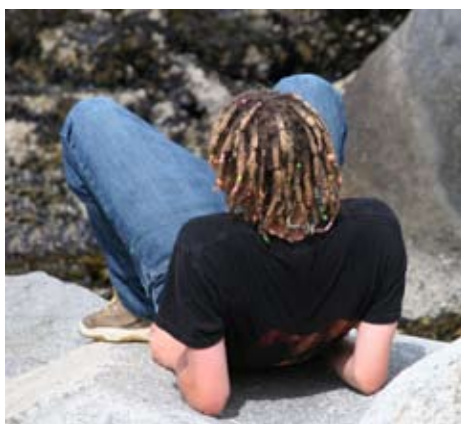


# CHAPTER 7

## OXIDATION-REDUCTION REACTIONS

In many important chemical reactions, electrons are transferred from atom to atom. We are surrounded by these reactions, commonly called oxidation-reduction (or redox) reactions, inside and out. Let's consider a typical "new millennium" family, sitting around the dining room table after the dishes have been cleared.

Mom, a computer programmer, is typing away on her portable computer. She's very anxious to see if the idea she got while on her drive home will fix a glitch in the accounting program at work. Christine, the thirteen-year-old, is fighting the bad guys on her video game. The electric currents from the batteries that power the computer and the game are generated by oxidation-reduction reactions. Buddy, who's 15, has recently become interested in studying Eastern Philosophy. Just now, he's gazing meditatively out into space, but redox reactions are powering his activity as well; they are important for the storage and release of energy in all our bodies. Dad's an engineer in charge of blasting a tunnel under the bay for the city's new rapid transit project. Each of the explosions that he triggers is created by oxidation-reduction reactions. The silverware he has just cleared from the table is tarnishing due to redox reactions, and the combustion of natural gas in the heater warming the room is a redox reaction as well...



Oxidation-reduction reactions power ourselves and many of our tools and toys.

7.1	An Introduction to Oxidation-Reduction Reactions
7.2	Oxidation Numbers
7.3	Types of Chemical Reactions
7.4	Voltaic Cells

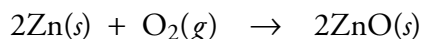
### Review Skills

The presentation of information in this chapter assumes that you can already perform the tasks listed below. You can test your readiness to proceed by answering the Review Questions at the end of the chapter. This might also be a good time to read the Chapter Objectives, which precede the Review Questions.

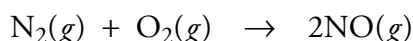
- Determine the charge on a monatomic ion in an ionic formula. (Section 4.5)
- Determine the formulas, including the charges, for common polyatomic ions. (Section 4.5)
- Identify a chemical formula as representing an element, a binary ionic compound, an ionic compound with one or two polyatomic ions, or a molecular compound. (Section 6.3)

## 7.1 An Introduction to Oxidation-Reduction Reactions

Zinc oxide is a white substance used as a pigment in rubber, sun-blocking ointments, and paint. It is added to plastics to make them less likely to be damaged by ultraviolet radiation and is also used as a dietary supplement. It can be made from the reaction of pure zinc and oxygen:



In a similar reaction that occurs every time you drive your car around the block, nitrogen monoxide is formed from some of the nitrogen and oxygen that are drawn into your car's engine:

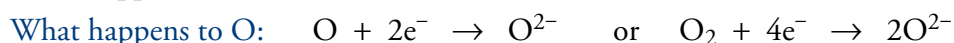
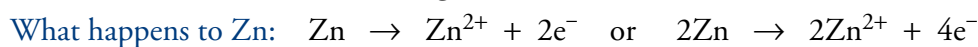
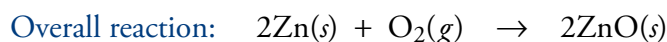


This nitrogen monoxide in turn produces other substances that lead to acid rain and help create the brown haze above our cities.

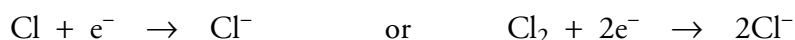
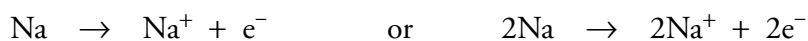
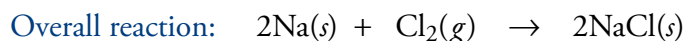
When an element, such as zinc or nitrogen, combines with oxygen, chemists say it is oxidized (or undergoes oxidation). They also use the term oxidation to describe many similar reactions that do not have oxygen as a reactant. This section explains the meaning of oxidation and shows why oxidation is coupled with a corresponding chemical change called reduction.

### Oxidation, Reduction, and the Formation of Binary Ionic Compounds

Zinc oxide is an ionic compound made up of zinc cations,  $\text{Zn}^{2+}$ , and oxide anions,  $\text{O}^{2-}$ . When uncharged zinc and oxygen atoms react to form zinc oxide, electrons are transferred from the zinc atoms to the oxygen atoms to form these ions. Each zinc atom loses two electrons, and each oxygen atom gains two electrons.



As we saw in Chapter 4, this transfer of electrons from metal atoms to nonmetal atoms is the general process for the formation of any binary ionic compound from its elements. For example, when sodium chloride is formed from the reaction of metallic sodium with gaseous chlorine, each sodium atom loses an electron, and each chlorine atom gains one.

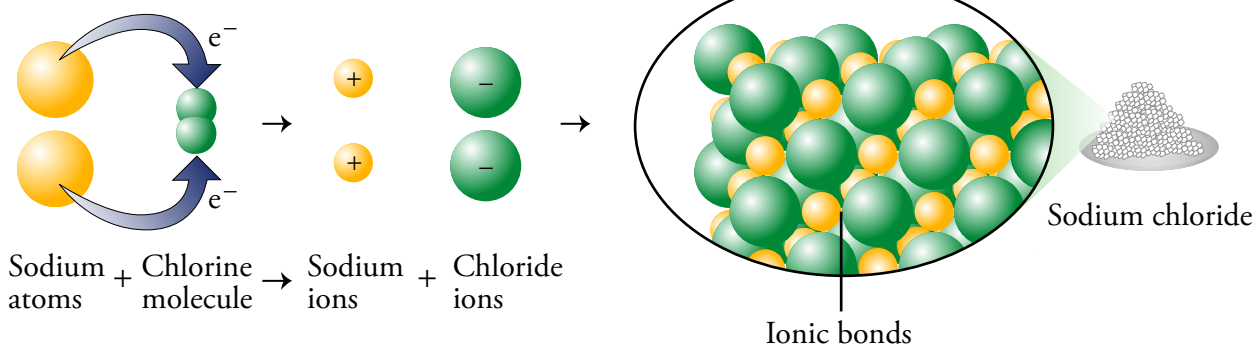


The reactions that form sodium chloride and zinc oxide from their elements are so similar that chemists find it useful to describe them using the same terms. Zinc atoms that lose electrons in the reaction with oxygen are said to be oxidized; therefore, when sodium atoms undergo a similar change in their reaction with chlorine, chemists say they too are *oxidized*, even though no oxygen is present. According to the modern convention, any chemical change in which an element loses electrons is called an **oxidation** (Figure 7.1).

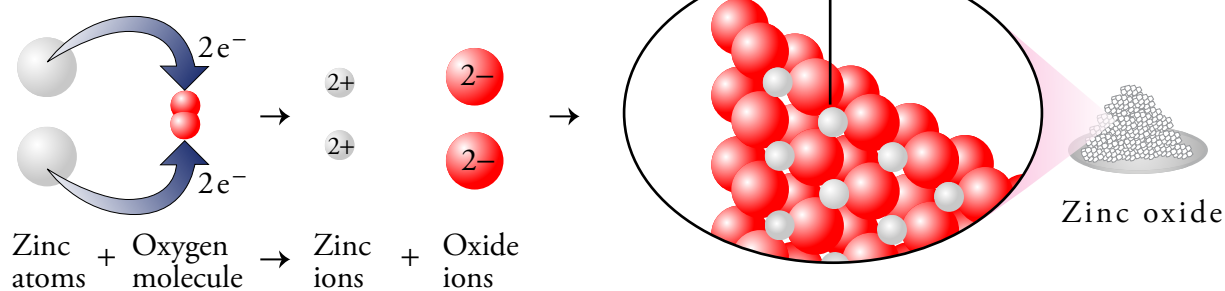
#### OBJECTIVE 2

#### OBJECTIVE 2

## Formation of NaCl



## Oxidation of zinc

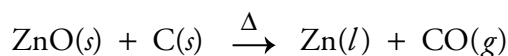


## OBJECTIVE 2

Figure 7.1

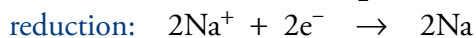
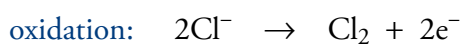
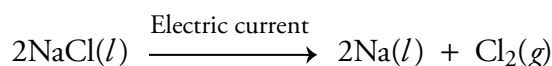
Oxidation and the Formation of Binary Ionic Compounds

The concept of reduction has undergone a similar evolution. At high temperature, zinc oxide, ZnO, reacts with carbon, C, to form molten zinc and carbon monoxide gas.



Bonds between zinc atoms and oxygen atoms are lost in this reaction, so chemists say the zinc has been *reduced*. Like the term oxidation, the term reduction has been expanded to include similar reactions, even when oxygen is not a participant. The zinc ions in zinc oxide have a +2 charge, and the atoms in metallic zinc are uncharged. Thus, in the conversion of zinc oxide to metallic zinc, each zinc ion must gain two electrons. According to the modern definition, any chemical change in which an element gains electrons is called a **reduction**. (Yes, *reduction* means a gain of electrons.) Because this can be confusing, some people use a memory aid to remember what oxidation and reduction mean in terms of the electron transfer. One device is the phrase *oil rig* which stands for oxidation *is* loss (of electrons) and reduction *is* gain (of electrons).

When an electric current passes through molten sodium chloride, the sodium ions,  $\text{Na}^+$ , are converted to uncharged sodium atoms, and the chloride ions,  $\text{Cl}^-$ , are converted to uncharged chlorine molecules,  $\text{Cl}_2$ . Because sodium ions gain one electron each, we say they are reduced. Chloride ions lose one electron each, so they are oxidized.



Electrons are rarely found unattached to atoms. Thus, for one element or compound

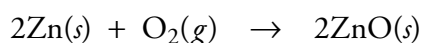
Oxidation  
LOSS  
Reduction  
GAIN

to lose electrons and be oxidized, another element or compound must be there to gain the electrons and be reduced. In other words, oxidation (loss of electrons) must be accompanied by reduction (gain of electrons). In the reaction that forms ZnO from Zn and O<sub>2</sub>, the uncharged zinc atoms cannot easily lose electrons and be oxidized unless something such as oxygen is there to gain the electrons and be reduced. In the reaction that converts NaCl to Na and Cl<sub>2</sub>, the chloride ions can lose electrons and be oxidized because the sodium ions are available to gain the electrons and be reduced.

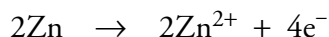
By similar reasoning, we can say that reduction requires oxidation. Because electrons are not likely to be found separated from an element or compound, a substance cannot gain electrons and be reduced unless there is another substance that is able to transfer the electrons and be oxidized. Oxidation and reduction take place together.

Reactions in which electrons are transferred, resulting in oxidation and reduction, are called **oxidation-reduction reactions**. Because the term *oxidation-reduction* is a bit cumbersome, we usually call these reactions **redox reactions**.

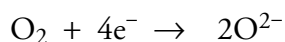
Even though the oxidation and reduction of a redox reaction take place simultaneously, each making the other possible, chemists often have reason to describe the reactions separately. The separate oxidation and reduction equations are called **half-reactions**. For example, in the reaction:



the oxidation half-reaction is



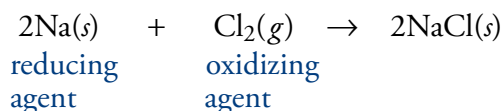
and the reduction half-reaction is



Because the zinc atoms lose the electrons that make it possible for the oxygen atoms to gain electrons and be reduced, the zinc is called the reducing agent. A **reducing agent** is a substance that loses electrons, making it possible for another substance to gain electrons and be reduced. *The oxidized substance is always the reducing agent.*

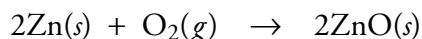
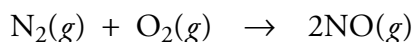
Because the oxygen atoms gain electrons and make it possible for the zinc atoms to lose electrons and be oxidized, the oxygen is called the oxidizing agent. An **oxidizing agent** is a substance that gains electrons, making it possible for another substance to lose electrons and be oxidized. *The reduced substance is always the oxidizing agent.*

In the reaction that forms sodium chloride from the elements sodium and chlorine, sodium is oxidized, and chlorine is reduced. Because sodium makes it possible for chlorine to be reduced, sodium is the reducing agent in this reaction. Because chlorine makes it possible for sodium to be oxidized, chlorine is the oxidizing agent.

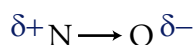


## Oxidation-Reduction and Molecular Compounds

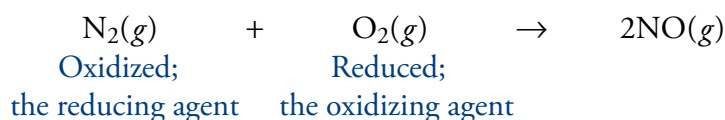
The oxidation of nitrogen to form nitrogen monoxide is very similar to the oxidation of zinc to form zinc oxide.



