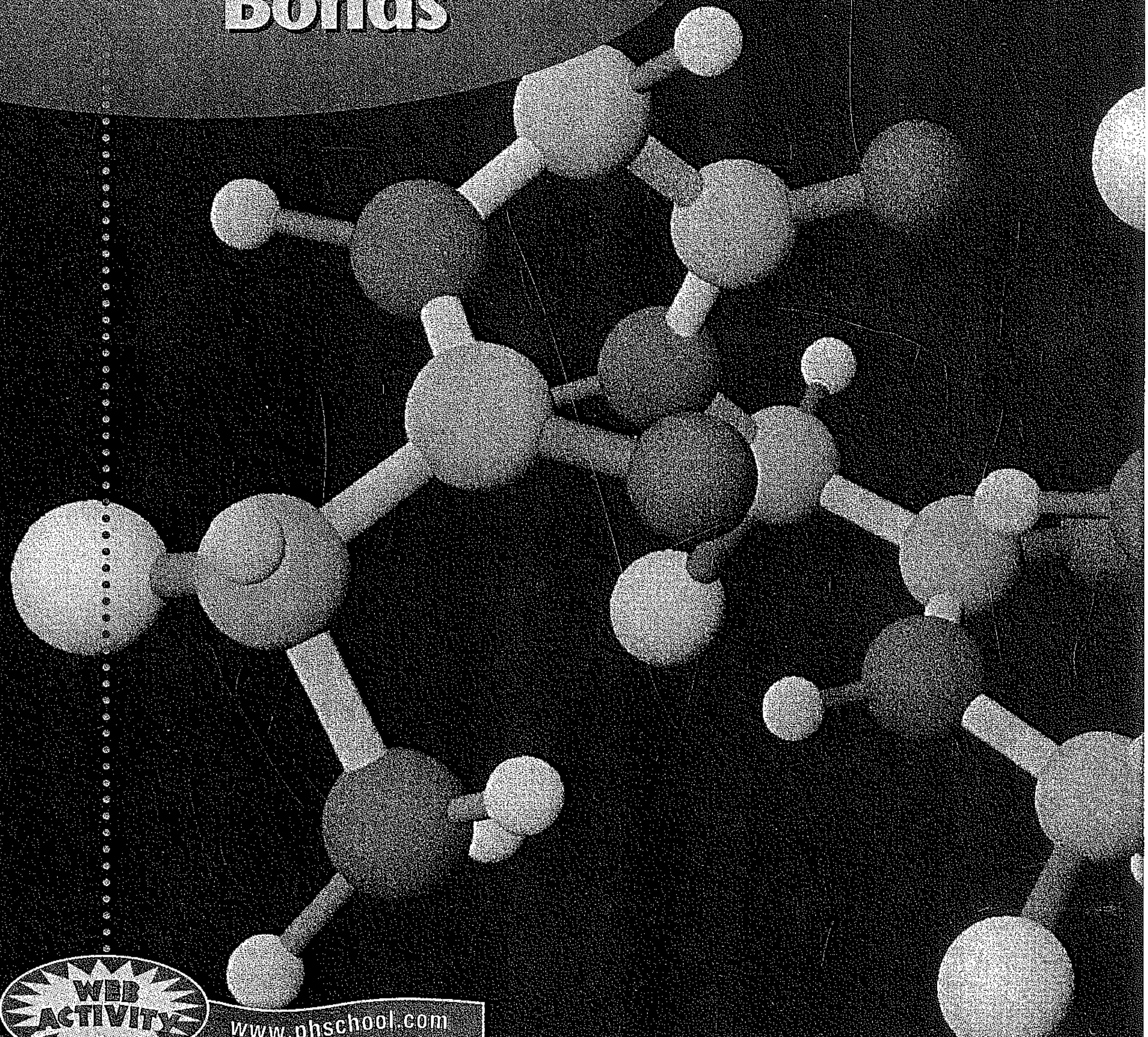
CHAPTER 4

Chemical Bonds



WEB ACTIVITY www.phschool.com

SECTION 1 Ionic Bonds

Discover How Do Ions Form?
Sharpen Your Skills Interpreting Data
Try This Crystal Shapes
Skills Lab Shape Up!

SECTION 2 Covalent Bonds

Discover Why Don't Water and Oil Mix?
Sharpen Your Skills Designing Experiments
Skills Lab Shedding Light on Chemical Bonds

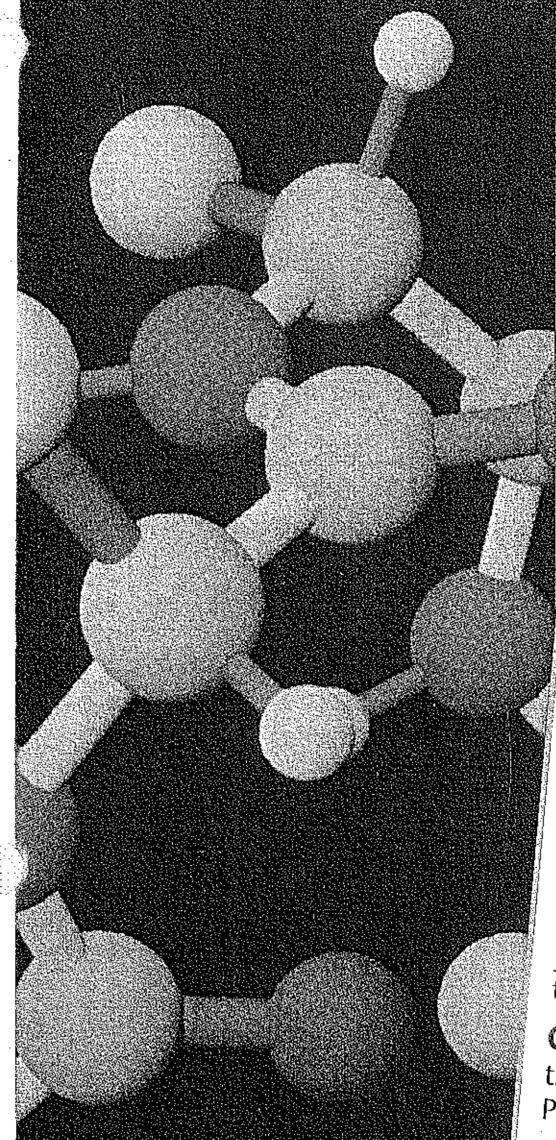
Integrating Earth Science

SECTION 3 Crystal Chemistry

Discover How Small Do They Get?
Science at Home Piling Up

CHAPTER 4 PROJECT

Model Compounds



A computer-made model of a protein shows the many atoms that are bonded together in the molecule.

Because atoms are very tiny, it can be hard to imagine how they join to form compounds. Models can help you understand what happens at the atomic level. The models can be made from common materials or made with computer software.

In this chapter, you will learn why atoms react with one another. You will also learn about the different types of chemical bonds that can hold atoms together. In your project, you can use familiar materials to show how atoms bond to form two different types of compounds.

Your Goal To make models showing how atoms bond in compounds that contain ionic and covalent bonds.

To complete the project you must

- select materials that can be used to represent atoms of different elements
- design ways to show the ionic and covalent bonds that form between atoms
- make and compare models of compounds that contain ionic and covalent bonds

Get Started Brainstorm with some of your classmates about materials you can use to represent atoms and chemical bonds. You may want to look ahead in the chapter to preview covalent and ionic bonding.

Check Your Progress You'll be working on this project as you study this chapter. To keep your project on track, look for Check Your Progress boxes at the following points.

Section 1 Review, page 119: Build models of ionic compounds.

Section 2 Review, page 125: Build models of compounds that contain covalent bonds.

Present Your Project At the end of the chapter (page 133), you will present and explain your models to the class.

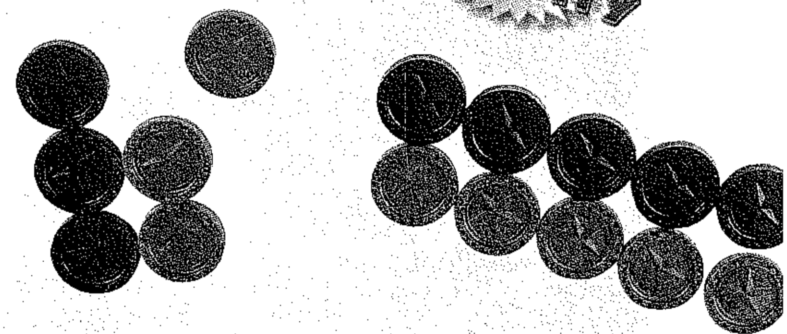
Chapter 4 - Physical Science

SECTION 1 Ionic Bonds

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

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

GUIDE FOR READING

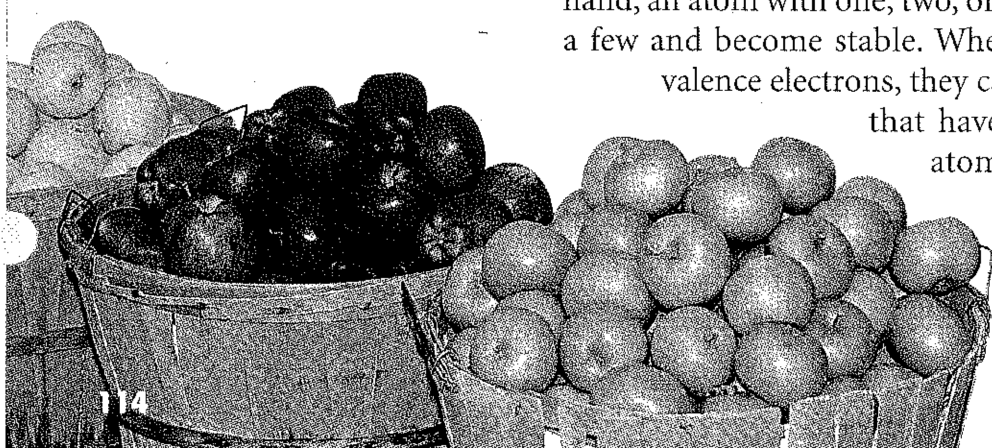
- **How does an atom become an ion?**
- **What are the properties of ionic compounds?**
- **How are the ions in an ionic compound held together?**

Reading Tip **As you read, make an outline describing the characteristics of compounds containing ionic bonds.**

Imagine you are walking down the street with your best friend. A market has a bin of apples for sale. A sign says that they cost 40 cents each. You both want an apple, but your friend has only 35 cents while you have 45 cents. What can you do? It doesn't take you long to figure out that if you give your friend a nickel, you can each buy an apple. Transferring the nickel to your friend gets both of you what you want. Your actions model, in a simple way, what can happen between atoms.

