UNIT

5

Electrochemistry

UNIT 5 CONTENTS

UNIT 5 OVERALL EXPECTATIONS

- What are oxidation-reduction reactions? How are they involved in the interconversion of chemical and electrical energy?
- How are galvanic and electrolytic cells built, and how do they function? What equations are used to describe these types of cells? How can you solve quantitative problems related to electrolysis?
- What are the uses of batteries and fuel cells? How is electrochemical technology used to produce and protect metals? How can you assess the environmental and safety issues associated with these technologies?

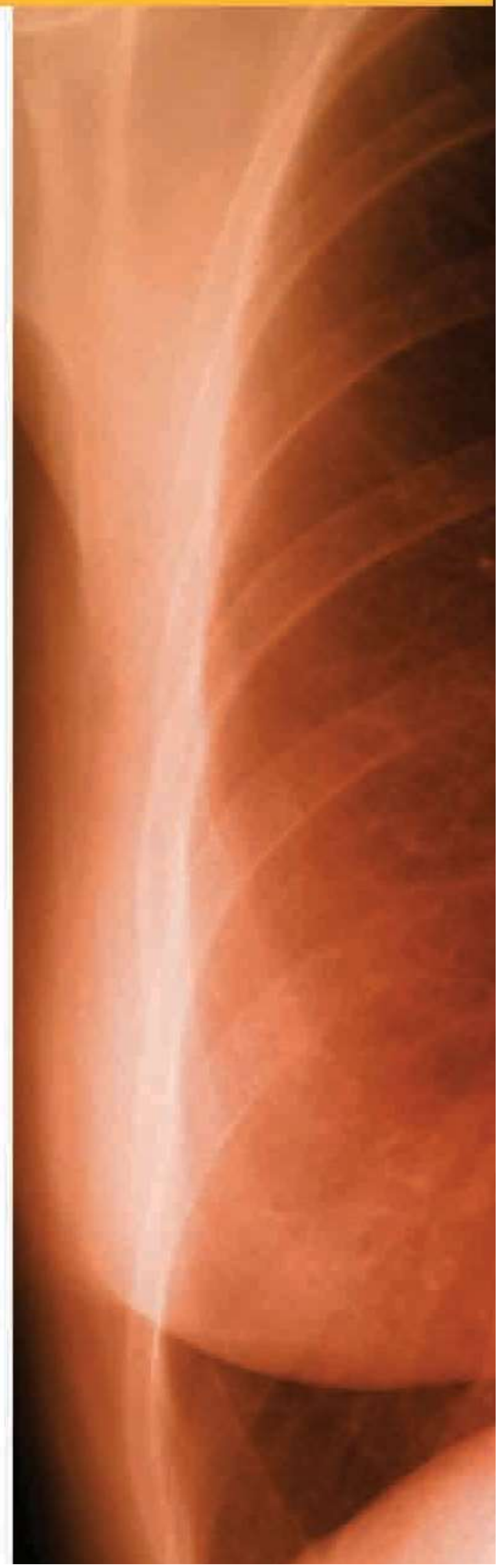
Unit Issue Prep

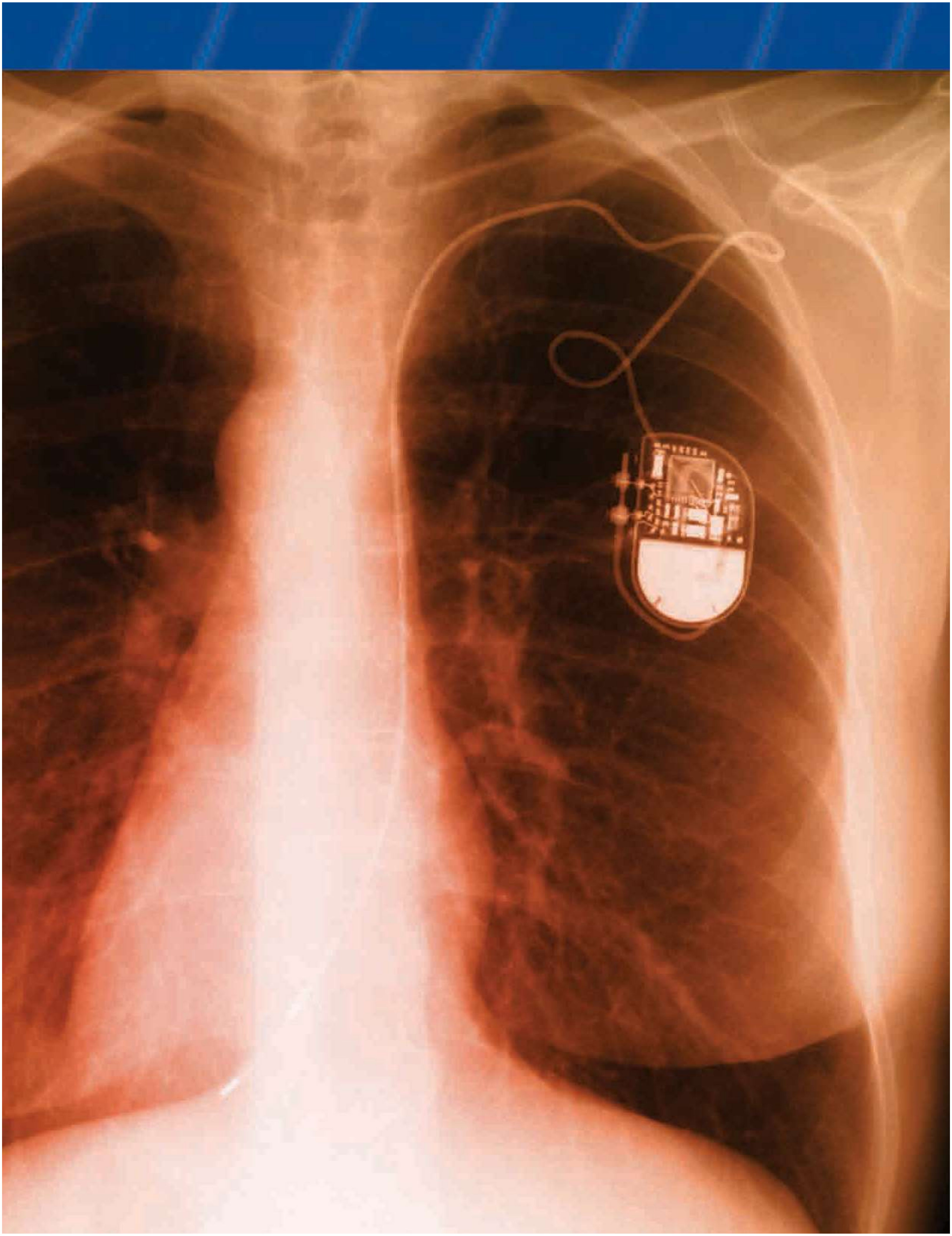
As you progress through Unit 5, look for the skills and information you will need for the investigation at the end of the unit.

Canadian engineer Dr. John Hopps, working with a team of medical researchers in the late 1940s, developed one of the most significant medical inventions of the twentieth century: the pacemaker. The photograph on the right shows a pacemaker embedded in the body of a heart patient. A modern pacemaker is essentially a tiny computer that monitors a person's heartbeat and corrects irregularities as needed. Pacemakers are particularly useful in correcting a heartbeat that is too slow.

The pacemaker device is surgically placed in a "pocket" of tissue near the patient's collarbone. One or more wires, called "leads," are connected to the pacemaker and threaded down through a major vein to the patient's heart. By sending electrical impulses along the leads to the heart, the pacemaker can induce a heartbeat.

A pacemaker obtains electrical energy from a tiny battery that lasts for about seven years before it must be replaced. But how do batteries supply electrical energy? The answer lies in a branch of chemistry known as electrochemistry. In this unit, you will learn about the connection between chemical reactions and electricity. You will also learn about the chemical reactions that take place inside batteries.





CHAPTER

10 Oxidation-Reduction Reactions

Chapter Preview

Prerequisite Concepts and Skills

Before you begin this chapter, review the following concepts and skills:

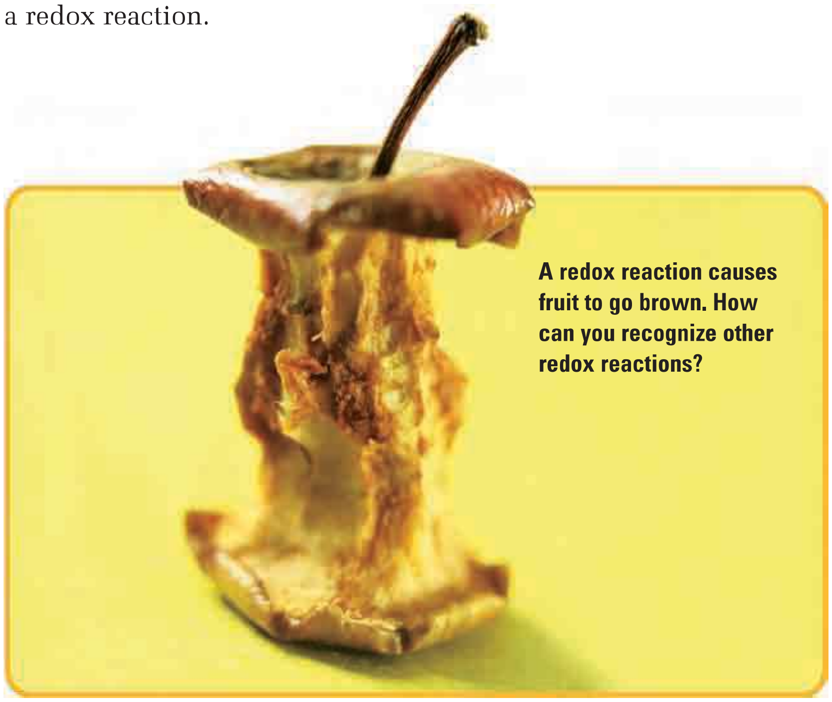
- balancing chemical, total ionic, and net ionic equations (Concepts and Skills Review)
- reaction types, including synthesis, decomposition, single displacement, and double displacement reactions (Concepts and Skills Review)
- the common ionic charges of metal ions and non-metal ions, and the formulas of common polyatomic ions
- drawing Lewis structures (Concepts and Skills Review)
- electronegativities and bond polarities (Chapter 4, section 4.1)

Kitchen chemistry is an important part of daily life. Cooks use chemistry all the time to prepare and preserve food. Even the simplest things you do in the kitchen can involve chemical reactions. For example, you have probably seen a sliced apple turn brown. The same thing happens to pears, bananas, avocados, and several other fruits. Slicing the fruit exposes the flesh to oxygen in the air. Compounds in the fruit react with oxygen to form brown products. An enzyme in the fruit acts as a catalyst, speeding up this reaction. How can you stop fruit from turning brown after it is sliced?

A Waldorf salad uses a simple method to prevent fruit from browning. This type of salad usually consists of diced apples, celery, and walnuts, covered with a mayonnaise dressing. The dressing keeps the air away from the food ingredients. Without air, the fruit does not turn brown.

Another way to solve this problem is to prevent the enzyme in the fruit from acting as a catalyst. Enzymes are sensitive to pH. Therefore, adding an acid such as lemon juice or vinegar to fruit can prevent the enzyme from acting. You may have noticed that avocado salad recipes often include lemon juice. In addition to hindering the enzyme, lemon juice contains vitamin C, which is very reactive toward oxygen. The vitamin C reacts with oxygen before the sliced fruit can do so.

In this chapter, you will be introduced to oxidation-reduction reactions, also called redox reactions. You will discover how to identify this type of reaction. You will also find out how to balance equations for a redox reaction.



A redox reaction causes fruit to go brown. How can you recognize other redox reactions?

Defining Oxidation and Reduction

10.1

The term *oxidation* can be used to describe the process in which certain fruits turn brown by reacting with oxygen. The original, historical definition of this term was "to combine with oxygen." Thus, oxidation occurred when iron rusted, and when magnesium was burned in oxygen gas. The term *reduction* was used historically to describe the opposite of oxidation, that is, the formation of a metal from its compounds. An **ore** is a naturally occurring solid compound or mixture of compounds from which a metal can be extracted. Thus, the process of obtaining a metal from an ore was known as a reduction. Copper ore was reduced to yield copper, and iron ore was reduced to yield iron.

As you will learn in this chapter, the modern definitions for oxidation and reduction are much broader. The current definitions are based on the idea of electron transfers, and can now be applied to numerous chemical reactions. In Unit 1, you saw the terms oxidation and reduction used to describe changes to carbon-hydrogen and carbon-oxygen bonds within organic compounds. These changes involve electron transfers, so the broader definitions that you will learn in this chapter still apply.

In your previous chemistry course, you compared the reactivities of metals. You may recall that, when a piece of zinc is placed in an aqueous solution of copper(II) sulfate, the zinc displaces the copper in a single displacement reaction. This reaction is shown in Figure 10.1. As the zinc dissolves, the zinc strip gets smaller. A dark red-brown layer of solid copper forms on the zinc strip, and some copper is deposited on the bottom of the beaker. The blue colour of the solution fades, as blue copper(II) ions are replaced by colourless zinc ions.

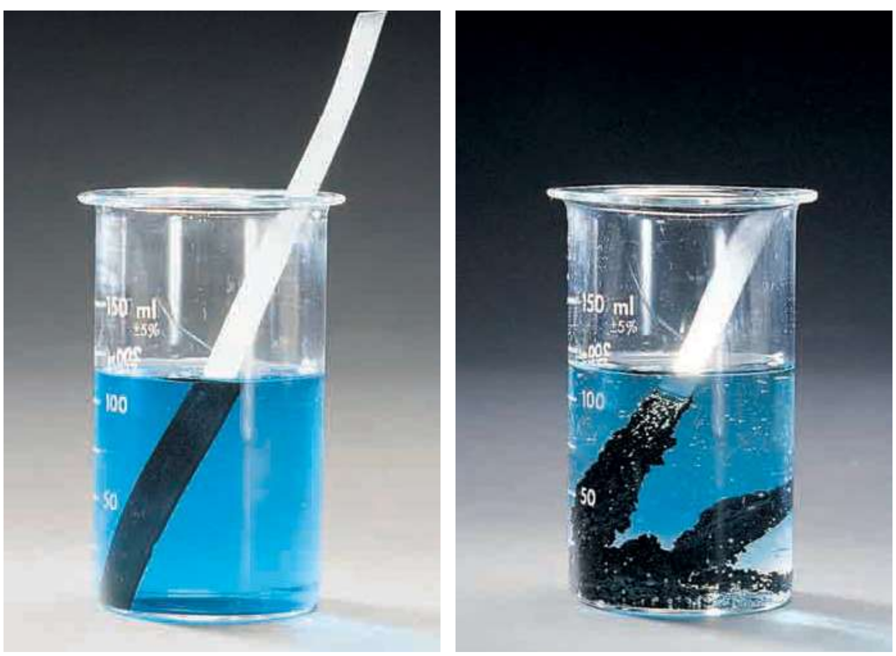


Figure 10.1 **A solid zinc strip reacts with a solution that contains blue copper(II) ions.**

Section Preview/ Specific Expectations

In this section, you will

- **describe** oxidation and reduction in terms of the loss and the gain of electrons
- **write** half-reactions from balanced chemical equations for oxidation-reduction systems
- **investigate** oxidation-reduction reactions by comparing the reactivities of some metals
- **communicate** your understanding of the terms *ore*, *oxidation*, *reduction*, *oxidation-reduction reaction*, *redox reaction*, *oxidizing agent*, *reducing agent*, *half-reaction*, *disproportionation*

CONCEPT CHECK

From your earlier work, you will recognize the sulfate ion, SO_4^{2-}, as a polyatomic ion. To review the names and formulas of common polyatomic ions, refer to Appendix E, Table E.5.

The reaction in Figure 10.1 is represented by the following equation.

$$Zn_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + ZnSO_{4(aq)}$$

This equation can be written as a total ionic equation.

