Chapter 6

6.1 Ionic Bonding

Reading Focus

Build Vocabulary

Word Forms Have students think of word forms related to crystals such as crystalline and crystallize. Have them discuss ways people commonly use these words—and the word crystal—and compare those usages to the way crystals is defined in Section 6.1.

Reading Strategy

a. Form a cation
b. Form an anion

INSTRUCT

Stable Electron Configurations

Integrate Social Studies

In 1902, G.N. Lewis proposed “the theory of the cubical atom.” He illustrated his theory with drawings of cubes with valence electrons placed at their corners. In his classic 1916 paper, “The Atom and the Molecule,” Lewis simplified his diagrams by using dots to represent electrons and a symbol to represent the kernel of an atom. Recreate the cube models for lithium and beryllium on the board or overhead projector. (Draw cubes with one and two corners circled, respectively.) Explain that each circle represents a valence electron. Then, have students refer to Figure 2 and draw their own cube models for boron, carbon, nitrogen, oxygen, and fluorine. Logical, Visual

Vocabulary

◆ electron dot diagram
◆ ion
◆ anion
◆ cation
◆ chemical bond
◆ ionic bond
◆ chemical formula
◆ crystals

Stable Electron Configurations

The highest occupied energy level of a noble gas atom is filled. When the highest occupied energy level of an atom is filled with electrons, the atom is stable and not likely to react. The noble gases have stable electron configurations with eight valence electrons (or two in the case of helium).

The chemical properties of an element depend on the number of valence electrons. Therefore, it is useful to have a model of atoms that focuses only on valence electrons. The models in Figure 2 are electron dot diagrams. An electron dot diagram is a model of an atom in which each dot represents a valence electron. The symbol in the center represents the nucleus and all the other electrons in the atom.

The handle of the titanium mug in Figure 1 was joined to the body by welding. The pieces were heated until their surfaces fused together. The welding of titanium does not take place in air. At the temperature at which welding occurs, titanium becomes hot enough to react with oxygen in the air, forming an oxide. The oxide makes the weld more brittle and likely to break. Because titanium does not react with a noble gas such as argon, the welding of titanium usually takes place in an argon atmosphere.

Argon’s name is a reminder of its inactivity. It comes from the Greek word argos, which means “idle” or “inert.” Why is argon very inactive yet oxygen is highly reactive? Chemical properties, such as reactivity, depend on an element’s electron configuration.

Stable Electron Configurations

The handle and body of this titanium mug were welded together in an argon atmosphere. If titanium is allowed to react with oxygen in air, the compound that forms makes the weld more brittle and more likely to break.

Figure 1 The handle and body of this titanium mug were welded together in an argon atmosphere. If titanium is allowed to react with oxygen in air, the compound that forms makes the weld more brittle and more likely to break.
An electron is transferred from each sodium atom to a chlorine atom. The charge on the atom is not balanced and the atom is not neutral. An electron is transferred from each sodium atom to a chlorine atom.

**Ionic Bonds**

Elements that do not have complete sets of valence electrons tend to react. By reacting, they achieve electron configurations similar to those of noble gases. Some elements achieve stable electron configurations through the transfer of electrons between atoms.

**Transfer of Electrons** Look at the electron dot diagram for chlorine in Figure 2. A chlorine atom has one electron fewer than an argon atom. If the chlorine atom were to gain a valence electron, it would have the same stable electron arrangement as argon. Look at the electron dot diagram for sodium. A sodium atom has one more electron than a neon atom. If a sodium atom were to lose this electron, its highest occupied energy level would have eight electrons. It would then have the same stable electron arrangement as neon.

What happens at the atomic level when sodium reacts with chlorine? An electron is transferred from each sodium atom to a chlorine atom. Each atom ends up with a more stable electron arrangement than it had before the transfer.

\[
\text{Na}^{+} + \text{Cl}^{-} \rightarrow \text{Na}^{+} : \text{Cl}^{-}
\]

**Formation of Ions** When an atom gains or loses an electron, the number of protons is no longer equal to the number of electrons. The charge on the atom is not balanced and the atom is not neutral. An atom that has a net positive or negative electric charge is called an ion. The charge on an ion is represented by a plus or a minus sign. Notice the plus sign next to the symbol for sodium and the minus sign next to the symbol for chlorine.

---

**Build Science Skills**

**Predicting** Emphasize that, except for hydrogen and helium, the dots in an electron dot diagram do not represent all of the electrons in an atom, just the valence electrons.

Have students look at Figure 2. Ask them to predict the electron dot diagrams for rubidium, strontium, indium, tin, antimony, tellurium, iodine, and xenon. (These elements—Rb, Sr, In, Sn, Sb, Te, I, and Xe—have the same valence electron configurations as the elements directly above them in the periodic table.)

Logical, Visual

---

**Go Online**

For: Links on ionic bonds

Visit: www.SciLinks.org

Web Code: ccm-1061

---

**Customize for Inclusion Students**

**Visually Impaired**

Have interested students listen to a recording of ionisation, a musical work scored for percussion and sires in 1931 by French-American composer Edgard Varése (1883–1965). There is no melody or harmony, just blocks of sound produced by instruments including cymbals, maracas, and drums. The noises are supposed to represent what happens to electrons as ionization occurs. Let students state their reactions to the music. Ask them to discuss how appropriate the title is for this piece of music.

---

**Answer to . . .**

Figure 2. Sodium has one valence electron. Chlorine has seven.
What Determines the Size of an Atom or Ion?

**Answers**

1. Within a period, the atomic radius decreases as the atomic number increases.
2. Within these groups, the atomic radius increases as the atomic number increases.
3. When the next higher energy level is occupied, there is a significant increase in the atomic radius.
4. The ionic radius for potassium is much smaller than its atomic radius. The ionic radius for bromine is much larger than its atomic radius.
5. With the loss of valence electrons, the radius decreases. With the addition of valence electrons, the radius increases.
6. An energy level that was occupied is no longer occupied and the size decreases.

For Extra Help

Explain that attractions between protons and electrons largely determine the atomic radius of a specific atom.

Verbal

**FYI**

Because an electron cloud does not have an outer edge, the radius for elements that form diatomic molecules is calculated by measuring the distance between the two nuclei in the molecule and dividing by two.

The ion that forms when a chlorine atom gains an electron has 17 protons and 18 electrons. This ion has a charge of 1− because it has one extra electron. The symbol for the ion is written Cl−, or Cl− for short. An ion with a negative charge is an anion (an eye-un). Anions like the Cl− ion are named by using part of the element name plus the suffix −ide. Thus, Cl− is called a chloride ion.

A sodium ion has 11 protons and 10 electrons. Because it has one extra proton, the sodium ion has a charge of 1+. The symbol for the ion is written Na1+, or Na+ for short. An ion with a positive charge is a cation (kat’ eye-un). Naming a cation is easy. You just use the element name, as in the sodium ion.

**Formation of Ionic Bonds** Remember that a particle with a negative charge will attract a particle with a positive charge. When an anion and a cation are close together, a chemical bond forms between them. A chemical bond is the force that holds atoms or ions together as a unit. An ionic bond is the force that holds cations and anions together. An ionic bond forms when electrons are transferred from one atom to another.

**Atomic and Ionic Radii** The radius of an atom decreases from left to right across a period because valence electrons are shielded from the nucleus by electrons in lower energy levels. The amount of positive nuclear charge experienced by valence electrons is mainly determined by the difference in charge between the nucleus and the inner (or core) electrons. Because the number of core electrons does not change across a period, as the number of protons increases the charge experienced by valence electrons increases and the atomic radius decreases.

The radius of an anion is larger than the radius of its corresponding atom. Adding electrons to the highest occupied energy level increases the repulsions among electrons. The increase in repulsions causes the electrons to spread out more in space.
mula for sodium chloride is $\text{NaCl}$. From the formula, you can tell that atoms or ions of these elements in the compound. The chemical formula shows what elements a compound contains and the ratio of the atom of an element in the formula, no subscript is needed. Generally decreases from left to right across a period.

Figure 3 Ionization energies generally increase from left to right across a period. Interpreting Diagrams What is the trend for ionization energy within a group?

Use Visuals

Figure 3 Emphasize the difference between the trend in ionization energy across a period and the trend within a group. Ask, According to Figure 3, which element has a greater ionization energy, sodium or magnesium? (Magnesium) Which element has a greater ionization energy, potassium or cesium? (Cesium) Logical

**Ionic Compounds**

Compounds that contain ionic bonds are ionic compounds, which can be represented by chemical formulas. A chemical formula is a notation that shows what elements a compound contains and the ratio of the atoms or ions of these elements in the compound. The chemical formula for sodium chloride is $\text{NaCl}$. From the formula, you can tell that there is one sodium ion for each chloride ion in sodium chloride.

Based on the diagram in Figure 4, what would the formula for magnesium chloride be? A magnesium atom cannot reach a stable electron configuration by reacting with just one chlorine atom. It must transfer electrons to two chlorine atoms. After the transfer, the charge on the magnesium ion is $2^+$ and its symbol is $\text{Mg}^{2+}$. The formula for the compound is $\text{MgCl}_2$. The 2 written to the right and slightly below the symbol for chlorine is a subscript. Subscripts are used to show the relative numbers of atoms of the elements present. If there is only one atom of an element in the formula, no subscript is needed.

**Facts and Figures**

**Ionization Energy** Ionization energy varies for each valence electron in an atom. These energies are referred to as first ionization energy, $I_1$, second ionization energy, $I_2$, and so forth. An $I_2$ is always greater than an $I_1$ for a given element because the second electron is being removed from a positively charged ion. Because the positive charge continues to increase with each subsequent removal of an electron, an $I_3$ is greater than an $I_2$.

**FYI**

Memory aids may help students distinguish cations and anions. Tell students that cation means “to go down,” which might make them think of losing an electron, and that anion means “to go up,” which might make them think of gaining an electron. Students could also think of related terms that start with the same letter, such as cast off for cation and accept for anion. Finally, students could associate the phrase “Negative ION with anion.”

**Use Visuals**

Figure 3 Emphasize the difference between the trend in ionization energy across a period and the trend within a group. Ask, According to Figure 3, which element has a greater ionization energy, sodium or magnesium? (Magnesium) Which element has a greater ionization energy, potassium or cesium? (Cesium) Logical

**Ionic Compounds**

Build Reading Literacy

KWL Refer to page 124D in Chapter 5, which provides the guidelines for a KWL. Before students read Ionic Compounds, have students construct a KWL chart with three columns entitled, What I Know, What I Want to Know, and What I Learned. Have them fill out the final column after they have read pp. 161–164. Intrapersonal

**Build Math Skills**

Ratios and Proportions When elements combine to form compounds, they do so in specific whole number ratios. For example, water, or $\text{H}_2\text{O}$, has a hydrogen to oxygen ratio of 2:1. Ask students to calculate the following ratios: sodium to chlorine in sodium chloride, $\text{NaCl}$; magnesium to chlorine in magnesium chloride, $\text{MgCl}_2$; and sodium to oxygen in sodium oxide, $\text{Na}_2\text{O}$. (1:1; 1:2; 2:1) Logical

Direct students to the Math Skills in the Skills and Reference Handbook at the end of the student text for additional help.

**Answer to . . .**

Figure 3 It generally decreases from top to bottom within a group. The amount of energy needed to remove an electron from an atom

**Chemical Bonds**

161
Integrate Earth Science

Geologists use a system to classify minerals similar to the one used by chemists to classify compounds. Ask students to use the library or Internet to find photographs of minerals that have cubic crystals (like sodium chloride) or hexagonal crystals (like ruby). Some students may also want to find examples of tetragonal, monoclinic, triclinic, and orthorhombic crystals.

Use Visuals

Figure 5 Have students examine Figure 5. Ask, How are sodium ions represented in the figure? (Sodium ions are represented by the smaller, orange spheres.) How are chloride ions represented in the figure? (Chloride ions are represented by the larger, green spheres.) What do you notice about the pattern of the locations of positive and negative ions in the diagram? (No two neighboring ions have the same charge.) What similarity do you notice between the diagram and the photograph of sodium chloride crystals? (The structures in both the diagram and the photograph have a cubic shape.)

Visual

Build Science Skills

Applying Concepts Help students understand the structure of a crystal lattice by encouraging them to think of three-dimensional analogies for lattices, such as scaffolding on a building or cups and saucers stacked on a tray in a restaurant.

