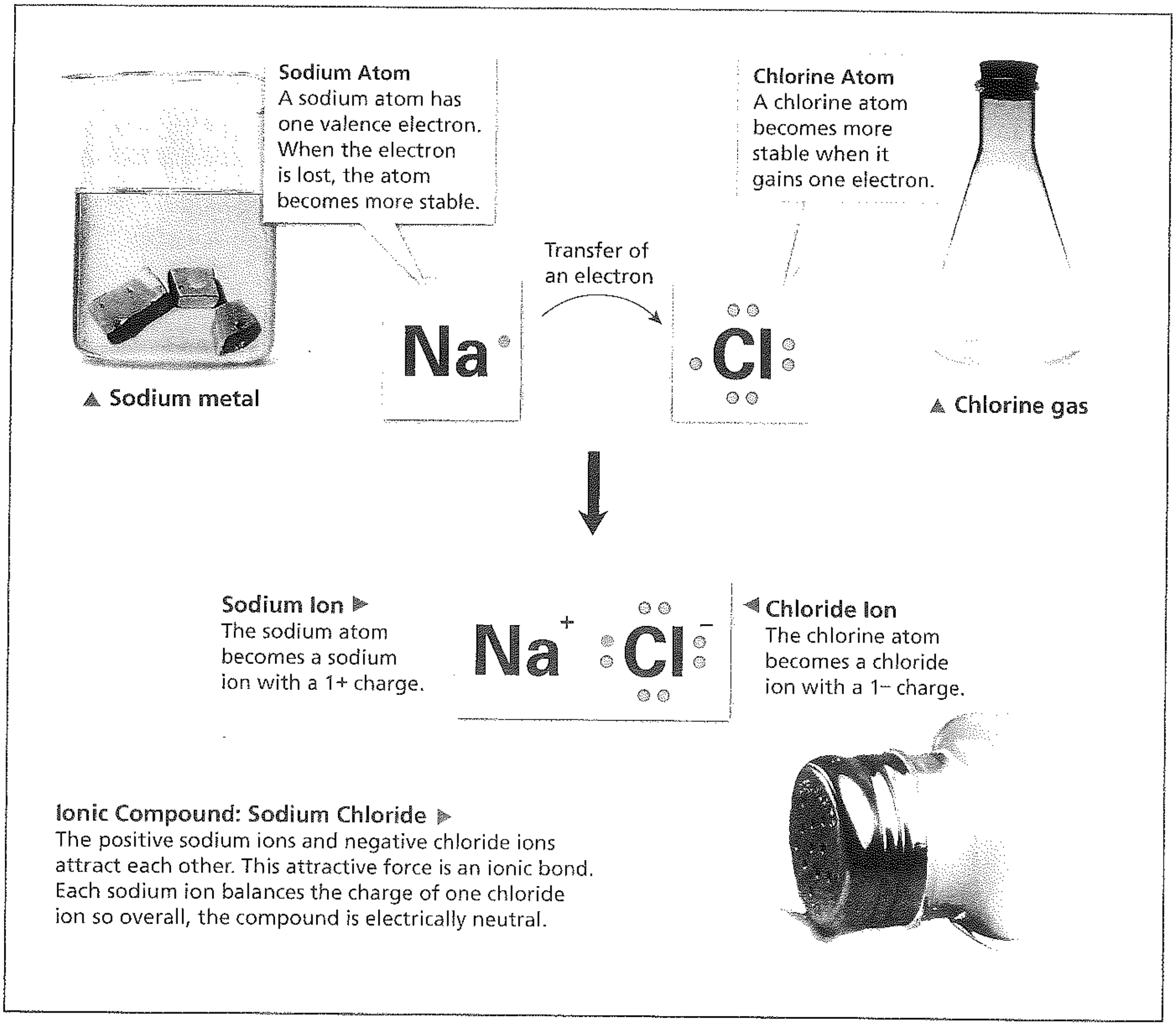
Ionic Bonds Look at Figure 9 to see how sodium atoms and chlorine atoms combine to form sodium chloride (table salt). Notice that sodium has one valence electron and chlorine has seven valence electrons. When sodium's valence electron is transferred to chlorine, both atoms become ions. The sodium atom 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?

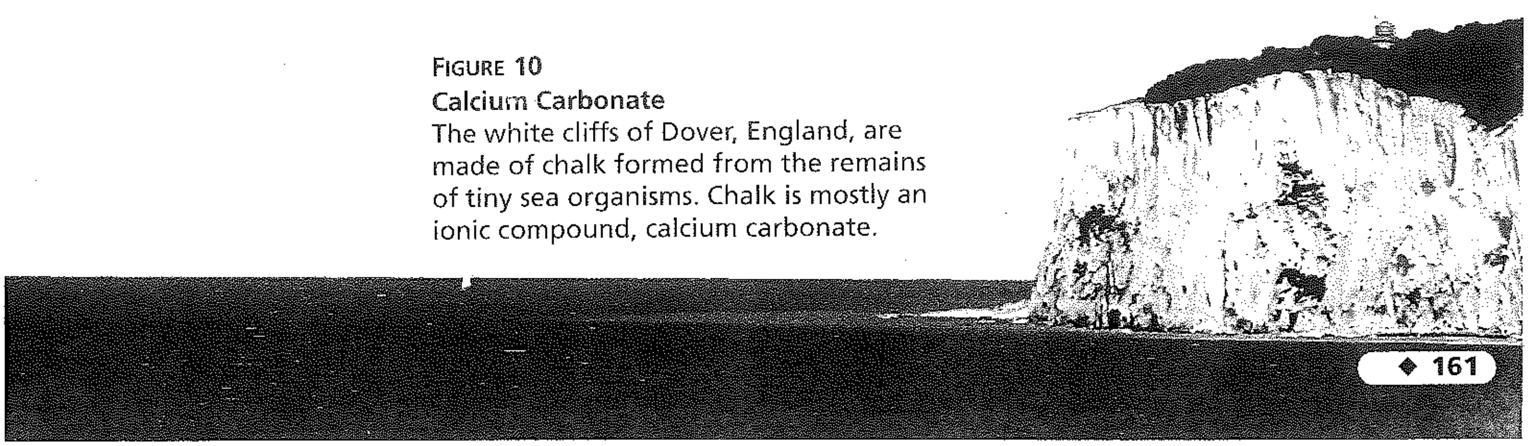


FIGURE 10
Calcium Carbonate
The white cliffs of Dover, England, are made of chalk formed from the remains of tiny sea organisms. Chalk is mostly an ionic compound, calcium carbonate.

Lab zone **Try This Activity**

Crystal Clear

Can you grow a salt crystal?

