



Ionic Compounds

What You'll Learn

- ▶ You will define a chemical bond.
- ▶ You will describe how ions form.
- ▶ You will identify ionic bonding and the characteristics of ionic compounds.
- ▶ You will name and write formulas for ionic compounds.
- ▶ You will relate metallic bonds to the characteristics of metals.

Why It's Important

The world around you is composed mainly of compounds. The properties of each compound are based on how the compound is bonded. The salts dissolved in Earth's oceans and the compounds that make up most of Earth's crust are held together by ionic bonds.



Visit the Chemistry Web site at chemistrymc.com to find links about ionic compounds.

The rock surface, the climbers' equipment, the atmosphere, and even the climbers are composed almost entirely of compounds and mixtures of compounds.





DISCOVERY LAB



An Unusual Alloy

Most metals that you encounter are solids. Can a metal melt at a temperature below the boiling point of water? You will use a metal alloy called Onion's Fusible Alloy to answer this question.

Safety Precautions



Use caution around the heat source and the heated beaker and its contents.

Procedure

1. Carefully place a small piece of Onion's Fusible Alloy into a 250-mL beaker. Add about 100 mL of water to the beaker.
2. Heat the beaker and its contents with a laboratory burner or a hot plate. Monitor the temperature with a thermometer. When the temperature rises above 85°C, carefully observe the Onion's Fusible Alloy. Record your observations. Remove the beaker from the heat when the water begins to boil. Allow the contents to cool before handling the Onion's Fusible Alloy.

Materials

Onion's Fusible Alloy
250-mL beaker
laboratory burner or
hot plate
Celsius thermometer

Analysis

What is unusual about Onion's Fusible Alloy compared to other metals? Onion's Fusible Alloy contains bismuth, lead, and tin. Compare the melting points of these metals to the melting point of Onion's Fusible Alloy.

Section

8.1

Forming Chemical Bonds

Objectives

- **Define** chemical bond.
- **Relate** chemical bond formation to electron configuration.
- **Describe** the formation of positive and negative ions.

Vocabulary

chemical bond
cation
anion

Ascending to the summit of a mountain peak, a rock climber can survey the surrounding world. This world is composed of many different kinds of compounds, ranging from simple ones such as the sodium chloride found in the perspiration on the climber's skin to more complex ones such as the calcite or pyrite found in certain rocks. How do these and thousands of other compounds form from the relatively few elements known to exist?

Chemical Bonds

The answer to this question lies in the electron structure of the atoms of the elements involved and the nature of the attractive forces between these atoms. The force that holds two atoms together is called a **chemical bond**. Chemical bonds may form by the attraction between a positive nucleus and negative electrons or the attraction between a positive ion and a negative ion.

In previous chapters, you learned about atomic structure, electron arrangement, and periodic properties of the elements. The elements within a group on the periodic table have similar properties. Many of these properties are due to the number of valence electrons. These same electrons are involved in the formation of chemical bonds between two atoms.

**Table 8-1**

Electron-Dot Structures								
Group	1A	2A	3A	4A	5A	6A	7A	8A
Diagram	Li·	·Be·	·B·	·C·	·N·	·O·	·F·	·Ne·

Recall that an electron-dot structure is a type of diagram used to keep track of valence electrons and is especially useful when illustrating the formation of chemical bonds. **Table 8-1** shows several examples of electron-dot structures. For example, carbon has an electron configuration of $1s^2 2s^2 2p^2$. Its valence electrons are those in the second energy level, as can be seen in the electron-dot structure for carbon in the table.

Recall from Chapter 6 that ionization energy refers to how easily an atom loses an electron. The term electron affinity indicates how much attraction an atom has for electrons. Noble gases, having high ionization energies and low electron affinities, show a general lack of chemical reactivity. Other elements on the periodic table react with each other, forming numerous compounds. The difference in reactivity is directly related to the valence electrons.

All atoms have valence electrons. Why does this difference in reactivity of elements exist? Noble gases have electron configurations that have a full outermost energy level. This level is full with two electrons for helium ($1s^2$). The other noble gases have electron configurations consisting of eight electrons in the outermost energy level, $ns^2 np^6$. As you will recall, the presence of eight valence electrons in the outer energy level is chemically stable and is called a stable octet. Elements tend to react to acquire the stable electron structure of a noble gas.

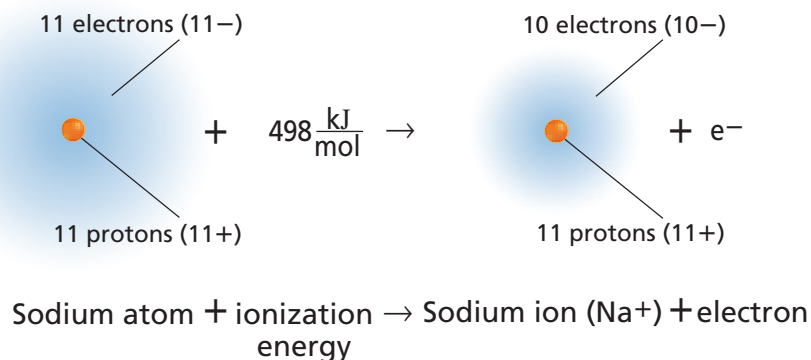
Formation of positive ions Recall that a positive ion forms when an atom loses one or more valence electrons in order to attain a noble gas configuration. To understand the formation of a positive ion, compare the electron configurations of the noble gas neon, atomic number 10, and the alkali metal sodium, atomic number 11.

Neon	$1s^2 2s^2 2p^6$
Sodium	$1s^2 2s^2 2p^6 3s^1$

Note that the sodium atom has one 3s valence electron; it differs from the noble gas neon by that single valence electron. If sodium loses this outer valence electron, the resulting electron configuration will be identical to that of neon. **Figure 8-1** shows how a sodium atom loses its valence electron to become a positive sodium ion. A positively charged ion is called a **cation**.

Figure 8-1

In the formation of a positive ion, a neutral atom loses one or more valence electrons. Note that the number of protons is equal to the number of electrons in the uncharged atom, but the ion contains more protons than electrons, making the overall charge on this ion positive.





By losing an electron, the sodium atom acquired the stable outer electron configuration of neon. It is important to understand that although sodium now has the electron configuration of neon, it is not neon. It is a sodium ion with a single positive charge. The 11 protons that establish the character of sodium still remain within its nucleus.

Reactivity of metals is based on the ease with which they lose valence electrons to achieve a stable octet, or noble gas configuration. Group 1A elements, [noble gas] ns^1 , lose their one valence electron, forming an ion with a 1+ charge. Group 2A elements, [noble gas] ns^2 , lose their two valence electrons and form ions with a 2+ charge. For example, potassium, a group 1A element, forms a K^+ ion; magnesium, a group 2A element, forms a Mg^{2+} ion. These two groups contain the most active metals on the periodic table. Some elements in group 3A, [noble gas] ns^2np^1 , also lose electrons and form positive ions. What is the charge on these ions? What is the formula for the aluminum ion?

Recall that, in general, transition metals have an outer energy level of ns^2 . Going from left to right across a period, atoms of each element are filling an inner d sublevel. When forming positive ions, transition metals commonly lose their two valence electrons, forming 2+ ions. However, it is also possible for d electrons to be lost. Thus transition elements also commonly form ions of 3+ or greater, depending on the number of d electrons in the electron structure. It is difficult to predict the number of electrons lost by transition elements. A useful rule of thumb for these metals is that they form ions with a 2+ or 3+ charge.

Although the formation of an octet is the most stable electron configuration, other electron configurations provide some stability. For example, elements in groups 1B through 4A in periods 4 through 6 lose electrons to form an outer energy level containing full s, p, and d sublevels. These relatively stable electron arrangements are referred to as *pseudo-noble gas configurations*. Let's examine the formation of the zinc ion, which is shown in **Figure 8-2**. The zinc atom has the electron configuration of $1s^22s^22p^63s^23p^64s^23d^{10}$. When forming an ion, the zinc atom loses the two 4s electrons in the outer energy level, and the stable configuration of $1s^22s^22p^63s^23p^63d^{10}$ results in a pseudo-noble gas configuration.

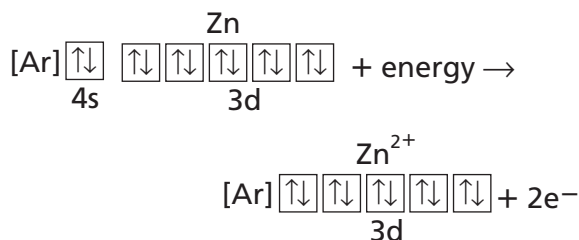
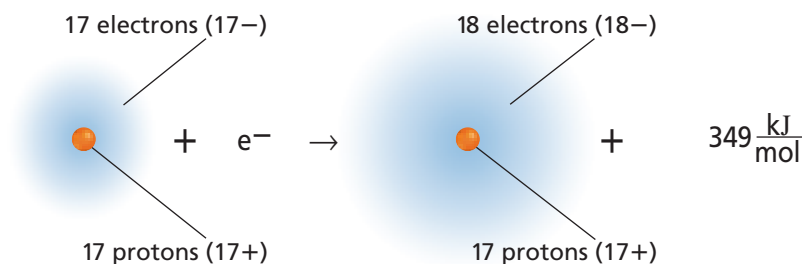


Figure 8-2

Orbital notation provides a convenient way to visualize the loss or gain of valence electrons. When zinc metal reacts with sulfuric acid, the zinc forms a Zn^{2+} ion with a pseudo-noble gas configuration.

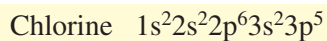
**Figure 8-3**

In the formation of a negative ion, a neutral atom gains one or more electrons. Again, note that in the neutral atom the number of protons equals the number of electrons. However, the ion contains more electrons than protons, making this overall charge on this ion negative.

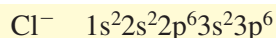
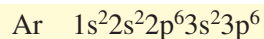


Chlorine atom + electron \rightarrow Chloride ion (Cl^-) + electron affinity

Formation of negative ions Recall that nonmetals, located on the right side of the periodic table, have a great attraction for electrons and form a stable outer electron configuration by gaining electrons. The chlorine atom, a halogen from group 7A, provides a good example.



Examine **Figure 8-3**. To attain a noble gas configuration, chlorine gains one electron, forming a negative ion with a $1-$ charge. By gaining the single electron, the chlorine atom now has the electron configuration of argon.