Visual, Logical

Crystal Lattices A chemical formula for an ionic compound tells you the ratio of the ions in the compound. But it does not tell you how the ions are arranged in the compound. If you looked at a sample of sodium chloride with a hand lens or microscope, you would be able to see that the pieces of salt are shaped like cubes. This shape is a clue to how the sodium and chloride ions are arranged in the compound.

Figure 5A shows that the ions in sodium chloride are arranged in an orderly, three-dimensional structure. Each chloride ion is surrounded by six sodium ions and each sodium ion is surrounded by six chloride ions. Each ion is attracted to all the neighboring ions with an opposite charge. This set of attractions keeps the ions in fixed positions in a rigid framework, or lattice. The repeating pattern of ions in the lattice is like the repeating pattern of designs on the wallpaper in Figure 6.

Solids whose particles are arranged in a lattice structure are called crystals. Compare the cubic shape of the sodium chloride crystals in Figure 5B to the arrangement of ions in Figure 5A. The shape of an ionic crystal depends on the arrangement of ions in its lattice. In turn, the arrangement of the ions depends on the ratio of ions and their relative sizes. Crystals are classified into groups based on the shape of their crystals. Crystals of ruby have a six-sided, hexagonal shape. The How It Works box on page 163 describes one way to make rubies.

What shape are sodium chloride crystals?

Facts and Figures

Crystal Systems Crystals are classified into seven different crystal systems, which are described by geometric figures with six faces. The figures are distinguished by the angles at which the faces meet and by how many edges on a face are equal in length. Cubic: three equal edges, three 90° angles; tetragonal: two equal edges, three 90° angles; orthorhombic: no equal edges, three 90° angles; monoclinic: no equal edges, two 90° angles; triclinic: no equal edges, no 90° angles; hexagonal: two equal edges, two 90° angles, one 120° angle; rhombohedral: three equal edges, two 90° angles.
Synthetic Rubies
Rubies are mainly aluminum oxide, which is white. The substitution of a small percentage of chromium ions for aluminum ions gives rubies their distinctive red color. Because natural rubies are rare, rubies are often manufactured.

**Interpreting Diagrams** What substances are in the mixture used to make rubies?

**Making synthetic rubies**
One way of making synthetic rubies is called the pulled-growth method. It was invented by Polish scientist Jan Czochralski (1885–1953).

1. **Seed crystal**
   - Aluminum oxide and chromium(VI) oxide are melted. A tiny piece of ruby, called a seed crystal, is attached to a rod and placed above the molten mixture (melt).

2. **Lowering into the melt**
   - The rod is lowered until the seed crystal touches the melt. The rod is slowly lifted, and ions in the melt begin to attach themselves to the seed crystal to form a ruby.

3. **Forming a boule**
   - As the rod is lifted higher, an oblong-shaped crystal called a boule grows from the end. Once cooled, the boule can be cut into different shapes.

**Synthetic ruby**
A synthetic ruby boule has a hexagonal crystal structure identical to the natural ruby gemstone. Its shape is determined by the arrangement of ions in the crystal.

**NATURAL RUBY**

**HEXAGONAL STRUCTURE CRYSTAL**

**SYNTHETIC RUBY BOULE**

**FORMING A BOULE**

**Answer to . . .**

*Figure 6* The pattern of design elements in the wallpaper repeats the way the arrangement of ions repeats in a crystal lattice.

They are cubes.
Section 6.1 (continued)

F.Y.I.

Real crystals are not perfect. Sites in a lattice may be vacant, sites may be occupied by impurities, and occupied sites may be squeezed in between regular sites in the lattice (an interstitial defect). In an ionic compound, a cation vacancy is balanced by a nearby anion vacancy or by an interstitial cation (which maintains an overall balance of charge). Crystal defects are largely responsible for how crystals fracture under stress.

Use Visuals

Figure 7 Have students compare the before-and-after diagrams. Ask, When the hammer hits the crystal, what happens to the positions of the ions? (Ions with similar charge are pushed near one another.) How do objects with the same charge behave? (They repel.) Visual

ASSESS

Evaluate Understanding

Have students describe the formation of anions, cations, and ionic bonds.

Reteach

Use the diagram at the bottom of p. 161 to review the formation of cations, anions, and ionic bonds.

Connecting Concepts

Potassium is more reactive than calcium because the amount of energy needed to remove a single valence electron from a potassium atom is much smaller than the amount of energy needed to remove two valence electrons from a calcium atom.

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 6.1.

Properties of Ionic Compounds The properties of sodium chloride are typical of an ionic compound. It has a high melting point (801°C). In its solid state, sodium chloride is a poor conductor of electric current. But when melted, it is a good conductor of electric current. Sodium chloride crystals shatter when struck with a hammer. The properties of an ionic compound can be explained by the strong attractions among ions within a crystal lattice.

Recall that the arrangement of particles in a substance is the result of two opposing factors. The first factor is the attractions among particles in the substance. The second factor is the kinetic energy of the particles. The stronger the attractions among the particles, the more kinetic energy the particles must have before they can separate.

For an electric current to flow, charged particles must be able to move from one location to another. The ions in a solid crystal lattice have fixed positions. However, when the solid melts, the lattice breaks apart and the ions are free to flow. Melted, or molten, sodium chloride is an excellent conductor of electric current.

Rock salt contains large crystals of sodium chloride. If you tapped a crystal of rock salt sharply with a hammer, it would shatter into many smaller crystals. Figure 7 shows what happens to the positions of the ions when the crystal is struck. Negative ions are pushed into positions near negative ions, and positive ions are pushed into positions near positive ions. Ions with the same charge repel one another and cause the crystal to shatter.

Section 6.1 Assessment

Reviewing Concepts

1. How is an atom least likely to react?
2. Describe one way an element can achieve a stable electron configuration.
3. What characteristic of ionic bonds can be used to explain the properties of ionic compounds?
4. Use ionization energy to explain why metals lose electrons more easily than nonmetals.
5. Why is a rock salt crystal likely to shatter when struck?

Critical Thinking

6. Making Generalizations What will the ratio of ions be in any compound formed from a Group 1A metal and a Group 7A nonmetal? Explain your answer.

7. Drawing Conclusions Why do ionic compounds include at least one metal?
8. Predicting Based on their chemical formulas, which of these compounds is not likely to be an ionic compound: KBr, SO2, or FeCl3? Explain your answer.

Reactivity of Metals Use what you know about how ionic bonds form to explain the difference in reactivity between potassium and calcium. If necessary, reread the description of Group 1A and Group 2A properties in Section 5.3.

Section 6.1 Assessment

1. When the highest occupied energy level of an atom is filled with electrons
2. Through the transfer of electrons between atoms
3. The strong attractions among ions within a crystal lattice
4. Metals have lower ionization energies than nonmetals. The lower the ionization energy, the easier it is to remove an electron from an atom.

5. When ions with the same charge are pushed close together, they repel one another.
6. The ratio will be one to one because a Group 1A metal loses one electron and a Group 7A nonmetal gains one electron to achieve a stable electron configuration.
7. One element in an ionic compound must form cations, but nonmetals tend to form anions.
8. SO2 because sulfur and oxygen are both nonmetals and unlikely to form cations.
6.2 Covalent Bonding

Key Concepts
- How are atoms held together in a covalent bond?
- What happens when atoms don't share electrons equally?
- What factors determine whether a molecule is polar?
- How do attractions between polar molecules compare to attractions between nonpolar molecules?

Vocabulary
- covalent bond
- molecule
- polar covalent bond

Reading Strategy
Relating Text and Visuals
Copy the table. As you read, look closely at Figure 9. Complete the table by describing each type of model shown.

<table>
<thead>
<tr>
<th>Model</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron dot</td>
<td>a</td>
</tr>
<tr>
<td>Structural formula</td>
<td>b</td>
</tr>
<tr>
<td>Space-filling</td>
<td>c</td>
</tr>
<tr>
<td>Electron cloud</td>
<td>d</td>
</tr>
</tbody>
</table>

Plants absorb water through their roots from soil or from a solution containing nutrients, as in Figure 8. Carbon dioxide from the air enters the plants through small openings in their leaves. The plants use the energy from sunlight to convert water and carbon dioxide into a sugar. Energy is stored in the chemical bonds of the sugar.

The elements in sugar are carbon, oxygen, and hydrogen. All three are nonmetals, which have relatively high ionization energies. A transfer of electrons does not tend to occur between nonmetal atoms. So, how are two nonmetals able to form bonds?

Covalent Bonds
You and a friend are participating in a treasure hunt. The rules state that the first person to find all eight items on a list will win a 21-speed bicycle. After about an hour, you have found six of the items on the list and your friend has found the other two. You and your friend have incomplete sets of items. But if you are willing to share your items with your friend, together you will have a complete set of items and qualify for the prize. Of course, you will have to be willing to share the bicycle, too.

When nonmetals join together, they display a similar sharing strategy.
As a space shuttle lifts off, it leaves a water vapor trail. A reaction of hydrogen and oxygen produces the water.

Using Models How is the bond between hydrogen atoms represented in each model of a hydrogen molecule?

Figure 9 As a space shuttle lifts off, it leaves a water vapor trail. A reaction of hydrogen and oxygen produces the water. Using Models How is the bond between hydrogen atoms represented in each model of a hydrogen molecule?

Figure 10 These space-filling models represent diatomic molecules of five elements. Using Models How many atoms are in a diatomic molecule?

Sharing Electrons A hydrogen atom has one electron. If it had two electrons, it would have the same electron configuration as a helium atom. Two hydrogen atoms can achieve a stable electron configuration by sharing their electrons and forming a covalent bond. A covalent bond is a chemical bond in which two atoms share a pair of valence electrons. When two atoms share one pair of electrons, the bond is called a single bond.

In the electron dot model, the bond is shown by a pair of dots in the space between the symbols for the hydrogen atoms. In the structural formula, the pair of dots is replaced by a line. The electron cloud model and the space-filling model show that orbitals of atoms overlap when a covalent bond forms.

Molecules of Elements Two hydrogen atoms bonded together form a unit called a molecule. A molecule is a neutral group of atoms that are joined together by one or more covalent bonds. The hydrogen molecule is neutral because it contains two protons (one from each atom) and two electrons (one from each atom). What keeps the hydrogen atoms together in the molecule? The attractions between the shared electrons and the protons in each nucleus hold the atoms together in a covalent bond.

A chemical formula can be used to describe the molecules of an element as well as a compound. The element hydrogen has the chemical formula H₂. The subscript 2 indicates that there are two atoms in a molecule of hydrogen.

Many nonmetal elements exist as diatomic molecules. Diatomic means “two atoms.” Four of the models in Figure 10 are of halogens. A halogen atom has seven valence electrons. If two halogen atoms share a valence electron from each atom, both atoms have eight valence electrons.

Customize for English Language Learners

Think-Pair-Share Have students work in pairs to think of situations that serve as analogies for ionic and covalent bonding. For example, if a jeweler lends an expensive piece of jewelry to a presenter at an awards show, a guard will accompany the presenter and stay as close as possible. This is similar to an ionic bond, in which a cation stays close to an anion to which it donates an electron. By contrast, students reading from one copy of the same book are like two atoms sharing a pair of electrons in a covalent bond. Strengthen discussion skills by having students share analogies with the class.
**Chemical Bonds 167**

### Unequal Sharing of Electrons

**Multiple Covalent Bonds** Nitrogen has five valence electrons. If two nitrogen atoms shared a pair of electrons, each would have only six valence electrons. If they shared two pairs of electrons, each atom would have only seven valence electrons. When the atoms in a nitrogen molecule (N₂) share three pairs of electrons, each atom has eight valence electrons. Each pair of shared electrons is represented by a long dash in the structural formula N≡N. When two atoms share three pairs of electrons, the bond is called a triple bond. When two atoms share two pairs of electrons, the bond is called a double bond.

**Unequal Sharing of Electrons** In general, elements on the right of the periodic table have a greater attraction for electrons than elements on the left have (except for noble gases). In general, elements at the top of a group have a greater attraction for electrons than elements at the bottom of a group have. Fluorine is on the far right and is at the top of its group. It has the strongest attraction for electrons and is the most reactive nonmetal.