1. Add 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

Table salt, baking soda, and iron rust are different compounds with different properties. You wouldn't want to season your food with either iron rust or baking soda. However, these compounds are alike in some ways because they are all ionic compounds. **In general, ionic compounds are hard, brittle crystals that have high melting points. When dissolved in water or melted, they conduct electricity.**

Ionic Crystals Figure 11 shows a chunk of halite, or table salt, NaCl. 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. It is attracted to ions above, below, and to all sides. 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 What happens when you heat an ionic compound such as table salt? When you heat a substance, its energy increases. When ions have enough energy to overcome the attractive forces between them, they break away from each other. In other words, the crystal melts to 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. They are all solids at room temperature. Table salt must be heated to 801°C before the crystal melts.

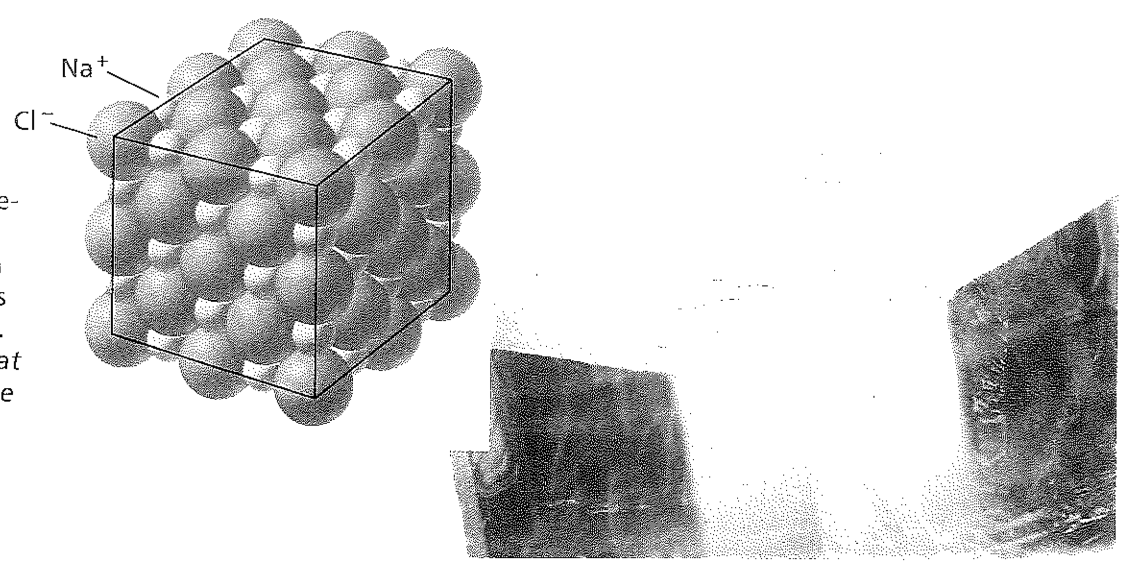


FIGURE 11
Ionic Crystals
The ions in ionic compounds are arranged in specific three-dimensional shapes called crystals. Some crystals have a cube shape like these crystals of halite, or sodium chloride.
Making Generalizations *What holds the ions together in the crystal?*

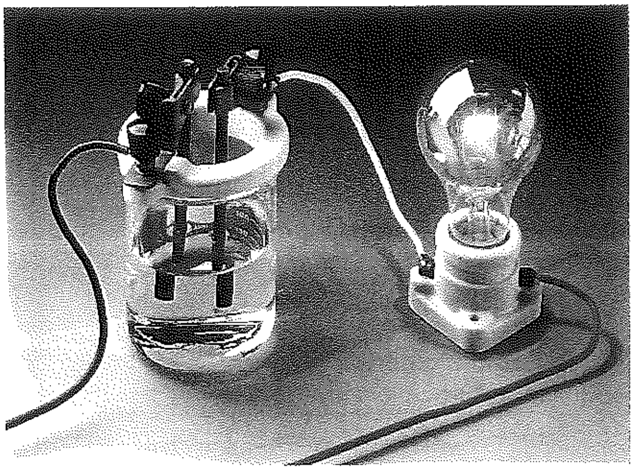


FIGURE 12
Ions in Solution
A solution of sodium chloride conducts current across the gap between the two black rods of a conductivity tester. As a result, the bulb lights up.

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.

Reading Checkpoint **What is a crystal?**

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

Section 2 Assessment

Target Reading Skill Previewing Visuals Compare your questions and answers about Figure 9 with those of a partner.

Reviewing Key Concepts

1. a. Reviewing What are the two basic ways in which ions form from atoms?
   b. Comparing and Contrasting Contrast sodium and chloride ions, including how they form. Write the symbol for each ion.
   c. Relating Cause and Effect What holds the ions together in sodium chloride? Indicate the specific charges that are involved.
2. a. Identifying What information is given by the formula of an ionic compound?
   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.
3. a. Listing List three properties of ionic compounds.
   b. Making Generalizations Relate each property that you listed to the characteristics of ionic bonds.

Writing in Science

Firsthand Account Pretend that you are the size of an atom, 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

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

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.

Data Table

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

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.
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.

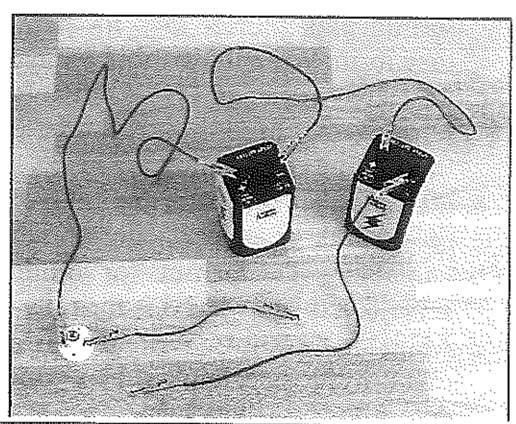
Making a Conductivity Tester

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.