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

The sulfate ions are *spectator ions*, meaning ions that are not involved in the chemical reaction. By omitting the spectator ions, you obtain the following net ionic equation.

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$$

Notice what happens to the reactants in this equation. The zinc atoms *lose* electrons to form zinc ions. The copper ions *gain* electrons to form copper atoms.

gains $2e^-$

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$$

loses $2e^-$

The following chemical definitions describe these changes.

- **Oxidation** is the loss of electrons.
- **Reduction** is the gain of electrons.

CHEM FACT

Try using a mnemonic to remember the definitions for oxidation and reduction. For example, in "LEO the lion says GER," LEO stands for "Loss of Electrons is Oxidation." GER stands for "Gain of Electrons is Reduction." The mnemonic "OIL RIG" stands for "Oxidation Is Loss. Reduction Is Gain." Make up your own mnemonic to help you remember these definitions.

In the reaction of zinc atoms with copper(II) ions, the zinc atoms lose electrons and undergo oxidation. In other words, the zinc atoms are *oxidized*. The copper(II) ions gain electrons and undergo reduction. In other words, the copper(II) ions are *reduced*. Because oxidation and reduction both occur in the reaction, it is known as an **oxidation-reduction reaction** or **redox reaction**.

Notice that electrons are transferred from zinc atoms to copper(II) ions. The copper(II) ions are responsible for the oxidation of the zinc atoms. A reactant that oxidizes another reactant is called an **oxidizing agent**. The oxidizing agent accepts electrons in a redox reaction. In this reaction, copper(II) is the oxidizing agent. The zinc atoms are responsible for the reduction of the copper(II) ions. A reactant that reduces another reactant is called a **reducing agent**. The reducing agent gives or donates electrons in a redox reaction. In this reaction, zinc is the reducing agent.

A redox reaction can also be defined as a reaction between an oxidizing agent and a reducing agent, as illustrated in Figure 10.2.

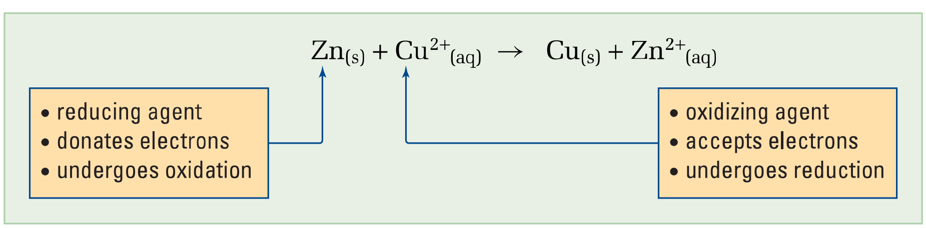


Figure 10.2 **In a redox reaction, the reducing agent is oxidized, and the oxidizing agent is reduced. Note that the oxidizing agent *does not* undergo oxidation, and that the reducing agent *does not* undergo reduction.**

Try the following practice problems to review your understanding of net ionic equations, and to work with the new concepts of oxidation and reduction.

Practice Problems

1. Write a balanced net ionic equation for the reaction of zinc with aqueous iron(II) chloride. Include the physical states of the reactants and products.
2. Write a balanced net ionic equation for each reaction, including physical states.
   (a) magnesium with aqueous aluminum sulfate
   (b) a solution of silver nitrate with metallic cadmium
3. Identify the reactant oxidized and the reactant reduced in each reaction in question 2.
4. Identify the oxidizing agent and the reducing agent in each reaction in question 2.

Half-Reactions

To monitor the transfer of electrons in a redox reaction, you can represent the oxidation and reduction separately. A **half-reaction** is a balanced equation that shows the number of electrons involved in either oxidation or reduction. Because a redox reaction involves both oxidation and reduction, two half-reactions are needed to represent a redox reaction. One half-reaction shows oxidation, and the other half-reaction shows reduction.

As you saw earlier, the reaction of zinc with aqueous copper(II) sulfate can be represented by the following net ionic equation.

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$$

Each neutral Zn atom is oxidized to form a Zn^{2+} ion. Thus, each Zn atom must lose two electrons. You can write an oxidation half-reaction to show this change.

$$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$$

Each Cu^{2+} ion is reduced to form a neutral Cu atom. Thus, each Cu^{2+} ion must gain two electrons. You can write a reduction half-reaction to show this change.

$$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$

If you look again at each half-reaction above, you will notice that the atoms and the charges are balanced. Like other types of balanced equations, half-reactions are balanced using the smallest possible whole-number coefficients. In the following equation, the atoms and charges are balanced, but the coefficients can all be divided by 2 to give the usual form of the half-reaction.

$$2Cu^{2+}{}_{(aq)} + 4e^- \rightarrow 2Cu_{(s)}$$

CONCEPT CHECK

You can write separate oxidation and reduction half-reactions to represent a redox reaction, but one half-reaction cannot occur on its own. Explain why this statement must be true.

In most redox reactions, one substance is oxidized and a different substance is reduced. In a **disproportionation** reaction, however, a single element undergoes both oxidation and reduction in the same reaction. For example, a copper(I) solution undergoes disproportionation in the following reaction.

$$2Cu^{+}_{(aq)} \rightarrow Cu_{(s)} + Cu^{2+}_{(aq)}$$

In this reaction, some copper(I) ions gain electrons, while other copper(I) ions lose electrons.

gains $1e^-$

$$Cu^{+}_{(aq)} + Cu^{+}_{(aq)} \rightarrow Cu_{(s)} + Cu^{2+}_{(aq)}$$

loses $1e^-$

The two half-reactions are as follows.

Oxidation: $Cu^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 1e^-$

Reduction: $Cu^{+}_{(aq)} + 1e^- \rightarrow Cu_{(s)}$

You have learned that half-reactions can be used to represent oxidation and reduction separately. Half-reactions always come in pairs: an oxidation half-reaction is always accompanied by a reduction half-reaction, and vice versa. Try writing and balancing half-reactions using the following practice problems.

Practice Problems

5. Write balanced half-reactions from the net ionic equation for the reaction between solid aluminum and aqueous iron(III) sulfate. The sulfate ions are spectator ions, and are not included.

$Al_{(s)} + Fe^{3+}_{(aq)} \rightarrow Al^{3+}_{(aq)} + Fe_{(s)}$

6. Write balanced half-reactions from the following net ionic equations.

(a) $Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

(b) $Cd_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cd^{2+}_{(aq)} + 2Ag_{(s)}$

7. Write balanced half-reactions for each of the following reactions.

(a) $Sn_{(s)} + PbCl_{2(aq)} \rightarrow SnCl_{2(aq)} + Pb_{(s)}$

(b) $Au(NO_3)_{3(aq)} + 3Ag_{(s)} \rightarrow 3AgNO_{3(aq)} + Au_{(s)}$

(c) $3Zn_{(s)} + Fe_2(SO_4)_{3(aq)} \rightarrow 3ZnSO_{4(aq)} + 2Fe_{(s)}$

8. Write the net ionic equation and the half-reactions for the disproportionation of mercury(I) ions in aqueous solution to give liquid mercury and aqueous mercury(II) ions. Assume that mercury(I) ions exist in solution as Hg_2^{2+}.

You already know that some metals are more reactive than others. You may also have carried out an investigation on the metal activity series in a previous course. In Investigation 10-A, located on page 470, you will discover how this series is related to oxidation and reduction. You will write chemical equations, ionic equations, and half-reactions for the single displacement reactions of several metals.

COURSE CHALLENGE

The Chemistry Bulletin, on the next page, introduces you to the terms *oxidant* and *antioxidant*. How may oxidants and antioxidants affect human health? Consider this question to prepare for your Chemistry Course Challenge.

Chemistry Bulletin

Science | Technology | Society | Environment

Aging: Is Oxidation a Factor?

Why do we grow old? Despite advances in molecular biology and medical research, the reasons for aging remain mysterious. One theory suggests that aging may be influenced by oxidizing agents, also known as *oxidants*.

Oxidants are present in the environment and in foods. Nitrogen oxides are oxidants present in cigarette smoke and urban smog. Other oxidants include the copper and iron salts in meat and some plants. Inhaling and ingesting oxidants such as these can increase the level of oxidants in our bodies.

Oxidants are also naturally present in the body, where they participate in important redox reactions. For example, mitochondria consume oxygen during aerobic respiration, and cells ingest and destroy bacteria. Both these processes involve oxidation and reduction.

As you have just seen, redox reactions are an essential part of your body's processes. However, these reactions can produce *free radicals*, which are highly reactive atoms or molecules with one or more unpaired electrons. Because they are so reactive, free radicals can oxidize surrounding molecules by robbing them of electrons. This process can damage DNA, proteins, and other macromolecules. Such damage may contribute to aging, and to diseases that are common among the aging, such as cancer, cardiovascular disease, and cataracts.

The study of oxidative damage has sparked a debate about the role that antioxidants might play in illness and aging. *Antioxidants* are reducing agents. They donate electrons to substances that have been oxidized, decreasing the damage caused by free radicals. Dietary antioxidants include vitamins C and E, beta-carotene, and carotenoids.

Most medical researchers agree that people with diets rich in fruits and vegetables have a lower incidence of cardiovascular disease, certain cancers, and cataracts. Although fruits and vegetables are high in antioxidants, they also contain fibre and many different vitamins

Carotenoids are pigments found in some fruits and vegetables, including spinach.

and plant chemicals. It is hard to disentangle the effects of antioxidants from the beneficial effects of these other substances.

As a result, the benefits of antioxidant dietary supplements are under debate. According to one study, vitamin E supplements may lower the risk of heart disease. Another study, however, concludes that taking beta-carotene supplements does *not* reduce the risk of certain cancers.

We can be sure that a balanced diet including fruits and vegetables is beneficial to human health. Whether antioxidants confer these benefits, and whether these benefits include longevity, remain to be seen.

Making Connections

1. Research vitamins C, E, alpha- and beta-carotenes, and folic acid. How do they affect our health? What fruits and vegetables contain these vitamins?
2. Lycopene is a carotenoid that has been linked to a decreased risk of pancreatic, cervical, and prostate cancer. Find out what fruits and vegetables contain lycopene. What colour are these fruits and vegetables?

MICROSCALE

Investigation 10-A

SKILL FOCUS

- Predicting
- Performing and recording
- Analyzing and interpreting

Single Displacement Reactions

The metal activity series is shown in the table below. The more reactive metals are near the top of the series, and the less reactive metals are near the bottom. In this investigation, you will relate the activity series to the ease with which metals are oxidized and metal ions are reduced.

Activity Series of Metals

Metal	
lithium	Most Reactive
potassium	
barium	
calcium	
sodium	
magnesium	
aluminum	
zinc	
chromium	
iron	
cadmium	
cobalt	
nickel	
tin	
lead	
copper	
mercury	
silver	
platinum	
gold	Least Reactive

Question

How is the order of the metals in the activity series related to the ease with which metals are oxidized and metal ions are reduced?

Predictions

Predict the relative ease with which the metals aluminum, copper, iron, magnesium, and zinc can be oxidized. Predict the relative ease with which the ions of these same metals can be reduced. Explain your reasoning in both cases.

Materials

well plate
test tube rack
4 small test tubes
4 small pieces of each of these metals:
aluminum foil, thin copper wire or tiny copper beads, iron filings, magnesium, and zinc
dropper bottles containing dilute solutions of aluminum sulfate, copper(II) sulfate, iron(II) sulfate, magnesium sulfate, and zinc nitrate

Safety Precautions

- Wear goggles, gloves, and an apron for all parts of this investigation.

Procedure

1. Place the well plate on a white piece of paper. Label it to match the table on the next page.
2. In each well plate, place a small piece of the appropriate metal, about the size of a grain of rice. Cover each piece with a few drops of the appropriate solution. Wait 3–5 min to observe if a reaction occurs.
3. Look for evidence of a chemical reaction in each mixture. Record the results, using “y” for a reaction, or “n” for no reaction. If you are unsure, repeat the process on a larger scale in a small test tube.

Metal \ Compound	$Al_2(SO_4)_3$	$CuSO_4$	$FeSO_4$	$MgSO_4$	$Zn(NO_3)_2$
Al					
Cu					
Fe					
Mg					
Zn					

4. Discard the mixtures in the waste beaker supplied by your teacher. Do not pour anything down the drain.

Analysis

1. For each single displacement reaction you observed, write
   (a) a balanced chemical equation
   (b) a total ionic equation
   (c) a net ionic equation
2. Write an oxidation half-reaction and a reduction half-reaction for each net ionic equation you wrote in question 1. Use the smallest possible whole-number coefficients in each half-reaction.
3. Look at each balanced net ionic equation. Compare the total number of electrons lost by the reducing agent with the total number of electrons gained by the oxidizing agent.
4. List the different oxidation half-reactions. Start with the half-reaction for the most easily oxidized metal, and end with the half-reaction for the least easily oxidized metal. Explain your reasoning. Compare your list with your first prediction from the beginning of this investigation.
5. List the different reduction half-reactions. Start with the half-reaction for the most easily reduced metal ion, and end with the half-reaction for the least easily reduced metal ion. Explain your reasoning. Compare your list with your second prediction from the beginning of this investigation.

Conclusions

6. Which list from questions 4 and 5 puts the metals in the same order as they appear in the activity series?
7. How is the order of the metals in the activity series related to the ease with which metals are oxidized and metal ions are reduced?

Applications

8. Use the activity series to choose a reducing agent that will reduce aqueous nickel(II) ions to metallic nickel. Explain your reasoning.
9. Use the activity series to choose an oxidizing agent that will oxidize metallic cobalt to form aqueous cobalt(II) ions. Explain your reasoning.

Section Summary

In this section, you learned to define and recognize redox reactions, and to write oxidation and reduction half-reactions. In Investigation 10-A, you observed the connection between the metal activity series and redox reactions. However, thus far, you have only worked with redox reactions that involve atoms and ions as reactants or products. In the next section, you will learn about redox reactions that involve covalent reactants or products.

Section Review

1. K/U Predict whether each of the following single displacement reactions will occur. If so, write a balanced chemical equation, a balanced net ionic equation, and two balanced half-reactions. Include the physical states of the reactants and products in each case.
   - (a) aqueous silver nitrate and metallic cadmium
   - (b) gold and aqueous copper(II) sulfate
   - (c) aluminum and aqueous mercury(II) chloride

2. (a) K/U On which side of an oxidation half-reaction are the electrons? Why?
   - (b) K/U On which side of a reduction half-reaction are the electrons? Why?

3. C Explain why, in a redox reaction, the oxidizing agent undergoes reduction.

4. C In a combination reaction, does metallic lithium act as an oxidizing agent or a reducing agent? Explain.

5. I Write a net ionic equation for a reaction in which
   - (a) Fe^{2+} acts as an oxidizing agent
   - (b) Al acts as a reducing agent
   - (c) Au^{3+} acts as an oxidizing agent
   - (d) Cu acts as a reducing agent
   - (e) Sn^{2+} acts as an oxidizing agent and as a reducing agent

6. MC The element potassium is made industrially by the single displacement reaction of molten sodium with molten potassium chloride.
   - (a) Write a net ionic equation for the reaction, assuming that all reactants and products are in the liquid state.
   - (b) Identify the oxidizing agent and the reducing agent in the reaction.
   - (c) Explain why the reaction is carried out in the liquid state and not in aqueous solution.

Unit Investigation Prep

In the end-of-unit investigation, you will be working with the metals zinc and copper. Which metal is more easily oxidized? Which is more easily reduced?

Oxidation Numbers

10.2

Section Preview/ Specific Expectations

In this section, you will

- **describe** oxidation and reduction in terms of changes in oxidation number
- **assign** oxidation numbers to elements in covalent molecules and polyatomic ions
- **identify** redox reactions using oxidation numbers
- **communicate** your understanding of the terms *oxidation numbers, oxidation, reduction*

Redox reactions are very common. Some of them produce light in a process known as *chemiluminescence*. In living things, the production of light in redox reactions is known as *bioluminescence*. You can actually see the light from redox reactions occurring in some organisms, such as glowworms and fireflies, as shown in Figure 10.3.