With the addition of one electron, chlorine becomes an **anion**, which is another name for a negative ion. To designate an anion, the ending *-ide* is added to the root name of the element. Thus the anion of chlorine is called the chloride ion. What is the name of the anion formed from nitrogen?

Nonmetals gain the number of electrons that, when added to their valence electrons, equals eight. Phosphorus, a group 5A element with the electron configuration of $[\text{Ne}]3s^2 3p^3$, has five valence electrons. To form a stable octet, the phosphorus atom may gain three electrons and form the phosphide ion with a $3-$ charge. If an oxygen atom, a group 6A element, gains two electrons, the oxide ion with a charge of $2-$ results.

Some nonmetals can lose or gain other numbers of electrons to form an octet. For example, in addition to gaining three electrons, phosphorus can lose five. However, in general, group 5A elements gain three electrons, group 6A gain two, and group 7A gain one to achieve an octet.

Section

8.1

Assessment

1. What is a chemical bond?
2. Why do ions form?
3. What family of elements is relatively unreactive and why?
4. Describe the formation of both positive and negative ions.
5. **Thinking Critically** Predict the change that must occur in the electron configuration if each of the

following atoms is to achieve a noble gas configuration.

- | | |
|-------------|------------|
| a. nitrogen | c. barium |
| b. sulfur | d. lithium |

6. **Formulating Models** Draw models to represent the formation of the positive calcium ion and the negative bromide ion.



Section

8.2

The Formation and Nature of Ionic Bonds

Look at the photos in **Figure 8-4a** and **b**. What do these reactions have in common? As you can see, in both cases, elements react with each other to form a compound. What happens in the formation of a compound?

Formation of an Ionic Bond

Figure 8-4a shows the reaction between the elements sodium and chlorine. During this reaction, a sodium (Na) atom transfers its valence electron to a chlorine (Cl) atom and becomes a positive ion. The chlorine atom accepts the electron into its outer energy level and becomes a negative ion. The compound sodium chloride forms because of the attraction between oppositely charged sodium and chloride ions. The electrostatic force that holds oppositely charged particles together in an ionic compound is referred to as an **ionic bond**. Compounds that contain ionic bonds are ionic compounds. If ionic bonds occur between metals and the nonmetal oxygen, oxides form. Most other ionic compounds are called salts.

Hundreds of compounds contain ionic bonds. Many ionic compounds are binary, which means that they contain only two different elements. Binary ionic compounds contain a metallic cation and a nonmetallic anion. Magnesium oxide, MgO , is a binary compound because it contains the two different elements magnesium and oxygen. However, CaSO_4 is not a binary compound. Can you explain why?

Consider the formation of the ionic compound calcium fluoride from calcium (Ca) and fluorine (F). Calcium, a group 2A metal with the electron configuration $[\text{Ar}]4s^2$, has two valence electrons. Fluorine, a group 7A nonmetal with the electron configuration $[\text{He}]2s^22p^5$, must gain one electron to attain the noble gas configuration of neon.

Because the number of electrons lost must equal the number of electrons gained, it will take two fluorine atoms to gain the two electrons lost from one

Objectives

- **Describe** the formation of ionic bonds.
- **Account** for many of the physical properties of an ionic compound.
- **Discuss** the energy involved in the formation of an ionic bond.

Vocabulary

ionic bond
electrolyte
lattice energy



Figure 8-4

These chemical reactions that produce ionic compounds also release a large amount of energy.

- a** The reaction that occurs between elemental sodium and chlorine gas produces a white crystalline solid.
- b** This sparkler contains iron, which burns in air to produce an ionic compound that contains iron and oxygen.





calcium atom. The compound formed will contain one calcium ion with a charge of 2+ for every two fluoride ions, each with a charge of 1-. Note that the overall charge on one unit of this compound is zero.

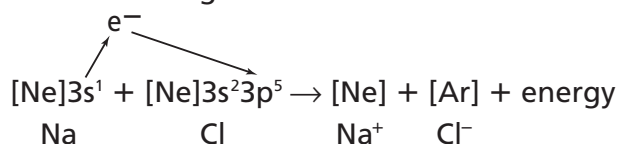
$$1 \text{ Ca ion} \left(\frac{2+}{\text{Ca ion}} \right) + 2 \text{ F ions} \left(\frac{1-}{\text{F ion}} \right) = (2+) + 2(1-) = 0$$

Figure 8-5

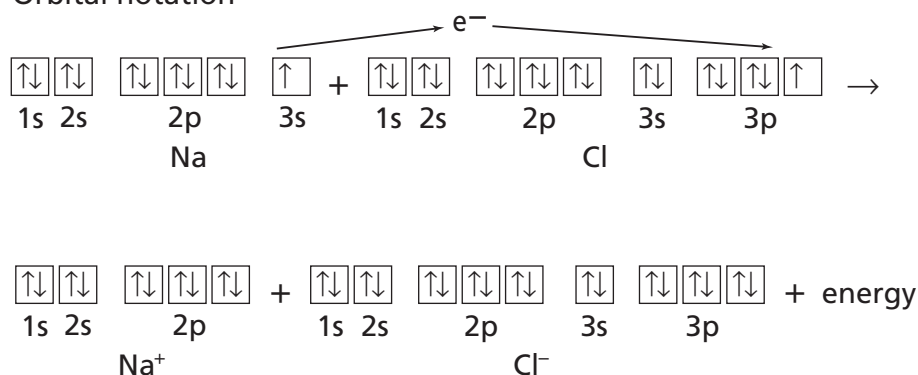
Several methods are used to show how an ionic compound forms.

Figure 8-5 summarizes the formation of an ionic compound from the elements sodium and chlorine using four different methods: electron configuration, orbital notation, electron-dot structures, and atomic models.

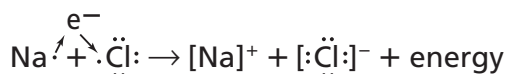
Electron configuration



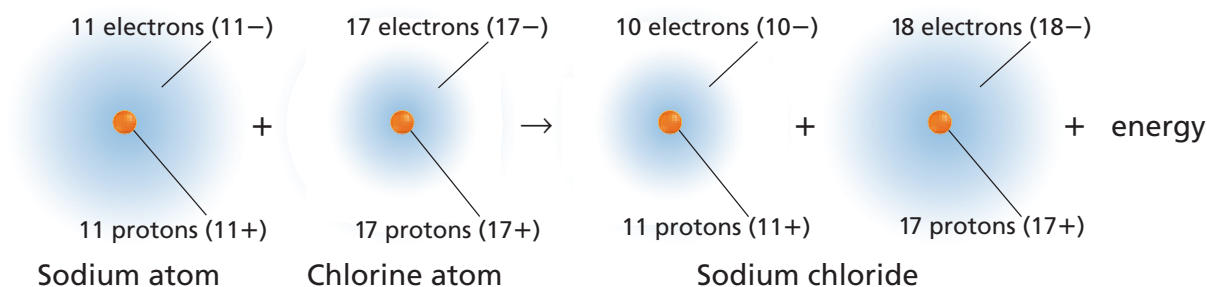
Orbital notation



Electron-dot structures



Atomic models





EXAMPLE PROBLEM 8-1

Formation of an Ionic Compound

Unprotected aluminum metal reacts with oxygen in air, forming the white coating you can observe on aluminum objects such as lawn furniture. Explain the formation of an ionic compound from the elements aluminum and oxygen.

1. Analyze the Problem

You are given that aluminum and oxygen react to form an ionic compound. Aluminum is a group 3A element with three valence electrons, and oxygen is a group 6A element with six valence electrons. To acquire a noble gas configuration, each aluminum atom must lose three electrons and each oxygen atom must gain two electrons.

2. Solve for the Unknown

Remember that the number of electrons lost must equal the number of electrons gained. The smallest number evenly divisible by the three electrons lost by aluminum and the two gained by oxygen is six. Three oxygen atoms are needed to gain the six electrons lost by two aluminum atoms.

3. Evaluate the Answer

The overall charge on one unit of this compound is zero.

$$2 \text{ Al ions } \left(\frac{3+}{\text{Al ion}} \right) + 3 \text{ O ions } \left(\frac{2-}{\text{O ion}} \right) = 2(3+) + 3(2-) = 0$$



Topic: Ionic Compounds

To learn more about ionic compounds, visit the Chemistry Web site at chemistrymc.com

Activity: Research the colors of minerals and their chemical formulas. Which elements seem to be in the most colored compounds? What are their uses? Make a chart to report your findings.

PRACTICE PROBLEMS

Explain the formation of the ionic compound composed of each pair of elements.

- | | |
|---------------------------|---------------------------|
| 7. sodium and nitrogen | 10. aluminum and sulfur |
| 8. lithium and oxygen | 11. cesium and phosphorus |
| 9. strontium and fluorine | |



For more practice with forming ionic compounds, go to **Supplemental Practice Problems** in Appendix A.

Properties of Ionic Compounds

The chemical bonds that occur between the atoms in a compound determine many of the physical properties of the compound. During the formation of an ionic compound, the positive and negative ions are packed into a regular repeating pattern that balances the forces of attraction and repulsion between the ions. This particle packing forms an ionic crystal, as shown in **Figure 8-6**. No single unit consisting of only one ion attracting one other ion is formed. Large numbers of positive ions and negative ions exist together in a ratio determined by the number of electrons transferred from the metal to the nonmetal.

Examine the pattern of the ions in the sodium chloride crystal shown in the figure. What shape would you expect a large crystal of this compound to be? This one-to-one ratio of ions produces a cubic crystal. Examine some table salt (NaCl) under a magnifying glass. What shape are these small salt crystals?