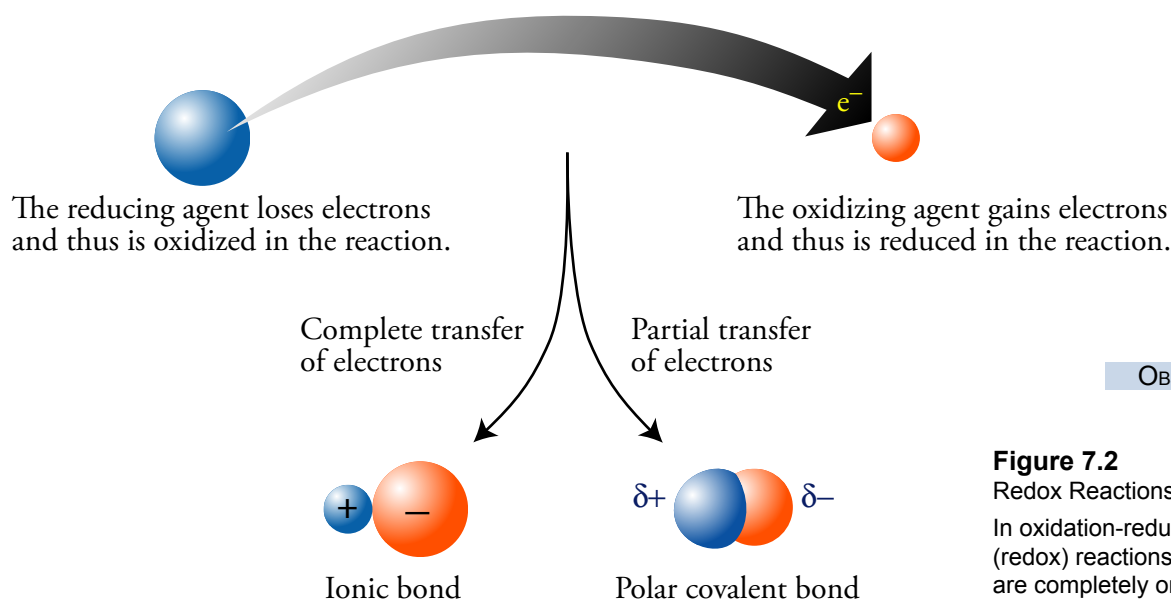
The main difference between these reactions is that as the nitrogen monoxide forms, electrons are not transferred completely, as occurs in the formation of zinc oxide, and no ions are formed. Nitrogen monoxide is a molecular compound, and the bonds between the nitrogen and the oxygen are covalent bonds, in which electrons are shared. Because the oxygen atoms attract electrons more strongly than nitrogen atoms, there is a partial transfer of electrons from the nitrogen atoms to the oxygen atoms in the formation of NO molecules, leading to polar bonds with a partial negative charge on each oxygen atom and a partial positive charge on each nitrogen atom.



Because the reactions are otherwise so much alike, chemists have expanded the definition of oxidation-reduction reactions to include *partial* as well as *complete* transfer of electrons. Thus **oxidation** is defined as the complete or partial loss of electrons, **reduction** as the complete or partial gain of electrons. The nitrogen in the reaction that forms NO from nitrogen and oxygen is oxidized, and the oxygen is reduced. Because the nitrogen makes it possible for the oxygen to be reduced, the nitrogen is the reducing agent. The oxygen is the oxidizing agent (Figure 7.2).



Special Topic 7.1 *Oxidizing Agents and Aging* describes how oxidizing agents might play a role in aging and how a good healthy diet might slow the aging process.



**Figure 7.2**  
Redox Reactions  
In oxidation-reduction (redox) reactions, electrons are completely or partially transferred.

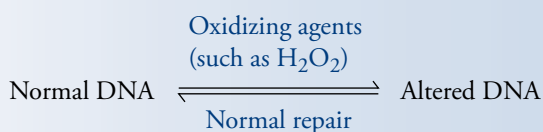
OBJECTIVE 2

OBJECTIVE 2

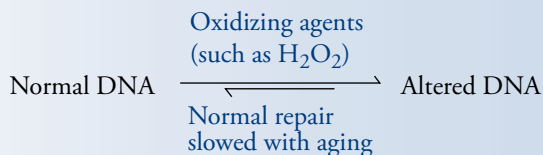
OBJECTIVE 2

## SPECIAL TOPIC 7.1 Oxidizing Agents and Aging

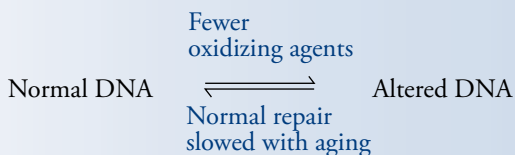
In some of the normal chemical reactions that take place in the human body, strong oxidizing agents, such as hydrogen peroxide, are formed. These highly reactive substances cause chemical changes in cell DNA that can be damaging unless the changes are reversed. Fortunately, in healthy cells, normal repair reactions occur that convert the altered DNA back to its normal form



The repair mechanisms are thought to slow down with age. Some medical researchers believe that this slowing down of DNA repair is connected to certain diseases associated with aging, such as cancer, heart disease, cataracts, and brain dysfunction.



Substances called *antioxidants* that are found in food react with oxidizing agents (such as hydrogen peroxide) and thus remove them from our system. This is believed to slow the alteration of DNA, so the slower rate of normal repair can balance it.



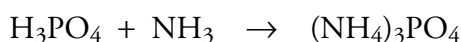
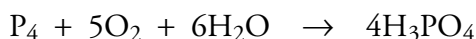
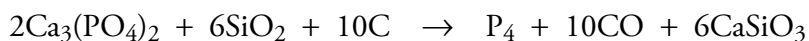
Vitamins C and E are antioxidants, and foods that contain relatively high amounts of them are considered important in slowing some of the medical problems that come from aging. Five servings of fruits and vegetables per day are thought to supply enough antioxidants to provide reasonable protection from the damage done by oxidizing agents.



Five portions of fruits and vegetables per day have enough of the antioxidants vitamin C and vitamin E to mitigate some of the problems of aging.

## 7.2 Oxidation Numbers

Phosphates, like ammonium phosphate, are important components of fertilizers used to stimulate the growth of agricultural crops and to make our gardens green. Their commercial synthesis requires elemental phosphorus, which can be acquired by heating phosphate rock (containing calcium phosphate) with sand (containing silicon dioxide) and coke (a carbon-rich mixture produced by heating coal). This method for isolating phosphorus, called the *furnace process*, is summarized in the first equation below. The other equations show how phosphorus can be converted into ammonium phosphate.



Are these reactions oxidation-reduction reactions? Are electrons transferred? Simply reading a chemical equation does not always tell us whether oxidation and reduction have occurred, so chemists have developed a numerical system to help identify a reaction as redox. For redox reactions, this system also shows us which element is oxidized, which is reduced, what the oxidizing agent is, and what the reducing agent is.

The first step in this system is to assign an **oxidation number** to each atom in the reaction equation. As you become better acquainted with the procedure, you will gain a better understanding of what the numbers signify, but for now, just think of them as tools for keeping track of the flow of electrons in redox reactions. Oxidation numbers are also called **oxidation states**.

If any element undergoes a change of oxidation number in the course of a reaction, the reaction is a redox reaction. If an element's oxidation number increases in a reaction, that element is oxidized. If an element's oxidation number decreases in a reaction, that element is reduced. The reactant containing the element that is oxidized is the reducing agent. The reactant containing the element that is reduced is the oxidizing agent (Table 7.1).

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

**Table 7.1**

Questions Answered by the Determination of Oxidation Numbers

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

Question	Answer
Is the reaction redox?	If any atoms change their oxidation number, the reaction is redox.
Which element is oxidized?	The element that increases its oxidation number is oxidized.
Which element is reduced?	The element that decreases its oxidation number is reduced.
What's the reducing agent?	The reactant that contains the element that is oxidized is the reducing agent.
What's the oxidizing agent?	The reactant that contains the element that is reduced is the oxidizing agent.

Study Sheet 7.1 on the next page describes how you can assign oxidation numbers to individual atoms.

## Sample Study Sheet 7.1

### Assignment of Oxidation Numbers

#### OBJECTIVE 3

**TIP-OFF** You are asked to determine the oxidation number of an atom, or you need to assign oxidation numbers to atoms to determine whether a reaction is a redox reaction, and if it is, to identify which element is oxidized, which is reduced, what the oxidizing agent is, and what the reducing agent is.

#### GENERAL STEPS

Use the following guidelines to assign oxidation numbers to as many atoms as you can. (Table 7.2 provides a summary of these guidelines with examples.)

- The oxidation number for each atom in a pure element is zero.
- The oxidation number of a monatomic ion is equal to its charge.
- When fluorine atoms are combined with atoms of other elements, their oxidation number is  $-1$ .
- When oxygen atoms are combined with atoms of other elements, their oxidation number is  $-2$ , except in peroxides, such as hydrogen peroxide,  $\text{H}_2\text{O}_2$ , where their oxidation number is  $-1$ . (*There are other exceptions that you will not see in this text.*)
- The oxidation number for each hydrogen atom in a molecular compound or a polyatomic ion is  $+1$ .

If a compound's formula contains one element for which you cannot assign an oxidation number using the guidelines listed above, calculate the oxidation number according to the following rules.

- The sum of the oxidation numbers for the atoms in an uncharged formula is equal to zero.
- The sum of the oxidation numbers for the atoms in a polyatomic ion is equal to the overall charge on the ion.

**EXAMPLE** See Example 7.1.

**Table 7.2**  
Guidelines for Assigning Oxidation Numbers

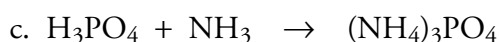
	Oxidation number	Examples	Exceptions
Pure element	0	The oxidation number for each atom in Zn, $\text{H}_2$ , and $\text{S}_8$ is zero.	none
Monatomic ions	charge on ion	Cd in $\text{CdCl}_2$ is $+2$ . Cl in $\text{CdCl}_2$ is $-1$ . H in LiH is $-1$ .	none
Fluorine in the combined form	$-1$	F in $\text{AlF}_3$ is $-1$ . F in $\text{CF}_4$ is $-1$ .	none
Oxygen in the combined form	$-2$	O in ZnO is $-2$ . O in $\text{H}_2\text{O}$ is $-2$ .	O is $-1$ in peroxides, such as $\text{H}_2\text{O}_2$
Hydrogen in the combined form	$+1$	H in $\text{H}_2\text{O}$ is $+1$ .	H is $-1$ when combined with a metal.



Example 7.1 shows how we can use our new tools.

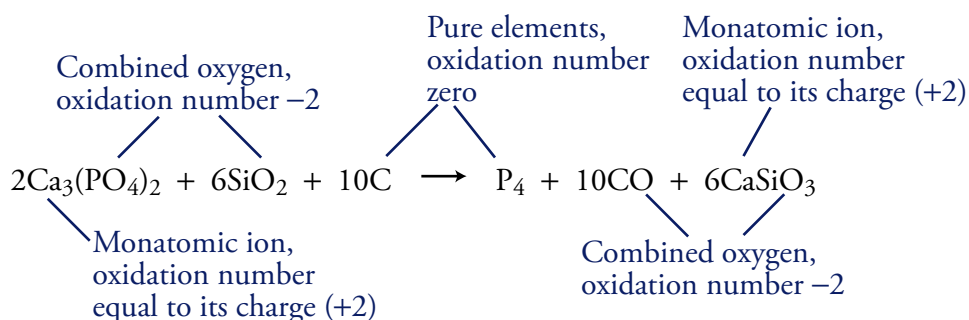
### EXAMPLE 7.1 - Oxidation Numbers and Redox Reactions

The following equations represent the reactions that lead to the formation of ammonium phosphate for fertilizers. Determine the oxidation number for each atom in the formulas. Decide whether each reaction is a redox reaction, and if it is, identify what element is oxidized, what is reduced, what the oxidizing agent is, and what the reducing agent is.



*Solution*

- a. The first step is to determine the oxidation number for each atom in the reaction. Let's consider the first equation above:



Because the sum of the oxidation numbers for the atoms in an uncharged molecule is zero, the oxidation number of the carbon atom in CO is +2:

$$(\text{ox \# C}) + (\text{ox \# O}) = 0$$

$$(\text{ox \# C}) + (-2) = 0$$

$$(\text{ox \# C}) = +2$$

Using a similar process, we can assign a +4 oxidation number to the silicon atom in SiO<sub>2</sub>:

$$(\text{ox \# Si}) + 2(\text{ox \# O}) = 0$$

$$(\text{ox \# Si}) + 2(-2) = 0$$

$$(\text{ox \# Si}) = +4$$

Calcium phosphate, Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>, is an ionic compound that contains monatomic calcium ions, Ca<sup>2+</sup>, and polyatomic phosphate ions, PO<sub>4</sub><sup>3-</sup>. The oxidation number of each phosphorus atom can be determined in two ways. The following shows how it can be done considering the whole formula.

$$3(\text{ox \# Ca}) + 2(\text{ox \# P}) + 8(\text{ox \# O}) = 0$$

$$3(+2) + 2(\text{ox \# P}) + 8(-2) = 0$$

$$(\text{ox \# P}) = +5$$

OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

The oxidation number for the phosphorus atom in  $\text{PO}_4^{3-}$  is always the same, no matter what the cation is that balances its charge. Thus we could also have determined the oxidation number of each phosphorus atom by considering the phosphate ion separately from the calcium ion.

$$(\text{ox \# P}) + 4(\text{ox \# O}) = -3$$

$$(\text{ox \# P}) + 4(-2) = -3$$

$$(\text{ox \# P}) = +5$$

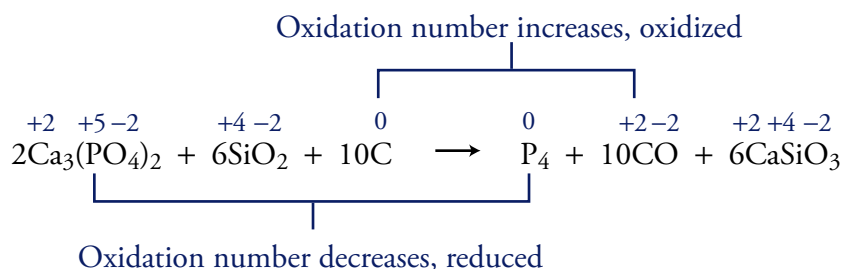
The silicon atoms in  $\text{CaSiO}_3$  must have an oxidation number of +4.

$$(\text{ox \# Ca}) + (\text{ox \# Si}) + 3(\text{ox \# O}) = 0$$

$$(+2) + (\text{ox \# Si}) + 3(-2) = 0$$

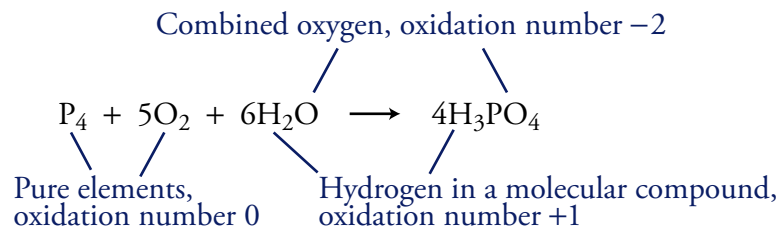
$$(\text{ox \# Si}) = +4$$

The oxidation numbers for the individual atoms in the first reaction are below.