Figure 10.3 **Fireflies use flashes of light produced by redox reactions to attract a mate.**

Not all redox reactions give off light, however. How can you recognize a redox reaction, and how can you identify the oxidizing and reducing agents? In section 10.1, you saw net ionic equations with monatomic elements, such as Cu and Zn, and with ions containing a single element, such as Cu^{2+} and Zn^{2+}. In these cases, you could use ionic charges to describe the transfer of electrons. However, many redox reactions involve reactants or products with covalent bonds, including elements that exist as covalent molecules, such as oxygen, O_2; covalent compounds, such as water, H_2O; or polyatomic ions that are not spectator ions, such as permanganate, MnO_4^-. For reactions involving covalent reactants and products, you cannot use ionic charges to describe the transfer of electrons.

Oxidation numbers are actual or hypothetical charges, assigned using a set of rules. They are used to describe redox reactions with covalent reactants or products. They are also used to identify redox reactions, and to identify oxidizing and reducing agents. In this section, you will see how oxidation numbers were developed from Lewis structures, and then learn the rules to assign oxidation numbers.

CHEM FACT

Oxidation numbers are just a bookkeeping method used to keep track of electron transfers. In a covalent molecule or a polyatomic ion, the oxidation number of each element does *not* represent an ionic charge, because the elements are not present as ions. However, to assign oxidation numbers to the elements in a covalent molecule or polyatomic ion, you can *pretend* the bonds are ionic.

Oxidation Numbers from Lewis Structures

You are probably familiar with the Lewis structure of water, shown in Figure 10.4A. From the electronegativities on the periodic table in Figure 10.5, on the next page, you can see that oxygen (electronegativity 3.44) is more electronegative than hydrogen (electronegativity 2.20). The electronegativity difference is less than 1.7, so the two hydrogen-oxygen bonds are polar covalent, not ionic. In each bond, the electrons are more strongly attracted to the oxygen atom than to the hydrogen atom.

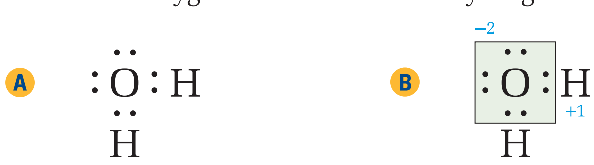


Figure 10.4 **(A) The Lewis structure of water; (B) The formal counting of electrons with the more electronegative element assigned a negative charge**

To assign oxidation numbers to the atoms in a water molecule, you can consider all the bonding electrons to be "owned" by the more electronegative oxygen atom, as shown in Figure 10.4B. Thus, each hydrogen atom in a water molecule is considered to have no electrons, as hydrogen would in a hydrogen ion, H^+. Therefore, the element hydrogen is assigned an oxidation number of +1 in water. On the other hand, the oxygen atom in a water molecule is considered to have a filled octet of electrons, as oxygen would in an oxide ion, O^{2-}. Therefore, the element oxygen is assigned an oxidation number of −2 in water. (**Note:** These are *not* ionic charges, since water is a covalent molecule. Also, note that the plus or minus sign in an oxidation number, such as −2, is written *before* the number. The plus or minus sign in an ionic charge, such as 2−, is written *after* the number.)

H 2.20																	He -
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne -
Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar -
K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr -
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 2.10	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe -
Cs 0.79	Ba 0.89	Lu 1.0	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn -
Fr 0.7	Ra 0.9	Lr -	Rf -	Db -	Sg -	Bh -	Hs -	Mt -	Uun -	Uuu -	Uub -	-	Uuq -	-	Uuh -	-	Uuo -

La 1.10	Ce 1.12	Pr 1.13	Nd 1.14	Pm -	Sm 1.17	Eu -	Gd 1.20	Tb -	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb -
Ac 1.1	Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am -	Cm -	Bk -	Cf -	Es -	Fm -	Md -	No -

Figure 10.5 **The periodic table, showing electronegativity values**

In a chlorine molecule, Cl_2, each atom has the same electronegativity, so the bond is non-polar covalent. Because the electrons are equally shared, you can consider each chlorine atom to "own" one of the shared electrons, as shown in Figure 10.6. Thus, each chlorine atom in the molecule is considered to have the same number of electrons as a neutral chlorine atom. Each chlorine atom is therefore assigned an oxidation number of 0.

Figure 10.6 **(A) The Lewis structure of a chlorine molecule; (B) The formal counting of electrons in a chlorine molecule for oxidation number purposes**

Figure 10.7 shows how oxidation numbers are assigned for the polyatomic cyanide ion, CN^-. The electronegativity of nitrogen (3.04) is greater than the electronegativity of carbon (2.55). Thus, the three shared pairs of electrons are all considered to belong to the nitrogen atom. As a result, the carbon atom is considered to have two valence electrons, which is two electrons less than the four valence electrons of a neutral carbon atom. Therefore, the carbon atom in CN^- is assigned an oxidation number of +2. The nitrogen atom is considered to have eight valence electrons, which is three electrons more than the five valence electrons of a neutral nitrogen atom. Therefore, the nitrogen atom in CN^- is assigned an oxidation number of −3.

Figure 10.7 **(A) The Lewis structure of a cyanide ion; (B) The formal counting of electrons in a cyanide ion for oxidation number purposes**

You have seen examples of how Lewis structures can be used to assign oxidation numbers for polar molecules such as water, non-polar molecules such as chlorine, and polar polyatomic ions such as the cyanide ion. In the following ThoughtLab, you will use Lewis structures to assign oxidation number values, and then look for patterns in your results.

ThoughtLab Finding Rules for Oxidation Numbers

Procedure

1. Use Lewis structures to assign an oxidation number to each element in the following covalent molecules.
   (a) HI **(b)** O_2 **(c)** PCl_5 **(d)** BBr_3
2. Use Lewis structures to assign an oxidation number to each element in the following polyatomic ions.
   (a) OH^- **(b)** NH_4^+ **(c)** CO_3^{2-}
3. Assign an oxidation number to each of the following atoms or monatomic ions. Explain your reasoning.
   (a) Ne **(b)** K **(c)** I^- **(d)** Mg^{2+}

Analysis

1. For each molecule in question 1 of the procedure, find the sum of the oxidation numbers of all the atoms present. What do you notice? Explain why the observed sum must be true for a neutral molecule.
2. For each polyatomic ion in question 2 of the procedure, find the sum of the oxidation numbers of all the atoms present. Describe and explain any pattern you see.

Extension

3. Predict the sum of the oxidation numbers of the atoms in the hypochlorite ion, OCl^-.
4. Test your prediction from question 3.

Using Rules to Find Oxidation Numbers

Drawing Lewis structures to assign oxidation numbers can be a very time-consuming process for large molecules or large polyatomic ions. Instead, the results from Lewis structures have been summarized to produce a more convenient set of rules, which can be applied more quickly. Table 10.1 summarizes the rules used to assign oxidation numbers. You may have discovered some of these rules for yourself in the ThoughtLab you just completed.

Table 10.1 Oxidation Number Rules

Rules	Examples
1. A pure element has an oxidation number of 0.	Na in $Na_{(s)}$, Br in $Br_{2(\ell)}$, and P in $P_{4(s)}$ all have an oxidation number of 0.
2. The oxidation number of an element in a monatomic ion equals the charge of the ion.	The oxidation number of Al in Al^{3+} is +3. The oxidation number of Se in Se^{2-} is −2.
3. The oxidation number of hydrogen in its compounds is +1, except in metal hydrides, where the oxidation number of hydrogen is −1.	The oxidation number of H in H_2S or CH_4 is +1. The oxidation number of H in NaH or in CaH_2 is −1.
4. The oxidation number of oxygen in its compounds is usually −2, but there are exceptions. These include peroxides, such as H_2O_2, and the compound OF_2.	The oxidation number of O in Li_2O or in KNO_3 is −2.
5. In covalent compounds that do not contain hydrogen or oxygen, the more electronegative element is assigned an oxidation number that equals the negative charge it usually has in its ionic compounds.	The oxidation number of Cl in PCl_3 is −1. The oxidation number of S in CS_2 is −2.
6. The sum of the oxidation numbers of all the elements in a compound is 0.	In CF_4, the oxidation number of F is −1, and the oxidation number of C is +4. $(+4) + 4(-1) = 0$
7. The sum of the oxidation numbers of all the elements in a polyatomic ion equals the charge on the ion.	In NO_2^-, the oxidation number of O is −2, and the oxidation number of N is +3. $(+3) + 2(-2) = -1$

Some oxidation numbers found using these rules are not integers. For example, an important iron ore called magnetite has the formula Fe_3O_4. Using the oxidation number rules, you can assign oxygen an oxidation number of −2, and calculate an oxidation number of $+\frac{8}{3}$ for iron. However, magnetite contains no iron atoms with this oxidation number. It actually contains iron(III) ions and iron(II) ions in a 2:1 ratio. The formula of magnetite is sometimes written as $Fe_2O_3 \bullet FeO$ to indicate that there are two different oxidation numbers. The value $+\frac{8}{3}$ for the oxidation number of iron is an average value.

$$\frac{2(+3) + (+2)}{3} = +\frac{8}{3}$$

Even though some oxidation numbers found using these rules are averages, the rules are still useful for monitoring electron transfers in redox reactions.

In the following Sample Problem, you will find out how to apply these rules to covalent molecules and polyatomic ions.

Sample Problem

Assigning Oxidation Numbers

Problem

Assign an oxidation number to each element.

(a) $SiBr_4$ **(b)** $HClO_4$ **(c)** $Cr_2O_7^{2-}$

Solution

(a)
- Because the compound $SiBr_4$ does not contain hydrogen or oxygen, rule 5 applies. Because $SiBr_4$ is a compound, rule 6 also applies.
- Silicon has an electronegativity of 1.90. Bromine has an electronegativity of 2.96. From rule 5, therefore, you can assign bromine an oxidation number of −1.
- The oxidation number of silicon is unknown, so let it be x. You know from rule 6 that the sum of the oxidation numbers is 0. Then,

$$x + 4(-1) = 0$$
$$x - 4 = 0$$
$$x = 4$$

The oxidation number of silicon is +4. The oxidation number of bromine is −1.

(b)
- Because the compound $HClO_4$ contains hydrogen and oxygen, rules 3 and 4 apply. Because $HClO_4$ is a compound, rule 6 also applies.
- Hydrogen has its usual oxidation number of +1. Oxygen has its usual oxidation number of −2. The oxidation number of chlorine is unknown, so let it be x. You know from rule 6 that the sum of the oxidation numbers is 0. Then,

$$(+1) + x + 4(-2) = 0$$
$$x - 7 = 0$$
$$x = 7$$

The oxidation number of hydrogen is +1. The oxidation number of chlorine is +7. The oxidation number of oxygen is −2.

(c)
- Because the polyatomic ion $Cr_2O_7^{2-}$ contains oxygen, rule 4 applies. Because $Cr_2O_7^{2-}$ is a polyatomic ion, rule 7 also applies.
- Oxygen has its usual oxidation number of −2.
- The oxidation number of chromium is unknown, so let it be x. You know from rule 7 that the sum of the oxidation numbers is −2. Then,

$$2x + 7(-2) = -2$$
$$2x - 14 = -2$$
$$2x = 12$$
$$x = 6$$

The oxidation number of chromium is +6. The oxidation number of oxygen is −2.

PROBLEM TIP

When finding the oxidation numbers of elements in ionic compounds, you can work with the ions separately. For example, $Na_2Cr_2O_7$ contains two Na^+ ions, and so sodium has an oxidation number of +1. The oxidation numbers of Cr and O can then be calculated as shown in part (c) of the Sample Problem.

Practice Problems

9. Determine the oxidation number of the specified element in each of the following.

(a) N in NF_3 **(b)** S in S_8 **(c)** Cr in CrO_4^{2-}

(d) P in P_2O_5 **(e)** C in $C_{12}H_{22}O_{11}$ **(f)** C in $CHCl_3$

10. Determine the oxidation number of each element in each of the following.

(a) H_2SO_3 **(b)** OH^- **(c)** HPO_4^{2-}

11. As stated in rule 4, oxygen does not always have its usual oxidation number of −2. Determine the oxidation number of oxygen in each of the following.

(a) the compound oxygen difluoride, OF_2 **(b)** the peroxide ion, O_2^{2-}

12. Determine the oxidation number of each element in each of the following ionic compounds by considering the ions separately. **Hint:** One formula unit of the compound in part (c) contains two identical monatomic ions and one polyatomic ion.

(a) $Al(HCO_3)_3$ **(b)** $(NH_4)_3PO_4$ **(c)** $K_2H_3IO_6$

Applying Oxidation Numbers to Redox Reactions

You have seen that the single displacement reaction of zinc with copper(II) sulfate is a redox reaction, represented by the following chemical equation and net ionic equation.

$$Zn_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + ZnSO_{4(aq)}$$
$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$$

Each atom or ion shown in the net ionic equation can be assigned an oxidation number. Zn has an oxidation number of 0; Cu^{2+} has an oxidation number of +2; Cu has an oxidation number of 0; and Zn^{2+} has an oxidation number of +2. Thus, there are changes in oxidation numbers in this reaction. The oxidation number of zinc increases, while the oxidation number of copper decreases.

oxidation number increases
(loss of electrons)

$$\underset{0}{Zn} + \underset{+2}{Cu^{2+}} \rightarrow \underset{+2}{Zn^{2+}} + \underset{0}{Cu}$$

oxidation number decreases
(gain of electrons)

In the oxidation half-reaction, the element zinc undergoes an increase in its oxidation number from 0 to +2.

$$\underset{0}{Zn} \rightarrow \underset{+2}{Zn^{2+}} + 2e^-$$

In the reduction half-reaction, the element copper undergoes a decrease in its oxidation number from +2 to 0.

$$\underset{+2}{Cu^{2+}} + 2e^- \rightarrow \underset{0}{Cu}$$

Therefore, you can describe oxidation and reduction as follows. (Also see Figure 10.8.)

- **Oxidation** is an increase in oxidation number.
- **Reduction** is a decrease in oxidation number.

You can also monitor changes in oxidation numbers in reactions that involve covalent molecules. For example, oxidation number changes occur in the reaction of hydrogen and oxygen to form water.

$$\underset{0}{2H_{2(g)}} + \underset{0}{O_{2(g)}} \rightarrow \underset{+1\ -2}{2H_2O_{(\ell)}}$$

Because hydrogen combines with oxygen in this reaction, hydrogen undergoes oxidation, according to the historical definition given at the beginning of section 10.1. Hydrogen also undergoes oxidation according to the modern definition, because the oxidation number of hydrogen increases from 0 to +1. Hydrogen is the reducing agent in this reaction. The oxygen undergoes reduction, because its oxidation number decreases from 0 to −2. Oxygen is the oxidizing agent in this reaction.

The following Sample Problem illustrates how to use oxidation numbers to identify redox reactions, oxidizing agents, and reducing agents.

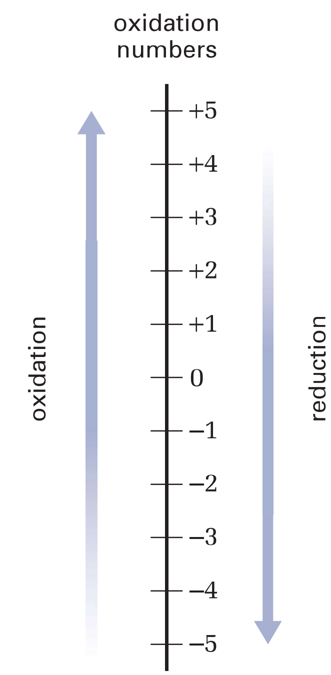


Figure 10.8 Oxidation and reduction are directly related to changes in oxidation numbers.

Sample Problem

Identifying Redox Reactions

Problem

Determine whether each of the following reactions is a redox reaction. If so, identify the oxidizing agent and the reducing agent.

(a) $CH_{4(g)} + Cl_{2(g)} \rightarrow CH_3Cl_{(g)} + HCl_{(g)}$

(b) $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(\ell)} + CO_{2(g)}$

Solution

Find the oxidation number of each element in the reactants and products. Identify any elements that undergo an increase or a decrease in oxidation number during the reaction.

