3.5 Ionic Compounds

Ionic compounds are substances composed of ions attracted to each other by ionic bonds. Let's consider how they play a part in a "typical" family's Fourth of July.

Before the family leaves the house to go to the holiday picnic, the kids are sent off to brush their teeth and change into clean clothes. Their toothpaste contains sodium fluoride, a common cavity-fighting ionic compound. The white shirts in their red, white, and blue outfits were bleached with the ionic compound sodium hypochlorite, the stains on their red pants were removed by potassium oxalate, and dyes were fixed to their blue hats by aluminum nitrate.

While the kids are getting ready, dad and mom get the picnic dinner together. The hot dogs they are packing are cured with the ionic compound sodium nitrite, and the buns contain calcium acetate as a mold inhibitor and calcium iodate as a dough conditioner. The soft drinks have potassium hydrogen carbonate to help trap the bubbles, the mineral water contains magnesium sulfate, and the glass for the bottles was made with a variety of ionic compounds. Because it will be dark before they get home, Mom packs a flashlight as well. Its rechargeable batteries contain the ionic compounds cadmium hydroxide and nickel hydroxide.

Ionic compounds play an important role in creating the excitement of a Fourth of July celebration.

When our family gets to the park, they find themselves a place on the lawn, which was fertilized with a mixture of ionic compounds, including iron(II) sulfate. They eat their dinner and play in the park until it's time for the fireworks. The safety matches used to light the rockets contain barium chromate, and ionic compounds in the fireworks provide the colors: red from strontium chlorate, white from magnesium nitrate, and blue from copper(II) chloride.

Cations and Anions

Remember the sodium fluoride in the kids' toothpaste? It could be made from the reaction of sodium metal with the nonmetallic atoms in fluorine gas. As you discovered in Section 3.2, metallic atoms hold some of their electrons relatively loosely, and as a result, they tend to lose electrons and form cations. In contrast, nonmetallic atoms attract electrons more strongly than metallic atoms, and so nonmetals tend to gain electrons and form anions. Thus, when a metallic element and a nonmetallic element combine, the nonmetallic atoms often pull one or more electrons far enough away from the metallic atoms to form ions. The positive cations and the negative anions then attract each other to form ionic bonds. In the formation of sodium fluoride from sodium metal and fluorine gas, each sodium atom donates one electron to a fluorine atom to form a Na^+ cation and an F^- anion. The F^- anions in toothpaste bind to the surface of your teeth, making them better able to resist tooth decay. This section provides you with more information about other cations and anions, including how to predict their charges and how to convert between their names and formulas.

Objective 29

Sodium fluoride is added to toothpaste to help fight tooth decay.

Predicting Ionic Charges

It is useful to be able to predict the charges the atoms of each element are most likely to attain when they form ions. Because the periodic table can be used to predict ionic charges, it is a good idea to have one in front of you when you study this section.

We discovered in Chapter 2 that the atoms of the noble gases found in nature are uncombined with other atoms. The fact that the noble gas atoms do not gain, lose, or share their electrons suggests there must be something especially stable about having 2 (helium, He), 10 (neon, Ne), 18 (argon, Ar), 36 (krypton, Kr), 54 (xenon, Xe), or 86 (radon, Rn) electrons. This stability is reflected in the fact that nonmetallic atoms form anions in order to get the same number of electrons as the nearest noble gas.

Objective 30

All of the halogens (group 17) have one less electron than the nearest noble gas. When halogen atoms combine with metallic atoms, they tend to gain one electron each and form −1 ions (Figure 3.15). For example, uncharged chlorine atoms have 17 protons and 17 electrons. If a chlorine atom gains one electron, it will have 18 electrons like an uncharged argon atom. With a −18 charge from the electrons and a +17 charge from the protons, the resulting chlorine ion has a −1 charge. The symbol for this anion is Cl^-. Notice that the negative charge is indicated with a "−" not "−1" or " 1−".

$$\underset{17p/17e^-}{Cl} \quad + \quad 1e^- \quad \rightarrow \quad \underset{17p/18e^-}{Cl^-}$$

Objective 30

The nonmetallic atoms in group 16 (oxygen, O, sulfur, S, and selenium, Se) have two fewer electrons than the nearest noble gas. When atoms of these elements combine with metallic atoms, they tend to gain two electrons and form −2 ions (Figure 3.15). For example, oxygen, in group 16, has atoms with eight protons and eight electrons.

Each oxygen atom can gain two electrons to achieve ten, the same number as its nearest noble gas, neon. The symbol for this anion is O^{2-}. Notice that the charge is indicated with "2–" not "–2".

$$\underset{8p/8e^-}{O} + 2e^- \rightarrow \underset{8p/10e^-}{O^{2-}}$$

Objective 30

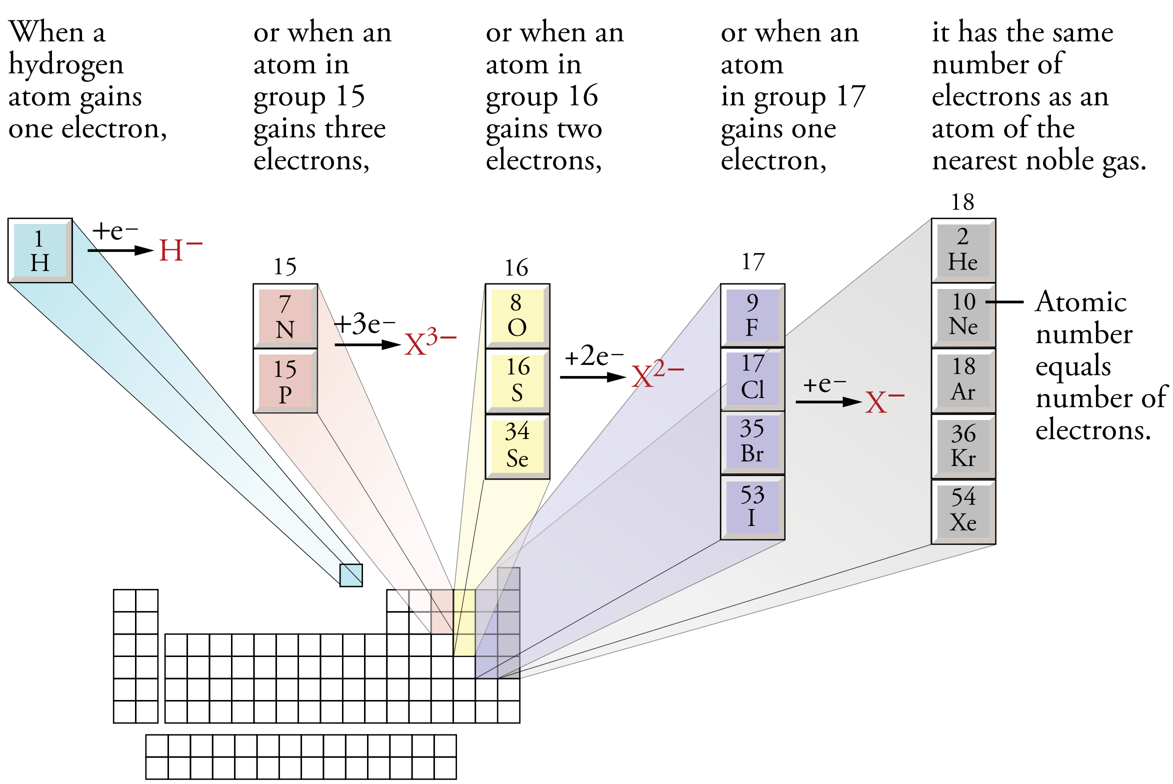
Nitrogen, N, and phosphorus, P, have three fewer electrons than the nearest noble gas. When atoms of these elements combine with metallic atoms, they tend to gain three electrons and form −3 ions (Figure 3.15). For example, nitrogen atoms have seven protons and seven electrons. Each nitrogen atom can gain three electrons to achieve ten, like neon, forming a N^{3-} anion.

