CHAPTER

3

Ionic Compounds: Opposites Do Attract

LEARNING OBJECTIVES

After reading this chapter, you will be able to:

1. Use the Lewis dot structure of an atom to predict its charge when it becomes an ion.
2. Relate the periodic table to the formula writing and naming of binary ionic compounds.
3. Discuss the role of ionic compounds with metal ions whose charge cannot be predicted from the periodic table.
4. Apply ionic compound properties to everyday uses.

At a Glance

Photograph by Tylar Mueller

Gallons of a popular sports drink douse a coach as part of a victory celebration.

Photographs by Tylar Mueller

FIGURE 3.1 Electrolytes in our water supply are a serious problem for many water districts.

Sports drinks were introduced to the public in 1965 as a means to support the body during strenuous athletic exercise. The secret of their success was in the electrolytes they contain.

The idea caught on quickly, and soon Gatorade and its cousins became omnipresent items in sports clubs, at sporting events, and anywhere someone wanted something deemed healthier than water. Soon the drinks became so strongly associated with athletic performance that the "Gatorade shower" became a playful means to congratulate a winning coach.

At the same time, water districts across the nation are struggling with a grave concern: electrolytes in our groundwater (Fig. 3.1). Nitrates from fertilizer and dairy farm runoff can cause infants to turn blue from lack of oxygen; perchlorates in the groundwater from industrial manufacturing attack thyroid function. Perhaps the best-known electrolytic contamination is hexavalent chromium, which was the subject of the film *Erin Brockovich*, starring Julia Roberts, and the object of one of the largest legal settlements in U.S. history.

How can something good for us (electrolytes) also be harmful? What are electrolytes? Where are they found? What do they have to do with ionic compounds, the topic of this chapter?

This chapter takes a closer look at compounds formed through ionic bonding. You will discover:

- The properties of ionic compounds.
- The wide range of substances formed through ionic bonds, both useful and dangerous.
- How ions and ionic compounds are named.
- Those elements on the periodic table most likely to form ionic compounds.

Let's get started.

3.1 Evidence of Compound Structure

As we move forward in our study of chemistry, it is worth taking another look back at those early chemists whose work laid the foundation for our understanding of compounds. We've already met the scientists who developed—and then proved—atomic theory. The path to developing that theory focused on the composition of various substances, using observation and logic to develop hypotheses that could then be tested.

3.1.1 Law of Constant Composition

In the late 1700s, French chemist Joseph Proust was experimenting with copper carbonate. He discovered that no matter where he obtained his samples of copper carbonate—in a laboratory or from natural sources—they always separated into their components of copper, carbon, and oxygen in the exact same proportions. He followed up with experiments on metal oxides with similar results.

The conclusion he reached was that "elements combine in specific mass ratios to form compounds." This became known as the **law of constant composition**, sometimes called the *law of definite proportions*.

For instance, no matter where you find water (H_2O) or how much of it you sample, if you break it down into its components, the ratio of hydrogen to oxygen is always the same. For every 8 grams of oxygen, 1 gram of hydrogen is present. It never varies (Fig. 3.3).

Although this may seem obvious to us today, the consistency this idea presented was controversial. Up until then, there was no distinction between mixtures and compounds. It was common knowledge that one could mix sugar and water, for instance, in many different proportions—you might put one, two, three, four, or even more teaspoons of sugar in a cup of tea. Chemists thought that all combinations of elements worked in the same way and could exist in an indefinite number of proportions.

Proust's law of constant composition was so revolutionary that it took years for it to become accepted. Once it was, it created an important distinction between mixtures and chemical compounds.

© Michael Nicholson/Corbis

FIGURE 3.2 Joseph Louis Proust (1754–1826).

law of constant composition The scientific law that elements combine in specific mass ratios to form compounds.

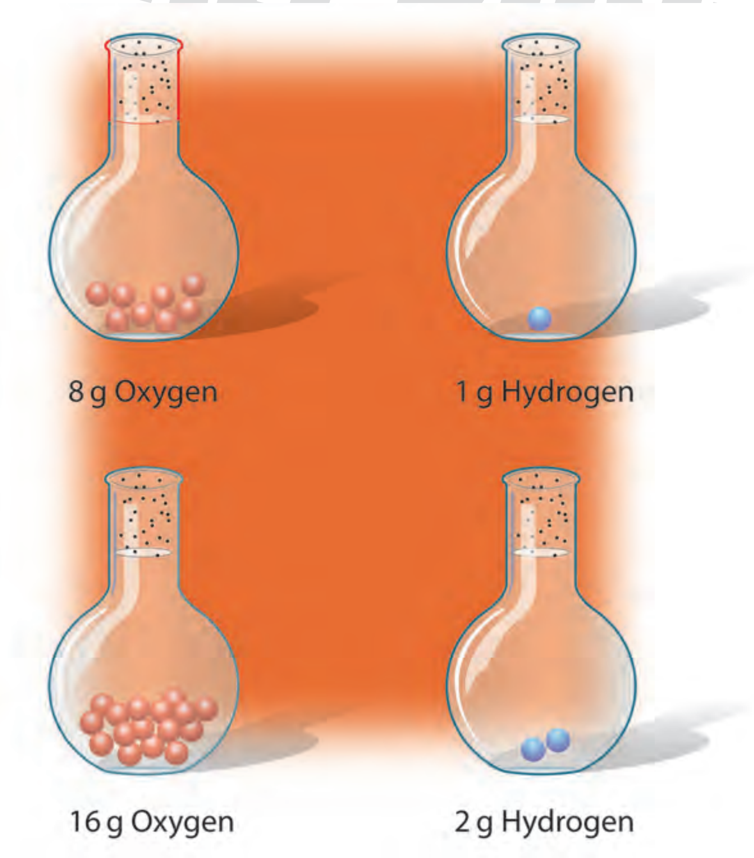


Illustration by Don Vierstra

FIGURE 3.3 If you break water down into its components, the ratio between the mass of oxygen and the mass of hydrogen is always 8:1 mass ratio regardless of how much water there was to begin with.


© 2009, JupiterImages Corporation

FIGURE 3.4
John Dalton (1766–1844).

3.1.2 Law of Multiple Proportions

John Dalton (Fig. 3.4), who was introduced in Chapter 1, took Proust's thinking one step further in proposing the **law of multiple proportions**. He stated that not only do elements combine in consistent ratios, but that *these ratios are always whole numbers.*

For instance, atoms of sodium and chlorine always combine in a 1:1 ratio to form table salt. Likewise, atoms of hydrogen and oxygen combine in a 2:1 ratio to form water (H_2O), and although there is an ionic formula of H_3O^+ that looks similar to water (about which you will learn more in Chapter 11), its ratio of 3:1 still represents whole numbers. There cannot be a molecule with a ratio that is $H_{2.5}O$, ionic or otherwise.

Furthermore, atoms of the same elements can form multiple molecules with each other. Atoms of carbon and oxygen react in different ratios to form carbon dioxide, CO_2, and carbon monoxide, CO. These compounds have different structures and different physical properties. Carbon monoxide is a poisonous gas, while carbon dioxide gas is what we exhale!

The idea of whole number ratios implies that there is a limit to how many times a sample of an element can be divided. This laid the groundwork for Dalton's atomic theory, which is, as we discussed in Chapter 1, the foundation for chemistry.

law of multiple proportions The scientific law stating that the ratios in which elements combine to form compounds are always whole number ratios.

3.2 Properties of Ionic Compounds

Chapter 2 explained that elements combine to form an ionic compound when electrons from the atoms of one element transfer to the atoms of another so that both will have the configuration of the noble gas nearest to each. This is a stable configuration.

You also learned that when an electron transfers, the atom that loses it becomes positively charged (the cation) and the atom that gains it becomes negatively charged (the anion). When atoms gain a charge because they add or lose electrons, we say that they **ionize**. The ionization is the act of gaining a charge. Because opposite charges attract, it is almost inevitable that the newly ionized cation and anion will then bond together to form an ionic compound.

ionize The act of gaining a charge when an atom loses electrons or adds electrons to become an ion.

Let's take a look at what happens next, using our earlier example of sodium chloride (NaCl), common table salt. Sodium loses an electron to chlorine; the two atoms become charged and bond together. But sodium and chlorine atoms rarely travel alone, the reaction rapidly spreads to other sodium and chlorine atoms, and more salt is formed (Fig. 3.5).


