CHAPTER

5

Ions and Ionic Compounds

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The photograph provides a striking view of an ordinary substance—sodium chloride, more commonly known as table salt. Sodium chloride, like thousands of other compounds, is usually found in the form of crystals. These crystals are made of simple patterns of ions that are repeated over and over, and the result is often a beautifully symmetrical shape.

Ionic compounds share many interesting characteristics in addition to the tendency to form crystals. In this chapter you will learn about ions, the compounds they form, and the characteristics that these compounds share.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Hard Water

PROCEDURE

1. Fill two **14 × 100 test tubes** halfway with **distilled water** and a third test tube with **tap water.**
2. Add about 1 tsp **Epsom salts** to one of the test tubes containing distilled water to make "hard water." Label the appropriate test tubes "Distilled water," "Tap water," and "Hard water."
3. Add a squirt of **liquid soap** to each test tube. Take one test tube, stopper it with a **cork,** and shake vigorously for 15 s. Repeat with the other two test tubes.
4. Observe the suds produced in each test tube.

ANALYSIS

1. Which water sample produces the most suds? Which produces the least suds?
2. What is meant by the term "hard water"? Is the water from your tap "hard water"?

Pre-Reading Questions

1. **What is the difference between an atom and an ion?**
2. **How can an atom become an ion?**
3. **Why do chemists call table salt sodium chloride?**
4. **Why do chemists write the formula for sodium chloride as NaCl?**

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SECTION

1 Simple Ions

KEY TERMS

- octet rule
- ion
- cation
- anion

OBJECTIVES

1. **Relate** the electron configuration of an atom to its chemical reactivity.
2. **Determine** an atom's number of valence electrons, and use the octet rule to predict what stable ions the atom is likely to form.
3. **Explain** why the properties of ions differ from those of their parent atoms.

Chemical Reactivity

Some elements are highly reactive, while others are not. For example, **Figure 1** compares the difference in reactivity between oxygen and neon. Notice that oxygen reacts readily with magnesium, but neon does not. Why is oxygen so reactive while neon is not? How much an element reacts depends on the electron configuration of its atoms. Examine the electron configuration for oxygen.

$$[\mathrm{O}] = 1s^2 2s^2 2p^4$$

Notice that the $2p$ orbitals, which can hold six electrons, have only four. The electron configuration of a neon atom is shown below.

$$[\mathrm{Ne}] = 1s^2 2s^2 2p^6$$

Notice that the $2p$ orbitals in a neon atom are full with six electrons.

a magnesium in oxygen

b magnesium in neon

Figure 1 Because of its electron configuration, oxygen reacts readily with magnesium (**a**). In contrast, neon's electron configuration makes it unreactive (**b**).

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Noble Gases Are the Least Reactive Elements

Neon is a member of the noble gases, which are found in Group 18 of the periodic table. The noble gases show almost no chemical reactivity. Because of this, noble gases have a number of uses. For example, helium is used to fill balloons that float in air, which range in size from party balloons to blimps. Like neon, helium will not react with the oxygen in the air. The electron configuration for helium is $1s^2$. The two electrons fill the first energy level, making helium stable.

The other noble gases also have filled outer energy levels. This electron configuration can be written as ns^2np^6 where n represents the outer energy level. Notice that this level has eight electrons. These eight electrons fill the *s* and *p* orbitals, making these noble gases stable. In most chemical reactions, atoms tend to match the *s* and *p* electron configurations of the noble gases. This tendency is called the **octet rule.**

Topic Link

Refer to the "Periodic Table" chapter for a discussion of the stability of the noble gases.

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octet rule

a concept of chemical bonding theory that is based on the assumption that atoms tend to have either empty valence shells or full valence shells of eight electrons

Alkali Metals and Halogens Are the Most Reactive Elements

Based on the octet rule, an atom whose outer *s* and *p* orbitals do not match the electron configurations of a noble gas will react to lose or gain electrons so the outer orbitals will be full. This prediction holds true for the alkali metals, which are some of the most reactive elements. **Figure 2** shows what happens when potassium, an alkali metal, is dropped into water. An explosive reaction occurs immediately, releasing heat and light.

As members of Group 1, alkali metals have only one electron in their outer energy level. When added to water, a potassium atom gives up this electron in its outer energy level. Then, potassium will have the *s* and *p* configuration of a noble gas.

$$1s^22s^22p^63s^23p^64s^1 \longrightarrow 1s^22s^22p^63s^23p^6$$

The halogens are also very reactive. As members of Group 17, they have seven electrons in their outer energy level. By gaining just one electron, a halogen will have the *s* and *p* configuration of a noble gas. For example, by gaining one electron, chlorine's electron configuration becomes $1s^22s^22p^63s^23p^6$.

Figure 2
Alkali metals, such as potassium, react readily with a number of substances, including water.

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Valence Electrons

Topic Link

Refer to the "Periodic Table" chapter for more about valence electrons.

You may have noticed that the electron configuration of potassium after it loses one electron is the same as that of chlorine after it gains one. Also, both configurations are the same as that of the noble gas argon.

$$[\text{Ar}] = 1s^22s^22p^63s^23p^6$$

After reacting, both potassium and chlorine have become stable. The atoms of many elements become stable by achieving the electron configuration of a noble gas. These electrons in the outer energy level are known as valence electrons.

Periodic Table Reveals an Atom's Number of Valence Electrons

It is easy to find out how many valence electrons an atom has. All you have to do is check the periodic table.

For example, **Figure 3** highlights the element magnesium, Mg. The periodic table lists its electron configuration.

$$[\text{Mg}] = [\text{Ne}]3s^2$$

This configuration shows that a magnesium atom has two valence electrons in the 3*s* orbital.

Now check the electron configuration of phosphorus, which is also highlighted in **Figure 3.**

$$[\text{P}] = [\text{Ne}]3s^23p^3$$

This configuration shows that a phosphorus atom has five valence electrons. Two valence electrons are in the 3*s* orbital, and three others are in the 3*p* orbitals.

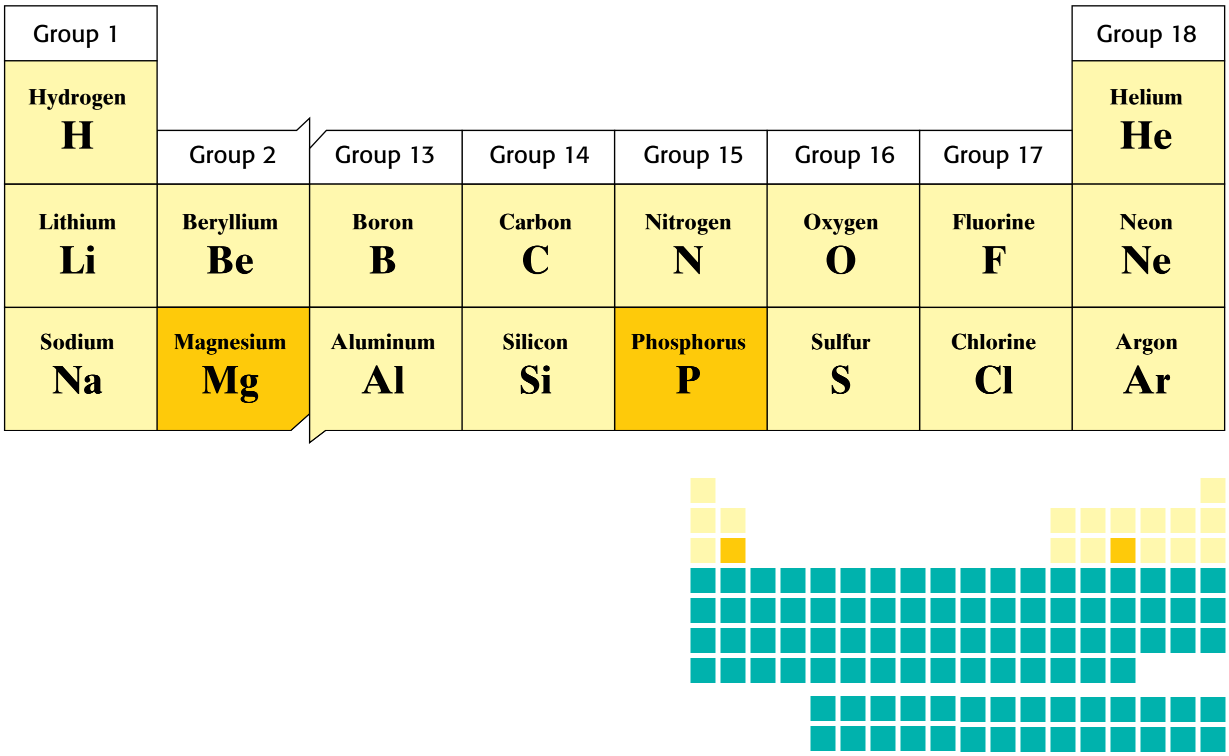

Figure 3
The periodic table shows the electron configuration of each element. The number of electrons in the outermost energy level is the number of valence electrons.

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Atoms Gain Or Lose Electrons to Form Stable Ions

Recall that potassium loses its one valence electron so it will have the electron configuration of a noble gas. But why doesn't a potassium atom gain seven more electrons to become stable instead? The reason is the energy that is involved. Removing one electron requires far less energy than adding seven more.

