CHAPTER 19

Redox Reactions

BIG Idea **Oxidation-reduction reactions—among the most-common chemical processes in both nature and industry—involve the transfer of electrons.**

19.1 Oxidation and Reduction

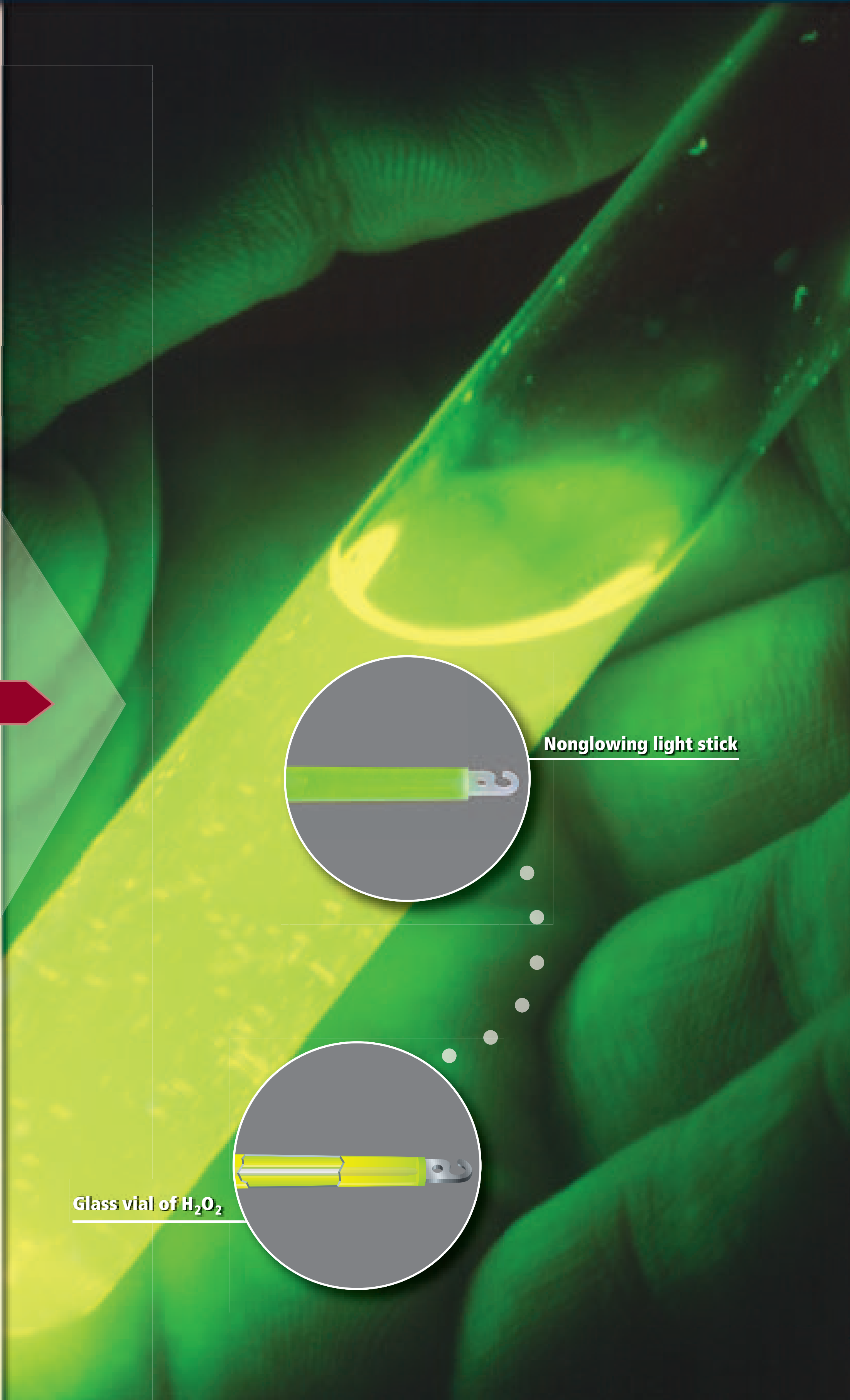
MAIN Idea Oxidation and reduction are complementary—as an atom is oxidized, another atom is reduced.