Courtesy of Lonnie Martin

FIGURE 3.5 Electrons transfer from the sodium atoms to the chlorine atoms at an astonishing rate, releasing heat and light in the process.

Phosphorus: The Limiting Factor

Living organisms grow, die, and decompose in a series of chemical interactions in which their elements are continuously recycled and reused. Six elements—carbon, oxygen, hydrogen, nitrogen, sulfur, and phosphorus—are primarily responsible for the form and structure of living tissues. All except phosphorus are abundant in the atmosphere, oceans, and lakes as N_2, O_2, H_2O, CO_2, and SO_2. But phosphorus does not occur in the atmosphere, and its salts are largely insoluble. Most phosphorus is recycled through living organisms by way of food chains. In a lake, for example, microorganisms feed on dead material containing phosphorus, while large organisms feed on each other. In this way, the limited supply of phosphorus is used over and over again.

Because the availability of phosphorus in this environment is limited, the amount of life a lake can support is thereby determined by its phosphorus content. In this respect the phosphorus content of a lake serves as a brake against runaway growth. However, when this limitation is removed the delicate balance of life is disrupted and spirals out of control. Phosphorus containing detergents are finding their way into our wastewater and subsequently into these lakes, further disrupting this balance.

Phosphates are polyatomic ions that contain phosphorus and oxygen and sometimes hydrogen and form ionic bonds with metals. Some phosphate compounds such as those containing group IA metals are soluble and break down into ions when dissolved in water, some are not so soluble such as the group IIA metal phosphate compounds.

Group IA metal phosphates such as sodium or potassium phosphates are found in some laundry detergents because they enhance their "cleaning power" with promises of whiter and brighter laundry, manufacturers added phosphates to more and more detergents setting off a chain of events leading to another ecological disaster. Since algae feed on phosphorus, algaeal blooms began to grow. Before too long, scientists noticed that this increased growth of algae was reducing the amount of sunlight that lake bottom plants received. Without light, these plants cannot produce oxygen. As a result, the oxygen content of the lakes began to dwindle. Without enough oxygen in the water fish can no longer survive and began to die and of course decay. This of course decreases the oxygen content further and more fish died. Soon the lakes can no longer support any life and everything in the lake dies, a catastrophe called eutrophication.

Phosphates are just one group of polyatomic ions increasingly found in our water supplies as we make better detergents, more powerful rocket fuel, better computers and pipes that do not corrode. For some of these ions, the impact is not fully understood,and may not be for many years to come.

A lake in the process of eutrophication. An excess of nutrients such as phosphorus encourages an overgrowth of green algae, which leads to a cycle of death and decay.

Immediately, all the newly formed sodium cations attract the new chloride anions around them, and vice versa. They bond together, not into individual molecules but into a single, repeating, three-dimensional structure called a **crystal lattice structure**. Each atom in the structure is surrounded by atoms of the opposite charge (Fig. 3.6).

If you have ever looked closely at table salt, you will have noticed that each particle is a separate cubic crystal (Fig. 3.7). Each of these crystals contains billions and billions of ions or charged atoms bonded together in a lattice structure. In fact, no matter where you find solid salt—or how large the sample is—it will exist in the same cubic crystalline form.

This lattice structure is typical of ionic compounds, although the exact shape varies. For instance, calcium fluoride (CaF_2), better known as fluorite, has two fluoride atoms for each calcium atom (Fig. 3.8). It forms a cubic crystal that often assumes an octahedral shape.

Other crystal lattice shapes exist as well. The point is that ionic compounds do not exist as individual units, as covalent molecules do. The attraction between ions is so strong that millions of them invariably bond together in larger structures.

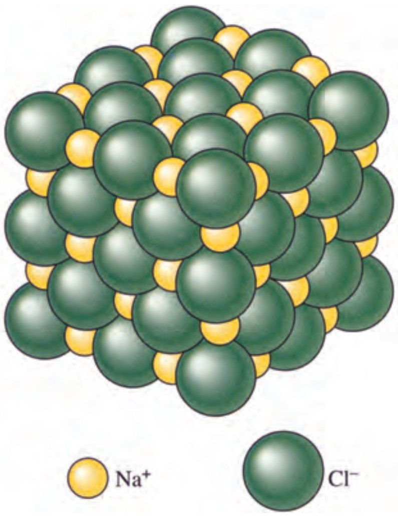


FIGURE 3.6 Sodium chloride consists of sodium and chloride ions in an alternating cubic array.

American Museum of Natural History

FIGURE 3.7 Crystals of halite (sodium chloride) in a rock matrix.

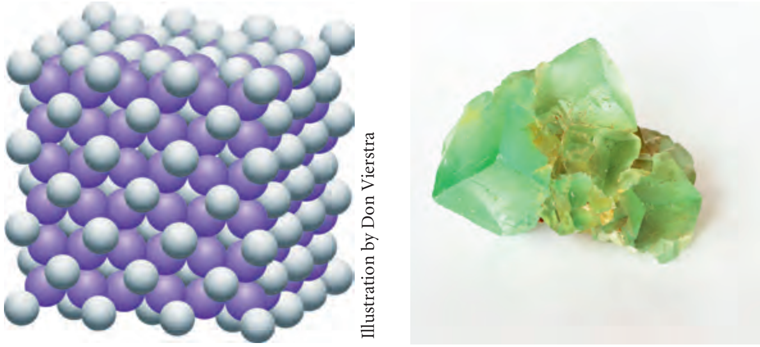


Illustration by Don Vierstra

© Laitr Keiows, 2009, Shutterstock Inc.

FIGURE 3.8 Flurorite (CaF_2) forms a crystal lattice in which there are two fluoride ions (white) to every calcium ion (blue).

crystal lattice structure A repeating three-dimensional arrangement of atoms/ions in a solid.

soluble Able to dissolve in water.

electrolytes Substances that dissolve in water to form solutions that conduct electricity.

Besides forming crystalline structures when solid, ionic compounds share other properties:

- They are solid under the "standard conditions" of room temperature and sea-level pressure.
- They have high melting and boiling points—ionic bonds are so strong that it takes an enormous amount of energy to break them apart into a liquid. For instance, the melting point of table salt is 800.8°C, or 1473.4°F.
- Many ionic compounds are **soluble** in water, meaning that they are able to dissolve in water.
- When molten or dissolved, they are **electrolytes**, substances that conduct electricity when dissolved in water.

We will delve a little deeper into some of these properties later in this chapter. First, let's take a closer look at ionic bonding and the elements that enter into such bonds. To do so, we need to once again revisit the periodic table—as we will do many, many more times!

3.2.1 A Positive or Negative Charge?

As you recall from the previous chapters, the organization of the periodic table gives much information about each element. Some of this information will help you to predict which elements are likely to form ionic bonds. Some of this is simply logical. You already know that the elements on the left side of the periodic table—the metals—want to give up electrons to achieve the configuration of the nearest noble gas. Meanwhile, the elements on the right side of the table—the nonmetals—are anxious to take on electrons to achieve the same stable configuration. Consequently, it stands to reason that *ionic bonds are between metals and nonmetals*.

We can also predict the ratio of metal and nonmetal elements in ionic compounds by using their group numbers. Remember, the group numbers indicate the number of valence electrons. Take a look at the periodic table. With which group are the elements in the first column, Group 1A, likely to bond in a 1:1 ratio?

If you said Group 7A, you are getting the idea. Elements in Group 1A want to lose an electron; elements in Group 7A want to gain one. The transfer of a single electron completes the valence shell of both atoms. So Group 1A elements will combine with Group 7A elements in a 1:1 ratio. For instance, as you well know by now, sodium combines with chlorine in a 1:1 ratio, and lithium combines with fluorine in a 1:1 ratio.

But what happens when elements in Group 1A meet up with elements from Group 6A? The latter needs to take on two electrons to fill their valence shells, but the former have only one electron to lose.

The answer is that electrons from *two* atoms of a first column element must transfer their electrons to fill the valence shell of *one* sixth column atom. Therefore, Group 1A elements react in a 2:1 ratio with elements in Group 6A.

As an example, sodium bonds to sulfur in a 1:2 ratio to form sodium sulfide (Na_2S). There are two sodium cations to one sulfur anion.

In ionic compounds, positively charged metal cations always react in simple ratios with the negatively charged anions. The periodic table and Lewis dot structures give us the tools we need to predict those ratios.