$$\underset{7p/7e^-}{N} + 3e^- \rightarrow \underset{7p/10e^-}{N^{3-}}$$

Hydrogen has one less electron than helium, so when it combines with metallic atoms, it forms a −1 ion, H^- (Figure 4.15). Anions like H^-, Cl^-, O^{2-}, and N^{3-}, which contain single atoms with a negative charge, are called **monatomic anions**.

Figure 3.15
The Making of an Anion

Objective 30



Objective 30

Some metallic atoms lose enough electrons to create a cation that has the same number of electrons as the nearest smaller noble gas. For example, the alkali metals in group 1 all have one more electron than the nearest noble gas. When they react with nonmetallic atoms, they lose one electron and form +1 ions (Figure 3.16). For example, sodium has atoms with 11 protons and 11 electrons. If an atom of sodium loses one electron, it will have ten electrons like uncharged neon. With a −10 charge from the electrons and a +11 charge from the protons, the sodium ions have a +1 overall charge. The symbol for this cation is Na^+. Notice that the charge is indicated with a "+" instead of "+1" or "1+".

$$\begin{array}{cccccc} Na & \rightarrow & Na^+ & + & 1e^- \\ 11p/11e^- & & 11p/10e^- & & \end{array}$$

Objective 30

The alkaline earth metals in group 2 all have two more electrons than the nearest noble gas. When they react with nonmetallic atoms, they tend to lose two electrons and form +2 ions (Figure 3.16). For example, calcium has atoms with 20 protons and 20 electrons. Each calcium atom can lose two electrons to achieve 18, the same number as its nearest noble gas, argon. The symbol for this cation is Ca^{2+}. Note that the charge is indicated with "2+" not "+2".

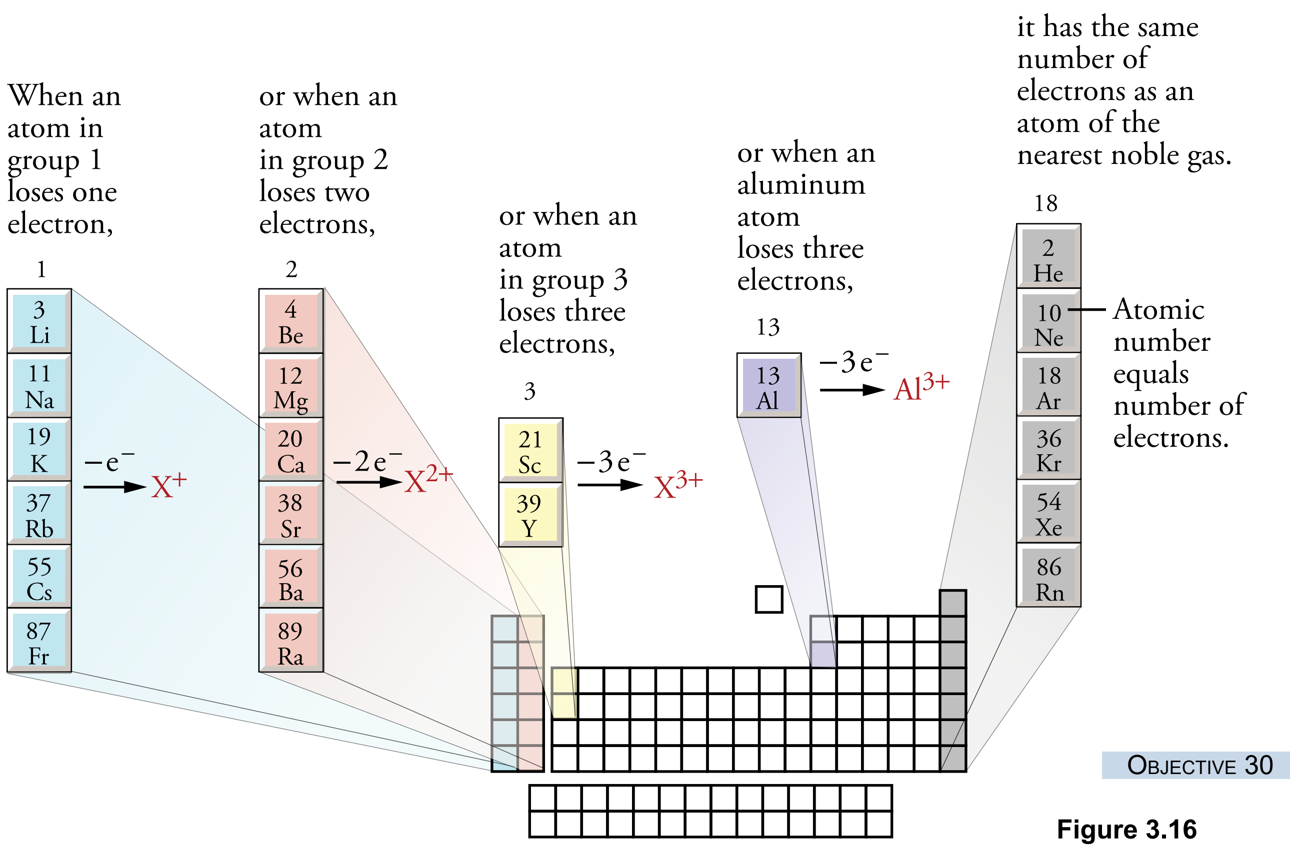
$$\begin{array}{cccccc} Ca & \rightarrow & Ca^{2+} & + & 2e^- \\ 20p/20e^- & & 20p/18e^- & & \end{array}$$

Objective 30

Aluminum atoms and the atoms of the group 3 metals have three more electrons than the nearest noble gas. When they react with nonmetallic atoms, they tend to lose three electrons and form +3 ions (Figure 3.16). For example, uncharged aluminum atoms have 13 protons and 13 electrons. Each aluminum atom can lose three electrons to achieve ten, like neon, forming an Al^{3+} cation.

$$\begin{array}{cccccc} Al & \rightarrow & Al^{3+} & + & 3e^- \\ 13p/13e^- & & 13p/10e^- & & \end{array}$$

Cations like Na^+, Ca^{2+}, and Al^{3+}, which are single atoms with a positive charge, are called **monatomic cations**.