When it gives up one electron to be more stable, a potassium atom also changes in another way. Recall that all atoms are uncharged because they have equal numbers of protons and electrons. For example, a potassium atom has 19 protons and 19 electrons. After giving up one electron, potassium still has 19 protons but only 18 electrons. Because the numbers are not the same, there is a net electrical charge. So the potassium atom becomes an **ion** with a 1+ charge, as shown in **Figure 4.** The following equation shows how a potassium atom forms an ion.

$$K \longrightarrow K^+ + e^-$$

An ion with a positive charge is called a **cation.** A potassium cation has an electron configuration just like the noble gas argon.

$$[K^+] = 1s^22s^22p^63s^23p^6 \qquad [Ar] = 1s^22s^22p^63s^23p^6$$

In the case of chlorine, far less energy is required for an atom to gain one electron rather than give up its seven valence electrons. By gaining an electron to be more stable, a chlorine atom becomes an ion with a 1– charge, as illustrated in **Figure 4.** The following equation shows the formation of a chlorine ion from a chlorine atom.

$$Cl + e^- \longrightarrow Cl^-$$

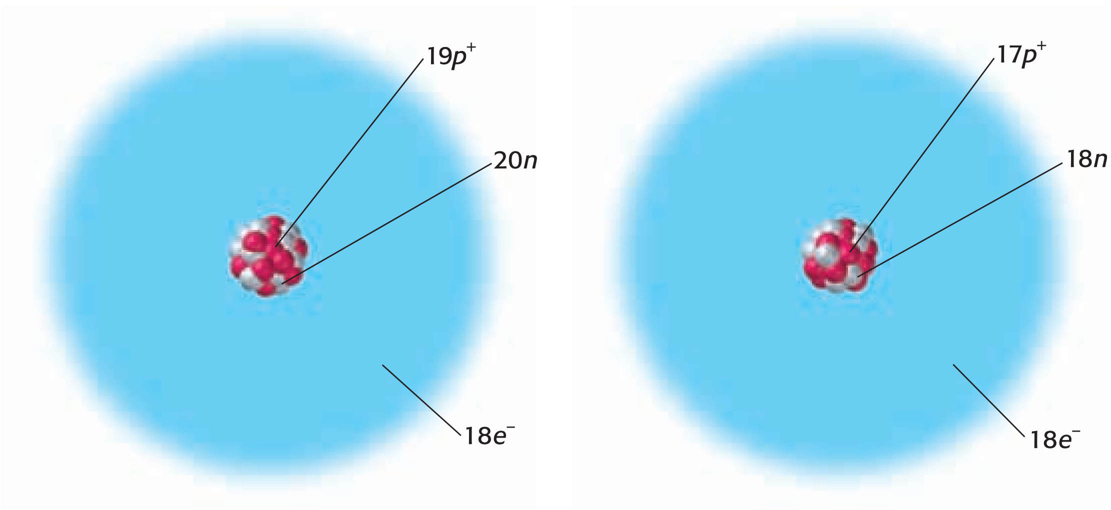
An ion with a negative charge is called an **anion.** A chlorine anion has an electron configuration just like the noble gas argon.

$$[Cl^-] = 1s^22s^22p^63s^23p^6 \qquad [Ar] = 1s^22s^22p^63s^23p^6$$

ion
an atom, radical, or molecule that has gained or lost one or more electrons and has a negative or positive charge

cation
an ion that has a positive charge

anion
an ion that has a negative charge



a potassium cation, K^+

b chloride anion, Cl^-

Figure 4
A potassium atom can lose an electron to become a potassium cation (**a**) with a 1+ charge. After gaining an electron, a chlorine atom becomes a chlorine anion (**b**) with a 1– charge.

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Characteristics of Stable Ions

How does an atom compare to the ion that it forms after it loses or gains an electron? Use of the same name for the atom and the ion that it forms indicates that the nucleus is the same as it was before. Both the atom and the ion have the same number of protons and neutrons. When an atom becomes an ion, it only involves loss or gain of electrons.

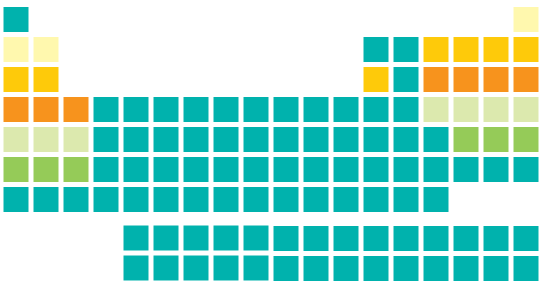
Recall that the chemical properties of an atom depend on the number and configuration of its electrons. Therefore, an atom and its ion have different chemical properties. For example, a potassium cation has a different number of electrons from a neutral potassium atom, but the same number of electrons as an argon atom. A chlorine anion also has the same number of electrons as an argon atom. However, it is important to realize that an ion is still quite different from a noble gas. An ion has an electrical charge, so therefore it forms compounds, and also conducts electricity when dissolved in water. Noble gases are very unreactive and have none of these properties.

Figure 5
These are examples of some stable ions that have an electron configuration like that of a noble gas.

Some Ions with Noble-Gas Configurations

Group 1	Group 2	Group 3	Group 13	Group 15	Group 16	Group 17	Group 18 Noble Gases
							Helium He $1s^2$
Li^+ $1s^2$	Be^{2+} $1s^2$			N^{3-} $[He]2s^22p^6$	O^{2-} $[He]2s^22p^6$	F^- $[He]2s^22p^6$	Neon Ne $[He]2s^22p^6$
Na^+ $[He]2s^22p^6$	Mg^{2+} $[He]2s^22p^6$		Al^{3+} $[He]2s^22p^6$	P^{3-} $[Ne]3s^23p^6$	S^{2-} $[Ne]3s^23p^6$	Cl^- $[Ne]3s^23p^6$	Argon Ar $[Ne]3s^23p^6$
K^+ $[Ne]3s^23p^6$	Ca^{2+} $[Ne]3s^23p^6$	Sc^{3+} $[Ne]3s^23p^6$		As^{3-} $[Ar]3d^{10}4s^24p^6$	Se^{2-} $[Ar]3d^{10}4s^24p^6$	Br^- $[Ar]3d^{10}4s^24p^6$	Krypton Kr $[Ar]3d^{10}4s^24p^6$
Rb^+ $[Ar]3d^{10}4s^24p^6$	Sr^{2+} $[Ar]3d^{10}4s^24p^6$	Y^{3+} $[Ar]3d^{10}4s^24p^6$			Te^{2-} $[Kr]4d^{10}5s^25p^6$	I^- $[Kr]4d^{10}5s^25p^6$	Xenon Xe $[Kr]4d^{10}5s^25p^6$
Cs^+ $[Kr]4d^{10}5s^25p^6$	Ba^{2+} $[Kr]4d^{10}5s^25p^6$	La^{3+} $[Kr]4d^{10}5s^25p^6$					

Each color denotes ions and a noble gas that have the same electron configurations. The small table at right shows the periodic table positions of the ions listed above.



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Many Stable Ions Have Noble-Gas Configurations

Potassium and chlorine are not the only atoms that form stable ions with a complete octet of valence electrons. **Figure 5** lists examples of other atoms that form ions with a full octet. For example, examine how calcium, Ca, forms a stable ion. The electron configuration of a calcium atom is written as follows.

$$[\mathrm{Ca}] = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$$

By giving up its two valence electrons in the 4*s* orbital, calcium forms a stable cation with a 2+ charge that has an electron configuration like that of argon.

$$[\mathrm{Ca}^{2+}] = 1s^2 2s^2 2p^6 3s^2 3p^6$$

Topic Link

Refer to the "Atoms" chapter for more about electron configuration.

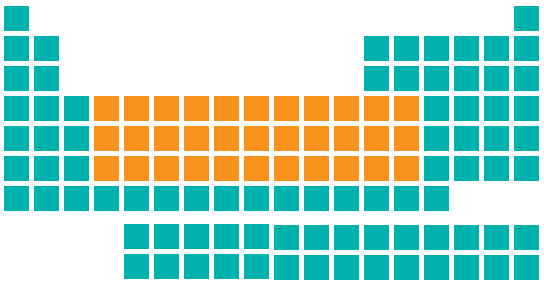
Some Stable Ions Do Not Have Noble-Gas Configurations

Not all atoms form stable ions with an electron configuration like those of noble gases. As illustrated in **Figure 6,** transition metals often form ions without complete octets. Notice that these stable ions are all cations. Also notice in **Figure 6** that some elements, mostly transition metals, can form several stable ions that have different charges. For example, copper, Cu, can give up one electron, forming a Cu^+ cation. It can also give up two electrons, forming a Cu^{2+} cation. Both the Cu^+ and Cu^{2+} cations are stable even though they do not have noble-gas configurations.

Figure 6
Some stable ions do not have electron configurations like those of the noble gases.

Stable Ions Formed by the Transition Elements and Some Other Metals

Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14
Ti^{2+} Ti^{3+}	V^{2+} V^{3+}	Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^{+} Cu^{2+}	Zn^{2+}	Ga^{2+} Ga^{3+}	Ge^{2+}
		Mo^{3+}	Tc^{2+}			Pd^{2+}	Ag^{+} Ag^{2+}	Cd^{2+}	In^{+} In^{2+} In^{3+}	Sn^{2+}
Hf^{4+}			Re^{4+} Re^{5+}			Pt^{2+} Pt^{4+}	Au^{+} Au^{3+}	Hg_2^{2+} Hg^{2+}	Tl^{+} Tl^{3+}	Pb^{2+}



The small table at left shows the periodic table positions of the ions listed above.

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Atoms and Ions

Many atoms form stable ions that have noble-gas configurations. It is important to remember that these elements do not actually *become* noble gases. Having identical electron configurations does not mean that a sodium cation is a neon atom. The sodium cation still has 11 protons and 12 neutrons, like a sodium atom that has not reacted to form an ion. But like a noble-gas atom, a sodium ion is very unlikely to gain or lose any more electrons.