Phosphorus atoms and carbon atoms change their oxidation numbers, so **the reaction is redox**. Each phosphorus atom changes its oxidation number from +5 to zero, so the **phosphorus atoms in  $\text{Ca}_3(\text{PO}_4)_2$  are reduced, and  $\text{Ca}_3(\text{PO}_4)_2$  is the oxidizing agent**. Each carbon atom changes its oxidation number from zero to +2, so the **carbon atoms are oxidized, and carbon is the reducing agent**.

b. Now, let's consider the second reaction.



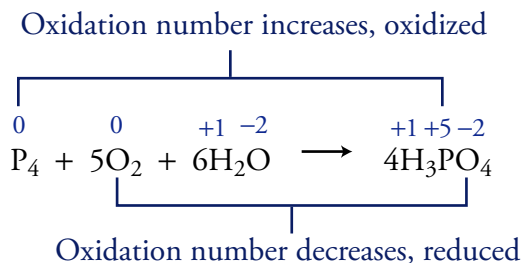
The following shows how we can determine the oxidation number of the phosphorus atom in  $\text{H}_3\text{PO}_4$ :

$$3(\text{ox \# H}) + (\text{ox \# P}) + 4(\text{ox \# O}) = 0$$

$$3(+1) + (\text{ox \# P}) + 4(-2) = 0$$

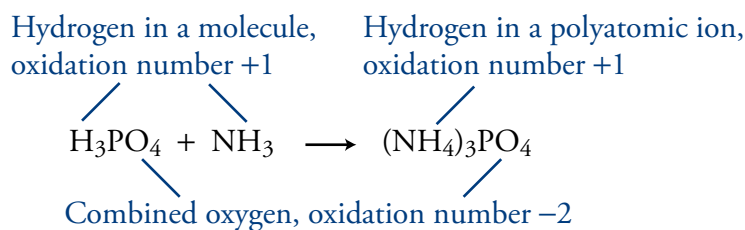
$$(\text{ox \# P}) = +5$$

The oxidation numbers for the individual atoms in the second reaction are below.



Phosphorus atoms and oxygen atoms change their oxidation numbers, so **the reaction is redox**. Each phosphorus atom changes its oxidation number from zero to +5, so the **phosphorus atoms in  $\text{P}_4$  are oxidized, and  $\text{P}_4$  is the reducing agent**. Each oxygen atom in  $\text{O}_2$  changes its oxidation number from zero to  $-2$ , so the **oxygen atoms in  $\text{O}_2$  are reduced, and  $\text{O}_2$  is the oxidizing agent**.

c. Finally, let's consider the third reaction.



We know from part (b) that the oxidation number of the phosphorus atoms in  $\text{H}_3\text{PO}_4$  is +5.

The oxidation number of the nitrogen atom in  $\text{NH}_3$  is calculated below.

$$(\text{ox \# N}) + 3(\text{ox \# H}) = 0$$

$$(\text{ox \# N}) + 3(+1) = 0$$

$$(\text{ox \# N}) = -3$$

We can determine the oxidation number of each nitrogen atom in  $(\text{NH}_4)_3\text{PO}_4$  in two ways, either from the whole formula or from the formula for the ammonium ion alone.

$$3(\text{ox \# N}) + 12(\text{ox \# H}) + (\text{ox \# P}) + 4(\text{ox \# O}) = 0$$

$$3(\text{ox \# N}) + 12(+1) + (+5) + 4(-2) = 0$$

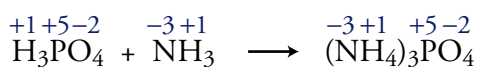
$$(\text{ox \# N}) = -3$$

$$\text{or } (\text{ox \# N}) + 4(\text{ox \# H}) = +1$$

$$(\text{ox \# N}) + 4(+1) = +1$$

$$(\text{ox \# N}) = -3$$

The oxidation numbers for the individual atoms in this reaction are below.



None of the atoms change their oxidation number, so **the reaction is not redox**.

### EXERCISE 7.1 - Oxidation Numbers

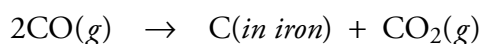
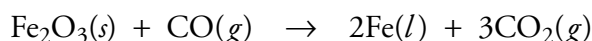
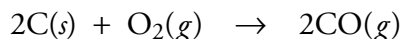
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

In one part of the steel manufacturing process, carbon is combined with iron to form pig iron. Pig iron is easier to work with than pure iron because it has a lower melting point (about 1130 °C compared to 1539 °C for pure iron) and is more pliable. The following equations describe its formation. Determine the oxidation number for each atom in the formulas. Decide whether each reaction is a redox reaction, and if it is, identify what is oxidized, what is reduced, what the oxidizing agent is, and what the reducing agent is.



Equations for redox reactions can be difficult to balance, but your ability to determine oxidation numbers can help. You can find a description of the process for balancing redox equations at the textbook's Web site.

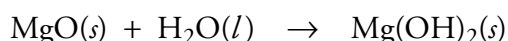
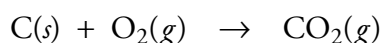
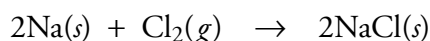
## 7.3 Types of Chemical Reactions

Chemists often group reactions into general categories, rather than treating each chemical change as unique. For example, you saw in Chapter 5 that many chemical changes can be assigned to the category of precipitation reactions. Understanding the general characteristics of this type of reaction helped you to learn how to predict products and write equations for specific precipitation reactions. You developed similar skills in Chapter 6 for neutralization reactions. Because several types of chemical reactions can also be redox reactions, we continue the discussion of types of chemical reactions here.

### Combination Reactions

OBJECTIVE 7

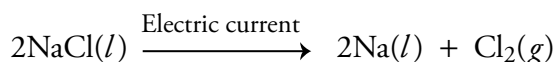
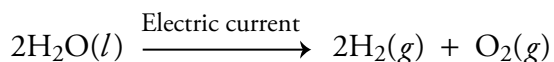
In **combination reactions**, two or more elements or compounds combine to form one compound. Combination reactions are also called synthesis reactions. The following are examples of combination reactions.



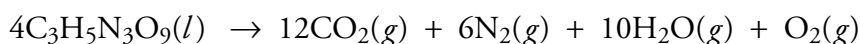
Two of the above combination reactions are also redox reactions. Can you tell which one is not?

## Decomposition Reactions

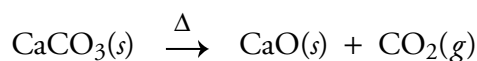
In **decomposition reactions**, one compound is converted into two or more simpler substances. The products can be either elements or compounds. For example, when an electric current is passed through liquid water or molten sodium chloride, these compounds decompose to form their elements.



Another example is the decomposition of nitroglycerin. When this compound decomposes, it produces large amounts of gas and heat, making nitroglycerin a dangerous explosive.

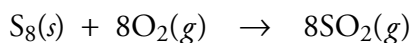
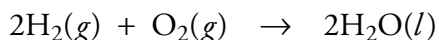
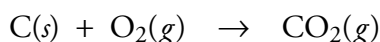


As is true of combination reactions, not all decomposition reactions are redox reactions. The following equation represents a decomposition reaction that is not a redox reaction.

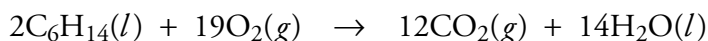
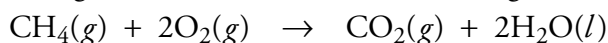


## Combustion Reactions

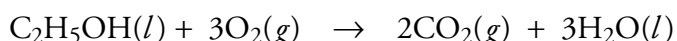
A log burns in the fireplace as a result of a **combustion reaction**, a redox reaction in which oxidation is very rapid and is accompanied by heat and usually light. The combustion reactions that you will be expected to recognize have oxygen,  $\text{O}_2$ , as one of the reactants. For example, the elements carbon, hydrogen, and sulfur react with oxygen in combustion reactions.



When any substance that contains carbon is combusted (or burned) completely, the carbon forms carbon dioxide. When a substance that contains hydrogen is burned completely, the hydrogen forms water. Therefore, when hydrocarbons found in natural gas, gasoline, and other petroleum products burn completely, the only products are  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . The equations below represent the combustion reactions for methane, the primary component of natural gas, and hexane, which is found in gasoline.



The complete combustion of a substance, such as ethanol,  $\text{C}_2\text{H}_5\text{OH}$ , that contains carbon, hydrogen, and oxygen also yields carbon dioxide and water.



When any substance that contains sulfur burns completely, the sulfur forms sulfur dioxide. For example, when methanethiol,  $\text{CH}_3\text{SH}$ , burns completely, it forms carbon dioxide, water, and sulfur dioxide. Small amounts of this strong-smelling substance are

### OBJECTIVE 7



Dynamite, which has nitroglycerine as its active component, can be used to blast tunnels through solid rock.

### OBJECTIVE 7

### OBJECTIVE 8



Combustion reactions include oxygen as a reactant and are accompanied by heat and usually light.

### OBJECTIVE 8

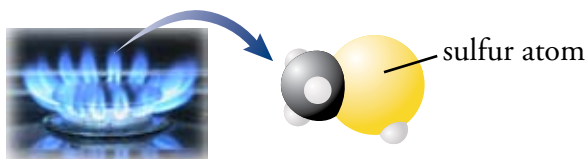
added to natural gas to give the otherwise odorless gas a smell that can be detected in case of leaks (Figure 7.3).



**Figure 7.3**

**Odor of Natural Gas**

The methane thiol added to natural gas warns us when there is a leak.



The following sample study sheet lists the steps for writing equations for combustion reactions.

## Sample Study Sheet 7.2

### Writing Equations for Combustion Reactions

OBJECTIVE 8

**TIP-OFF** You are asked to write an equation for the complete combustion of a substance composed of one or more of the elements carbon, hydrogen, oxygen, and sulfur.

**GENERAL STEPS**

**STEP 1** Write the formula for the substance combusted.

**STEP 2** Write  $\text{O}_2(g)$  for the second reactant.

**STEP 3** Predict the products using the following guidelines.

- If the compound contains carbon, one product will be  $\text{CO}_2(g)$ .
- If the compound contains hydrogen, one product will be  $\text{H}_2\text{O}(l)$ . (Even though water may be gaseous when it is first formed in a combustion reaction, we usually describe it as a liquid in the equation. By convention, we usually describe the state of each reactant and product as their state at room temperature and pressure. When water returns to room temperature, it is a liquid.)
- If the compound contains sulfur, one product will be  $\text{SO}_2(g)$ .
- Any oxygen in the combusted substance would be distributed between the products already mentioned.

**STEP 4** Balance the equation.

**EXAMPLE** See Example 7.2.

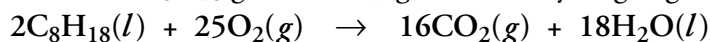
### EXAMPLE 7.2 - Combustion Reactions

OBJECTIVE 8

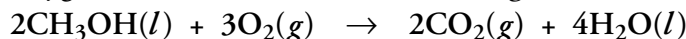
Write balanced equations for the complete combustion of (a)  $\text{C}_8\text{H}_{18}(l)$ , (b)  $\text{CH}_3\text{OH}(l)$ , and (c)  $\text{C}_3\text{H}_7\text{SH}(l)$ .

*Solution*

- a. The carbon in  $\text{C}_8\text{H}_{18}$  goes to  $\text{CO}_2(g)$ , and the hydrogen goes to  $\text{H}_2\text{O}(l)$ .



- b. The carbon in  $\text{CH}_3\text{OH}$  goes to  $\text{CO}_2(g)$ , the hydrogen goes to  $\text{H}_2\text{O}(l)$ , and the oxygen is distributed between the  $\text{CO}_2(g)$  and the  $\text{H}_2\text{O}(l)$ .



- c. The carbon in  $\text{C}_3\text{H}_7\text{SH}$  goes to  $\text{CO}_2(g)$ , the hydrogen goes to  $\text{H}_2\text{O}(l)$ , and the sulfur goes to  $\text{SO}_2(g)$ .



**EXERCISE 7.2 - Combustion Reactions**

Write balanced equations for the complete combustion of (a)  $C_4H_{10}(g)$ , (b)  $C_3H_7OH(l)$ , and (c)  $C_4H_9SH(l)$ .

**OBJECTIVE 8**

If there is insufficient oxygen to burn a carbon-containing compound completely, the incomplete combustion converts some of the carbon into carbon monoxide, CO. Because the combustion of gasoline in a car's engine is not complete, the exhaust that leaves the engine contains both carbon dioxide and carbon monoxide. However, before this exhaust escapes out the tail pipe, it passes through a catalytic converter, which further oxidizes much of the carbon monoxide to carbon dioxide. See the Special Topic 7.2: *Air Pollution and Catalytic Converters*.

**SPECIAL TOPIC 7.2 Air Pollution and Catalytic Converters**

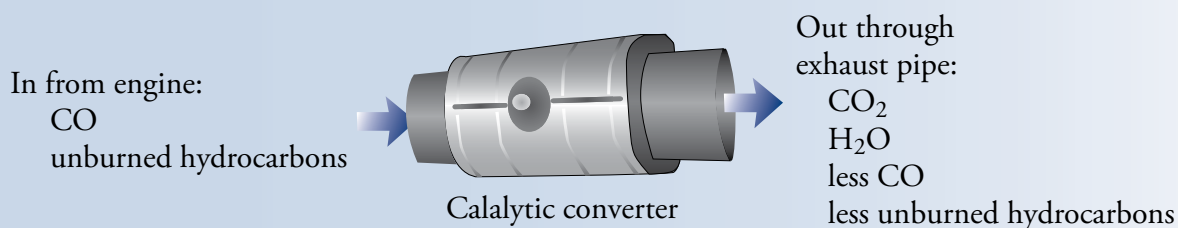
When gasoline, which is a mixture of hydrocarbons, burns in the cylinders of a car engine, the primary products are carbon dioxide and water. Unfortunately, the hydrocarbons are not burned completely, so the exhaust leaving the cylinders also contains some unburned hydrocarbons and some carbon monoxide, which are serious pollutants. When the unburned hydrocarbon molecules escape into the atmosphere, they combine with other substances to form eye irritants and other problematic compounds.