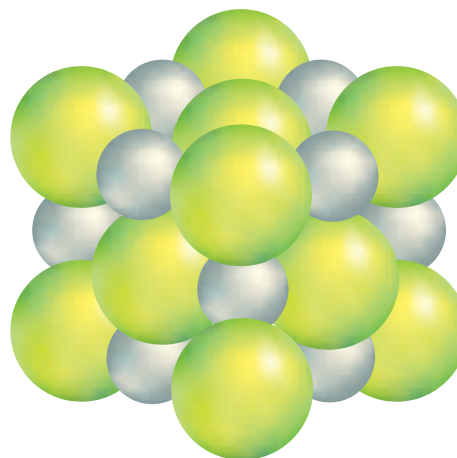


Figure 8-6

In this ionic compound, each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

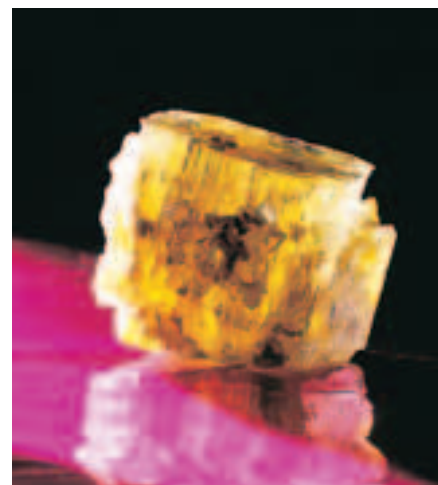


Figure 8-7

Aragonite (CaCO_3), barite (BaSO_4), and beryl ($\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$) are examples of minerals that are ionic compounds. The ions that form them are bonded together in a crystal lattice.


The strong attraction of positive ions and negative ions in an ionic compound results in a crystal lattice. A crystal lattice is a three-dimensional geometric arrangement of particles. In a crystal lattice, each positive ion is surrounded by negative ions and each negative ion is surrounded by positive ions. Ionic crystals vary in shape due to the sizes and relative numbers of the ions bonded, as shown in **Figure 8-7**.

Melting point, boiling point, and hardness are physical properties that depend on how strongly the particles are attracted to each other. Because ionic bonds are relatively strong, the crystals that result require a large amount of energy to be broken apart. Therefore, ionic crystals have high melting points and boiling points, as shown in **Table 8-2**. Their color may be related to their structure. See the **problem-solving LAB** on the next page and **Everyday Chemistry** at the end of this chapter. These crystals are also hard, rigid, and brittle solids due to the strong attractive forces that hold the ions in place. When an external force large enough to overcome the attraction of ions in the crystal is applied, the crystal cracks. The force repositions the like-charged ions next to each other, and the repulsive force cracks the crystal.

Table 8-2

Melting and Boiling Points of Some Ionic Compounds		
Compound	Melting point ($^{\circ}\text{C}$)	Boiling point ($^{\circ}\text{C}$)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
CaCl_2	782	>1600
CaI_2	784	1100
NaCl	801	1413
MgO	2852	3600

Charged particles must be free to move for a material to conduct an electric current. In the solid state, ionic compounds are nonconductors of electricity because of the fixed positions of the ions. However, in a liquid state or when dissolved in water, ionic compounds are electrical conductors because the ions are free to move. An ionic compound whose aqueous solution conducts an electric current is called an **electrolyte**. You will learn more about solutions of electrolytes in Chapter 15.

 **LAB**

See page 955 in Appendix E for
Comparing Sport Drink Electrolytes





Energy and the ionic bond During any chemical reaction, energy is either absorbed or released. When energy is absorbed during a chemical reaction, the reaction is endothermic. If energy is released, it is exothermic. As you will see in the **CHEMLAB** at the end of this chapter, energy is released when magnesium reacts with oxygen.

Energy changes also occur during the formation of ionic bonds from the ions formed during a chemical reaction. The formation of ionic compounds from positive and negative ions is always exothermic. The attraction of the positive ion for the negative ions close to it forms a more stable system that is lower in energy than the individual ions. If the amount of energy released during bond formation is added to an ionic compound, the bonds that hold the positive and negative ions together break.

You just learned that the ions in an ionic compound are arranged in a pattern in a crystal lattice. The energy required to separate one mole of the ions of an ionic compound is referred to as the **lattice energy**. The strength of the forces holding ions in place is reflected by the lattice energy. The more negative the lattice energy, the stronger the force of attraction.

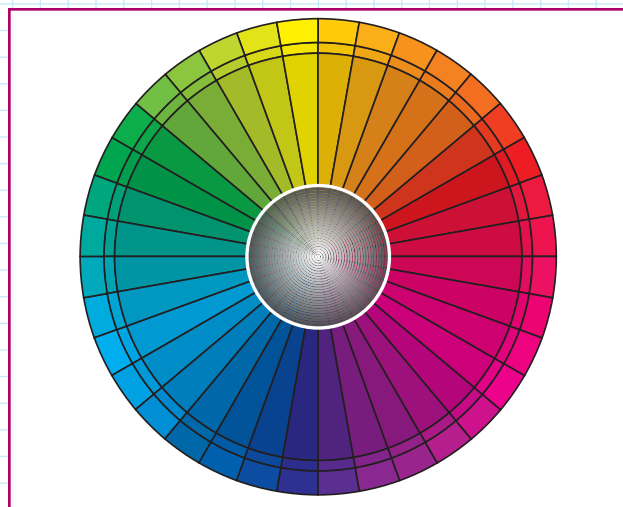
problem-solving LAB

How is color related to a transferred electron?

Factors that Affect Color of an Ionic Compound

Compound	Color	Anion radius (Å)	Visible absorption band
AgF	Yellow	1.36	Blue-violet
AgCl	White	1.81	None
AgBr	Cream	1.95	Violet
AgI	Yellow	2.16	Blue-violet
Ag ₂ S	Black	1.84	All
Al ₂ O ₃	White	1.40	None
Sb ₂ O ₃	White	1.40	None
Bi ₂ O ₃	?	1.40	Violet

Predicting Once an ionic bond is formed, the cation has a tendency to pull the transferred electron toward the nucleus. The appearance of color is directly related to the strength of the pull, which depends upon the size of the ions involved and their oxidation numbers. If visible light of a certain color can send the electron back to the cation momentarily, then the light reflected from a crystal of the compound will be missing this color from its spectrum. The resulting color of the crystal will be the complement of this color of light and can be predicted using the color wheel that is shown here. Complementary colors are across from each other on the color wheel.



Analysis

The information in the table can be used to make some general conclusions about a compound's color and the strength of the cation's pull on the transferred electron. Using this information, the color of a compound can be predicted.

Thinking Critically

1. A larger anion radius results in a more pronounced color. What reason can you give for this fact?
2. Which do you think produces a more pronounced color, a high oxidation state for the anions or a low one? Explain.
3. Use the color wheel to predict the color of Bi₂O₃, the last compound on the list.



Lattice energy is directly related to the size of the ions bonded. Smaller ions generally have a more negative value for lattice energy because the nucleus is closer to and thus has more attraction for the valence electrons. Thus, the lattice energy of a lithium compound is more negative than that of a potassium compound containing the same anion because the lithium ion is smaller than the potassium ion. Which would have a more negative lattice energy, lithium chloride or lithium bromide?

The value of lattice energy is also affected by the charge of the ion. The ionic bond formed from the attraction of ions with larger positive or negative charges generally has a more negative lattice energy. The lattice energy of MgO is almost four times greater than the lattice energy of NaF because the charge of the ions is greater. The lattice energy of SrCl₂ is between the lattice energies of MgO and NaF because SrCl₂ contains ions with both higher and lower charges.

Table 8-3 shows the lattice energies of some ionic compounds. Examine the lattice energies of RbF and KF. How do they confirm that lattice energy is related to ion size? Look at the lattice energies of SrCl₂ and AgCl. How do they show the relationship between lattice energy and the charge of the ions involved?

Table 8-3

Lattice Energies of Some Ionic Compounds			
Compound	Lattice energy (kJ/mol)	Compound	Lattice energy (kJ/mol)
KI	−632	KF	−808
KBr	−671	AgCl	−910
RbF	−774	NaF	−910
NaI	−682	LiF	−1030
NaBr	−732	SrCl ₂	−2142
NaCl	−769	MgO	−3795

Section

8.2

Assessment

- What is an ionic bond?
- How does an ionic bond form?
- List three physical properties associated with an ionic bond.
- Describe the arrangement of ions in a crystal lattice.
- What is lattice energy and how is it involved in an ionic bond?
- Thinking Critically** Using the concepts of ionic radii and lattice energy, account for the trend in melting points shown in the following table.

Trend in Melting Points	
Ionic compound	Melting point in °C
KF	858
KCl	770
KBr	734
KI	681

- Formulating Models** Use electron configurations, orbital notation, and electron-dot structures to represent the formation of an ionic compound from the metal strontium and the nonmetal chlorine.





Section

8.3

Names and Formulas for Ionic Compounds

One of the most important requirements of chemistry is communicating information to others. Chemists discuss compounds by using both chemical formulas and names. The chemical formula and the name for the compound must be understood universally. Therefore, a set of rules is used in the naming of compounds. This system of naming allows everyone to write a chemical formula when given a compound name and to name the compound from a given chemical formula.

Formulas for Ionic Compounds

Recall from Section 8.2 that a sample of an ionic compound contains crystals formed from many ions arranged in a pattern. Because no single particle of an ionic compound exists, ionic compounds are represented by a formula that provides the simplest ratio of the ions involved. The simplest ratio of the ions represented in an ionic compound is called a **formula unit**. For example, the formula KBr represents a formula unit for potassium bromide because potassium and bromide ions are in a one-to-one ratio in the compound. A formula unit of magnesium chloride is MgCl_2 because two chloride ions exist for each magnesium ion in the compound. In the compound sodium phosphide, three sodium ions exist for every phosphide ion. What is the formula unit for sodium phosphide?

Because the total number of electrons gained by the nonmetallic atoms must equal the total number of electrons lost by the metallic atoms, the overall charge of a formula unit is zero. The formula unit for MgCl_2 contains one Mg^{2+} ion and two Cl^- ions, for a total charge of zero.

Determining charge Binary ionic compounds are composed of positively charged monatomic ions of a metal and negatively charged monatomic ions of a nonmetal. A **monatomic ion** is a one-atom ion, such as Mg^{2+} or Br^- . **Table 8-4** indicates the charges of common monatomic ions according to the location of their atoms on the periodic table. What is the formula for the beryllium ion? The iodide ion? The nitride ion? Transition metals, which are in groups 3B through 2B, and metals in groups 3A and 4A are not included in this table because of the variance in ionic charges of atoms in the groups. Most transition metals and those in groups 3A and 4A can form several different positive ions.

Table 8-4

Common Ions Based on Groups		
Group	Atoms that commonly form ions	Charge on ions
1A	H, Li, Na, K, Rb, Cs	1+
2A	Be, Mg, Ca, Sr, Ba	2+
5A	N, P, As	3-
6A	O, S, Se, Te	2-
7A	F, Cl, Br, I	1-

Objectives

- **Write** formulas for ionic compounds and oxyanions.
- **Name** ionic compounds and oxyanions.