Electron Transfer

Like your friend with not quite enough money to buy an apple, an atom with five, six, or seven valence electrons has not quite enough to total the more stable number of eight. On the other hand, an atom with one, two, or three valence electrons can lose a few and become stable. When atoms have fewer than four valence electrons, they can transfer these to other atoms that have more than four. In this way, atoms either gain electrons or lose electrons, becoming more stable.



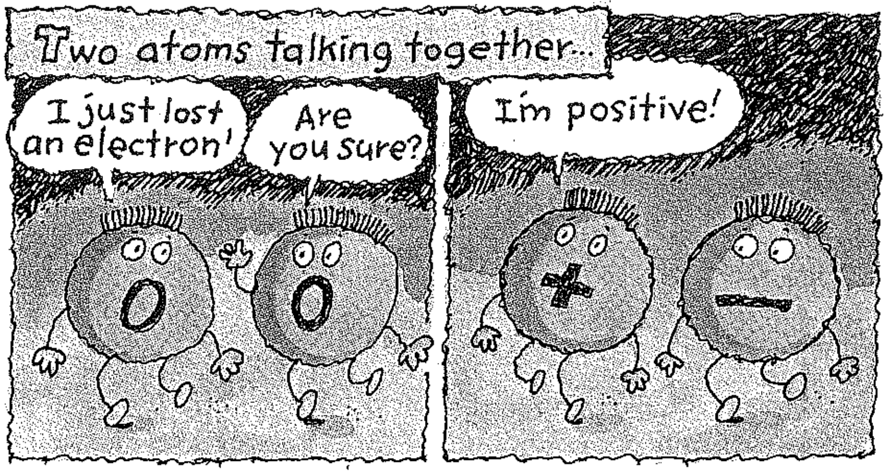


Figure 1 When an atom loses one of its negatively charged electrons, it becomes a positively charged ion.

An **ion** (EYE ahn) is an atom or group of atoms that has become electrically charged. **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.**

Forming an Ionic Bond

Consider what can happen to sodium and chlorine atoms. Note how many valence electrons each atom has. Suppose sodium's valence electron is transferred to chlorine. The transfer changes both atoms into ions. The sodium atom becomes a positive ion (Na^+). The chlorine atom becomes a negative ion (Cl^-). Negative and positive electric charges attract each other, so the oppositely charged Na^+ and Cl^- ions come together. They form sodium chloride, which you know as table salt.

$$Na\cdot + \cdot\ddot{\underset{\cdot\cdot}{Cl}}\colon \rightarrow Na^+\colon\ddot{\underset{\cdot\cdot}{Cl}}\colon^-$$

An **ionic bond** is the attraction between two oppositely charged ions. This attraction is similar to the attraction between the opposite poles of two magnets. When the two ions come together, the opposite charges cancel out. Every sodium ion (with a charge written as 1+) is balanced by a chloride ion (with a charge written as 1–). The formula for sodium chloride, NaCl, shows you this 1 : 1 ratio.

Compounds are electrically neutral. When ions come together, they do so in a way that balances out the charges on the ions. Figure 2 lists some common ions. Look at the charge of the magnesium ion. How many chloride ions would be needed to cancel out the 2+ charge of magnesium in the compound magnesium chloride? The formula for magnesium chloride, $MgCl_2$, tells you the answer is two. *Exploring Ionic Bonds* on page 116 reviews how ionic bonds are formed.

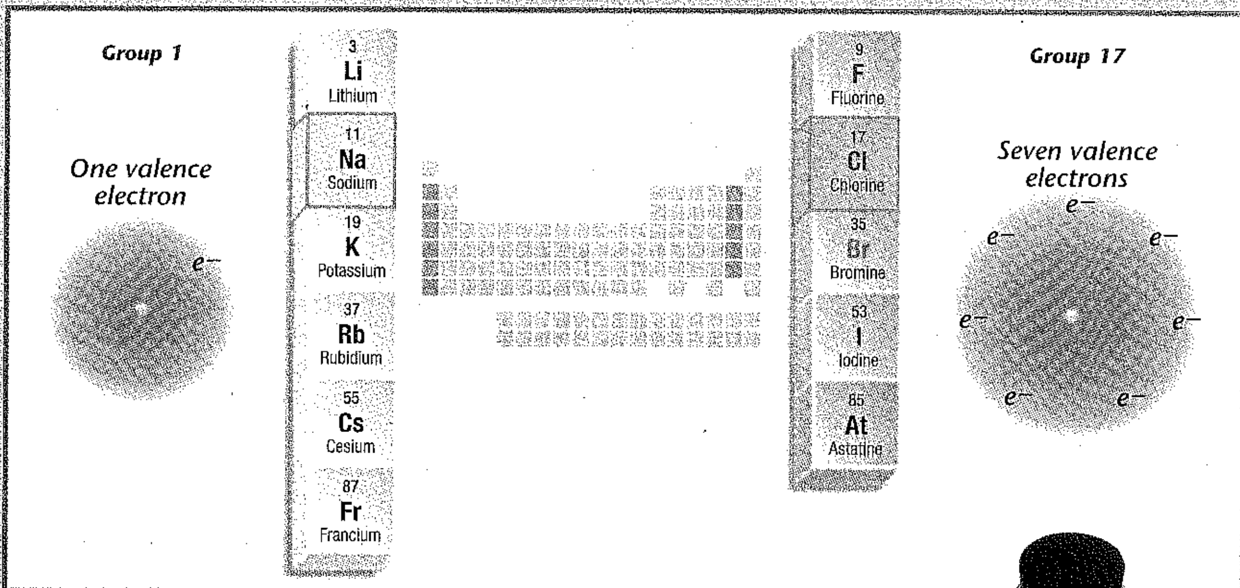
Figure 2 Positively charged ions have lost one or more electrons. Negatively charged ions have gained one or more electrons. ***Classifying*** *Which ions in the table are positively charged and which are negatively charged?*

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-}

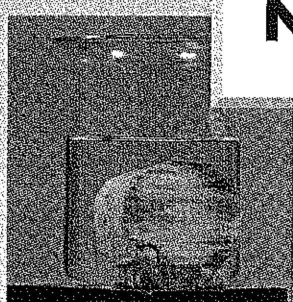
EXPLORING *Ionic Bonds*

Reactions between metals and nonmetals often form ionic compounds. These reactions occur easily between the metals in Group 1 and the halogens in Group 17. Review what happens when an ionic bond forms between a sodium atom and a chlorine atom.

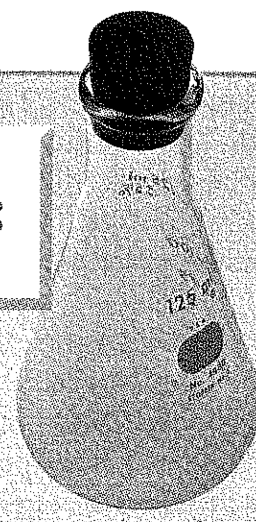


$\mathrm{Na}\cdot$ $\longrightarrow$ $\cdot\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}:$

One electron is transferred from a sodium atom to a chlorine atom.



Sodium metal



Chlorine gas

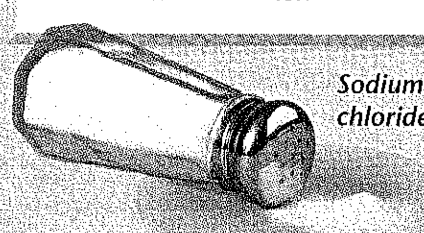
$\mathrm{Na}^{+}\ :\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}:^{-}$

Sodium ion — Chloride ion

A sodium ion has a 1+ charge.

A chloride ion has a 1− charge.

Oppositely charged ions attract each other. This attractive force is an ionic bond.



Sodium chloride

In an ionic compound, the positive ions balance the charge of the negative ions. Overall, the compound is electrically neutral.