3.2.2 Fixed Charge Metals and Lewis Dot Structures

Using the periodic table, we can often, but not always, make predictions about ratios and the resulting chemical formulas by using the group numbers for the elements in an ionic compound. Whereas some metals are **fixed charge metals**, whose ions always have the same charge, other metals, called the **variable charge metals**, form multiple ions with different charges. Some variable charge metals are introduced later in this chapter; meanwhile, following are the rules for fixed charge metals:

fixed charge metals Metals that always take the same charge when ionized.

variable charge metals Metals that can take different charges when ionized. This includes all metals that do not have a fixed charge.

- Metals in Group 1A will always take a +1 charge because they have one electron to lose.
- Metals in Group 2A always take a +2 charge because they have two electrons to lose.

In addition, five other metals will always take a constant charge:

- Aluminum (Al) and gallium (Ga) in Group 3A will always have the +3 charge predicted by their group number.
- Zinc (Zn) and cadmium (Cd) in Group 12B will always have a +2 charge.
- Silver (Ag) in Group 11B will always have a +1 charge.

Memorizing these five elements and their charges is a good idea that will streamline solving ionic bonding problems, such as the ones you will face in chemistry exams.

Let's try a few examples, using Lewis structures to help us identify the ratios in compounds containing fixed charge metals.

Example

PROBLEM: Show how electrons are lost and gained to form the following compounds: (a) KBr, (b) $AlCl_3$, and (c) Mg_3N_2 .

Solution on page 88.

Once the manner in which electrons transfer is understood, we can easily derive the chemical formula for an ionic compound by using a shortcut: the "criss-cross method." Because "electrons lost equal electrons gained," we can criss-cross our ion charges and then reduce the subscripts, if necessary, to get the simplest ratio of ions in the compound.

For instance, in a bond between magnesium (Mg) and oxygen (O), we know from the periodic table that magnesium forms a 2+ cation, whereas oxygen, as a member

Group 6A, will form a 2− anion. We criss-cross these charges and then reduce the subscripts as follows:

$$Mg^{2+} \; O^{2-} \longrightarrow Mg_2O_2 \longrightarrow MgO$$

Concept Check Using charges from the periodic table, predict formulas for the following compounds by criss-crossing charges and reducing to the simplest ratio:

a. Na and S

b. Al and Br

Solution on page 89.

3.2.3 Transition or Variable Charge Metals in Ionic Compounds

Now that you are adept at using Lewis structures to figure out the chemical formulas for ionic compounds that contain fixed charge metals, it's time to meet the variable charge metals. For that, let's once again go back to the periodic table.

You have already memorized which metals have fixed charges. Look for the locations of the fixed charge metals on the periodic table. Then find the remainder of the metals; they are the variable charge metals. *All metals that are not fixed charge metals are variable charge metals*. Variable charge metals include all transition metals except for zinc, cadmium, and silver, and all metals in Groups 3A, 4A, 5A, and 6A except for aluminum and gallium. In fact, there are more variable charge metals than there are fixed charge metals—although, as it turns out, most of the ionic compounds we will explore contain ions of fixed charge metals.

The variable charge metals form multiple ions because the octet rule no longer applies to them. Their atoms may give up between 1 and 7 electrons; the exact range varies from element to element. For instance, iron (Fe) will lose either two or three electrons. When it loses two electrons, it becomes a cation with a 2+ charge; when it loses three electrons, it becomes a cation with a 3+ charge. Copper (Cu) will lose one or two electrons to take on a 1+ or 2+ charge.

Chromium (Cr)—which was introduced at the beginning of this chapter, will lose two, three, or six electrons to form ions with a 2+, 3+, or 6+ charge.

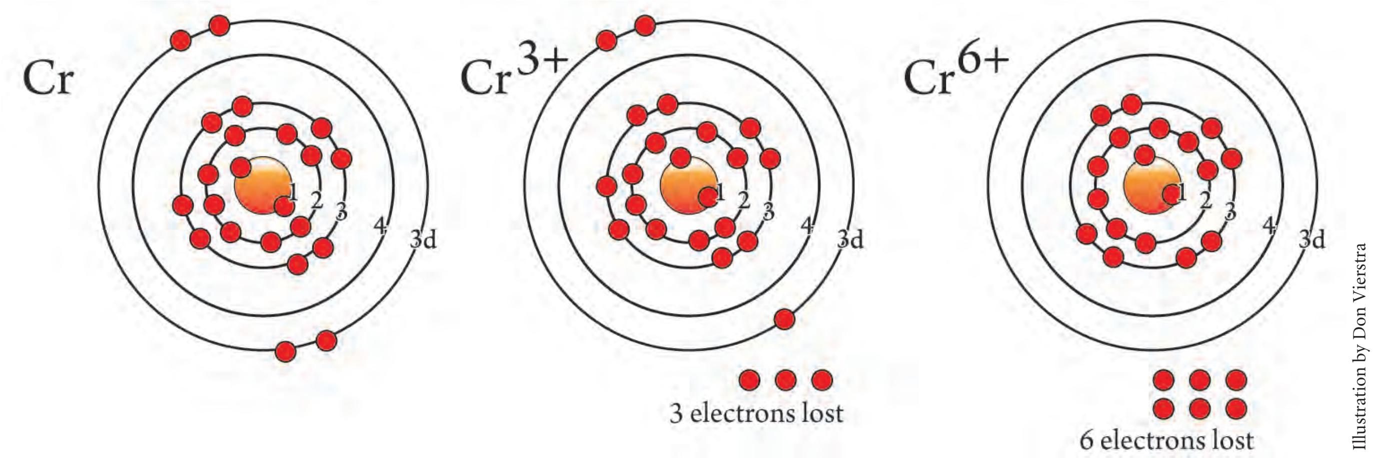


FIGURE 3.9 When a chromium atom loses electrons, it becomes a cation because it now has more protons than electrons. Cr(II) has a 2+ charge because it has lost two electrons, Cr(III) has a 3+ charge because it has lost three electrons, and hexavalent chromium, Cr(VI) has a 6+ charge because it has lost six electrons.

When writing formulas, the different charges a single element takes are indicated by Roman numerals. An iron cation with a 3+ charge is written as Fe(III); a copper cation with a 2+ charge is Cu(II), and chromium with a 6+ charge is written as Cr(VI).

When you know the charge of a variable charge metal in a compound, writing the chemical formula or drawing a Lewis structure is easy; just use the same procedures you have been using for compounds with fixed charged metals.

But what if you don't know the charge of a metal in a compound? How can you predict how many electrons will transfer? Chemists can and do make such predictions, but the details are better suited to the advanced study of chemistry than an introductory course.

However, if we already know the chemical formula for an ionic compound, we can determine the charge of its metal ion by "doing the math." Because the number of electrons lost must equal the number of electrons gained, we can use the periodic table to determine what charge the nonmetal ion will take and use that to tell us what charge the metal will take.

For example, one of the compounds iron forms with oxygen is FeO. We can see that there is only one oxygen ion present with a charge of −2. Therefore, because there is only one iron ion, it must take a +2 charge to equal the number of electrons gained by oxygen. Therefore, the metal ion must be Fe(II).

Iron also forms another compound with oxygen: Fe_2O_3. We can calculate the charge of the metal ion in this compound as follows:

3 oxygen ions are present, each has a charge of −2.

$$3 \times -2 = -6$$

There are 2 Fe ions to compensate for the −6 charge, so each will have a charge of +3.

$$+6/2 \text{ Fe ions} = +3$$

Simple!

Example

Use the anion in each of the following formulas to determine the charge on the variable charge metals.

a. $FeCl_3$ b. SnS_2 c. Cr_3O_3

Solution on page 89.

Concept Check

Use the anion in the following compounds to determine the charge on each of the variable charge metals.

a. PbI_2 b. CoN   c. NiO_2

Solution on page 89.

3.3 Naming Ionic Compounds

Pick up a bottle of your favorite sports drink and take a look at the ingredient list (Fig. 3.10). Some of the polysyllabic terms you'll find there can be downright intimidating. By the end of this section, however, you'll not only be able to interpret some of them but write the formulas based on their names! You'll learn others when molecular compounds are discussed in Chapter 4.

Photograph by Tylar Mueller

FIGURE 3.10 Sports drinks contain a number of ionic compounds with intimidating names—until you learn how to "translate" them!

binary compound Compound that contains only two elements; for example, NaBr, KCl, HCl.