Carbon monoxide is a colorless, odorless, and poisonous gas. When inhaled, it deprives the body of oxygen. Normally, hemoglobin molecules in the blood carry oxygen throughout the body, but when carbon monoxide is present, it combines with hemoglobin and prevents oxygen from doing so. As little as 0.001% CO in air can lead to headache, dizziness, and nausea. Concentrations of 0.1% can cause death.

Catalytic converters in automobile exhaust systems were developed to remove some of the carbon monoxide and unburned hydrocarbons from automobile exhaust. A catalyst is any substance that speeds a chemical reaction without being permanently altered itself. Some of the transition metals, such as platinum, palladium, iridium, and rhodium,

and some transition metal oxides, such as  $V_2O_5$ , CuO, and  $Cr_2O_3$ , will speed up oxidation-reduction reactions. An automobile's catalytic converter contains tiny beads coated with a mixture of transition metals and transition metal oxides. When the exhaust passes over these beads, the oxidation of CO to  $CO_2$  and the conversion of unburned hydrocarbons to  $CO_2$  and  $H_2O$  are both promoted.

The conversion of CO to  $CO_2$  and of unburned hydrocarbons to  $CO_2$  and  $H_2O$  is very slow at normal temperatures, but it takes place quite rapidly at temperatures around 600 °F. This means that the removal of the CO and hydrocarbon pollutants does not take place efficiently when you first start your engine. The pollutants released when a car is first started are called "cold-start emissions." Recent research has focused on developing a way to insulate the catalytic converter so it retains its heat even when the car is not running. Some of the prototypes can maintain their operating temperatures for almost a day while the engine is shut off. With these new, insulated catalytic converters installed in most cars, cold-start emissions could be reduced by at least 50%. This could have a significant effect on air pollution.

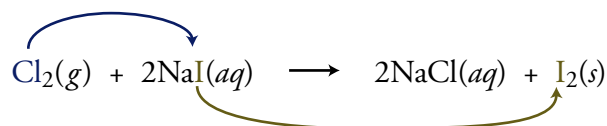
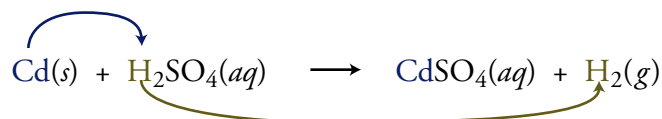
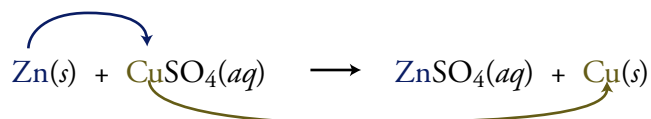


## Single-Displacement Reactions

### OBJECTIVE 7

Now let's consider another type of chemical reaction. In single-displacement reactions, atoms of one element in a compound are displaced (or replaced) by atoms from a pure element. These reactions are also called single-replacement reactions. All of the following are single-displacement reactions.

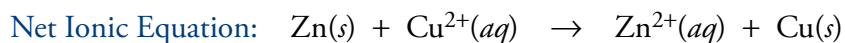
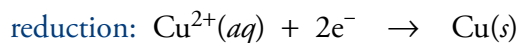
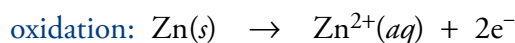
Pure element displaces element in compound



### OBJECTIVE 9

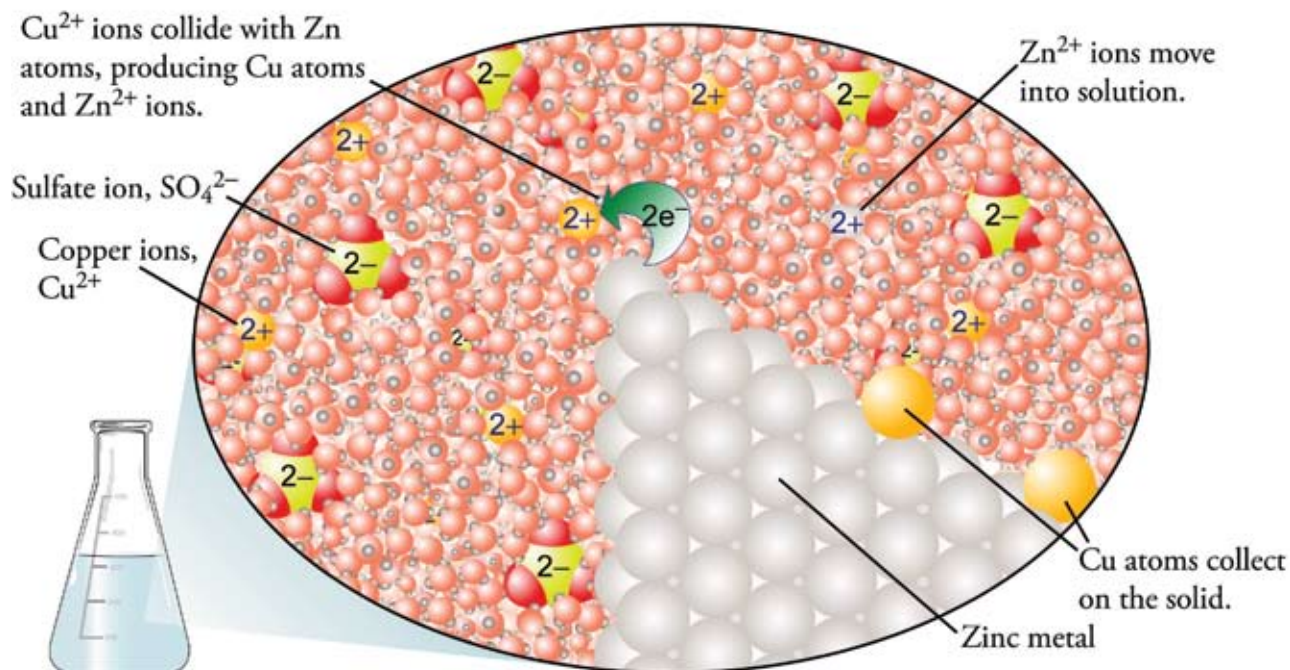
Our model of particle behavior enables us to visualize the movements of atoms and ions participating in single-displacement reactions. For example, consider the first equation above, a reaction in which atoms of zinc replace ions of copper in a copper sulfate solution. Because copper(II) sulfate is a water-soluble ionic compound, the  $\text{CuSO}_4$  solution consists of free  $\text{Cu}^{2+}$  ions surrounded by the negatively charged oxygen ends of water molecules and free  $\text{SO}_4^{2-}$  ions surrounded by the positively charged hydrogen ends of water molecules. These ions move throughout the solution, colliding with each other, with water molecules, and with the walls of their container. Now imagine that a lump of solid zinc is added to the solution. Copper ions begin to collide with the surface of the zinc. When the  $\text{Cu}^{2+}$  ions collide with the uncharged zinc atoms, two electrons are transferred from the zinc atoms to the copper(II) ions. The resulting zinc ions move into solution, where they become surrounded by the negatively charged ends of water molecules, and the uncharged copper solid forms on the surface of the zinc (Figure 7.4).

Because the zinc atoms lose electrons in this reaction and change their oxidation number from 0 to +2, they are oxidized, and zinc is the reducing agent. The  $\text{Cu}^{2+}$  ions gain electrons and decrease their oxidation number from +2 to 0, so they are reduced and act as the oxidizing agent. The half reaction equations and the net ionic equation for this reaction are below.



You can see an animation that illustrates the reaction above at the textbook's Web site.





**Figure 7.4**  
Single-Displacement Reaction  
Between Copper(II) Sulfate and  
Solid Zinc

OBJECTIVE 9

Example 7.3 shows how chemical changes can be classified with respect to the categories described in this section.

### EXAMPLE 7.3 - Classification of Chemical Reactions

Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.

OBJECTIVE 7

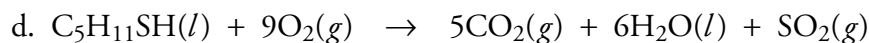
- $2\text{Fe}_2\text{O}_3(s) + 3\text{C}(s) \xrightarrow{\Delta} 4\text{Fe}(l) + 3\text{CO}_2(g)$
- $\text{Cl}_2(g) + \text{F}_2(g) \rightarrow 2\text{ClF}(g)$
- $2\text{Pb}(\text{NO}_3)_2(s) \xrightarrow{\Delta} 4\text{NO}_2(g) + 2\text{PbO}(s) + \text{O}_2(g)$
- $\text{C}_5\text{H}_{11}\text{SH}(l) + 9\text{O}_2(g) \rightarrow 5\text{CO}_2(g) + 6\text{H}_2\text{O}(l) + \text{SO}_2(g)$

*Solution*

- $2\text{Fe}_2\text{O}_3(s) + 3\text{C}(s) \xrightarrow{\Delta} 4\text{Fe}(l) + 3\text{CO}_2(g)$   
Because the atoms of the pure element carbon are displacing (or replacing) the atoms of iron in iron(III) oxide, this is a **single-displacement reaction**. This particular reaction is used to isolate metallic iron from iron ore.
- $\text{Cl}_2(g) + \text{F}_2(g) \rightarrow 2\text{ClF}(g)$   
Because two substances are combining to form one substance, this is a **combination reaction**.



Because one substance is being converted into more than one substance, this is a **decomposition reaction**.

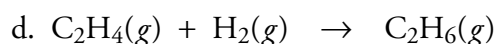
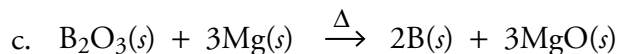
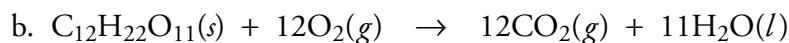
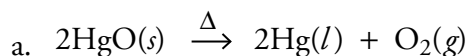


Because a substance combines with oxygen, and because we see carbon in that substance going to  $\text{CO}_2$ , hydrogen going to  $\text{H}_2\text{O}$ , and sulfur going to  $\text{SO}_2$ , we classify this reaction as a **combustion reaction**. The compound  $\text{C}_5\text{H}_{11}\text{SH}$  is 3-methyl-1-butanethiol, a component of the spray produced by skunks.

### EXERCISE 7.3 - Classification of Chemical Reactions

#### OBJECTIVE 7

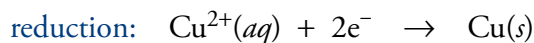
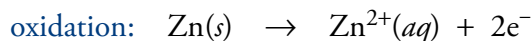
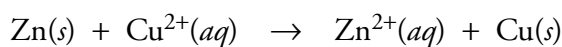
Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.



## 7.4 Voltaic Cells

We're people on the go...with laptop computers, portable drills, and electronic games that kids can play in the car. To keep all of these tools and toys working, we need batteries, and because the newer electronic devices require more power in smaller packages, scientists are constantly searching for stronger and more efficient batteries. The goal of this section is to help you understand how batteries work and to describe some that may be familiar to you and some that may be new.

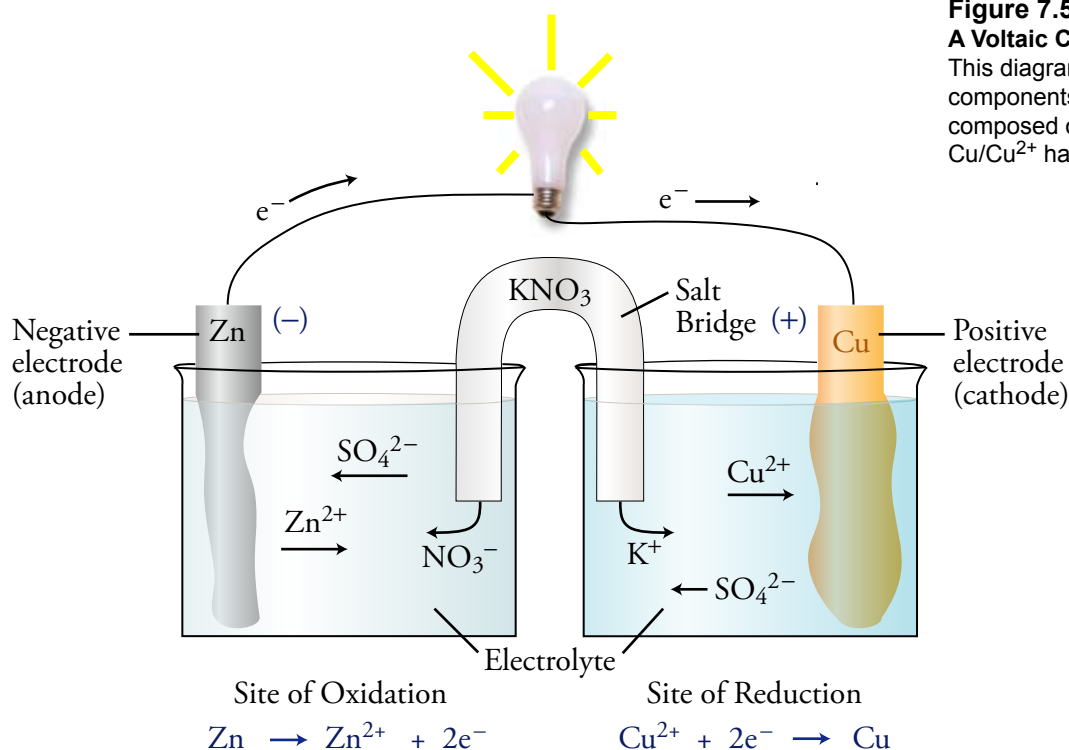
A **battery** is a device that converts chemical energy into electrical energy using redox reactions. To discover what this means and how batteries work, let's examine a simple system that generates an electric current using the reaction between zinc metal and copper(II) ions described in Figure 7.4. In this redox reaction, uncharged zinc atoms are oxidized to zinc ions, and copper(II) ions are reduced to uncharged copper atoms.