(a) The oxidation number of each element in the reactants and products is as shown.

$$\underset{-4\ +1}{CH_{4(g)}} + \underset{0}{Cl_{2(g)}} \rightarrow \underset{-2\ +1\ -1}{CH_3Cl_{(g)}} + \underset{+1\ -1}{HCl_{(g)}}$$

- The oxidation number of hydrogen is +1 on both sides of the equation, so hydrogen is neither oxidized nor reduced.
- Both carbon and chlorine undergo changes in oxidation number, so the reaction is a redox reaction.
- The oxidation number of carbon increases from −4 to −2. The carbon atoms on the reactant side exist in methane molecules, $CH_{4(g)}$, so methane is oxidized. Therefore, methane is the reducing agent.
- The oxidation number of chlorine decreases from 0 to −1, so elemental chlorine, $Cl_{2(g)}$, is reduced. Therefore, elemental chlorine is the oxidizing agent.

(b) Because this reaction involves ions, write the equation in its total ionic form.

$$CaCO_{3(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^-_{(aq)} + H_2O_{(\ell)} + CO_{2(g)}$$

Continued ...

PROBLEM TIPS

- Use the fact that the sum of the oxidation numbers in a molecule is zero to check the assignment of the oxidation numbers.
- Make sure that a reaction does not include only a reduction or only an oxidation. Oxidation and reduction must occur together in a redox reaction.

CONCEPT CHECK

In part (b) of the Sample Problem, you can assign oxidation numbers to each element in the given chemical equation *or* in the net ionic equation. What are the advantages and the disadvantages of each method?

CONCEPT CHECK

In your previous chemistry course, you classified reactions into four main types: synthesis, decomposition, single displacement, and double displacement. You also learned to recognize combustion reactions and neutralization reactions. You have now learned to classify redox reactions. In addition, you have also learned about a special type of redox reaction known as a disproportionation reaction.

1. Classify each reaction in two ways.
   (a) magnesium reacting with a solution of iron(II) nitrate
   $Mg + Fe(NO_3)_2 \rightarrow Fe + Mg(NO_3)_2$
   (b) hydrogen sulfide burning in oxygen
   $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$
   (c) calcium reacting with chlorine
   $Ca + Cl_2 \rightarrow CaCl_2$
2. Classify the formation of water and oxygen from hydrogen peroxide in three ways.
   $2H_2O_2 \rightarrow 2H_2O + O_2$

Continued ...

The chloride ions are spectator ions, which do not undergo oxidation or reduction. The net ionic equation is as follows.

$CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + H_2O_{(\ell)} + CO_{2(g)}$

For the net ionic equation, the oxidation number of each element in the reactants and products is as shown.

$$\underset{+2\ +4\ -2}{CaCO_{3(s)}} + \underset{+1}{2H^+_{(aq)}} \rightarrow \underset{+2}{Ca^{2+}_{(aq)}} + \underset{+1\ -2}{H_2O_{(\ell)}} + \underset{+4\ -2}{CO_{2(g)}}$$

No elements undergo changes in oxidation numbers, so the reaction is not a redox reaction.

Practice Problems

13. Determine whether each reaction is a redox reaction.
    (a) $H_2O_2 + 2Fe(OH)_2 \rightarrow 2Fe(OH)_3$
    (b) $PCl_3 + 3H_2O \rightarrow H_3PO_3 + 3HCl$
14. Identify the oxidizing agent and the reducing agent for the redox reaction(s) in the previous question.
15. For the following balanced net ionic equation, identify the reactant that undergoes oxidation and the reactant that undergoes reduction.
    $Br_2 + 2ClO_2^- \rightarrow 2Br^- + 2ClO_2$
16. Nickel and copper are two metals that are important to the Ontario economy, particularly in the Sudbury area. Nickel and copper ores usually contain the metals as sulfides, such as NiS and Cu_2S. Do the extractions of these pure elemental metals from their ores involve redox reactions? Explain your reasoning.

Section Summary

In this section, you extended your knowledge of redox reactions to include covalent reactants and products. You did this by learning how to assign oxidation numbers and how to use them to recognize redox reactions, oxidizing agents, and reducing agents. In the next section, you will extend your knowledge further by learning how to write balanced equations that represent redox reactions.

Section Review

1 C At the beginning of section 10.1, it was stated that oxidation originally meant "to combine with oxygen." Explain why a metal that combines with the element oxygen undergoes oxidation as we now define it. What happens to the oxygen in this reaction? Write a balanced chemical equation for a reaction that illustrates your answer.

2 K/U Determine whether each of the following reactions is a redox reaction.

(a) $H_2 + I_2 \rightarrow 2HI$

(b) $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

(c) $2HBr + Ca(OH)_2 \rightarrow CaBr_2 + 2H_2O$

(d) $PCl_5 \rightarrow PCl_3 + Cl_2$

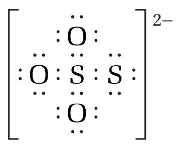
3 K/U Write three different definitions for a redox reaction.

4 C Explain why fluorine has an oxidation number of −1 in all its compounds.

5 C When an element combines with another element, is the reaction a redox reaction? Explain your answer.

6 **(a)** I Use the oxidation number rules to find the oxidation number of sulfur in a thiosulfate ion, $S_2O_3{}^{2-}$.

(b) The Lewis structure of a thiosulfate ion is given here. Use the Lewis structure to find the oxidation number of each sulfur atom.



(c) Compare your results from parts (a) and (b) and explain any differences.

(d) What are the advantages and disadvantages of using Lewis structures to assign oxidation numbers?

(e) What are the advantages and disadvantages of using the oxidation number rules to assign oxidation numbers?

7 **(a)** MC The Haber Process for the production of ammonia from nitrogen and hydrogen is a very important industrial process. Write a balanced chemical equation for the reaction. Use oxidation numbers to identify the oxidizing agent and the reducing agent.

(b) When ammonia is reacted with nitric acid to make the common fertilizer ammonium nitrate, is the reaction a redox reaction? Explain. (**Hint:** Consider the two polyatomic ions in the product separately.)

8 MC Historically, the extraction of a metal from its ore was known as reduction. One way to reduce iron ore on an industrial scale is to use a huge reaction vessel, 30 m to 40 m high, called a blast furnace. The reactants in a blast furnace are an impure iron ore, such as Fe_2O_3, mixed with limestone, $CaCO_3$, and coke, C, which is made from coal. The solid mixture is fed into the top of the blast furnace. A blast of very hot air, at about 900°C, is blown in near the bottom of the furnace. The following reactions occur.

$2C + O_2 \rightarrow 2CO$

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

The limestone is present to convert sand or quartz, SiO_2, which is present as an impurity in the ore, to calcium silicate, $CaSiO_3$.

$CaCO_3 \rightarrow CaO + CO_2$

$CaO + SiO_2 \rightarrow CaSiO_3$

(a) Which of the four reactions above are redox reactions?

(b) For each redox reaction that you identified in part (a), name the oxidizing agent and the reducing agent.

CHEM FACT

Redox reactions are involved in some very important industrial processes, such as iron and steel production. However, the widespread use of metals has occupied a relatively small part of human history. In the Stone Age, humans relied on stone, wood, and bone to make tools and weapons. The Stone Age ended in many parts of the world with the start of the Bronze Age, which was marked by the use of copper and then bronze (an alloy of copper and tin). In the Iron Age, bronze was replaced by the use of iron. The dates of the Bronze Age and the Iron Age vary for different parts of the world.

10.3 The Half-Reaction Method for Balancing Equations

Section Preview/ Specific Expectations

In this section, you will

- **investigate** oxidation-reduction reactions by reacting metals with acids and by combusting hydrocarbons
- **write** balanced equations for redox reactions using the half-reaction method

Did you know that redox reactions are an important part of CD manufacturing? The CDs you buy at a music store are made of Lexan®, the same plastic used for riot shields and bulletproof windows. The CDs are coated with a thin aluminum film. They are copies of a single master disc, which is made of glass coated with silver, as seen in Figure 10.9. Silver is deposited on a glass disc by the reduction of silver ions with methanal, HCHO, also known as formaldehyde. In the same reaction, formaldehyde is oxidized to methanoic acid, HCOOH, also known as formic acid. The redox reaction occurs under acidic conditions.

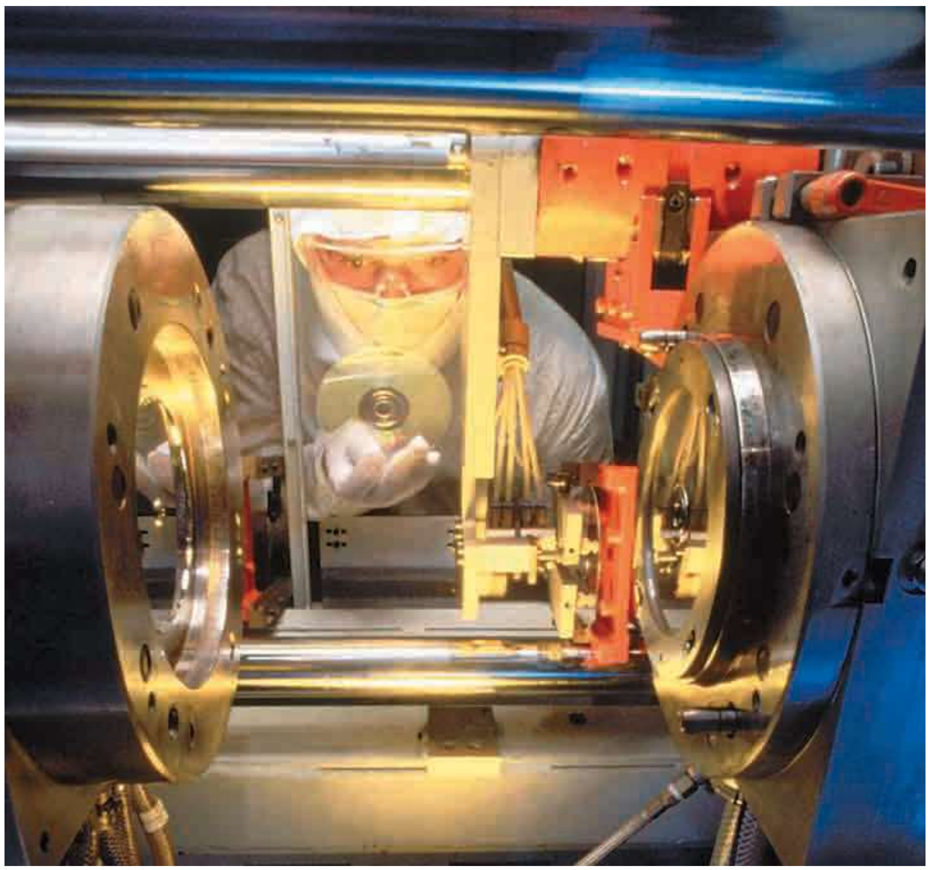


Figure 10.9 **The production of CDs depends on a redox reaction used to coat the master disc with silver.**

You have seen many balanced chemical equations and net ionic equations that represent redox reactions. There are specific techniques for balancing these equations. These techniques are especially useful for reactions that take place under acidic or basic conditions, such as the acidic conditions used in coating a master CD with silver.

In section 10.1, you learned to divide the balanced equations for some redox reactions into separate oxidation and reduction half-reactions. You will now use the reverse approach, and discover how to write a balanced equation by combining two half-reactions. To do this, you must first understand how to write a wide range of half-reactions.

Balancing Half-Reactions

In the synthesis of potassium chloride from its elements, metallic potassium is oxidized to form potassium ions, and gaseous chlorine is reduced to form chloride ions. This reaction is shown in Figure 10.10. Each half-reaction can be balanced by writing the correct formulas for the reactant and product, balancing the numbers of atoms, and then adding the correct number of electrons to balance the charges. For the oxidation half-reaction,

$$K \rightarrow K^+ + e^-$$

The atoms are balanced. The net charge on each side is 0. For the reduction half-reaction,

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

The atoms are balanced. The net charge on each side is −2.

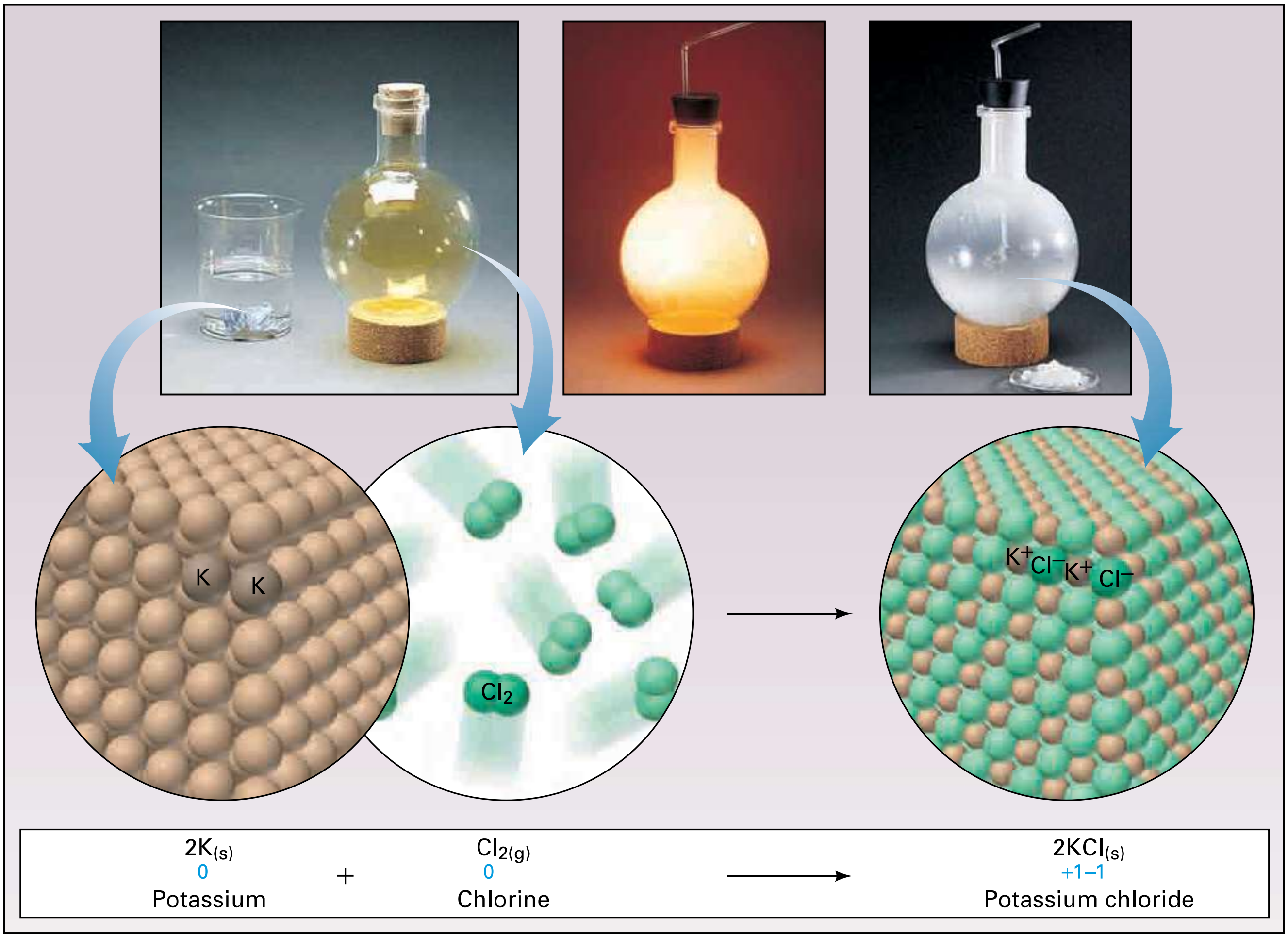


Figure 10.10 Grey potassium metal, which is stored under oil, reacts very vigorously with greenish-yellow chlorine gas to form white potassium chloride. The changes in oxidation numbers show that this synthesis reaction is also a redox reaction.