19.2 Balancing Redox Equations

MAIN Idea Redox equations are balanced when the total increase in oxidation numbers equals the total decrease in oxidation numbers of the atoms involved in the reaction.

ChemFacts

- The glow of an activated light stick can be made brighter by warming it, although the glow will not last as long.
- Light generated by redox reactions doesn't generally result in the formation of heat.
- About 90% of marine life uses some form of bioluminescence—generating light through redox reactions.



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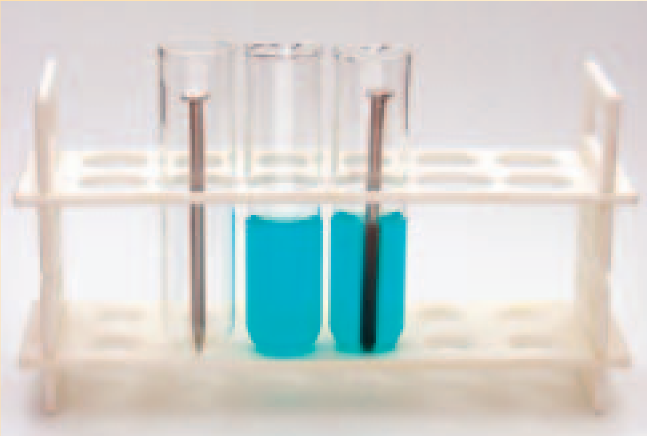
Start-Up Activities

LAUNCH Lab

What happens when iron and copper(II) sulfate react?

Rust is the product of a reaction between iron and oxygen. Iron can also react with substances other than oxygen.



Procedure

1. Read and complete the lab safety form.
2. Use a piece of **steel wool** to polish the end of an **iron nail.**
3. Add about 3 mL of **1.0*M* copper (II) sulfate ($CuSO_4$)** solution to a **test tube.** Place the polished end of the nail into the $CuSO_4$ solution. Let the test tube stand in a **test-tube rack,** and observe it for about 10 min. Record your observations.

Analysis

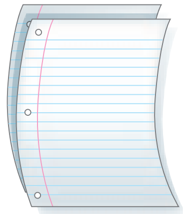
1. **Explain** what happened to the color of the copper(II) sulfate solution.
2. **Identify** the substance clinging to the nail.
3. **Write** the balanced chemical equation for the reaction you observed.

Inquiry What do you think would happen if copper was placed in an iron sulfate solution? Design an investigation to test your hypothesis.

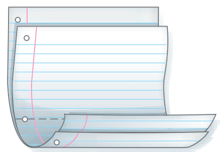
FOLDABLES™ Study Organizer

Balancing Redox Equations Make the following Foldable to help you summarize information about the different methods of balancing redox equations.

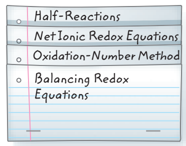
STEP 1 Collect two sheets of paper, and layer them about 2 cm apart vertically.



STEP 2 Fold up the bottom edges of the paper to form three equal tabs. Crease the fold to hold the tabs in place.



STEP 3 Staple along the fold. Label as follows: *Balancing Redox Equations*, *Oxidation-Number Method*, *Net Ionic Redox Equations*, and *Half-Reactions*.



FOLDABLES **Use this Foldable with Section 19.2.** As you read about balancing redox equations, summarize and provide an example of each method.

Chemistry Online

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Section 19.1

Objectives

- **Describe** the processes of oxidation and reduction.
- **Identify** oxidizing and reducing agents.
- **Determine** the oxidation number of an element in a compound.
- **Interpret** redox reactions in terms of change in oxidation state.

Review Vocabulary

spectator ion: an ion that does not participate in a reaction and is not usually shown in an ionic equation

New Vocabulary

oxidation-reduction reaction
redox reaction
oxidation
reduction
oxidizing agent
reducing agent

Oxidation and Reduction

MAIN Idea Oxidation and reduction are complementary—as an atom is oxidized, another atom is reduced.

Real-World Reading Link The light produced by a light stick is the result of a chemical reaction. When you snap the glass capsule inside the plastic case, two chemicals are mixed and electron transfer occurs. As the electrons are transferred, chemical energy is converted into light energy.