Vocabulary

formula unit
monatomic ion
oxidation number
polyatomic ion
oxyanion



Careers Using Chemistry

Wastewater Treatment Operator

Would you be interested in a job that assures your community of a safe water supply? Then consider a career in wastewater treatment.

Not only does our water supply have to be safe for human consumption, the water that is returned to rivers and streams must be cleaned of pathogens and suspended solids so it can be used over and over. This career involves testing to identify chemicals, pathogens, and materials in the wastewater, as well as monitoring the multi-step process of their removal.

Table 8-5

Common Ions of Transition Metals and Groups 3A and 4A	
Group	Common ions
3B	Sc^{3+} , Y^{3+} , La^{3+}
4B	Ti^{2+} , Ti^{3+}
5B	V^{2+} , V^{3+}
6B	Cr^{2+} , Cr^{3+}
7B	Mn^{2+} , Mn^{3+} , Tc^{2+}
8B	Fe^{2+} , Fe^{3+}
8B	Co^{2+} , Co^{3+}
8B	Ni^{2+} , Pd^{2+} , Pt^{2+} , Pt^{4+}
1B	Cu^{+} , Cu^{2+} , Ag^{+} , Au^{+} , Au^{3+}
2B	Zn^{2+} , Cd^{2+} , Hg_2^{2+} , Hg^{2+}
3A	Al^{3+} , Ga^{2+} , Ga^{3+} , In^{+} , In^{2+} , In^{3+} , Tl^{+} , Tl^{3+}
4A	Sn^{2+} , Sn^{4+} , Pb^{2+} , Pb^{4+}

The charge of a monatomic ion is its **oxidation number**. Most transition metals and group 3A and 4A metals have more than one oxidation number, as shown in **Table 8-5**. The oxidation numbers given in the table are the most common ones for many of the elements listed but might not be the only ones possible.

The term *oxidation state* is sometimes used and means the same thing as oxidation number. The oxidation number, or oxidation state, of an element in an ionic compound equals the number of electrons transferred from an atom of the element to form the ion. For example, when sodium and chlorine atoms react, the sodium atom transfers one electron to the chlorine atom, forming Na^{+} and Cl^{-} . Thus, in the compound formed, the oxidation state of sodium is $1+$ because one electron is transferred from the sodium atom. The oxidation state of chlorine is $1-$. One electron is transferred, and the negative sign shows that the electron transferred to, not from, the chlorine atom.

The oxidation numbers of ions are used to determine the formulas for the ionic compounds they form. Recall that in ionic compounds, oppositely charged ions combine chemically in definite ratios to form a compound that has no charge. If you add the oxidation number of each ion multiplied by the number of these ions in a formula unit, the total must be zero.

In the chemical formula for any ionic compound, the symbol of the cation is always written first, followed by the symbol of the anion. Subscripts, which are small numbers to the lower right of a symbol, are used to represent the number of ions of each element in an ionic compound. If no subscript is written, it is assumed to be one.

Suppose you need to determine the formula for one formula unit of the compound that contains sodium and chloride ions. Write the symbol and charge for each ion.



The ratio of ions must be such that the number of electrons lost by the metal is equal to the number of electrons gained by the nonmetal. Because the sum of the oxidation numbers of these ions is zero, these ions must be present in a one-to-one ratio. One sodium ion transfers one electron to one chloride ion, and the formula unit is NaCl .



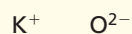
EXAMPLE PROBLEM 8-2

Determining the Formula for an Ionic Compound

The ionic compound formed from potassium and oxygen is used as a dehydrating agent because it reacts readily with water. Determine the correct formula for the ionic compound formed from potassium and oxygen.

1. Analyze the Problem

It is given that potassium and oxygen ions form an ionic compound. The first thing to do is determine the symbol and oxidation number for each ion involved in the ionic compound and write them as shown.



If the charges are not the same, subscripts must be determined to indicate the ratio of positive ions to negative ions.

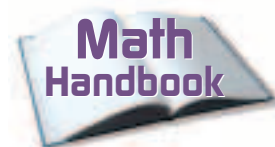
2. Solve for the Unknown

A potassium atom loses one electron while an oxygen atom gains two electrons. If combined in a one-to-one ratio, the number of electrons lost by potassium will not balance the number of electrons gained by oxygen. To have the same number of electrons lost and gained, you must have two potassium ions for every oxide ion. The correct formula is K_2O .

3. Evaluate the Answer

The overall charge on one formula unit of this compound is zero.

$$2 \text{ K-ions } \left(\frac{1+}{\text{K-ion}} \right) + 1 \text{ O-ion } \left(\frac{2-}{\text{O-ion}} \right) = 2(1+) + 1(2-) = 0$$



Review least common multiple in the **Math Handbook** on page 909 of this text.

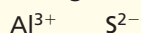
EXAMPLE PROBLEM 8-3

Determining the Formula for an Ionic Compound

Determine the correct formula for the yellowish-gray compound formed from aluminum ions and sulfide ions. This compound decomposes in moist air.

1. Analyze the Problem

You are given that aluminum and sulfur ions form an ionic compound. First, determine the charge of each ion involved.



Each aluminum atom loses three electrons while each sulfur atom gains two. The number of electrons lost must equal the number of electrons gained.

2. Solve for the Unknown

The smallest number that both two and three divide into evenly is six. Therefore, a total of six electrons was transferred. Three sulfur atoms accept the six electrons lost by two aluminum atoms. The correct formula will show two aluminum ions bonded to three sulfur ions, or Al_2S_3 .

3. Evaluate the Answer

The overall charge on one formula unit of this compound is zero.

$$2 \text{ Al-ions } \left(\frac{3+}{\text{Al-ion}} \right) + 3 \text{ S-ions } \left(\frac{2-}{\text{S-ion}} \right) = 2(3+) + 3(2-) = 0$$



Practice!

For more practice with writing formulas for ionic compounds, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Write the correct formula for the ionic compound composed of the following pairs of ions.

- | | |
|----------------------------|------------------------|
| 19. potassium and iodide | 22. cesium and nitride |
| 20. magnesium and chloride | 23. barium and sulfide |
| 21. aluminum and bromide | |

Compounds that contain polyatomic ions Many ionic compounds contain **polyatomic ions**, which are ions made up of more than one atom. **Table 8-6** lists the formulas and the charges for several polyatomic ions.

The charge given to a polyatomic ion applies to the entire group of atoms. Although an ionic compound containing one or more polyatomic ions contains more than two atoms, the polyatomic ion acts as an individual ion. Therefore, the chemical formula for the compound can be written following the same rules used for a binary compound.

Because a polyatomic ion exists as a unit, never change subscripts of the atoms within the ion. If more than one polyatomic ion is needed, place parentheses around the ion and write the appropriate subscript outside the parentheses. For example, the formula for magnesium chlorate is $\text{Mg}(\text{ClO}_3)_2$. Note that the ammonium ion is the only common polyatomic cation.

How can you determine the formula unit for an ionic compound containing a polyatomic ion? Chemists use a naming system called the Stock System, after the German chemist Alfred Stock. Let's consider the compound formed from the ammonium ion and the chloride ion.



Because the sum of the charges on the ions is zero, the ions are in a one-to-one ratio. The correct formula unit for this compound is NH_4Cl .

Table 8-6

Common Polyatomic Ions			
Ion	Name	Ion	Name
NH_4^+	ammonium	IO_4^-	periodate
NO_2^-	nitrite	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
NO_3^-	nitrate	H_2PO_4^-	dihydrogen phosphate
HSO_4^-	hydrogen sulfate	CO_3^{2-}	carbonate
OH^-	hydroxide	SO_3^{2-}	sulfite
CN^-	cyanide	SO_4^{2-}	sulfate
MnO_4^-	permanganate	$\text{S}_2\text{O}_3^{2-}$	thiosulfate
HCO_3^-	hydrogen carbonate	O_2^{2-}	peroxide
ClO^-	hypochlorite	CrO_4^{2-}	chromate
ClO_2^-	chlorite	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
ClO_3^-	chlorate	HPO_4^{2-}	hydrogen phosphate
ClO_4^-	perchlorate	PO_4^{3-}	phosphate
BrO_3^-	bromate	AsO_4^{3-}	arsenate
IO_3^-	iodate		



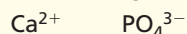
EXAMPLE PROBLEM 8-4

Determining the Formula for an Ionic Compound Containing a Polyatomic Ion

The ionic compound formed from the calcium ion and the phosphate ion is a common ingredient in fertilizers. Write the formula for this compound.

1. Analyze the Problem

It is given that calcium and phosphate ions form an ionic compound. You should first write each ion along with its charge.



Because the numerical values of the charges differ, a one-to-one ratio is not possible.

2. Solve for the Unknown

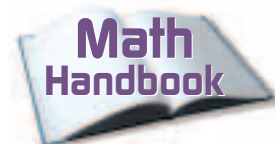
Six is the smallest number evenly divisible by both ionic charges. Therefore, a total of six electrons were transferred. The amount of negative charge of two phosphate ions equals the amount of positive charge of three calcium ions. To use a subscript to indicate more than one unit of a polyatomic ion, you must place the polyatomic ion in parentheses and add the subscript to the outside. The correct formula is $\text{Ca}_3(\text{PO}_4)_2$.

3. Evaluate the Answer

The overall charge on one formula unit of calcium phosphate is zero.

$$3 \text{ calcium ions } \left(\frac{2+}{\text{calcium ion}} \right) + 2 \text{ phosphate ions } \left(\frac{3-}{\text{phosphate ion}} \right) =$$

$$3(2+) + 2(-3) = 0$$



Review positive and negative numbers in the **Math Handbook** on page 887 of this text.

PRACTICE PROBLEMS

Determine the correct formula for the ionic compound composed of the following pairs of ions.

24. sodium and nitrate

27. potassium and chromate

25. calcium and chlorate

28. magnesium and carbonate

26. aluminum and carbonate



For more practice with writing formulas for ionic compounds that contain polyatomic ions, go to **Supplemental Practice Problems** in Appendix A.

Naming Ions and Ionic Compounds

You already know how to name monatomic ions. How do you name polyatomic ions? Most polyatomic ions are oxyanions. An **oxyanion** is a polyatomic ion composed of an element, usually a nonmetal, bonded to one or more oxygen atoms. Many oxyanions contain the same nonmetal and have the same charges but differ in the number of oxygen atoms. More than one oxyanion exists for some nonmetals, such as nitrogen and sulfur. These ions are easily named using the following conventions.