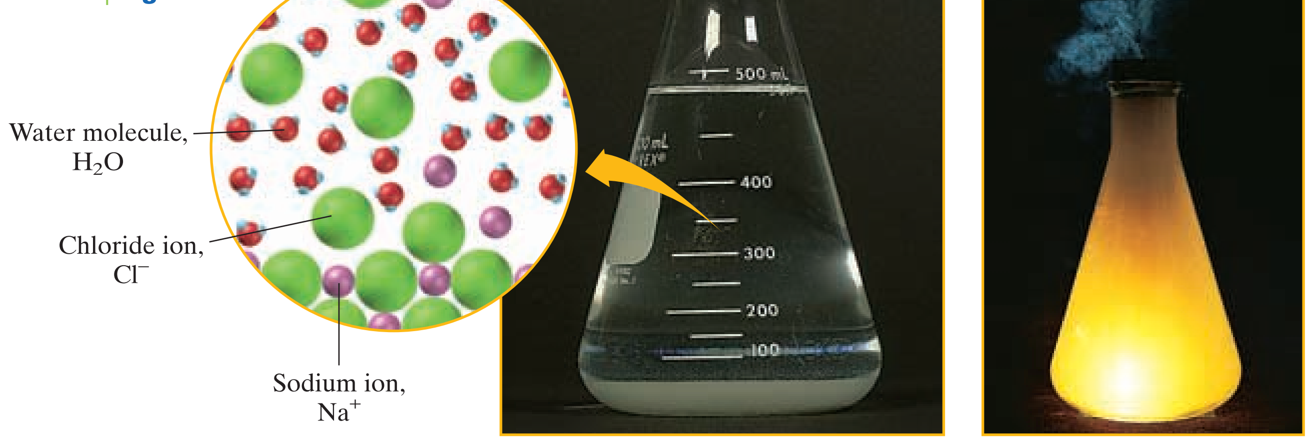
Ions and Their Parent Atoms Have Different Properties

Like potassium and all other alkali metals of Group 1, sodium is extremely reactive. When it is placed in water, a violent reaction occurs, producing heat and light. Like all halogens of Group 17, chlorine is extremely reactive. In fact, atoms of chlorine react with each other to form molecules of chlorine, Cl_2, a poisonous, yellowish green gas. As a pure element, chlorine is almost always found in nature as Cl_2 molecules rather than as individual Cl atoms.

Because both sodium and chlorine are very reactive, you might expect a violent reaction when these two are brought together. This is exactly what happens. If a small piece of sodium is lowered into a flask filled with chlorine gas, there is a violent reaction that releases both heat and light. After the reaction is complete, all that remains is a white solid. Even though it is formed from two dangerous elements, it is something you probably eat every day—table salt.

Chemists call this salt *sodium chloride.* Sodium chloride is made from sodium cations and chloride anions. As illustrated in **Figure 7,** these ions have very different properties than those of their parent atoms. That is why salt is not as dangerous to have around your house as the elements that make it up. It does not react with water like sodium metal does because salt contains stable sodium ions, not reactive sodium atoms.

Figure 7



a Sodium chloride dissolves in water to produce unreactive sodium cations and chlorine anions.

b In contrast, the elements sodium and chlorine are very reactive when they are brought together.

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Atoms of Metals and Nonmetal Elements Form Ions Differently

Nearly all metals form cations, as can be seen by examining their electron configuration. For example, consider the configuration for the Group 2 metal magnesium, Mg.

$$[\text{Mg}] = 1s^22s^22p^63s^2$$

To have a noble-gas configuration, the atom must either gain six electrons or lose two. Losing two electrons requires less energy than gaining six. Similarly, for all metals, the energy required to remove electrons from atoms to form ions with a noble-gas configuration is always less than the energy required to add more electrons. As a result, the atoms of metals form cations.

In contrast, the atoms of nonmetal elements form anions. Consider the example of oxygen, whose electron configuration is written as follows.

$$[\text{O}] = 1s^22s^22p^4$$

To have a noble-gas configuration, an oxygen atom must either gain two electrons or lose six. Acquiring two electrons requires less energy than losing six. For other nonmetals, the energy required to add electrons to atoms of nonmetals so that their ions have a noble-gas configuration is always less than the energy required to remove enough electrons. As a result, the atoms of nonmetal elements form anions.

1 Section Review

UNDERSTANDING KEY IDEAS

1. Explain why the noble gases tend not to react.

2. Where are the valence electrons located in an atom?

3. How does a cation differ from an anion?

4. State the octet rule.

5. Why do the properties of an ion differ from those of its parent atom?

6. Explain why alkali metals are extremely reactive.

7. How can you determine the number of valence electrons an atom has?

8. Explain why almost all metals tend to form cations.

9. Explain why, as a pure element, oxygen is usually found in nature as O_2.

CRITICAL THINKING

10. How could each of the following atoms react to achieve a noble-gas configuration?

a. iodine

b. strontium

c. nitrogen

d. krypton

11. Write the electron configuration for each of the following ions.

a. Al^{3+}

b. Se^{2-}

c. Sc^{3+}

d. As^{3-}

12. In what way is an ion the same as its parent atom?

13. To achieve a noble-gas configuration, a phosphorus atom will form a P^{3-} anion rather than forming a P^{5+} cation. Why?

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SECTION

2 Ionic Bonding and Salts

KEY TERMS

- salt
- lattice energy
- crystal lattice
- unit cell

OBJECTIVES

1. **Describe** the process of forming an ionic bond.
2. **Explain** how the properties of ionic compounds depend on the nature of ionic bonds.
3. **Describe** the structure of salt crystals.

Ionic Bonding

You may think that the material shown in **Figure 8** is very valuable. If you look closely, you will see what appear to be chunks of gold. The object shown in **Figure 8** is actually a mineral called *pyrite,* which does not contain any gold. However, the shiny yellow flakes make many people believe that they have discovered gold. All they have really discovered is a mineral that is made of iron cations and sulfur anions.

Because opposite charges attract, cations and anions should attract one another. This is exactly what happens when an ionic bond is formed. In the case of pyrite, the iron cations and sulfur anions attract one another to form an ionic compound.

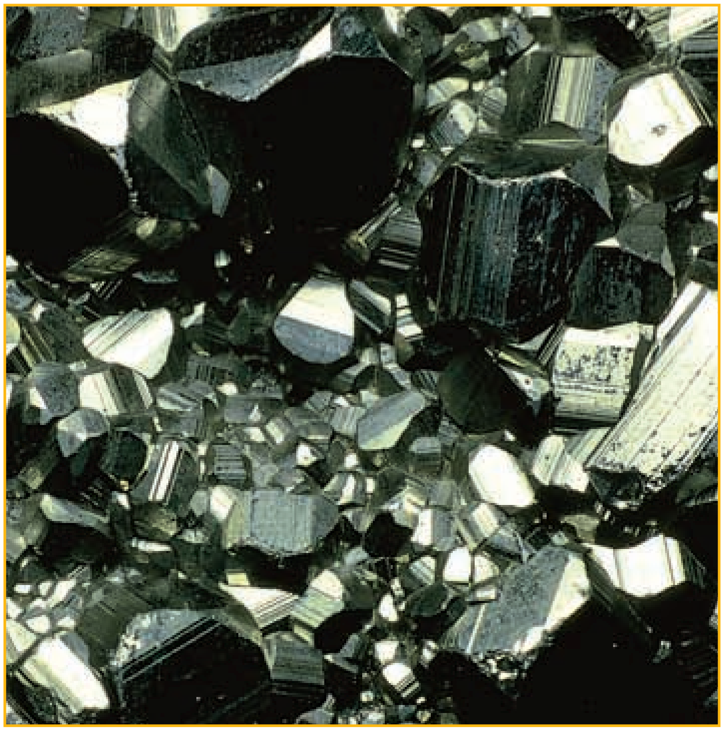

Figure 8
The mineral pyrite is commonly called *fool's gold*. Unlike real gold, pyrite is actually quite common in Earth's crust.

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Ionic Bonds Form Between Ions of Opposite Charge

To understand how an ionic bond forms, take another look at what happens when sodium and chlorine react to form sodium chloride. Recall that sodium gives up its only valence electron to form a stable Na^+ cation. Chlorine, with seven valence electrons, acquires that electron. As a result, a chlorine atom becomes a stable Cl^- anion.

The force of attraction between the 1+ charge on the sodium cation and the 1– charge on the chloride anion creates the ionic bond in sodium chloride. Recall that sodium chloride is the scientific name for table salt. Chemists call table salt by its scientific name because the word **salt** can actually be used to describe any one of thousands of different ionic compounds. Other salts that are commonly found in a laboratory include potassium chloride, magnesium oxide, and calcium iodide.

salt

an ionic compound that forms when a metal atom or a positive radical replaces the hydrogen of an acid

All these salts are ionic compounds that are electrically neutral. They are made up of cations and anions that are held together by ionic bonds in a simple, whole-number ratio. For example, sodium chloride consists of sodium cations and chloride anions bonded in a 1:1 ratio. To show this 1:1 ratio, chemists write the formula for sodium chloride as NaCl.

However, the attractions between the ions in a salt do not stop with a single cation and a single anion. These forces are so far reaching that one cation attracts several different anions. At the same time, each anion attracts several different cations. In this way, many ions are pulled together into a tightly packed structure. The tight packing of the ions causes any salt, such as sodium chloride, to have a distinctive crystal structure. The smallest crystal of table salt that you could see would still have more than a billion billion sodium and chloride ions.

Transferring Electrons Involves Energy Changes

Recall that ionization energy is the energy that it takes to remove the outermost electron from an atom. In other words, moving a negatively charged electron away from an atom that will become a positively charged ion requires an input of energy before it will take place. In the case of sodium, this process can be written as follows.

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$$Na + \text{energy} \longrightarrow Na^+ + e^-$$

Recall that electron affinity is the energy needed to add an electron onto a neutral atom. However, some elements, such as chlorine, easily accept extra electrons. For elements like this, energy is released when an electron is added. This process can be written as follows.

$$Cl + e^- \longrightarrow Cl^- + \text{energy}$$

But this energy released is less than the energy required to remove an electron from a sodium atom. Then why does an ionic bond form if these steps do not provide enough energy? Adding and removing electrons is only part of forming an ionic bond. The rest of the process of forming a salt supplies more than enough energy to make up the difference so that the overall process releases energy.