**Analyzing Inks**

**Materials** test paper, metric ruler, felt-tip markers, stapler, beaker, alcohol-water mixture, Petri dish

**Procedure**

1. Place the test paper on a clean surface. Use the ruler to draw the pencil line shown in the drawing. Use your markers to place color dots at the locations shown in the drawing.
2. With the ink marks on the outside, staple the two ends of the paper together to form a tube.
3. Pour the alcohol-water mixture into the beaker to a depth of 0.5 cm. Stand the paper in the beaker so that the dots are at the bottom. The paper should not touch the sides of the beaker. Invert the Petri dish over the beaker.
4. When the mixture reaches the top of the paper, remove the paper from the beaker. Unstaple the paper and lay it flat. Make a drawing of the results with each colored area labeled.

**Analyze and Conclude**

1. **Observing** Which markers contained inks that were mixtures of colored substances?
2. **Formulating Hypotheses** How did some molecules in the ink move up the paper?
3. **Predicting** Assume that molecules in the test paper are more polar than molecules in the alcohol-water mixture. Would you expect the most polar molecules in ink to stick tightly to the paper or to move with the liquid? Explain.
4. **Designing Experiments** How could the procedure from this lab be used to identify a black ink whose composition is unknown?

**Materials**

- test paper
- metric ruler
- felt-tip markers
- stapler
- beaker
- alcohol-water mixture
- Petri dish

**Procedure**

1. Place the test paper on a clean surface. Use the ruler to draw the pencil line shown in the drawing. Use your markers to place color dots at the locations shown in the drawing.
2. With the ink marks on the outside, staple the two ends of the paper together to form a tube.
3. Pour the alcohol-water mixture into the beaker to a depth of 0.5 cm. Stand the paper in the beaker so that the dots are at the bottom. The paper should not touch the sides of the beaker. Invert the Petri dish over the beaker.
4. When the mixture reaches the top of the paper, remove the paper from the beaker. Unstaple the paper and lay it flat. Make a drawing of the results with each colored area labeled.

**Analyze and Conclude**

1. **Observing** Which markers contained inks that were mixtures of colored substances?
2. **Formulating Hypotheses** How did some molecules in the ink move up the paper?
3. **Predicting** Assume that molecules in the test paper are more polar than molecules in the alcohol-water mixture. Would you expect the most polar molecules in ink to stick tightly to the paper or to move with the liquid? Explain.
4. **Designing Experiments** How could the procedure from this lab be used to identify a black ink whose composition is unknown?

**Unequal Sharing of Electrons**

In general, elements on the right of the periodic table have a greater attraction for electrons than elements on the left have (except for noble gases). In general, elements at the top of a group have a greater attraction for electrons than elements at the bottom of a group have. Fluorine is on the far right and is at the top of its group. It has the strongest attraction for electrons and is the most reactive nonmetal.
Chapter 6

[Page not visible]

Section 6.2 (continued)

Modeling Overall Polarity

Purpose
Students examine a model for molecular polarity.

Materials
4 12-inch pieces of string or yarn, tape, overhead projector

Procedure
Show students ball-and-stick models of a carbon dioxide molecule and a water molecule. Have them compare the linear shape of CO₂ with the bent shape of H₂O. Tie or tape one piece of string onto each of the oxygen atoms of the CO₂ model and onto each of the hydrogen atoms of the H₂O model. Place the CO₂ model on the overhead and demonstrate the effect of pulling the strings gently in opposite directions. Explain that this represents the canceling effect of opposing polar bonds (dipoles). Place the H₂O model on the overhead and demonstrate the effect of pulling the strings gently away from the oxygen atom in the direction of the bond angles. Explain that this represents the additive effect of polar bonds that are at an angle.

Expected Outcome
The CO₂ model will stay in one place. The H₂O model will move in the direction of the hydrogen atoms.

Visual

Build Science Skills

Using Models
Show students a ball-and-stick molecular model of ammonia, NH₃. Ask students to predict whether the molecule is polar or nonpolar and to give a reason for their choice. (Ammonia is a polar molecule because its three bonds are oriented to one side of the central nitrogen atom. The polar bonds do not cancel out.) Now show students a molecular model of sulfur trioxide, SO₃. Ask students to predict whether this molecule is polar or nonpolar and to explain their reasoning. (Sulfur trioxide is a nonpolar molecule because its three bonds are oriented symmetrically in a plane around the central sulfur atom. The polar bonds cancel each other out.) Logical, Visual

Facts and Figures

Surface Tension
Refer back to the discussion in Section 3.3 about the effect of intermolecular attractions on the evaporation of water. Hydrogen bonds explain other properties of water, including its high surface tension. Because molecules on the surface are drawn inward by attractions from molecules beneath the surface, surface area is reduced. Surface tension explains why water drops bead up on a clean, waxed surface.
Attraction Between Molecules

In a molecular compound, there are forces of attraction between molecules. These attractions are not as strong as ionic or covalent bonds, but they are strong enough to hold molecules together in a liquid or solid. Attraction between polar molecules are stronger than attractions between nonpolar molecules.

Water molecules are similar in mass to methane (CH₄) molecules. Yet, methane boils at -161.5°C and water boils at 100°C because methane molecules are nonpolar and water molecules are polar. Each dashed line in Figure 13 represents an attraction between a partially positive hydrogen atom in one water molecule and a partially negative oxygen atom in another. Molecules on the surface of a water sample are attracted to molecules that lie below the surface and are pulled toward the center of the sample. These attractions increase the energy required for water molecules to evaporate. They raise the temperature at which vapor pressure equals atmospheric pressure—the boiling point.

Attraction among nonpolar molecules are weaker than attractions among polar molecules, but they do exist. After all, carbon dioxide can exist as solid dry ice. Attractions among nonpolar molecules explain why nitrogen can be stored as a liquid at low temperatures and high pressures. Because electrons are constantly in motion, there are times when one part of a nitrogen molecule has a small positive charge and one part has a small negative charge. At those times, one nitrogen molecule can weakly attracted to another nitrogen molecule.

Section 6.2 Assessment

Reviewing Concepts
1. What attractions hold atoms together in a covalent bond?
2. What happens to the charge on atoms when they form a polar covalent bond?
3. Name the two factors that determine whether a molecule is polar.
4. Compare the strength of attractions between polar molecules to the strength of attractions between nonpolar molecules.
5. What is a molecule?
6. Applying Concepts Which of these elements does not bond to form molecules: oxygen, chlorine, neon, or sulfur?
7. Inferring Why is the boiling point of water higher than the boiling point of chlorine?
8. Using Diagrams Based on their electron dot diagrams, what is the formula for the covalently bonded compound of nitrogen and hydrogen?

Connecting Concepts
Viscosity Review the description of the physical property viscosity in Section 2.2. Then write a paragraph explaining how attractions between molecules might affect the viscosity of a liquid.

Critical Thinking
6. Applying Concepts Which of these elements does not bond to form molecules: oxygen, chlorine, neon, or sulfur?
4. Attraction between polar molecules are stronger than attractions between nonpolar molecules.
5. A neutral group of atoms that are joined together by one or more covalent bonds
6. Neon
7. Attraction between polar water molecules are stronger than attractions between nonpolar chlorine molecules.
8. NH₃

Surfacing Tension

Purpose Students observe how surface tension can support a needle.
Materials 200-mL beaker, water, sewing needle, tweezers, dropper pipet
Procedure Fill the beaker with water. Using tweezers, gently place the needle on the water’s surface so that surface tension supports it. Remove the needle, and place it in the water vertically so that it sinks. Pour the water out and collect the needle.
Expected Outcome Surface tension supports the needle.
Logical, Visual

Assess

Evaluate Understanding
Have students make a chart that compares and contrasts polar covalent bonds and nonpolar covalent bonds. Be sure students discuss the bonding atoms’ attractions for electrons, partial charges, and attractions between molecules.
Reteach
Use Figures 11, 12, and 13 as visual aids while reviewing polar covalent bonds, nonpolar and polar molecules, and attractions between molecules.

A greater attraction between molecules is likely to produce an increase in viscosity because the attractions would act in opposition to the motion of the molecules and reduce their ability to flow.

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 6.2.

Answer to . . .

Figure 13 Chlorine

Figure 13 The oxygen atom has a partial negative charge. The hydrogen atoms have partial positive charges.
6.3 Naming Compounds and Writing Formulas

**Objectives**

6.3.1 Recognize and describe binary ionic compounds, metals with multiple ions, and polyatomic ions.

6.3.2 Name and determine chemical formulas for ionic and molecular compounds.

**Reading Focus**

**Key Concepts**
- What information do the name and formula of an ionic compound provide?
- What information do the name and formula of a molecular compound provide?

**Vocabulary**
- polyatomic ion

**Reading Strategy**

Predicting: Copy the table. Before you read, predict the meaning of the term polyatomic ion. After you read, if your prediction was incorrect, revise your definition.

<table>
<thead>
<tr>
<th>Vocabulary Term</th>
<th>Before You Read</th>
<th>After You Read</th>
</tr>
</thead>
<tbody>
<tr>
<td>Polyatomic ion</td>
<td>a. ?</td>
<td>b. ?</td>
</tr>
</tbody>
</table>

**Build Vocabulary**

**Word-Part Analysis** Ask students what words they know that have the prefix poly-. (Polygon, polysyllabic, polyglot, polytechnic, and polygraph). Give a definition of the prefix. (Poly- means “many.”) Have students predict the meaning of the term polyatomic ion. (A polyatomic ion is a covalently bonded group of atoms that has a positive or negative charge and acts as a unit.)

**Reading Strategy**

A and B. Students should assume that any particle described as an ion has a charge. If they know the meaning of poly-, they may conclude that the ion contains three or more atoms.

**INSTRUCT**

**Integrate Language Arts**

The production of lime through the decomposition of limestone (or any form of calcium carbonate) has been known for millennia, which is why there is a word for lime in many ancient languages. In Latin, this word is calx, which is the source for the name calcium. Have students research the origin of the phrase “in the limelight.” (Drummond developed limelight first as an aid to surveying. When lime was heated in a hydrogen-oxygen flame, it produced a bright, white light. In 1825, a light that Drummond placed on top of a hill in Belfast could be seen in Donegal about 105 km (66 miles) away. Limelight was first used in a theater in 1856, when a lens was placed in front of the limelight to produce a spotlight.) Logical

**Section Resources**

**Print**
- Laboratory Manual, Investigation 6A
- Reading and Study Workbook With Math Support, Section 6.3 and Math Skill: Writing Formulas for Ionic Compounds
- Math Skills and Problem Solving Workbook, Section 6.3
- Transparencies, Section 6.3

**Technology**
- Interactive Textbook, Section 6.3
- Presentation Pro CD-ROM, Section 6.3
- Go Online, NSTA SciLinks, Chemical formulas
Describing Ionic Compounds

Both of the objects in Figure 15 are coated with compounds of copper and oxygen. Based on the two colors of the coatings, copper and oxygen must form at least two compounds. One name cannot describe all the compounds of copper and oxygen. There must be at least two names to distinguish red copper oxide from black copper oxide.

The name of an ionic compound must distinguish the compound from other ionic compounds containing the same elements. The formula of an ionic compound describes the ratio of the ions in the compound.

**Binary Ionic Compounds** A compound made from only two elements is a binary compound. (The Latin prefix bi- means “two,” as in bicycle or bisect.) Naming binary ionic compounds, such as sodium chloride and cadmium iodide, is easy. The names have a predictable pattern: the name of the cation followed by the name of the anion. Remember that the name for the cation is the name of the metal without any change; sodium atom and sodium ion. The name for the anion uses part of the name of the nonmetal with the suffix -ide to the stem of the name of the nonmetal. Figure 16 shows the names and charges for eight common anions.

![Figure 16](image)

**Common Anions**

<table>
<thead>
<tr>
<th>Element Name</th>
<th>Ion Name</th>
<th>Ion Symbol</th>
<th>Ion Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fluorine</td>
<td>Fluoride</td>
<td>F&lt;sup&gt;-&lt;/sup&gt;</td>
<td>1&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Chloride</td>
<td>Cl&lt;sup&gt;-&lt;/sup&gt;</td>
<td>1&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Bromine</td>
<td>Bromide</td>
<td>Br&lt;sup&gt;-&lt;/sup&gt;</td>
<td>1&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Iodine</td>
<td>Iodide</td>
<td>I&lt;sup&gt;-&lt;/sup&gt;</td>
<td>1&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Oxygen</td>
<td>Oxide</td>
<td>O&lt;sup&gt;2-&lt;/sup&gt;</td>
<td>2&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Sulfur</td>
<td>Sulfide</td>
<td>S&lt;sup&gt;2-&lt;/sup&gt;</td>
<td>2&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>Nitride</td>
<td>N&lt;sup&gt;3-&lt;/sup&gt;</td>
<td>3&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>Phosphate</td>
<td>P&lt;sup&gt;3-&lt;/sup&gt;</td>
<td>3&lt;sup&gt;-&lt;/sup&gt;</td>
</tr>
</tbody>
</table>

**Describing Ionic Compounds**

Some students may think that a material’s particles possess the same properties as the material. For example, they may think that the atoms that compose the copper oxides in Figure 15 are red or black. Challenge this misconception by noting that the black and red copper oxides both contain the same two elements—copper and oxygen—but they are different colors.