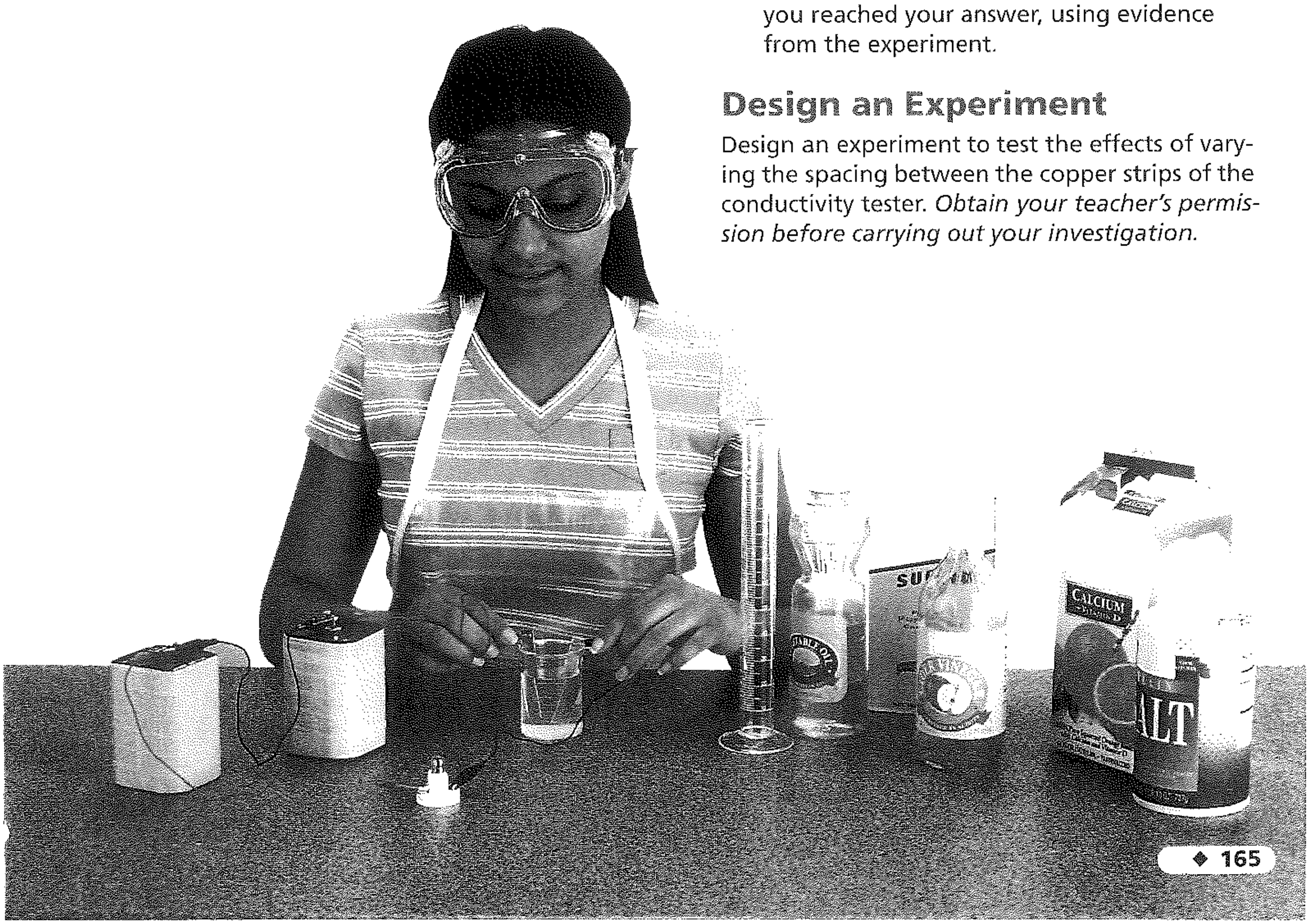
6. Add 1 spoonful of sodium chloride to the distilled water in the beaker. Stir with the spoon until the salt dissolves. Repeat the conductivity test and record your observations.
7. Disconnect the conductivity tester and rinse the beaker, spoon, and copper strips with distilled water. Dry the beaker 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. Controlling Variables Why did you test both tap water and distilled water before testing the sodium chloride solution?
2. Interpreting Data Could you have used tap water in your tests instead of distilled water? Explain.
3. Drawing Conclusions Based on your observations, add a column to your data table indicating whether each substance produced ions in solution.
4. Inferring Sodium chloride is an ionic compound. How can you account for any observed differences in conductivity between dry and dissolved sodium chloride?
5. Communicating Based on your observations, decide whether or not you think sucrose (table sugar) is made up of ions. Explain how you reached your answer, using evidence from the experiment.

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

Reading Preview

Key Concepts

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

Key Terms

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

Target Reading Skill

Asking Questions Before you read, preview the red headings. In a graphic organizer like the one below, ask a *what* or *how* question for each heading. As you read, answer your questions.

Covalent Bonds

Question	Answer
How do covalent bonds form?	Covalent bonds form when...

Lab zone Discover Activity

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.

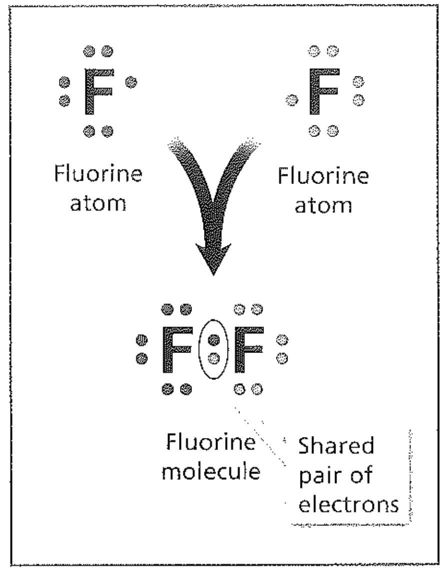
FIGURE 13
Sharing Skills
One student is a skilled artist, while the other is a skilled writer. By pooling their skills, the students can complete their project.

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 Recall that the noble gases are not very reactive. In contrast, all other nonmetals, including hydrogen, can bond to other nonmetals by sharing electrons. Most nonmetals can even bond with another atom of the same element, as is the case with fluorine in Figure 14. When you count the electrons on each atom, count the shared pair each time. By sharing electrons, each 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.

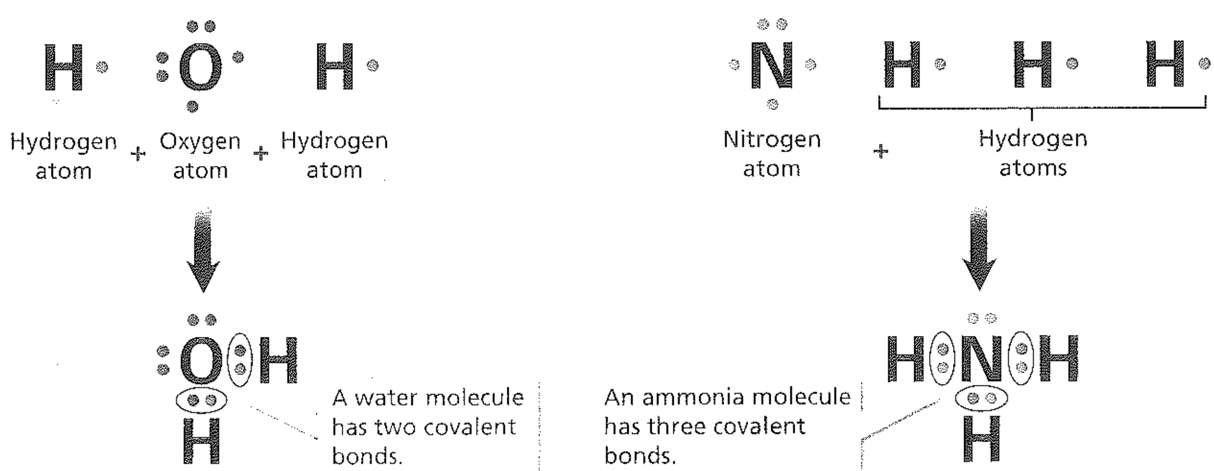
FIGURE 14
Sharing Electrons
By sharing electrons in a covalent bond, each fluorine atom has a stable set of eight valence electrons.