- The ion with more oxygen atoms is named using the root of the nonmetal plus the suffix *-ate*.
- The ion with fewer oxygen atoms is named using the root of the nonmetal plus the suffix *-ite*.

For example:

NO_3^-
nitrate

NO_2^-
nitrite

SO_4^{2-}
sulfate

SO_3^{2-}
sulfite





Earth Science

CONNECTION

Mineralogists, the scientists who study minerals, use various classification schemes to organize the thousands of known minerals. Color; crystal structure; hardness; chemical, magnetic, and electrical properties; and numerous other characteristics are used to classify minerals.

The types of anions minerals contain also can identify them. For example, more than one-third of all known minerals are silicates, which are minerals that contain an anion that is a combination of silicon and oxygen. Halides contain fluoride, chloride, bromide, or iodide ions. Minerals in which the anions contain boron are called borates.



Chlorine in group 7A, the halogens, forms four oxyanions. These oxyanions are named according to the number of oxygen atoms present. The following conventions are used to name these oxyanions.

- The oxyanion with the greatest number of oxygen atoms is named using the prefix *per-*, the root of the nonmetal, and the suffix *-ate*.
- The oxyanion with one less oxygen atom is named with the root of the nonmetal and the suffix *-ate*.
- The oxyanion with two fewer oxygen atoms is named using the root of the nonmetal plus the suffix *-ite*.
- The oxyanion with three fewer oxygen atoms is named using the prefix *hypo-*, the root of the nonmetal, and the suffix *-ite*.

ClO_4^-	ClO_3^-	ClO_2^-	ClO^-
perchlorate	chlorate	chlorite	hypochlorite

Other halogens form oxyanions that are named similarly to the oxyanions chlorine forms. Bromine forms BrO_3^- , the bromate ion. Iodine forms the periodate ion (IO_4^-) and the iodate ion (IO_3^-).

Naming ionic compounds Chemical nomenclature is a systematic way of naming compounds. Now that you are familiar with writing chemical formulas, you will use the following general rules in naming ionic compounds when their formulas are known.

1. *Name the cation first and the anion second.* Remember that the cation is always written first in the formula. For example, CsBr is a compound used in X-ray fluorescent screens. In the formula CsBr , Cs^+ is the cation and is named first. The anion is Br^- and is named second.
2. *Monatomic cations use the element name.* The name of the cation Cs^+ is cesium, the name of the element.
3. *Monatomic anions take their name from the root of the element name plus the suffix *-ide*.* The compound CsBr contains the bromide anion.
4. *Group 1A and group 2A metals have only one oxidation number. Transition metals and metals on the right side of the periodic table often have more than one oxidation number.* To distinguish between multiple oxidation numbers of the same element, the name of the chemical formula must indicate the oxidation number of the cation. The oxidation number is written as a Roman numeral in parentheses after the name of the cation. For example, the compound formed from Fe^{2+} and O^{2-} has the formula FeO and is named iron(II) oxide. The compound formed from Fe^{3+} and O^{2-} has the formula Fe_2O_3 and is named iron(III) oxide.
5. *If the compound contains a polyatomic ion, simply name the ion.* The name of the compound that contains the sodium cation and the polyatomic hydroxide anion, NaOH , is sodium hydroxide. The compound $(\text{NH}_4)_2\text{S}$ is ammonium sulfide.

PRACTICE PROBLEMS

Name the following compounds.

29. NaBr

30. CaCl_2

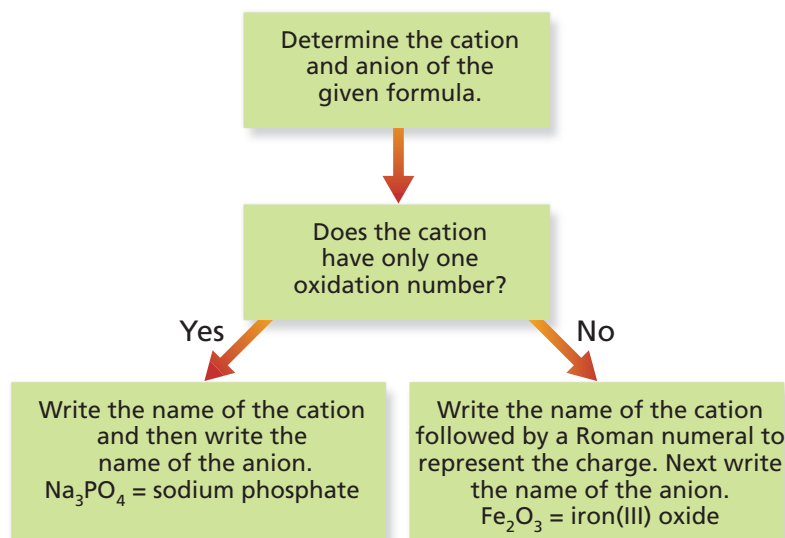
31. KOH

32. $\text{Cu}(\text{NO}_3)_2$

33. Ag_2CrO_4

Practice!

For more practice with naming ionic compounds, go to **Supplemental Practice Problems** in Appendix A.

**Figure 8-8**

This diagram summarizes how to name ionic compounds from their formulas.

Figure 8-8 reviews the steps used in naming ionic compounds if the formula is known. Naming ionic compounds is important in communicating the cation and anion present in a crystalline solid or aqueous solution. How might you change the diagram to help you write the formulas for ionic compounds if you know their names?

All the ion-containing substances you have investigated so far have been ionic compounds. Do any other substances contain ions? Can certain elements contain ions and still be electrically neutral? Do the properties of other ion-containing substances differ from the properties of ionic compounds? In the next section, you will learn the answers to these questions by examining how ions relate to the structure and properties of metals.

Section 8.3 Assessment

34. What is the difference between a monatomic ion and a polyatomic ion? Give an example of each.
35. How do you determine the correct subscripts in a chemical formula?
36. How are metals named in an ionic compound? Nonmetals? Polyatomic ions?
37. What is an oxyanion and how is it named?
38. **Thinking Critically** What subscripts would most likely be used if the following substances formed an ionic compound?
 - a. an alkali metal and a halogen
 - b. an alkali metal and a nonmetal from group 6A
 - c. an alkaline earth metal and a halogen
 - d. an alkaline earth metal and a nonmetal from group 6A
 - e. a metal from group 3A and a halogen

39. **Making and Using Tables** Complete the table below by providing the correct formula for each compound formed from the listed ions.

Formulas for Some Ionic Compounds				
	Oxide	Chloride	Sulfate	Phosphate
Potassium				
Barium				
Aluminum				
Ammonium				





Section 8.4

Metallic Bonds and Properties of Metals

Objectives

- **Describe** a metallic bond.
- **Explain** the physical properties of metals in terms of metallic bonds.
- **Define** and **describe** alloys.

Vocabulary

electron sea model
delocalized electrons
metallic bond
alloy

Although metals are not ionic, they share several properties with ionic compounds. Properties of materials are based on bonding, and the bonding in both metals and ionic compounds is based on the attraction of particles with unlike charges.

Metallic Bonds

Although metals do not bond ionically, they often form lattices in the solid state. These lattices are similar to the ionic crystal lattices that were discussed in Section 8.2. In such a lattice, eight to 12 other metal atoms surround each metal atom. Although metal atoms have at least one valence electron, they do not share these electrons with neighboring atoms nor do they lose electrons to form ions.

Instead, in this crowded condition, the outer energy levels of the metal atoms overlap. The **electron sea model** proposes that all the metal atoms in a metallic solid contribute their valence electrons to form a “sea” of electrons. The electrons present in the outer energy levels of the bonding metallic atoms are not held by any specific atom and can move easily from one atom to the next. Because they are free to move, they are often referred to as **delocalized electrons**. When the atom’s outer electrons move freely throughout the solid, a metallic cation is formed. Each such ion is bonded to all neighboring metal cations by the “sea” of valence electrons shown in **Figure 8-9**. A **metallic bond** is the attraction of a metallic cation for delocalized electrons.

Properties of metals The typical physical properties of metals can be explained by metallic bonding. These properties provide evidence of the strength of metallic bonds.

The melting points of metals vary greatly. Mercury is a liquid at room temperature, which makes it useful in scientific instruments such as thermometers and barometers. On the other hand, tungsten has a melting point of 3422°C, which makes it useful by itself or in combination with other metals

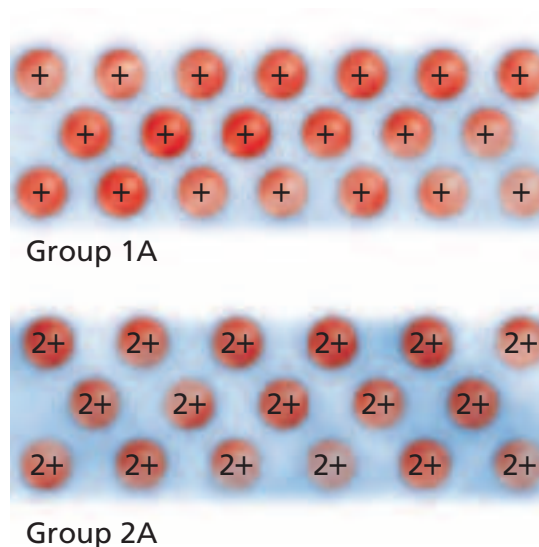


Figure 8-9

The valence electrons in metals (shown in blue) are evenly distributed among the metallic cations (shown in red). Attractions between the positive cations and negative “sea” hold the metal atoms together in a lattice.

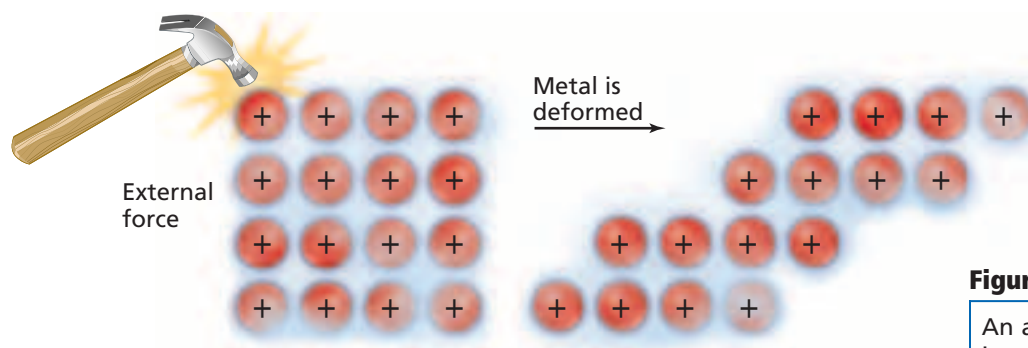


Figure 8-10

An applied force causes metal ions to move through delocalized electrons, making metals malleable and ductile.

for purposes that involve high temperatures or strength. Lightbulb filaments are usually made from tungsten, as are certain spacecraft parts. In general, metals have moderately high melting points and high boiling points, as shown in **Table 8-7**. The melting points are not as extreme as the boiling points because the cations and electrons are mobile in a metal. It does not take an extreme amount of energy for them to be able to move past each other. However, during boiling, atoms must be separated from the group of cations and electrons, which requires much more energy.