**Logical**

**FYI**

The copper oxides differ in more than color. Copper(I) oxide melts at 1235°C and has a density of 6.0 g/cm<sup>3</sup>. Copper(II) oxide melts at 1446°C and has a density of 6.31 g/cm<sup>3</sup>.

**Build Reading Literacy**

Outline Refer to page 156D in this chapter, which provides the guidelines for an outline.

Have students read the text on pp. 171–175 related to describing ionic and molecular compounds. Then, have students use the headings as major divisions in an outline. Have students refer to their outlines when answering the questions in the Section 6.3 Assessment.

**Visual**

![Figure 15](image)

They must be different compounds because their colors vary, and the properties of a compound should be consistent.
Section 6.3 (continued)

Use Visuals

**Figure 18** Have students compare the two models of an ammonium ion. Have them discuss the advantages of each type of model. (Both models show the number and types of atom in the ion. The electron dot diagram shows the valence electrons. The space-filling model shows the relative sizes of the atoms and how they are arranged in space.) Point out that the brackets in the electron dot diagram indicate that the group of atoms as a whole, not any specific atom, has a positive charge. Ask, What is the charge on an ammonium ion? (1 +) How many covalent bonds are in an ammonium ion? (Four) How many valence electrons are involved in the bonds in the ammonium ion? (Eight)

**Visual**

Paint is a mixture of nonvolatile ingredients (the pigment and the binder) that are dispersed in a volatile liquid. A pigment is a substance that provides the color.

**FYI**

Many paint pigments contain compounds of transition metals. These metals often form more than one type of ion. The ion names must contain a Roman numeral. Using Tables How is the Roman numeral in the name related to the charge on the ion?

<table>
<thead>
<tr>
<th>Ion Name</th>
<th>Ion Symbol</th>
<th>Ion Name</th>
<th>Ion Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper(I)</td>
<td>Cu⁺</td>
<td>Chromium(II)</td>
<td>Cr²⁺</td>
</tr>
<tr>
<td>Copper(I)</td>
<td>Cu⁺⁺</td>
<td>Chromium(II)</td>
<td>Cr²⁺</td>
</tr>
<tr>
<td>Iron(III)</td>
<td>Fe³⁺</td>
<td>Titanium(III)</td>
<td>Ti³⁺</td>
</tr>
<tr>
<td>Iron(III)</td>
<td>Fe³⁺</td>
<td>Titanium(III)</td>
<td>Ti³⁺</td>
</tr>
<tr>
<td>Lead(II)</td>
<td>Pb²⁺</td>
<td>Titanium(IV)</td>
<td>Pb⁴⁺</td>
</tr>
<tr>
<td>Lead(IV)</td>
<td>Pb⁴⁺</td>
<td>Mercury(II)</td>
<td>Hg²⁺</td>
</tr>
</tbody>
</table>

**Figure 17** Many paint pigments contain compounds of transition metals. These metals often form more than one type of ion. The ion names must contain a Roman numeral. Using Tables How is the Roman numeral in the name related to the charge on the ion?

**Metals With Multiple Ions** The alkali metals, alkaline earth metals, and aluminum form ions with positive charges equal to the group number. For example, the symbol for a potassium ion is K⁺, the symbol for a calcium ion is Ca²⁺, and the symbol for an aluminum ion is Al³⁺.

Many transition metals form more than one type of ion. Notice the two copper ions listed in Figure 17, a copper(I) ion with a 1+ charge and a copper(II) ion with a 2+ charge. When a metal forms more than one ion, the name of the ion contains a Roman numeral to indicate the charge on the ion. These ion names can distinguish red copper(I) oxide from black copper(II) oxide. The formula for “copper one oxide” is Cu₂O because it takes two Cu⁺ ions to balance the charge on an O²⁻ ion. The formula for “copper two oxide” is CuO because it takes only one Cu²⁺ ion to balance the charge on an O²⁻ ion.

**Polyatomic Ions** The electron dot diagram in Figure 18 describes a group of atoms that includes one nitrogen and four hydrogen atoms. It is called an ammonium ion. The atoms are joined by covalent bonds. Why does the group have a positive charge? The nitrogen atom has seven protons, and each hydrogen atom has one proton—eleven in total. But the group has only ten electrons to balance the charge on the protons—eight valence electrons and nitrogen’s two inner electrons. A covalently bonded group of atoms that has a positive or negative charge and acts as a unit is a polyatomic ion. The prefix poly- means “many.” Most simple polyatomic ions are anions. Figure 19 lists the names and formulas for some polyatomic ions. Sometimes there are parentheses in a formula that includes polyatomic ions. For example, the formula for iron(III) hydroxide is Fe(OH)₃. The subscript 3 indicates that there are three hydroxide ions for each iron(III) ion.

**When are Roman numerals used in compound names?**
Using Models

Materials
blue plastic-foam ball, black plastic-foam ball,
7 white gumdrops, toothpicks

Procedure
1. To make a model of an ammonia molecule (NH₃), insert a toothpick in each of 3 gumdrops. The gumdrops represent hydrogen atoms and the toothpicks represent bonds.
2. An ammonia molecule is like a pyramid with the nitrogen at the top and the hydrogen atoms at the corners of the base. Insert the toothpicks in the blue foam ball (nitrogen) so that each gumdrop is the same distance from the ball.

Writing Formulas for Ionic Compounds
If you know the name of an ionic compound, you can write its formula. Place the symbol of the cation first, followed by the symbol of the anion. Use subscripts to show the ratio of the ions in the compound. Because all compounds are neutral, the total charges on the cations and anions must add up to zero.

Suppose an atom that gains two electrons, such as sulfur, reacts with an atom that loses one electron, such as sodium. There must be a subscript to show the ratio of the ions in the compound. Because all compounds are neutral, the total charges on the cations and anions must add up to zero.

Some Polyatomic Ions

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonium</td>
<td>NH₄⁺</td>
<td>Acetate</td>
<td>CH₃CO₂⁻</td>
</tr>
<tr>
<td>Hydroxide</td>
<td>OH⁻</td>
<td>Peroxide</td>
<td>O₂⁻</td>
</tr>
<tr>
<td>Nitrate</td>
<td>NO₃⁻</td>
<td>Permanganate</td>
<td>MnO₄⁻</td>
</tr>
<tr>
<td>Sulfate</td>
<td>SO₄²⁻</td>
<td>Hydrogen sulfate</td>
<td>HSO₄⁻</td>
</tr>
<tr>
<td>Carbonate</td>
<td>CO₃²⁻</td>
<td>Hydrogen carbonate</td>
<td>HCO₃⁻</td>
</tr>
<tr>
<td>Phosphate</td>
<td>PO₄³⁻</td>
<td>Hydrogen phosphate</td>
<td>HPO₄²⁻</td>
</tr>
<tr>
<td>Chromate</td>
<td>CrO₄²⁻</td>
<td>Dichromate</td>
<td>Cr₂O₇²⁻</td>
</tr>
<tr>
<td>Silicate</td>
<td>SiO₃²⁻</td>
<td>Hypochlorite</td>
<td>OCI⁻</td>
</tr>
</tbody>
</table>

Figure 15 This table lists the names and formulas of some polyatomic ions. Except for the ammonium ion, all the ions listed are anions. Using Tables Which element is found in all the anions whose names end in -ate?

Analyzing and Concluding
1. Comparing and Contrasting: Compare the shapes of the methane and ammonia molecules.
2. Using Models: Why is carbon in the center of the methane molecule?

Modeling Molecules

Objective
After completing this activity, students will be able to:
- use physical models to compare the shapes of molecules.

Skills Focus Using Models

Prep Time 10 minutes

Advance Prep Plastic-foam balls are available from craft supply stores. If black and blue foam balls are not available, wrap other foam balls in colored foil, glue colored tissue paper around them, or spray-paint them to match the colors used in this book to represent atoms of carbon and nitrogen.

Class Time 20 minutes

Safety Tell students not to eat the gumdrops or anything else in the laboratory.

Expected Outcome Students will make tetrahedral (pyramidal) models of methane molecules.

Analyze and Conclude
1. The methane molecule is a tetrahedron (equal-sided triangular pyramid) and the ammonia molecule has a triangular shape.
2. The carbon atom is in the center because each of the four hydrogen atoms must bond to one carbon atom.

Build Science Skills

Inferring Molecules of NH₃ and NF₃ have the same shape, but bonds in NH₃ are shorter than bonds in NF₃. In other words, the hydrogen atoms are closer to the nitrogen atom in an ammonia molecule than the fluorine atoms are to the nitrogen atom in nitrogen trifluoride. Ask, What does this data imply about the relative sizes of hydrogen and fluorine atoms? (The data confirms that a hydrogen atom has a smaller atomic radius than does a nitrogen atom.) Logical

Answer to . . .

Figure 17 The Roman numeral is equal to the charge on the ion.

Figure 19 Oxygen

When the metal in the compound can form more than one type of ion.
Domestic Skills

Positive and Negative Numbers

Remind students that the charges in an ionic compound must cancel each other. Ask, How many atoms of a halogen would combine with one atom of an alkaline earth metal? Why? (Two atoms; the alkaline earth metal atom loses two electrons but each halogen atom needs only one electron to become stable.) Logical

Direct students to the Math Skills in the Skills and Reference Handbook at the end of the student text for additional help.

Solutions

1. It takes one calcium ion with a charge of 2+ to balance one oxide ion with a charge of 2-. The formula is CaO.
2. Two copper(I) ions, each with a charge of 1+, balance one sulfide ion with a charge of 2-. The formula is Cu2S.
3. Two sodium ions, each with a charge of 1+, balance one sulfate ion with a charge of 2-. The formula is Na2SO4.
4. In the formula, Na represents the sodium ion and OH represents the hydroxide ion. The name of the compound is sodium hydroxide. Logical

For Extra Help

Make sure the first step students take is to find the symbols and charges on the ions. Then, check that they are able to balance the charges. Logical

Additional Problems

1. Write the formula for lithium oxide. (Li2O)
2. Write the formula for iron(III) oxide. (Fe2O3)

Logical

Describing Molecular Compounds

FYI

There are two exceptions to the general rule for naming molecular compounds. Hydrogen is treated as though it were positioned between Group 5A and Group 6A, and oxygen is treated as though it were positioned after chlorine but before fluorine.

For: chemical links

Visit: www.SciLinks.org

Web Code: cco-1063

Writing Formulas for Ionic Compounds

What is the formula for the ionic compound calcium chloride?

1. Write and understand

What information are you given?

The name of the compound is calcium chloride.

2. Plan and solve

List the symbols and charges for the cation and anion.

Ca with a charge of 2+ and Cl with a charge of 1-

Determine the ratio of ions in the compound.

It takes two 1- charges to balance the 2+ charge. There will be two chloride ions for each calcium ion.

Write the formula for calcium chloride.

CaCl2

3. Look back and check

Is your answer reasonable?

Each calcium atom loses two electrons and each chlorine atom gains one electron. So there should be a 1-to-2 ratio of calcium ions to chloride ions.

Describing Molecular Compounds

Like ionic compounds, molecular compounds have names that identify specific compounds, and formulas that match those names. With molecular compounds, the focus is on the composition of molecules. Logical

The name and formula of a molecular compound describe the type and number of atoms in a molecule of the compound.

Naming Molecular Compounds The general rule is that the most metallic element appears first in the name. These elements are farther to the left in the periodic table. If both elements are in the same group, the more metallic element is closer to the bottom of the group. The name of the second element is changed to end in the suffix -ide, as in carbon dioxide.

Download a worksheet on chemical formulas for students to complete, and find additional teacher support from NSTA SciLinks.
Two compounds that contain nitrogen and oxygen have the formulas \( \text{N}_2\text{O}_3 \) and \( \text{NO}_2 \). The names of these two compounds reflect the actual number of atoms of nitrogen and oxygen in a molecule of each compound. You can use the Greek prefixes in Figure 20 to describe the number of nitrogen and oxygen atoms in each molecule.