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Salt Formation Involves Endothermic Steps

The process of forming the salt sodium chloride can be broken down into five steps as shown in **Figure 9** on the following page. Keep in mind that these steps do not really take place in this order. However, these steps, do model what must happen for an ionic bond to form between sodium cations and chloride anions.

The starting materials are sodium metal and chlorine gas. Energy must be supplied to make the solid sodium metal into a gas. This process takes energy and can be written as follows.

$$\mathrm{Na(solid)} + \text{energy} \longrightarrow \mathrm{Na(gas)}$$

Recall that energy is also required to remove an electron from a gaseous sodium atom.

$$\mathrm{Na(gas)} + \text{energy} \longrightarrow \mathrm{Na^{+}(gas)} + e^{-}$$

No energy is required to convert chlorine into the gaseous state because it is already a gas. However, chlorine gas consists of two chlorine atoms that are bonded to one another. Therefore, energy must be supplied to separate these chlorine atoms so that they can react with sodium. This third process can be written as follows.

$$\mathrm{Cl{-}Cl(gas)} + \text{energy} \longrightarrow \mathrm{Cl(gas)} + \mathrm{Cl(gas)}$$

To this point, the first three steps have all been endothermic. These steps have produced sodium cations and chlorine atoms.

STUDY TIP

READING TABLES AND GRAPHS

Tables and graphs organize data into an easy-to-see form that is also easy to understand. This text is full of these tools to help you organize and clarify information.

When reading them, be sure to pay close attention to the headings and units of measurement. To get useful information from a table, you must understand how it is organized. Also look for trends or patterns in the table values or graph lines. You may want to design your own tables and graphs to help you understand and remember certain topics as you prepare for a chapter test.

Salt Formation Also Involves Exothermic Steps

As **Figure 9** illustrates, the next step adds an electron to a chlorine atom to form an anion. This is the first step that releases energy. Recall that this step cannot supply enough energy to remove an electron from a sodium atom. Obviously, this step cannot produce nearly enough energy to drive the first three steps.

The chief driving force for the formation of the salt is the last step, in which the separated ions come together to form a crystal held together by ionic bonds. When a cation and anion form an ionic bond, it is an exothermic process. Energy is released.

$$\mathrm{Na^{+}(gas)} + \mathrm{Cl^{-}(gas)} \longrightarrow \mathrm{NaCl(solid)} + \text{energy}$$

lattice energy

the energy associated with constructing a crystal lattice relative to the energy of all constituent atoms separated by infinite distances

The energy released when ionic bonds are formed is called the **lattice energy.** This energy is released when the crystal structure of a salt is formed as the separated ions bond. In the case of sodium chloride, the lattice energy is greater than the energy needed for the first three steps. Without this energy, there would not be enough energy to make the overall process spontaneous. Lattice energy is the key to salt formation.

The value of the lattice energy is different if other cations and anions form the salt. For example, Na^{+} ions can form salts with anions of any of the halogens. The lattice energy values for each of these salts are about the same. However, when magnesium cations, Mg^{2+}, form salts, these values

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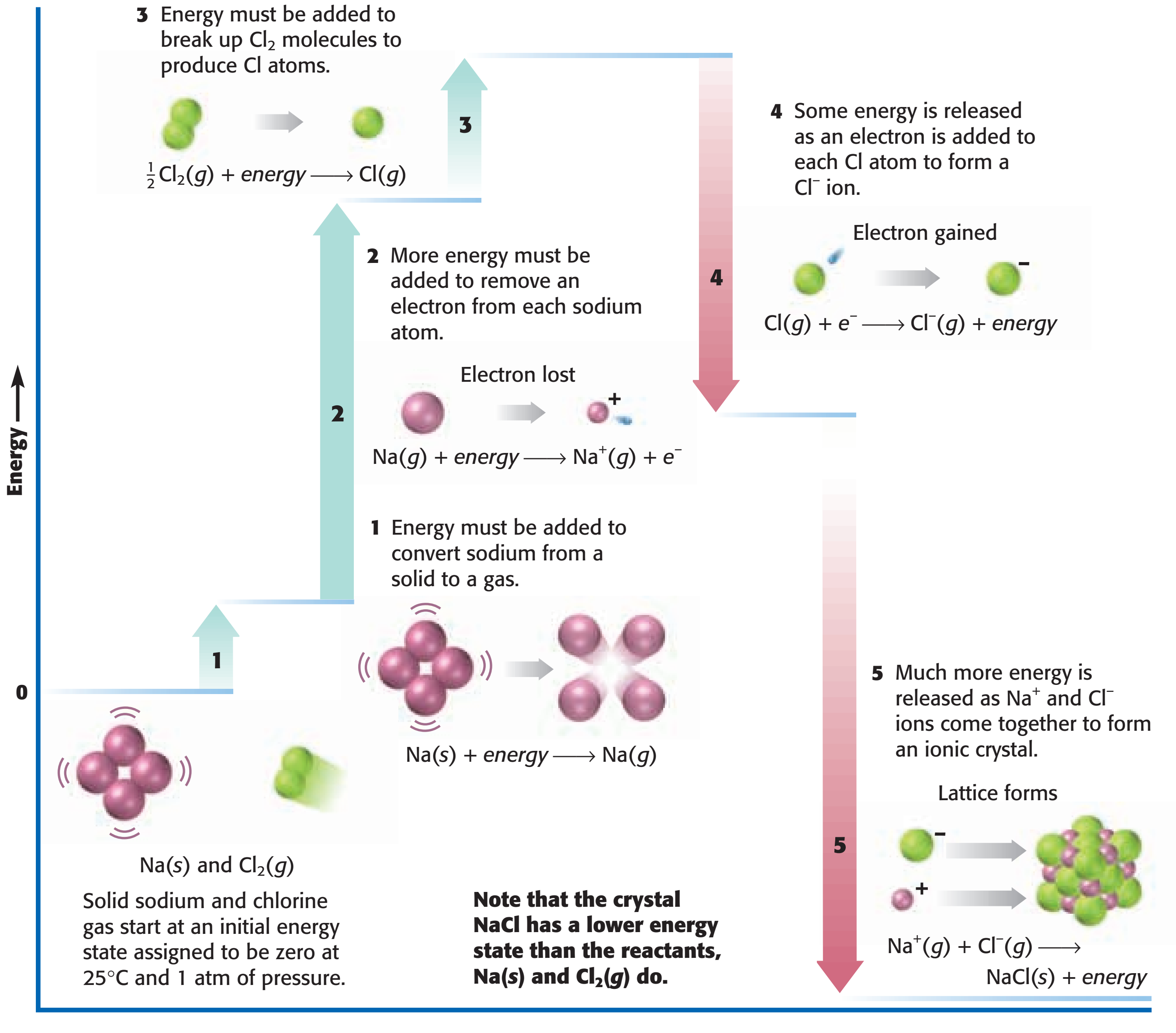
are much higher than the values for salts of sodium. This large difference in lattice energy is due to the fact that ions with greater charge are more strongly attracted to the oppositely charged ions in the crystal. The lattice energy value for magnesium oxide is almost five times greater than that for sodium chloride.

If energy is released when ionic bonds are formed, then energy must be supplied to break these bonds and separate the ions. In the case of sodium chloride, the needed energy can come from water. As a result, a sample of sodium chloride dissolves when it is added to a glass of water. As the salt dissolves, the Na^+ and Cl^- ions separate as the ionic bonds between them are broken. Because of its much higher lattice energy, magnesium oxide does not dissolve well in water. In this case, the energy that is available in a glass of water is significantly less than the lattice energy of the magnesium oxide. There is not enough energy to separate the Mg^{2+} and O^{2-} ions from one another.

Figure 9
The reaction between Na(*s*) and Cl_2(*g*) to form sodium chloride can be broken down into steps. More energy is released overall than is absorbed.



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Ionic Compounds

Recall that salts are ionic compounds made of cations and anions. Many of the rocks and minerals in Earth's crust are made of cations and anions held together by ionic bonds. The ratio of cations to anions is always such that an ionic compound has no overall charge. For example, in sodium chloride, for every Na^+ cation, there is a Cl^- anion to balance the charge. In magnesium oxide, for every Mg^{2+} cation, there is an O^{2-} anion. Ionic compounds also share certain other chemical and physical properties.

Ionic Compounds Do Not Consist of Molecules

Figure 10 shows sodium chloride, an ionic compound, being added to water, a molecular compound. If you could look closely enough into the water, you would find individual water molecules, each made of two hydrogen atoms and one oxygen atom. The pot would be filled with many billions of these individual H_2O molecules.

Recall that the smallest crystal of table salt that you could see contains many billions of sodium and chloride ions all held together by ionic bonds. However, if you could look closely enough into the salt, all you would see are many Na^+ and Cl^- ions all bonded together to form a crystal. There are no NaCl molecules.

Elements in Groups 1 and 2 reacting with elements in Groups 16 and 17 will almost always form ionic compounds and not molecular compounds. Therefore, the formula CaO likely indicates an ionic compound because Ca is a Group 2 metal and O is a Group 16 nonmetal. In contrast, the formula ICl likely indicates a molecular compound because both I and Cl are members of Group 17.

However, you cannot be absolutely sure that something is made of ions or molecules just by looking at its formula. That determination must be made in the laboratory.

Figure 10
Salt is often added to water for flavor when pasta is being cooked.

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Ionic Bonds Are Strong

Both repulsive and attractive forces exist within a salt crystal. The repulsive forces include those between like-charged ions. Within the crystal, each Na^+ ion repels the other Na^+ ions. The same is true for the Cl^- ions. Another repulsive force exists between the electrons of ions that are close together, even if the ions have opposite charges.