3.3.1 Naming Binary Ionic Compounds

A **binary compound** is a compound that contains only two elements. The first rule for naming binary ionic compounds follows the first rule for writing their formulas: *the cation is always listed first.* (If the cation is a variable charge metal, the charge should follow.) For sodium chloride (NaCl), sodium is the cation. For potassium sulfide (K_2S), potassium is the cation. When a variable charge metal is the cation, we indicate which ion it is in the compound's name; for example, FeO is iron(II) oxide (spoken as "iron 2 oxide").

Now let's look at naming conventions for the anion. In a simple ionic compound, the anion ends with *-ide.* You may have noticed that while we talk about chlor*ine* atoms, we refer to table salt as sodium chlor*ide*. Now you know why.

Following is a list of some common ionic compounds ending with *-ide*:

$NaCl$ = sodium chloride
MgO = magnesium oxide
$CaBr_2$ = calcium bromide
K_2S = potassium sulfide
Mg_3N_2 = magnesium nitride
FeO = iron(II) oxide
PbI_2 = lead(II) iodide
$AsCl_3$ = arsenic(III) chloride
$AsCl_5$ = arsenic(V) chloride

3.3.2 Naming Polyatomic Ionic Compounds

Names also tell us when the anion is a polyatomic ion. Following is a list of some polyatomic ions with their names and charges. You need to memorize only a few of these—the ones you will use most often: ammonium (NH_4^+), acetate ($C_2H_3O_2^-$), and hydroxide (OH^-). As you use the others, however, you may find that you are inadvertently memorizing more than you intended to!

Nitrogen	Nitrite	NO_2^-
	Nitrate	NO_3^+
	Ammonium	NH_4^+
Sulfur	Sulfite	SO_3^{2-}
	Sulfate	SO_4^{2-}
	Hydrogen sulfate (or bisulfate)	HSO_4^-
Phosphorus	Phosphite	PO_3^{3-}
	Phosphate	PO_4^{3-}
	Hydrogen phosphate	HPO_4^{2-}
	Dihydrogen phosphate	$H_2PO_4^-$
Chlorine	Hypochlorite	ClO^-
	Chlorite	ClO_2^-
	Chlorate	ClO_3^-
	Perchlorate	ClO_4^-
Carbon	Acetate	$C_2H_3O_2^-$
	Carbonate	CO_3^{2-}
	Hydrogen carbonate or bicarbonate	HCO_3^-
	Cyanide	CN^-
Oxygen and hydrogen	Hydroxide	OH^-
	Peroxide	O_2^{2-}

A Case of Heavy Metals: The Diary of a Chemical Detective

The email from the Source Control Administrator read: "High Nickel." Somehow, somewhere the waste water flowing into our facility had picked up heavy metal ions that we would need to clean up before our waste water reached the ocean. But where was it coming from?

The dumping of heavy metals into wastewater poses a problem for our clean-up systems, which is why it is illegal. I had to act fast. As a Source Control Inspector, I'm a chemist responsible for inspecting industries whose wastewater flows into the Orange County Sanitation District's Treatment Plant 2. I headed for the Plant 2 Control Center where the Plant Operator and I looked at the pH charts for the sewer trunks. We noticed a smooth undulating pattern resulting from occasional high-acidity events on one of the sewer trunks. That identified the problem trunk.

The next question was critical. I asked the Plant Operator if there was evidence that the nickel ions had reached our secondary "Activated Sludge Process" that uses microorganisms to clean our waste water. The nonverbal communication I received back told me I needed to solve this fast. Heavy metal ions can zap out the "good bugs" needed to treat the wastewater, one by one, lowering efficiency of the process. This meant more work for Plant Operators to keep the microorganism population balanced.

Driving back to my office at OCSD's Plant 1, my head was swirling with questions: How much acid was in the waste water? How strong would it have to be to produce the test results I witnessed? How could heavy metal ions, particularly nickel and presumably in illegally dumped acid, get past the primary treatment at Plant 2? And finally, what kind of 'industrial fingerprint' did the undulating pattern represent? I'd get to work right away to answer my questions. None of my fellow Source Control Inspectors were assigned heavy metal industries on the trunk sewer. Nevertheless, they'd be doing their own precautionary detective work on the other sewer trunks.

Back at my desk, I calculated pH ranges and volumes, and narrowed down the list of suspects from my known list of industries on that trunk. None of the usual "heavy metal" suspects used the type of nickel that could pass through our primary treatment. It wasn't them. But there was one suspect with a nitric acid metal strip process that would allow nickel ions to do just that. Furthermore, this suspect had two distinctly separate wastewater processes under the roof of one large facility. One process had no heavy metals or acids, and its alkaline wastewater pumped out periodically "against" the gravity flow of the second process, which did include heavy metals and strong acids. Aha! That was the source of the undulating pattern! It looked like an industrial fingerprint match, including the nickel!

For this case, I contracted a criminal investigator at the Orange County District Attorney's Environmental Crimes Unit. That night, we began our joint investigation of the suspect industry, including the collection of criminal evidence of illegal dumping to both the sewer and the storm drain! The investigation ultimately resulted in a Superior Court sentence of the company owner to 3-years criminal probation and a very large fine. That company never violated in Orange County again.

—Deon Carrico, OCSD

Let's take a closer look at the list of polyatomic ions. First, note that all of the polyatomic ions on the list are anions except one: the ammonium ion (NH_4^+) is a cation! (Remember, this means that it will be listed first in a name.)

Now look at the charges.

- Among the nitrogen polyatomic ions, note that both the nitrite and nitrate ions have a 1− charge.
- Among polyatomic ions with sulfur, the sulfite and sulfate ions both have a 2− charge (just as sulfide does).
- In the phosphorus group of polyatomic ions, both the phosphite and phosphate have a 3− charge (just as phosphide does).
- In the group of polyatomic ions that contain chlorine, all the polyatomic ions have a 1− charge (just as chloride does).

Within each of the groups in the previous list, the charge changes only when a hydrogen atom (with its own 1+ charge) is added.

Now that you are more familiar with the list of polyatomic ions, let's take another look at our sports drink. You should be able to pick out which additives are ionic compounds. In this example, potassium phosphate stands out (Fig. 3.11). What would the formula for this ionic compound be?

When determining the formula for a compound containing a polyatomic ion, the key is to consider the ion as a single entity even though it contains more than one atom.

Photograph by Tylar Mueller

FIGURE 3.11 The ingredient list for a familiar sports drink. Can you recognize the ionic compounds? (Hint: Check the suffixes!)

The chart earlier in this section indicates that the phosphate ion has three extra electrons. If we make a compound from potassium and phosphate, we need to get those three electrons from potassium. Remember each potassium loses one electron. The formula for potassium phosphate is, therefore, K_3PO_4.

When we name ionic compounds that contain polyatomic ions, we keep the name of the polyatomic ion as it is, we do not use the "ide" ending. Let's take a look at some additional examples of ionic compounds.

Example

PROBLEM: Predict the names and chemical formulas of the ionic compounds formed between the following ions.

a. Na^+ and CN^- b. Ba^{2+} and CO_3^{2-} c. Ca^{2+} and NO_3^-

Solution on page 89.

Concept Check

PROBLEM: Determine the formulas and names for the following ionic compounds.

a. NH_4^+ and S^{2-} b. Na^+ and PO_4^{3-} c. Al^{3+} and CO_3^{2-}

Solution on page 90.

3.4 Ionic Compounds Are Everywhere

You don't have to be very far into your day before you have a personal interaction with an ionic compound. Ionic compounds are all around you, from the fluoride in your toothpaste to the salt on your eggs and the smelly stuff in your window cleaner. Their uses in household products make life easier; ingestion of these substances can support your life processes, as sports drinks do, or harm your health, as hexavalent chromium does.

Let's take a look at where some of the most common ionic compounds put in their appearance.

dissociate To break down into essential particles. In ionic compounds, the essential particles are ions; in molecular compounds, the essential particles are molecules.

3.4.1 Electrolytes

nonelectrolytes Solutes that dissolve in water to form solutions that do not conduct electricity.

Now that you can decipher some of the ingredients in a sports drink, we can delve into them a little further to find out why these beverages are generally considered to be good for you. As the advertising has told you repeatedly, sports drinks contain electrolytes. Let's take a closer look at what that means.

When an ionic compound, like NaCl, is put into water, it dissolves. Water molecules surround the solid and pull it apart. The compound **dissociates**, or breaks apart into its essential particles, the $Na^+$and Cl^- ions, as it dissolves. When a substance dissociates into ions, it is often said to ionize because the neutral solid is breaking down into charged particles.