The last section showed how this reaction takes place when zinc metal is added to a solution of copper(II) sulfate. When a  $\text{Cu}^{2+}$  ion collides with the zinc metal, two electrons are transferred from a zinc atom directly to the copper(II) ion.

A clever arrangement of the reaction components allows us to harness this reaction to produce electrical energy. The setup, shown in Figure 7.5, keeps the two half-reactions separated, causing the electrons released in the oxidation half of the reaction to pass through a wire connecting the two halves of the apparatus. The proper name for a set-up of this type is a **voltaic cell**. Strictly speaking, a **battery** is a series of voltaic cells joined in such a way that they work together.

OBJECTIVE 10

**Figure 7.5****A Voltaic Cell**

This diagram shows the key components of a voltaic cell composed of the  $\text{Zn}/\text{Zn}^{2+}$  and  $\text{Cu}/\text{Cu}^{2+}$  half-cells.

A voltaic cell, then, is composed of two separate half-cells. In the case of the zinc and copper(II) redox reaction, the first half-cell consists of a strip of zinc metal in a solution of zinc sulfate. The second half-cell consists of a strip of copper metal in a solution of copper(II) sulfate. In the  $\text{Zn}/\text{Zn}^{2+}$  half-cell, zinc atoms lose two electrons and are converted to zinc ions. The two electrons pass through the wire to the  $\text{Cu}/\text{Cu}^{2+}$  half-cell, where  $\text{Cu}^{2+}$  ions gain the two electrons to form uncharged copper atoms. The zinc metal and copper metal strips are called **electrodes**, the general name for electrical conductors placed in half-cells of voltaic cells.



Messages were originally sent through telegraph wires by voltaic cells composed of  $\text{Zn}/\text{Zn}^{2+}$  and  $\text{Cu}/\text{Cu}^{2+}$  half-cells.

The electrode at which oxidation occurs in a voltaic cell is called the **anode**. Because electrons are lost in oxidation, the anode is the source of electrons. For this reason, the anode of a voltaic cell is designated the negative electrode. Because electrons are lost, forming more positive (or less negative) species at the anode, the anode surroundings tend to become more positive. Thus anions are attracted to the anode. In our voltaic cell, Zn is oxidized to  $\text{Zn}^{2+}$  at the zinc electrode, so this electrode is the anode. The solution around the zinc metal becomes positive due to excess  $\text{Zn}^{2+}$  ions, so anions are attracted to the zinc electrode.

OBJECTIVE 11

OBJECTIVE 10

The **cathode** is the electrode in a voltaic cell at which reduction occurs. By convention, the cathode is designated the positive electrode. Because electrons flow along the wire

OBJECTIVE 10

to the cathode, and because substances gain those electrons to become more negative (or less positive), the cathode surroundings tend to become more negative. Thus cations are attracted to the cathode. In our voltaic cell,  $\text{Cu}^{2+}$  ions are reduced to uncharged copper atoms at the copper strip, so metallic copper is the cathode. Because cations are removed from the solution there and anions are not, the solution around the copper cathode tends to become negative and attracts cations.

The component of the voltaic cell through which ions are able to flow is called the **electrolyte**. For our voltaic cell, the zinc sulfate solution is the electrolyte in the anode half-cell, and the copper(II) sulfate solution is the electrolyte in the cathode half-cell.

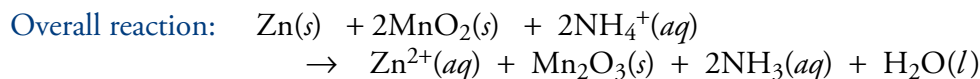
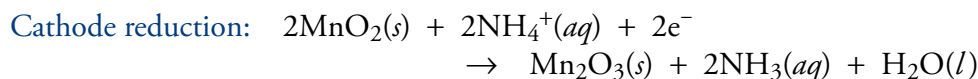
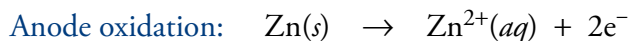
As described above, the  $\text{Zn}/\text{Zn}^{2+}$  half-cell tends to become positive due to the loss of electrons as uncharged zinc atoms are converted into  $\text{Zn}^{2+}$  ions, and the  $\text{Cu}/\text{Cu}^{2+}$  half-cell tends to become negative due to the gain of electrons and the conversion of  $\text{Cu}^{2+}$  ions into uncharged copper atoms. This charge imbalance would block the flow of electrons and stop the redox reaction if something were not done to balance the growing charges. One way to keep the charge balanced is to introduce a device called a **salt bridge**.

#### OBJECTIVE 10

One type of salt bridge is a tube connecting the two solutions and filled with an unreactive ionic compound such as potassium nitrate in a semisolid support such as gelatin or agar. For each negative charge lost at the anode due to the loss of an electron, an anion, such as  $\text{NO}_3^-$ , moves from the salt bridge into the solution to replace it. For example, when one zinc atom oxidizes at the anode and loses two electrons to form  $\text{Zn}^{2+}$ , two nitrates enter the solution to keep the solution electrically neutral overall. For each negative charge gained at the cathode due to the gain of an electron, a cation, such as  $\text{K}^+$  ion, moves into the solution to keep the solution uncharged. For example, each time a copper ion gains two electrons and forms an uncharged copper atom, two potassium ions enter the solution to replace the  $\text{Cu}^{2+}$  lost.

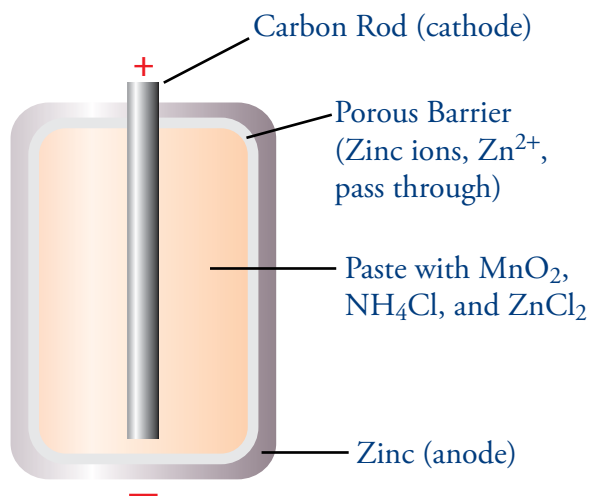
## Dry Cells

Although a voltaic cell of the kind described above, using zinc and a solution of copper ions, was used in early telegraph systems, there are problems with this sort of cell. The greatest problem is that the cell cannot be easily moved because the electrolyte solutions are likely to spill. The Leclanché cell, or dry cell, was developed in the 1860s to solve this problem. It contained a paste, or semisolid, electrolyte. The reactions in the dry cell can be thought of as consisting of the following half-reactions (although they are a bit more complicated than described here):



These inexpensive and reliable cells have served as the typical “flashlight battery” for many years. Their outer wrap surrounds a zinc metal cylinder that acts as the anode (Figure 7.6). Inside the zinc cylinder is a porous barrier that separates the zinc metal

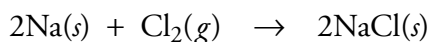
**Figure 7.6**  
Leclanché (or Dry) Cell



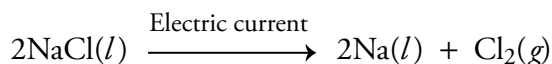
from a paste containing  $\text{NH}_4\text{Cl}$ ,  $\text{ZnCl}_2$ , and  $\text{MnO}_2$ . The porous barrier allows  $\text{Zn}^{2+}$  ions to pass through, but it keeps the  $\text{MnO}_2$  from coming into direct contact with the zinc metal. In the center of the cell, a carbon rod that can conduct an electric current acts as the cathode.

## Electrolysis

Voltage, a measure of the strength of an electric current, represents the force that moves electrons from the anode to the cathode in a voltaic cell. When a greater force (voltage) is applied in the opposite direction, electrons can be pushed from what would normally be the cathode toward the voltaic cell's anode. This process is called **electrolysis**. In a broader sense, electrolysis is the process by which a redox reaction is made to occur in the nonspontaneous direction. For example, sodium metal reacts readily with chlorine gas to form sodium chloride, but we do not expect sodium chloride, as it sits in our saltshakers, to decompose into sodium metal and chlorine gas. We say the forward reaction below is spontaneous, and the reverse reaction is nonspontaneous.



Sodium metal and chlorine gas can, however, be formed from the electrolysis of salt, in which an electric current is passed through the molten sodium chloride.

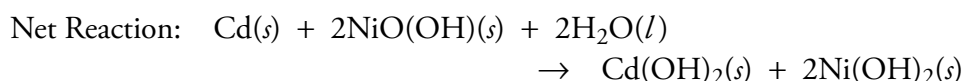
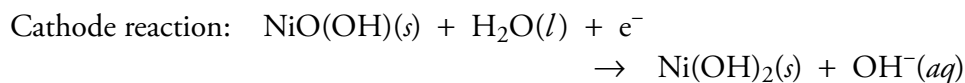


Electrolysis is used in industry to purify metals, such as copper and aluminum, and in electroplating, the process used, for example, to deposit the chrome on the bumper of a 1955 Chevy.

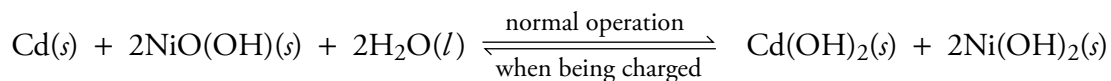
As you will see below, a similar process is used to refresh rechargeable batteries.

## Nickel-Cadmium Batteries

Leclanché cells are not rechargeable; once the reactants are depleted, the battery must be discarded and replaced. Batteries that are not rechargeable are called **primary batteries**. A rechargeable battery is called a **secondary battery** or a **storage battery**. The nickel-cadmium (NiCd) battery is a popular rechargeable battery that uses the following redox reaction:



To recharge a secondary battery, an opposing external voltage is applied that is greater than the voltage of the cell, pushing the electrons in the opposite direction from the way they move in the normal operation of the cell. In this process, the original chemical reaction is reversed. Because the  $\text{Cd}(\text{OH})_2$  and  $\text{Ni}(\text{OH})_2$  produced during the normal operation of the nickel-cadmium battery are solids, they stay at the electrodes where they are formed and are available to be converted back to the original reactants.



Leclanché cells and nickel-cadmium batteries both have their drawbacks. One problem with the Leclanché cell is that over time the zinc metal and the ammonium ions react to corrode the anode, so dry cell batteries have a shorter “shelf life” than other batteries do. Another problem is that Leclanché cells are not rechargeable. Nickel-cadmium batteries are rechargeable, but they are too bulky to be used for many purposes. The manufacturers of electronic devices are constantly trying to make their products smaller and more powerful. Thus they want batteries that are smaller, lighter, and more powerful than dry cell and nickel-cadmium batteries. They want batteries with the highest possible energy-to-mass ratio and energy-to-volume ratio, and dry cells and NiCd batteries have very low ratios. To solve these and other problems, many additional types of batteries have been developed, including the lithium battery (Table 7.3).

Extensive research is being done to develop new types of batteries to power electric cars and trucks. See Special Topic 7.3 *Zinc-Air Batteries* for an example.

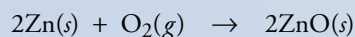
**Table 7.3**

Common Voltaic Cells (Batteries)

Voltaic Cell	Anode Reaction	Cathode Reaction	Pros	Cons
Dry cell	$\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^{-}$	$2\text{MnO}_2(s) + 2\text{NH}_4^{+}(aq) + 2e^{-} \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{NH}_3(aq) + \text{H}_2\text{O}(l)$	Inexpensive; reliable	Short shelf-life; large and heavy; not rechargeable
Alkaline battery	$\text{Zn}(s) + 2\text{OH}^{-}(aq) \rightarrow \text{ZnO}(s) + \text{H}_2\text{O}(l) + 2e^{-}$	$2\text{MnO}_2(s) + \text{H}_2\text{O}(l) + 2e^{-} \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{OH}^{-}(aq)$	Longer lasting and slightly higher energy density than dry cells	More expensive than dry cells; not rechargeable
Nickel-cadmium battery	$\text{Cd}(s) + 2\text{OH}^{-}(aq) \rightarrow \text{Cd}(\text{OH})_2(s) + 2e^{-}$	$\text{NiO}(\text{OH})(s) + \text{H}_2\text{O}(l) + e^{-} \rightarrow \text{Ni}(\text{OH})_2(s) + \text{OH}^{-}(aq)$	Rechargeable; lower price than other rechargeable batteries; long life	Relatively low energy density; less efficient with repeated use
Lead-Acid battery	$\text{Pb}(s) + \text{HSO}_4^{-}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{PbSO}_4(s) + \text{H}_3\text{O}^{+}(aq) + 2e^{-}$	$\text{PbO}_2(s) + \text{HSO}_4^{-}(aq) + 3\text{H}_3\text{O}^{+}(aq) + 2e^{-} \rightarrow \text{PbSO}_4(s) + 5\text{H}_2\text{O}(l)$	Rechargeable; very reliable; inexpensive; durable	Very heavy, yielding a very low energy density
One type of lithium battery	$\text{Li}_n\text{C}_6 \rightarrow n\text{Li}^{+} + 6\text{C} + ne^{-}$	$\text{CoO}_2 + n\text{Li}^{+} + ne^{-} \rightarrow \text{Li}_n\text{CoO}_2$	Rechargeable; reliable; very high energy density	Expensive

### SPECIAL TOPIC 7.3 Zinc-Air Batteries

One type of battery being considered for use in electric cars is the zinc-air battery, in which zinc acts as the anode and air provides the oxygen for the cathode reaction. This battery is low in mass because of its use of oxygen as the cathode reactant. The overall reaction is



Most of the batteries being considered for electric cars take several hours to recharge, a major drawback in comparison to the ease of refilling a gas tank. It is possible that the use of zinc-air batteries will avoid this problem. The reactions are not directly reversible in the zinc-air batteries, so the

materials would need to be removed to be recharged.

You could stop at your local zinc station, have the spent electrolyte and zinc oxide particles quickly suctioned from the cells and replaced with a slurry of fresh particles and electrolyte, and be ready to go another 200-300 miles. The zinc oxide would be taken to a central plant where it would be converted back to zinc metal and returned to the stations.