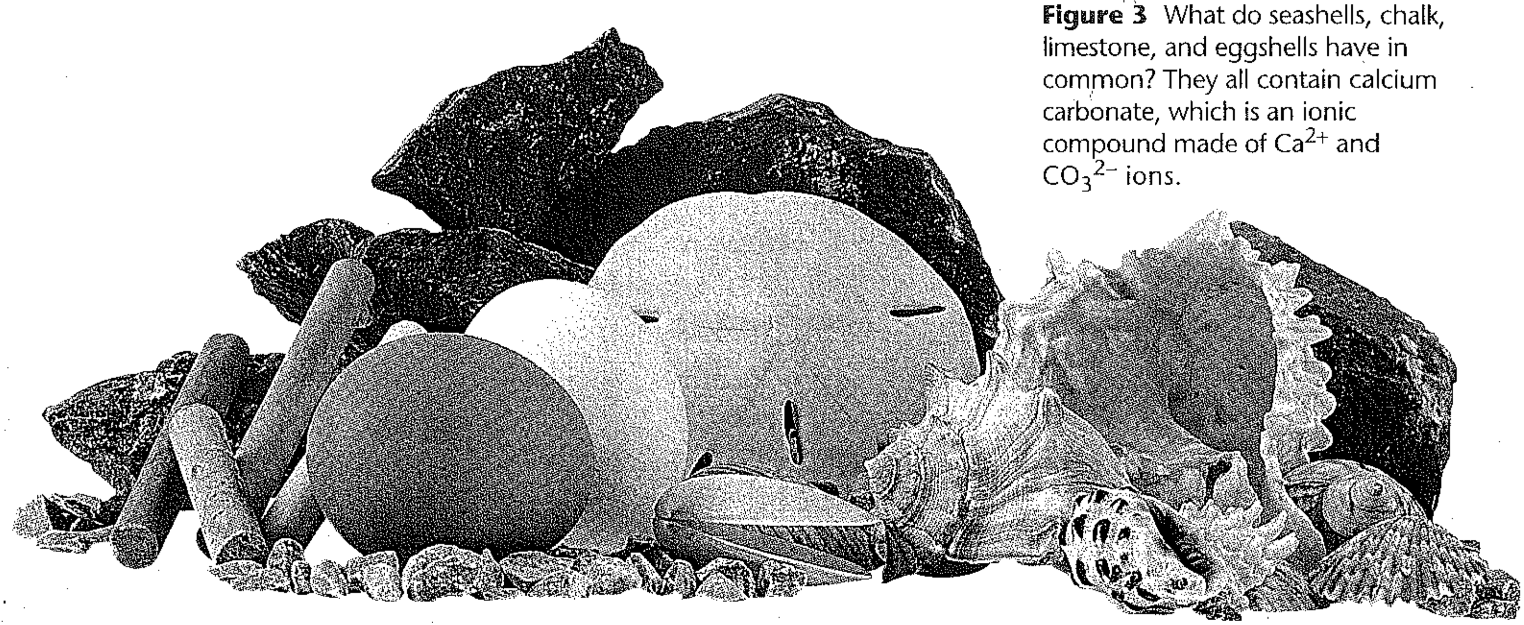


Figure 3 What do seashells, chalk, limestone, and eggshells have in common? They all contain calcium carbonate, which is an ionic compound made of Ca^{2+} and CO_3^{2-} ions.

Polyatomic Ions

Some ions are made of more than one atom. Ions that are made of more than one atom are examples of **polyatomic ions** (pahl ee uh TAHM ik). The prefix *poly* means "many," so the word *polyatomic* means "many atoms." You can think of a polyatomic ion as a group of atoms that react as one. Each polyatomic ion has an overall positive or negative charge. If a polyatomic ion combines with another ion of opposite charge, an ionic compound forms. Think, for example, about the carbonate ion (CO_3^{2-}). It is made of one carbon atom and three oxygen atoms and has an overall charge of 2–. This ion can combine with a calcium ion (Ca^{2+}) to form calcium carbonate ($CaCO_3$). Calcium carbonate is the main compound in limestone.

Naming Ionic Compounds

Calcium chloride, potassium iodide, 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. It may also be the name of a positive polyatomic ion, such as ammonium. If the negative ion is an element, the end of its name changes to *-ide*. For example, MgO is magnesium oxide. If the negative ion is polyatomic, its name is unchanged. For example, the chemical name for washing soda (Na_2CO_3) is sodium carbonate.

Checkpoint *What kind of atom has a name change when it becomes an ion?*

Sharpen your Skills

Interpreting Data

ACTIVITY

Look at the list of compounds below. Use the periodic table and Figure 2 to identify the charges of the ions in each compound. Then write the formula for each compound.

- sodium fluoride
- lithium oxide
- magnesium sulfide
- boron chloride
- aluminum sulfide

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

Try This

Crystal Shapes

ACTIVITY

Compare the shapes of crystals of different ionic compounds.

1. Use a spoon to place a small amount of halite crystals (NaCl) on a piece of black paper.
2. With a hand lens, carefully examine the structure of the crystals.
3. On a separate piece of paper, draw and label a picture of what you see.
4. Repeat Steps 1–3 with samples of other crystals provided by your teacher.

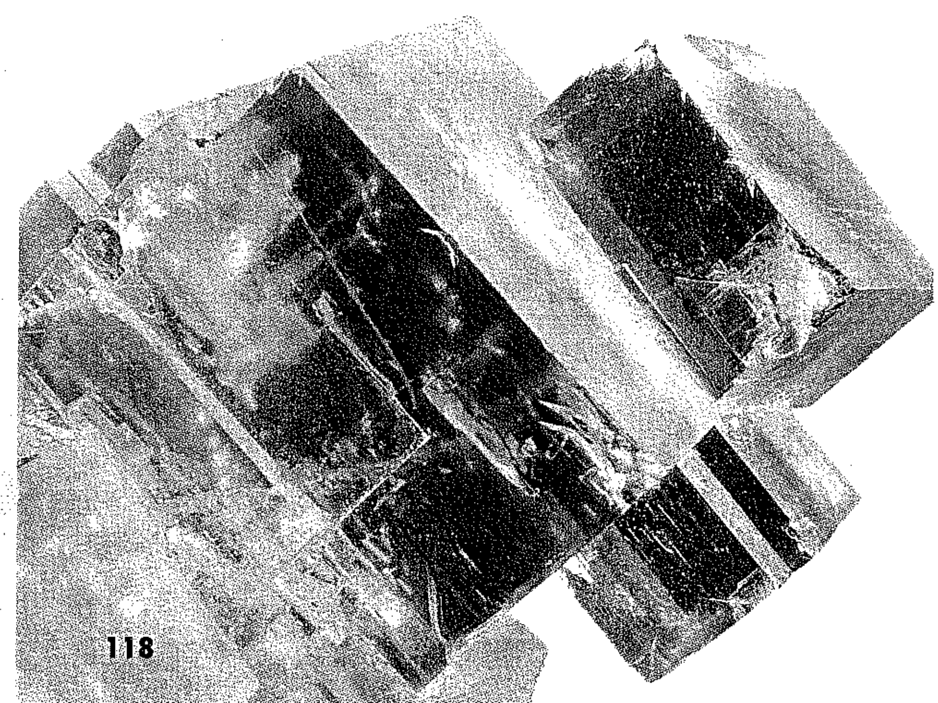
Observing Do the shapes of the crystals vary within a sample? Do the shapes vary from one sample to another? Explain.

Properties of Ionic Compounds

Do you think table salt, iron rust, baking soda, and limestone are very much alike? If you answer no, you're right. If you answer yes, you're right, too! You wouldn't want to season your food with rust, or construct a building out of baking soda. But despite their differences, these compounds share some similarities because they all contain ionic bonds. **The characteristic properties of ionic compounds include crystal shape, high melting points, and electrical conductivity.**

Crystal Shape The object in Figure 4 that looks like a glass sculpture is really a chunk of halite, or table salt. Halite is an ionic compound. All halite samples have sharp edges, corners, and flat surfaces. These properties result from how the ions are arranged. In solid sodium chloride, the Na^+ and Cl^- ions come together 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 near it that have an opposite charge. Positive ions tend to be near negative ions and farther from other positive ions. As a result, a positive sodium ion doesn't bond with just one negative chloride ion. It bonds with ions above, below, and to all sides. Because chloride ions bond with sodium ions in the same way, a crystal forms. This pattern continues no matter what the size of the crystal. In a single grain of salt, the crystal can extend for millions of ions in every direction. The number of sodium ions and chloride ions in the crystal is equal. The formula for sodium chloride, NaCl, represents this 1 : 1 ratio.



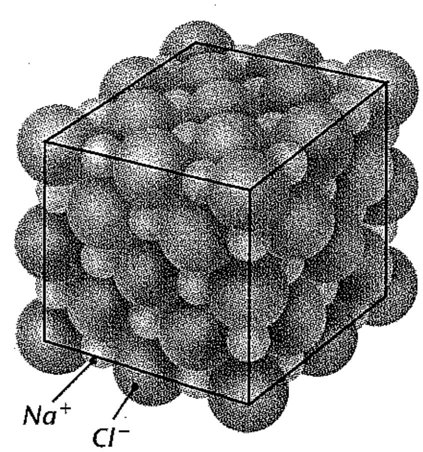


Figure 4 A halite crystal contains sodium and chloride ions in an alternating pattern.
Making Generalizations *What general characteristics of crystals can you observe in the photograph of halite?*

High Melting Points What happens when you heat an ionic compound such as table salt? Remember, the ions are held together in a crystal by attractions between oppositely charged particles. When the particles have enough energy to overcome the attractive forces between them, they break away from each other. It takes a temperature of 801°C to reach this energy for table salt. Ionic bonds are strong enough to cause all ionic compounds to be solids at room temperature.