Objective 30

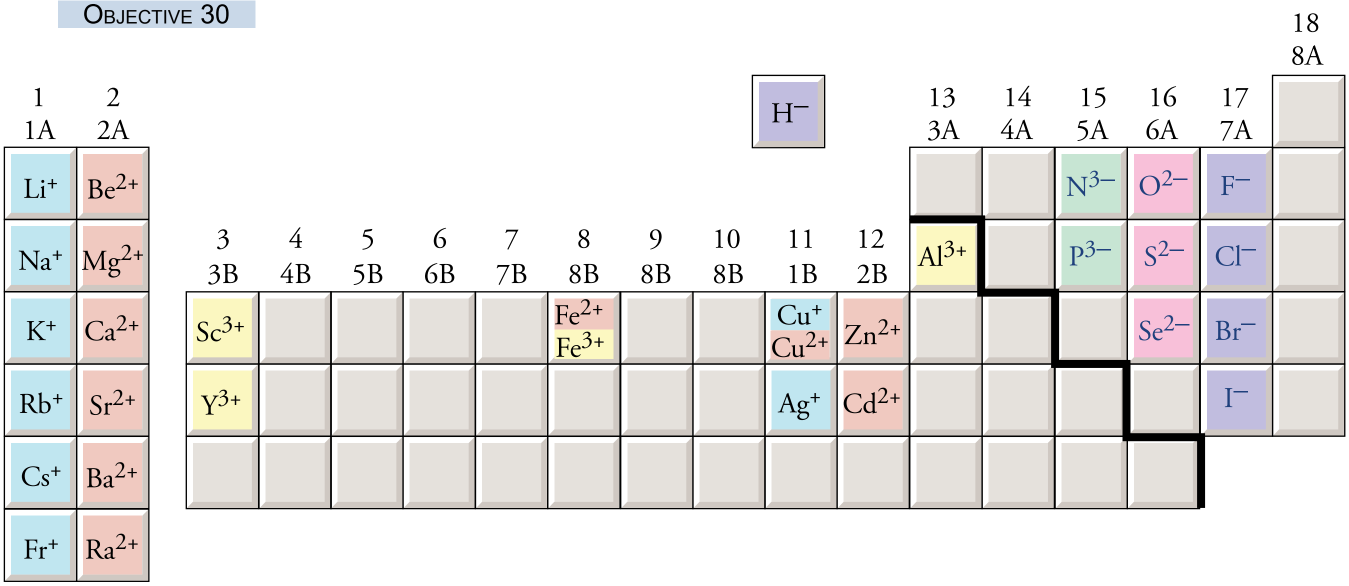
Figure 3.16
The Making of a Cation

The metallic elements in groups other than 1, 2, or 3 also lose electrons to form cations, but they do so in less easily predicted ways. It will be useful to memorize some of the charges for these metals. Ask your instructor which ones you will be expected to know. To answer the questions in this text, you will need to know that iron atoms form both Fe^{2+} and Fe^{3+}, copper atoms form Cu^{+} and Cu^{2+}, zinc atoms form Zn^{2+}, cadmium atoms form Cd^{2+}, and silver atoms form Ag^{+}. Figure 3.17 summarizes the charges of the ions that you should know at this stage.

OBJECTIVE 30

Figure 3.17
Common Monatomic Ions

OBJECTIVE 30



Naming Monatomic Anions and Cations

OBJECTIVE 31

The monatomic anions are named by adding -ide to the root of the name of the nonmetal that forms the anion. For example, N^{3-} is the nitride ion. The roots of the nonmetallic atoms are listed in Table 3.4, and the names of the anions are displayed in Table 3.5.

OBJECTIVE 31

Table 3.5
Names of the Monatomic Anions

Anion	Name	Anion	Name	Anion	Name
N^{3-}	nitride	O^{2-}	oxide	H^{-}	hydride
P^{3-}	phosphide	S^{2-}	sulfide	F^{-}	fluoride
		Se^{2-}	selenide	Cl^{-}	chloride
				Br^{-}	bromide
				I^{-}	iodide

The names of monatomic cations always start with the name of the metal, sometimes followed by a Roman numeral to indicate the charge of the ion. For example, Cu^{+} is copper(I), and Cu^{2+} is copper(II). The Roman numeral in each name represents the charge on the ion and allows us to distinguish between more than one possible charge. Notice that there is no space between the end of the name of the metal and the parentheses with the Roman numeral.

If the atoms of an element always have the same charge, the Roman numeral is unnecessary (and considered to be incorrect). For example, all cations formed from sodium atoms have a +1 charge, so Na^{+} is named sodium ion, without the Roman numeral for the charge. The following elements have only one possible charge, so it would be incorrect to put a Roman numeral after their name.

Objective 31

- The alkali metals in group 1 are always +1 when they form cations.
- The alkaline earth metals in group 2 are always +2 when they form cations.
- Aluminum and the elements in group 3 are always +3 when they form cations.
- Zinc and cadmium always form +2 cations.

Although silver can form both +1 and +2 cations, the +2 is so rare that we usually name Ag^{+} as *silver ion*, not *silver(I) ion.* Ag^{2+} is named silver(II) ion.

We will assume that all of the metallic elements other than those mentioned above can have more than one charge, so their cation names will include a Roman numeral. For example, Mn^{2+} is named manganese(II). We know to put the Roman numeral in the name because manganese is not on our list of metals with only one charge.

Example 3.6 - Naming Monatomic Ions

Objective 31

Write names that correspond to the following formulas for monatomic ions: (a) Ba^{2+}, (b) S^{2-}, and (c) Cr^{3+}.

Solution

a. Because barium is in group 2, the only possible charge is +2. When there is only one possible charge, metallic ions are named with the name of the metal. Therefore, Ba^{2+} is **barium ion.**
b. Monatomic anions are named with the root of the nonmetal and -ide, so S^{2-} is **sulfide ion.**
c. Because chromium is not on our list of metals with only one possible charge, we need to indicate the charge with a Roman numeral. Therefore, Cr^{3+} is **chromium(III) ion**.