Vans powered by zinc-air batteries are already being used by the European postal services. In 1997, one of them made a 152-mile trip from Chambria, France, over the Alps to Turin, Italy. It not only succeeded in crossing the 6874-ft peak but also averaged 68-mph on the highway.

## Chapter Glossary

**Oxidation** Any chemical change in which at least one element loses electrons, either completely or partially.

**Reduction** Any chemical change in which at least one element gains electrons, either completely or partially.

**Oxidation-reduction reactions** The chemical reactions in which there is a complete or partial transfer of electrons, resulting in oxidation and reduction. These reactions are also called **redox** reactions.

**Half-reactions** Separate oxidation and reduction reaction equations in which electrons are shown as a reactant or product.

**Reducing agent** A substance that loses electrons, making it possible for another substance to gain electrons and be reduced.

**Oxidizing agent** A substance that gains electrons, making it possible for another substance to lose electrons and be oxidized.

**Oxidation number** A tool for keeping track of the flow of electrons in redox reactions (also called **oxidation state**).

**Combination or synthesis reaction** The joining of two or more elements or compounds into one product.

**Decomposition reaction** The conversion of one compound into two or more simpler substances.

**Combustion reaction** Rapid oxidation accompanied by heat and usually light.

**Single-displacement reaction** Chemical change in which atoms of one element displace (or replace) atoms of another element in a compound.

**Voltaic cell** A system in which two half-reactions for a redox reaction are separated, allowing the electrons transferred in the reaction to be passed between them through a wire.

**Battery** A device that has two or more voltaic cells connected together. The term is also used to describe any device that converts chemical energy into electrical energy using redox reactions.

**Electrodes** The electrical conductors placed in the half-cells of a voltaic cell.

**Anode** The electrode at which oxidation occurs in a voltaic cell. It is the source of electrons and is the negative electrode.

**Cathode** The electrode at which reduction occurs in a voltaic cell. It is the positive electrode.

**Salt bridge** A device used to keep the charges in a voltaic cell balanced.

**Electrolyte** The portion of a voltaic cell that allows ions to flow.

**Primary battery** A battery that is not rechargeable.

**Secondary battery or storage battery** A rechargeable battery.

**Electrolysis** The process by which a redox reaction is pushed in the nonspontaneous direction or the process of applying an external voltage to a voltaic cell, causing electrons to move from what would normally be the cell's cathode toward its anode.

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



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

## Chapter Objectives

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

### Section 7.1 An Introduction to Oxidation-Reduction Reactions

2. Describe the difference between the redox reactions that form binary ionic compounds, such as zinc oxide, from their elements and the similar redox reactions that form molecular compounds, such as nitrogen monoxide, from their elements. (Your description should include the degree to which the electrons are transferred in the reactions.)

### Section 7.2 Oxidation Numbers

3. Given a formula for a substance, determine the oxidation number (or oxidation state) for each atom in the formula.
4. Given an equation for a chemical reaction, identify whether the equation represents a redox reaction or not.
5. Given an equation for a redox reaction, identify the substance that is oxidized and the substance that is reduced.
6. Given an equation for a redox reaction, identify the substance that is the reducing agent and the substance that is the oxidizing agent.

### Section 7.3 Types of Reactions

7. Given an equation for a reaction, identify whether it represents a combination reaction, a decomposition reaction, a combustion reaction, or a single-displacement reaction.
8. Given the formula for a substance that contains one or more of the elements carbon, hydrogen, oxygen, and sulfur, write the balanced equation for the complete combustion of the substance.
9. Describe the reaction that takes place between an uncharged metal, such as zinc, and a cation, such as the copper(II) ion, in a water solution. Your description should include the nature of the particles in the system before the reaction takes place, the nature of the reaction itself, and the nature of the particles in the system after the reaction. Your description should also include the equations for the half-reactions and the net ionic equation for the overall reaction.

### Section 7.4 Voltaic Cells

10. Describe how a voltaic cell can be made using the redox reaction between a metal, such as zinc, and an ionic substance containing a metallic ion, such as the copper(II) ion. Your description should include a sketch that shows the key components of the two half-cells, labels to indicate which electrode is the cathode and which is the anode, signs to indicate which electrode is negative and which is positive, arrows to show the direction of movement of the electrons in the wire between the electrodes, and arrows to show the direction of movement of the ions in the system. (See Figure 7.5.)
11. Given a description of a voltaic cell, including the nature of the half-reactions, identify which electrode is the cathode and which is the anode.

## Review Questions

- For each of the following ionic formulas, write the formula for the cation and the formula for the anion.
  - $\text{FeBr}_3$
  - $\text{Co}_3(\text{PO}_4)_2$
  - $\text{AgCl}$
  - $(\text{NH}_4)_2\text{SO}_4$
- Classify each of the following formulas as representing a binary ionic compound, an ionic compound with polyatomic ions, or a molecular compound.
  - $\text{CF}_4$
  - $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$
  - $\text{CoCl}_2$
  - $\text{C}_2\text{H}_5\text{OH}$
  - $\text{H}_2\text{S}$
  - $\text{ClF}$
  - $\text{Cr}(\text{OH})_3$
  - $\text{H}_3\text{PO}_4$
- Balance the following equations. ( $\text{C}_8\text{H}_{18}$  is a component of gasoline, and  $\text{P}_2\text{S}_5$  is used to make the insecticides parathion and malathion.)
  - $\text{C}_8\text{H}_{18}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)$
  - $\text{P}_4(s) + \text{S}_8(s) \rightarrow \text{P}_2\text{S}_5(s)$

## Key Ideas

Complete the following statements by writing one of these words or phrases in each blank.

carbon dioxide  
change  
decreases  
flow of electrons  
gains  
half-reactions  
heat  
increases  
light  
loses  
one  
oxidation  
oxidized

oxidizing  
oxidizing agent  
partial  
pure element  
rarely  
reduced  
reducing  
reducing agent  
reduction  
sulfur dioxide  
transferred  
two or more  
water

- According to the modern convention, any chemical change in which an element \_\_\_\_\_ electrons is called an oxidation.
- According to the modern definition, any chemical change in which an element \_\_\_\_\_ electrons is called a reduction.

6. Electrons are \_\_\_\_\_ found unattached to atoms. Thus, for one element or compound to lose electrons and be \_\_\_\_\_, another element or compound must be there to gain the electrons and be \_\_\_\_\_. In other words, \_\_\_\_\_ (loss of electrons) must be accompanied by \_\_\_\_\_ (gain of electrons).
7. Reactions in which electrons are \_\_\_\_\_, resulting in oxidation and reduction, are called oxidation-reduction reactions.
8. The separate oxidation and reduction equations are called \_\_\_\_\_.
9. A(n) \_\_\_\_\_ is a substance that loses electrons, making it possible for another substance to gain electrons and be reduced.
10. A(n) \_\_\_\_\_ is a substance that gains electrons, making it possible for another substance to lose electrons and be oxidized.
11. Oxidation is defined as the complete or \_\_\_\_\_ loss of electrons, reduction as the complete or partial gain of electrons.
12. Just think of oxidation numbers as tools for keeping track of the \_\_\_\_\_ in redox reactions.
13. If any element undergoes a(n) \_\_\_\_\_ of oxidation number in the course of a reaction, the reaction is a redox reaction. If an element's oxidation number \_\_\_\_\_ in a reaction, that element is oxidized. If an element's oxidation number \_\_\_\_\_ in a reaction, that element is reduced. The reactant containing the element that is oxidized is the \_\_\_\_\_ agent. The reactant containing the element that is reduced is the \_\_\_\_\_ agent.
14. In combination reactions, \_\_\_\_\_ elements or compounds combine to form one compound.
15. In decomposition reactions, \_\_\_\_\_ compound is converted into two or more simpler substances.
16. In a combustion reaction, oxidation is very rapid and is accompanied by \_\_\_\_\_ and usually \_\_\_\_\_.
17. When any substance that contains carbon is combusted (or burned) completely, the carbon forms \_\_\_\_\_.
18. When a substance that contains hydrogen is burned completely, the hydrogen forms \_\_\_\_\_.
19. When any substance that contains sulfur burns completely, the sulfur forms \_\_\_\_\_.
20. In single-displacement reactions, atoms of one element in a compound are displaced (or replaced) by atoms from a(n) \_\_\_\_\_.

Complete the following statements by writing one of these words or phrases in each blank.

chemical energy	not rechargeable
electrical conductors	oxidation
electrical energy	positive
electrolysis	positive electrode
flow	rechargeable
force	reduction
half-reactions	voltaic cells

21. Strictly speaking, a battery is a series of \_\_\_\_\_ joined in such a way that they work together. A battery can also be described as a device that converts \_\_\_\_\_ into \_\_\_\_\_ using redox reactions.
22. A voltaic cell keeps two oxidation-reduction \_\_\_\_\_ separated, causing the electrons released in the oxidation half of the reaction to pass through a wire connecting the two halves of the apparatus.
23. Metal strips in voltaic cells are called electrodes, which is the general name for \_\_\_\_\_ placed in half-cells of voltaic cells.
24. The electrode at which \_\_\_\_\_ occurs in a voltaic cell is called the anode. Because electrons are lost, forming more positive (or less negative) species at the anode, the anode surroundings tend to become more \_\_\_\_\_.
25. The cathode is the electrode in a voltaic cell at which \_\_\_\_\_ occurs. By convention, the cathode is designated the \_\_\_\_\_. Because electrons flow along the wire to the cathode, and because substances gain those electrons to become more negative (or less positive), the cathode surroundings tend to become more negative. Thus cations are attracted to the cathode.
26. The component of the voltaic cell through which ions are able to \_\_\_\_\_ is called the electrolyte.
27. Voltage, a measure of the strength of an electric current, represents the \_\_\_\_\_ that moves electrons from the anode to the cathode in a voltaic cell. When a greater voltage is applied in the opposite direction, electrons can be pushed from what would normally be the cathode toward the voltaic cell's anode. This process is called \_\_\_\_\_.
28. Batteries that are \_\_\_\_\_ are called primary batteries. A(n) \_\_\_\_\_ battery is called a secondary battery or a storage battery.

## Chapter Problems

### OBJECTIVE 3

### Section 7.1 An Introduction to Oxidation-Reduction Reactions

29. Describe the difference between the redox reactions that form binary ionic compounds, such as zinc oxide, from their elements and the similar redox reactions that form molecular compounds, such as nitrogen monoxide, from their elements.
30. Are the electrons in the following redox reactions transferred completely from the atoms of one element to the atoms of another or are they only partially transferred?
  - a.  $4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s)$
  - b.  $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$

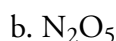
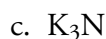
31. Are the electrons in the following redox reactions transferred completely from the atoms of one element to the atoms of another or are they only partially transferred?
- $2\text{K}(s) + \text{F}_2(g) \rightarrow 2\text{KF}(s)$
  - $2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l)$
32. Are the electrons in the following redox reactions transferred completely from the atoms of one element to the atoms of another or are they only partially transferred?
- $\text{S}_8(s) + 8\text{O}_2(g) \rightarrow 8\text{SO}_2(g)$
  - $\text{P}_4(s) + 6\text{Cl}_2(g) \rightarrow 4\text{PCl}_3(l)$
33. Are the electrons in the following redox reactions transferred completely from the atoms of one element to the atoms of another or are they only partially transferred?
- $\text{Ca}(s) + \text{Cl}_2(g) \rightarrow \text{CaCl}_2(s)$
  - $4\text{Cu}(s) + \text{O}_2(g) \rightarrow 2\text{Cu}_2\text{O}(s)$
34. Aluminum bromide,  $\text{AlBr}_3$ , which is used to add bromine atoms to organic compounds, can be made by passing gaseous bromine over hot aluminum. Which of the following half-reactions for this oxidation-reduction reaction describes the oxidation, and which one describes the reduction?
- $$2\text{Al} \rightarrow 2\text{Al}^{3+} + 6e^-$$
- $$3\text{Br}_2 + 6e^- \rightarrow 6\text{Br}^-$$
35. Iodine,  $\text{I}_2$ , has many uses, including the production of dyes, antiseptics, photographic film, pharmaceuticals, and medicinal soaps. It forms when chlorine,  $\text{Cl}_2$ , reacts with iodide ions in a sodium iodide solution. Which of the following half-reactions for this oxidation-reduction reaction describes the oxidation, and which one describes the reduction?
- $$\text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^-$$
- $$2\text{I}^- \rightarrow \text{I}_2 + 2e^-$$

### Section 7.2 Oxidation Numbers

36. Determine the oxidation number for the atoms of each element in the following formulas. OBJECTIVE 3
- $\text{S}_8$
  - $\text{S}^{2-}$
  - $\text{Na}_2\text{S}$
  - $\text{FeS}$
37. Determine the oxidation number for the atoms of each element in the following formulas. OBJECTIVE 3
- $\text{P}_4$
  - $\text{PF}_3$
  - $\text{PH}_3$
  - $\text{P}_2\text{O}_3$
  - $\text{H}_3\text{PO}_4$
38. Determine the oxidation number for the atoms of each element in the following formulas. OBJECTIVE 3
- $\text{Sc}_2\text{O}_3$
  - $\text{RbH}$
  - $\text{N}_2$
  - $\text{NH}_3$

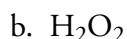
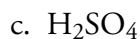
## OBJECTIVE 3

39. Determine the oxidation number for the atoms of each element in the following formulas.



## OBJECTIVE 3

40. Determine the oxidation number for the atoms of each element in the following formulas.



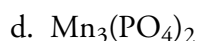
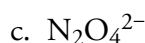
## OBJECTIVE 3

41. Determine the oxidation number for the atoms of each element in the following formulas.



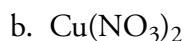
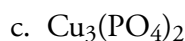
## OBJECTIVE 3

42. Determine the oxidation number for the atoms of each element in the following formulas.



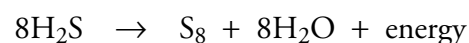
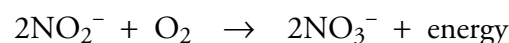
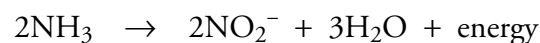
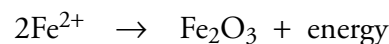
## OBJECTIVE 3

43. Determine the oxidation number for the atoms of each element in the following formulas.



## OBJECTIVE 3

44. The following partial reactions represent various means by which bacteria obtain energy. Determine the oxidation number for each atom other than oxygen and hydrogen atoms, and decide whether the half-reaction represents oxidation or reduction. None of the oxygen or hydrogen atoms are oxidized or reduced.