Redox reactions do not always take place under neutral conditions. Balancing half-reactions is more complicated for reactions that take place in acidic or basic solutions. When an acid or base is present, H^+ or OH^- ions must also be considered. However, the overall approach is similar. This approach involves writing the correct formulas for the reactants and products, balancing the atoms, and adding the appropriate number of electrons to one side of the half-reaction to balance the charges.

Balancing Half-Reactions for Acidic Solutions

The following steps are used to balance a half-reaction for an acidic solution. The Sample Problem that follows applies these steps.

Step 1 Write an unbalanced half-reaction that shows the formulas of the given reactant(s) and product(s).

Step 2 Balance any atoms other than oxygen and hydrogen first.

Step 3 Balance any oxygen atoms by adding water molecules.

Step 4 Balance any hydrogen atoms by adding hydrogen ions.

Step 5 Balance the charges by adding electrons.

Sample Problem

Balancing a Half-Reaction in Acidic Solution

Problem

Write a balanced half-reaction that shows the reduction of permanganate ions, MnO_4^-, to manganese(II) ions in an acidic solution.

Solution

Step 1 Represent the given reactant and product with correct formulas.
$MnO_4^- \rightarrow Mn^{2+}$

Step 2 Balance the atoms, starting with the manganese atoms. Here, the manganese atoms are already balanced.

Step 3 The reduction occurs in aqueous solution, so add water molecules to balance the oxygen atoms.
$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

Step 4 The reaction occurs in acidic solution, so add hydrogen ions to balance the hydrogen atoms.
$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$

Step 5 The atoms are now balanced, but the net charge on the left side is 7+, whereas the net charge on the right side is 2+. Add five electrons to the left side to balance the charges.
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Practice Problems

17. Write a balanced half-reaction for the reduction of cerium(IV) ions to cerium(III) ions.

18. Write a balanced half-reaction for the oxidation of bromide ions to bromine.

19. Balance each of the following half-reactions under acidic conditions.
(a) $O_2 \rightarrow H_2O_2$ **(b)** $H_2O \rightarrow O_2$ **(c)** $NO_3^- \rightarrow N_2$

20. Balance each of the following half-reactions under acidic conditions.
(a) $ClO_3^- \rightarrow Cl^-$ **(b)** $NO \rightarrow NO_3^-$ **(c)** $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

CONCEPT CHECK

The ability to balance a single half-reaction as a bookkeeping exercise does not mean that a single half-reaction can occur on its own. In a redox reaction, oxidation and reduction must both occur.

Balancing Half-Reactions for Basic Solutions

The following steps are used to balance a half-reaction for a basic solution. The Sample Problem that follows applies these steps.

Step 1 Write an unbalanced half-reaction that shows the formulas of the given reactant(s) and product(s).

Step 2 Balance any atoms other than oxygen and hydrogen first.

Step 3 Balance any oxygen and hydrogen atoms as if the conditions are acidic.

Step 4 Adjust for basic conditions by adding to both sides the same number of hydroxide ions as the number of hydrogen ions already present.

Step 5 Simplify the half-reaction by combining the hydrogen ions and hydroxide ions on the same side of the equation into water molecules.

Step 6 Remove any water molecules present on both sides of the half-reaction.

Step 7 Balance the charges by adding electrons.

Sample Problem

Balancing a Half-Reaction in Basic Solution

Problem

Write a balanced half-reaction that shows the oxidation of thiosulfate ions, $S_2O_3^{2-}$, to sulfite ions, SO_3^{2-}, in a basic solution.

Solution

Step 1 Represent the given reactant and product with correct formulas.

$S_2O_3^{2-} \rightarrow SO_3^{2-}$

Step 2 Balance the atoms, beginning with the sulfur atoms.

$S_2O_3^{2-} \rightarrow 2SO_3^{2-}$

Step 3 Balance the oxygen and hydrogen atoms as if the solution is acidic.

$S_2O_3^{2-} + 3H_2O \rightarrow 2SO_3^{2-}$

$S_2O_3^{2-} + 3H_2O \rightarrow 2SO_3^{2-} + 6H^+$

Step 4 There are six hydrogen ions present, so adjust for basic conditions by adding six hydroxide ions to each side.

$S_2O_3^{2-} + 3H_2O + 6OH^- \rightarrow 2SO_3^{2-} + 6H^+ + 6OH^-$

Step 5 Combine the hydrogen ions and hydroxide ions on the right side into water molecules.

$S_2O_3^{2-} + 3H_2O + 6OH^- \rightarrow 2SO_3^{2-} + 6H_2O$

Step 6 Remove three water molecules from each side.

$S_2O_3^{2-} + 6OH^- \rightarrow 2SO_3^{2-} + 3H_2O$

Step 7 The atoms are now balanced, but the net charge on the left side is 8–, whereas the net charge on the right side is 4–. Add four electrons to the right side to balance the charges.

$S_2O_3^{2-} + 6OH^- \rightarrow 2SO_3^{2-} + 3H_2O + 4e^-$

Practice Problems

21. Write a balanced half-reaction for the oxidation of chromium(II) ions to chromium(III) ions.

22. Write a balanced half-reaction for the reduction of oxygen to oxide ions.

23. Balance each of the following half-reactions under basic conditions.

(a) $Al \rightarrow Al(OH)_4^-$ **(b)** $CN^- \rightarrow CNO^-$ **(c)** $MnO_4^- \rightarrow MnO_2$

(d) $CrO_4^{2-} \rightarrow Cr(OH)_3$ **(e)** $CO_3^{2-} \rightarrow C_2O_4^{2-}$

24. Balance each of the following half-reactions.

(a) $FeO_4^{2-} \rightarrow Fe^{3+}$ (acidic conditions)

(b) $ClO_2^- \rightarrow Cl^-$ (basic conditions)

Half-Reaction Method for Balancing Redox Reactions

Recall that, if you consider a redox reaction as two half-reactions, electrons are lost in the oxidation half-reaction, and electrons are gained in the reduction half-reaction. For example, you know the reaction of zinc with aqueous copper(II) sulfate.

$$Zn_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + ZnSO_{4(aq)}$$

Removing the spectator ions leaves the following net ionic equation.

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$$

- You can break the net ionic equation into two half-reactions:
  Oxidation half-reaction: $Zn \rightarrow Zn^{2+} + 2e^-$
  Reduction half-reaction: $Cu^{2+} + 2e^- \rightarrow Cu$
- You can also start with the half-reactions and use them to produce a net ionic equation. If you add the two half-reactions, the result is as follows.

$$Zn + Cu^{2+} + 2e^- \rightarrow Cu + Zn^{2+} + 2e^-$$

Removing the two electrons from each side results in the original net ionic equation.

As shown above, you can use half-reactions to write balanced net ionic equations for redox reactions. In doing so, you use the fact that *no electrons are created or destroyed in a redox reaction*. Electrons are transferred from one reactant (the reducing agent) to another (the oxidizing agent).

Balancing a Net Ionic Equation

You know from Investigation 10-A that magnesium metal, $Mg_{(s)}$, displaces aluminum from an aqueous solution of one of its compounds, such as aluminum nitrate, $Al(NO_3)_{3(aq)}$. To obtain a balanced net ionic equation for this reaction, you can start by looking at the half-reactions. Magnesium atoms undergo oxidation to form magnesium ions, which have a 2+ charge. The oxidation half-reaction is as follows.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

Aluminum ions, which have a 3+ charge, undergo reduction to form aluminum atoms. The reduction half-reaction is as follows.

$$Al^{3+} + 3e^- \rightarrow Al$$

To balance the net ionic equation for this redox reaction, you can combine the two half-reactions in such a way that the number of electrons lost through oxidation equals the number of electrons gained through reduction. In other words, you can model the transfer of a certain number of electrons from the reducing agent to the oxidizing agent.

For the reaction of magnesium metal with aluminum ions, the two balanced half-reactions include different numbers of electrons, 2 and 3. The least common multiple of 2 and 3 is 6. To combine the half-reactions and give a balanced net ionic equation, multiply the balanced half-reactions by different numbers so that the results both include six electrons, as shown below.

- Multiply the oxidation half-reaction by 3. Multiply the reduction half-reaction by 2.

$$3Mg \rightarrow 3Mg^{2+} + 6e^-$$
$$2Al^{3+} + 6e^- \rightarrow 2Al$$

- Add the results.

$$3Mg + 2Al^{3+} + 6e^- \rightarrow 3Mg^{2+} + 2Al + 6e^-$$

- Remove $6e^-$ from each side to obtain the balanced net ionic equation.

$$3Mg + 2Al^{3+} \rightarrow 3Mg^{2+} + 2Al$$

To produce the balanced chemical equation, you can include the spectator ions, which are nitrate ions in this example. Include the states, if necessary. The balanced chemical equation is:

$$3Mg_{(s)} + 2Al(NO_3)_{3(aq)} \rightarrow 3Mg(NO_3)_{2(aq)} + 2Al_{(s)}$$

Steps for Balancing by the Half-Reaction Method

You could balance the chemical equation for the reaction of magnesium with aluminum nitrate by inspection, instead of writing half-reactions. However, many redox equations are difficult to balance by the inspection method. In general, you can balance the net ionic equation for a redox reaction by a process known as the half-reaction method. The preceding example of the reaction of magnesium with aluminum nitrate illustrates this method. Specific steps for following the half-reaction method are given below.

Step 1 Write an unbalanced net ionic equation, if it is not already given.

Step 2 Divide the unbalanced net ionic equation into an oxidation half-reaction and a reduction half-reaction. To do this, you may need to assign oxidation numbers to all the elements in the net ionic equation to determine what is oxidized and what is reduced.

Step 3 Balance the oxidation half-reaction and the reduction half-reaction independently.

Step 4 Determine the least common multiple (LCM) of the numbers of electrons in the oxidation half-reaction and the reduction half-reaction.

Continued on the next page

Math LINK

The lowest or least common multiple (LCM) of two numbers is the smallest multiple of each number. For example, the LCM of 2 and 1 is 2; the LCM of 3 and 6 is 6; and the LCM of 2 and 5 is 10. One way to find the LCM of two numbers is to list the multiples of each number and to find the smallest number that appears in both lists. For the numbers 6 and 8,

- the multiples of 6 are: 6, 12, 18, **24**, 30,…
- the multiples of 8 are: 8, 16, **24**, 32, 40,…

Thus, the LCM of 6 and 8 is 24.

What is the LCM of the numbers 4 and 12?

What is the LCM of the numbers 7 and 3?

Step 5 Use coefficients to write each half-reaction so that it includes the LCM of the numbers of electrons.

Step 6 Add the balanced half-reactions that include the equal numbers of electrons.

Step 7 Remove the electrons from both sides of the equation.

Step 8 Remove any identical molecules or ions that are present on both sides of the equation.

Step 9 If you require a balanced chemical equation, include any spectator ions in the chemical formulas.

Step 10 If necessary, include the states.

When using the half-reaction method, keep in mind that, in a redox reaction, *the number of electrons lost through oxidation must equal the number of electrons gained through reduction*. Figure 10.11 provides another example.

Figure 10.11 **Lithium displaces hydrogen from water to form lithium hydroxide.**

Oxidation half-reaction:

$Li \rightarrow Li^+ + e^-$

Reduction half-reaction:

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

Multiply the oxidation half-reaction by 2, add the half-reactions, and simplify the result to obtain the balanced net ionic equation.

$2Li + 2H_2O \rightarrow 2Li^+ + 2OH^- + H_2$

Balancing Redox Reactions in Acidic and Basic Solutions

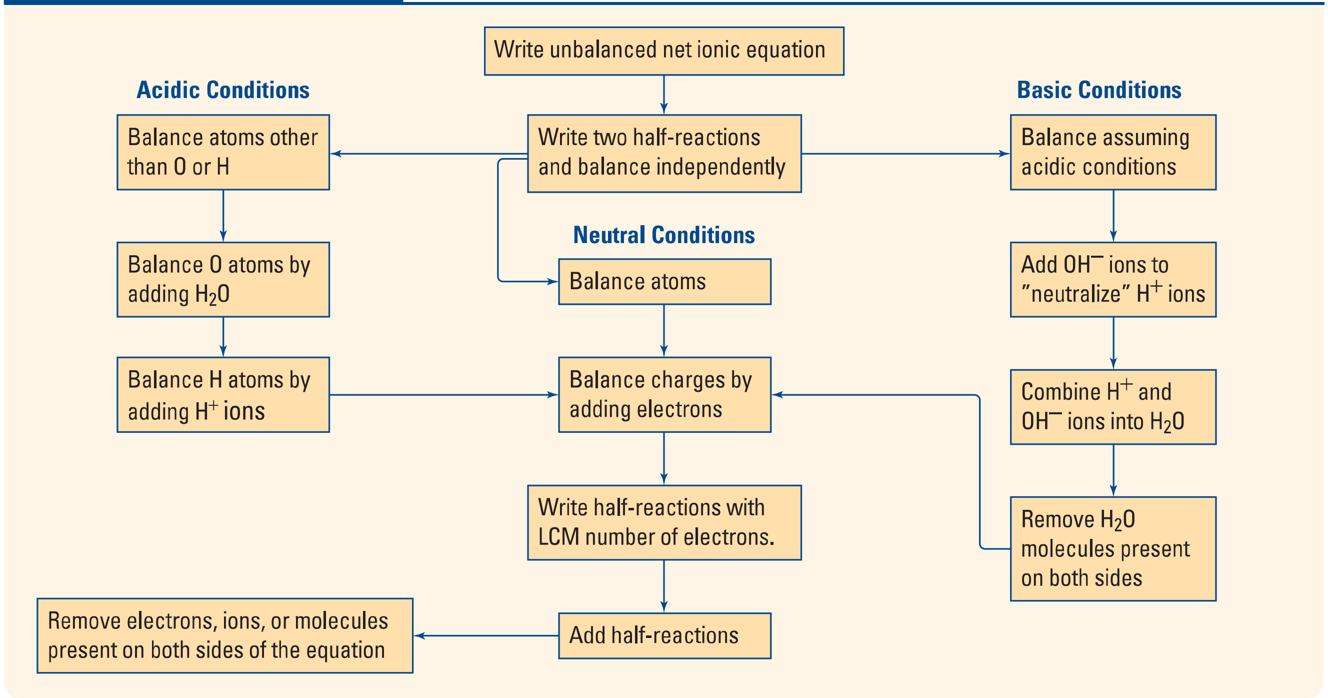
The half-reaction method of balancing equations can be more complicated for reactions that take place under acidic or basic conditions. The overall approach, however, is the same. You need to balance the two half-reactions, find the LCM of the numbers of electrons, and then multiply by coefficients to equate the number of electrons lost and gained. Finally, add the half-reactions and simplify to give a balanced net ionic equation for the reaction. The ten steps listed above show this process in more detail.

The Sample Problem on the next page illustrates the use of these steps for an acidic solution. To balance a net ionic equation for basic conditions by the half-reaction method, balance each half-reaction for acidic conditions, adjust for basic conditions, and then combine the half-reactions to obtain the balanced net ionic equation. The following Concept Organizer summarizes how to use the half-reaction method in both acidic and basic conditions.

CONCEPT CHECK

Explain why a balanced chemical equation or net ionic equation for a redox reaction does not include any electrons.

Concept Organizer — The Half-Reaction Method for Balancing Redox Equations



Sample Problem

Balancing a Redox Equation in Acidic Solution

Problem

Write a balanced net ionic equation to show the reaction of perchlorate ions, ClO_4^-, and nitrogen dioxide in acidic solution to produce chloride ions and nitrate ions.

What Is Required?

You need to write a balanced net ionic equation for the given reaction.

What Is Given?

You know the identities of two reactants and two products, and that the reaction takes place in acidic solution.

Plan Your Strategy

- Write an unbalanced ionic equation.
- Determine whether the reaction is a redox reaction.
- If it is not a redox reaction, balance by inspection.
- If it is a redox reaction, follow the steps for balancing by the half-reaction method.