Exercise 3.6 - Naming Monatomic Ions

Objective 31

Write names that correspond to the following formulas for monatomic ions: (a) Mg^{2+}, (b) F^{-}, and (c) Sn^{2+}.

Example 3.7 - Formulas for Monatomic Ions

Objective 31

There are two tutorials on the textbook's Web site that provide practice converting between names and formulas of cations and anions.

Write formulas that correspond to the following names for monatomic ions: (a) phosphide ion, (b) lithium ion, and (c) cobalt(II) ion.

Solution

a. We know this is a monatomic anion because it has the form, *(nonmetal root) ide*. Phosphorus atoms gain three electrons to get 18 electrons like the noble gas argon, Ar. Phosphide ion is $\mathbf{P^{3-}}$.
b. Lithium atoms lose one electron to get two electrons, like the noble gas helium. Lithium ion is $\mathbf{Li^{+}}$.
c. The Roman numeral indicates that the cobalt ion has a +2 charge. Note that we would not have been able to determine this from cobalt's position on the periodic table. Cobalt(II) is $\mathbf{Co^{2+}}$.

Exercise 3.7 - Formulas for Monatomic Ions

Objective 31

Write formulas that correspond to the following names for monatomic ions: (a) bromide ion, (b) aluminum ion, and (c) gold(I) ion.

Several of the monatomic cations play important roles in our bodies. For example, we need calcium ions in our diet for making bones and teeth. Iron(II) ions are found in hemoglobin molecules in red blood cells that carry oxygen from our lungs to the tissues of our bodies. Potassium, sodium, and chloride ions play a crucial role in the transfer of information between nerve cells. Enzymes (chemicals in the body that increase the speed of chemical reactions) often contain metallic cations, such as manganese(II) ions, iron(III) ions, copper(II) ions, and zinc ions. For example, Zn^{2+} ions are in the center of the enzyme alcohol dehydrogenase, which is the enzyme in our livers that accelerates the breakdown of the ethanol consumed in alcoholic beverages.

Structure of Ionic Compounds

Objective 31

Figure 3.18 shows the solid structure of the ionic compound sodium chloride, NaCl. We have already seen that the particles that form the structure of ionic compounds are cations and anions, and the attractions that hold them together are ionic bonds. When atoms gain electrons and form anions, they get larger. When atoms lose electrons and form cations, they get significantly smaller. Thus the chloride ions are larger than the sodium ions. The ions take the arrangement that provides the greatest cation-anion attraction while minimizing the anion-anion and cation-cation repulsions. Each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions.

Any ionic compound that has the same arrangement of cations and anions as NaCl is said to have the sodium chloride crystal structure. The ionic compounds in this category include AgF, AgCl, AgBr, and the oxides and sulfides of the alkaline earth metals, such as MgO, CaS, etc. The sodium chloride crystal structure is just one of many different possible arrangements of ions in solid ionic compounds.

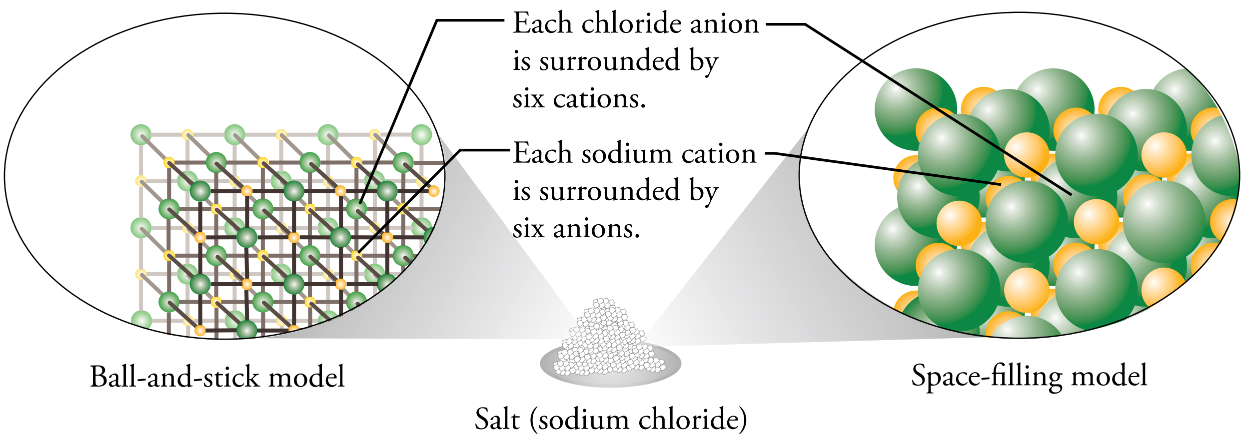


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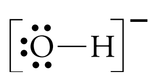
Figure 3.18
Sodium Chloride Structure

Polyatomic Ions

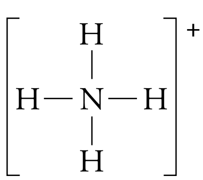
When an electric current is run through a purified saltwater solution (brine), hydrogen gas, chlorine gas, and an ionic compound called sodium hydroxide, NaOH, form. The sodium hydroxide, commonly called caustic soda or lye, is a very important compound that is used in paper production, vegetable oil refining, and to make many different compounds, such as soap and rayon. Like sodium chloride, NaCl, sodium hydroxide, NaOH, contains a cation and an anion, but unlike the monatomic Cl^- anion in NaCl, the hydroxide ion, OH^-, in NaOH is a **polyatomic ion**, a charged *collection* of atoms held together by covalent bonds. To show the charge, Lewis structures of polyatomic ions are often enclosed in brackets, with the charge indicated at the top right. The Lewis structure for hydroxide is



hydroxide

Note in the Lewis structure above that the oxygen atom does not have its most common bonding pattern, two bonds and two lone pairs. The gain or loss of electrons in the formation of polyatomic ions leads to one or more atoms in the ions having a different number of bonds and lone pairs than is predicted on Table 3.1.