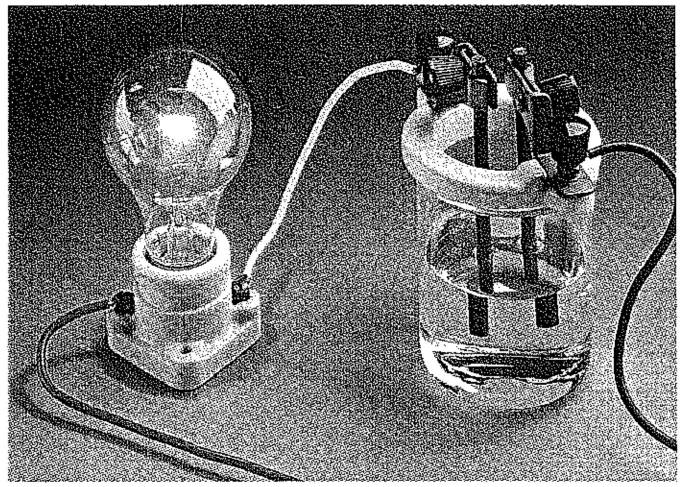


Figure 5 A conductivity tester shows that a solution of salt in water conducts electricity. The bulb lights up because the ions in the salt solution complete the circuit for the flow of electricity.

Electrical Conductivity When ionic compounds dissolve in water, the solution conducts electricity.

INTEGRATING PHYSICS

Electricity is the flow of electric charge, and ions have electric charges. However, if you connect wires from a salt crystal to a battery and a light bulb, don't expect anything to happen. A *solid* ionic compound does not conduct electricity very well. The ions in the crystal are tightly bound to each other. If the charged particles do not move, electricity does not flow. But what if the ions are broken apart? When ionic compounds dissolve in water, the ions separate. These ions then move freely, and the solution conducts electricity.

Melting ionic compounds also allows them to conduct electricity. Can you figure out why? Think about the difference between the particles in a solid and a liquid. In a solid, the particles do not move from place to place. But in a liquid, the particles slip and slide past each other. As long as the ions can move around, electricity can flow.

Section 1 Review

1. How does an ion form from an atom?
2. What properties may be used to identify ionic compounds?
3. Why are ions in an ionic compound attracted to each other?
4. Name these compounds: NaF, BeI_2, K_2SO_4, CaO, H_2S, $MgCO_3$.
5. Solid table salt does not conduct electricity. How does dissolving salt in water allow electricity to flow?
6. **Thinking Critically Problem Solving** The metal scandium (Sc) has three valence electrons. What is the formula of the ionic compound formed when scandium reacts with iodine?

CHAPTER PROJECT

Check Your Progress

Use your materials to make models of compounds containing ionic bonds, such as sodium chloride (NaCl), magnesium chloride ($MgCl_2$), or potassium oxide (K_2O). (*Hint:* Figure out whether each atom forms a positive or negative ion. Then use combinations that result in a neutral compound.)

Skills Lab

Drawing Conclusions

Shape Up!

As an ionic solid, table salt—sodium chloride—is a crystal. In this lab, you will investigate the structure of that crystal.

Problem

What is the structure of sodium chloride crystals?

Materials

sodium chloride	100-mL graduated cylinder
plastic spoon, small	250-ml beaker
Petri dishes, 2	water (80°C)
black paper	hand lens

Procedure

1. Read the procedure, and write a prediction about the appearance of the crystals in Step 7. Then create a data table like the one above.
2. Your teacher will give you a sample of sodium chloride. Use a spoon to sprinkle some grains of the salt in a Petri dish. Put the dish on black paper. Observe the salt with a hand lens. Record your observations, and draw what you see.
3. Pour 100 mL of hot water (80°C) into a 250-mL beaker. Add about half a spoonful of salt to the water and stir until it dissolves. Observe the solution, and record your observations.
4. Add the rest of the salt and stir well. Let any undissolved salt particles settle at the bottom of the beaker.
5. Carefully pour off 50 mL of the solution into a Petri dish labeled with your name.
6. Let the uncovered Petri dish sit overnight, or until all the water has evaporated.
7. Put the Petri dish on black paper and observe the new crystals with a hand lens. Record your observations, and draw what you see.

DATA TABLE

Substance	Observations
Original NaCl crystals	
NaCl solution	
New NaCl crystals	

Analyze and Conclude

1. How did the appearance of the sodium chloride change between Steps 2 and 3? Explain that change in terms of the ions that make up the compound.
2. Describe the shapes of the crystals you observed in Step 7, along with any patterns you see within any of the crystals.
3. Compare and contrast the crystals you observed in Step 2 and Step 7.
4. Use what you know about ionic compounds to explain the shapes of the new crystals. Would you expect all sodium chloride crystals to have the same shapes? Explain.
5. **Think About It** Could you use the results of this experiment to draw conclusions about the crystals formed by other ionic compounds? Explain.

Design an Experiment

Does the original temperature of the sodium chloride solution affect crystal formation? Design a safe experiment comparing solutions of ice water, tap water, and hot water. Obtain your teacher's approval before trying this experiment.

Section 2 Covalent Bonds

Discover Activity

Why Don't Water and Oil Mix?

1. Pour water into a small jar that has a tight-fitting top until the jar is about a third full.
2. Add an equal amount of vegetable oil to the water and cover the jar tightly.
3. Shake the jar with vigor for about 15–20 seconds. Observe.
4. Allow the jar to sit undisturbed for about 1 minute. Observe again.
5. Remove the top and add 2–3 drops of liquid soap. Repeat Steps 3 and 4.

Think It Over

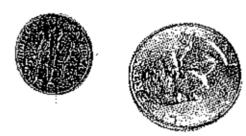
Inferring Describe how adding soap affected the mixing of the oil and water. How might what you observed depend on chemical bonds in the soap, oil, and water molecules?

Remember the market with apples selling for 40 cents each? On another day, the apples are put on sale at two for 70 cents. You and your friend check your pockets and find 35 cents each. What can you do? You could give your friend a nickel to make enough money for one apple. Then you would have only 30 cents, not enough to get one for yourself. But if you share your money, together you can buy two apples.

Guide for Reading

- What happens to electrons in a covalent bond?
- Why do some atoms in covalent bonds have slight negative or positive charges?
- How are polar and nonpolar compounds different?

Reading Tip Before you read, preview the illustrations in the section. Predict how covalent bonds differ from ionic bonds.



Electron Sharing

Just as you and your friend can buy apples by sharing money, atoms can become more stable by sharing valence electrons. A chemical bond formed when two atoms share electrons is called a **covalent bond.**

Unlike ionic bonds, which form between metals and nonmetals, covalent bonds often form between two or more nonmetals. Oxygen, carbon, nitrogen, and the halogens are examples of elements that frequently bond to other nonmetals by sharing electrons.

The element fluorine forms molecules made of two fluorine atoms. Each fluorine atom shares one of its seven valence electrons with the other atom. When you count the number of electrons on one atom, you count the shared pair each time. By sharing, both atoms have eight valence electrons. **In a covalent bond, both atoms attract the two shared electrons at the same time.**

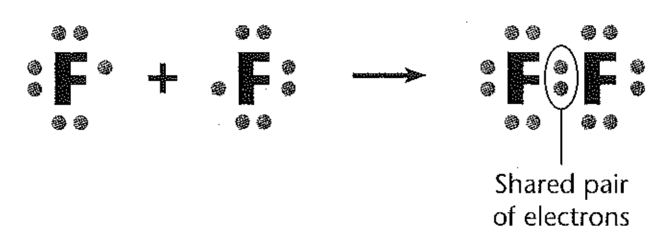
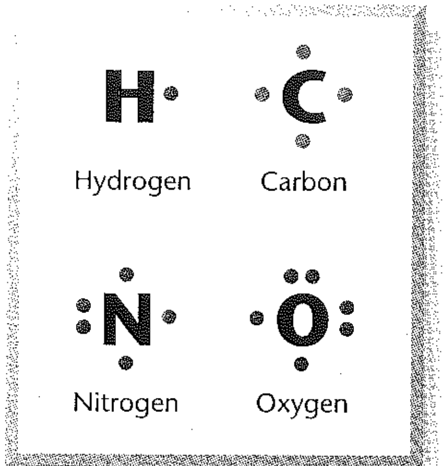


Figure 6 The shared pair of electrons in a molecule of fluorine is a single covalent bond.