How Many Bonds? Look at the electron dot diagrams in Figure 15. Count the valence electrons around each atom. Except for hydrogen, the number of covalent bonds that nonmetal atoms can form equals the number of electrons needed to make a total of eight. Hydrogen needs only two electrons.

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. Do you see why water's formula is H_2O, instead of H_3O, H_4O, or just HO?

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?*



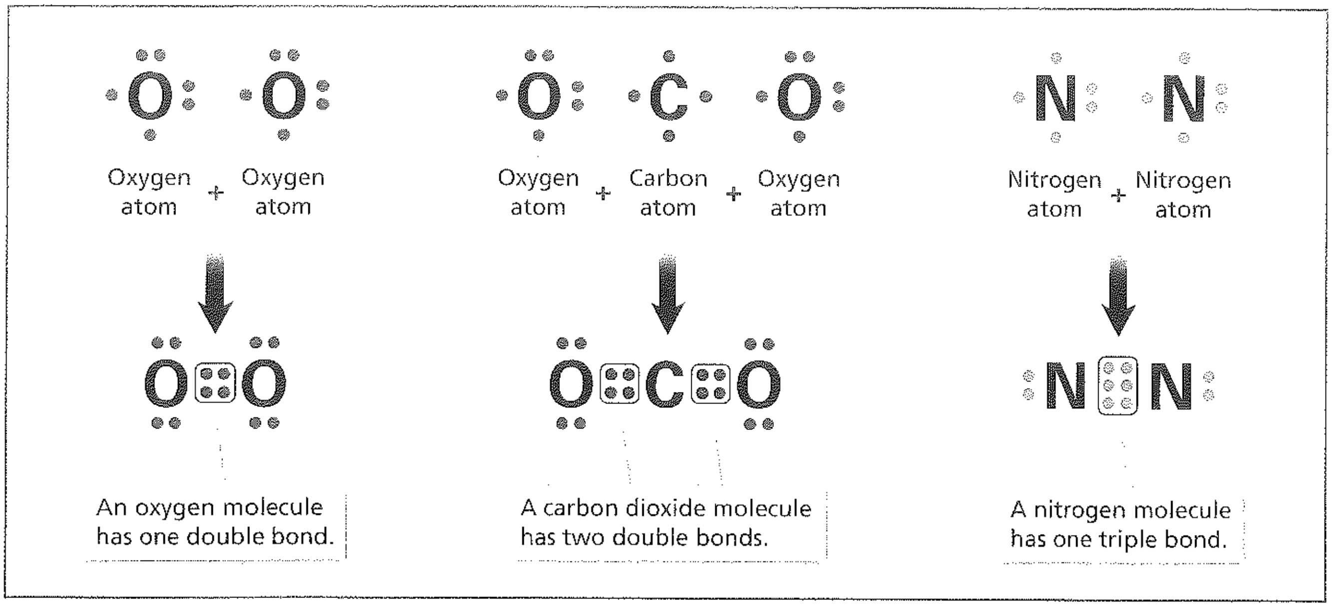


FIGURE 16
Double and Triple Bonds
Double and triple bonds can form when atoms share more than one pair of electrons.
Interpreting Diagrams *In a nitrogen molecule, how many electrons does each nitrogen atom share with the other?*

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. Molecular compounds have very different properties than ionic compounds. **Compared to ionic compounds, molecular compounds generally have lower melting points and boiling points, and they do not conduct electricity when 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.

Go Online
SCI LINKS NSTA
For: Links on molecular compounds
Visit: www.SciLinks.org
Web Code: scn-1214

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. Put the molecular compounds on the left and the ionic compounds on the right. 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

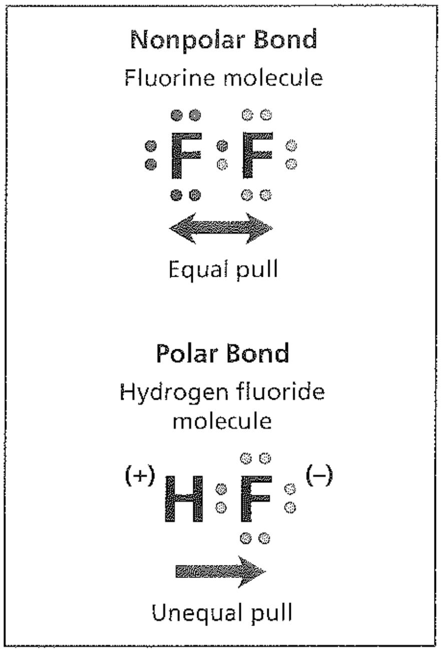
5. **Predicting** Ammonia's melting point is −78°C and its boiling point is −34°C. Is ammonia a molecular compound or an ionic compound? Explain.

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 on the rope with unequal force? 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 pulled more toward one atom, causing the bonded atoms to have slight electrical charges.** These charges are not as strong as the charges on ions.

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.



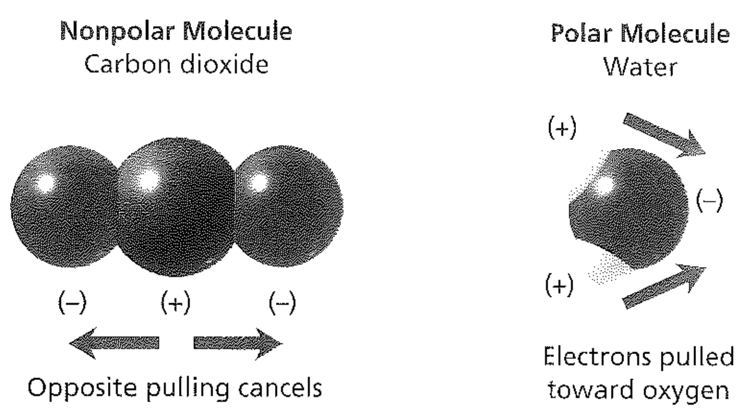
Polar Bonds and Nonpolar Bonds The unequal sharing of electrons is enough to make the atom with the stronger pull slightly negative and the atom with the weaker pull slightly positive. A covalent bond in which electrons are shared unequally is called a **polar bond.** Of course, if two atoms pull equally 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 It makes sense that a molecule with nonpolar bonds will itself be nonpolar. But a molecule may contain polar bonds and still be nonpolar overall. 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 shape like a straight line. So, the two oxygen atoms pull with equal strength in opposite directions. In a sense, the attractions cancel out, and the molecule is nonpolar.

In contrast, other molecules that have polar covalent bonds are themselves polar. In a water molecule, the two hydrogen atoms are at one end of the molecule, while the oxygen atom is 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 Between 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 between 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 Discover 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

Target Reading Skill Asking Questions Use the answers to the questions you wrote about the headings to help you answer the questions below.

Reviewing Key Concepts

1. a. Identifying What is the attraction that holds two covalently bonded atoms together?
   b. Inferring A carbon atom can form four covalent bonds. How many valence electrons does it have?
   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.
2. a. Reviewing How are the properties of molecular compounds different from those of ionic compounds?
   b. Relating Cause and Effect Why are most molecular compounds poor conductors of electricity?
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?
   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?
   c. Predicting Predict whether carbon dioxide or water would have a higher boiling point. Explain your prediction in terms of the attractions between molecules.