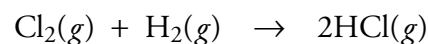
## OBJECTIVE 3

## OBJECTIVE 4

## OBJECTIVE 5

## OBJECTIVE 6

45. About 47% of the hydrochloric acid produced in the U.S. is used for cleaning metallic surfaces. Hydrogen chloride,  $\text{HCl}$ , which is dissolved in water to make the acid, is formed in the reaction of chlorine gas and hydrogen gas, displayed below. Determine the oxidation number for each atom in the equation, and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.



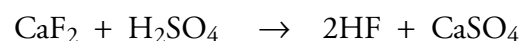
## OBJECTIVE 3

## OBJECTIVE 4

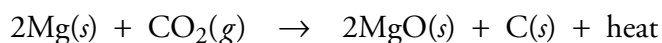
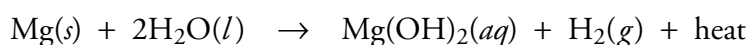
## OBJECTIVE 5

## OBJECTIVE 6

46. The following equation describes the reaction that produces hydrofluoric acid, which is used to make chlorofluorocarbons (CFCs). Determine the oxidation number for each atom in the equation, and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.



47. Water and carbon dioxide fire extinguishers should not be used on magnesium fires because both substances react with magnesium and generate enough heat to intensify the fire. Determine the oxidation number for each atom in the equations that describe these reactions (displayed below), and decide whether each reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.



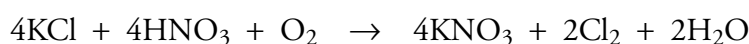
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

48. Potassium nitrate is used in the production of fireworks, explosives, and matches. It is also used in curing foods and to modify the burning properties of tobacco. The reaction for the industrial production of  $\text{KNO}_3$  is summarized below. Determine the oxidation number for each atom, and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.



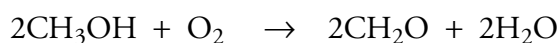
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

49. Formaldehyde,  $\text{CH}_2\text{O}$ , which is used in embalming fluids, is made from methanol in the reaction described below. Determine the oxidation number for each atom in this equation, and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.



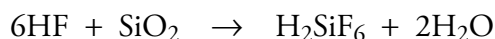
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

50. The weak acid hydrofluoric acid,  $\text{HF}(aq)$ , is used to frost light bulbs. It reacts with the silicon dioxide in glass on the inside of light bulbs to form a white substance,  $\text{H}_2\text{SiF}_6$ , that deposits on the glass and reduces the glare from the bulb. The same reaction is run on a larger scale to produce  $\text{H}_2\text{SiF}_6$  used for fluoridating drinking water. Determine the oxidation number for each atom in the reaction, and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.



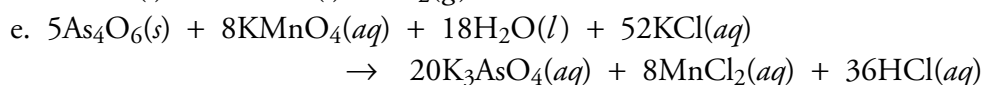
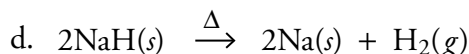
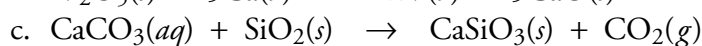
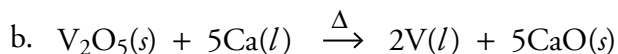
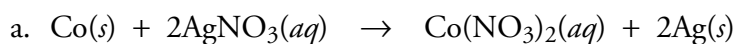
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

51. For each of the following equations, determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.



OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

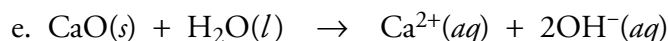
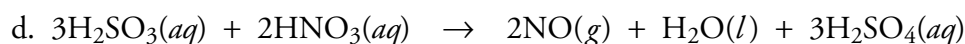
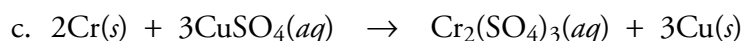
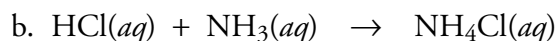
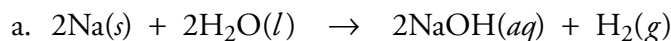
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

52. For each of the following equations, determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.



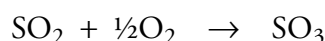
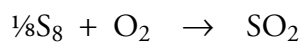
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

53. The following equations summarize the steps in the process used to make most of the sulfuric acid produced in the United States. Determine the oxidation number for each atom in each of the following equations, and identify whether each reaction is a redox reaction or not. For the redox reactions, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.



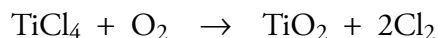
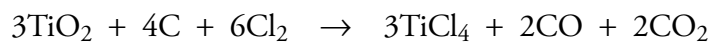
OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 5

OBJECTIVE 6

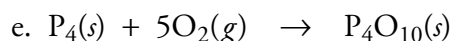
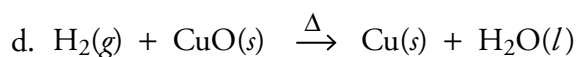
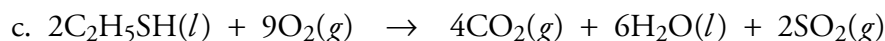
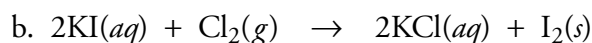
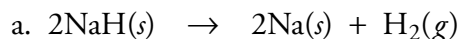
54. Because of its superior hiding ability, titanium dioxide has been the best selling white pigment since 1939. In 1990, there were 2.16 billion pounds of it sold in the United States for a variety of purposes, including surface coatings (paint), plastics, and paper. The following equations show how impure  $\text{TiO}_2$  is purified. Determine the oxidation number for each atom in them and identify whether each reaction is a redox reaction or not. For the redox reactions, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.



### Section 7.3 Types of Redox Reactions

OBJECTIVE 7

55. Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.





56. Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.

OBJECTIVE 7

- $\text{Fe}_2(\text{CO}_3)_2(s) \xrightarrow{\Delta} \text{Fe}_2\text{O}_3(s) + 2\text{CO}_2(g)$
- $2\text{C}_6\text{H}_{11}\text{OH}(l) + 17\text{O}_2(g) \rightarrow 12\text{CO}_2(g) + 12\text{H}_2\text{O}(l)$
- $\text{P}_4\text{O}_{10}(s) + 6\text{H}_2\text{O}(l) \rightarrow 4\text{H}_3\text{PO}_4(aq)$
- $2\text{C}(s) + \text{MnO}_2(s) \xrightarrow{\Delta} \text{Mn}(s) + 2\text{CO}(g)$
- $2\text{NaClO}_3(s) \xrightarrow{\Delta} 2\text{NaCl}(s) + 3\text{O}_2(g)$

57. Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.

OBJECTIVE 7

- $4\text{B}(s) + 3\text{O}_2(g) \rightarrow 2\text{B}_2\text{O}_3(s)$
- $(\text{C}_2\text{H}_5)_2\text{O}(l) + 6\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 5\text{H}_2\text{O}(l)$
- $2\text{Cr}_2\text{O}_3(s) + 3\text{Si}(s) \xrightarrow{\Delta} 4\text{Cr}(s) + 3\text{SiO}_2(s)$
- $\text{C}_6\text{H}_{11}\text{SH}(l) + 10\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) + \text{SO}_2(g)$
- $2\text{NaHCO}_3(s) \xrightarrow{\Delta} \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$

58. Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.

OBJECTIVE 7

- $3\text{H}_2(g) + \text{WO}_3(s) \xrightarrow{\Delta} \text{W}(s) + 3\text{H}_2\text{O}(l)$
- $2\text{I}_4\text{O}_9(s) \rightarrow 2\text{I}_2\text{O}_6(s) + 2\text{I}_2(s) + 3\text{O}_2(g)$
- $2\text{NaNO}_3(s) \xrightarrow{\Delta} 2\text{NaNO}_2(s) + \text{O}_2(g)$
- $\text{Cl}_2(g) + 2\text{KBr}(aq) \rightarrow 2\text{KCl}(aq) + \text{Br}_2(l)$

59. Write balanced equations for the complete combustion of each of the following substances.

OBJECTIVE 8

- $\text{C}_3\text{H}_8(g)$
- $\text{C}_4\text{H}_9\text{OH}(l)$
- $\text{CH}_3\text{COSH}(l)$

60. Write balanced equations for the complete combustion of each of the following substances.

OBJECTIVE 8

- $\text{C}_{13}\text{H}_{28}(l)$
- $\text{C}_{12}\text{H}_{22}\text{O}_{11}(s)$
- $\text{C}_2\text{H}_5\text{SO}_3\text{H}(l)$

## OBJECTIVE 9

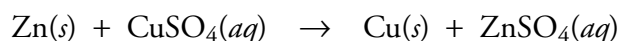
61. The following pairs react in single-displacement reactions that are similar to the reaction between uncharged zinc metal and the copper(II) ions in a copper(II) sulfate solution. Describe the changes in these reactions, including the nature of the particles in the system before the reaction takes place, the nature of the reaction itself, and the nature of the particles in the system after the reaction. Your description should also include the equations for the half-reactions and the net ionic equation for the overall reaction.
- Magnesium metal and nickel(II) nitrate,  $\text{Ni}(\text{NO}_3)_2(aq)$
  - Calcium metal and cobalt(II) chloride,  $\text{CoCl}_2(aq)$

## Section 7.4 Voltaic Cells

## OBJECTIVE 10

## OBJECTIVE 11

62. We know that the following reaction can be used to generate an electric current in a voltaic cell.



Sketch similar voltaic cells made from each of the reactions presented below, showing the key components of the two half-cells and indicating the cathode electrode and the anode electrode, the negative and positive electrodes, the direction of movement of the electrons in the wire between the electrodes, and the direction of movement of the ions in the system. Show a salt bridge in each sketch, and show the movement of ions out of the salt bridge.

- $\text{Mn}(s) + \text{PbCl}_2(aq) \rightarrow \text{Pb}(s) + \text{MnCl}_2(aq)$
- $\text{Mg}(s) + 2\text{AgNO}_3(aq) \rightarrow 2\text{Ag}(s) + \text{Mg}(\text{NO}_3)_2(aq)$

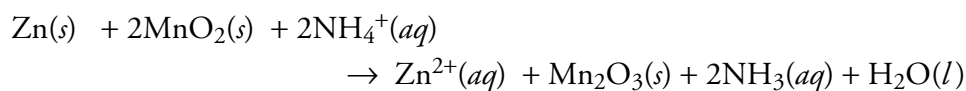
## OBJECTIVE 3

## OBJECTIVE 4

## OBJECTIVE 5

## OBJECTIVE 6

63. The following equation summarizes the chemical changes that take place in a typical dry cell.



Determine the oxidation number for each atom in the equation, and identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.

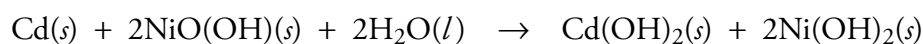
## OBJECTIVE 3

## OBJECTIVE 4

## OBJECTIVE 5

## OBJECTIVE 6

64. The following equation summarizes the chemical changes that take place in a nickel-cadmium battery. Determine the oxidation number for each atom in the equation, and identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.



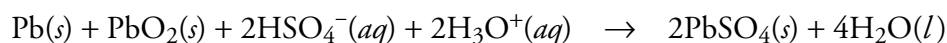
## OBJECTIVE 3

## OBJECTIVE 4

## OBJECTIVE 5

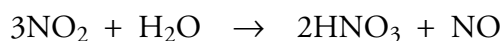
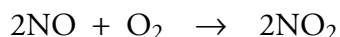
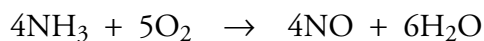
## OBJECTIVE 6

65. The following equation summarizes the chemical changes that take place in a lead-acid battery. Determine the oxidation number for each atom in the equation, and identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.

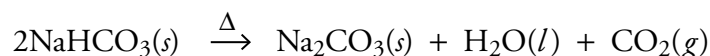


**Additional Problems**

66. Nitric acid, which is used to produce fertilizers and explosives, is made in the following three steps. Determine the oxidation number for each atom in the three equations and identify whether each reaction is a redox reaction or not. If a reaction is redox, identify what is oxidized and what is reduced.



67. Sodium hydrogen carbonate,  $\text{NaHCO}_3$ , best known as the active ingredient in baking soda, is used in several ways in food preparation. It is also added to animal feeds and used to make soaps and detergents. Baking soda can also be used to put out small fires on your stovetop. The heat of the flames causes the  $\text{NaHCO}_3$  to decompose to form carbon dioxide, which displaces the air above the flames depriving the fire of the oxygen necessary for combustion. The equation for the reaction is below.



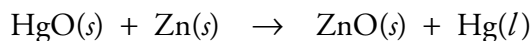
Determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized and what is reduced.

68. Swimming pools can be chlorinated by adding either calcium hypochlorite,  $\text{Ca}(\text{OCl})_2$ , or sodium hypochlorite,  $\text{NaOCl}$ . The active component is the hypochlorite ion,  $\text{OCl}^-$ . Because the rate of the following reaction is increased by ultraviolet radiation in sunlight, it is best to chlorinate pools in the evening to avoid the hypochlorite ion's decomposition.



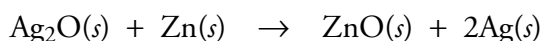
Determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized and what is reduced.