In an \( \text{N}_2\text{O}_3 \) molecule, there are two nitrogen atoms and four oxygen atoms. The Greek prefixes for two and four are \( \text{di-} \) and \( \text{tetra-} \).

The name for the compound with the formula \( \text{N}_2\text{O}_3 \) is dinitrogen trioxide. In an \( \text{NO}_2 \) molecule, there are one nitrogen atom and two oxygen atoms. The Greek prefixes for one and two are \( \text{mono-} \) and \( \text{di-} \).

So a name for the compound with the formula \( \text{NO}_2 \) is mononitrogen dioxide. However, the prefix \( \text{mono-} \) often is not used for the first element in the name. A more common name for the compound with the formula \( \text{NO}_2 \) is nitrogen dioxide.

Writing Molecular Formulas Writing the formula for a molecular compound is easy. Write the symbols for the elements in the order the elements appear in the name. The prefixes indicate the number of atoms of each element in the molecule. The prefixes appear as subscripts in the formulas. If there is no prefix for an element in the name, there is only one atom of that element in the molecule.

What is the formula for diphosphorus tetrafluoride? Because the compound is molecular, look for elements on the right side of the periodic table. Phosphorus has the symbol \( \text{P} \). Fluorine has the symbol \( \text{F} \). \( \text{Di-} \) indicates two phosphorus atoms and \( \text{tetra-} \) indicates four fluorine atoms. The formula for the compound is \( \text{P}_2\text{F}_4 \).

Figure 20 These Greek prefixes are used to name molecular compounds. The prefix \( \text{tetra-} \) means “eight,” as in the eight tentacles of an octopus.

Integrate Language Arts
To help students learn the prefixes used in molecular compounds, have them think of words they know that contain the Greek prefixes listed in Figure 20. If they are having trouble, encourage them to use a dictionary to find words. Examples include monochrome, dichotomy, tricycle, tetrahedron, pentagon, hexadecimal, heptad, octave, nonagenarian, and decathlon.

Verbal

Evaluate Understanding
Note that the process of writing molecular formulas is the reverse of the process for naming them. Have students write chemical formulas for three substances and chemical names for another three substances. Have students exchange the formulas and names with a partner to check and review their work.

Reteach
Use the tables on pp. 171–173 and the Math Skills on p. 174 to review naming and writing formulas for ionic compounds. Use the table on p. 175 to review naming and writing formulas for molecular compounds.

Section 6.3 Assessment

7. Calculating How many potassium ions are needed to bond with a phosphate ion?

8. What are the names of these ionic compounds: \( \text{LiCl} \), \( \text{BaO} \), \( \text{Na}_3\text{N} \), and \( \text{PbS}_3 \)?

9. Name the molecular compounds with these formulas: \( \text{P}_2\text{O}_5 \) and \( \text{CO} \).

10. What is the formula for the ionic compound formed from potassium and sulfur?

Solutions

8. Lithium chloride, barium oxide, sodium nitride, lead sulfate

9. Diphosphorus pentaoxide (pentoxide) and carbon monoxide

10. K$_2$S because it takes two potassium ions, each with a charge of +1, to balance one sulfide ion with a charge of 2–
6.4 The Structure of Metals

**Key Concepts**
- What are the forces that give a metal its structure as a solid?
- How do metallic bonds produce some of the typical properties of metals?
- How are the properties of alloys controlled?

**Vocabulary**
- metallic bond
- alloy

**Reading Focus**

**Objectives**
- 6.4.1 Describe the structure and strength of bonds in metals.
- 6.4.2 Relate the properties of metals to their structure.
- 6.4.3 Define an alloy and demonstrate how the composition of an alloy affects its properties.

**Build Vocabulary**

**Vocabulary Knowledge Rating Chart**
Have students construct a chart with four columns labeled Term, Can Define It/Use It, Heard It/Seen It, and Don’t Know. Have students copy the terms metallic bond, metal lattice, alloy, and metallurgy into the first column and rate their term knowledge by putting a check in one of the other columns. Ask how many students actually know each term. Have them share their knowledge. To provide a purpose for reading, ask focused questions to help students predict text content based on each term. After students have read the section, have them rate their knowledge again.

**Reading Strategy**

**a. and b. Conductivity or malleability**

**INSTRUCT**

**Metallic Bonds**

**Integrate Social Studies**

In the first light bulbs, air was removed from the light bulb to prevent combustion as the filament heated up. This solution was not ideal because atoms can sublime from the hot filament at very low pressures. With almost no air, atoms of the vaporized filament have uninterrupted paths to the inner walls of the bulb where they are deposited. The modern light bulb uses an argon atmosphere and a tungsten filament. This prolongs the life of the bulb because collisions between tungsten atoms and argon atoms can redirect tungsten atoms back toward the filament. Have interested students research the search for an effective filament. Students can present their findings in a poster or other visual display.

**Visual**

176 Chapter 6

**Print**
- *Reading and Study Workbook With Math Support*, Section 6.4
- *Transparencies*, Section 6.4

**Technology**
- *Interactive Textbook*, Section 6.4
- *Presentation Pro CD-ROM*, Section 6.4
- *Go Online*, Science News, Metals

**Figure 21** This photograph of the tungsten filament from a light bulb was taken with a scanning electron microscope. Color was added to the photo. The filament is magnified more than 100 times. The diameter of the wire is about 15 μm, or 0.0015 cm.

Light bulbs are easy to ignore unless a bulb burns out and you are searching for a replacement in the dark. But in the decades just before the year 1900, light bulbs were an exciting new technology. One challenge for researchers was to find the best material for the filaments in light bulbs. The substance had to be ductile enough to be drawn into a narrow wire. It could not melt at the temperatures produced when an electric current passes through a narrow wire. It had to have a low vapor pressure so that particles on the surface were not easily removed by sublimation.

The substance the researchers found was tungsten (W), a metal whose name means “heavy stone” in Swedish. Figure 21 shows a magnified view of the narrow coils in a tungsten filament. Tungsten has the highest melting point of any metal—3410°C—and it has the lowest vapor pressure. The properties of a metal are related to bonds within the metal.

**Metallic Bonds**

Metal atoms achieve stable electron configurations by losing electrons. But what happens if there are no nonmetal atoms available to accept the electrons? There is a way for metal atoms to lose and gain electrons at the same time. In a metal, valence electrons are free to move among the atoms. In effect, the metal atoms become cations surrounded by a pool of shared electrons. A metallic bond is the attraction between a metal cation and the shared electrons that surround it.
The cations in a metal form a lattice that is held in place by strong metallic bonds between the cations and the surrounding valence electrons. Although the electrons are moving among the atoms, the total number of electrons does not change. So, overall, the metal is neutral.

The metallic bonds in some metals are stronger than in other metals. The more valence electrons an atom can contribute to the shared pool, the stronger the metallic bonds will be. The bonds in an alkali metal are relatively weak because alkali metals contribute only a single valence electron. The result is that alkali metals, such as sodium, are soft enough to cut with a knife and have relatively low melting points. Sodium melts at 97.8°C. Transition metals, such as tungsten, have more valence electrons to contribute and, therefore, are harder and have higher melting points. Recall that tungsten melts at 3410°C.

Explaining Properties of Metals

The structure within a metal affects the properties of metals. The mobility of electrons within a metal lattice explains some of the properties of metals. The ability to conduct an electric current and malleability are two important properties of metals.

Recall that a flow of charged particles is an electric current. A metal has a built-in supply of charged particles that can flow from one location to another—the pool of shared electrons. An electric current can be carried through a metal by the free flow of the shared electrons.

The lattice in a metal is flexible compared to the rigid lattice in an ionic compound. Figure 22 is a model of what happens when someone strikes a metal with a hammer. The metal ions shift their positions and the shape of the metal changes. But the metal does not shatter because ions are still held together by the metallic bonds between the ions and the electrons. Metallic bonds also explain why metals, such as tungsten and copper, can be drawn into thin wires without breaking.

What two important properties of metals can be explained by their structure?

Comparing Bond Types

Purpose
Students observe differences in the properties of substances with ionic bonds and metallic bonds.

Materials
salt lick (or rock salt), copper wire, hammer, goggles

Procedure
Take students outside to an open area. Allow them to examine the samples of rock salt and copper. Have students stand back a safe distance. While wearing goggles, hit each sample with a hammer against a hard surface, such as concrete. Allow students to observe how each sample looks after being pounded with a hammer.

Expected Outcome
The rock salt shatters because sodium chloride is an ionic substance. The end of the copper wire can be pounded flat with a hammer because metals are malleable.

Visual

Explaining Properties of Metals

Use Visuals

Figure 22 Have students examine Figure 22. Ask, How are the cations affected when the hammer strikes the metal? (The cations shift positions.) How are the metallic bonds between the cations and electrons affected when the hammer strikes the metal? (The metallic bonds are unaffected. Each cation is still surrounded by electrons.)

Visual

Address Misconceptions

Many students think that particles in solids cannot move. Challenge this misconception by reminding students that the kinetic theory of matter says that all particles of matter are in constant motion. Ask, Describe the motion of cations in a metal when the metal is not being struck by a hammer. (The cations vibrate, or move repeatedly back and forth, around fixed locations.)

Logical

Answer to ... Figure 22 Malleability

The ability to conduct an electric current and malleability

Customize for English Language Learners

Use a Cloze Strategy

Use a Cloze strategy for students with very limited English proficiency. Have students fill in the blanks in the following sentences while reading Explaining Properties of Metals. The of within a metal lattice explains some of the properties of metal. The lattice in a metal is compared to the lattice in an ionic compound. (mobility; electrons; flexible; rigid)
Alloys

Use Community Resources

Arrange for your class to visit the workshop of a jewelry maker, metalworker, blacksmith, or welder. Have students observe the types of equipment used to work with different metals. Ask questions regarding the artisan’s choice of materials for different projects. Ask about the use of different alloys in metallurgy.

Interpersonal

FYI
There are different categories of alloys: true solutions, heterogeneous mixtures with two phases (a pure element and a compound), or intermetallic compounds with definite compositions.

Milestones in Metallurgy

Ask students to choose one of the significant events in metallurgy history described in the timeline. Have them do research in the library or on the Internet to find out more about the materials, applications, and people involved in their chosen topic. Have them present their findings to the class. Be sure that part of their presentation focuses on the specific properties of metals and alloys.

Verbal

Henry Ford saw an automobile part made from vanadium steel when a French racer crashed at a European race meeting in 1905. He was impressed with the strength and low weight of the steel. American foundries did not have furnaces that could achieve the temperature required to produce the alloy. Ford found a small steel company in Ohio that was willing to experiment with the process and produce the alloy exclusively for Ford.

Verbal

Alloys

A friend shows you a beautiful ring that she says is made from pure gold. Your friend is lucky to have such a valuable object. The purity of gold is expressed in units called karats. Gold that is 100 percent pure is labeled 24-karat gold. Gold jewelry that has a 12-karat label is only 50 percent gold. Jewelry that has an 18-karat label is 75 percent gold.

The surface of an object made from pure gold can easily be worn away by contact with other objects or dented because gold is a soft metal. When silver, copper, nickel, or zinc is mixed with gold, the gold is harder and more resistant to wear. These gold mixtures are alloys. An alloy is a mixture of two or more elements, at least one of which is a metal. Alloys have the characteristic properties of metals.

Milestones in Metallurgy

The science of metallurgy includes ways to extract metals from ores, refine metals, and use metals. Described here are some advances in metallurgy since 1850.

1856 Henry Bessemer develops an efficient process for producing steel by blowing air through molten iron.

1886 Charles Hall and Paul Heroult independently develop a method for using electricity to obtain aluminum from aluminum oxide.

1908 Henry Ford uses vanadium steel (an alloy of iron with carbon and vanadium) extensively in his Model T Fords.

1914 The start of World War I leads to the widespread use of welding techniques, such as gas welding with acetylene, for ship building.
Copper Alloys  The first important alloy was bronze, whose name is associated with an important era in history—the Bronze Age. Metalworkers in Thailand may have been the first to make bronze. But people in other locations probably thought they were the first to make bronze. News didn’t travel quickly in that era.