Metals are malleable, which means they can be hammered into sheets, and they are ductile, which means they can be drawn into wire. **Figure 8-10** shows how the mobile particles involved in metallic bonding can be pushed or pulled past each other, making metals malleable and ductile.

Metals are generally durable. Although metallic cations are mobile in a metal, they are strongly attracted to the electrons surrounding them and aren't easily removed from the metal.

Delocalized electrons in a metal are free to move, keeping metallic bonds intact. The movement of mobile electrons around positive metallic cations explains why metals are good conductors. The delocalized electrons move heat from one place to another much more quickly than the electrons in a material that does not contain mobile electrons. Mobile electrons easily move as a part of an electric current when electrical potential is applied to a metal. These same delocalized electrons interact with light, absorbing and releasing photons, thereby creating the property of luster in metals.

The mobile electrons in transition metals consist not only of the two outer s electrons but also the inner d electrons. As the number of delocalized electrons increases, so do the properties of hardness and strength. For example, strong metallic bonds are found in transition metals such as chromium, iron, and nickel, whereas alkali metals are considered soft because they have only one delocalized electron, ns^1 .

Table 8-7

Melting Points and Boiling Points of Some Metals		
Element	Melting point (°C)	Boiling point (°C)
Lithium	180	1347
Tin	232	2623
Aluminum	660	2467
Barium	727	1850
Silver	961	2155
Copper	1083	2570



Metal Alloys

Due to the nature of a metallic bond, it is relatively easy to introduce other elements into a metallic crystal, forming an alloy. An **alloy** is a mixture of elements that has metallic properties. **Table 8-8** lists some commercially important alloys and their uses. A company that manufactures trophies probably would use which alloy listed in the table?

The properties of alloys differ somewhat from the properties of the elements they contain. For example, steel is iron mixed with at least one other element. Some properties of iron are present, but steel has additional properties, such as increased strength. Some alloys, such as that used in the **miniLAB**, vary in properties depending on how they are manufactured.

miniLAB

Heat Treatment of Steel

Recognizing Cause and Effect People have treated metals with heat for many centuries. Different properties result when the metal is slowly or rapidly cooled. Can you determine how and why the properties change?

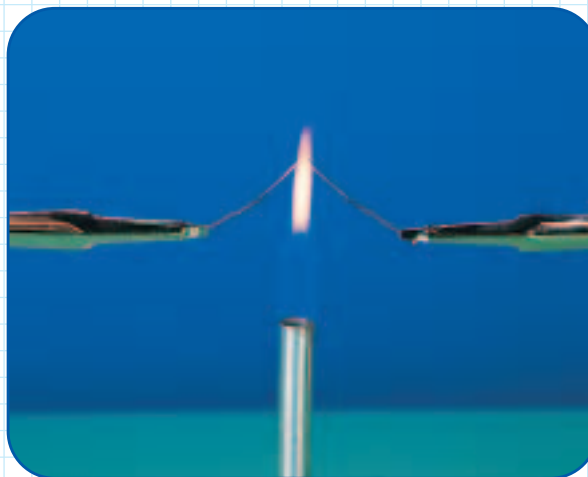
Materials laboratory burner, forceps (2), hairpins (3), 250-mL beaker

Procedure 

1. Examine a property of spring steel by trying to bend open one of the hairpins. Record your observations.
2. Hold each end of a hairpin with forceps. Place the curved central loop in the top of the burner's flame. When it turns red, pull it open into a straight piece of metal. Allow it to cool as you record your observations. Repeat this procedure for the remaining two hairpins.

CAUTION: Do not touch the hot metal.

3. To make softened steel, use forceps to hold all three hairpins vertically in the flame until they glow red all over. Slowly raise the three hairpins straight up and out of the flame so they cool slowly. Slow cooling results in the formation of large crystals.
4. After cooling, bend each of the three hairpins into the shape of the letter J. Record how the metal feels as you bend it.
5. To harden the steel, use tongs to hold two of the bent hairpins in the flame until they are glowing red all over. Quickly plunge the hot metals into a 250-mL beaker containing approximately 200 mL of cold water. Quick-cooling causes the crystal size to be small.
6. Attempt to straighten one of the bends. Record your observations.



7. To temper the steel, use tongs to briefly hold the remaining hardened metal bend above the flame. Slowly move the metal back and forth just above the flame until the gray metal turns to an iridescent blue-gray color. Do not allow the metal to glow red. Slowly cool the metal and then try to unbend it using the end of your finger. Record your observations.

Analysis

1. State a use for spring steel that takes advantage of its unique properties.
2. What are the advantages and disadvantages of using softened steel for body panels on automobiles?
3. What is the major disadvantage of hardened steel? Do you think this form of iron would be wear resistant and retain a sharpened edge?
4. Which two types of steel appear to have their properties combined in tempered steel?
5. State a hypothesis that explains how the different properties you have observed relate to crystal size.



Table 8-8

Some Commercially Important Alloys		
Common name	Composition	Uses
Alnico	Fe 50%, Al 20%, Ni 20%, Co 10%	Magnets
Brass	Cu 67-90%, Zn 10-33%	Plumbing, hardware, lighting
Bronze	Cu 70-95%, Zn 1-25%, Sn 1-18%	Bearings, bells, medals
Cast iron	Fe 96-97%, C 3-4%	Casting
Dental amalgam	Hg 50%, Ag 35%, Sn 15%	Dental fillings
Gold, 10 carat	Au 42%, Ag 12-20%, Cu 38-46%	Jewelry
Lead shot	Pb 99.8%, As 0.2%	Shotgun shells
Pewter	Sn 70-95%, Sb 5-15%, Pb 0-15%	Tableware
Stainless steel	Fe 73-79%, Cr 14-18%, Ni 7-9%	Instruments, sinks
Sterling silver	Ag 92.5%, Cu 7.5%	Tableware, jewelry

Alloys most commonly form when the elements involved are either similar in size or the atoms of one element are considerably smaller than the atoms of the other. Thus, two basic types of alloys exist, substitutional and interstitial, and many industries depend on their production. A substitutional alloy has atoms of the original metallic solid replaced by other metal atoms of similar size. Sterling silver is an example of a substitutional alloy. When copper atoms replace silver atoms in the original metallic crystal, a solid with properties of both silver and copper is formed. Brass, pewter, and 10-carat gold are all examples of substitutional alloys.

An interstitial alloy is formed when the small holes (interstices) in a metallic crystal are filled with smaller atoms. Forming this type of alloy is similar to pouring sand into a bucket of gravel. Even if the gravel is tightly packed, holes exist between the pieces. The sand does not replace any of the gravel but fills in the spaces. The best-known interstitial alloy is carbon steel. Holes in the iron crystal are filled with carbon atoms, and the physical properties of iron are changed. Iron is relatively soft and malleable. However, the presence of carbon makes the solid harder, stronger, and less ductile than pure iron, increasing its uses.

Section 8.4 Assessment

40. What is a metallic bond?
41. Explain how conductivity of electricity and high melting point of metals are explained by metallic bonding.
42. What is an alloy?
43. How does a substitutional alloy differ from an interstitial alloy?
44. **Thinking Critically** In the laboratory, how could you determine if a solid has an ionic bond or a metallic bond?
45. **Formulating Models** Draw a model to represent the ductility of a metal using the electron sea model shown in **Figure 8-10**.



Making Ionic Compounds

Elements combine to form compounds. If energy is released as the compound is formed, the resulting product is more stable than the reacting elements. In this investigation you will react elements to form two compounds. You will test the compounds to determine several of their properties. Ionic compounds have properties that are different from those of other compounds. You will decide if the products you formed are ionic compounds.

Problem

What are the formulas and names of the products that are formed? Do the properties of these compounds classify them as having ionic bonds?

Objectives

- **Observe** evidence of a chemical reaction.
- **Acquire** and **analyze** information that will enable you to decide if a compound has an ionic bond.
- **Classify** the products as ionic or not ionic.

Materials

magnesium ribbon	crucible tongs
crucible	centigram balance
ring stand and ring	100-mL beaker
clay triangle	distilled water
laboratory burner	conductivity tester
stirring rod	

Safety Precautions



- Always wear safety glasses and a lab apron.
- Do not look directly at the burning magnesium. The intensity of the light can damage your eyes.
- Avoid handling heated materials until they have cooled.

Pre-Lab

1. Read the entire procedure. Identify the variables. List any conditions that must be kept constant.
2. Write the electron configuration of the magnesium atom.
 - a. Based on this configuration, will magnesium lose or gain electrons to become a magnesium ion?
 - b. Write the electron configuration of the magnesium ion.
 - c. The magnesium ion has an electron configuration like that of which noble gas?
3. Repeat question 2 for oxygen and nitrogen.
4. Prepare your data table.
5. In your data table, which mass values will be measured directly? Which mass values will be calculated?
6. Explain what must be done to calculate each mass value that is not measured directly.

Mass Data	
Material(s)	Mass (g)
Empty crucible	
Crucible and Mg ribbon before heating	
Magnesium ribbon	
Crucible and magnesium products after heating	
Magnesium products	

Procedure

1. Arrange the ring on the ring stand so that it is about 7 cm above the top of the Bunsen burner. Place the clay triangle on the ring.
2. Measure the mass of the clean, dry crucible, and record the mass in the data table.
3. Roll 25 cm of magnesium ribbon into a loose ball. Place it in the crucible. Measure the mass of the magnesium and crucible and record this mass in the data table.
4. Place the crucible on the clay ring. Heat the crucible with a hot flame, being careful to position the crucible near the top of the flame.