Act on Your Strategy

- The unbalanced ionic equation is: $ClO_4^- + NO_2 \rightarrow Cl^- + NO_3^-$
- Assign oxidation numbers to all the elements to determine which reactant, if any, is oxidized or reduced.

$$\underset{+7\ -2}{ClO_4^-} + \underset{+4\ -2}{NO_2} \rightarrow \underset{-1}{Cl^-} + \underset{+5\ -2}{NO_3^-}$$

Continued ...

Continued ...

The oxidation number of chlorine decreases, so perchlorate ions are reduced to chloride ions.

The oxidation number of nitrogen increases, so nitrogen dioxide is oxidized to nitrate ions.

- This is a redox reaction. Use the half-reaction method to balance the equation.

Step 1 The unbalanced net ionic equation is already written.
$ClO_4^- + NO_2 \rightarrow Cl^- + NO_3^-$

Step 2 Write two unbalanced half-reactions.
Oxidation: $NO_2 \rightarrow NO_3^-$
Reduction: $ClO_4^- \rightarrow Cl^-$

Step 3 Balance the two half-reactions for acidic conditions.

Oxidation	Reduction
$NO_2 \rightarrow NO_3^-$	$ClO_4^- \rightarrow Cl^-$
$NO_2 + H_2O \rightarrow NO_3^-$	$ClO_4^- \rightarrow Cl^- + 4H_2O$
$NO_2 + H_2O \rightarrow NO_3^- + 2H^+$	$ClO_4^- + 8H^+ \rightarrow Cl^- + 4H_2O$
$NO_2 + H_2O \rightarrow NO_3^- + 2H^+ + e^-$	$ClO_4^- + 8H^+ + 8e^- \rightarrow Cl^- + 4H_2O$

Step 4 The LCM of 1 and 8 is 8.

Step 5 Multiply the oxidation half-reaction by 8, so that equal numbers of electrons are lost and gained.
$8NO_2 + 8H_2O \rightarrow 8NO_3^- + 16H^+ + 8e^-$

Step 6 Add the half reactions.

$$\begin{array}{rl} 8NO_2 + 8H_2O & \rightarrow 8NO_3^- + 16H^+ + 8e^- \\ ClO_4^- + 8H^+ + 8e^- & \rightarrow Cl^- + 4H_2O \\ \hline \end{array}$$

$$8NO_2 + 8H_2O + ClO_4^- + 8H^+ + 8e^- \rightarrow 8NO_3^- + 16H^+ + 8e^- + Cl^- + 4H_2O$$

Step 7 Simplify by removing 8 electrons from both sides.
$8NO_2 + 8H_2O + ClO_4^- + 8H^+ \rightarrow 8NO_3^- + 16H^+ + Cl^- + 4H_2O$

Step 8 Simplify by removing 4 water molecules, and 8 hydrogen ions from each side.
$8NO_2 + ClO_4^- + 4H_2O \rightarrow 8NO_3^- + 8H^+ + Cl^-$

(Steps 9 and 10 are not required for this problem.)

Check Your Solution

- The atoms are balanced.
- The charges are balanced.

Practice Problems

25. Balance each of the following redox equations by inspection. Write the balanced half-reactions in each case.

(a) $Na + F_2 \rightarrow NaF$

(b) $Mg + N_2 \rightarrow Mg_3N_2$

(c) $HgO \rightarrow Hg + O_2$

26. Balance the following equation by the half-reaction method.
$Cu^{2+} + I^- \rightarrow CuI + I_3^-$

27. Balance each of the following ionic equations for acidic conditions. Identify the oxidizing agent and the reducing agent in each case.
   (a) $MnO_4^- + Ag \rightarrow Mn^{2+} + Ag^+$
   (b) $Hg + NO_3^- + Cl^- \rightarrow HgCl_4^{2-} + NO_2$
   (c) $AsH_3 + Zn^{2+} \rightarrow H_3AsO_4 + Zn$
   (d) $I_3^- \rightarrow I^- + IO_3^-$
28. Balance each of the following ionic equations for basic conditions. Identify the oxidizing agent and the reducing agent in each case.
   (a) $CN^- + MnO_4^- \rightarrow CNO^- + MnO_2$
   (b) $H_2O_2 + ClO_2 \rightarrow ClO_2^- + O_2$
   (c) $ClO^- + CrO_2^- \rightarrow CrO_4^{2-} + Cl_2$
   (d) $Al + NO_2^- \rightarrow NH_3 + AlO_2^-$

In the next investigation, you will carry out several redox reactions, including reactions of acids with metals, and the combustion of hydrocarbons.

Tools & Techniques

The Breathalyzer Test: A Redox Reaction

The police may pull over a driver weaving erratically on the highway on suspicion of drunk driving. A police officer must confirm this suspicion by assessing whether the driver has a blood alcohol concentration over the "legal limit." The "Breathalyzer" test checks a person's breath using a redox reaction to determine blood alcohol concentration. This test was invented in 1953 by Robert Borkenstein, a former member of the Indiana State Police, and a professor of forensic studies.

What does a person's breath have to do with the alcohol in his or her blood? In fact, there is a direct correlation between the concentration of alcohol in an exhaled breath and the concentration of alcohol in the blood.

As blood moves through the lungs, it comes in close contact with inhaled gases. If the blood contains alcohol, the concentration of alcohol in the blood quickly reaches equilibrium with the concentration of alcohol in each inhaled breath. Thus, the alcohol content in an exhaled breath is a measure of the alcohol concentration in the blood itself. For example, if a person has been drinking alcohol, every 2100 mL of air exhaled contains about the same amount of alcohol as 1 mL of blood.

In the Breathalyzer test, the subject blows into a tube connected to a vial. The exhaled air collects in the vial, which already contains a mixture of sulfuric acid, potassium dichromate, water, and the catalyst silver nitrate. The alcohol reacts with the dichromate ion in the following redox reaction.

$$16H^+_{(aq)} + \underset{\text{(orange)}}{2Cr_2O_7^{2-}{}_{(aq)}} + 3C_2H_5OH_{(\ell)} \rightarrow \underset{\text{(green)}}{4Cr^{3+}{}_{(aq)}} + 3C_2H_4O_{2(aq)} + 11H_2O_{(\ell)}$$

This reaction is accompanied by a visible colour change, as orange dichromate ions become green chromium(III) ions. The concentration of alcohol in the blood is determined by measuring the intensity of the final colour.

A recent modification of the Breathalyzer test prevents drivers from starting their cars if they have been drinking. Alcohol ignition locks involve a type of Breathalyzer test that is linked to the car's ignition system. Until the driver passes the test, the car will not start. This test is useful in regulating the driving habits of people who have been previously convicted of drinking and driving.

Investigation 10-B

SKILL FOCUS

- Predicting
- Performing and recording
- Analyzing and interpreting

Redox Reactions and Balanced Equations

A redox reaction involves the transfer of electrons between reactants. A reactant that loses electrons is oxidized and acts as a reducing agent. A reactant that gains electrons is reduced and acts as an oxidizing agent. Redox reactions can be represented by balanced equations.

Questions

How can you tell if a redox reaction occurs when reactants are mixed? Can you observe the transfer of electrons in the mixture?

Predictions

- Predict which of the metals magnesium, zinc, copper, and aluminum can be oxidized by aqueous hydrogen ions. Explain your reasoning.
- Predict whether metals that cannot be oxidized by hydrogen ions can dissolve in acids. Explain your reasoning.
- Predict whether the combustion of a hydrocarbon is a redox reaction. What assumptions have you made about the products?

Materials

well plate
4 small test tubes
test tube rack
small pieces of each of the metals magnesium, zinc, copper, and aluminum
dilute hydrochloric acid (1 mol/L)
dilute sulfuric acid (1 mol/L)
Bunsen burner
candle

Safety Precautions

- The acid solutions are corrosive. Handle them with care.
- If you accidentally spill a solution on your skin, wash the area immediately with copious amounts of cool water. If you get any acid in your eyes, wash at the eye wash station. Inform your teacher.
- Before lighting a Bunsen burner or candle, make sure that there are no flammable liquids nearby. Also, tie back long hair, and confine any loose clothing.

Procedure

Part 1 Reactions of Acids

1. Place a small piece of each metal on the well plate. Add a few drops of hydrochloric acid to each metal. Record your observations. If you are unsure of your observations, repeat the procedure on a larger scale in a small test tube.
2. Place another small piece of each metal on clean sections of the well plate. Add a few drops of sulfuric acid to each metal. Record your observations. If you are unsure of your observations, repeat the procedure on a larger scale in a small test tube.
3. Dispose of the mixtures in the beaker supplied by your teacher.

Part 2 Combustion of Hydrocarbons

4. Observe the combustion of natural gas in a Bunsen burner. Adjust the colour of the flame by varying the quantity of oxygen admitted to the burner. How does the colour depend on the quantity of oxygen?
5. Observe the combustion of a candle. Compare the colour of the flame with the colour of the Bunsen burner flame. Which adjustment of the burner makes the colours of the two flames most similar?

Analysis

Part 1 Reactions of Acids

1. Write a balanced chemical equation for each of the reactions of an acid with a metal.
2. Write each equation from question 1 in net ionic form.
3. Determine which of the reactions from question 1 are redox reactions.
4. Write each redox reaction from question 3 as two half-reactions.
5. Explain any similarities in your answers to question 4.
6. In the reactions you observed, are the hydrogen ions acting as an oxidizing agent, a reducing agent, or neither?
7. In the neutralization reaction of hydrochloric acid and sodium hydroxide, do the hydrogen ions behave in the same way as you found in question 6? Explain.
8. Your teacher may demonstrate the reaction of copper with concentrated nitric acid to produce copper(II) ions and brown, toxic nitrogen dioxide gas. Write a balanced net ionic equation for this reaction. Do the hydrogen ions behave in the same way as you found in question 6? Identify the oxidizing agent and the reducing agent in this reaction.
9. From your observations of copper with hydrochloric acid and nitric acid, can you tell whether hydrogen ions or nitrate ions are the better oxidizing agent? Explain.

Part 2 Combustion of Hydrocarbons

10. The main component of natural gas is methane, CH_4. The products of the combustion of this gas in a Bunsen burner depend on how the burner is adjusted. A blue flame indicates complete combustion. What are the products in this case? Write a balanced chemical equation for this reaction.
11. A yellow or orange flame from a Bunsen burner indicates incomplete combustion and the presence of carbon in the flame. Write a balanced chemical equation for this reaction.
12. Name another possible carbon-containing product from the incomplete combustion of methane. Write a balanced chemical equation for this reaction.
13. The fuel in a burning candle is paraffin wax, $C_{25}H_{52}$. Write a balanced chemical equation for the complete combustion of paraffin wax.
14. Write two balanced equations that represent the incomplete combustion of paraffin wax.
15. How do you know that at least one of the incomplete combustion reactions is taking place when a candle burns?
16. Are combustion reactions also redox reactions? Does your answer depend on whether the combustion is complete or incomplete? Explain.

Conclusion

17. How could you tell if a redox reaction occurred when reactants were mixed? Could you observe the transfer of electrons in the mixture?

Applications

18. Gold is very unreactive and does not dissolve in most acids. However, it does dissolve in *aqua regia* (Latin for "royal water"), which is a mixture of concentrated hydrochloric and nitric acids. The unbalanced ionic equation for the reaction is as follows.

$$Au + NO_3^- + Cl^- \rightarrow AuCl_4^- + NO_2$$

Balance the equation, and identify the oxidizing agent and reducing agent.

19. Natural gas is burned in gas furnaces. Give at least three reasons why this combustion reaction should be as complete as possible. How would you try to ensure complete combustion?

Section Summary

In this section, you learned the half-reaction method for balancing equations for redox reactions. You investigated the redox reactions of metals with acids, and the combustion of two hydrocarbons. After applying the half-reaction method in the following review problems, you will learn a different method in section 10.4. This method will make greater use of oxidation numbers.

Section Review

1 K/U Balance each half-reaction. Identify it as an oxidation or reduction half-reaction.

(a) $C \rightarrow C_2^{2-}$

(b) $S_2O_3^{2-} \rightarrow S_4O_6^{2-}$

(c) $AsO_4^{3-} \rightarrow As_4O_6$ (acidic conditions)

(d) $Br_2 \rightarrow BrO_3^-$ (basic conditions)

2 K/U Balance each equation.

(a) $Co^{3+} + Au \rightarrow Co^{2+} + Au^{3+}$

(b) $Cu + NO_3^- \rightarrow Cu^{2+} + NO$ (acidic conditions)

(c) $NO_3^- + Al \rightarrow NH_3 + AlO_2^-$ (basic conditions)

3 MC This section began with a description of the use of a redox reaction to make the master disc in the production of CDs. Write a balanced net ionic equation for the reaction of silver ions with methanal under acidic conditions to form metallic silver and methanoic acid.

4 I In basic solution, ammonia, NH_3, can be oxidized to dinitrogen monoxide, N_2O.

(a) Try to balance the half-reaction by adding water molecules, hydroxide ions, and electrons, without first assuming acidic conditions. Describe any difficulties you encounter.

(b) Balance the half-reaction by first assuming acidic conditions and then adjusting to introduce the hydroxide ions. Compare your findings with those from part (a).

5 MC A mixture of liquid hydrazine, N_2H_4, and liquid dinitrogen tetroxide can be used as a rocket fuel. The products of the reaction are nitrogen gas and water vapour.

(a) Write a balanced chemical equation for the reaction by inspection.

(b) Identify the oxidizing agent and the reducing agent.

(c) Hydrazine is made in the Raschig Process. In this process, ammonia reacts with hypochlorite ions in a basic solution to form hydrazine and chloride ions. Write the balanced net ionic equation.

6 C Ben and Larissa were working together to balance the following equation for a redox reaction.

$Zn + SO_4^{2-} + H^+ \rightarrow Zn^{2+} + S + H_2O$

Ben suggested balancing by inspection, with the following result.

$Zn + SO_4^{2-} + 8H^+ \rightarrow Zn^{2+} + S + 4H_2O$

Larissa said: "That's not balanced."

(a) Was Larissa right? Explain.

(b) How would you balance the equation, and what would the result be?

The Oxidation Number Method for Balancing Equations

10.4

Section Preview/ Specific Expectations

In this section, you will

- **write** balanced equations for redox reactions using the oxidation number method

In section 10.2, you learned that a redox reaction involves changes in oxidation numbers. If an element undergoes oxidation, its oxidation number increases. If an element undergoes reduction, its oxidation number decreases. When balancing equations by the half-reaction method in section 10.3, you sometimes used oxidation numbers to determine the reactant(s) and product(s) in each half-reaction.

In fact, you can use oxidation numbers to balance a chemical equation by a new method. The oxidation number method is a method of balancing redox equations by ensuring that *the total increase in the oxidation numbers of the oxidized element(s) equals the total decrease in the oxidation numbers of the reduced element(s)*.

For example, the combustion of ammonia in oxygen produces nitrogen dioxide and water.

$$\underset{-3\ +1}{NH_3} + \underset{0}{O_2} \rightarrow \underset{+4\ -2}{NO_2} + \underset{+1\ -2}{H_2O}$$

The oxidation number of nitrogen increases from −3 to +4, an increase of 7. The oxidation number of oxygen decreases from 0 to −2, a decrease of 2. The least common multiple of 7 and 2 is 14. In this case, two nitrogen atoms must react for every seven oxygen atoms so that the total increase and decrease in oxidation numbers both equal 14.

$$2NH_3 + \frac{7}{2}O_2 \rightarrow NO_2 + H_2O$$

Complete the equation by inspection. If necessary, eliminate the fraction.

$$2NH_3 + \frac{7}{2}O_2 \rightarrow 2NO_2 + 3H_2O$$

$$4NH_3 + 7O_2 \rightarrow 4NO_2 + 6H_2O$$

A summary of the steps of the oxidation number method is given below. The following Sample Problem shows how these steps are applied.

Step 1 Write an unbalanced equation, if it is not given.