The Lewis structure of the ammonium ion, NH_4^+, the only common polyatomic cation, is



ammonium ion

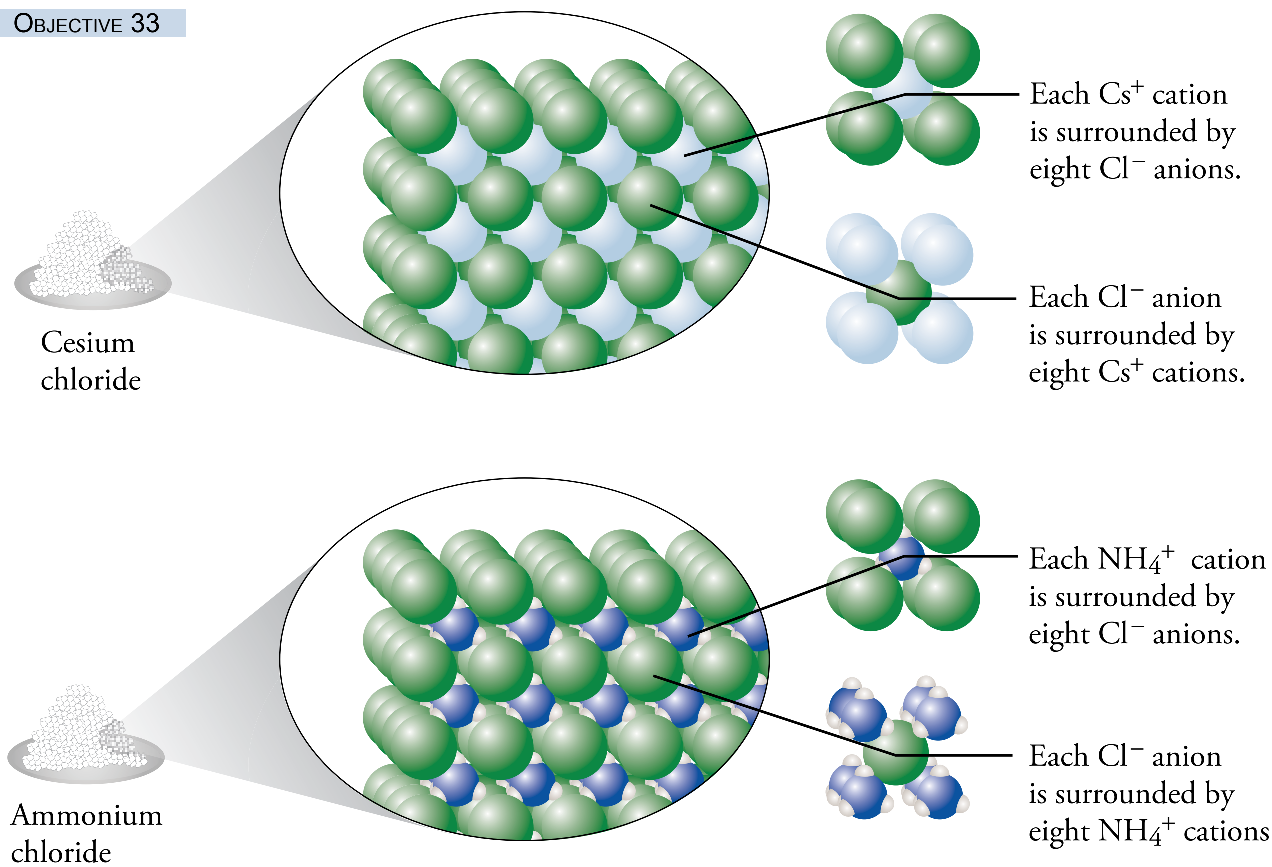
The ammonium ion can take the place of a monatomic cation in an ionic crystal structure. For example, the crystal structure of ammonium chloride, NH_4Cl, which is found in fertilizers, is very similar to the crystal structure of cesium chloride, CsCl, which is used in brewing, mineral waters, and to make fluorescent screens. In each structure, the chloride ions form a cubic arrangement with chloride ions at the corners of each

OBJECTIVE 33

cube. In cesium chloride, the cesium ions sit in the center of each cube, surrounded by eight chloride ions. Ammonium chloride has the same general structure as cesium chloride, with ammonium ions playing the same role in the NH_4Cl structure as cesium ions play in CsCl. The key idea is that because of its overall positive charge, the polyatomic ammonium ion acts like the monatomic cesium ion, Cs^+ (Figure 3.19).

Figure 3.19
Cesium Chloride Crystal Structure

Objective 33



There are many polyatomic anions that can take the place of monatomic anions. For example, zinc hydroxide, used as an absorbent in surgical dressings, has a similar structure to zinc chloride, which is used in embalming and taxidermist's fluids. The hydroxide ion, OH^-, plays the same role in the structure of $Zn(OH)_2$ as the chloride ion, Cl^-, plays in $ZnCl_2$. (Note that to show that there are two hydroxide ions for each zinc ion, the OH is in parentheses, with a subscript of 2.)

Objective 34

It is very useful to be able to convert between the names and formulas of the common polyatomic ions listed in Table 3.6. Check with your instructor to find out which of these you will be expected to know and whether there are others you should know as well.

Table 3.6
Common Polyatomic Ions

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Ion	Name	Ion	Name	Ion	Name
NH_4^+	ammonium	PO_4^{3-}	phosphate	SO_4^{2-}	sulfate
OH^-	hydroxide	NO_3^-	nitrate	$C_2H_3O_2^-$	acetate
CO_3^{2-}	carbonate				

Objective 35

Some polyatomic anions are formed by the attachment of one or more hydrogen atoms. In fact, it is common for hydrogen atoms to be transferred from one ion or molecule to another ion or molecule. When this happens, the hydrogen atom is usually transferred without its electron, as H^+. If an anion has a charge of −2 or −3, it can gain one or two H^+ ions and still retain a negative charge. For example, carbonate, CO_3^{2-}, can gain an H^+ ion to form HCO_3^-, which is found in baking soda. The sulfide ion, S^{2-}, can gain one H^+ ion to form HS^-. Phosphate, PO_4^{3-}, can gain one H^+ ion and form HPO_4^{2-}, or it can gain two H^+ ions to form $H_2PO_4^-$. Both HPO_4^{2-} and $H_2PO_4^-$ are found in flame retardants. These polyatomic ions are named with the word *hydrogen* in front of the name of the anion if there is one H^+ ion attached and *dihydrogen* in front of the name of the anion if two H^+ ions are attached.

HCO_3^- is hydrogen carbonate ion.

HS^- is hydrogen sulfide ion.

HPO_4^{2-} is hydrogen phosphate ion.

$H_2PO_4^-$ is dihydrogen phosphate ion.

Some polyatomic ions also have nonsystematic names that are often used. For example, HCO_3^- is often called bicarbonate instead of hydrogen carbonate. You should avoid using this less accepted name, but because many people still use it, you should know it.