Electron Transfer and Redox Reactions

In Chapter 9, you learned that a chemical reaction can usually be classified as one of five types—synthesis, decomposition, combustion, single-replacement, or double-replacement. A defining characteristic of combustion and single-replacement reactions is that they always involve the transfer of electrons from one atom to another, as do many synthesis and decomposition reactions. For example, in the synthesis reaction in which sodium (Na) and chlorine (Cl_2) react to form the ionic compound sodium chloride (NaCl), an electron from each of two sodium atoms is transferred to the Cl_2 molecule to form two Cl^- ions.

Complete chemical equation: $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

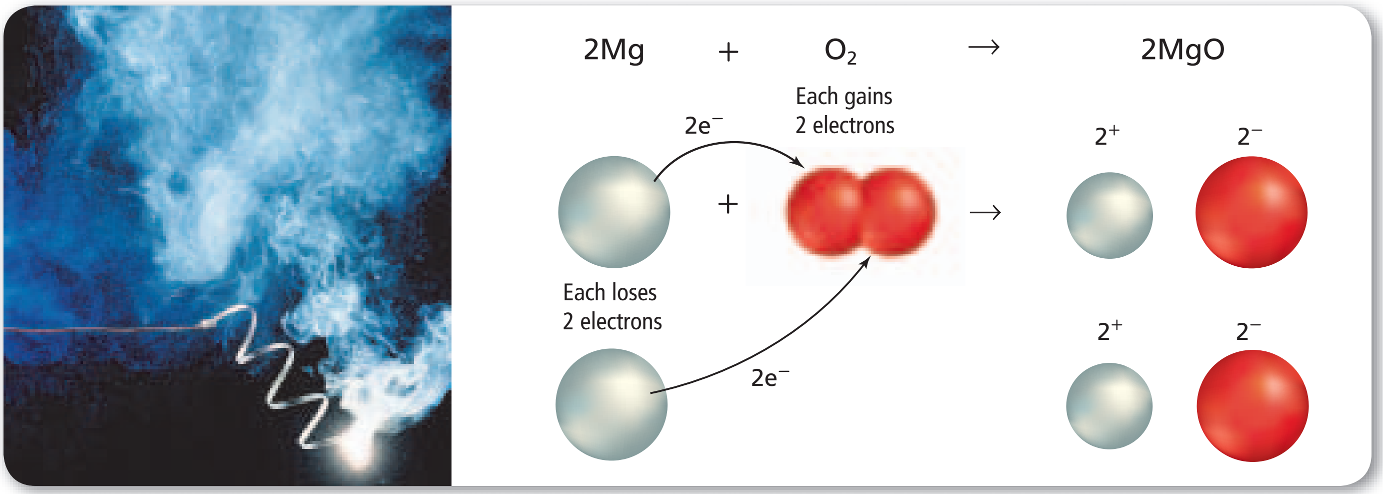
Net ionic equation: $2Na(s) + Cl_2(g) \longrightarrow 2Na^+ + 2Cl^-$ (ions in crystal)

An example of a combustion reaction is the burning of magnesium in air, which involves the transfer of electrons.

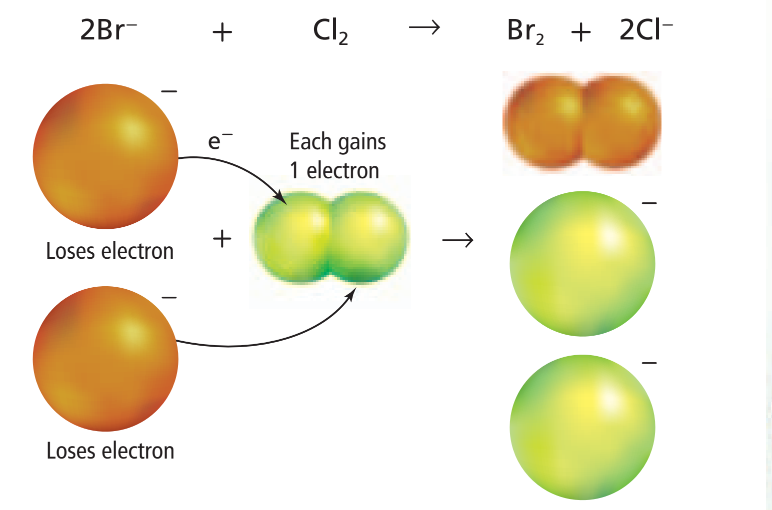
Complete chemical equation: $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