Metalworkers might have noticed that the metal they extracted by heating deposits of copper was not always the same. The difference in properties could be traced to the presence of tin. In its simplest form, bronze contains only copper and tin, which are relatively soft metals. Mixed together in bronze, the metals are much harder and stronger than either metal alone. Scientists can design alloys with specific properties by varying the types and amounts of elements in an alloy.

Steel-framed structure  The Empire State Building is supported by a framework of steel columns and beams that weigh 60,000 tons. The 102-story Empire State Building in New York City is completed in 1931.

Metal parts, such as the gears in this gold watch, are made by applying heat and pressure to powdered metal in a mold. Superalloys containing rhenium are used in jet engines.

Facts and Figures

Bronze  Bronze jewelry found in graves beneath the town of Ban Chiang in northeast Thailand has been dated to 3600 B.C., but this date is controversial. If correct, the site predates sites in Mesopotamia by several hundred years. Some copper alloys that contain little tin but are similar in color to bronze are labeled as bronzes to take advantage of the reputation of bronze as a hard, durable material.
Bronze and Brass Tones  

**Purpose** Students observe the difference in tone between a brass bell and a bronze bell.  
**Materials** brass bell, bronze bell  
**Procedure** Allow students to examine each bell to see if they appear different in color or shininess. Ring each bell and allow students to note the difference in tone.  
**Expected Outcome** In general, brass is shinier than bronze. A brass bell will have a duller tone that does not last as long as the clear, loud tone of a bronze bell. (Bell bronze contains about 80% copper and 20% tin, which is just the right composition to produce a hard, resonant material.)  
**Visual, Musical**  

---  

**Go Online**  
Visit: PhSchool.com  
Web Code: ccs-1064  

---  

**Build Science Skills**  
**Applying Concepts** After students read about copper alloys, present the following problem to the class. Imagine that you are going to make a wind chime in shop class as a gift. You have two choices of metals to use, brass or bronze. Which do you choose? Explain your answer. (Some students may choose brass because it is shinier and easier to shape into forms than bronze. Others may choose bronze because it has a clearer, louder tone and weathers better than brass.)  
**Intrapersonal**

---  

**Facts and Figures**  
**Bronze Horses** The origin of the gilded bronze horses of St. Mark’s is unclear. The sculptures are not similar in style to other Greek or Roman sculptures of horses. Nor are they similar in composition. From chemical analyses, art historians know that the bronze is 98% copper, 1% tin, and 1% lead, which is an unusually high percent of copper. Art historians do know that the horses were cast in pieces that were welded together. The pieces were made using the lost wax method described in Chapter 2.  

The horses can be traced from a triumphal arch in Rome to one in Constantinople. They were taken from Constantinople to Venice in 1204 (where they were placed on a balcony of the cathedral), then to Paris in 1797, and then back to Venice in 1815. For protection from air pollution, the horses are now stored inside the cathedral. A copy appears on the balcony.
An aluminum-magnesium alloy keeps the advantages of magnesium pure magnesium is soft enough to cut with a knife, and it burns in air.

Other Alloys Airplane parts are made of many different alloys that are suited to particular purposes. The body of a plane is large and needs to be made from a lightweight material. Pure aluminum is lighter than most metals, but it bends and dents too easily. If a small amount of copper or manganese is added to aluminum, the result is a stronger material that is still lighter than steel.

For certain aircraft parts, even lighter materials are needed. Alloys of aluminum and magnesium are used for these parts. Magnesium is much less dense than most metals used to build structures. However, pure magnesium is soft enough to cut with a knife, and it burns in air. An aluminum-magnesium alloy keeps the advantages of magnesium without the disadvantages.

Section 6.4 Assessment

Reviewing Concepts
1. What holds metal ions together in a metal lattice?
2. What characteristic of a metallic bond explains some of the properties of metals?
3. How can scientists design alloys with specific properties?
4. Explain why the metallic bonds in some metals are stronger than the bonds in other metals.
5. Why are metals good conductors of electric current?
6. How does adding carbon to steel make the steel harder and stronger?

Critical Thinking
7. Predicting Which element has a higher melting point, potassium in Group 1A or calcium in Group 1B? Give a reason for your answer.
8. Applying Concepts Can two different elements form a metallic bond together?

Writing in Science: Compare-Contrast Paragraphs Write a paragraph comparing the properties of ionic compounds and alloys. Relate their properties to the structure of their lattices.

Integrate Materials Science
L2
The properties of an alloy that are determined by its composition include malleability, ductility, hardness, corrosion resistance, tensile strength (the ability to resist being pulled apart), shear strength (the ability to resist opposing forces that are not acting in a straight line), compressive strength (the ability to withstand pressures acting on a given plane), and elasticity (the ability to return to its original size and shape). Ask, What properties of steel make it useful for the cables and towers of the Golden Gate Bridge? (Its high tensile strength makes steel a useful material for the cables, and its high compressive strength makes it a useful material for the towers.) Logical

Evaluate Understanding L2
Have students write the names of the following elements on index cards: copper, gold, silver, nickel, zinc, tin, iron, carbon, and aluminum. On the other side of the cards, have students write common uses for these elements. (For example, alloys of aluminum and copper are used to make airplane parts.)

Reteach L1
Use Figure 22 to review how the structure within a metal affects the properties of a metal.

In the rigid network of a crystal lattice, the charged particles are not free to flow unless the solid melts. When struck, the crystal shatters because ions with the same charge shift position in the lattice and repel. Thus, ionic solids are brittle and are poor conductors. In a metal lattice, the electrons are free to move and conduct a current. The lattice is malleable when struck because if a cation moves it is still surrounded by particles with an opposite charge.

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 6.4.

Chemical Bonds 181

Answer to . . .

Figure 24 The compositions would probably be different because the forces acting on the cables and towers are different—tensile versus compressive, respectively.
Chipping In

Background
The silicon crystal shown in the photo is called an ingot. Pure silicon ingots can be produced by the Czochralski pulled-growth method described on p. 163. Wafers cut from the ingot can be ground, polished, cleaned, etched, doped with elements such as phosphorus and boron, and sliced with a diamond saw to produce microchips.

Build Science Skills

Applying Concepts
Purpose Students observe the importance of transistors in their lives.
Materials access to a library or the Internet, notepad for data collection
Class Time 20–30 minutes of research, one day of data collection, 20 minutes of class discussion
Procedure Have students use the library or the Internet to come up with a list of electronic devices that contain transistors. Have students record every time they use a device that contains a transistor, and the type of device. Have students compare their data in a class discussion.
Expected Outcome Devices that contain transistors include alarm clocks, watches, refrigerators, microwave ovens, washers, dryers, calculators, computers, radios, CD players, televisions, VCRs, DVD players, and video game consoles.
Intrapersonal, Kinesthetic

Chipping In

Tasks done by computers that filled an entire room in the 1950s are now done by devices the size of a credit card. This miniaturization in the electronics industry is due to semiconductors.

Semiconductors are solid substances, such as silicon, that have poor electrical conductivity at ordinary temperatures. Silicon has four valence electrons. In pure silicon, each atom forms single bonds with four other atoms. This arrangement leaves no electrons free to move through the silicon. The conductivity of silicon is greatly improved by adding small amounts of other elements to silicon, a process called doping.

Doping
An element with five valence electrons, such as phosphorus, can be added to silicon. After a phosphorus atom bonds with four atoms, there is an extra electron that is free to move. Silicon doped with phosphorus is called n-type silicon because electrons have a negative charge. An element with three valence electrons, such as boron, can be added to silicon. Adding boron leaves holes to which electrons can move from neighboring atoms. Because the lack of an electron has the effect of a positive charge, silicon with boron is called p-type silicon.

Silicon with phosphorus
Silicon with boron
Silicon wafer
Wafers and chips
Silicon crystals are cut into wafer-thin slices. Each wafer can be made into hundreds of silicon chips.
Glass vacuum tubes used in early computers were fragile and took up space.

**Transistor**

In a transistor, there are three layers of doped silicon. There is a layer of p-type silicon sandwiched between two layers of n-type silicon or a layer of n-type silicon between two layers of p-type silicon. Transistors are used to amplify current. A small current applied to the central layer of a transistor can produce a larger current.

**Diode**

In a diode, a layer of p-type silicon is joined to a layer of n-type silicon. When the leads on a diode are correctly connected in a circuit, electrons flow toward the junction between the two types of silicon. Electrons from the n-type silicon fill holes in the p-type silicon. In devices that use batteries, a diode can keep electrons from flowing if the batteries are not inserted correctly. Some diodes emit light when the circuit is complete.

**Computer chip**

In 1974, the first computer microchip contained 6000 transistors. Today, more than 40 million transistors can be placed on a single computer chip.

**Vacuum tube**

Vacuum tubes used in early computers were fragile and took up space.

**Integrated circuit on a silicon chip**

Going Further

- Research and write about the development of transistors. Why were researchers looking for a replacement for vacuum tubes? How did replacing vacuum tubes with transistors affect the size of radios and computers?
- Take a Discovery Channel Video Field Trip by watching “Good Conduct.”
Chapter 6

Chemical Bonds

How do science concepts apply to your world? Here are some questions you’ll be able to answer after you read this chapter.

- Why is titanium metal welded in an argon atmosphere? (Section 6.1)
- What causes a crystal of rock salt to shatter when it is struck? (Section 6.1)
- Why is water a liquid while carbon dioxide is a gas at room temperature? (Section 6.2)
- What advantage does jewelry made from a gold silver alloy have over jewelry made from pure gold? (Section 6.4)
- How does mixing other elements with silicon make silicon a better conductor of electric current? (Page 182)

Chapter Pretest

1. Describe the structure of atoms. (*An atom consists of a dense, positively charged nucleus containing protons and neutrons, surrounded by space in which negatively charged electrons move.*)
2. True or False: Objects with opposite charges attract one another. (*True*)
3. What are valence electrons? (*Electrons in the highest occupied energy level of an atom*)
4. Which group in the periodic table contains elements that hardly react at all? (*The noble gases*)
5. Where on the periodic table are nonmetals generally found? (*The right side*)
6. How do the compositions of mixtures differ from those of substances? (*The composition of a mixture can vary. The composition of a substance is fixed.*)
7. What property is being described when someone says that a solid is easily hammered into sheets? (*b*)
   a. Conductivity
   b. Malleability
   c. Melting point
   d. Density
Chapter Preview
6.1 Ionic Bonding  6.4 The Structure of Metals
6.2 Covalent Bonding
6.3 Naming Compounds and Writing Formulas

Chapter Preview
What Can the Shape of a Material Tell You About the Material?

Purpose
In this activity, students begin to recognize that the shape of pieces of a compound could be related to its structure at the atomic level.

Skills Focus
Observing, Comparing and Contrasting

Prep Time
10 minutes

Materials
4 wood splints or small lab spatulas, sodium chloride, black construction paper, hand lens, alum, Epsom salts, sucrose

Class Time
15 minutes

Safety
Students should wear safety goggles, lab aprons, and disposable plastic gloves.

Teaching Tips
• Point out that if all the samples of a substance have a similar shape, the shape may be a property of the substance.
• Encourage students to think about the spatial arrangement of the particles in a substance.

Expected Outcome
Each substance forms crystals with a particular and different shape.

Think About It
1. Sodium chloride crystals are cubes. Alum crystals are small and irregular “cuboids” shaped like a cube that is missing part of a side. Crystals of Epsom salts look like cylinders with one or both ends pointed. Sucrose crystals may appear oblong with slightly slanted ends.
2. Although the samples have distinctive shapes, the shapes are not sufficiently distinct to help a person identify each substance. (Many crystals have the same general shape.)
3. Students may hypothesize that the arrangement of particles at the atomic level might be responsible for the different crystal shapes.

Visual, Logical

Encourage students to view the Video Field Trip “Good Conduct.”
Improving the Dyeing of Nonpolar Fabrics

**Objective**
After completing this activity, students will be able to
- relate the ability of fabrics to absorb and retain dyes to the polar or nonpolar character of their fibers.

**Skills Focus** Observing

**Prep Time** 40 minutes

**Advance Prep** Wear plastic gloves, a lab apron, and a dust mask when preparing the solutions. To prepare methyl orange dye solution, add 2.1 g Na₂SO₄ and 6–9 drops concentrated H₂SO₄ to 450 mL of water. Then, add 1.50 g methyl orange powder to the solution. Stir the solution. To prepare iron(II) sulfate solution, add 12.6 g FeSO₄·H₂O to 450 mL water. Stir the solution. Place the solutions in 1-L beakers on hot plates and heat them in advance to near boiling. Arrange hot plates with the two solutions at several lab stations. Test fabrics that consist of a repeating pattern of wool, polyester, nylon, cotton, and various other natural and synthetic fibers may be obtained from a commercial fabric company. You may also make up your own test strips from fabric remnants.