The attractive forces include those between the positively charged nuclei of one ion and the electrons of other nearby ions. In addition, attractive forces exist between oppositely charged ions. These forces involve more than a single Na^+ ion and a single Cl^- ion. Within the crystal, each sodium cation is surrounded by six chloride anions. At the same time, each chloride anion is surrounded by six sodium cations. As a result, the attractive force between oppositely charged ions is significantly greater in a crystal than it would be if the sodium cations and chloride anions existed only in pairs.

Overall, the attractive forces are significantly stronger than the repulsive forces, so ionic bonds are very strong.

Ionic Compounds Have Distinctive Properties

All ionic compounds share certain properties because of the strong attraction between their ions. Compare the boiling point of sodium chloride (1413°C) with that of water, a molecular compound (100°C). Similarly, most other ionic compounds have high melting and boiling points, as you can see in **Table 1.**

To melt, ions cannot be in fixed locations. Because each ion in these compounds forms strong bonds to neighboring ions, considerable energy is required to free them. Still more energy is needed to move ions out of the liquid state and cause boiling.

As a result of their high boiling points, ionic compounds are rarely gaseous at room temperature, while many molecular compounds are. Ice, for example, will eventually melt and then vaporize. In contrast, salt will remain a solid no matter how long it remains at room temperature.

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Topic: Ionic Compounds
SciLinks code: HW4072

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Table 1 Melting and Boiling Points of Compounds

Compound name	Formula	Type of compound	Melting point °C	Melting point K	Boiling point °C	Boiling point K
Magnesium fluoride	MgF_2	ionic	1261	1534	2239	2512
Sodium chloride	NaCl	ionic	801	1074	1413	1686
Calcium iodide	CaI_2	ionic	784	1057	1100	1373
Iodine monochloride	ICl	covalent	27	300	97	370
Carbon tetrachloride	CCl_4	covalent	–23	250	77	350
Hydrogen fluoride	HF	covalent	–83	190	20	293
Hydrogen sulfide	H_2S	covalent	–86	187	–61	212
Methane	CH_4	covalent	–182	91	–164	109

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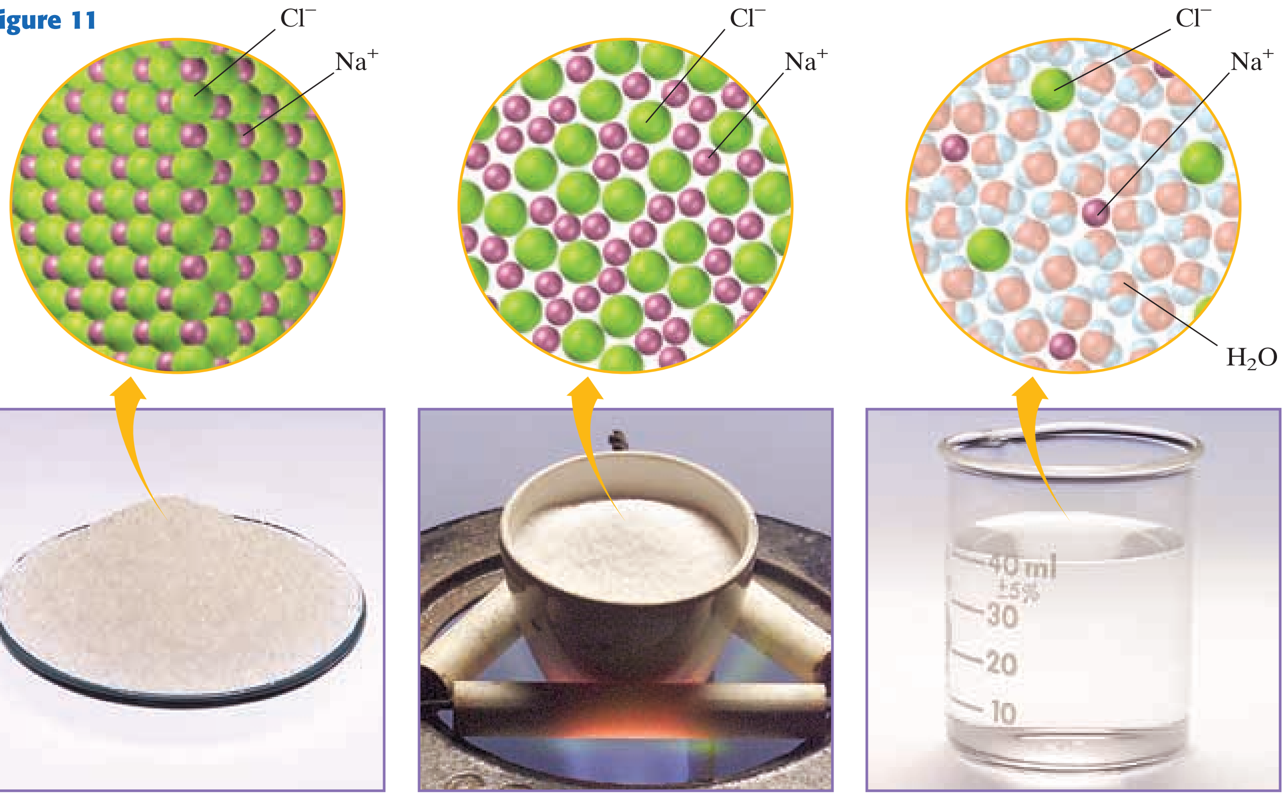
Liquid and Dissolved Salts Conduct Electric Current

To conduct an electric current, a substance must satisfy two conditions. First, the substance must contain charged particles. Second, those particles must be free to move. Because ionic compounds are composed of charged particles, you might expect that they could be good conductors. While particles in a solid have some vibrational motion, they remain in fixed locations, as shown by the model in **Figure 11a.** Therefore, ionic solids, such as salts, generally are not conductors of electric current because the ions cannot move.

However, when the ions can move about, salts are excellent electrical conductors. This is possible when a salt melts or dissolves. When a salt melts, the ions that make up the crystal can freely move past each other, as **Figure 11b** illustrates. Molten salts are good conductors of electric current, although they do not conduct as well as metals. Similarly, if a salt dissolves in water, its ions are no longer held tightly in a crystal. Because the ions are free to move, as shown by the model in **Figure 11c,** the solution can conduct electric current.

As often happens in chemistry, there are exceptions to this rule. There is a small class of ionic compounds that can allow charges to move through their crystals. The lattices of these compounds have an unusually open structure, so certain ions can move past others, jumping from one site to another. One of these salts, zirconium oxide, is used in a device that controls emissions from the exhaust of automobiles.

Figure 11



a As a solid, an ionic compound has charged particles that are held in fixed positions and cannot conduct electric current.

b When melted, an ionic compound conducts electric current because its charged particles move about more freely.

c When dissolved, an ionic compound conducts electric current because its charged particles move freely.

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Salts Are Hard and Brittle

Like most other ionic compounds, table salt is fairly hard and brittle. *Hard* means that the crystal is able to resist a large force applied to it. *Brittle* means that when the applied force becomes too strong to resist, the crystal develops a widespread fracture rather than a small dent. Both of these properties can be attributed to the patterns in which the cations and anions are arranged in all salt crystals.

The ions in a crystal are arranged in a repeating pattern, forming layers. Each layer is positioned so that a cation is next to an anion in the next layer. As long as the layers stay in a fixed position relative to one another, the attractive forces between oppositely charged ions will resist motion. As a result, the ionic compound will be hard, and it will take a lot of energy to break all the bonds between layers of ions.

However, if a force causes one layer to move slightly, ions with the same charge will be positioned next to each other. The cations in one layer are now lined up with other cations in a nearby layer. In the same way, anions from one layer are lined up with other anions in a nearby layer. Because the anions are next to each other, the like charges will repel each other and the layers will split apart. This is why all salts shatter along a line extending through the crystal known as a cleavage plane.

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SKILLS Toolkit 1

How to Identify a Compound as Ionic

You can carry out the following procedures in a laboratory to determine if a substance is an ionic compound.

- Examine the substance. All ionic compounds are solid at room temperature. If the substance is a liquid or gas, then it is not an ionic compound. However, if it is a solid, then it *may* or *may not* be an ionic compound.
- Tap the substance gently. Ionic compounds are hard and brittle. If it is an ionic compound, then it should not break apart easily. If it does break apart, the substance should fracture into tinier crystals and not crumble into a powder.
- Heat a sample of the substance. Ionic compounds generally have high melting and boiling points.
- If the substance melts, use a conductivity apparatus to determine if the melted substance conducts electric current. Ionic compounds are good conductors of electric current in the liquid state.
- Dissolve a sample of the substance in water. Use a conductivity apparatus to see if it conducts electric current. Ionic compounds conduct electric current when dissolved in water.

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Salt Crystals

The ions in a salt crystal form repeating patterns, with each ion held in place because there are more attractive forces than repulsive ones. The way the ions are arranged is the same in a number of different salts. Not all salts, however, have the same crystal structure as sodium chloride. Despite their differences, the crystals of all salts are made of simple repeating units. These repeating units are arranged in a salt to form a **crystal lattice.** These arrangements of repeating units within a salt are the reason for the crystal shape that can be seen in most salts.

crystal lattice

the regular pattern in which a crystal is arranged

Crystal Structure Depends on the Sizes and Ratios of Ions

As the formula for sodium chloride, NaCl, indicates, there is a 1:1 ratio of sodium cations and chlorine anions. Recall that the attractions in sodium chloride involve more than a single cation and a single anion. **Figure 12a** illustrates the crystal lattice structure of sodium chloride. Within the crystal, each Na^+ ion is surrounded by six Cl^- ions, and, in turn, each Cl^- ion is surrounded by six Na^+ ions. Because this arrangement does not hold for the edges of the crystal, the edges are locations of weak points.