Net ionic equation: $2Mg(s) + O_2(g) \longrightarrow 2Mg^{2+} + 2O^{2-}$ (ions in crystal)

When magnesium reacts with oxygen, as illustrated in **Figure 19.1,** each magnesium atom transfers two electrons to each oxygen atom. The two magnesium atoms become magnesium ions (Mg^{2+}), and the two oxygen atoms become oxide ions (O^{2-}). A reaction in which electrons are transferred from one atom to another is called an **oxidation-reduction reaction,** which is also called a **redox reaction.**



■ **Figure 19.1** The reaction of magnesium and oxygen involves a transfer of electrons from magnesium to oxygen. Therefore, this reaction is an oxidation-reduction reaction.
Classify *the reaction between magnesium and oxygen.*



■ **Figure 19.2** The reaction between aqueous bromide ions and chlorine gas is a redox reaction. Here, electrons are transferred from bromide ions to chlorine.

Concepts In Motion

Interactive Figure To see an animation of a redox reaction, visit glencoe.com.

Consider the single-replacement reaction in which chlorine in an aqueous solution reacts with bromide ions from an aqueous solution of potassium bromide, which is shown in **Figure 19.2.**

Complete chemical equation: $2KBr(aq) + Cl_2(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$

Net ionic equation: $2Br^-(aq) + Cl_2(aq) \longrightarrow Br_2(aq) + 2Cl^-(aq)$

Note that chlorine "takes" electrons from bromide ions to become chloride ions. When the two bromide ions lose electrons, the two bromine atoms form a covalent bond with each other to produce Br_2 molecules. The formation of the covalent bond by sharing of electrons is also an oxidation-reduction reaction.

Oxidation and reduction Originally, the word *oxidation* referred only to reactions in which a substance combined with oxygen. Today, **oxidation** is defined as the loss of electrons from atoms of a substance. Look again at the net ionic equation for the reaction of sodium and chlorine. Sodium is oxidized because it loses an electron.

$$\text{Oxidation: } Na \longrightarrow Na^+ + e^-$$

For oxidation to occur, the electrons lost by the substance that is oxidized must be accepted by atoms or ions of another substance. In other words, there must be an accompanying process that involves the gain of electrons. **Reduction** is the gain of electrons by atoms of a substance. Following the sodium chloride example further, the reduction reaction that accompanies the oxidation of sodium is the reduction of chlorine.

$$\text{Reduction: } Cl_2 + 2e^- \longrightarrow 2Cl^-$$

Oxidation and reduction are complementary processes; oxidation cannot occur unless reduction also occurs. It is important to recognize and distinguish between oxidation and reduction. A memory aid might help you remember the distinction. The phrase **L**oss of **E**lectrons is **O**xidation, and **G**ain of **E**lectrons is **R**eduction is shortened to **LEO GER.**