5. When the magnesium metal ignites and begins to burn with a bright white light, immediately turn off the laboratory burner. **CAUTION:** Do not look directly at the burning magnesium. After the magnesium product and crucible have cooled, measure their mass and record it in the data table.
6. Place the dry solid product in a small beaker for further testing.
7. Add 10 mL of distilled water to the dry magnesium product in the beaker and stir. Check the mixture with a conductivity checker, and record your results.

Cleanup and Disposal

1. Wash out the crucible with water.
2. Dispose of the product as directed by your teacher.
3. Return all lab equipment to its proper place.

Analyze and Conclude

1. **Analyzing Data** Use the masses in the table to calculate the mass of the magnesium ribbon and the mass of the magnesium product. Record these masses in the table.
2. **Classifying** What kind of energy was released by the reaction? What can you conclude about the product of this reaction?
3. **Using Numbers** How do you know that the magnesium metal reacts with certain components of the air?
4. **Predicting** Magnesium reacts with both oxygen and nitrogen from the air at the high temperature of the crucible. Predict the binary formulas for both products. Write the names of these two compounds.
5. **Analyzing and Concluding** The product formed from magnesium and oxygen is white, and the product formed from magnesium and nitrogen is yellow. From your observations, which compound makes up most of the product?
6. **Analyzing and Concluding** Did the magnesium compounds and water conduct an electric current? Do the results indicate whether or not the compounds are ionic?
7. **Error Analysis** If the magnesium lost mass instead of gaining mass, what do you think was a possible source of the error?

Real-World Chemistry

1. The magnesium ion plays an important role in a person's biochemistry. Research the role of this electrolyte in your physical and mental health. Is magnesium listed as a component in a multi-vitamin and mineral tablet?
2. Research the use of $\text{Mg}(\text{OH})_2$ in everyday products. What is $\text{Mg}(\text{OH})_2$ commonly called in over-the-counter drugs?

Everyday Chemistry

Colors of Gems

Have you ever wondered what produces the gorgeous colors in a stained-glass window or in the rubies, emeralds, and sapphires mounted on a ring? Compounds of transition elements are responsible for creating the entire spectrum of colors.

Transition elements color gems and glass

Transition elements have many important uses, but one that is often overlooked is their role in giving colors to gemstones and glass. Although not all compounds of transition elements are colored, most inorganic colored compounds contain a transition element such as chromium, iron, cobalt, copper, manganese, nickel, cadmium, titanium, gold, or vanadium. The color of a compound is determined by the identity of the metal, its oxidation number, and the negative ion combined with it.

Impurities give gemstones their color

Crystals have fascinating properties. A clear, colorless quartz crystal is pure silicon dioxide (SiO_2). But a crystal that is colorless in its pure form may exist as a variety of colored gemstones when tiny amounts of transition element compounds, usually oxides, are present. Amethyst (purple), citrine (yellow-brown), and rose quartz (pink) are quartz crystals with transition element impurities scattered throughout. Blue sapphires are composed of aluminum oxide (Al_2O_3) with the impurities iron(II) oxide (FeO) and titanium(IV) oxide (TiO_2). If trace amounts of chromium(III) oxide (Cr_2O_3) are

present in the Al_2O_3 , the resulting gem is a red ruby. A second kind of gemstone is one composed entirely of a colored compound. Most are transition element compounds, such as rose-red rhodochrosite (MnCO_3), black-grey hematite (Fe_2O_3), or green malachite ($\text{CuCO}_3 \cdot \text{Cu(OH)}_2$).

How metal ions interact with light to produce color

Why does the presence of Cr_2O_3 in Al_2O_3 make a ruby red? The Cr^{3+} ion absorbs yellow-green colors from white light striking the ruby, and the remaining red-blue light is transmitted, resulting in a deep red color. This same process occurs in all gems. Trace impurities absorb certain colors of light from

white light striking or passing through the stone. The remaining colors of light that are reflected or transmitted produce the color of the gem.

Adding transition elements to molten glass for color

Glass is colored by adding transition element compounds to the glass while it is molten. This process is used for stained glass, glass used in glass blowing, and even glass in the form of ceramic glazes. Most of the coloring agents are oxides. When oxides of

copper or cobalt are added to molten glass, the glass is blue; oxides of manganese produce purple glass; iron oxides, green; gold oxides, deep ruby red; copper or selenium oxides, red; and antimony oxides, yellow. Some coloring compounds are not oxides. Chromates, for example, produce green glass, and iron sulfide gives a brown color.



Testing Your Knowledge

- 1. Applying** Explain why iron(III) sulfate is yellow, iron(II) thiocyanate is green, and iron(III) thiocyanate is red.
- 2. Acquiring Information** Find out what impurities give amethyst, rose quartz, and citrine their colors.
- 3. Comparing and Contrasting** Conduct research to find the similarities and differences between synthetic and natural gemstones.

Summary

8.1 Forming Chemical Bonds

- A chemical bond is the force that holds two atoms together.
- Atoms that form ions gain or lose valence electrons to achieve the same electron arrangement as that of a noble gas, which is a stable configuration. This noble gas configuration involves a complete outer electron energy level, which usually consists of eight valence electrons.
- A positive ion, or cation, forms when valence electrons are removed and a stable electron configuration is obtained.
- A negative ion, or anion, forms when valence electrons are added to the outer energy level, giving the ion a stable electron configuration.

8.2 The Formation and Nature of Ionic Bonds

- An ionic bond forms when anions and cations close to each other attract, forming a tightly packed geometric crystal lattice.
- Lattice energy is needed to break the force of attraction between oppositely charged ions arranged in a crystal lattice.
- The physical properties of ionic solids, such as melting point, boiling point, hardness, and the ability to conduct electricity in the molten state and as an aqueous solution, are related to the strength of the ionic bonds and the presence of ions.
- An ionic compound is an electrolyte because it conducts an electric current when it is liquid or in aqueous solution.

8.3 Names and Formulas for Ionic Compounds

- Subscripts in an ionic compound indicate the ratio

of cations and anions needed to form electrically neutral compounds. The formula unit represents the ratio of these ions in the crystal lattice.

- If the element that forms the cation has more than one possible oxidation number, Roman numerals are used to indicate the oxidation number present for that element in the compound.
- Ions formed from only one atom are monatomic ions. The charge on a monatomic ion is its oxidation number, or oxidation state.
- Polyatomic ions are two or more atoms bonded together that act as a single unit with a net charge. Many polyatomic ions are oxyanions, containing an atom, usually a nonmetal, and oxygen atoms.
- In a chemical formula, polyatomic ions are placed inside parentheses when using a subscript.
- Ionic compounds are named by the name of the cation followed by the name of the anion.

8.4 Metallic Bonds and Properties of Metals

- Metallic bonds are formed when metal cations attract free valence electrons. A “sea” of electrons moves throughout the entire metallic crystal, producing this attraction.
- The electrons involved in metallic bonding are called delocalized electrons because they are free to move throughout the metal and are not attached to a particular atom.
- The electron sea model can explain the melting point, boiling point, malleability, conductivity, and ductility of metallic solids.
- Metal alloys are formed when a metal is mixed with one or more other elements. The two common types of alloys are substitutional and interstitial.

Vocabulary

- | | | |
|----------------------------------|-------------------------------|-----------------------------|
| • alloy (p. 230) | • electrolyte (p. 218) | • metallic bond (p. 228) |
| • anion (p. 214) | • electron sea model (p. 228) | • monatomic ion (p. 221) |
| • cation (p. 212) | • formula unit (p. 221) | • oxidation number (p. 222) |
| • chemical bond (p. 211) | • ionic bond (p. 215) | • oxyanion (p. 225) |
| • delocalized electrons (p. 228) | • lattice energy (p. 219) | • polyatomic ion (p. 224) |

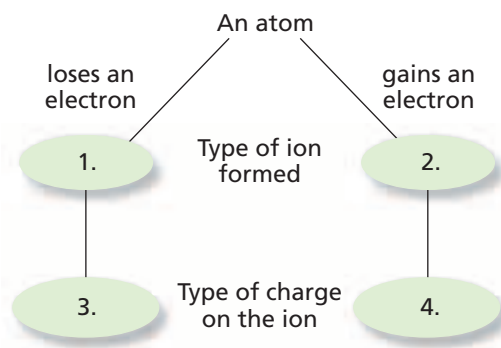




Go to the Chemistry Web site at chemistrymc.com for additional Chapter 8 Assessment.

Concept Mapping

46. Complete the concept map, showing what type of ion is formed in each case and what type of charge the ion has.



Mastering Concepts

47. When do chemical bonds form? (8.1)
48. Why do positive ions and negative ions form? (8.1)
49. Why are halogens and alkali metals likely to form ions? Explain your answer. (8.1)
50. Discuss the importance of electron affinity and ionization energy in the formation of ions. (8.1)
51. Discuss the formation of ionic bonds. (8.2)
52. Briefly discuss three physical properties of ionic solids that are linked to ionic bonds. (8.2)
53. What does the term *electrically neutral* mean when discussing ionic compounds? (8.2)
54. What information is needed to write a correct chemical formula to represent an ionic compound? (8.3)
55. When are subscripts used in formulas for ionic compounds? (8.3)
56. Discuss how an ionic compound is named. (8.3)
57. Describe a metallic bond. (8.4)
58. Briefly explain how malleability and ductility of metals are explained by metallic bonding. (8.4)
59. Compare and contrast the two types of metal alloys. (8.4)

Mastering Problems

Ion Formation (8.1)

60. Explain why noble gases are not likely to form chemical bonds.
61. Give the number of valence electrons in an atom of each of the following:
 - a. cesium
 - b. rubidium
 - c. gallium
 - d. zinc
 - e. strontium
62. Discuss the formation of the barium ion.
63. Explain how an anion of nitrogen forms.
64. The more reactive an atom, the higher its potential energy. Which atom has higher potential energy, neon or fluorine? Explain.
65. Predict the reactivity of the following atoms based on their electron configurations.
 - a. potassium
 - b. fluorine
 - c. neon
66. Discuss the formation of the iron ion that has a 3+ oxidation number.