The arrangement of cations and anions to form a crystal lattice depends on the size of the ions. Another factor that affects how the crystal forms is the ratio of cations to anions. Not all salts have a 1:1 ratio of cations to anions as found in sodium chloride. For example, the salt calcium fluoride has one Ca^{2+} ion for every two F^- ions. A Ca^{2+} ion is larger than an Na^+ ion, and an F^- ion is smaller than a Cl^- ion. Because of the size differences of its ions and their ratio in the salt, the crystal lattice structure of calcium fluoride is different from that of sodium chloride. As illustrated in **Figure 12b,** each calcium ion is surrounded by eight fluoride ions. At the same time, each fluoride ion is surrounded by four calcium ions. This is very different from the arrangement of six oppositely charged ions around any given positive or negative ion in a crystal of NaCl.

Figure 12 The crystal structure of sodium chloride (**a**) is not the same as that of calcium fluoride (**b**) because of the differences in the sizes of their ions and the cation-anion ratio making up each salt.

a **sodium chloride** **b** **calcium fluoride**

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Salts Have Ordered Packing Arrangements

Salts vary in the types of ions from which they are made. Salts also vary in the ratio of the ions that make up the crystal lattice. Despite these differences, all salts are made of simple repeating units. The smallest repeating unit in a crystal lattice is called a **unit cell.**

unit cell
the smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire lattice

The ways in which a salt's unit cells are arranged are determined by a technique called X-ray diffraction crystallography. First, a salt is bombarded with X rays. Then, the X rays that strike ions in the salt are deflected, while X rays that do not strike ions pass straight through the crystal lattice without stopping. The X rays form a pattern on exposed film. By analyzing this pattern, scientists can calculate the positions that the ions in the salt must have in order to cause the X rays to make such a pattern. After this work, scientists can then make models to show how the ions are arranged in the unit cells of the salt.

Analysis of many different salts show that the salts all have ordered packing arrangements, such as those described earlier for NaCl and CaF_2. Another example is the salt cesium chloride, where the ratio of cations to anions is 1:1 just as it is in sodium chloride. However, the size of a cesium cation is larger than that of a sodium cation. As a result, the structure of the crystal lattice is different. In sodium chloride, a sodium cation is surrounded by six chloride anions. In cesium chloride, a cesium cation is surrounded by eight chloride anions. The bigger cation has more room around it, so more anions can cluster around it.

2 Section Review

UNDERSTANDING KEY IDEAS

1. What force holds together the ions in a salt?
2. Describe how an ionic bond forms.
3. Why are ionic solids hard?
4. Why are ionic solids brittle?
5. Explain why lattice energy is the key to the formation of a salt.
6. Why do ionic crystals conduct electric current in the liquid phase or when dissolved in water but do not conduct electric current in the solid phase?

CRITICAL THINKING

7. Crystals of the ionic compound calcium fluoride have a different structure from that of the ionic compound calcium chloride. Suggest a reason for this difference.
8. Explain why each of the following pairs is not likely to form an ionic bond.
   a. chlorine and bromine
   b. potassium and helium
   c. sodium and lithium
9. The lattice energy for sodium iodide is 700 kJ/mol, while that for calcium sulfide is 2775 kJ/mol. Which of these salts do you predict has the higher melting point? Explain.
10. The electron affinity for chlorine has a negative value, indicating that the atom readily accepts another electron. Why does a chlorine atom readily accept another electron?
11. Use **Figure 9** on page 169 to describe how the formation of calcium chloride would differ from that of sodium chloride. (Hint: Compare the electron configurations of each atom.)

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SECTION 3

Names and Formulas of Ionic Compounds

KEY TERMS

- **polyatomic ion**

OBJECTIVES

1. **Name** cations, anions, and ionic compounds.
2. **Write** chemical formulas for ionic compounds such that an overall neutral charge is maintained.
3. **Explain** how polyatomic ions and their salts are named and how their formulas relate to their names.

Naming Ionic Compounds

You may recall that chemists call table salt *sodium chloride.* In fact, they have a name for every salt. With thousands of different salts, you might think that it would be hard to remember the names of all of them. But naming salts is very easy, especially for those that are made of a simple cation and a simple anion. These kinds of salts are known as binary ionic compounds. The adjective *binary* indicates that the compound is made up of just two elements.

Rules for Naming Simple Ions

Simple cations borrow their names from the names of the elements. For example, K^+ is known as the potassium ion, and Zn^{2+} is known as the zinc ion. When an element forms two or more ions, the ion names include roman numerals to indicate charge. In the case of copper, Cu, the names of the two ions are written as follows.

Cu^+ copper(I) ion $\quad$ Cu^{2+} copper(II) ion

When we read the names of these ions out loud, we say "copper one ion" or "copper two ion."

The name of a simple anion is also formed from the name of the element, but it ends in *-ide*. Thus, Cl^- is the chloride ion, O^{2-} is the oxide ion, and P^{3-} is the phosphide ion.

The Names of Ions Are Used to Name an Ionic Compound

Naming binary ionic compounds is simple. The name is made up of just two words: the name of the cation followed by the name of the anion.

NaCl sodium chloride	$CuCl_2$ copper(II) chloride
ZnS zinc sulfide	Mg_3N_2 magnesium nitride
K_2O potassium oxide	Al_2S_3 aluminum sulfide

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Writing Ionic Formulas

Ionic compounds never have an excess of positive or negative charges. To maintain this balance the total positive and negative charges must be the same. Because both ions in sodium chloride carry a single charge, this compound is made up of equal numbers of the ions Na^+ and Cl^-. As you have seen, the formula for sodium chloride is written as NaCl to show this one-to-one ratio. The cation in zinc sulfide has a 2+ charge and the anion has a 2– charge. Again there is a one-to-one ratio in the salt. Zinc sulfide has the formula ZnS.

Compounds Must Have No Overall Charge

You must take care when writing the formula for an ionic compound where the charges of the cation and anion differ. Consider the example of magnesium nitride. The magnesium ion, Mg^{2+}, has two positive charges, and the nitride ion, N^{3-}, has three negative charges. The cations and anions must be combined in such a way that there are the same number of negative charges and positive charges. Three Mg^{2+} cations are needed for every two N^{3-} anions for electroneutrality. That way, there are six positive charges and six negative charges. Subscripts are used to denote the three magnesium ions and two nitride ions. Therefore, the formula for magnesium nitride is Mg_3N_2.

SKILLS Toolkit 2

Writing the Formula of an Ionic Compound

Follow the following steps when writing the formula of a binary ionic compound, such as iron(III) oxide.

- Write the symbol and charges for the cation and anion. Refer to **Figures 5** and **6** earlier in the chapter for the charges on the ions. The roman numeral indicates which cation iron forms.

  symbol for iron(III): Fe^{3+} symbol for oxide: O^{2-}

- Write the symbols for the ions side by side, beginning with the cation.

$$Fe^{3+}O^{2-}$$

- To determine how to get a neutral compound, look for the lowest common multiple of the charges on the ions. The lowest common multiple of 3 and 2 is 6. Therefore, the formula should indicate six positive charges and six negative charges.

For six positive charges, you need two Fe^{3+} ions because $2 \times 3+ = 6+$.

For six negative charges, you need three O^{2-} ions because $3 \times 2- = 6-$.

Therefore the ratio of Fe^{3+} to O^{2-} is 2Fe:3O. The formula is written as follows.

$$Fe_2O_3$$

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Polyatomic Ions

Fertilizers have potassium, nitrogen, and phosphorus in a form that dissolves easily in water so that plants can absorb them. The potassium in fertilizer is in an ionic compound called *potassium carbonate.* Two ionic compounds in the fertilizer contain the nitrogen—ammonium nitrate and ammonium sulfate. The phosphorus supplied is in another ionic compound, calcium dihydrogen phosphate.

These compounds in fertilizer are made of cations and anions in a ratio so there is no overall charge, like all other ionic compounds. But instead of ions made of a single atom, these compounds contain groups of atoms that are ions.

Many Atoms Can Form One Ion

The adjective *simple* describes an ion formed from a single atom. A simple ion could also be called *monatomic,* which means "one-atom." Just as the prefix *mon-* means "one," the prefix *poly-* means "many." The term **polyatomic ion** means a charged group of two or more bonded atoms that can be considered a single ion. A polyatomic ion as a whole forms ionic bonds in the same way that simple ions do.

polyatomic ion
an ion made of two or more atoms

Unlike simple ions, most polyatomic ions are made of atoms of several elements. However, polyatomic ions are like simple ones in that their charge is either positive or negative. Consider the polyatomic ion ammonium, NH_4^+, found in many fertilizers. Ammonium is made of one nitrogen and four hydrogen atoms. These atoms have a combined total of 11 protons but only 10 electrons. So the ammonium ion has a 1+ charge overall. This charge is not found on any single atom. Instead, it is spread across this group of atoms, which are bonded together.

The Names of Polyatomic Ions Can Be Complicated

Naming polyatomic ions is not as easy as naming simple cations and anions. Even so, there are rules you can follow to help you remember how to name some of them.

Many polyatomic ions contain oxygen. The endings *-ite* and *-ate* indicate the presence of oxygen. Examples include sulfite, nitrate, and acetate. Often there are several polyatomic ions that differ only in the number of oxygen atoms present. For example, the formulas for two polyatomic ions made from sulfur and oxygen are SO_3^{2-} and SO_4^{2-}. In such cases, the one with less oxygen takes the *-ite* ending, so SO_3^{2-} is named *sulfite*. The ion with more oxygen takes the *-ate* ending, so SO_4^{2-} is named *sulfate*. For the same reason, NO_2^- is named *nitrite,* and NO_3^- is named *nitrate*.

The presence of hydrogen is often indicated by an ion's name starting with *hydrogen*. The prefixes *mono-* and *di-* are also used. Thus, HPO_4^{2-} and $H_2PO_4^-$ are monohydrogen phosphate and dihydrogen phosphate ions, respectively. The prefix *thio-* means "replace an oxygen with a sulfur" in the formula, as in potassium thiosulfate, $K_2S_2O_3$, compared with potassium sulfate, K_2SO_4. **Table 2** lists the names and formulas for some common polyatomic ions. Notice that some are made of more than one atom of the same element, such as peroxide, O_2^{2-}.