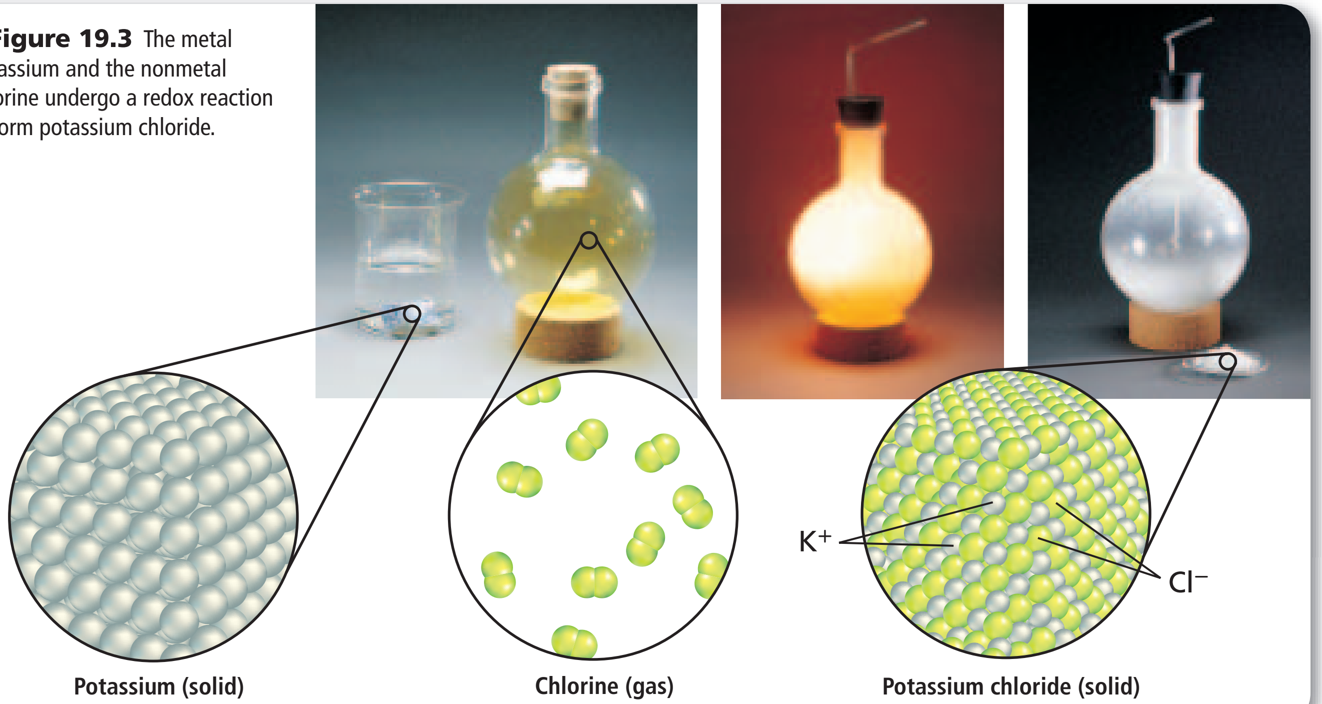
LEO the lion says **GER** or, for short, **LEO GER.**

VOCABULARY

WORD ORIGIN

Reduction

comes from the Latin *re*, meaning *back*, and *ducere*, meaning *to lead*



■ **Figure 19.3** The metal potassium and the nonmetal chlorine undergo a redox reaction to form potassium chloride.

Changes in oxidation number You might recall from previous chapters that the oxidation number of an atom in an ionic compound is the number of electrons lost or gained by the atom when it forms ions. The reaction of potassium with chlorine, shown in **Figure 19.3,** is a redox reaction. The equation for the reaction of potassium metal with chloride vapor is as follows.

Complete chemical equation: $2K(s) + Cl_2(g) \longrightarrow 2KCl(s)$

Net ionic equation: $2K(s) + Cl_2(g) \longrightarrow 2K^+(s) + 2Cl^-(s)$

Potassium, a group 1 element that tends to lose one electron in reactions because of its low electronegativity, is assigned an oxidation number of +1. On the other hand, chlorine, a group 17 element that tends to gain one electron in reactions because of its high electronegativity, is assigned an oxidation number of −1. In redox terms, you would say that potassium atoms are oxidized from 0 to the +1 state because each atom loses an electron, and chlorine atoms are reduced from 0 to the −1 state because each atom gains an electron. When an atom or ion is reduced, the numerical value of its oxidation number decreases. Conversely, when an atom or ion is oxidized, its oxidation number increases.

Oxidation numbers are tools that scientists use in written chemical equations to help them keep track of the movement of electrons in a redox reaction. Like some of the other tools you have learned about, oxidation numbers have a specific notation. Oxidation numbers are written with the positive or negative sign before the number (+3, +2), whereas ionic charge is written with the sign after the number (3+, 2+).

Oxidation number: +3 Ionic charge: 3+

Reading Check **Determine** Which element is more likely to gain electrons, potassium or chlorine?

CAREERS IN CHEMISTRY

Potter A potter is an artist who makes pottery. He or she uses glazes containing metallic ions that have multiple oxidation states to achieve a variety of colors on ceramics. Glazes that contain copper ions produce a green-to-blue color when oxidized, and they produce a reddish color when reduced in a kiln. For more information on chemistry and careers, visit glencoe.com.