**Class Time** 45 minutes, if different lab groups perform Parts A and B simultaneously. Alternatively, the work can be done by all groups over the course of two days.

**Safety** Review the safety information on the MSDS for methyl orange and iron(II) sulfate with students. Make sure students wear their goggles, gloves, and lab aprons at all times during the investigation. Emphasize the importance of protective clothing for this investigation. Any spills should be wiped up promptly to avoid slips and falls. The hot plates and furniture should be placed so that students will not bump into them as they move around. Do not use flames to heat the solutions. Advise students to wear old clothes suitable for painting on the day of this investigation. Stains from methyl orange may be permanent.

---

**Problem** How can you increase the dye-holding capacity of nonpolar fibers?

**Materials**
- tongs
- 2 fabric test strips
- hot dye bath containing methyl orange
- clock or watch
- paper towels
- scissors
- soap
- hot iron(II) sulfate solution

**Skills** Observing, Drawing Conclusions

**Procedure**

**Part A: Dyeing Without Treatment**
1. On a sheet of paper, copy the data table shown.
2. Use the tongs to immerse a fabric test strip in the methyl orange dye bath. **CAUTION:** The dye bath is hot. Do not touch the glass. The dye will stain skin and clothing.

<table>
<thead>
<tr>
<th>Data Table</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dye Treatment</td>
</tr>
<tr>
<td>Methyl orange</td>
</tr>
</tbody>
</table>

3. After 7 minutes, remove the strip from the dye bath. Allow as much of the dye solution as possible to drip back into the bath as shown on page 185. Rinse off the excess dye with water in the sink.
4. Place the strip on a paper towel to dry. Be careful to avoid splashes when transferring the strip between the dye bath and paper towel. Record your observations in your data table.
5. After the fabric strip is dry, test it for colorfastness, or the ability to hold dye. Cut the strip in half lengthwise and wash one half of the strip in the sink with soap and water.

---

**Part B: Dyeing With Treatment**

1. Place 12.6 g FeSO₄·H₂O in a 1-L beaker on a hot plate. Heat it to near boiling. Allow the dye solution to cool to room temperature. Then, dip the fabric test strips into the iron(II) sulfate solution.
2. Use the tongs to immerse a fabric test strip in the methyl orange dye solution. **CAUTION:** The dye bath is hot. Do not touch the glass. The dye will stain skin and clothing.

<table>
<thead>
<tr>
<th>Data Table</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dye Treatment</td>
</tr>
<tr>
<td>Methyl orange</td>
</tr>
</tbody>
</table>

3. After 7 minutes, remove the strip from the dye bath. Allow as much of the dye solution as possible to drip back into the bath as shown on page 185. Rinse off the excess dye with water in the sink.
4. Place the strip on a paper towel to dry. Be careful to avoid splashes when transferring the strip between the dye bath and paper towel. Record your observations in your data table.
5. After the fabric strip is dry, test it for colorfastness, or the ability to hold dye. Cut the strip in half lengthwise and wash one half of the strip in the sink with soap and water.

---

**Conclusion**
A comparison of the dyeing with and without treatment will be made based on colorfastness and color depth. The results will be recorded in the data tables.
6. Allow the washed half-strip to dry and then compare the washed half to the unwashed half. Record your observations in your data table. Staple the half-strips to a sheet of paper and label each half-strip to indicate how you treated it.

Part B: Dyeing With Treatment

7. Use the tongs to place the second fabric strip in the iron(II) sulfate solution for 25 minutes. Then use tongs to lift the strip and allow it to drain into the iron(II) sulfate solution. Wiring the strip as dry as possible over the solution.

CAUTION The strip will be hot. Allow it to cool before touching it. Wear plastic gloves.

8. Repeat Steps 2, 3, and 4 using the strip that you treated with iron(II) sulfate.

9. To test the strip for colorfastness, repeat Steps 5 and 6.

10. Clean up your work area and wash your hands thoroughly with warm water and soap before leaving the laboratory.

Analyse and Conclude

1. **Comparing and Contrasting** How did the color of the untreated strip compare with the color of the treated strip?

2. **Comparing and Contrasting** How did the colorfastness of the untreated strip compare to the colorfastness of the treated strip?

3. **Applying Concepts** Silk blouses and shirts can be purchased in many intense colors. Why do you think silk is able to hold a variety of intense dyes?

4. **Drawing Conclusions** How does iron(II) sulfate affect the ability of a fabric to absorb dyes? (Hint: What kind of compound is iron(II) sulfate?)

5. **Predicting** A care label might say Wash in cold water only. What might happen to the color of a piece of clothing with this label if you washed the clothing in hot water?

Teaching Tips

- Explain the role of ionic and polar regions in the dyeing process. In Part A, students perform direct dyeing, which involves the ionic sites of the dye molecules attaching to ionic sites in the fabric. In Part B, ionic sites are introduced into the synthetic fibers so that more dye can bond to the fibers.

- Identify the different fibers in the fabric test strips for students.

- Make sure that the dye and iron(II) sulfate baths stay close to the boiling point, but do not let them boil, as they may spatter.

**Expected Outcome**

Natural fibers absorb dye more easily and are more colorfast than synthetic fibers. Iron(II) sulfate treatment improves the dyeing and colorfastness of synthetic fibers.

**Analyze and Conclude**

1. Treatment increased the intensity of the colors for synthetic fibers.

2. In general, treatment improved colorfastness, especially in synthetic fabrics.

3. Silk is a natural fiber. Like most natural fibers, it has many polar regions that allow the dye to attach to the fibers.

4. Iron(II) sulfate is an ionic compound. Treatment with this compound adds ionic sites to the fabric, increasing its ability to bind dyes.

5. The hot water would remove more dye than would cold water, and the colors would run.

Visual, Kinesthetic

---

### Sample Data Table

<table>
<thead>
<tr>
<th>Dye Treatment</th>
<th>Dyeing of Fibers</th>
<th>Colorfastness of Fibers</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methyl orange</td>
<td>Intensity of color varies; wool is strongest, synthetics weakest.</td>
<td>Wool and cotton are colorfast; synthetic fibers are not.</td>
</tr>
<tr>
<td>Iron sulfate and methyl orange</td>
<td>Intensity of color varies, but increase in intensity is strongest in synthetics.</td>
<td>All fibers are colorfast.</td>
</tr>
</tbody>
</table>
### Chapter 6 Planning Guide

<table>
<thead>
<tr>
<th>SECTION OBJECTIVES</th>
<th>STANDARDS NATIONAL</th>
<th>ACTIVITIES and LABS</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>6.1 Ionic Bonding, pp. 158–164</strong></td>
<td>A-1, B-1, B-2, G-1</td>
<td><strong>SE</strong> Inquiry Activity: What Can the Shape of a Material Tell You About the Material? p. 157</td>
</tr>
<tr>
<td>6.1.1 Recognize stable electron configurations.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.1.2 Predict an element’s chemical properties using number of valence electrons and electron dot diagrams.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.1.3 Describe how an ionic bond forms and how ionization energy affects the process.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.1.4 Predict the composition of an ionic compound from its chemical formula.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.1.5 Relate the properties of ionic compounds to the structure of crystal lattices.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.2.1 Describe how covalent bonds form and the attractions that keep atoms together in molecules.</td>
<td></td>
<td><strong>SE</strong> Consumer Lab: Improving the Dyeing of Nonpolar Fabrics, pp. 184–185</td>
</tr>
<tr>
<td>6.2.2 Compare polar and nonpolar bonds, and demonstrate how polar bonds affect the polarity of a molecule.</td>
<td></td>
<td><strong>TE</strong> Teacher Demo: Modeling Overall Polarity, p. 168</td>
</tr>
<tr>
<td>6.2.3 Compare the attractions between polar and nonpolar molecules.</td>
<td></td>
<td><strong>TE</strong> Teacher Demo: Surface Tension, p. 169</td>
</tr>
<tr>
<td><strong>6.3 Naming Compounds and Writing Formulas, pp. 170–175</strong></td>
<td>A-1, A-2, B-2, G-2</td>
<td><strong>SE</strong> Quick Lab: Modeling Molecules, p. 173</td>
</tr>
<tr>
<td>6.3.1 Recognize and describe binary ionic compounds, metals with multiple ions, and polyatomic ions.</td>
<td></td>
<td><strong>LM</strong> Investigation 6A: Playing the Ionic Compounds Card Game</td>
</tr>
<tr>
<td>6.3.2 Name and determine chemical formulas for ionic and molecular compounds.</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>6.4 The Structure of Metals, pp. 176–181</strong></td>
<td>A-1, A-2, B-2, B-6, E-2, G-1, G-2, G-3</td>
<td><strong>TE</strong> Teacher Demo: Comparing Bond Types, p. 177</td>
</tr>
<tr>
<td>6.4.1 Describe the structure and strength of bonds in metals.</td>
<td></td>
<td><strong>TE</strong> Teacher Demo: Bronze and Brass Tones, p. 180</td>
</tr>
<tr>
<td>6.4.2 Relate the properties of metals to their structure.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.4.3 Define an alloy and demonstrate how the composition of an alloy affects its properties.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Materials for Activities and Labs

Quantities for each group

**STUDENT EDITION**

- Inquiry Activity, p. 157
  4 wood splints or small lab spatulas, sodium chloride, black construction paper, hand lens, alum, Epsom salts, sucrose
- Quick Lab, p. 167
  test paper, metric ruler, felt-tip markers, stapler, beaker, alcohol-water mixture, Petri dish
- Quick Lab, p. 173
  blue plastic-foam ball, black plastic-foam ball, 7 white gumdrops, toothpicks
- Consumer Lab, pp. 184–185
  tongs, 2 fabric test strips, hot dye bath containing methyl orange, clock or watch, paper towels, scissors, soap, hot iron(II) sulfate solution

**TEACHER’S EDITION**

- Teacher Demo, p. 168
  molecular model kit, 4 12-inch pieces of string or yarn, tape, overhead projector
- Teacher Demo, p. 169
  200-mL beaker, water, sewing needle, tweezers, dropper pipet
- Teacher Demo, p. 177
  salt lick (or rock salt), copper wire, hammer, goggles
- Teacher Demo, p. 180
  brass bell, bronze bell
- Build Science Skills, p. 182
  access to a library or the Internet, notepad for data collection
A bond in an ionic compound forms when electrons are transferred from metal atoms to nonmetal atoms. The ions that form are in a three-dimensional array called a crystal lattice. Within the lattice, there are strong attractions between an ion and neighboring ions of opposite charge. The strength of the bonds depends on the arrangement of the ions, their sizes, and their charges. Crystals are classified by the angles at which the faces (sides) of the crystal meet and by how many of the edges on a face are of equal length. Most binary ionic compounds, such as sodium chloride, have cube-shaped crystals. In a cubic crystal, all of the edges on a face are of equal length and all the faces meet at a 90° angle. A unit cell (like the one shown below) is the smallest unit of a crystal that shows the arrangement of ions. For a given arrangement, the bond strength increases as the charges on the ions increase and as the distance between ions decreases (with smaller ionic radii).

Electronegativity is a quantity that describes an atom's ability to attract the electrons within a bond. Values for electronegativity are often based on values for ionization energy and electron affinity. Ionization energy is a measure of an atom's ability to lose electrons. Electron affinity is a measure of an atom's ability to gain electrons.
Differences in electronegativity are used to predict the type of bond that will form between atoms of two elements. As the difference in electronegativity increases, the bond becomes more polar until the electrons are transferred and an ionic bond forms.

### Naming Compounds 6.3

The preferred method for naming ionic compounds of metals that have variable charge is the Stock system, which uses Roman numerals to represent the charge on the ion. However, an older system that uses a root word combined with different suffixes still persists for metals that have only two charges. In this older system, the suffix -ous is used for the ion with the lesser charge and the suffix -ic is used for the ion with the greater charge. Thus, an iron(II) ion is a ferrous ion and an iron(III) ion is a ferric ion. The older system does not indicate the actual charge on the ion.