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.

Section 4 · Tech & Design ·

Bonding in Metals

Reading Preview

Key Concepts

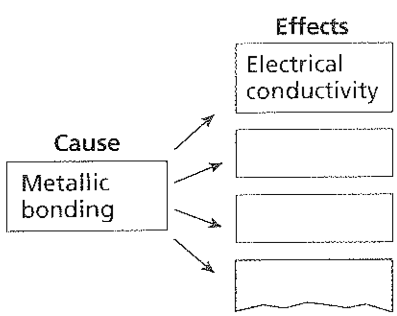
- How do the properties of metals and alloys compare?
- How are metal atoms bonded in solid metal?
- How does metallic bonding result in useful properties of metals?

Key Terms

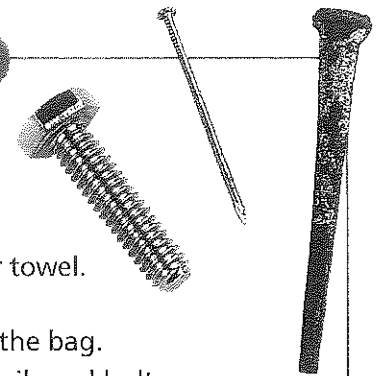
- alloy
- metallic bond

Target Reading Skill

Relating Cause and Effect As you read, identify the properties of metals that result from metallic bonding. Write the information in a graphic organizer like the one below.



Lab zone Discover Activity



Are They "Steel" the Same?

1. Wrap a cut nail (low-carbon steel), a wire nail (high-carbon steel), and a stainless steel bolt 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 metal an ideal material for making furniture, musical instruments, electrical wire, pots and pans, eating utensils, and strong beams for buildings? Why do metals have these physical properties?

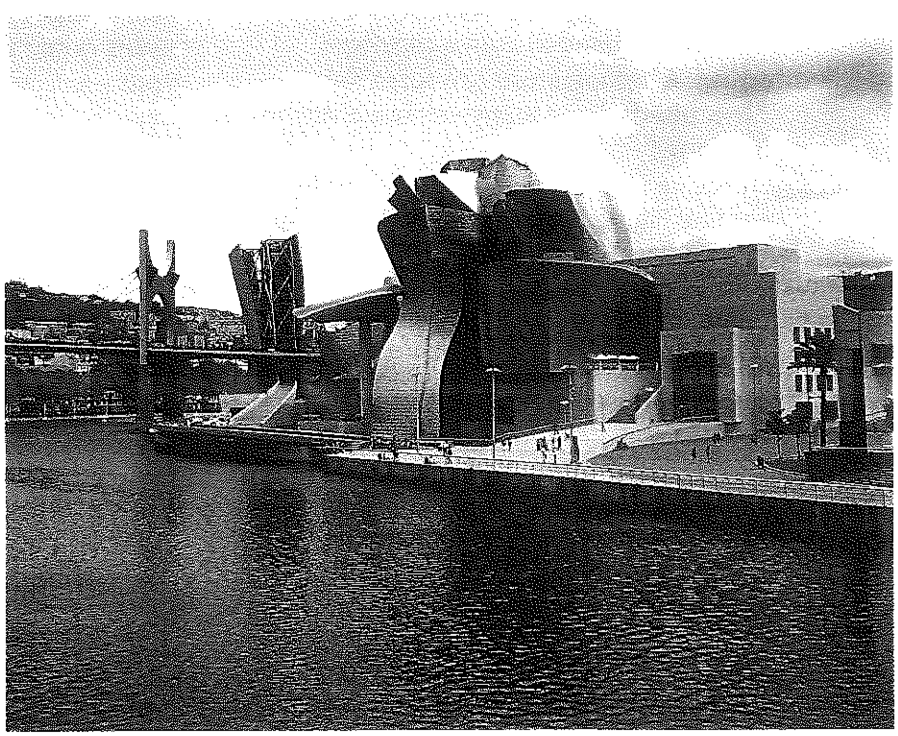


FIGURE 20
Metal in Architecture
The Guggenheim Museum in Bilbao, Spain, makes dramatic use of some properties of metals. The museum's shiny outer "skin" is made of the lightweight metal titanium, which can be pressed into large, thin, flexible sheets.

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

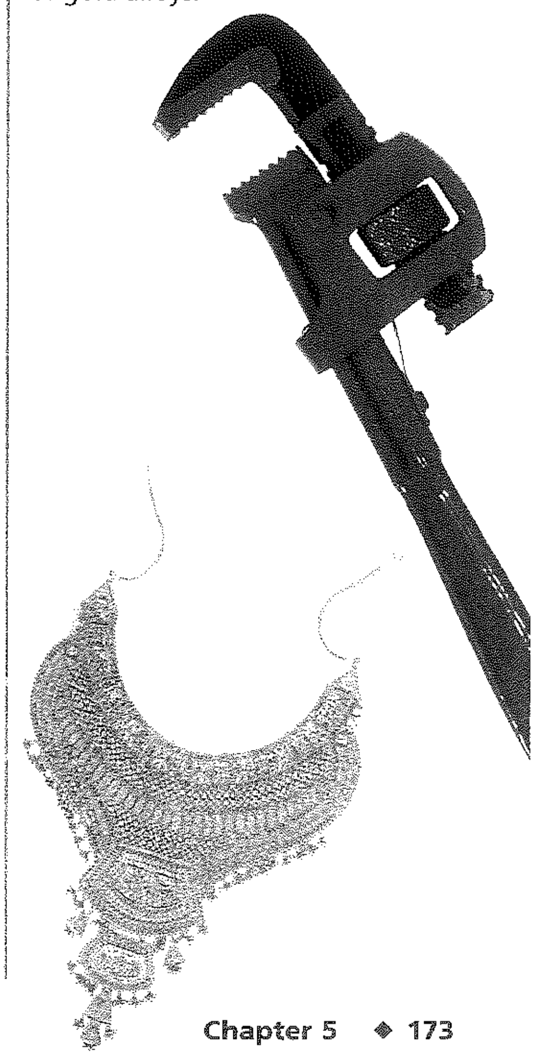
You know a piece of metal when you see it. It's usually hard and shiny. At room temperature, most metals are solids. They can be hammered flat or drawn out into thin wire. Electronics such as stereos, computers, and MP3 players have metal parts because metals conduct electric current well. However, very few of the "metals" you use every day are made of one element. Instead, metals are usually used in the form of an alloy. An **alloy** is a mixture made of two or more elements that has the properties of metal. In every alloy, at least one of the elements is a metal. **Alloys are generally stronger and less likely to react with air or water than are 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 an extremely strong metal that 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 nearly as strong as iron but 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.