Objective 36

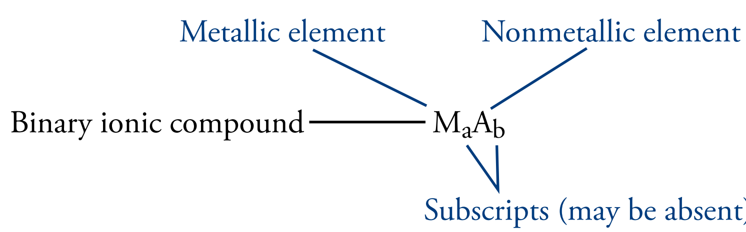
You can find a more comprehensive description of polyatomic ions, including a longer list of their names and formulas at the textbook's Web site.

Converting Formulas to Names

Objective 37

As we noted earlier, chemists have established different sets of rules for writing the names and formulas of different types of chemical compounds, so the first step in writing a name from a chemical formula is to decide what type of compound the formula represents. A chemical formula for an ionic compound will have one of the following forms.

- **Metal-nonmetal:** Ionic compounds whose formula contains one symbol for a metal and one symbol for a nonmetal are called **binary ionic compounds**. Their general formula is M_aA_b, with *M* representing the symbol of a metallic element, *A* representing the symbol for a nonmetallic element, and lowercase *a* and *b* representing subscripts in the formula (unless one or more of the subscripts are assumed to be 1). For example, NaF (used to fluoridate municipal waters), $MgCl_2$ (used in floor sweeping compounds), and Al_2O_3 (in ceramic glazes) represent binary ionic compounds.



- **Metal-polyatomic ion:** Polyatomic ions can take the place of monatomic anions, so formulas that contain a symbol for a metallic element and the formula for a polyatomic ion represent ionic compounds. For example, $NaNO_3$ (in solid rocket propellants) and $Al_2(SO_4)_3$ (a foaming agent in fire foams) represent ionic compounds.

- **Ammonium-nonmetal or ammonium-polyatomic ion:** Ammonium ions, NH_4^+, can take the place of metallic cations in an ionic compound, so chemical formulas that contain the formula for ammonium with either a symbol for a nonmetallic element or a formula for a polyatomic ion represent ionic compounds. For example, NH_4Cl (in dry cell batteries), $(NH_4)_2S$ (used to color brass), and $(NH_4)_2SO_4$ (in fertilizers) represent ionic compounds.

Objective 37

The names of ionic compounds consist of the name for the cation followed by the name for the anion. Tables 3.7 and 3.8 summarize the ways cations and anions are named.

Table 3.7
Summary of the Ways Cations Are Named

Objective 37

Type of cation	General name	Example
metal with one possible charge (groups 1, 2, 3–Al, Zn, and Cd)	name of metal	Mg^{2+} is magnesium.
metal with more than one possible charge (the rest of the metallic elements)	name of metal (Roman numeral)	Cu^{2+} is copper(II).
polyatomic cations (for us, only ammonium)	name of polyatomic cation	NH_4^+ is ammonium.

Table 3.8
Summary of the Ways Anions Are Named

Objective 37

Type of anion	General name	Example
monatomic anions	(root of nonmetal)ide	O^{2-} is oxide.
polyatomic anions	name of polyatomic anion	NO_3^- is nitrate.

As an example of the thought-process for naming ionic compounds, let's write the name for MnO (used as a food additive and dietary supplement). Our first step is to identify the type of compound it represents. Because it is composed of a metallic element and a nonmetallic element, we recognize it as ionic. Thus we must write the name of the cation followed by the name of the anion. Manganese is not on our list of metallic elements with only one possible charge, so the cation's name is the name of the element followed by a Roman numeral that represents the charge. Therefore, our next step is to determine the charge on each manganese cation in MnO.

When the cation in an ionic formula is created from a metallic element whose atoms can have more than one charge, you can discover the cation's charge by identifying the charge on the anion and then figuring out what charge the cation must have to yield a formula that is uncharged overall. To discover the charge on the manganese ions in MnO, you first determine the charge on the anions. A glance at the periodic table shows oxygen to be in group 16, or 6A, whose nonmetallic members always form -2 ions. With x representing the charge on the manganese ion, the charge on the manganese cation can be figured out as follows:

$$\text{total cation charge} + \text{total anion charge} = 0$$
$$x + (-2) = 0$$
$$x = +2$$

Each manganese cation must therefore be +2 to balance the -2 of the oxide to yield an uncharged ionic formula. The systematic name for Mn^{2+} is *manganese(II)*. Monatomic anions are named with the root of the nonmetal followed by –ide, so O^{2-} is *oxide*. MnO is named manganese(II) oxide. Example 3.8 provides other examples.

Example 3.8 - Naming Ionic Compounds

Objective 37

Write the names that correspond to the following formulas: (a) MgO (used to make aircraft windshields), (b) $CoCl_2$ (used to manufacture vitamin B-12), (c) NH_4NO_3 (used to make explosives), and (d) Fe_2O_3 (in paint pigments).

$CoCl_2$ is used to make vitamin B-12.

Solution

a. The compound MgO includes the cation Mg^{2+} and the anion O^{2-}. Magnesium cations are always +2, so the name of the cation is the same as the name of the metallic element. The anion O^{2-} is monatomic, so it is named by combining the root of the name of the nonmetal with -ide. Therefore, MgO is **magnesium oxide**.

b. Cobalt is not on our list of metallic elements that form ions with only one charge, so we assume it can form ions with more than one possible charge. Therefore, we need to show the charge on the cobalt ion with a Roman numeral in parentheses after the name cobalt. The cobalt ion must be +2 to balance the −2 from the two −1 chloride ions.

$$\text{Total cation charge} + \text{total anion charge} = 0$$
$$x + 2(-1) = 0$$
$$x = +2$$

The anion Cl^- is monatomic, so its name includes the root of the name of the nonmetal and -ide. Therefore, $CoCl_2$ is **cobalt(II) chloride**.

c. The compound NH_4NO_3 includes the cation NH_4^+ and the anion NO_3^-. Both of these ions are polyatomic ions with names you should memorize. The name of NH_4NO_3 is **ammonium nitrate**.

d. Iron is not on our list of metallic elements that form ions with only one charge, so we need to show the charge on the iron ion with a Roman numeral. Because oxygen atoms have two fewer electrons than the nearest noble gas, neon, they form −2 ions. In the following equation, x represents the charge on each iron ion. Because there are two iron ions, $2x$ represents the total cation charge. Likewise, because there are three oxygen ions, the total anion charge is three times the charge on each oxygen ion.

$$\text{total cation charge} + \text{total anion charge} = 0$$
$$2x + 3(-2) = 0$$
$$x = +3$$

The iron ions must be +3 in order for them to balance the −6 from three −2 oxide ions, so the cation name is iron(III). Because O^{2-} is a monatomic anion, its name includes the root of the name of the nonmetal and -ide. Therefore, Fe_2O_3 is **iron(III) oxide**.