Table 2
Some Polyatomic Ions

Ion name	Formula
Acetate	CH_3COO^-
Ammonium	NH_4^+
Carbonate	CO_3^{2-}
Chromate	CrO_4^{2-}
Cyanide	CN^-
Dichromate	$Cr_2O_7^{2-}$
Hydroxide	OH^-
Nitrate	NO_3^-
Nitrite	NO_2^-
Permanganate	MnO_4^-
Peroxide	O_2^{2-}
Phosphate	PO_4^{3-}
Sulfate	SO_4^{2-}
Sulfite	SO_3^{2-}
Thiosulfate	$S_2O_3^{2-}$

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SKILLS Toolkit 3

Naming Compounds with Polyatomic Ions

Follow these steps when naming an ionic compound that contains one or more polyatomic ions, such as K_2CO_3.

- Name the cation. Recall that a cation is simply the name of the element. In this formula, K is potassium that forms a singly charged cation, K^+, of the same name.
- Name the anion. Recall that salts are electrically neutral. Because there are two K^+ cations present in this salt, these two positive charges must be balanced by two negative charges. Therefore, the polyatomic anion in this salt must be CO_3^{2-}. You may find it helpful to think of the formula as follows, although it is not written this way.

$$(K^+)_2(CO_3^{2-})$$

If you check **Table 2,** you will see that the CO_3^{2-} polyatomic ion is called carbonate.

- Name the salt. Recall that the name of a salt is just the names of the cation and anion. The salt K_2CO_3 is potassium carbonate.

SAMPLE PROBLEM A

Formula of a Compound with a Polyatomic Ion

What is the formula for iron(III) chromate?

1 Gather information.

- Use **Figure 6,** found earlier in this chapter, to determine the formula and charge for the iron(III) cation.

$$Fe^{3+}$$

- Use **Table 2,** found earlier in this chapter, to determine the formula and charge for the chromate polyatomic ion.

$$CrO_4^{2-}$$

2 Plan your work.

- Because all ionic compounds are electrically neutral, the total charges of the cations and anions must be equal. To balance the charges, find the least common multiple of the ions' charges. The least common multiple of 2 and 3 is 6. To get six positive charges, you need two Fe^{3+} ions.

$$2 \times 3 = 6+$$

To get six negative charges, you need three CrO_4^{2-} ions.

$$3 \times 2 = 6-$$

continued on next page

PRACTICE HINT

Sometimes parentheses must be placed around the polyatomic cation, as in the formula $(NH_4)_2CrO_4$.

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3 Calculate.

- The formula must indicate that two Fe^{3+} ions and three CrO_4^{2-} ions are present. Parentheses are used whenever a polyatomic ion is present more than once. The formula for iron(III) chromate is written as follows.

$$Fe_2(CrO_4)_3$$

Notice that the parentheses show that everything inside the parentheses is tripled by the subscript *3* outside.

4 Verify your result.

- The formula includes the correct symbols for the cation and polyatomic anion.
- The formula reflects that the salt is electrically neutral.

PRACTICE

PROBLEM SOLVING SKILL

1 Write the formulas for the following ionic compounds.

a. calcium cyanide
b. rubidium thiosulfate
c. calcium acetate
d. ammonium sulfate

3 Section Review

UNDERSTANDING KEY IDEAS

1. In what ways are polyatomic ions like simple ions? In what ways are they different?

2. Why must roman numerals be used when naming certain ionic compounds?

3. What do the endings *-ite* and *-ate* indicate about a polyatomic ion?

4. Explain how calcium, Ca^{2+}, and phosphate, PO_4^{3-}, can make a compound with electroneutrality.

CRITICAL THINKING

5. Name the compounds represented by the following formulas.

a. $Ca(NO_2)_2$
b. $Fe(OH)_3$
c. $(NH_4)_2Cr_2O_7$
d. $CuCH_3COO$

6. Write the formulas for the following ionic compounds made of simple ions.

a. sodium oxide
b. magnesium phosphide
c. silver(I) sulfide
d. niobium(V) chloride

7. Name the following binary ionic compounds. If the metal forms more than one cation, be sure to denote the charge.

a. Rb_2O **b.** FeF_2 **c.** K_3N

PRACTICE PROBLEMS

8. Write formulas for the following compounds.

a. mercury(II) sulfate
b. lithium thiosulfate
c. ammonium phosphate
d. potassium permanganate

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11
Na
Sodium
22.989 770
[Ne]$3s^1$

Element Spotlight

Where Is Na?

Earth's Crust
2.36% by mass
Seventh most abundant element
Fifth most abundant metal

Sea Water
30.61% of all dissolved materials
1.03% by mass, taking the water into account

A Major Nutritional Mineral

Sodium is important in the regulation of fluid balance within the body. Most sodium in the diet comes from the use of table salt, NaCl, to season and preserve foods. Sodium is also supplied by compounds such as sodium carbonate and sodium hydrogen carbonate in baked goods. Sodium benzoate is a preservative in carbonated beverages. Sodium citrate and sodium glutamate are used in packaged foods as flavor additives.

In ancient Rome, salt was so scarce and highly prized that it was used as a form of payment. Today, however, salt is plentiful in the diet. Many people must limit their intake of sodium as a precaution against high blood pressure, heart attacks, and strokes.

Food labels list the amount of sodium contained in each serving.

Industrial Uses

- Common table salt is the most important commercial sodium compound. It is used in ceramic glazes, metallurgy, soap manufacture, home water softeners, highway de-icing, herbicides, fire extinguishers, and resins.
- The United States produces about 42.1 million metric tons of sodium chloride per year.
- Elemental sodium is used in sodium vapor lamps for lighting highways, stadiums, and other buildings.
- Liquid sodium is used to cool liquid-metal fast-breeder nuclear reactors.

Real-World Connection For most people, the daily intake of sodium should not exceed 2400 mg.

A Brief History

1200 — 1800 — 1900 — 2000

1251: The Wieliczka Salt Mine, located in Krakow, Poland, is started. The mine is still in use today.

1807: Sir Humphry Davy isolates sodium by the electrolysis of caustic soda (NaOH) and names the metal.

1930: Sodium vapor lamps are first used for street lighting.

1940: The Food and Nutrition Board of the National Research Council develops the first Recommended Dietary Allowances.

1990: The Nutrition Labeling and Education Act defines a Daily Reference Value for sodium to be listed in the Nutrition Facts portion of a food label.

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Questions

1. What is one possible consequence of too much sodium in the diet?

2. What is the chemical formula of the most important commercial sodium compound?

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5 Chapter Highlights

KEY IDEAS

SECTION ONE Simple Ions

- Atoms may gain or lose electrons to achieve an electron configuration identical to that of a noble gas.
- Alkali metals and halogens are very reactive when donating and accepting electrons from one another.
- Electrons in the outermost energy level are known as valence electrons.
- Ions are electrically charged particles that have different chemical properties than their parent atoms.

KEY TERMS

octet rule
ion
cation
anion

SECTION TWO Ionic Bonding and Salts

- The opposite charges of cations and anions attract to form a tightly packed substance of bonded ions called a *crystal lattice.*
- Salts have high melting and boiling points and do not conduct electric current in the solid state, but they do conduct electric current when melted or when dissolved in water.
- Salts are made of unit cells that have an ordered packing arrangement.

KEY TERMS

salt
lattice energy
crystal lattice
unit cell

SECTION THREE Names and Formulas of Ionic Compounds

- Ionic compounds are named by joining the cation and anion names.
- Formulas for ionic compounds are written to show their balance of overall charge.
- A polyatomic ion is a group of two or more atoms bonded together that functions as a single unit.
- Parentheses are used to group polyatomic ions in a chemical formula with a subscript.

KEY TERMS

polyatomic ion

KEY SKILLS

How to Identify a Compound as Ionic
Skills Toolkit 1 p. 173

Writing the Formula of an Ionic Compound
Skills Toolkit 2 p. 177

Naming Compounds with Polyatomic Ions
Skills Toolkit 3 p. 179

Formula of a Compound with a Polyatomic Ion
Sample Problem A p. 179

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Chapter Review 5

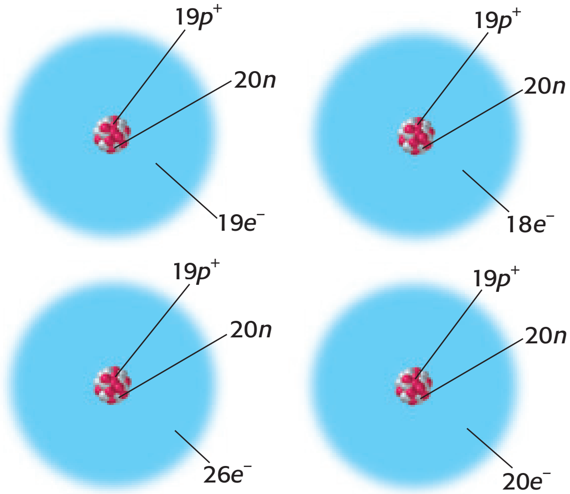
USING KEY TERMS

1. How is an ion different from its parent atom?
2. What does a metal atom need to do in order to form a cation?
3. What does a nonmetal element need to do to form an anion?
4. Explain how the octet rule describes how atoms form stable ions.
5. Why is lattice energy the key to forming an ionic bond?
6. Explain why it is appropriate to group a polyatomic ion in parentheses in a chemical formula, if more than one of that ion is present in the formula.

UNDERSTANDING KEY IDEAS

Simple Ions

7. The electron configuration for arsenic, As, is [Ar]$3d^{10}4s^24p^3$. How many valence electrons does an As atom have? Write the symbol for the ion it forms to achieve a noble-gas configuration.
8. Explain why the properties of an ion differ from its parent atom.
9. How does the octet rule help predict the chemical reactivity of an element?
10. Why are the halogens so reactive?
11. If helium does not obey the octet rule, then why do its atoms not react?
12. Explain why metals tend to form cations, while nonmetals tend to form anions.
13. Which of the following diagrams illustrates the electron diagram for a potassium ion found in the nerve cells of your body? (Hint: potassium's atomic number is 19.)