The loosely held electrons in metal atoms result in a type of bonding that is characteristic of metals. Like many solids, metals exist as crystals. The metal atoms are very close together and in specific arrangements. These atoms are actually positively charged ions. Their valence electrons are free to drift among the ions. Each metal ion is held in the crystal by a **metallic bond**—an attraction between a positive metal ion and the electrons surrounding it. Look at Figure 23. **A metal or metal alloy consists of positively charged metal ions 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?

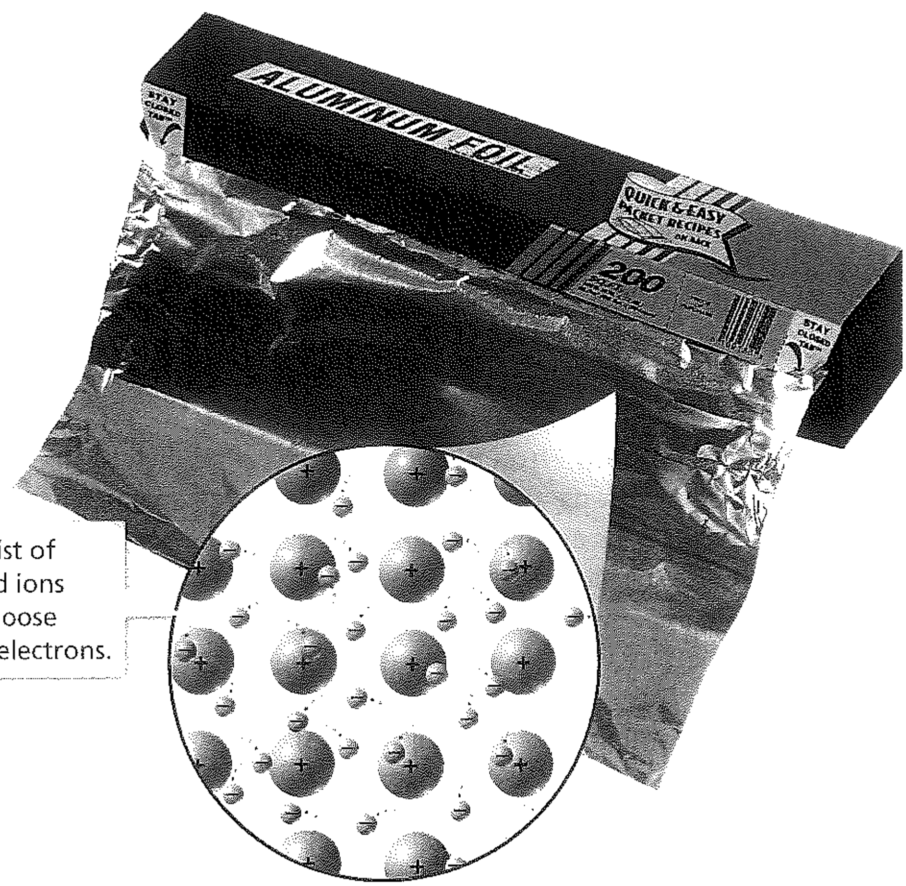


FIGURE 23
Metallic Bonding
The type of bonding in metals is the result of loosely held electrons. **Problem Solving** *Why would nonmetals be unlikely to have the type of bonding shown here?*

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 would not be very useful if the metal couldn't be bent or hammered into a useful shape.

Metallic bonding explains many of the common physical properties of metals and their alloys. **The "sea of electrons" model of solid metals explains the ease with which they can change shape, their ability to conduct electric current, their luster, and their ability to conduct heat.**

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For: Links on metallic bonding
Visit: www.SciLinks.org
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Changes in Shape 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.

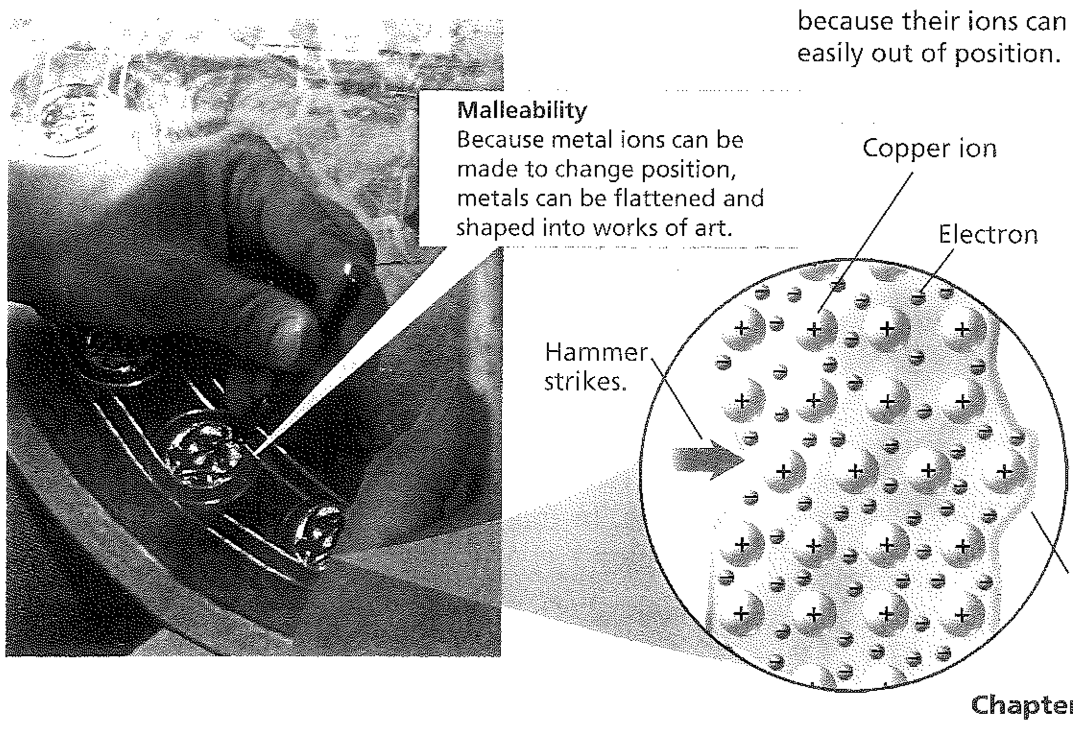
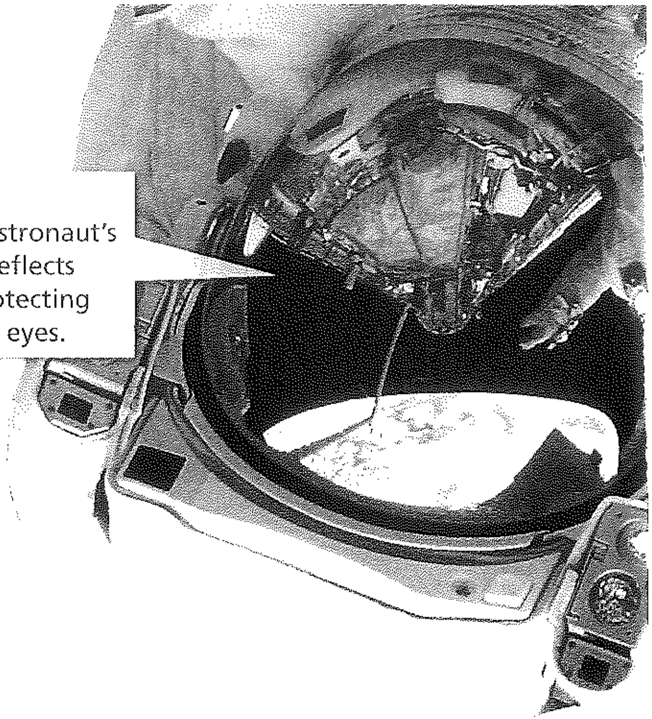
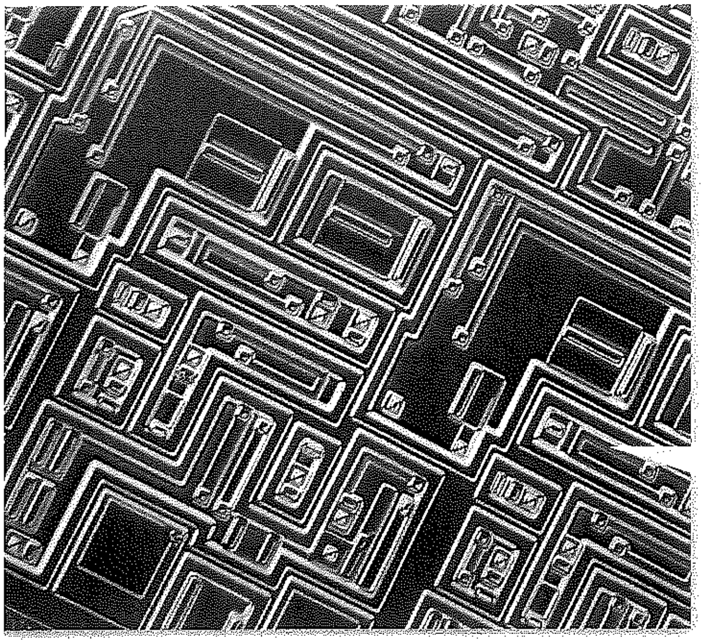