Exercise 3.8 - Naming Ionic Compounds

Objective 37

Write the names that correspond to the following formulas: (a) $LiCl$ (used in soft drinks to help reduce the escape of bubbles), (b) $Cr_2(SO_4)_3$ (used in chrome plating), and (c) NH_4HCO_3 (used as a leavening agent for cookies, crackers, and cream puffs).

Converting Names of Ionic Compounds to Formulas

Before you can write a chemical formula from the name of a compound, you need to recognize what type of compound the name represents. For binary ionic compounds, the first part of the name is the name of a metallic cation. This may include a Roman numeral in parentheses. The anion name starts with the root of the name of a nonmetal and ends with *-ide*.

OBJECTIVE 37

(name of metal)(maybe Roman numeral) (root of nonmetal)ide

For example, aluminum fluoride (used in the production of aluminum) and tin(II) chloride (used in galvanizing tin) are binary ionic compounds.

You can identify other names as representing ionic compounds by recognizing that they contain the names of common polyatomic ions. For example, ammonium chloride and iron(III) hydroxide are both ionic compounds. Many of the polyatomic ions that you will be expected to recognize end in *-ate*, so this ending tells you that the name represents an ionic compound. Copper(II) sulfate is an ionic compound.

Follow these steps to write formulas for ionic compounds.

OBJECTIVE 37

STEP 1 Write the formula, including the charge, for the cation. (See Figure 3.17 to review the charges on monatomic cations.)

STEP 2 Write the formula, including the charge, for the anion. (See Figure 3.17 to review the charges on monatomic anions. See Table 3.6 to review polyatomic ion formulas.)

STEP 3 Write subscripts for each formula so as to yield an uncharged compound. Table 3.9 shows examples.

- Use the lowest whole number ratio for the subscripts.
- If the subscript for a polyatomic ion is higher than one, place the formula for the polyatomic ion in parentheses and put the subscript outside the parentheses.

Table 3.9
Possible Cation-Anion Ratios (X represents the cation, and Y represents the anion.)

OBJECTIVE 37

Ionic charges	General formula	Example ions	Example formula
X^+ and Y^-	XY	Na^+ and Cl^-	NaCl
X^+ and Y^{2-}	X_2Y	NH_4^+ and SO_4^{2-}	$(NH_4)_2SO_4$
X^+ and Y^{3-}	X_3Y	Li^+ and PO_4^{3-}	Li_3PO_4
X^{2+} and Y^-	XY_2	Mg^{2+} and NO_3^-	$Mg(NO_3)_2$
X^{2+} and Y^{2-}	XY	Ca^{2+} and CO_3^{2-}	$CaCO_3$
X^{2+} and Y^{3-}	X_3Y_2	Ba^{2+} and N^{3-}	Ba_3N_2
X^{3+} and Y^-	XY_3	Al^{3+} and F^-	AlF_3
X^{3+} and Y^{2-}	X_2Y_3	Sc^{3+} and S^{2-}	Sc_2S_3
X^{3+} and Y^{3-}	XY	Fe^{3+} and PO_4^{3-}	$FePO_4$

Example 3.9 - Formulas of Ionic Compounds

Objective 37

Write the chemical formulas that correspond to the following names: (a) aluminum chloride (used in cosmetics), (b) chromium(III) oxide (a pigment for coloring pottery glazes), (c) calcium nitrate (provides a red-orange color in fireworks), and (d) ammonium sulfide (used to make synthetic flavors).

Solution

a. Aluminum chloride has the form *(name of metal) (root of nonmetal)ide*, so we recognize it as a binary ionic compound. Because aluminum atoms have three more electrons than the nearest noble gas, neon, they lose three electrons and form +3 ions. Because chlorine atoms have one fewer electron than the nearest noble gas, argon, they gain one electron to form −1 ions. The formulas for the individual ions are Al^{3+} and Cl^-. It would take three chlorides to neutralize the +3 aluminum ion, so the formula for the compound is **$AlCl_3$**.

b. Chromium(III) oxide has the form *(name of metal)(Roman numeral) (root of nonmetal)ide*, so it represents a binary ionic compound. The *(III)* tells you that the chromium ions have a +3 charge. Because oxygen atoms have two fewer electrons than the nearest noble gas, neon, they gain two electrons to form −2 ions. The formulas for the ions are Cr^{3+} and O^{2-}. When the ionic charges are +3 and −2 (or +2 and −3), a simple procedure will help you to determine the subscripts in the formula. Disregarding the signs of the charges, use the superscript on the anion as the subscript on the cation, and use the superscript on the cation as the subscript on the anion.

$$Cr^{3+}\ O^{2-}$$

Chromium(III) oxide is **Cr_2O_3**.

c. Calcium nitrate has the form *(name of metal) (name of a polyatomic ion)*, so it represents an ionic compound. (The *-ate* at the end of nitrate tells us that it is a polyatomic ion.) Calcium is in group 2 on the periodic table. Because all metals in group 2 have two more electrons than the nearest noble gas, they all lose two electrons and form +2 ions. Nitrate is NO_3^-, so two nitrate ions are needed to neutralize the charge on the Ca^{2+}. Calcium nitrate is **$Ca(NO_3)_2$**. Notice that in order to show that there are two nitrate ions, the formula for nitrate is placed in parentheses.

d. Ammonium sulfide has the form *ammonium (root of a nonmetal)ide*, so it represents an ionic compound. You should memorize the formula for ammonium, NH_4^+. Sulfur has two fewer electrons than the noble gas, argon, so it gains two electrons and forms a −2 anion. Two ammonium ions would be necessary to neutralize the −2 sulfide. Ammonium sulfide is **$(NH_4)_2S$**.

Cr_2O_3 is a pigment used for coloring pottery glazes.

There is a tutorial on the textbook's Web site that provides practice converting between names and formulas of ionic compounds.

Exercise 3.9 - Formulas of Ionic Compounds

Objective 37

Write the formulas that correspond to the following names: (a) aluminum oxide (used to waterproof fabrics), (b) cobalt(III) fluoride (used to add fluorine atoms to compounds), (c) iron(II) sulfate (in enriched flour), (d) ammonium hydrogen phosphate (coats vegetation to retard forest fires), and (e) potassium bicarbonate (in fire extinguishers).

Chapter Glossary

Compound A substance that contains two or more elements, the atoms of these elements always combining in the same whole-number ratio.

Chemical formula A concise written description of the components of a chemical compound. It identifies the elements in the compound by their symbols and indicates the relative number of atoms of each element with subscripts.

Pure substance A sample of matter that has constant composition. There are two types of pure substances, elements and compounds.

Mixture A sample of matter that contains two or more pure substances and has variable composition.