A solution containing ions with positive and negative charges will conduct electricity, so we refer to it as an **electrolyte solution**. A sports drink is an electrolyte solution, so when you gulp one down, you are ingesting ions!

Some covalently bonded compounds, such as sugar, also dissolve in water. When they dissolve, the molecules do not break up into ions but remain as individual molecules with no charge. Therefore, these solutions either do not conduct electricity or do so very weakly; substances that do not conduct electricity when dissolved are **nonelectrolytes**; they dissolve to form **nonelectrolyte solutions** (Fig. 3.12).

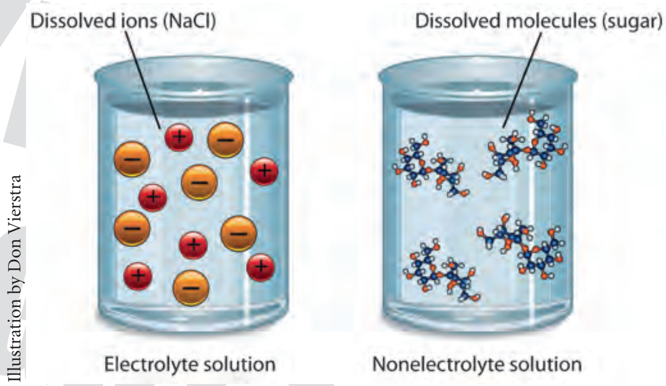


FIGURE 3.12 Dissolved ions form an electrolyte solution, sugar molecules dissolve, but they do not break up into ions and are nonelectrolytes.

Scientists use the conduction of electricity to differentiate between ionic substances and those of molecular compounds. Ionic compounds dissolve to form solutions that conduct electricity, whereas solutions of molecular compounds do not. Ionic or electrolytic solutions are used to facilitate the transfer of electrons in batteries, whereas covalently bonded substances do not break up into ions and, therefore, cannot complete a circuit. The role of electrolytes in batteries is discussed further in Chapter 13.

You may be wondering: If ionic bonds are so strong that it takes enormous energy to melt ionic compounds, why do they break up so readily in water? The answer lies in the polarity of the water molecule you learned about in Chapter 2. The partially positive hydrogen atoms in the water molecule tug at the chloride anion, whereas the partially negative oxygen atom tugs at the sodium cation. If there are enough water molecules surrounding the salt crystal, the attractions are strong enough to break the bonds in the NaCl and the ions separate (Fig. 3.13). It takes a lot of water molecules lining up next to one sodium ion and one chloride ion to overcome the forces that hold the Na^+ and Cl^- ions together.

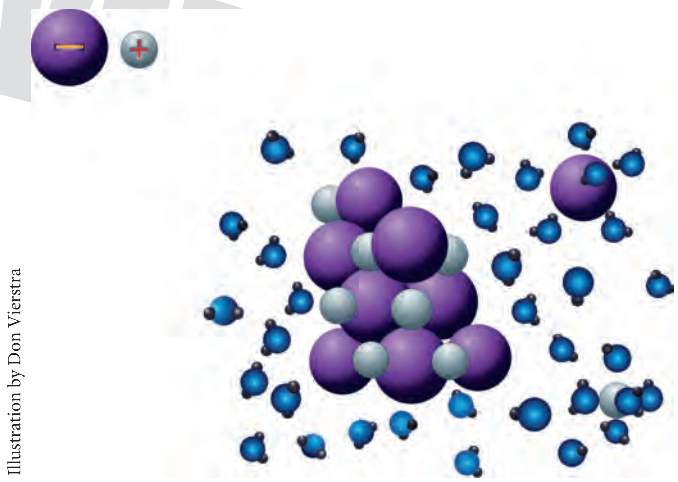


FIGURE 3.13 The charges in the polar water molecules pull apart ionic compounds, like NaCl, into individual ions.

Interestingly, most ionic compounds are most useful—or most harmful—when in solution. Most acids and bases are reactive only when in solution. Although there are some uses for salts in the solid phase (such as cleaning an aquarium because soap traces can kill fish), salts also are most useful when dissolved. Even the salt thrown on an icy road must dissolve in the top layer of water before it is effective at melting ice. This is because as it dissolves, it lowers the freezing point of the solution and prevents the ice from reforming at the present temperature.

Solutions are discussed in more detail in Chapter 9. Now, let's take closer look at some of the commonly found ionic compounds. You may be surprised at how commonplace they are!

3.4.2 Ionic Compounds in the Home

Photograph by Tylar Mueller

FIGURE 3.14 Many common household products either are ionic compounds or contain them.

A brief tour of the average home, particularly the kitchen and laundry room, easily yields a half-dozen ionic compounds that are part of daily living (Fig. 3.14). Many are known by other nonchemical names!

- Sodium bicarbonate, or "baking soda" ($NaHCO_3$), soothes upset stomachs, deodorizes refrigerators, and acts as a leavening for some baked goods.
- Epsom salts is a common household name for magnesium sulfate heptahydrate ($MgSO_4 \times 7\ H_2O$), and is used as a bath salt or to soak sore muscles. Ancient Egyptians used solid magnesium sulfate to dry tissues and preserve bodies because it readily soaks up water molecules. Using $MgSO_4$ in the garden will provide needed magnesium ions for chlorophyll production in your plants, as well as being a repellent for snails and slugs.
- Sodium hypochlorite (NaClO) is the active ingredient in bleach, a whitening agent for cotton. In a more concentrated form, it reacts with cotton fibers and plastics, leaving holes in the material. Sodium hypochlorite also destroys microorganisms, and thus is a sanitizing agent in countertop cleaners, food preparation equipment, and in bathroom cleaners.
- Household ammonia is a solution of NH_4OH. As a household item, it is most often used for cleaning, particularly windows. It is the active ingredient in many window-cleaning products.
- Sodium hydroxide (NaOH), better known as lye, has long been used for dissolving clogs in drainpipes, and is the active ingredient in oven cleaners. It is a caustic chemical and should be handled with extreme care.
- Sodium nitrate is one of the preservatives in cured meats, such as bacon. Both nitrites and nitrates are added to meats to preserve their color and prevent the growth of bacteria, although some controversy exists over the amount of these substances that is safe for humans to consume. Nitrates and nitrites may cause allergic reactions in some people and are implicated as possible triggers for migraine headaches. However, food preservatives do maintain the freshness of food and prevent such illnesses as botulism poisoning.
- Sulfites are naturally occurring preservatives found in many fruits, including grapes and strawberries. They are used as preservatives in dried fruits and wine because they block the action of enzymes that promote the ripening of fruits. Wines contain sodium sulfite (Na_2SO_3), an ionic compound that occurs naturally in grapes. Many years ago, winemakers discovered that adding more sulfites to wine increased the shelf life of the wine by inhibiting the growth of bacteria in the wine. This allowed wines to be aged for more than 18 months, enhancing their flavor. Like nitrates and nitrites, sulfites also may cause allergic reactions in some people.

And the list goes on. An interesting exercise is to read the labels of the various household products and processed foods to discover which ones have ingredients you now recognize as ionic compounds.

Note that most of these compounds must be dissolved in water for their usefulness to be released. Even the baking soda in cookie dough must have contact with liquid ingredients to do its work as a leavening agent to help the dough rise; it will not leaven a bowl full of dry flour and salt. You will learn more about electrolytic solutions in Chapter 9, and visit them again in Chapters 11 and 12.

3.4.3 Electrolytes in the Body

All living organisms depend on electrolytes to support the processes that sustain life. Ultimately, the substances we need to survive reach the individual cells while dissolved in water. Whether it is in the sap of a tree or the blood of a human being, water is the carrier for nutrients, trace elements, and wastes. Many of these substances exist as ions when they are in solution; they are electrolytes.

As we've seen, the nutritional label on a sports drink indicates that, besides vitamins, it contains sodium and potassium ions. Both play an important role in the transmission of the nerve impulses that tell the heart to beat, the legs to run, and the eyes to move across the page to read this paragraph. Sports drinks also contain magnesium ions, which is also involved in muscle contraction and relaxation. A lack of these ions—or an imbalance in the amounts present—can cause muscle cramps (Fig. 3.15).

© iStock, Inc./GAPS

FIGURE 3.15 Sweating washes electrolytes out of the body. Failing to replace them will limit your endurance and cause muscle cramps.