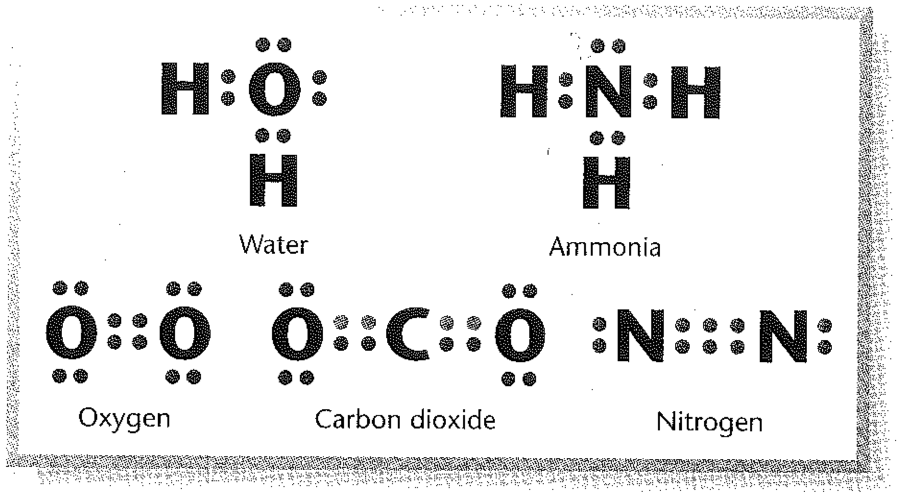


Figure 7 The electron dot diagrams for hydrogen, carbon, nitrogen, and oxygen (left) show the number of valence electrons for each. The diagrams of molecules (right) show how the electrons are shared in covalent bonds.
Interpreting Diagrams *How many bonds does each nitrogen atom form?*

How Many Bonds?

Look at the electron dot diagrams for oxygen, nitrogen, and carbon atoms in Figure 7. Count the dots on each atom. The number of bonds these atoms can form equals the number of valence 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 hydrogen atom. Since nitrogen has five valence electrons, it can form three bonds. In ammonia (NH_3), a nitrogen atom bonds with three hydrogen atoms.

Next, compare water to a molecule of oxygen. Can you find the two covalent bonds? This time *two* pairs of electrons are shared between the oxygen atoms, forming a **double bond.** In a carbon dioxide molecule, carbon forms a double bond with each oxygen atom. Elements such as nitrogen and carbon can even form triple bonds in which *three* pairs of electrons are shared.

Count the electrons around any atom in the molecules in Figure 7. Remember that shared pairs count for both atoms forming a bond. You'll find that each atom has eight valence electrons. The exception is hydrogen, which can have no more than two electrons and forms one bond.

Figure 8 Molecular compounds have much lower melting points than ionic compounds.

Melting and Boiling Points of Some Molecular Compounds

Compound	Formula	Melting Point (°C)	Boiling Point (°C)
Water	H_2O	0	100
Methane	CH_4	–182	–164
Carbon dioxide	CO_2	—	–78.6*
Ammonia	NH_3	–77.7	–33.6
Rubbing alcohol	C_3H_7OH	–89.5	82.4
Sugar	$C_{12}H_{22}O_{11}$	185–186	(decomposes)

*Carbon dioxide changes directly from a solid to a gas.

Properties of Molecular Compounds

Molecular compounds consist of molecules having covalently bonded atoms. Such compounds have very different properties from ionic compounds.

Look at Figure 8, which lists the melting and boiling points for some molecular compounds. Quite a difference from the

801°C and 1,413°C described for table salt! In molecular solids, the molecules are held close to each other. But the forces holding them are much weaker than those holding ions together in an ionic solid. Less heat is needed to separate molecules than is needed to separate ions. Some molecular compounds, such as sugar and water, do form crystals. But these compounds, like other molecular solids, melt and boil at much lower temperatures than ionic compounds do.

Most molecular compounds are poor conductors of electricity. No charged particles are available to move, and electricity does not flow. That's why molecular compounds, such as plastic and rubber, are used to insulate electric wires. Even as liquids, molecular compounds are poor conductors. Pure water, for example, does not conduct electricity. Neither does water with sugar dissolved in it.

Checkpoint *Why are molecular compounds poor conductors?*

Sharpen your Skills

Designing Experiments

ACTIVITY

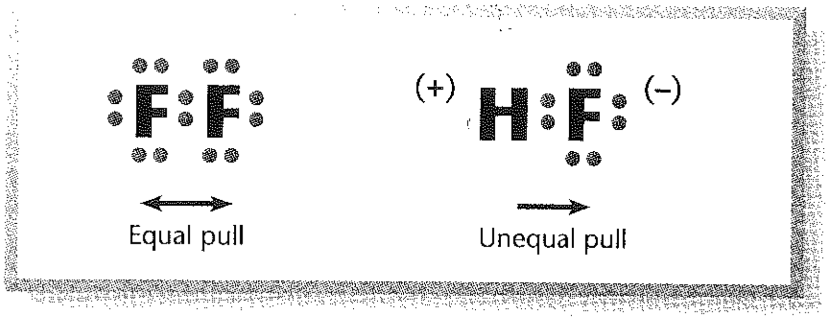
Suppose you have samples of two colorless, odorless gases. You are told that one gas is methane (CH_4) and the other is carbon dioxide (CO_2). How could you use the information in Figure 8 to find out which gas is which? Describe the experiment you would set up. Tell what conditions you would control and what you would change. What result would you look for to get an answer?

Unequal Sharing of Electrons

Have you ever played tug of war? If you have, you know that if both teams have equal strength, the contest is a tie. But what if the teams pull on the rope with unequal force? Then the rope moves closer to one side or the other. The same is true of electrons in a covalent bond. **Some atoms pull more strongly on the shared electrons than other atoms do. As a result, the electrons move closer to one atom, causing the atoms to have slight electrical charges.** These charges are not as strong as the charges on ions. But the unequal sharing is enough to make one atom slightly negative and the other atom slightly positive. A covalent bond in which electrons are shared unequally is **polar.**

Figure 9 The unequal sharing of the electrons in a polar covalent bond is like a tug of war in which one atom is slightly stronger than the other atom.

Figure 10 In the nonpolar bond in F_2, the two fluorine atoms pull equally on the shared electrons. In the polar bond in HF, fluorine pulls more strongly on the shared electrons than hydrogen does.



If two atoms pull equally on the electrons, neither atom becomes charged. This is the case when the two atoms are identical, as in fluorine gas (F_2). The valence electrons are shared equally and the bond is **nonpolar.** Compare the bond in F_2 with the polar bond in hydrogen fluoride (HF) in Figure 10.

Nonpolar Molecules Keep tug of war in mind as you look at the carbon dioxide (CO_2) molecule in Figure 11. Oxygen attracts electrons much more strongly than carbon, so bonds between oxygen and carbon are polar. But the two oxygen atoms are pulling with equal strength in opposite directions. In a sense, they cancel each other out. Overall, a carbon dioxide molecule is nonpolar even though it has polar bonds. A molecule is nonpolar if it contains polar bonds that cancel each other. As you might guess, molecules that contain only nonpolar bonds are also nonpolar.

Polar Molecules Water molecules are polar. As you can see in Figure 11, the shape of the molecule leaves the two hydrogen atoms more to one end and the oxygen atom toward the other. The oxygen atom pulls electrons closer to it from both hydrogen atoms. Overall, the molecule is polar. It has a slightly negative charge at the oxygen end and a slightly positive charge near the hydrogen atoms.

☑ ***Checkpoint*** *What makes a covalent bond polar?*

Language Arts CONNECTION

Breaking a word into its parts can help you understand its meaning. Take *covalent,* for example. The prefix *co-* means "together." The *-valent* part comes from "valence electrons." So "valence electrons together" can remind you that in a covalent bond, valence electrons are shared.

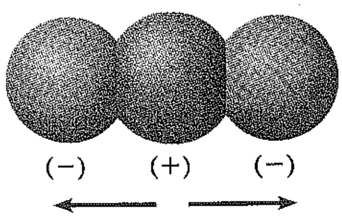
In Your Journal

The prefix *co-* is used in many other words—*coauthor, coexist,* and *cooperate* are just a few. Add five more *co-* words to this list and try to define them all without looking them up. Then check their meanings in a dictionary and write sentences that use each one.

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 and positively charged hydrogen ends behave like poles of a bar magnet. They attract the opposite ends of other water molecules. These attractions between positive and negative ends pull water molecules toward each other.

What about carbon dioxide? There is no pulling between these molecules. Remember, carbon dioxide molecules are nonpolar. No oppositely charged ends means there are no strong attractions between the molecules.



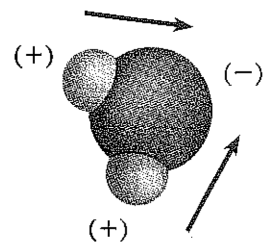
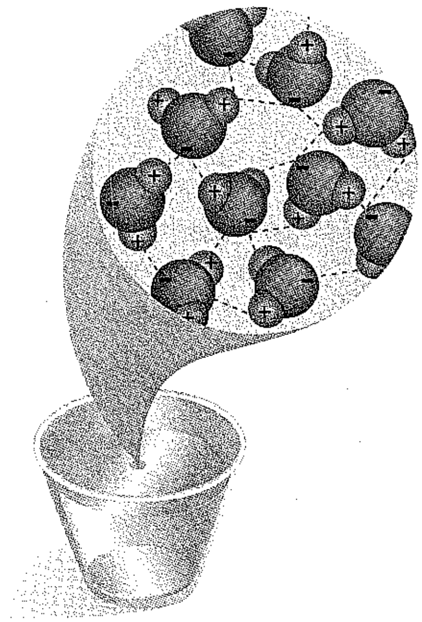


Figure 11 CO_2 molecules are nonpolar, and H_2O molecules are polar. Attractions between the slightly positive and slightly negative ends of water molecules pull the molecules toward each other (below).



Differences in the attractions between molecules lead to different properties in polar and nonpolar compounds. For example, water and vegetable oil don't mix. Oil is nonpolar, and nonpolar compounds do not dissolve well in water. The polar water molecules are attracted more strongly to each other than to the molecules of oil. Water stays with water and oil stays with oil.