Ionic Bonding and Salts

14. Why do most ionic compounds have such high melting and boiling points?
15. Explain the importance of lattice energy in the formation of a salt.
16. Why can't an ionic bond form between potassium and magnesium?

Names and Formulas of Ionic Compounds

17. What is the difference between the chlorite ion and the chlorate ion?
18. Identify and name the cations and anions that make up the following ionic compounds and indicate the charge on each ion.
    a. $NaNO_3$
    b. K_2SO_3
    c. $(NH_4)_2CrO_4$
    d. $Al_2(SO_4)_3$
19. Name the compounds represented by the following formulas.
    a. $Cu_3(PO_4)_2$
    b. $Fe(NO_3)_3$
    c. Cu_2O
    d. CuO

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PRACTICE PROBLEMS

PROBLEM SOLVING SKILL

20. Write formulas for the following ionic compounds.
a. lithium sulfate
b. strontium nitrate
c. ammonium acetate
d. titanium(III) sulfate

21. Complete the table below, and then use it to answer the questions that follow.

Element	Ion	Name of ion
Barium	Ba^{2+}	
Chlorine		chloride
Chromium	Cr^{3+}	
Fluorine	F^{-}	
Manganese		manganese(II)
Oxygen		oxide

Write the formula for the following substances:
a. manganese chloride
b. chromium(III) fluoride
c. barium oxide

MIXED REVIEW

22. Name the following polyatomic ions.
a. O_2^{2-}
b. CrO_4^{2-}
c. NH_4^{+}
d. CO_3^{2-}

23. Complete the table below.

Atom	Ion	Noble-gas configuration of ion
S		
Be		
I		
Rb		
O		
Sr		
F		

24. Write formulas for the following polyatomic ions.
a. cyanide
b. sulfate
c. nitrite
d. permanganate

25. Determine the number of valence electrons in the following atoms.
a. Al
b. Rb
c. Si
d. F

CRITICAL THINKING

26. Why are most metals found in nature as ores and not as pure metals?

27. Why can't sodium gain a positive charge by acquiring a proton in its nucleus?

28. Why are there no rules for naming Group 18 ions?

29. Compound B has lower melting and boiling points than compound A does. At the same temperature, compound B vaporizes faster and to a greater extent than compound A. If only one of these compounds is ionic, which one would you expect it to be? Why?

ALTERNATIVE ASSESSMENT

WRITING SKILLS

30. A number of homes have "hard water," which, as you learned in the Start-Up Activity, does not produce as many soap suds as water that contains fewer ions. Such homes often have water conditioners that remove the ions from the water, making it "softer" and more likely to produce soapsuds. Research how such water softeners operate by checking the Internet or by contacting a company that sells such devices. Design an experiment to test the effectiveness of the softener in removing ions from water.

CONCEPT MAPPING

concept map

31. Use the following terms to create a concept map: *atoms, valence electrons, ions, cations, anions,* and *ionic compounds.*

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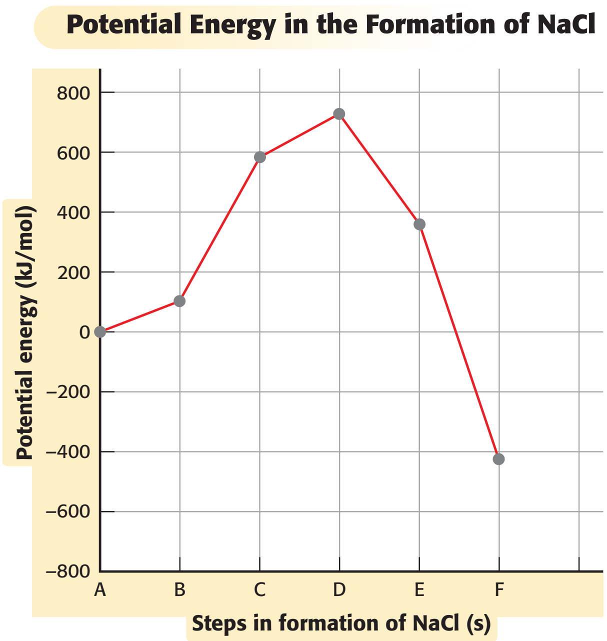
FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."



The graph shows the changes in potential energy that occur when an ionic bond forms between Na(*s*) and Cl_2(*g*). The reactants, solid sodium and chlorine gas, start at an initial energy state that is assigned a value of zero at 25°C and 1 atm of pressure.

32. In terms of energy, what do the steps from point A to point D have in common?

33. What do the steps from point D to point F have in common?

34. What is occurring between points D and E?

35. Write the word equation to show what happens between points B and C when electrons are removed from 1 mol of sodium atoms.

36. Which portion of this graph represents the lattice energy involved in the formation of an ionic bond between sodium and chlorine?

37. Calculate the quantity of energy released when 2.5 mol of NaCl form.

TECHNOLOGY AND LEARNING

38. Graphing Calculator

Calculating the Number of Valence Electrons

The graphing calculator can run a program that can determine the number of valence electrons in an atom, given its atomic number.

Go to Appendix C. If you are using a TI-83 Plus, you can download the program VALENCE and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes to use. After you have run the program, answer these questions.

How many valence electrons are there in the following atoms?

a. Rutherfordium, Rf, atomic number 104

b. Gold, Au, atomic number 79

c. Molybdenum, Mo, atomic number 42

d. Indium, In, atomic number 49

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5 STANDARDIZED TEST PREP

UNDERSTANDING CONCEPTS

***Directions (1–4):* For *each* question, write on a separate sheet of paper the letter of the correct answer.**

1. Which of the following can achieve the same electron configuration as a noble gas when the atom forms an ion?
   A. argon
   B. iron
   C. nickel
   D. potassium

2. Why is an input of energy needed when forming NaCl?
   F. to change chlorine to a gas
   G. to add an electron to the chlorine atom
   H. to remove an electron from the sodium atom
   I. to bring together the sodium and the chloride ions

3. Which of the following is a characteristic of a salt?
   A. bends but does not shatter when struck sharply
   B. has the ability to conduct electric current in the solid state
   C. has the ability to conduct electric current in the liquid state
   D. melts at temperatures that are slightly higher than room temperature

4. Which of the following pairs of elements are **most** likely to form an ionic bond?
   F. Br and Ca
   G. Br and N
   H. Ca and Mg
   I. Ca and Fe

***Directions (5–6):* For *each* question, write a short response.**

5. Explain why only a few metals are found in nature in their pure form, while most exist only as ores, which are metal-containing compounds.

6. How can you tell from the number of valence electrons whether an element is more likely to form a cation or an anion?

READING SKILLS

***Directions (7–8):* Read the passage below. Then answer the questions.**

In 1980 an oil drilling rig in Lake Peignur in Louisiana opened a hole from the lake to a salt mine 1,300 feet below ground. As the lake water flowed into the mine, it dissolved the salt pillars that were left behind to hold up the ceiling. When the entire mine collapsed, the resulting whirlpool swallowed a number of barges, a tugboat, trucks, and a large portion of an island in the middle of the lake. Eventually, the hole filled with water from a canal, leaving a much deeper lake.

7. What was the most likely cause of the collapse of the salt mine?
   A. The salt melted due to the temperature of the water.
   B. Water dissolved the ionic sodium chloride, leaving no supports.
   C. Water is denser than salt, so the salt began to float, moving the columns.
   D. The open hole exposed the salt pillars to the air and they had a chemical reaction with oxygen.

8. When there is no water present, the pillars in a salt mine are capable of holding the weight of the ceiling because
   F. Salt is held together by strong ionic bonds.
   G. Salt melts as it is mined and then reforms to a hard crystal.
   H. Salt contains sodium, which gives it the properties of metal.
   I. Salt does not crumble due to the low temperatures found below ground level.

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INTERPRETING GRAPHICS

Directions (9-12): **For *each* question below, record the correct answer on a separate sheet of paper.**

Many transition metals are capable of forming more than one type of stable ion. The properties of compounds formed by one ion are often very different from those formed by an ion of the same element having a different charge. Use the table below to answer questions 9 through 12.

Stable Ions Formed by the Transition Elements and Some Other Metals

Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14
Ti^{2+} Ti^{3+}	V^{2+} V^{3+}	Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^{+} Cu^{2+}	Zn^{2+}	Ga^{2+} Ga^{3+}	Ge^{2+}
		Mo^{3+}	Tc^{2+}			Pd^{2+}	Ag^{+} Ag^{2+}	Cd^{2+}	In^{+} In^{2+} In^{3+}	Sn^{2+}
Hf^{4+}			Re^{4+} Re^{5+}			Pt^{2+} Pt^{4+}	Au^{+} Au^{3+}	Hg_2^{2+} Hg^{2+}	Tl^{+} Tl^{3+}	Pb^{2+}

9 How do the cations formed by transition metals differ from those formed by metals in the first two columns of the periodic table?
A. Transition metals lose more electrons.
B. All of the transition metal ions have a positive charge.
C. Transition metals generally do not ionize to a noble gas configuration.
D. All of the transition metals are capable of forming several different ions.

10 Which of these metals forms ions with a noble gas electron configuration?
F. copper
G. germanium
H. hafnium
I. platinum

11 Based on the stable ions in the illustration, which of these compounds is most likely to exist?
A. Fe_2O
B. FeO_2
C. Hg_2O
D. Mo_3O_2

12 How many different ionic compounds exist that consist of only iron and chlorine?

Test TIP

When possible, use the text in the test to answer other questions. For example, use a multiple-choice answer to "jump start" your thinking about another question.

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CHAPTER

6

COVALENT COMPOUNDS

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