When you sweat during physical activity, some of these ions are eliminated from the body because they are dissolved in the sweat. This is why sweat tastes so salty. Sports drinks replace these lost electrolytes, allowing you to have greater endurance with less pain for your gain—as long as you don't drink too much. Then another dissolved substance, sugar, becomes a concern!

Another important electrolyte to the proper functioning of bodies is bicarbonate, HCO_3^-, which, as you will discover in Chapter 11, helps your body to maintain the proper acid-base balance.

Calcium ions (Ca^{2+}) not only build strong bones but, like sodium and potassium, facilitate the transmission of electrical signals through nerves.

Then there's Cr(III), trivalent chromium (the *tri*- prefix refers to the number 3 in its 3+ charge). Trivalent chromium is a necessary nutrient that helps to control blood sugar and is contained in chromium picolinate, which is sold as a nutritional supplement to encourage weight loss.

3.4.4 Industrial Products

Flying in an airplane, looking out over the wing, the shiny metal glistens in the sun. As a chemistry student, you might wonder, "Why doesn't the metal airplane skin rust or corrode?"

Airplane skins are made from aluminum metal because it is strong, malleable, and very light. In addition, aluminum has another benefit: it reacts with oxygen in the air to produce aluminum oxide, Al_2O_3. Once a layer of aluminum oxide forms, water molecules can no longer react with the metal. Consequently, it cannot corrode or rust as steel does (Fig. 3.16).

© Maxim Tupikov, 2009, Shutterstock, Inc.

FIGURE 3.16 Airplane wings are protected from corrosion by a layer of aluminum oxide that forms shortly after the wings are manufactured.

Another anticorrosive agent is hexavalent chromium (the prefix *hexa*- refers to the number 6), Cr(VI). It is a key ingredient in paints and coatings where protection against rust and other corrosion is desired. When Cr(VI) comes into contact with most other metals, it forms an oxide layer just as aluminum does. If something scratches this layer, the chromium seeps from the pigment in the coating to heal the damage.

At the same time, hexavalent chromium (Fig. 3.17) is quite toxic, which is why, as noted at the beginning of this chapter, it was the object of a legal settlement.

Perchlorates in Water

More than 22 states have identified a contaminant in their drinking water known as *perchlorate* (ClO_4^-) present in high enough concentrations to cause health issues. Perchlorates have been linked to a condition known as hypothyroidism because they can interfere with the uptake of iodine by the thyroid gland, resulting in decreased levels of thyroid hormones. This interferes with developing nervous systems of very young babies and fetuses, and affects the metabolism of children and adults.

Perchlorate is mostly a man-made polyatomic ion associated with ammonium, potassium, and sodium salts. Perchlorates are used as an oxygen source in solid fuels used in missiles and rockets, and are also components of fireworks. They are extremely soluble in water, and generally inert or unreactive as long as they remain in the water. For this reason, they are difficult to remove from water supplies. Current research using bacteria and ion exchange columns to remove perchlorate are expensive and inefficient at best.

Companies manufacturing or using perchlorates are responsible for their contamination of drinking water. A single chemical company near Las Vegas, Nevada, has been identified as the point source of contamination in the lower Colorado River. The perchlorates from this plant trickled down into the groundwater below this plant, which feeds water into Lake Mead through the Hoover Dam and into the lower Colorado River. From this single source, perchlorates have trickled into the water supplies of Southern California and Arizona, affecting 15 million people.

Courtesy of Nancy Gardner

Perchlorates trickle into the groundwater near Las Vegas (Nevada), find their way to Lake Mead and through the Hoover Dam to the lower Colorado River, and eventually the communities of San Diego (California), Coachella Valley (Palm Springs, California), Tuscon (Arizona), and Phoenix (Arizona).

Government agencies, including the Environmental Protection Agency (EPA), have been debating what concentration of perchlorates in drinking water can be deemed safe. In 1999, the EPA set 4–18 (ppb) as an interim range until further research could be completed. In 2002, they changed their recommendation to 1 ppb. Although the debate continues, perchlorates are still present in the water, and no effective methods for their removal have been identified.

From *Chemical and Engineering News,* August 18, 2003, p. 37.

Courtesy of Paul Lioy

FIGURE 3.17 Hexavalent chromium (Cr^{6+}) and the alchemy of contamination: Chromium in the soil of this New Jersey industrial parking lot has dissolved in a pool of standing water. Because chromium can go into solution and move through soil, chromium pools and blooms (the crystallized chromium left on the surface when the water evaporates) may occur some distance from the original site of contamination.

3.4.5 Ionic Compounds in Our Water

With a 6+ charge, hexavalent chromium is an oxidant—it readily attracts electrons and strips them away from other elements and substances, including those that make up our cells (Fig. 3.17). So it is an extremely toxic form of chromium, causing significant damage to lung tissue when inhaled.

When hexavalent chromium is ingested, as in drinking water, the body converts it to trivalent chromium, Cr(III), the essential nutrient contained in chromium picolinate supplements. However, the by-products of this conversion interfere with the ability of DNA to correctly duplicate itself and direct life processes. It has also been linked with several forms of cancer.

In 1987, Pacific Gas and Electric Company (PG&E) detected the presence of chromium(VI) in the groundwater at a concentration nearly 10 times the amount allowed by law in the State of California. PG&E was using chromium(VI) as an anticorrosive in its cooling towers when it seeped into the groundwater in the Mojave Desert town of Hinkley, California.

The strong anticorrosive properties of chromium(VI) make it an ideal chemical for many industrial uses. However, to the residents of Hinkley, the high concentration of chromium(VI) in their water proved to be the cause of multiple health problems, including brain, kidney, breast, and gastrointestinal cancers, as well as Hodgkin's disease.

The problems in Hinkley, California, were exposed by a young woman, Erin Brockovich, who worked for an attorney. Brockovich, with no formal legal experience, was successful in rallying the town members to sue PG&E for health problems that resulted from their exposure to hexavalent chromium. The courts awarded the residents of Hinkley the largest settlement of its kind in 1993, and ordered PG&E to clean up the hexavalent chromium.

Cr(VI) is just one electrolytic contaminant in our groundwater. As of 2007, 36 cities across the United States were struggling to find a solution to perchlorate contamination.

Other common ionic contaminants include nitrates and phosphates (Fig. 3.18). Phosphates, once a key ingredient in laundry detergents, stimulate the growth of plants and algae that choke freshwater streams and rivers.

Photograph by Tylar Mueller

FIGURE 3.18 Nitrates and phosphates are found in water in many areas of North America. Note the suds in the lower left corner typical of the presence of phosphates.

The nitrate ion, found in fertilizers and food preservatives (see earlier), is another common groundwater contaminant, particularly in the Midwest region of the United States. High nitrate levels are dangerous to infants. Bacteria convert ingested nitrate to nitrite, which then converts hemoglobin in red blood cells to a form that cannot carry oxygen. The result is "blue baby syndrome" because of the lack of oxygen in the infant's bloodstream. Increased nitrate levels have also been linked to non-Hodgkin's lymphoma.

SUMMARY

1. **Evidence of Compound Structure (Section 3.1)**
   - Early chemists identified two laws that govern the formation of compounds. The law of constant composition states that elements combine in specific mass ratios to form compounds. The law of multiple proportions states that the ratios in which elements combine to form compounds are always whole numbers.
2. **Properties of Ionic Compounds (Section 3.2)**
   - Ionic compounds are composed of cations, ions with a positive charge, and anions, ions with a negative charge. The attractive forces between positive and negative charges keep the ions in a lattice structure. Ionic compounds are solid under standard conditions, have high melting and boiling points, often dissolve readily in water, and their solutions conduct electricity.
   - Simple ionic compounds contain a metal cation and a nonmetal anion. They combine in a ratio such that the same number of electrons lost by the cation is gained by the anion. Group numbers can predict these ratios based on valence electrons; for example, Groups 1A and 7A always combine in a 1:1 ratio whereas Groups 1A and 6A combine in a 2:1 ratio. Drawing Lewis structures can be helpful in determining ratios between elements in ionic compounds.
   - Combining ratios can be predicted only for compounds that contain fixed charge metals, which always take the same charge. Fixed charge metals include the Group 1A and Group 2A metals, as well as Al, Ga, Zn, Cd and Ag. Group 1A metals and Ag always have a +1 charge. Group 2A and Zn and Cd always have

a +2 charge, and Al and Ga always have a +3 charge. "Criss-crossing" charges with subscripts, then reducing the subscripts to the simplest terms, is a simple way to derive the proper ratios in a formula.