INTEGRATING TECHNOLOGY These differences in attractions come in handy when you wash laundry. Many kinds of dirt—for example, grease—are nonpolar compounds. Their molecules won't mix with plain water. So how can you wash dirt out of your clothes?

As you found if you did the Discover activity, adding soap helped the oil and water to mix. When you do laundry, detergent causes the nonpolar dirt to mix with the polar water. Soaps and detergents have long molecules. One end of a soap molecule is polar, and the other end is nonpolar. Soaps and detergents dissolve in water because the polar ends of their molecules are attracted to water molecules. Meanwhile, their nonpolar ends mix easily with the dirt. When the water washes down the drain, the soap and the dirt go with it.

Section 2 Review

1. How are valence electrons involved in the formation of a covalent bond?
2. How do atoms in covalent bonds become slightly negative or slightly positive?
3. Explain how attractions between molecules could cause water to have a higher boiling point than carbon dioxide.
4. **Thinking Critically Comparing and Contrasting** In terms of electrons, how is a covalent bond different from an ionic bond?

CHAPTER PROJECT

Check Your Progress

Use your materials to build molecules with single covalent bonds. Also make models of molecules containing double or triple bonds. (*Hint:* After you make bonds, each atom should have a total of eight valence electrons or, in the case of hydrogen, two valence electrons.)

Skills Lab

Interpreting Data

Shedding Light on Chemical Bonds

Electricity is the flow of electric charges. In this lab, you will interpret data about which compounds conduct electricity in order to determine the nature of their bonds.

Problem

How can you use a conductivity tester to determine whether a compound contains ionic or covalent bonds?

Materials

2 dry cells, 1.5-V
small light bulb and socket
4 lengths of wire with insulation scraped off the ends
} or conductivity probe
small beaker
small plastic spoon
sodium chloride
100-mL graduated cylinder
additional substances supplied by your teacher

DATA TABLE

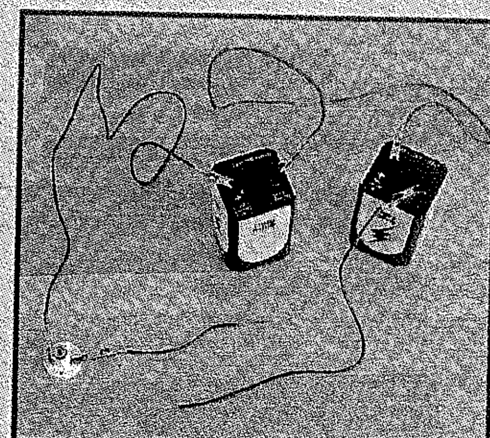
Sample	Observations
Water	
Sodium chloride in water	

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 water into a small beaker. Place the free ends of the two wires of the conductivity tester into the water. Be sure the ends are close but not touching each other. Record your observations.

Making a Conductivity Tester

A. Use wire to connect the positive terminal of a dry cell to a lamp socket. **CAUTION:** *The bulb is fragile and can break.*
B. Use another wire to connect the negative terminal to the positive terminal of a second dry cell.
C. Connect a third wire to the negative terminal of the second dry cell.
D. Connect a fourth wire to the other terminal of the lamp socket.



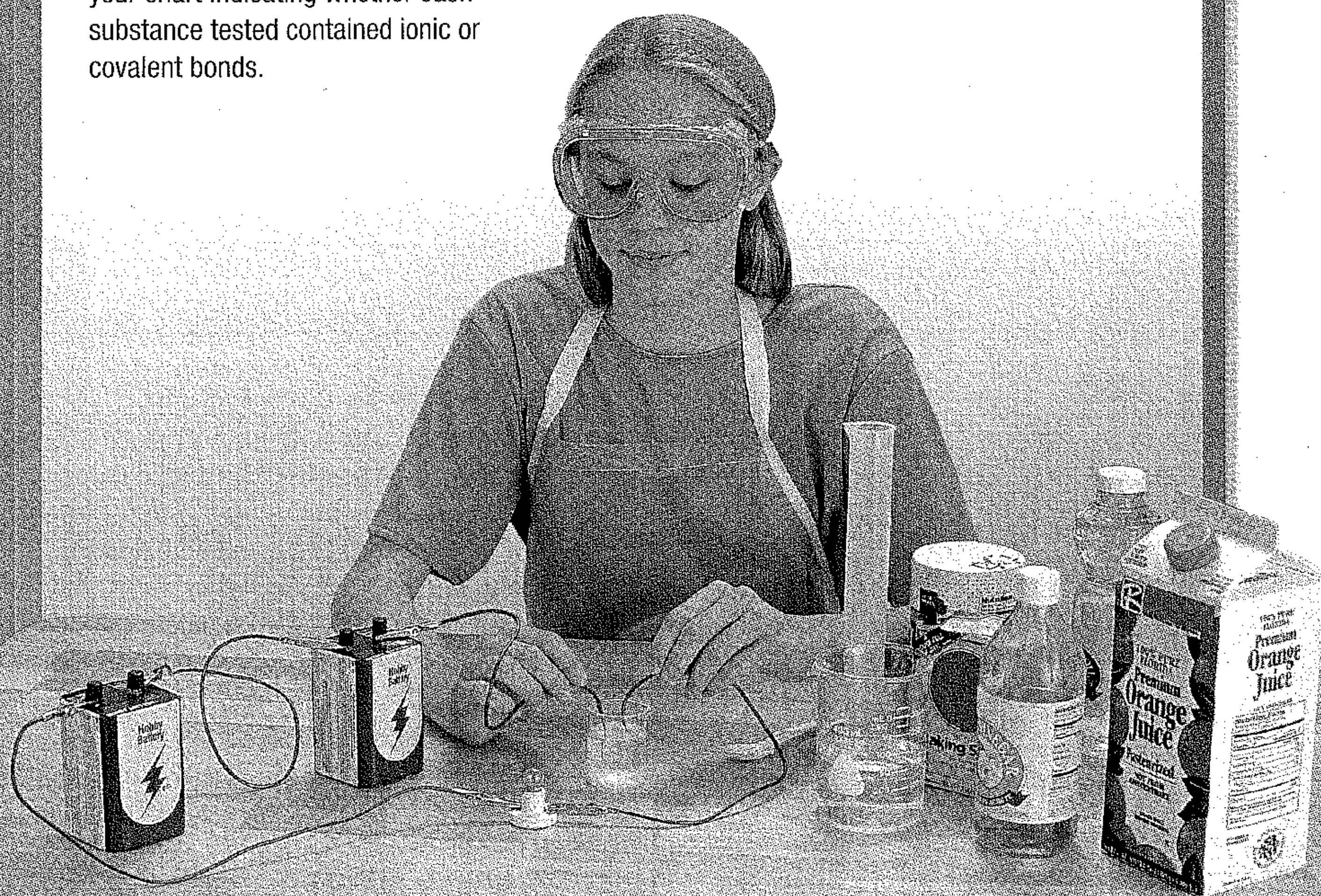
3. Remove the conductivity tester and add a small sample (about 3 spoonfuls) of sodium chloride to the water in the small beaker. Stir with the spoon until mixed.
4. Repeat the conductivity test and record your observations in your data table.
5. Rinse the beaker, spoon, and wire ends with clear water. Then repeat Steps 3 and 4 for each substance provided by your teacher.
   - If the substance is a solid, mix 1 to 3 spoonfuls of it with about 50 mL of fresh water. Test the resulting mixture.
   - If the substance is a liquid, simply pour about 50 mL into the beaker. Test it as you did the solutions.

Analyze and Conclude

1. Why did you test plain water first?
2. Based on your observations, add a column to your chart indicating whether each substance tested contained ionic or covalent bonds.
3. Explain why one substance is a better conductor of electricity than another.
4. Did all the substances that conducted electricity show the same amount of conductivity? How do you know?
5. **Think About It** How might varying the amount of each substance added to the water have affected your results? How could you better control the amount of each substance?

Design an Experiment

Design another experiment to compare a different property of compounds containing ionic and covalent bonds. You might want to examine properties such as the ability to dissolve in water or in some other liquid. Present your experimental plan to your teacher before proceeding.



SECTION 3 Crystal Chemistry

DISCOVER ACTIVITY

How Small Do They Get?

1. Place a piece of rock salt on a hard surface. Make a rough sketch of the shape of your sample.
2. Put on your goggles. Cover the salt with a paper towel. Then break the salt sample into several smaller pieces with the back of a metal spoon.
3. Look at these smaller pieces with a hand lens. Then draw a picture of the shapes you see.
4. Crush a few of these smaller pieces with the spoon. Repeat Step 3.

Think It Over

Predicting What would the crystals look like if you crushed them into such small pieces that you needed a microscope to see them?

GUIDE FOR READING

- How are the properties of a mineral related to chemical bonds?

Reading Tip As you read, make a list of the ways in which a mineral can be described or identified.

A class of earth science students gathers rock samples on a field trip. They want to know whether the rocks contain any of the minerals they have been studying. The teacher takes a hammer and strikes one rock. It cracks open to reveal a few small crystals peeking out of the new surface. The crystals are mostly the same shape and have a metallic shine. The teacher tries to scratch one crystal, first with her fingernail and then with a copper penny. Only the penny leaves a mark. By now, the students have enough information to make an inference about the identity of the crystals. They'll do more tests back in their classroom to be sure their inference is correct.