Step 2 Determine whether the reaction is a redox reaction by assigning an oxidation number to each element wherever it appears in the equation.

Step 3 If the reaction is a redox reaction, identify the element(s) that undergo an increase in oxidation number and the element(s) that undergo a decrease in oxidation number.

Step 4 Find the numerical values of the increase and the decrease in oxidation numbers.

Step 5 Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.

Step 6 Use the smallest whole-number ratio to balance the numbers of atoms of the element(s) oxidized and the element(s) reduced.

continued on the next page

Step 7 Balance the other elements by inspection, if possible.

Step 8 For reactions that occur in acidic or basic solutions, include water molecules, hydrogen ions, or hydroxide ions as needed to balance the equation.

Sample Problem

Balancing a Redox Equation in Basic Solution

Problem

Write a balanced net ionic equation to show the formation of iodine by bubbling oxygen gas through a basic solution that contains iodide ions.

Solution

Step 1 Write an unbalanced equation from the given information.

$O_2 + I^- \rightarrow I_2$

Step 2 Assign oxidation numbers to see if it is a redox reaction.

$\underset{0}{O_2} + \underset{-1}{I^-} \rightarrow \underset{0}{I_2}$

Because iodide is oxidized to iodine, the reaction is a redox reaction. Though the product that contains oxygen is unknown at this stage, oxygen must be reduced.

Step 3 Iodine is the element that undergoes an increase in oxidation number. Oxygen is the element that undergoes a decrease in oxidation number.

Step 4 Iodine undergoes an increase in its oxidation number from −1 to 0, an increase of 1. Assume that the oxidation number of oxygen after reduction is its normal value, that is, −2. Thus, oxygen undergoes a decrease in its oxidation number from 0 to −2, a decrease of 2.

Step 5 A 2:1 ratio of iodine atoms to oxygen atoms ensures that the total increase in oxidation numbers and the total decrease in oxidation numbers are both equal to 2. This is the smallest whole-number ratio.

Step 6 Use the ratio to balance the numbers of atoms of iodine and oxygen. Make sure there are two iodine atoms for every oxygen atom.

$O_2 + 4I^- \rightarrow 2I_2$

Step 7 No other reactants or products can be balanced by inspection.

Step 8 The reaction occurs in basic solution. As you learned in section 10.3, for basic conditions, start by assuming that the conditions are acidic. Add water molecules and hydrogen ions as necessary to balance the atoms.

$O_2 + 4I^- \rightarrow 2I_2 + 2H_2O$

$O_2 + 4I^- + 4H^+ \rightarrow 2I_2 + 2H_2O$

Add hydroxide ions to adjust for basic conditions. Simplify the resulting equation.

$O_2 + 4I^- + 4H^+ + 4OH^- \rightarrow 2I_2 + 2H_2O + 4OH^-$

$O_2 + 4I^- + 4H_2O \rightarrow 2I_2 + 2H_2O + 4OH^-$

$O_2 + 4I^- + 2H_2O \rightarrow 2I_2 + 4OH^-$

CONCEPT CHECK

Explain why none of the steps in the oxidation number method result in equations that include electrons, e^-.

Practice Problems

29. Use the oxidation number method to balance the following equation for the combustion of carbon disulfide.

$CS_2 + O_2 \rightarrow CO_2 + SO_2$

30. Use the oxidation number method to balance the following equations.

(a) $B_2O_3 + Mg \rightarrow MgO + Mg_3B_2$

(b) $H_2S + H_2O_2 \rightarrow S_8 + H_2O$

31. Use the oxidation number method to balance each ionic equation in acidic solution.

(a) $Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$

(b) $I_2 + NO_3^{-} \rightarrow IO_3^{-} + NO_2$

(c) $PbSO_4 \rightarrow Pb + PbO_2 + SO_4^{2-}$

32. Use the oxidation number method to balance each ionic equation in basic solution.

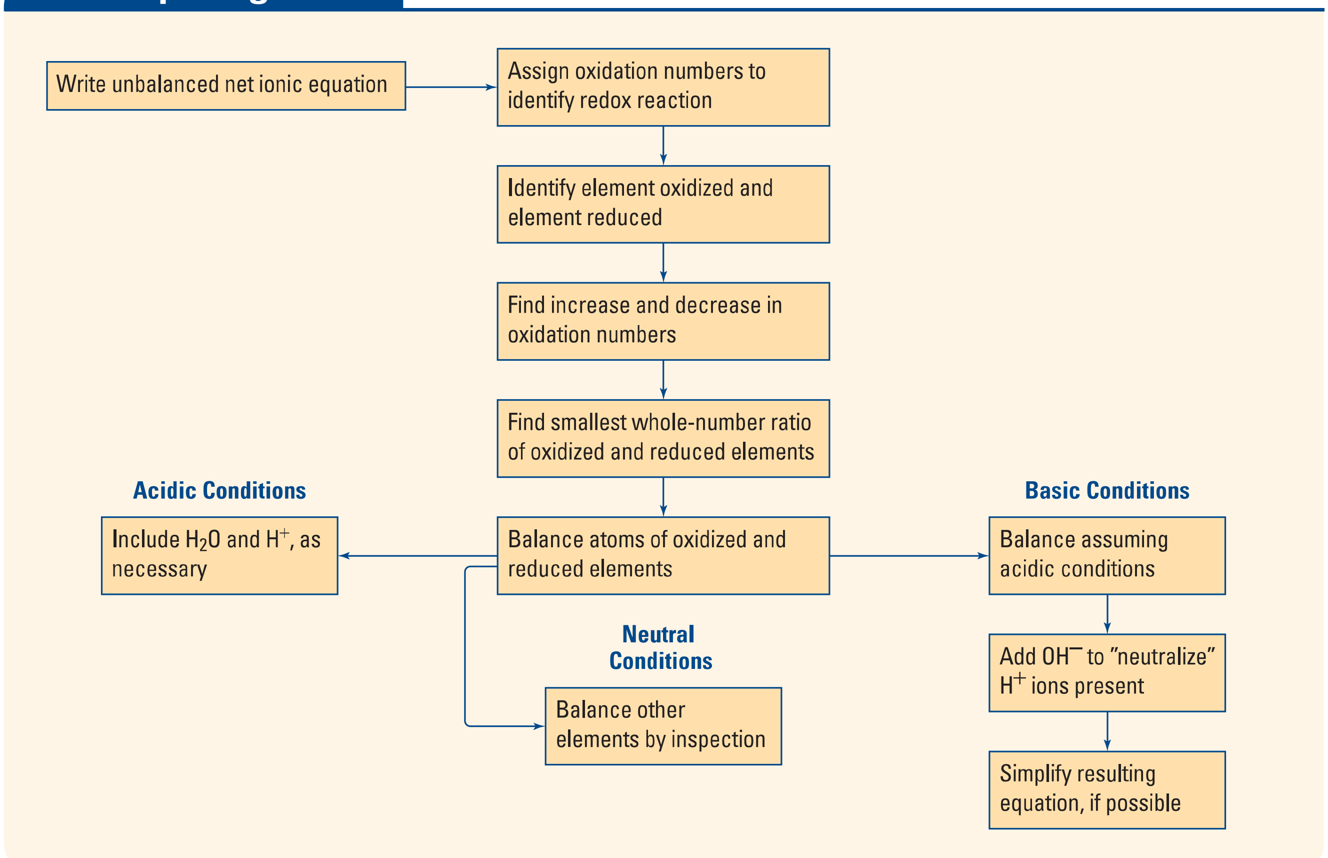
(a) $Cl^{-} + CrO_4^{2-} \rightarrow ClO^{-} + CrO_2^{-}$

(b) $Ni + MnO_4^{-} \rightarrow NiO + MnO_2$

(c) $I^{-} + Ce^{4+} \rightarrow IO_3^{-} + Ce^{3+}$

Concept Organizer

The Oxidation Number Method for Balancing Redox Equations



Section Summary

In this section, you learned how to use the oxidation number method to balance redox equations. You now know various techniques for recognizing and representing redox reactions. In Chapter 11, you will use these techniques to examine specific applications of redox reactions in the business world and in your daily life.

Section Review

1 K/U Is it possible to use the half-reaction method or the oxidation number method to balance the following equation? Explain your answer.

$Al_2S_3 + H_2O \rightarrow Al(OH)_3 + H_2S$

2 I Balance each equation by the method of your choice. Explain your choice of method in each case.

(a) $CH_3COOH + O_2 \rightarrow CO_2 + H_2O$

(b) $O_2 + H_2SO_3 \rightarrow HSO_4^-$ (acidic conditions)

3 K/U Use the oxidation number method to balance the following equations.

(a) $NH_3 + Cl_2 \rightarrow NH_4Cl + N_2$

(b) $Mn_3O_4 + Al \rightarrow Al_2O_3 + Mn$

4 C Explain why, in redox reactions, the total increase in the oxidation numbers of the oxidized elements must equal the total decrease in the oxidation numbers of the reduced elements.

5 I The combustion of ammonia in oxygen to form nitrogen dioxide and water vapour involves covalent molecules in the gas phase. The oxidation number method for balancing the equation was shown in an example in this section. Devise a half-reaction method for balancing the equation. Describe the assumptions you made in order to balance the equation. Also, describe why these assumptions did not affect the final result.

CHAPTER 10 Review

Reflecting on Chapter 10

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Give two different definitions for the term oxidation.
- Give two different definitions for the term reduction.
- Define a half-reaction. Give an example of an oxidation half-reaction and a reduction half-reaction.
- Compare the half-reaction and oxidation number methods of balancing equations.
- Practise balancing equations using both methods.
- Write an example of a balanced chemical equation for a redox reaction. Assign oxidation numbers to each element in the equation, then explain how you know it is a redox reaction.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

ore	oxidation
reduction	oxidation-reduction reaction
redox reaction	reducing agent
oxidizing agent	disproportionation
half-reaction	
oxidation numbers	

Knowledge/Understanding

1. For each reaction below, write a balanced chemical equation by inspection.
 (a) zinc metal with aqueous silver nitrate
 (b) aqueous cobalt(II) bromide with aluminum metal
 (c) metallic cadmium with aqueous tin(II) chloride

2. For each reaction in question 1, write the total ionic and net ionic equations.

3. For each reaction in question 1, identify the oxidizing agent and reducing agent.

4. For each reaction in question 1, write the two half-reactions.

5. When a metallic element reacts with a non-metallic element, which reactant is
 (a) oxidized?
 (b) reduced?
 (c) the oxidizing agent?
 (d) the reducing agent?

6. Use a Lewis structure to assign an oxidation number to each element in the following compounds.
 (a) $BaCl_2$
 (b) CS_2
 (c) XeF_4

7. Determine the oxidation number of each element present in the following substances.
 (a) BaH_2
 (b) Al_4C_3
 (c) KCN
 (d) $LiNO_2$
 (e) $(NH_4)_2C_2O_4$
 (f) S_8
 (g) AsO_3^{3-}
 (h) VO_2^+
 (i) XeO_3F^-
 (j) $S_4O_6^{2-}$

8. Identify a polyatomic ion in which chlorine has an oxidation number of +3.

9. Determine which of the following balanced chemical equations represent redox reactions. For each redox reaction, identify the oxidizing agent and the reducing agent.
 (a) $2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$
 (b) $CaO + SO_2 \rightarrow CaSO_3$
 (c) $H_2 + I_2 \rightarrow 2HI$
 (d) $KMnO_4 + 5CuCl + 8HCl \rightarrow KCl + MnCl_2 + 5CuCl_2 + 4H_2O$

10. Determine which of the following balanced net ionic equations represent redox reactions. For each redox reaction, identify the reactant that undergoes oxidation and the reactant that undergoes reduction.
 (a) $2Ag^+_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)}$
 (b) $Pb^{2+}_{(aq)} + S^{2-}_{(aq)} \rightarrow PbS_{(s)}$
 (c) $2Mn^{2+} + 5BiO_3^- + 14H^+ \rightarrow 2MnO_4^- + 5Bi^{3+} + 7H_2O$

11. (a) Examples of molecules and ions composed only of vanadium and oxygen are listed below. In this list, identify molecules and ions in which the oxidation number of vanadium is the same.
V_2O_5
V_2O_3
VO_2
VO
VO_2^{+}
VO^{2+}
VO_3^{-}
VO_4^{3-}
$V_3O_9^{3-}$
(b) Is the following reaction a redox reaction?
$2NH_4VO_3 \rightarrow V_2O_5 + 2NH_3 + H_2O$

12. The method used to manufacture nitric acid involves the following three steps.
Step 1 $4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$
Step 2 $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$
Step 3 $3NO_{2(g)} + H_2O_{(\ell)} \rightarrow 2HNO_{3(aq)} + NO_{(g)}$
(a) Which of these steps are redox reactions?
(b) Identify the oxidizing agent and the reducing agent in each redox reaction.

13. In a synthesis reaction involving elements A and B, the oxidation number of element A increases. What happens to the oxidation number of element B? How do you know?

14. Balance each of the following half-reactions.
(a) The reduction of iodine to iodide ions
(b) The oxidation of lead to lead(IV) ions
(c) The reduction of tetrachlorogold(III) ions, $AuCl_4^{-}$, to chloride ions and metallic gold
(d) $C_2H_5OH \rightarrow CH_3COOH$ (acidic conditions)
(e) $S_8 \rightarrow H_2S$ (acidic conditions)
(f) $AsO_2^{-} \rightarrow AsO_4^{3-}$ (basic conditions)

15. Use the half-reaction method to balance each of the following equations.
(a) $MnO_2 + Cl^{-} \rightarrow Mn^{2+} + Cl_2$ (acidic conditions)
(b) $NO + Sn \rightarrow NH_2OH + Sn^{2+}$ (acidic conditions)
(c) $Cd^{2+} + V^{2+} \rightarrow Cd + VO_3^{-}$ (acidic conditions)
(d) $Cr \rightarrow Cr(OH)_4^{-} + H_2$ (basic conditions)
(e) $S_2O_3^{2-} + NiO_2 \rightarrow Ni(OH)_2 + SO_3^{2-}$ (basic conditions)
(f) $Sn^{2+} + O_2 \rightarrow Sn^{4+}$ (basic conditions)

16. Use the oxidation number method to balance each of the following equations.
(a) $SiCl_4 + Al \rightarrow Si + AlCl_3$
(b) $PH_3 + O_2 \rightarrow P_4O_{10} + H_2O$
(c) $I_2O_5 + CO \rightarrow I_2 + CO_2$
(d) $SO_3^{2-} + O_2 \rightarrow SO_4^{2-}$

17. Complete and balance a net ionic equation for each of the following disproportionation reactions.
(a) $NO_2 \rightarrow NO_2^{-} + NO_3^{-}$ (acidic conditions), which is one of the reactions involved in acid rain formation
(b) $Cl_2 \rightarrow ClO^{-} + Cl^{-}$ (basic conditions), which is one of the reactions involved in the bleaching action of chlorine in basic solution

18. Balance each of the following net ionic equations. Then include the named spectator ions to write a balanced chemical equation. Include the states.
(a) $Co^{3+} + Cd \rightarrow Co^{2+} + Cd^{2+}$ (spectator ions NO_3^{-})
(b) $Ag^{+} + SO_2 \rightarrow Ag + SO_4^{2-}$ (acidic conditions; spectator ions NO_3^{-})
(c) $Al + CrO_4^{2-} \rightarrow Al(OH)_3 + Cr(OH)_3$ (basic conditions; spectator ions Na^{+})

19. If possible, give an example for each.
(a) a synthesis reaction that is a redox reaction
(b) a synthesis reaction that is not a redox reaction
(c) a decomposition reaction that is a redox reaction
(d) a decomposition reaction that is not a redox reaction
(e) a double displacement reaction that is a redox reaction
(f) a double displacement reaction that is not a redox reaction

20. Give an example of a reaction in which sulfur behaves as
(a) an oxidizing agent
(b) a reducing agent

21. Write a balanced equation for a synthesis reaction in which elemental oxygen acts as a reducing agent.

22. Phosphorus, $P_{4(s)}$, reacts with hot water to form phosphine, $PH_{3(g)}$, and phosphoric acid.
   (a) Write a balanced chemical equation for this reaction.
   (b) Is the phosphorus oxidized or reduced? Explain your answer.
23. The thermite reaction, which is highly exothermic, can be used to weld metals. In the thermite reaction, aluminum reacts with iron(III) oxide to form iron and aluminum oxide. The temperature becomes so high that the iron is formed as a liquid.
   (a) Write a balanced chemical equation for the reaction.
   (b) Is the reaction a redox reaction? If so, identify the oxidizing agent and the reducing agent.