- All metals not listed as fixed charge metals are variable charge metals, which means they can take more than one charge depending on how many electrons they lose. We can determine their charge based on the charge on the anion in the compound or by analyzing the chemical formula for the compound. The charge on variable charge metals is indicated by a Roman numeral placed in parentheses after the element name; for instance, Fe(III).

3. Naming Ionic Compounds (Section 3.3)

- Binary ionic compounds are named by first listing the metal cation followed by the name of the nonmetal anion with the ending of *-ide*. For instance, table salt (NaCl) is sodium chloride.
- More complex ionic compounds contain polyatomic ions, groups covalently bonded with an overall charge. Each polyatomic ion has a distinctive name; many end in *-ate* or *-ite*. Most polyatomic ions have a negative charge; one exception is ammonium (NH_4^+). If more than one polyatomic ion is present in the compound, they must be put in brackets with any subscripts after the bracket. The names of compounds that contain polyatomic ions begin with the name of the cation placed first, followed by the name of the negatively charged polyatomic ion. In the case of ammonium, it is placed first as it is the cation and the other atoms follow.

4. Ionic Compounds Are Everywhere (Section 3.4)

- Ionic compounds are omnipresent and are often most useful when dissolved in water. When ionic compounds dissolve, they dissociate into ions. The presence of charged particles makes it possible to pass an electric current through the solution; a solution that conducts electricity is called an *electrolyte solution*. Solutions that do not conduct electricity are nonelectrolyte solutions.
- Ionic compounds are found in household products such as drain cleaner, which contains sodium hydroxide (NaOH), or bleach, which contains sodium hypochlorite (NaOCl). Foods may also contain ionic compounds, as they are often added as preservatives such as the nitrates (NO_3^-) or nitrites (NO_2^-) found in pork products, or sulfites SO_3^{2-} found in wine. You may also find baking soda ($NaHCO_3$), in your kitchen as well as NaCl, known as table salt, on your dinner table.
- Our bodies contain electrolytes, including sodium, magnesium, and potassium, that facilitate nerve function. The electrolytes in sports drinks are used to replace electrolytes lost through sweat during physical activity.
- In industry, ionic compounds are used in the construction of airplane skins and as anticorrosive agents.
- Ionic compounds are found in our groundwater. Many are there naturally and are necessary to sustain life, but others are there as a result of human activity and pose a significant health hazard. Hexavalent chromium, perchlorates, and nitrates are three types of ionic compounds about which water districts across the United States are concerned.

KEY TERMS

binary compound, pg. 76
crystal lattice structure, pg. 72
dissociate, pg. 79
electrolytes, pg. 72
fixed charge metals, pg. 73
ionize, pg. 70
law of constant composition, pg. 69
law of multiple proportions, pg. 70
nonelectrolytes, pg. 79
soluble, pg. 72
variable charge metals, pg. 73

REVIEW EXERCISES

Evidence of Compound Structure

1. The law of constant composition states that elements combine in constant mass ratios to form compounds, whereas the law of multiple proportions states that the ratios that elements combine to form new compounds must be in whole-number ratios. Select the law that applies to each of the following statements.

   a. Carbon combines with hydrogen to form C_2H_6, C_3H_8 and other compounds.

   b. A total of 12 grams of carbon combines with 16 grams of oxygen to form a single carbon-oxygen compound.

   c. A total of 18.5 grams of oxygen will not combine with 2.02 grams of hydrogen to form an oxygen-hydrogen compound because this does not represent a whole-number ratio of hydrogen to oxygen.

2. Nitrogen and oxygen combine to make three different compounds: NO, NO_2, and N_2O. Explain how this is possible, and which law describes these phenomena.
3. Phosphorus and oxygen combine to make two compounds: P_2O_5 and P_4O_{10}. Explain the relationship between these compounds. Is the ratio of P and O the same in each?
4. Give the number of atoms for each element in the following compounds.

   a. Na_2SO_4 b. $Al(ClO_4)_3$ c. NH_4NO_3 d. C_2H_5OH

Properties of Ionic Compounds

5. Predict the ion charge when each of the following metals form a cation.

   a. Na b. K
   c. Al d. Zn
   e. Mg f. Ca

6. Predict the ion charge when each of the following elements forms an ion.

   a. Ba b. Cl
   c. O d. N
   e. Li f. Br

7. What charges have the metals formed in each of the following compounds?

   a. $MgCl_2$ b. BaS
   c. AlN d. BeF_2
   e. Ca_3N_2 f. KI

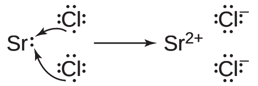
8. How many electrons are transferred from the metal to the anion in each of the following compounds?

   a. $MgCl_2$ b. BaS
   c. AlN d. BeF_2
   e. Ca_3N_2 f. KI

9. Draw the Lewis dot structure and show how many electrons are transferred from the metal to the nonmetal in each of the following compounds:

   a. CaO b. MgF_2 c. Al_2O_3

10. Given the Lewis structure for $SrCl_2$, answer the following questions.



   a. Strontium is losing how many electron(s) in this compound?
   b. What is the total number of electron(s) gained by each chloride ion?
   c. What is the total number of electon(s) transferred in this ionic formula?

11. Determine the charge on the metal in each compound from the chemical formula.

   a. FeS b. Cr_2O_3 c. BaF_2 d. AlN

12. Using the position of each atom in the periodic table, predict the chemical formulas formed between the following elements:

   a. silver and oxygen b. cadmium and chlorine c. aluminum and sulfur

13. Ionic compounds are formed when atoms lose or gain electrons, whereas molecular compounds are formed when atoms share electrons. Which of the following groups contain only ionic compounds?

   a. CaO, NH_3, H_2O b. $CaCO_3$, KI, $LiNO_3$
   c. NH_4Cl, CH_3OH, $CaCl_2$ d. ZnO, Ag_2S, NH_3

Naming Ionic Compounds

14. Predict the charges on the ions they will form and name the ions.

   a. chlorine b. oxygen c. fluorine d. phosphorus

15. Predict the charges on the ions they will form and name the ions.

   a. sodium b. potassium d. magnesium e. aluminum

16. Using the periodic table, in each of the following groups of elements, predict which elements will form identical charges?

   a. chlorine, fluorine, iodine, sulfur b. oxygen, sulfur
   c. phosphorus, nitrogen, arsenic d. magnesium, calcium, beryllium, strontium

17. In each of the following groups, which elements contain only fixed charge metals?

   a. barium, potassium, calcium b. lithium, silver, iron
   c. beryllium, phosphorus, strontium d. calcium, zinc, magnesium

18. Identify the following compounds as containing a fixed charge metal or a variable charge metal.

   a. $CaCl_2$ b. $BiCl_3$
   c. Ag_2S d. KF

19. Determine the charge on each of the following variable charge metals using the anions.

   a. FeS b. Hg_2O
   c. $CrCl_3$ d. CoN

20. Write the names of each of the following ionic compounds.

   a. Ag_2S b. NiO c. SnF_2 d. $PbCl_4$
   e. $BiCl_3$ f. Au_2S g. $CoBr_3$ h. MnO_2

21. Write the formula of each of the following compounds.

    a. strontium oxide b. iron(III) iodide c. nickel(IV) sulfide

    d. lead(II) chloride e. manganese(IV) oxide f. chromium(III) fluoride

22. Identify the charges and write the formulas for each of the following polyatomic ions.

    a. nitrate b. sulfite c. hydrogen phosphate d. carbonate

    e. perchlorate f. ammonium g. hypochlorite h. bicarbonate

23. Use the charges on each of the following polyatomic ions to determine the charge on the variable charge metals in each compound.