Mineral Properties

A **mineral** is a naturally occurring solid that has a crystal structure and a definite chemical composition. A few minerals, such as sulfur and gold, are elements. But most minerals are compounds.

Mineralogists, scientists who study minerals, identify minerals by looking at certain properties. These properties include color, shininess, density, crystal shape, hardness, and magnetism. Color and shininess can be seen just by looking at a mineral. Other properties, however, require measurements or testing. For example, scientists rate a mineral's hardness by comparing it with something harder or softer. You can scratch the softest mineral, talc, with your fingernail. Diamond is the hardest mineral. Other minerals are somewhere in between.

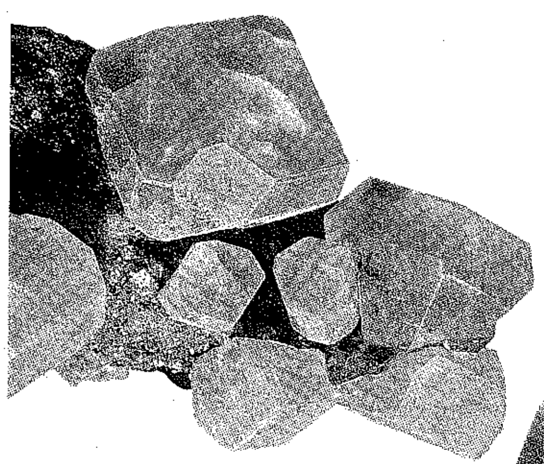
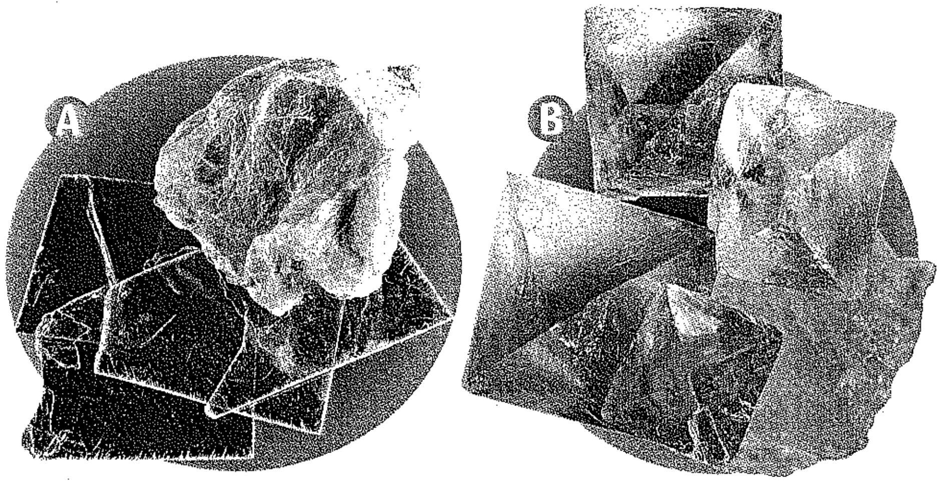
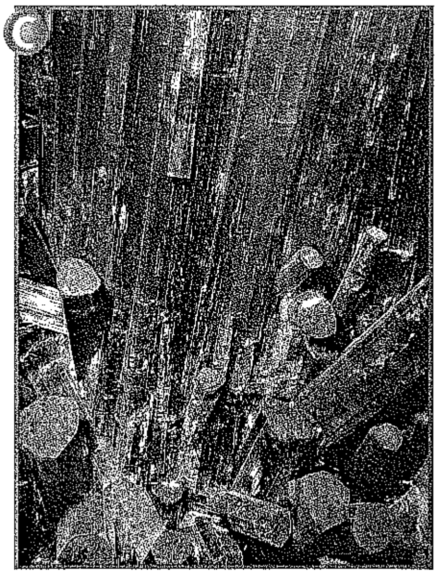


Figure 12 The mineral sulfur (above) is a pure element. The mineral galena (right) is a compound of sulfur and lead.





Another key property is the way a mineral breaks apart. Some minerals break into regular shapes. Mica, for example, splits easily along flat surfaces and at sharp angles. Crystals also grow in characteristic shapes. All the properties of a mineral depend on its chemical composition. Since each mineral has a different composition, its properties will not be exactly like those of any other mineral.

Figure 13 Mica's flakes **(A)** are a result of how the mineral splits when it breaks. The crystals of fluorite **(B)** and tourmaline **(C)** grew into the shapes you see. ***Observing*** *How do the shapes of fluorite crystals and tourmaline crystals differ?*

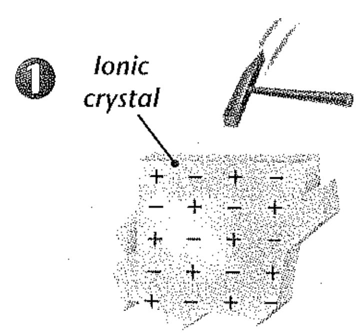
☑ *Checkpoint* *What is a mineral?*

Bonding in Mineral Crystals

Every mineral has a crystal structure. The repeating pattern of particles creates a shape that may be visible to your eye. Or, you may have to look under a microscope to see it. Either way, the structure of the crystal is a characteristic property of the mineral.

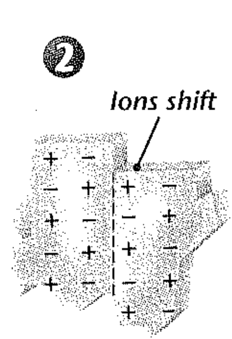
Mineral crystals may be made of ions, or they may contain atoms that are covalently bonded together. **The arrangement of particles in a mineral and the kind of bonds holding them together determine properties such as crystal shape, hardness, and the way the crystal breaks apart.**

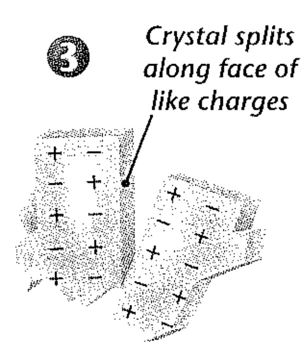
Figure 14 The particles in an ionic crystal such as halite can shift because of a blow or pressure.



An Ionic Crystal In Section 1, you read about halite, a mineral made of sodium chloride (NaCl). You can easily scratch halite with a steel knife. If you put a crystal of halite into water, it would dissolve. The oppositely charged sodium and chloride ions in a halite crystal alternate in every direction, making a pattern something like a three-dimensional checkerboard. This arrangement affects the shape into which halite crystals grow.

If you break a piece of halite, the smaller pieces of halite have the same shape as the bigger piece. When bonds in an ionic crystal break, they break along a face of ions. A blow or crushing action shifts the ions slightly so that positive ions are next to other positive ions and negative ions are next to other negative ions. The effect is the same as bringing the





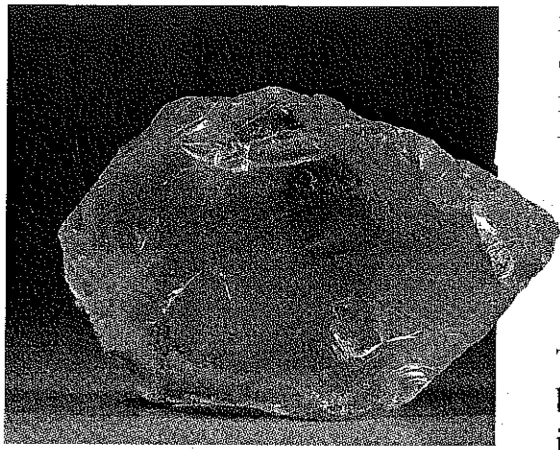


Figure 15 The uneven surfaces on this crystal are typical of broken quartz.
Comparing and Contrasting How does the way quartz breaks compare to the way mica breaks?

north ends of two magnets together. It creates a weakness in the crystal. The ions push each other away, breaking bonds along a flat surface or face. The result is that the smaller crystals retain the cube shape that is characteristic of halite.

A Covalent Crystal If you picked up a handful of sand, most likely you would be holding some quartz. Quartz is a compound made of silicon and oxygen atoms covalently bonded together to form the compound silicon dioxide (SiO_2). The covalent bonds in quartz are much stronger than the ionic bonds in halite. Quartz won't dissolve in water. You can't scratch it with a knife. In fact, you could use quartz to scratch steel! Because of its strong bonds, a quartz crystal doesn't have clear lines of weakness. You can't crush it into predictable shapes with a hammer. Instead, it breaks into smaller pieces with irregular shapes. The broken surfaces have shell-like ridges similar to chipped glass. These features help identify the mineral as quartz.

Comparing Crystals

Not all mineral crystals made of ions have the same properties as halite. Similarly, not all minerals made of covalently bonded atoms are like quartz. Properties such as hardness, for example, depend on the strength of the bonds in a crystal. The stronger bonds of quartz make it harder than halite. But other crystals with covalently bonded atoms are stronger than quartz. Still others have weaknesses in their bonds that cause the minerals to break apart the same way every time.

Experienced mineralogists can usually identify a mineral just by looking at it. But when there is a question, they test the sample for characteristics such as hardness and the way the crystals break. The results give the answer.

Section 3 Review

1. Name two properties of minerals that depend on chemical bonds.
2. What property of a mineral can be determined by scratching it?
3. How does the way in which a mineral crystal breaks apart help to identify it?
4. **Thinking Critically Comparing and Contrasting** Name three ways in which a halite crystal differs from a quartz crystal.