FIGURE 24
Flexibility of Metals
Most metals can be reshaped because their ions can be pushed easily out of position.

FIGURE 25
Conductivity and Luster 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.



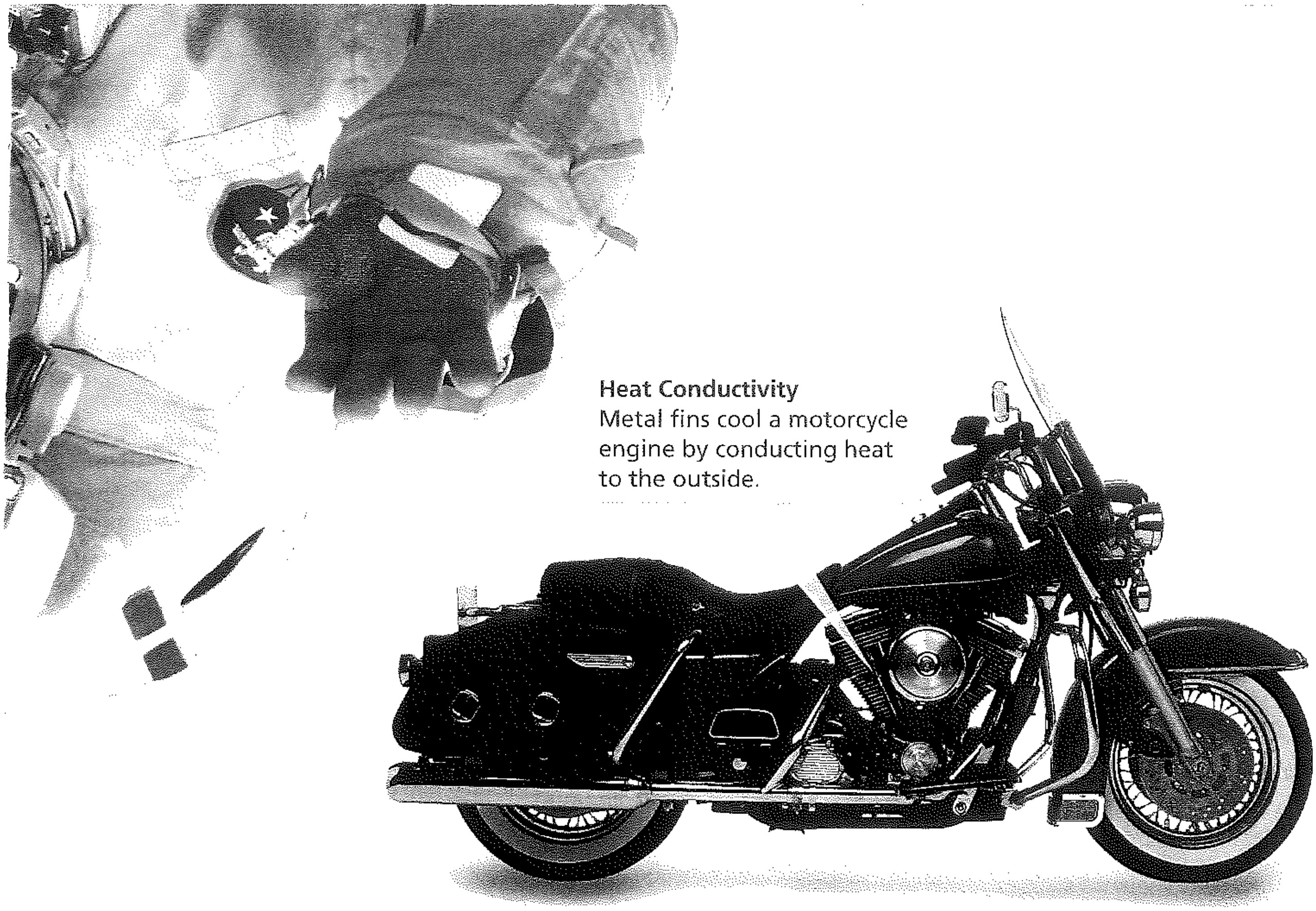
Electrical Conductivity
Metal strips on a circuit board conduct electric current throughout the circuit.

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, there is a current into the metal at one point and out at another point.

Luster Polished metals exhibit 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.

Heat Conductivity Heat causes particles of matter to move faster. If these particles collide with cooler particles, energy is transferred to the cooler particles. In a metal, the freely moving valence electrons transfer energy to nearby atoms and other electrons. In this way, heat travels easily through a metal or metal alloy.

Reading Checkpoint **Why are metals and their alloys shiny?**



Heat Conductivity
Metal fins cool a motorcycle engine by conducting heat to the outside.

Section 4 Assessment

Target Reading Skill Relating Cause and Effect Refer to your graphic organizer about metallic properties to help you answer Question 3 below.