Chemical bond An attraction between atoms or ions in chemical compounds. Covalent bonds and ionic bonds are examples.

Polar covalent bond A covalent bond in which electrons are shared unequally, leading to a partial negative charge on the atom that attracts the electrons more and to a partial positive charge on the other atom.

Nonpolar covalent bond A covalent bond in which the difference in electron-attracting ability of two atoms in a bond is negligible (or zero), so the atoms in the bond have no significant charges.

Ionic bond The attraction between a cation and an anion.

Molecular compound A compound composed of molecules. In such compounds, all of the bonds between atoms are covalent bonds.

Ionic Compound A compound that consists of ions held together by ionic bonds.

Valence electrons The electrons that are most important in the formation of chemical bonds. The number of valence electrons for the atoms of an element is equal to the element's A-group number on the periodic table. (A more comprehensive definition of valence electrons appears in Chapter 12.)

Electron-dot symbol A representation of an atom that consists of its elemental symbol surrounded by dots representing its valence electrons.

Lewis structure A representation of a molecule that consists of the elemental symbol for each atom in the molecule, lines to show covalent bonds, and pairs of dots to indicate lone pairs.

Lone pair Two electrons that are not involved in the covalent bonds between atoms but are important for explaining the arrangement of atoms in molecules. They are represented by pairs of dots in Lewis structures.

Hydrocarbons Compounds that contain only carbon and hydrogen.

Organic chemistry The branch of chemistry that involves the study of carbon-based compounds.

Double bond A link between atoms that results from the sharing of four electrons. It can be viewed as two 2-electron covalent bonds.

Triple bond A link between atoms that results from the sharing of six electrons. It can be viewed as three 2-electron covalent bonds.

Alcohols Compounds that contain a hydrocarbon group with one or more -OH groups attached.

Tetrahedral The molecular shape that keeps the negative charge of four electron groups as far apart as possible. This shape has angles of 109.5° between the atoms.

Bond angle The angle formed by straight lines (representing bonds) connecting the nuclei of three adjacent atoms.

Binary covalent compound A compound that consists of two nonmetallic elements.

Monatomic anions Negatively charged particles, such as Cl^-, O^{2-}, and N^{3-}, that contain single atoms with a negative charge.

Monatomic cations Positively charged particles, such as Na^+, Ca^{2+}, and Al^{3+}, that contain single atoms with a positive charge.

Polyatomic ion A charged collection of atoms held together by covalent bonds.

Binary ionic compound An ionic compound whose formula contains one symbol for a metal and one symbol for a nonmetal.

You can test yourself on the glossary terms at the textbook's Web site.

Chapter Objectives

The goal of this chapter is to teach you to do the following.

1. Define all of the terms in the Chapter Glossary.

Section 3.1 Classification of Matter

2. Convert between a description of the number of atoms of each element found in a compound and its chemical formula.
3. Given a description of a form of matter, classify it as a pure substance or mixture.
4. Given a description of a pure substance, classify it as an element or compound.

Section 3.2 Chemical Compounds and Chemical Bonds

5. Describe the polar covalent bond between two nonmetallic atoms, one of which attracts electrons more than the other one does. Your description should include a rough sketch of the electron-cloud that represents the electrons involved in the bond.
6. Describe the process that leads to the formation of ionic bonds between metallic and nonmetallic atoms.
7. Describe the difference between a nonpolar covalent bond, a polar covalent bond, and an ionic bond. Your description should include rough sketches of the electron-clouds that represent the electrons involved in the formation of each bond.
8. Given the names or formulas for two elements, identify the bond that would form between them as covalent or ionic.
9. Given a formula for a compound, classify it as either a molecular compound or an ionic compound.

Section 3.3 Molecular Compounds

10. Determine the number of valence electrons for the atoms of each of the nonmetallic elements.
11. Draw electron-dot symbols for the nonmetallic elements and use them to explain why these elements form the bonding patterns listed in Table 3.1.
12. Give a general description of the information provided in a Lewis structure.

13. Identify the most common number of covalent bonds and lone pairs for the atoms of each of the following elements: hydrogen, the halogens (group 17), oxygen, sulfur, selenium, nitrogen, phosphorus, and carbon.
14. Convert between the following systematic names and their chemical formulas: methanol, ethanol, and 2-propanol.
15. Given one of the following names for alcohols, write its chemical formula: methyl alcohol, ethyl alcohol, and isopropyl alcohol.
16. Given a chemical formula, draw a Lewis structure for it that has the most common number of covalent bonds and lone pairs for each atom.
17. Describe the tetrahedral molecular shape.
18. Explain why the atoms in the CH_4 molecule have a tetrahedral molecular shape.
19. Describe the information given by a space-filling model, a ball-and-stick model, and a geometric sketch.
20. Draw geometric sketches, including bond angles, for CH_4, NH_3, and H_2O.
21. Describe attractions between H_2O molecules.
22. Describe the structure of liquid water.

Section 3.4 Naming Binary Covalent Compounds

23. Convert between the names and chemical formulas for water, ammonia, methane, ethane, and propane.
24. Given a formula or name for a compound, identify whether it represents a binary covalent compound.
25. Write or identify prefixes for the numbers 1-10. (For example, mono- represents one, di- represents two, etc.)
26. Write or identify the roots of the names of the nonmetallic elements. (For example, the root for oxygen is ox-).
27. Convert among the complete name, the common name, and the chemical formula for HF, HCl, HBr, HI, and H_2S.
28. Convert between the systematic names and chemical formulas for binary covalent compounds.

Section 3.5 Ionic Compounds

29. Explain why metals usually combine with nonmetals to form ionic bonds.
30. Write the ionic charges acquired by the following elements:
    a. group 17 – halogens
    b. oxygen, sulfur, and selenium
    c. nitrogen and phosphorus
    d. hydrogen
    e. group 1 - alkali metals
    f. group 2 - alkaline earth metals
    g. group 3 elements
    h. aluminum
    i. iron, silver, copper, and zinc
31. Convert between the names and chemical formulas for the monatomic ions.

32. Describe the crystal structure of sodium chloride, NaCl.
33. Describe the similarities and differences between the ionic structure of cesium chloride and ammonium chloride.
34. Convert between the names and chemical formulas for common polyatomic ions such as hydroxide, ammonium, acetate, sulfate, nitrate, phosphate, and carbonate. *Be sure to check with your instructor to determine which polyatomic ions you will be expected to know for your exams.*
35. Convert between the names and chemical formulas for the polyatomic anions that are derived from the additions of H^+ ions to anions with –2 or –3 charges. For example, $H_2PO_2^-$ is dihydrogen phosphate.
36. Write the chemical formula corresponding to the common name bicarbonate.
37. Convert between the names and chemical formulas for ionic compounds.