Inquiry

24. Iodine reacts with concentrated nitric acid to form iodic acid, gaseous nitrogen dioxide, and water.
   (a) Write the balanced chemical equation.
   (b) Calculate the mass of iodine needed to produce 28.0 L of nitrogen dioxide at STP.
25. Describe a laboratory investigation you could perform to decide whether tin or nickel is the better reducing agent. Include in your description all the materials and equipment you would need, and the procedure you would follow.
26. The following table shows the average composition, by volume, of the air we inhale and exhale, as part of a biochemical process called respiration. (The values are rounded.)

Gas	Inhaled Air (% by volume)	Exhaled Air (% by volume)
Oxygen	21	16
Carbon dioxide	0.04	4
Nitrogen and other gases	79	80

   How do the data indicate that at least one redox reaction is involved in respiration?
27. Highly toxic phosphine gas, PH_3, is used in industry to produce flame retardants. One way to make phosphine on a large scale is by heating elemental phosphorus with a strong base.
   (a) Balance the following net ionic equation for the reaction under basic conditions.
   $P_4 \rightarrow H_2PO_2^- + PH_3$
   (b) Show that the reaction in part (a) is a disproportionation reaction.
   (c) Calculate the mass of phosphine that can theoretically be made from 10.0 kg of phosphorus by this method.

Communication

28. Explain why, in a redox reaction, the reducing agent undergoes oxidation.
29. Explain why you would not expect sulfide ions to act as an oxidizing agent.
30. Why can't the oxidation number of an element in a compound be greater than the number of valence electrons in one atom of that element?
31. Explain why the historical use of the word "reduction," that is, the production of a metal from its ore, is consistent with the modern definitions of reduction.
32. Organic chemists sometimes describe redox reactions in terms of the loss or gain of pairs of hydrogen atoms. Examples include the addition of hydrogen to ethene to form ethane, and the elimination of hydrogen from ethanol to form ethanal.
   (a) Write a balanced equation for each reaction.
   (b) Determine whether the organic reactant is oxidized or reduced in each reaction.
   (c) Write a definition of oxidation and a definition of reduction based on an organic reactant losing or gaining hydrogen.
   (d) Would your definitions be valid for the synthesis and decomposition of a metal hydride? Explain your answer.

Making Connections

33. The compound $NaAl(OH)_2CO_3$ is a component of some common stomach acid remedies.
    (a) Determine the oxidation number of each element in the compound.
    (b) Predict the products of the reaction of the compound with stomach acid (hydrochloric acid), and write a balanced chemical equation for the reaction.
    (c) Were the oxidation numbers from part (a) useful in part (b)? Explain your answer.
    (d) What type of reaction is this?
    (e) Check your medicine cabinet at home for stomach acid remedies. If possible, identify the active ingredient in each remedy.

34. Two of the substances on the head of a safety match are potassium chlorate and sulfur. When the match is struck, the potassium chlorate decomposes to give potassium chloride and oxygen. The sulfur then burns in the oxygen and ignites the wood of the match.
    (a) Write balanced chemical equations for the decomposition of potassium chlorate and for the burning of sulfur in oxygen.
    (b) Identify the oxidizing agent and the reducing agent in each reaction in part (a).
    (c) Does any element in potassium chlorate undergo disproportionation in the reaction? Explain your answer.
    (d) Research the history of the safety match to determine when it was invented, why it was invented, and what it replaced.

35. Ammonium ions, from fertilizers or animal waste, are oxidized by atmospheric oxygen. The reaction results in the acidification of soil on farms and the pollution of ground water with nitrate ions.
    (a) Write a balanced net ionic equation for this reaction.
    (b) Why do farmers use fertilizers? What alternative farming methods have you heard of? Which farming method(s) do you support, and why?

36. One of the most important discoveries in the history of the chemical industry in Ontario was accidental. Thomas "Carbide" Willson (1860–1915) was trying to make the element calcium from lime, CaO, by heating the lime with coal tar. Instead, he made the compound calcium carbide, CaC_2. This compound reacts with water to form a precipitate of calcium hydroxide and gaseous ethyne (acetylene). Willson's discovery led to the large-scale use of ethyne in numerous applications.
    (a) Was Willson trying to perform a redox reaction? How do you know? Why do you not need to know the substances in coal tar to answer this question?
    (b) Write a balanced chemical equation for the reaction of calcium carbide with water. Is this reaction a redox reaction?
    (c) An early use of Willson's discovery was in car headlights. Inside a headlight, the reaction of calcium carbide and water produced ethyne, which was burned to produce light and heat. Write a balanced chemical equation for the complete combustion of ethyne. Is this reaction a redox reaction?
    (d) Research the impact of Willson's discovery on society, from his lifetime to the present day.

Answers to Practice Problems and Short Answers to Section Review Questions

Practice Problems: 1. $Zn_{(s)} + Fe^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Fe_{(s)}$
2.(a) $3Mg_{(s)} + 2Al^{3+}{}_{(aq)} \rightarrow 3Mg^{2+}{}_{(aq)} + 2Al_{(s)}$
(b) $2Ag^{+}{}_{(aq)} + Cd_{(s)} \rightarrow 2Ag_{(s)} + Cd^{2+}{}_{(aq)}$
3.(a) Mg oxidized, Al^{3+} reduced
(b) Cd oxidized, Ag^{+} reduced
4.(a) Al^{3+} oxidizing agent, Mg reducing agent
(b) Ag^{+} oxidizing agent, Cd reducing agent
5. $Al_{(s)} \rightarrow Al^{3+}{}_{(aq)} + 3e^{-}$, $Fe^{3+}{}_{(aq)} + 3e^{-} \rightarrow Fe_{(s)}$
6.(a) $Fe_{(s)} \rightarrow Fe^{2+}{}_{(aq)} + 2e^{-}$, $Cu^{2+}{}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)}$
(b) $Cd_{(s)} \rightarrow Cd^{2+}{}_{(aq)} + 2e^{-}$, $Ag^{+}{}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$
7.(a) $Sn_{(s)} \rightarrow Sn^{2+}{}_{(aq)} + 2e^{-}$, $Pb^{2+}{}_{(aq)} + 2e^{-} \rightarrow Pb_{(s)}$
(b) $Ag_{(s)} \rightarrow Ag^{+} + e^{-}$, $Au^{3+}{}_{(aq)} + 3e^{-} \rightarrow Au_{(s)}$
(c) $Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^{-}$, $Fe^{3+}{}_{(aq)} + 3e^{-} \rightarrow Fe_{(s)}$
8. $Hg_2{}^{2+}{}_{(aq)} \rightarrow Hg_{(\ell)} + Hg^{2+}{}_{(aq)}$, $Hg_2{}^{2+}{}_{(aq)} + 2e^{-} \rightarrow 2Hg_{(\ell)}$, $Hg_2{}^{2+}{}_{(aq)} \rightarrow 2Hg^{2+}{}_{(aq)} + 2e^{-}$
9.(a) +3 **(b)** 0 **(c)** +6 **(d)** +5 **(e)** 0 **(f)** +2

10.(a) H, +1; S, +4; O, −2
(b) H, +1; O, −2 **(c)** H, +1; P, +5; O, −2
11.(a) +2 **(b)** −1 **12.(a)** Al, +3; H, +1; C, +4; O, −2
(b) N, −3; H, +1; P, +5; O, −2 **(c)** K, +1; H, +1; I, +7; O, −2
13.(a) yes **(b)** no
14.(a) H_2O_2 oxidizing agent, Fe^{2+} reducing agent
15. ClO_2^- oxidized, Br_2 reduced **16.** yes
17. $Ce^{4+} + e^- \rightarrow Ce^{3+}$ **18.** $2Br^- \rightarrow Br_2 + 2e^-$
19.(a) $O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$ **(b)** $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$
(c) $2NO_3^- + 12H^+ + 10e^- \rightarrow N_2 + 6H_2O$
20.(a) $ClO_3^- + 6H^+ + 6e^- \rightarrow Cl^- + 3H_2O$
(b) $NO + 2H_2O \rightarrow NO_3^- + 4H^+ + 3e^-$
(c) $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
21. $Cr^{2+} \rightarrow Cr^{3+} + e^-$ **22.** $O_2 + 4e^- \rightarrow 2O^{2-}$
23.(a) $Al + 4OH^- \rightarrow Al(OH)_4^- + 3e^-$
(b) $CN^- + 2OH^- \rightarrow CNO^- + H_2O + 2e^-$
(c) $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$
(d) $CrO_4^{2-} + 4H_2O + 3e^- \rightarrow Cr(OH)_3 + 5OH^-$
(e) $2CO_3^{2-} + 2H_2O + 2e^- \rightarrow C_2O_4^{2-} + 4OH^-$
24.(a) $FeO_4^{2-} + 8H^+ + 3e^- \rightarrow Fe^{3+} + 4H_2O$
(b) $ClO_2^- + 2H_2O + 4e^- \rightarrow Cl^- + 4OH^-$
25.(a) $2Na + F_2 \rightarrow 2NaF$
ox: $Na \rightarrow Na^+ + e^-$
red: $F_2 + 2e^- \rightarrow 2F^-$
(b) $3Mg + N_2 \rightarrow Mg_3N_2$
ox: $Mg \rightarrow Mg^{2+} + 2e^-$
red: $N_2 + 6e^- \rightarrow 2N^{3-}$
(c) $2HgO \rightarrow 2Hg + O_2$
ox: $2O^{2-} \rightarrow O_2 + 4e^-$
red: $Hg^{2+} + 2e^- \rightarrow Hg$
26. $2Cu^{2+} + 5I^- \rightarrow 2CuI + I_3^-$
27.(a) $MnO_4^- + 5Ag + 8H^+ \rightarrow Mn^{2+} + 5Ag^+ + 4H_2O$
oxidizing agent, MnO_4^-; reducing agent, Ag
(b) $Hg + 2NO_3^- + 4Cl^- + 4H^+ \rightarrow HgCl_4^{2-} + 2NO_2 + 2H_2O$
oxidizing agent, NO_3^-; reducing agent, Hg
(c) $AsH_3 + 4Zn^{2+} + 4H_2O \rightarrow H_3AsO_4 + 4Zn + 8H^+$
oxidizing agent, Zn^{2+}; reducing agent, AsH_3
(d) $3I_3^- + 3H_2O \rightarrow 8I^- + IO_3^- + 6H^+$
oxidizing agent, I_3^-; reducing agent, I_3^-
28.(a) $3CN^- + 2MnO_4^- + H_2O \rightarrow 3CNO^- + 2MnO_2 + 2OH^-$
oxidizing agent, MnO_4^-; reducing agent, CN^-
(b) $H_2O_2 + 2ClO_2 + 2OH^- \rightarrow 2ClO_2^- + O_2 + 2H_2O$
oxidizing agent, ClO_2; reducing agent, H_2O_2
(c) $6ClO^- + 2CrO_2^- + 2H_2O \rightarrow 3Cl_2 + 2CrO_4^{2-} + 4OH^-$
oxidizing agent, ClO^-; reducing agent, CrO_2^-
(d) $2Al + NO_2^- + H_2O + OH^- \rightarrow NH_3 + 2AlO_2^-$
oxidizing agent, NO_2^-; reducing agent, Al
29. $CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$
30.(a) $B_2O_3 + 6Mg \rightarrow 3MgO + Mg_3B_2$
(b) $8H_2S + 8H_2O_2 \rightarrow S_8 + 16H_2O$
31.(a) $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$
(b) $I_2 + 10NO_3^- + 8H^+ \rightarrow 2IO_3^- + 10NO_2 + 4H_2O$
(c) $2PbSO_4 + 2H_2O \rightarrow Pb + PbO_2 + 2SO_4^{2-} + 4H^+$
32.(a) $3Cl^- + 2CrO_4^{2-} + H_2O \rightarrow 3ClO^- + 2CrO_2^- + 2OH^-$
(b) $3Ni + 2MnO_4^- + H_2O \rightarrow 3NiO + 2MnO_2 + 2OH^-$
(c) $I^- + 6Ce^{4+} + 6OH^- \rightarrow IO_3^- + 6Ce^{3+} + 3H_2O$
Section Review 10.1: **1.(a)** yes;
$2AgNO_{3(aq)} + Cd_{(s)} \rightarrow 2Ag_{(s)} + Cd(NO_3)_{2(aq)}$;
$2Ag^+_{(aq)} + Cd_{(s)} \rightarrow 2Ag_{(s)} + Cd^{2+}_{(aq)}$;
$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$; $Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2e^-$
(b) no **(c)** yes; $2Al_{(s)} + 3HgCl_{2(aq)} \rightarrow 2AlCl_{3(aq)} + 3Hg_{(\ell)}$;
$2Al_{(s)} + 3Hg^{2+}_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3Hg_{(\ell)}$;
$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$; $Hg^{2+}_{(aq)} + 2e^- \rightarrow Hg_{(\ell)}$
2.(a) right side **(b)** left side **4.** reducing agent
6.(a) $K^+ + Na_{(\ell)} \rightarrow Na^+ + K_{(\ell)}$
(b) oxidizing agent, K^+; reducing agent, $Na_{(\ell)}$
10.2: **2.(a)** yes **(b)** no **(c)** no **(d)** yes
6.(a) +2 **(b)** +5 and −1
7.(a) $N_2 + 3H_2 \rightarrow 2NH_3$ (0, 0, −3, +1); N_2 is oxidizing agent, H_2 is reducing agent **(b)** no
8.(a) $2C + O_2 \rightarrow 2CO$, $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
(b) carbon—reducing agent, oxygen—oxidizing agent;
carbon monoxide—reducing agent,
iron(III) oxide—oxidizing agent
10.3: **1.(a)** $2C + 2e^- \rightarrow C_2^{2-}$, reduction
(b) $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$, oxidation
(c) $4AsO_4^{3-} + 20H^+ + 8e^- \rightarrow As_4O_6 + 10H_2O$, reduction
(d) $Br_2 + 12OH^- \rightarrow 2BrO_3^- + 6H_2O + 10e^-$, oxidation
2.(a) $3Co^{3+} + Au \rightarrow 3Co^{2+} + Au^{3+}$
(b) $3Cu + 2NO_3^- + 8H^+ \rightarrow 3Cu^{2+} + 2NO + 4H_2O$
(c) $3NO_3^- + 8Al + 5OH^- + 2H_2O \rightarrow 3NH_3 + 8AlO_2^-$
3. $2Ag^+ + HCHO + H_2O \rightarrow 2Ag + HCOOH + 2H^+$
4.(b) $2NH_3 + 8OH^- \rightarrow N_2O + 7H_2O + 8e^-$
5.(a) $2N_2H_4 + N_2O_4 \rightarrow 3N_2 + 4H_2O$
(b) N_2O_4 is oxidizing agent, N_2H_4 is reducing agent
(c) $2NH_3 + ClO^- \rightarrow N_2H_4 + Cl^- + H_2O$
6.(a) Larissa was right
(b) $3Zn + SO_4^{2-} + 8H^+ \rightarrow 3Zn^{2+} + S + 4H_2O$
10.4: **1.** no, not a redox reaction
2.(a) $CH_3COOH + 2O_2 \rightarrow 2CO_2 + 2H_2O$
(b) $O_2 + 2H_2SO_3 \rightarrow 2HSO_4^- + 2H^+$
3.(a) $8NH_3 + 3Cl_2 \rightarrow 6NH_4Cl + N_2$
(b) $3Mn_3O_4 + 8Al \rightarrow 4Al_2O_3 + 9Mn$