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. 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 can include illustrations on your label as well.

Chapter 5

Study Guide

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

1 Atoms, Bonding, and the Periodic Table

Key Concepts

- 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 gives you information about the arrangement of electrons in atoms.
- The elements within a group have similar properties because they all have the same number of valence electrons in their atoms.

Key Terms

valence electron
electron dot diagram
chemical bond

2 Ionic Bonds

Key Concepts

- When an atom loses an electron, it becomes a positive ion. When an atom gains an electron, it becomes a negative ion.
- 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.
- Ionic compounds are hard, brittle crystals that have high melting points and conduct electricity when dissolved in water.

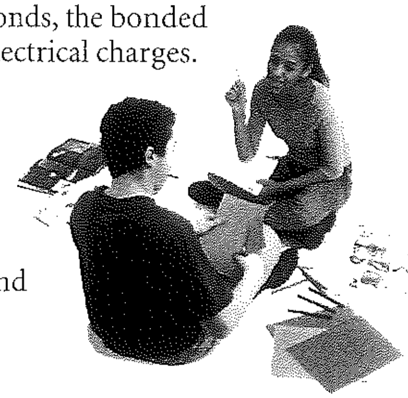
Key Terms

ion
polyatomic ion
ionic bond
ionic compound
chemical formula
subscript
crystal

3 Covalent Bonds

Key Concepts

- 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 electricity.
- In polar covalent bonds, the bonded atoms 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

- Alloys are generally stronger and less likely to react with air or water than are the pure metals from which they are made.
- A metal or metal alloy consists of positively charged metal ions in a "sea" of valence electrons.
- The "sea of electrons" model of solid metals explains the ease with which they can change shape, their ability to conduct electric current, their luster, and their ability to conduct heat.

Key Terms

alloy
metallic bond

Review and Assessment

Go Online
PHSchool.com
For: Self-Assessment
Visit: PHSchool.com
Web Code: cka-1050

Organizing Information

Comparing and Contrasting Copy the graphic organizer about chemical bonds onto a separate sheet of paper. Then complete it. (For more on Comparing and Contrasting, see the Skills Handbook.)

Types of Chemical Bonds

Feature	Ionic Bond	Polar Covalent Bond	Nonpolar Covalent Bond	Metallic Bond
How Bond Forms	a. ?	Unequal sharing of electrons	b. ?	c. ?
Charge on Bonded Atoms?	Yes; positive or negative	d. ?	e. ?	Yes; positive
Example	f. ?	g. ?	O_2 molecule	h. ?

Reviewing Key Terms

Choose the letter of the best answer.

1. Valence electrons in an atom are those that
   a. are held most loosely.
   b. have the lowest energy level.
   c. are always easily lost.
   d. are never easily lost.
2. An electron dot diagram shows an atom's number of
   a. protons. b. electrons.
   c. valence electrons. d. chemical bonds.
3. On the periodic table, elements with the same number of valence electrons are in the same
   a. square. b. period.
   c. block. d. group.
4. When an atom loses or gains electrons, it becomes a(n)
   a. ion. b. formula.
   c. crystal. d. subscript.
5. A covalent bond in which electrons are shared unequally is a
   a. double bond. b. triple bond.
   c. polar bond. d. nonpolar bond.
6. The metal atoms in stainless steel are held together by
   a. ionic bonds. b. polar bonds.
   c. covalent bonds. d. metallic bonds.

If the statement is true, write *true*. If it is false, change the underlined word or words to make the statement true.

7. A <u>chemical bond</u> is a force of attraction between two atoms.
8. A <u>polyatomic ion</u> is made up of more than one atom.
9. A neutral group of atoms joined by covalent bonds is <u>an ion</u>.
10. <u>An alloy</u> is a mixture of elements that has the properties of a metal.

Writing in Science

Travel Brochure Pretend you have just visited a city modeled on the periodic table. Write a travelogue about how the "city" is organized. Be sure to describe some of the elements you visited and how they are related to their neighbors.

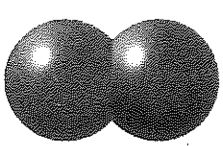
Review and Assessment

Checking Concepts

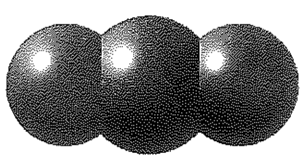
11. Which element is less reactive, an element whose atoms have seven valence electrons or an element whose atoms have eight valence electrons? Explain.
12. Why do ionic compounds generally have high melting points?
13. The formula of sulfuric acid is H_2SO_4. How many atoms of hydrogen, sulfur, and 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

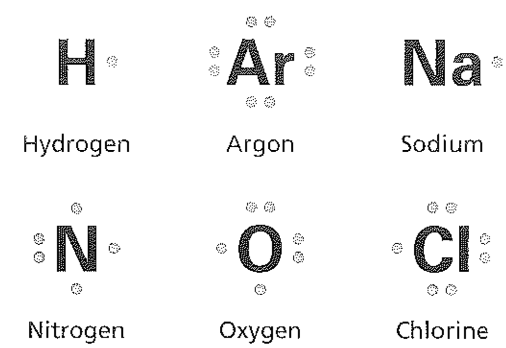


Carbon dioxide

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 Chapter Project

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?

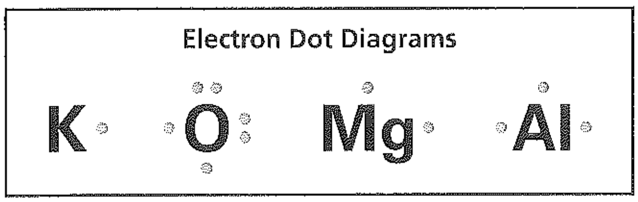
Standardized Test Prep

Test-Taking Tip

Interpreting Diagrams

When answering a question related to a diagram, examine the diagram carefully. Read any titles or labels included. Be sure you understand the meanings of the symbols used. For example, in the diagram below, the dots represent valence electrons. Study the diagram and answer the sample question.

Sample Question



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)

Answer

The correct answer is **C**. Atoms with low numbers of valence electrons are likely to lose electrons. Potassium, magnesium, and aluminum all have low numbers of valence electrons, but only magnesium will lose two electrons, producing an ion with a 2+ charge.

Choose the letter of the best answer.

Use the electron dot diagrams above to answer Questions 1–3.

1. 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

2. 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?
   F 1 : 1 G 1 : 2
   H 2 : 1 J 2 : 2

3. 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.

4. The chemical formula for a glucose molecule is $C_6H_{12}O_6$. The subscripts represent the
   F mass of each element.
   G number of atoms of each element in a glucose molecule.
   H total number of bonds made by each atom.
   J number of valence electrons.

5. 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.

Constructed Response

6. In a working light bulb, a thin tungsten wire filament that is wound in a coil conducts electric current. Describe two properties that make the metal tungsten a good material for the filament of a light bulb. Indicate how the type of bonding in tungsten contributes to these properties.