### Metallic Bonding 6.4

The valence electrons in a metal are often referred to as a “sea of electrons” because of their mobility. In general, the strength of a metallic bond increases with the number of valence electrons. A metal’s melting point increases with bond strength. However, melting points for transition metals peak at Group 6B even though the number of valence electrons is still increasing across the period. According to quantum theory, valence electrons in metals are in a band of overlapping atomic orbitals. Properties of metals depend on electrons moving from lower to higher energy orbitals when energy is absorbed. The configuration in Group 6B metals is ideal because all the lower energy orbitals are filled while all the higher energy orbitals are unoccupied.
### Study Guide

#### 6.1 Ionic Bonding

**Key Concepts**
- When the highest occupied energy level of an atom is filled with electrons, the atom is stable and not likely to react.
- Some elements achieve stable electron configurations through the transfer of electrons between atoms. An ionic bond forms when electrons are transferred from one atom to another.
- The properties of an ionic compound can be explained by the strong attractions among ions within a crystal lattice.

**Vocabulary**
- electron dot diagram, p. 158
- ion, p. 159
- anion, p. 160
- cation, p. 160
- chemical bond, p. 160
- ionic bond, p. 160
- chemical formula, p. 161
- crystals, p. 162

#### 6.2 Covalent Bonding

**Key Concepts**
- The attractions between the shared electrons and the protons in each nucleus hold the atoms together in a covalent bond.
- When atoms form a polar covalent bond, the atom with the greater attraction for electrons has a partial negative charge. The other atom has a partial positive charge.
- The type of atoms in a molecule and its shape are factors that determine whether a molecule is polar or nonpolar.
- Attractions between polar molecules are stronger than attractions between nonpolar molecules.

**Vocabulary**
- covalent bond, p. 166
- molecule, p. 166
- polar covalent bond, p. 168

#### 6.3 Naming Compounds and Writing Formulas

**Key Concepts**
- The name of an ionic compound must distinguish the compound from other ionic compounds containing the same elements. The formula of an ionic compound describes the ratio of the ions in the compound.
- The name and formula of a molecular compound describe the type and number of atoms in a molecule of the compound.

**Vocabulary**
- polyatomic ion, p. 172

#### 6.4 The Structure of Metals

**Key Concepts**
- The cations in a metal form a lattice that is held in place by strong metallic bonds between the cations and the surrounding valence electrons.
- The mobility of electrons within a metal lattice explains some of the properties of metals.
- Scientists can design alloys with specific properties by varying the types and amounts of elements in an alloy.

**Vocabulary**
- metallic bond, p. 176
- alloy, p. 178

**Thinking Visually**

**Concept Map** Use information from the chapter to complete the concept map below.

---

**Chapter Resources**

**Print**
- Chapter and Unit Tests, Chapter 6
- Test A and Test B
- Test Prep Resources, Chapter 6

**Technology**
- Computer Test Bank, Chapter 6
- Interactive Textbook, Chapter 6
- Go Online, PHSchool.com, Chapter 6
Reviewing Content

Choose the letter that best answers the question or completes the statement.

1. When an atom loses an electron, it forms a(n)
   a. anion. b. cation. c. polyatomic ion. d. neutral ion.

2. The charge on a chloride ion in AlCl₃ is
   a. 1+. b. 3+. c. 1-. d. 3-.

3. Which pair has the same electron configuration?
   a. Cl⁻ and Ar b. Cl⁻ and Ar²⁺ c. Cl⁻ and Ar²⁺ d. Cl and Ar²⁺

4. A chemical bond that forms when atoms share electrons is always a(n)

5. When two fluorine atoms share a pair of electrons, the bond that forms is a(n)

6. The chemical formula for magnesium bromide is
   a. MgBr₂ b. MgBr₃ c. Mg₂Br₃ d. Mg₂Br₄

7. The compound with the formula SiCl₄ is
   a. silicon chloride b. silicon tetrachloride c. silicon(IV) chloride d. silicon tetrachloride.

8. The attraction among water molecules is stronger than the attraction among
   a. sodium and chloride ions. b. carbon dioxide molecules. c. the atoms in a polyatomic ion. d. atoms in a diatomic molecule.

9. Which type of solid is likely to be the best conductor of electric current?
   a. metal element b. covalent compound c. ionic compound d. nonmetal element.

10. An alloy contains
    a. at least one metallic element. b. at least one nonmetallic element. c. only metallic elements. d. only nonmetallic elements.

Understanding Concepts

11. What is a stable electron configuration?
12. What does each dot in an electron dot diagram represent?
13. What process changes atoms into ions?
14. What keeps the ions in their fixed positions within a crystal lattice?
15. What are subscripts used for in chemical formulas?
16. Explain why a melted ionic compound is a good conductor of electric current, but a solid ionic compound is a poor conductor of electric current.
17. What distinguishes single, double, and triple covalent bonds?
18. Explain why the covalent bonds in molecules of elements are always nonpolar.
19. Explain why, in a covalent bond between oxygen and hydrogen, the hydrogen atom has a partial positive charge and the oxygen atom has a partial negative charge.
20. What is the name of the binary compound formed from potassium and iodine?
21. Write the formulas for the compounds called copper(I) chloride and copper(II) chloride.
22. Name the compounds represented by the space-filling models labeled A, B, and C.

Assessment

If your class subscribes to the Interactive Textbook, your students can go online to access an interactive version of the Student Edition and a self-test.

11. In a stable electron configuration, the highest occupied energy level is filled with electrons.
12. Each dot represents a valence electron.
13. The transfer of electrons
14. Attractions between neighboring cations and anions keep the ions in fixed positions within the lattice.
15. A subscript is used to show the number of atoms of an element in a molecule or the ratio of ions in a crystal lattice.
16. When an ionic compound melts, ions can move away from their fixed locations in the crystal lattice.
17. Two atoms share two electrons in a single bond, four in a double bond, and six in a triple bond.
18. The covalent bonds in molecules of elements are always nonpolar because the atoms have the same attraction for electrons.
19. The oxygen atom has a greater attraction for electrons than the hydrogen atom does.
20. Potassium iodide
21. CuCl and CuCl₂
22. A is sulfur trioxide, B is carbon monoxide, and C is nitrogen dioxide.
23. In general, the more valence electrons a metal has, the stronger the metal bonds are.
24. A mixture of copper and tin is harder and stronger than either metal in its pure form.
25. The advantage that is retained is that magnesium is a lightweight metal. The disadvantage that is reduced is that magnesium is a soft metal.
**Critical Thinking**

26. All three have the same electron configuration.
27. Molecules and polyatomic ions both contain covalent bonds.
29. Sulfur dichloride, silver(I) sulfate, lithium fluoride, carbon disulfide, calcium hydroxide
30. Q is a metal. X and Z are nonmetals.
31. QX and QZ
32. Cr₂Z₃
33. Z : Z
34. Nonpolar covalent bond

**Math Skills**

35. 8
36. The ratio is two to one anions to cations.
37. BaF₂, Na₂O, FeSO₄, and (NH₄)₂SO₄

**Concepts in Action**

38. The correct formula is B.
39. Molecules of carbon dioxide are nonpolar. Molecules of water are polar.
40. The carbonate ion is a polyatomic ion, which contains covalent bonds.
41. Phosphorus has five valence electrons compared to four valence electrons in silicon. The extra electrons are not needed to bond the atoms together and are free to move and carry the current.
42. Ideally, students should note that in both cases the electron will be sharing space with a second electron in orbitals that overlap. In the bond between hydrogen atoms, the electron will be equally attracted to both nuclei. In the bond between hydrogen and oxygen, the electron will have a greater attraction to the oxygen nucleus.

**Critical Thinking**

26. Classifying What does a fluoride ion have in common with a neon atom and a sodium ion?
27. Comparing and Contrasting How are molecules and polyatomic ions similar?
28. Classifying Classify the bonds in each of these compounds as ionic, polar covalent, or nonpolar covalent: SO₂, CaO, and I₂.
29. Applying Concepts Write the names for the compounds with these chemical formulas: Sc₂O₃, Ag₃SO₄, LiF, Cs₂, and Ca(OH)₂.

**Math Skills**

30. Using Models Which of the three elements are metals and which are nonmetals?
31. Applying Concepts Element Q forms compounds with element X and with element Z. Write the formulas for these two compounds.
32. Calculating What would the formula be for a compound containing chromium(III) ions and ions of element Z?
33. Applying Concepts Draw an electron dot structure for a compound of fluorine and Z.
34. Predicting If an atom of X reacts with an atom of Z, what kind of bond forms?

**Concepts In Action**

38. Using Models A solution of hydrogen peroxide (H₂O₂) and water is sometimes used to disinfect a cut. Which of the following formulas is the correct structural formula for hydrogen peroxide?

39. Relating Cause and Effect In a carbonated beverage, the main ingredients are water and carbon dioxide. Carbon dioxide gas is released when the bottle is opened. Why is water a liquid but carbon dioxide a gas at room temperature?
40. Classifying The shells shown on page 156 contain the compound calcium carbonate (CaCO₃). Explain how this compound can contain both ionic and covalent bonds.
41. Relating Cause and Effect How does adding some phosphorus to silicon make silicon a better conductor of electric current?
42. Writing in Science Compare what happens to the valence electron in a hydrogen atom when the atom bonds with another hydrogen atom and when the atom bonds with an oxygen atom.

**Performance-Based Assessment**
**Designing an Advertisement** You own a store that sells bronze bells. Design a quarter-page ad for your store to be published in your local directory of businesses. Write copy for your ad. Describe a photograph to use in the ad. Also supply a sketch showing how you want the copy and the photograph to be laid out on the page.

**Go Online**

Your students can independently test their knowledge of the chapter and print out their test results for your files.

**Performance-Based Assessment**

The ad could include properties of bronze, such as hardness and durability. The ad should stress the sound produced by a bronze bell. (Have examples of ads in a local directory of businesses for students to look at. Students may want to research the cost of a quarter-page ad in the directory.)
Test-Taking Tip

Paying Attention to the Details

Sometimes two or more answers to a question are almost identical. If you do not read the answers carefully, you may select an incorrect answer by mistake. In the question below, all the answers include the correct elements in the correct order—metal before nonmetal. However, only one of the answers uses the correct rules for naming CaCl2.

The name for the compound with the formula CaCl2 is

(A) calcium(II) chloride.
(B) calcium chlorine.
(C) calcium dichloride.
(D) calcium chloride.
(E) monocalcium dichloride.

(Answer: D)

Choose the letter that best answers the question or completes the statement.

1. How many electrons does a Group 7A atom need to gain in order to achieve a stable electron configuration?
   (A) 0
   (B) 1
   (C) 2
   (D) 7
   (E) 8

2. What type of bond forms when electrons are transferred from one atom to another?
   (A) nonpolar covalent bond
   (B) ionic bond
   (C) polar covalent bond
   (D) polyatomic bond
   (E) metallic bond

3. Metallic bonds form between
   (A) cations and protons.
   (B) cations and anions.
   (C) cations and anions.
   (D) cations and electrons.
   (E) anions and electrons.

4. What is the formula for copper(II) nitrate?
   (A) CuNO3
   (B) Cu2(NO3)2
   (C) Cu(NO3)2
   (D) Cu2NO3
   (E) CuNO2

5. In the compound iron(II) carbonate, the ratio of iron(II) ions to carbonate ions will be
   (A) one to one.
   (B) two to one.
   (C) three to one.
   (D) one to two.
   (E) one to three.

6. All steels contain
   (A) copper and zinc.
   (B) copper and tin.
   (C) iron and chromium.
   (D) chromium and carbon.
   (E) iron and carbon.

7. What is the reason that water has a higher boiling point than expected?
   (A) Attractions among nonpolar water molecules are strong.
   (B) Water molecules have a linear shape.
   (C) Water molecules are not very massive.
   (D) There are strong attractions among polar water molecules.
   (E) There are no attractions among water molecules.

<table>
<thead>
<tr>
<th>Ion Name</th>
<th>Ion Symbol</th>
<th>Ion Name</th>
<th>Ion Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper(I)</td>
<td>Cu⁺</td>
<td>Nitrate</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>Copper(II)</td>
<td>Cu²⁺</td>
<td>Sulfate</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>Iron(II)</td>
<td>Fe²⁺</td>
<td>Carbonate</td>
<td>CO₃²⁻</td>
</tr>
<tr>
<td>Iron(III)</td>
<td>Fe³⁺</td>
<td>Phosphate</td>
<td>PO₄³⁻</td>
</tr>
</tbody>
</table>

Some Ions and Their Symbols

Use the table to answer Questions 4 and 5.

Chemical Bonds 189