    a. $Cr(NO_3)_6$ b. $Au(CN)_3$ c. $Sn(SO_4)_2$ d. $Pb(CO_3)_2$

24. Write names for each of the following compounds that contain polyatomic ions.

    a. $Fe(NO_3)_3$ b. $Mn(OH)_3$ c. $SbSO_4$ d. $Bi_2(CO_3)_3$

    e. $Al(ClO_4)_3$ f. $Ni(NO_2)_4$ g. $PbSO_3$ h. $CoPO_3$

25. Write the formulas for each of the following compounds that contain polyatomic ions.

    a. sodium bicarbonate b. calcium carbonate

    c. potassium sulfite d. iron(II) acetate

    e. manganese(IV) carbonate f. potassium dihydrogen phosphate

    g. tin(II) hydroxide h. magnesium perchlorate

26. Write the formula or provide a name for each of the following compounds.

    a. silver chloride b. cobalt(III) carbonate c. iron(II) hydroxide

    d. AuCN e. sodium hypochlorite f. lead(II) sulfate

    g. manganese(IV) oxide h. bismuth(V) chlorate i. nickel(II) acetate

    j. $KHCO_3$ k. $MgHPO_4$ l. $Mn(CH_3COO)_3$

Ionic Compounds Are Everywhere

27. Which of the following compounds are likely to dissolve in water to form an electrolyte solution?

    a. NaBr b. CO c. NH_4NO_3 d. C_2H_6

28. Ionic compounds form solids composed of an extensive network of ions. The strong attractive forces between the positive and negative charged ions in the crystal are responsible for the large amount of energy required to melt the solid. Which of the following solids would you expect to have a very high melting point?

    a. KCl b. C_2H_6 c. NH_3 d. Na_{2S}

29. Use the Internet to research which metal might be light enough and yet be strong and flexible enough to be used in airplane skins?

    a. Pb b. Fe d. Al

30. Which of the following would you suspect to be made of ionic compounds?

    a. wood b. Epsom salts c. Tums d. paper

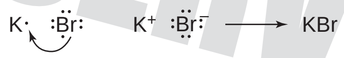
31. At your local market, pick a food label for the following: dried cranberries, wine, bleach, drain cleaner, and packaged bacon. List the ionic compounds found in each.
32. At your local market, review labels of eight different wines. Do they say no added sulfites or just no sulfites? Why might vineyards be misrepresenting the sulfite content in their wine?
33. Do an Internet search to learn about allergies to sulfites and nitrates; then answer the following questions about each.
    a. What are common symptoms of these allergies?
    b. List six different foods or beverages that contain nitrates or sulfites.
    c. Which wine has more sulfites in it: white or red?
    d. Which meat has more sulfites in it: beef, pork, chicken, or fish?
34. Do an Internet search about hexavalent chromium; then answer the following questions.
    a. Is this contaminant present in small amounts in your drinking water?
    b. What is this ion used for?
    c. How did Erin Brockovich learn about hexavalent chromium and its health effects?
35. List 15 of the ionic compounds found in your drinking water and their concentrations in parts per million (ppm) or parts per billion (ppb). This may be provided by your water company Web site on your March water bill.
36. Where do perchlorates come from? Why can't they be filtered out of drinking water?
37. Why are perchlorates in drinking water harmful?

SOLUTIONS TO EXAMPLES

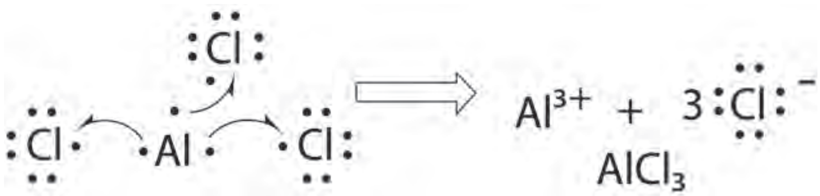
(pg. 73)

a. Bromine (Br) needs just one electron to get an octet, and potassium (K) has one valence electron to give up.

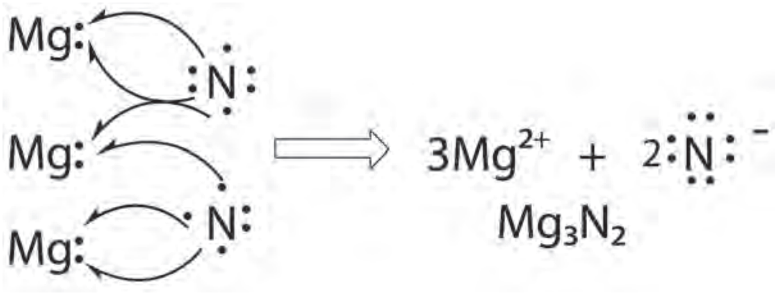
Thus, electrons lost equals electrons gained.



b. Chlorine (Cl) has seven valence electrons and needs one electron to get an octet. Aluminum (Al) has three electrons to lose; thus, it will require three chlorine atoms to remove aluminum's three valence electrons.



c. Magnesium (Mg) is a metal and has two valence electrons it can lose, whereas nitrogen (N) has five valence electrons and needs three to complete an octet.



Nitrogen takes two valence electrons from magnesium but needs one more, so we add another magnesium. Nitrogen now has eight electrons, but the second magnesium has one electron left over.

We add a second nitrogen; it takes the one electron left from magnesium. However, it needs two more electrons to get eight electrons, so we add yet another magnesium. Nitrogen takes the two electrons from this magnesium, and it now has eight.

Three magnesium atoms have lost two electrons each for a total of six electrons lost. Two nitrogen atoms have each gained three electrons, for a total of six electrons gained. The formula is thus Mg_3N_2.

(pg. 75)

a. The anion is chloride and has a charge of −1. Because we have three chloride ions, each with a −1 charge, the total negative charge will be −3. There is only one Fe ion to neutralize this charge, so the charge on Fe must be +3: iron(III) chloride.

b. The anion is sulfur and has a charge of −2. Because we have two sulfur ions, the total negative charge in this compound is −4. Therefore, the only Sn ion must have a charge of +4: tin(IV) sulfide.

c. The anion is oxygen and has a charge of −2. Because we have three oxygen ions, each with a −2 charge, the total negative charge will be −6. There are two Cr ions, therefore, each must have a charge of +3 to neutralize the −6 charge: chromium(III) oxide.

(pg. 78)

a. The charges on Na^{+} and CN^{-} ions are equal and opposite. Therefore, the ratio of ions is 1:1. So the formula for sodium cyanide, an extremely toxic chemical used in gold mining, is NaCN.

b. The charges on Ba^{2+} and CO_3^{2-} are equal and opposite. Therefore, the ratio of ions is also 1:1. The resulting formula is $BaCO_3$, barium carbonate, used in rat poison.

c. Calcium has lost two electrons, it has a +2 charge. Nitrate (NO_3^{-}) has a charge of −1. Therefore, two nitrate ions are needed to compensate for the +2 charge on calcium.

Note that the resulting formula for the fertilizer, calcium nitrate, is written as $Ca(NO_3)_2$, not $CaNO_{32}$. The nitrate ions are put in parentheses to indicate that there are 2 nitrate ions present, and not 1 nitrogen atom and 32 oxygen atoms!

SOLUTIONS TO CONCEPT CHECKS

(pg. 74)

a. Na is in group 1A and, therefore; forms a +1 ion. Sulfur is in Group 6A, needs two electrons to complete an octet, and therefore, will gain two electrons and have a charge of −2.

$$Na^{+} \; S^{2-} \longrightarrow Na_2S_1 \longrightarrow Na_2S$$

b. Al is in group 3A and, therefore, forms a +3 ion. Bromine is in Group 7A and, therefore, will gain one electron and have a charge of −1.

$$Al^{3+} \; Br^{-} \longrightarrow Al_1Br_3 \longrightarrow AlBr_3$$

(pg. 75)

a. The anion is iodine, and has a charge of −1. Because we have two iodine ions, each with a −1 charge, the total negative charge will be −2. There is only one Pb ion to neutralize this charge, so the charge on Pb must be +2: lead(II) iodide.

b. The anion is nitrogen, which has a charge of −3. Because there is one nitrogen ion, the total negative charge in this compound is −3. Therefore the only Co ion must have a charge of +3: cobalt(III) nitride

c. The anion is oxygen and has a charge of −2. Because we have two oxygen ions, each with a −2 charge, the total negative charge will be −4. There is only one Ni ion, so the charge on Ni must be +4: nickel(IV) oxide.

(pg. 78)

a. $(NH_4)_2S$, ammonium sulfide: The charge on ammonium is +1, and the charge on sulfur is −2. Two ammonium ions are needed to compensate for the −2 charge on sulfur.

b. Na_3PO_4, sodium phosphate: The charge on sodium is +1 and the charge on phosphate is −3. Three sodium ions are needed to compensate for the −3 charge on phosphate. Because sodium is not a polyatomic ion, do not put it in parentheses.

c. $Al_2(CO_3)_3$, aluminum carbonate:The charge on aluminum is +3 and the charge on carbonate is −2. Because one charge is +3 and the other −2, we will have a total of six electrons transferred, which will require two aluminum ions and three carbonate ions.