Ionic Bonds and Ionic Compounds (8.2)

67. Determine the ratio of cations to anions for the following ionic compounds.
 - a. potassium chloride, a salt substitute
 - b. calcium fluoride, used in the steel industry
 - c. aluminum oxide, known as corundum in the crystalline form
 - d. calcium oxide, used to remove sulfur dioxide from power plant exhaust
 - e. strontium chloride, used in fireworks
68. Using orbital notation, diagram the formation of an ionic bond between aluminum and fluorine.
69. Using electron configurations, diagram the formation of an ionic bond between barium and nitrogen.
70. Discuss the formation of an ionic bond between zinc and oxygen.
71. Under certain conditions, ionic compounds conduct an electric current. Describe these conditions and explain why ionic compounds are not always used as conductors.
72. Which of the following compounds are not likely to occur: CaKr, Na₂S, BaCl₃, MgF? Explain your choices.
73. Using Table 8-2, determine which of the following ionic compounds will have the highest melting point: MgO, KI, or AgCl. Explain your answer.

Formulas and Names for Ionic Compounds (8.3)

- 74.** Give the formula for each of the following ionic compounds.
- calcium iodide
 - silver(I) bromide
 - copper(II) chloride
 - potassium periodate
 - silver(I) acetate
- 75.** Name each of the following ionic compounds.
- K_2O
 - $CaCl_2$
 - Mg_3N_2
 - $NaClO$
 - KNO_3
- 76.** Complete **Table 8-9** by placing the symbols, formulas, and names in the blanks.

Table 8-9

Identifying Ionic Compounds			
Cation	Anion	Name	Formula
		ammonium sulfate	
			PbF_2
		lithium bromide	
			Na_2CO_3
Mg^{2+}	PO_4^{3-}		

- 77.** Chromium, a transition metal, forms both the Cr^{2+} and Cr^{3+} ions. Write the formulas for the ionic compounds formed when each of these ions react with
- fluorine
 - oxygen
- 78.** Which of the following are correct formulas for ionic compounds? For those that are not correct, give the correct formula and justify your answer.
- $AlCl$
 - Na_3SO_4
 - $MgCO_3$
 - $BaOH_2$
 - Fe_2O
- 79.** Write the formulas for all of the ionic compounds that can be formed by combining each of the cations with each of the anions listed below. Name each compound formed.

Table 8-10

Cations	Anions
K^+	SO_3^{2-}
NH_4^+	I^-
Fe^{3+}	NO_3^-

Metals and Metallic Bonds (8.4)

- 80.** How is a metallic bond different from an ionic bond?
- 81.** Briefly explain why silver is a good conductor of electricity.
- 82.** Briefly explain why iron is used in making the structures of many buildings.
- 83.** The melting point of beryllium is $1287^\circ C$, while that of lithium is $180^\circ C$. Account for the large difference in values.
- 84.** Describe the difference between the metal alloy sterling silver and carbon steel in terms of the types of alloys involved.

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 85.** Give the number of valence electrons for atoms of oxygen, sulfur, arsenic, phosphorus, and bromine.
- 86.** Explain why calcium can form a Ca^{2+} ion but not a Ca^{3+} ion.
- 87.** Which of the following ionic compounds would have the most negative lattice energy: $NaCl$, KCl , or $MgCl_2$? Explain your answer.
- 88.** Give the formula for each of the following ionic compounds.
- sodium sulfide
 - iron(III) chloride
 - sodium sulfate
 - calcium phosphate
 - zinc nitrate
- 89.** Cobalt, a transition metal, forms both the Co^{2+} and Co^{3+} ions. Write the correct formulas and give the name for the oxides formed by the two different ions.
- 90.** Briefly explain why gold can be used as both a conductor in electronic devices and in jewelry.
- 91.** Discuss the formation of the nickel ion with a 2+ oxidation number.
- 92.** Using electron-dot structure, diagram the formation of an ionic bond between potassium and iodine.
- 93.** Magnesium forms both an oxide and a nitride when burned in air. Discuss the formation of magnesium oxide and magnesium nitride when magnesium atoms react with oxygen and nitrogen atoms.
- 94.** An external force easily deforms sodium metal, while sodium chloride shatters when the same amount of force is applied. Why do these two solids behave so differently?



95. Name each of the following ionic compounds.

- a. CaO
- b. BaS
- c. AlPO_4
- d. $\text{Ba}(\text{OH})_2$
- e. $\text{Sr}(\text{NO}_3)_2$

96. Write the formulas for all of the ionic compounds that can be formed by combining each of the cations with each of the anions listed below. Name each compound formed.

Table 8-11

Cations	Anions
Ba^{2+}	$\text{S}_2\text{O}_3^{2-}$
Cu^+	Br^-
Al^{3+}	NO_2^-

Thinking Critically

- 97. Concept Mapping** Design a concept map to explain the physical properties of both ionic compounds and metallic solids.
- 98. Predicting** Predict which solid in each of the following will have the higher melting point. Explain your answer.
- a. NaCl or CsCl
 - b. Ag or Cu
 - c. Na_2O or MgO
- 99. Comparing and Contrasting** Compare and contrast cations and anions.
- 100. Observing and Inferring** From the following incorrect formulas and formula names, identify the mistakes and design a flow chart to prevent the mistakes.
- a. copper acetate
 - b. Mg_2O_2
 - c. Pb_2O_5
 - d. disodium oxide
 - e. Al_2SO_4
- 101. Hypothesizing** Look at the locations of potassium and calcium on the periodic table. Form a hypothesis as to why the melting point of calcium is considerably higher than the melting point of potassium.
- 102. Drawing a Conclusion** Explain why the term *delocalized* is an appropriate term for the electrons involved in metallic bonding.
- 103. Applying Concepts** All uncharged atoms have valence electrons. Explain why elements such as iodine and sulfur don't have metallic bonds.
- 104. Drawing a Conclusion** Explain why lattice energy is a negative quantity.

Writing in Chemistry

- 105.** Many researchers believe that free radicals are responsible for the effects of aging and cancer. Research free radicals and write about the cause and what can be done to prevent free radicals.
- 106.** Crystals of ionic compounds can be easily grown in the laboratory setting. Research the growth of crystals and try to grow one crystal in the laboratory.

Cumulative Review

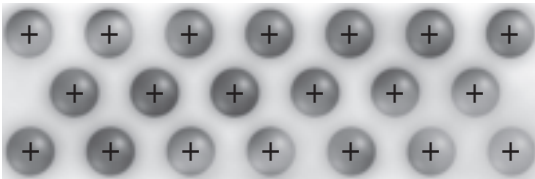
Refresh your understanding of previous chapters by answering the following.

- 107.** You are given a liquid of unknown density. The mass of a graduated cylinder containing 2.00 mL of the liquid is 34.68 g. The mass of the empty graduated cylinder is 30.00 g. What is the density of the liquid? (Chapter 2)
- 108.** A mercury atom drops from 1.413×10^{-18} J to 1.069×10^{-18} J. (Chapter 5)
- a. What is the energy of the photon emitted by the mercury atom?
 - b. What is the frequency of the photon emitted by the mercury atom?
 - c. What is the wavelength of the photon emitted by the mercury atom?
- 109.** Which element has the greater ionization energy, chlorine or carbon? (Chapter 6)
- 110.** Compare and contrast the way metals and nonmetals form ions and explain why they are different. (Chapter 6)
- 111.** What are transition elements? (Chapter 6)
- 112.** Write the symbol and name of the element that fits each description. (Chapter 6)
- a. the second-lightest of the halogens
 - b. the metalloid with the lowest period number
 - c. the only group 6A element that is a gas at room temperature
 - d. the heaviest of the noble gases
 - e. the group 5A nonmetal that is a solid at room temperature
- 113.** Which group 4A element is (Chapter 7)
- a. a metalloid that occurs in sand?
 - b. a nonmetal?
 - c. used in electrodes in car batteries?
 - d. a component in many alloys?

STANDARDIZED TEST PRACTICE

CHAPTER 8

Use these questions and the test-taking tip to prepare for your standardized test.

- Which of the following is NOT true of the Sc^{3+} ion?
 - It has the same electron configuration as Ar.
 - It is a scandium ion with three positive charges.
 - It is considered to be a different element than a neutral Sc atom.
 - It was formed by the removal of the valence electrons of Sc.
- Of the salts below, it would require the most energy to break the ionic bonds in
 - BaCl_2 .
 - LiF .
 - NaBr .
 - KI .
- What is the correct chemical formula for the ionic compound formed by the calcium ion (Ca^{2+}) and the acetate ion ($\text{C}_2\text{H}_3\text{O}_2^-$)?
 - $\text{CaC}_2\text{H}_3\text{O}_2$
 - $\text{CaC}_4\text{H}_6\text{O}_8$
 - $(\text{Ca})_2\text{C}_2\text{H}_3\text{O}_2$
 - $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$
- 

The model above has been proposed to explain why

 - metals are shiny, reflective substances.
 - metals are excellent conductors of heat and electricity.
 - ionic compounds are malleable compounds.
 - ionic compounds are good conductors of electricity.
- Yttrium, a metallic element with atomic number 39, will form
 - positive ions.
 - negative ions.
 - both positive and negative ions.
 - no ions at all.
- The high strength of its ionic bonds results in all of the following properties of NaCl EXCEPT
 - hard crystals.
 - high boiling point.
 - high melting point.
 - low solubility.

Interpreting Tables Use the table below to answer questions 7–10.

Properties of Some Ionic Compounds			
Compound	Lattice Energy (kJ/mol)	Melting Point ($^{\circ}\text{C}$)	Color
Ag_2Se	-2686	?	gray
AlPO_4	?	1460	white
FeI_2	-2439	?	reddish-purple
RbClO_4	?	281	white

- What is the correct name of the compound with the formula RbClO_4 ?
 - rubidium chlorine oxide
 - rubidium chloride tetroxide
 - rubidium perchlorate
 - rubidium chlorate
- Rank the compounds in order of increasing melting point.
 - Ag_2Se , AlPO_4 , FeI_2 , RbClO_4
 - RbClO_4 , FeI_2 , Ag_2Se , AlPO_4
 - AlPO_4 , Ag_2Se , FeI_2 , RbClO_4
 - RbClO_4 , AlPO_4 , Ag_2Se , FeI_2
- Which compound is expected to have the strongest attraction between its ions?
 - Ag_2Se
 - AlPO_4
 - FeI_2
 - RbClO_4
- What is the charge on the anion in AlPO_4 ?
 - 2+
 - 3+
 - 2-
 - 3-

TEST-TAKING TIP

Work Weak Muscles; Maintain Strong Ones

If you're preparing for a standardized test that covers many topics, it's sometimes difficult to focus on all the topics that require your attention. Ask yourself "What's my strongest area?" and "What's my weakest area?" Focus most of your energy on your weaker area and review your stronger topics less frequently.