69. In the past, mercury batteries were commonly used to power electronic watches and small appliances. The overall reaction for this type of battery is



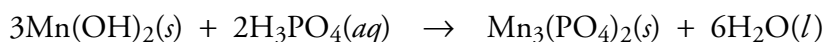
Determine the oxidation number for each atom in the equation and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

70. Silver batteries have been used to run heart pacemakers and hearing aids. The overall reaction for this type of battery is



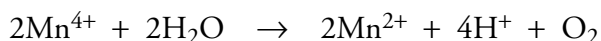
Determine the oxidation number for each atom in the equation and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

71.  $\text{Mn}_3(\text{PO}_4)_2$ , which is used to make corrosion resistant coatings on steel, aluminum, and other metals, is made from the reaction of  $\text{Mn}(\text{OH})_2$  with  $\text{H}_3\text{PO}_4$ .



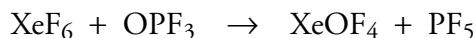
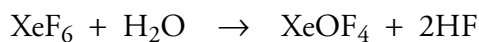
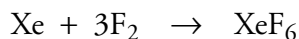
Determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If it is redox, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.

72. One of the ways that plants generate oxygen is represented by the following reaction.

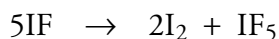


Determine the oxidation number for each atom in the equation and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

73. The *noble* gases in group 18 on the periodic table used to be called the *inert* gases because they were thought to be incapable of forming compounds. Their name has been changed to noble gases because although they resist combining with the more common elements to their left on the periodic table, they do mingle with them on rare occasions. The following equations describe reactions that form xenon compounds. Determine the oxidation number for each atom in the reactions, and identify each reaction as redox or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

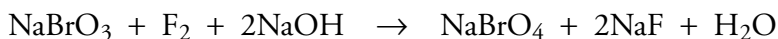
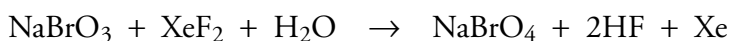


74. Sometimes one of the elements in a reactant appears in more than one product of a reaction and has in one product, a higher oxidation number than before the reaction and in the other product, a lower oxidation number than before the reaction. In this way, the same element is both oxidized and reduced, and the same compound is both the oxidizing agent and the reducing agent. This process is called disproportionation. For example, iodine monofluoride,  $\text{IF}$ , disproportionates into iodine and iodine pentafluoride in the following reaction:



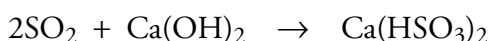
Determine the oxidation number for each atom in this equation and show that iodine is both oxidized and reduced and that iodine monofluoride is both the oxidizing agent and the reducing agent.

75. Sodium perbromate is an oxidizing agent that can be made in the two ways represented by the equations below. The first equation shows the way it was made in the past, and the second equation represents the technique used today.



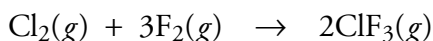
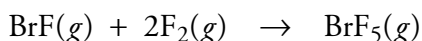
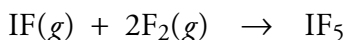
Determine the oxidation number for each atom in each of these equations, and decide whether each reaction is a redox reaction or not. If a reaction is a redox reaction, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

76. Calcium hydrogen sulfite,  $\text{Ca}(\text{HSO}_3)_2$ , which is used as a paper pulp preservative and as a disinfectant, is made by reacting sulfur dioxide with calcium hydroxide.



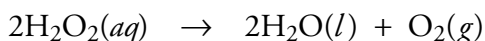
Determine the oxidation number for each atom in the equation and decide whether the reaction is a redox reaction or not. If the reaction is a redox reaction, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

77. The following equations represent reactions that involve only halogen atoms. Iodine pentafluoride,  $\text{IF}_5$ , is used to add fluorine atoms to other compounds, bromine pentafluoride,  $\text{BrF}_5$ , is an oxidizing agent in liquid rocket propellants, and chlorine trifluoride,  $\text{ClF}_3$ , is used to reprocess nuclear reactor fuels.



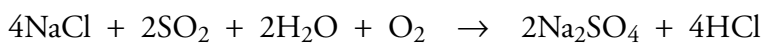
Determine the oxidation number for each atom in these equations, and decide whether each reaction is a redox reaction or not. If a reaction is a redox reaction, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

78. The water solution of hydrogen peroxide,  $\text{H}_2\text{O}_2$ , used as an antiseptic (3%) and bleach (6%) are stored in dark plastic bottles because the reaction below is accelerated by the metal ions found in glass and by light.



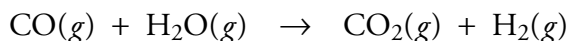
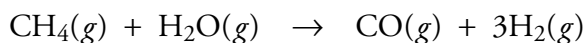
Determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized and what is reduced.

79. Sodium sulfate, which is used to make detergents and glass, is formed in the following reaction.



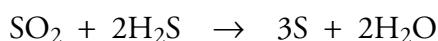
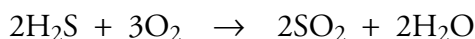
Determine the oxidation number for each atom in the equation and decide whether the reaction is a redox reaction or not. If a reaction is a redox reaction, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

80. Hydrogen gas can be made in two steps:



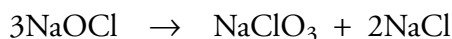
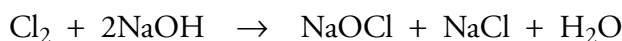
Determine the oxidation number for each atom in these equations and decide whether each reaction is a redox reaction or not. If a reaction is redox, identify which substance is oxidized and which substance is reduced. (You do not need to identify the oxidizing agent and the reducing agent.)

**81.** Elemental sulfur is produced by the chemical industry from naturally occurring hydrogen sulfide in the following steps.

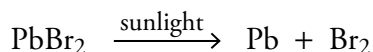


Determine the oxidation number for each atom in these equations and decide whether each reaction is a redox reaction or not. If a reaction is a redox reaction, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent.

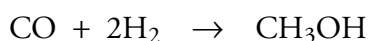
82. Sodium chlorate,  $\text{NaClO}_3$ , which is used to bleach paper, is made in the following reactions. Determine the oxidation number for each atom in the equations and identify whether each reaction is a redox reaction or not. If a reaction is redox, identify what is oxidized and what is reduced.



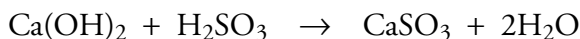
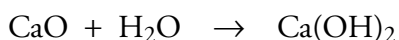
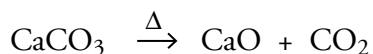
**83.** Leaded gasoline, originally developed to decrease pollution, is now banned because the lead(II) bromide,  $\text{PbBr}_2$ , emitted when it burns decomposes in the atmosphere into two serious pollutants, lead and bromine. The equation for this reaction is below. Determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized and what is reduced.



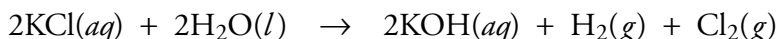
84. When leaded gasoline was banned, there was a rush to find safer ways to reduce emissions of unburned hydrocarbons from gasoline engines. One alternative is to add methyl t-butyl ether (MTBE). In 1990, about 25% of the methanol,  $\text{CH}_3\text{OH}$ , produced by the U.S. chemical industry was used to make methyl t-butyl ether. The equations below show the steps used to make methanol. Determine the oxidation number for each atom in the equation and identify whether the reactions are redox reactions or not. For each redox reaction, identify what is oxidized and what is reduced.



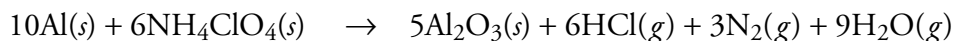
85. When the calcium carbonate,  $\text{CaCO}_3$ , in limestone is heated to a high temperature, it decomposes into calcium oxide (called lime or quick lime) and carbon dioxide. Lime was used by the early Romans, Greeks, and Egyptians to make cement and is used today to make over 150 different chemicals. In another reaction, calcium oxide and water form calcium hydroxide,  $\text{Ca(OH)}_2$  (called slaked lime), used to remove the sulfur dioxide from smoke stacks above power plants burning high-sulfur coal. The equations for all these reactions are below. Determine the oxidation number for each atom in the equation and identify whether the reactions are redox reaction or not. For each redox reaction, identify what is oxidized and what is reduced.



86. Potassium hydroxide, which is used to make fertilizers and soaps, is produced by running an electric current through a potassium chloride solution. The equation for this reaction is below. Is this a redox reaction? What is oxidized, and what is reduced?

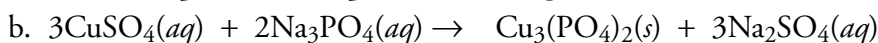
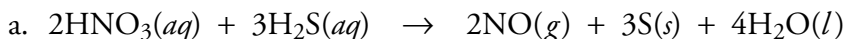


87. The space shuttle's solid rocket boosters get their thrust from the reaction of aluminum metal with ammonium perchlorate,  $\text{NH}_4\text{ClO}_4$ , which generates a lot of gas and heat. The billowy white smoke is due to the formation of very finely divided solid aluminum oxide. One of the reactions that takes place is

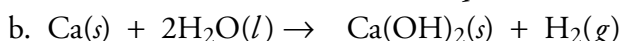
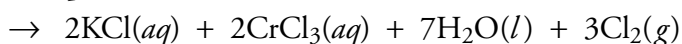
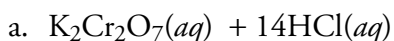


Is this a redox reaction? What is oxidized, and what is reduced?

88. For each of the following equations, determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.



89. Determine the oxidation number for each atom in the following equations and decide whether each reaction is a redox reaction or not. If a reaction is redox, identify which substance is oxidized, which is reduced, the oxidizing agent, and the reducing agent.



90. Determine the oxidation number for each atom in the following equations and decide whether each reaction is a redox reaction or not. If the reaction is redox, identify which substance is oxidized, which is reduced, the oxidizing agent, and the reducing agent.
- $2\text{Ag}_2\text{CrO}_4(s) + 4\text{HNO}_3(aq) \rightarrow 4\text{AgNO}_3(aq) + \text{H}_2\text{Cr}_2\text{O}_7(aq) + \text{H}_2\text{O}(l)$
  - $2\text{MnO}_4^-(aq) + 3\text{IO}_3^-(aq) + \text{H}_2\text{O}(l) \rightarrow 2\text{MnO}_2(s) + 3\text{IO}_4^-(aq) + 2\text{OH}^-(aq)$
91. For each of the following equations, determine the oxidation number for each atom in the equation and identify whether the reaction is a redox reaction or not. If the reaction is redox, identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.
- $\text{Ca}(s) + \text{F}_2(g) \rightarrow \text{CaF}_2(s)$
  - $2\text{Al}(s) + 3\text{H}_2\text{O}(g) \rightarrow \text{Al}_2\text{O}_3(s) + 3\text{H}_2(g)$
92. Determine the oxidation number for each atom in the following equations and decide whether each reaction is a redox reaction or not. If a reaction is redox, identify which substance is oxidized, which is reduced, the oxidizing agent, and the reducing agent.
- $\text{Cr}_2\text{O}_7^{2-}(aq) + 6\text{Cl}^-(aq) + 14\text{H}^+(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 3\text{Cl}_2(g) + 7\text{H}_2\text{O}(l)$
  - $5\text{H}_2\text{C}_2\text{O}_4(aq) + 2\text{KMnO}_4(aq) + 3\text{H}_2\text{SO}_4(aq) \rightarrow 10\text{CO}_2(g) + 2\text{MnSO}_4(aq) + 8\text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq)$
93. The following equations represent reactions used by the U.S. chemical industry. Classify each with respect to the following categories: combination reaction, decomposition reaction, and single-displacement reaction.
- $\text{P}_4 + 5\text{O}_2 + 6\text{H}_2\text{O} \rightarrow 4\text{H}_3\text{PO}_4$
  - $\text{TiCl}_4 + \text{O}_2 \rightarrow \text{TiO}_2 + 2\text{Cl}_2$
  - $\text{CH}_3\text{CH}_3(g) \xrightarrow{\Delta} \text{CH}_2\text{CH}_2(g) + \text{H}_2(g)$
94. The following equations represent reactions used by the U.S. chemical industry. Classify each with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.
- $2\text{HF} + \text{SiF}_4 \rightarrow \text{H}_2\text{SiF}_6$
  - $2\text{H}_2\text{S} + 3\text{O}_2 \rightarrow 2\text{SO}_2 + 2\text{H}_2\text{O}$
  - $\text{CH}_3\text{OH} \xrightarrow{\Delta} \text{CH}_2\text{O} + \text{H}_2$
95. Write a balanced equation for the redox reaction of carbon dioxide gas and hydrogen gas to form carbon solid and water vapor.
96. Phosphorus pentachloride, which is used to add chlorine atoms to other substances, can be made from the reaction of phosphorus trichloride and chlorine. The phosphorus pentachloride is the only product. Write a balance equation, without including states, for this redox reaction.



97. Titanium metal is used to make metal alloys for aircraft, missiles, and artificial hip joints. It is formed in the reaction of titanium(IV) chloride with magnesium metal. The other product is magnesium chloride. Write a balanced equation, without including states, for this redox reaction.
98. Dichlorine monoxide, which is used to add chlorine atoms to other substances, is made from mercury(II) oxide and chlorine. The products are dichlorine monoxide and mercury. Write a balanced equation, without including states, for this redox reaction.
99. Write a balanced equation for the redox reaction of solid potassium with liquid water to form aqueous potassium hydroxide and hydrogen gas.
100. Write a balanced equation for the redox reaction of aqueous chlorine with aqueous potassium iodide to form aqueous potassium chloride and solid iodine.
101. Write a balanced equation for the redox reaction of calcium metal and bromine liquid to form solid calcium bromide.
102. Write a balanced equation for the redox reaction of solid copper(II) sulfide with oxygen gas to form solid copper(II) oxide and sulfur dioxide gas.
103. Magnesium chloride is used to make disinfectants, fire extinguishers, paper, and floor sweeping compounds. It is made from the redox reaction of hydrochloric acid with solid magnesium hydroxide. Write a balanced equation for this reaction, which yields aqueous magnesium chloride and liquid water.
104. Write a balanced equation for the redox reaction at room temperature of chromium metal with hydrochloric acid to form aqueous chromium(III) chloride and hydrogen gas.

### Discussion Question

105. What makes one battery better than another? Find a reference book that tells you about the properties of the elements. Why do you think lithium batteries are superior to batteries that use lead? What other elements might be considered for new batteries?