Science at Home

Piling Up Construct a model of an ionic crystal. Place round objects of two different sizes (such as balls of clay) in a checkerboard pattern to make the first layer. Now place one smaller object on top of each larger one and vice versa to make the second layer. Continue until the first layer is completely covered. Construct a third layer in a similar way. Explain to your family how your model represents an ionic crystal.

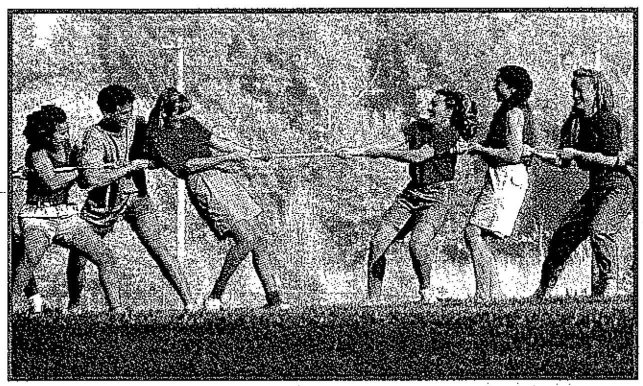
SECTION 1 Ionic Bonds

Key Ideas

- When atoms lose or gain electrons, they become positively or negatively charged ions.
- In an ionic compound, the ions are arranged in a three-dimensional structure called a crystal. Each ion in the crystal is attracted to nearby ions of opposite charge.
- Characteristic properties of ionic compounds include crystal shape, high melting points, and electrical conductivity.

Key Terms

ion
ionic bond
polyatomic ion
crystal



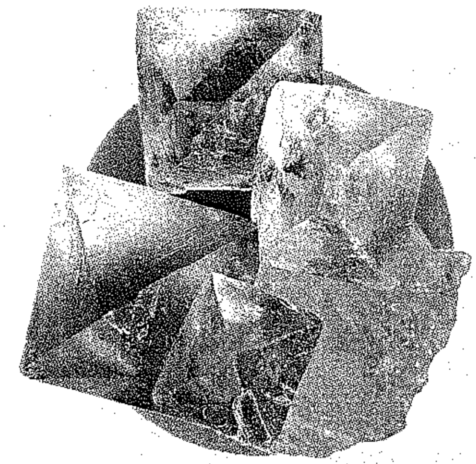
SECTION 2 Covalent Bonds

Key Ideas

- In covalent bonds, pairs of electrons are shared between atoms.
- In polar covalent bonds, the shared electrons are attracted more to one atom than the other.
- Attractions between polar molecules are stronger than attractions between nonpolar molecules, leading to differences in properties.

Key Terms

covalent bond
double bond
molecular compound
polar
nonpolar



SECTION 3 Crystal Chemistry

INTEGRATING EARTH SCIENCE

Key Ideas

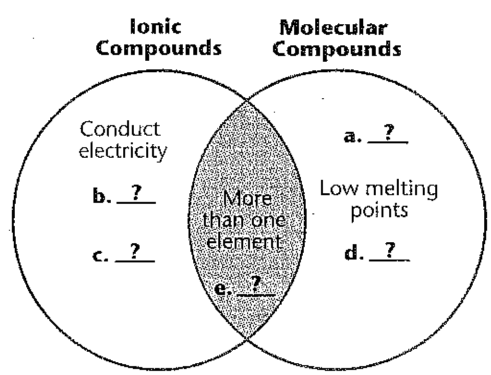
- Minerals have characteristic properties, such as hardness, density, color, crystal shape, and the way the crystal breaks.
- The properties of a mineral depend on its chemical composition and its bonding. Mineral crystals may contain ions or covalently bonded molecules.
- The stronger the chemical bonds in a mineral crystal, the harder the crystal is.

Key Term

mineral

Organizing Information

Venn Diagram Copy the Venn diagram comparing ionic and molecular compounds onto a separate sheet of paper. Then complete the diagram and add a title. (For more on Venn diagrams, see the Skills Handbook.)



Reviewing Content

For more review of key concepts, see the Interactive Student Tutorial CD-ROM.

Multiple Choice

Choose the letter of the best answer.

1. When an atom loses an electron, it
a. becomes a negative ion.
b. becomes a positive ion.
c. forms a covalent bond.
d. gains protons.

2. Which of these is a property of an ionic compound?
a. low melting point
b. poor conductor of electricity
c. crystal shape
d. shared electrons

3. A bond in which a pair of electrons is shared between two atoms is called
a. ionic. **b.** covalent.
c. polyatomic. **d.** triple.

4. Fluorine atoms cannot form a double or triple bond because fluorine
a. is a nonmetal.
b. strongly attracts electrons.
c. has 7 valence electrons.
d. is a gas at room temperature.

5. An ionic crystal splits along a face of
a. like charges. **b.** molecules.
c. unlike charges. **d.** random ions.

True or False

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

6. When a chlorine atom gains an electron, it becomes a <u>positive ion</u>.

7. When atoms share electrons unequally, a <u>polar</u> bond forms.

8. The attractions between polar molecules are <u>weaker</u> than those between nonpolar molecules.

9. <u>Hardness</u> is determined by how easily a mineral can be scratched.

10. The bonds in halite are <u>weaker</u> than the bonds in quartz.

Checking Concepts

11. Explain how the number of valence electrons an element has affects the type and number of bonds it can form.

12. Use the periodic table to identify what type of chemical bond is involved in each of these compounds: NaF, NO_2, CBr_4, MgS. Explain your reasoning.

13. How is a covalent bond between two atoms affected when each atom attracts electrons equally?

14. Of all the elements, fluorine atoms attract electrons most strongly. When fluorine atoms form covalent bonds with other kinds of atoms, are the bonds polar or nonpolar?

15. Writing to Learn Imagine you are a chlorine atom. Write a first-person description of the changes you undergo when forming an ionic bond with sodium. Compare these with what happens when you form a covalent bond with another chlorine atom.

Thinking Critically

16. Applying Concepts Use the periodic table to find the number of valence electrons for calcium (Ca), aluminum (Al), rubidium (Rb), oxygen (O), sulfur (S), and iodine (I).Then use that information to predict the formula for each of the following compounds: calcium oxide, aluminum iodide, rubidium sulfide, and aluminum oxide.

17. Inferring Element Z is a yellow solid that melts at about 100°C and does not conduct electricity. What type of bond holds the element's atoms together? Explain the reasoning for your answer.

18. Relating Cause and Effect Explain why a solid ionic compound does not conduct electricity, but the compound will do so when melted or dissolved in water.

19. Problem Solving Suppose you were given two mineral crystals that looked alike. How would you determine the identity of each?

Applying Skills

Element X exists as a nonpolar molecule made of two identical atoms. When individual atoms of element X react with sodium, they form ions with a 2– charge. Use the periodic table in Appendix D to answer Questions 20–24.

20. **Classifying** To what group of elements does element *X* belong?
21. **Inferring** How many valence electrons does an atom of element *X* have?
22. **Predicting** Sodium can react with element X to form a compound. How many atoms of sodium are needed for each atom of element X? Write the formula for the compound.
23. **Calculating** How many covalent bonds can element *X* form?
24. **Posing Questions** In order to identify element *X*, what additional questions would you need to ask?

Performance Assessment (Chapter Project)

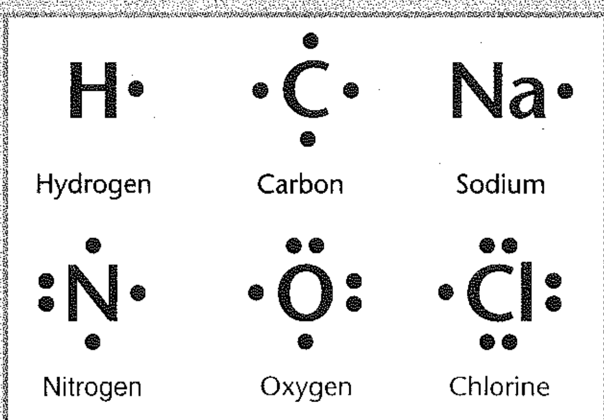
Present Your Project Before you present your models to the class, use an index card for each model to make a key telling what each part of the model represents. Explain why you chose particular items to model the atoms and the chemical bonds. How are your models alike or different from models of the same compounds made by other students?

Reflect and Record In your journal, compare your models containing ionic bonds with those containing covalent bonds. Which were easier to show? Why? What more would you like to know about bonding that could help improve your models?

Test Preparation

Use these questions to prepare for standardized tests.

Use the diagram to answer Questions 25–28.



25. When one nitrogen atom joins with another nitrogen atom forming a molecule of nitrogen gas, the atoms are held together by a(n)
    a. single bond
    b. double bond
    c. triple bond
    d. ionic bond
26. When nitrogen and hydrogen combine, the ratio of hydrogen atoms to nitrogen atoms in a molecule of the resulting compound is
    a. 2 to 1
    b. 3 to 1
    c. 1 to 3
    d. 1 to 1
27. Atoms of which pair of elements are most likely to form a polar covalent bond?
    a. H and O
    b. Na and O
    c. Na and Cl
    d. Cl and Cl
28. The correct symbol for an ion of oxygen is
    a. O^{+}
    b. O^{2+}
    c. O^{-}
    d. O^{2-}