Oxidizing and Reducing Agents

The potassium-chlorine reaction in **Figure 19.3** can also be described by saying that "potassium is oxidized by chlorine." This description is useful because it clearly identifies both the substance that is oxidized and the substance that does the oxidizing. The substance that oxidizes another substance by accepting its electrons is called an **oxidizing agent.** This term describes the substance that is reduced. The substance that reduces another substance by losing electrons is called a **reducing agent.** A reducing agent supplies electrons to the substance being reduced (gaining electrons). The reducing agent is oxidized because it loses electrons. The reducing agent in the potassium-chlorine reaction is potassium—the substance that is oxidized.

$$\overbrace{2\text{K}(s) + \underbrace{\text{Cl}_2(g) \longrightarrow}_{\text{reduced}}}^{\text{oxidized}} 2\text{KCl}(s)$$

Oxidizing agent: Cl_2

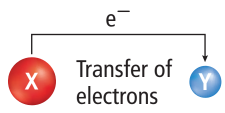
Reducing agent: K

A common application of redox chemistry is to remove tarnish from metal objects. Other oxidizing agents and reducing agents are useful in everyday life. For example, when you add chlorine bleach to your laundry to whiten clothes, you are using an aqueous solution of sodium hypochlorite (NaClO), an oxidizing agent. It oxidizes dyes, stains, and other materials that discolor clothes. **Table 19.1** summarizes the different ways to describe oxidation-reduction reactions.

Concepts In Motion

Table 19.1 Summary of Redox Reactions

Interactive Table Explore redox reactions at glencoe.com.

Process	
Oxidation • A reactant loses an electron. • Reducing agent is oxidized. • Oxidation number increases.	• X loses an electron. • X is the reducing agent and becomes oxidized. • The oxidation number of X increases.
Reduction • Other reactant gains an electron. • Oxidizing agent is reduced. • Oxidation number decreases.	• Y gains an electron. • Y is the oxidizing agent and becomes reduced. • The oxidation number of Y decreases.

MiniLab

Observe a Redox Reaction

How can tarnish be removed from silver?

Procedure

1. Read and complete the lab safety form.
2. Lightly buff a piece of **aluminum foil** with **steel wool** to remove any oxide coating.
3. Wrap a **small tarnished object** in the aluminum foil, making sure that the tarnished area makes firm contact with the foil.
4. Place the wrapped object in a **400-mL beaker** and add a sufficient volume of **tap water** to cover it completely.
5. Add about 1 spoonful of **baking soda** and about 1 spoonful of **table salt** to the beaker.
6. Using **beaker tongs,** set the beaker and its contents on a **hot plate,** and heat until the water is almost boiling. Maintain the heat for approximately 15 min, until the tarnish disappears.

Analysis

1. **Write** the equation for the reaction of silver with hydrogen sulfide that yields silver sulfide and hydrogen.
2. **Write** the equation for the reaction of the tarnish (silver sulfide) with the aluminum foil that yields aluminum sulfide and silver.
3. **Determine** which metal, aluminum or silver, is more reactive. How do you know this from your results?
4. **Explain** why you should not use an aluminum pan to clean silver objects.

Redox and Electronegativity

The chemistry of oxidation-reduction reactions is not limited to atoms of an element changing to ions or the reverse. Some redox reactions involve changes in molecular substances or polyatomic ions in which atoms are covalently bonded to other atoms. For example, the following equation represents the redox reaction used to manufacture ammonia (NH_3).

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

This process involves neither ions nor any obvious transfer of electrons. The reactants and products are all molecular compounds. Yet, it is still a redox reaction in which nitrogen is the oxidizing agent and hydrogen is the reducing agent.

In situations such as the formation of ammonia, where two atoms share electrons, how is it possible to say that one atom lost electrons and was oxidized, while the other atom gained electrons and was reduced? To answer this, you need to know which atom attracts electrons more strongly, or, in other words, which atom is more electronegative. You might find it helpful to review the discussion of electronegativity trends in Chapters 6 and 8. **Figure 19.4** shows that electronegativity increases left to right across a period and generally decreases down a group. Elements with low electronegativity (Groups 1 and 2) are strong reducing agents, and those with high electronegativity (Group 17 and oxygen in Group 16) are strong oxidizing agents.

$$\overbrace{N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)}^{\text{reduced (partial gain of } e^-)}$$

$$\underbrace{3H_2(g) \longrightarrow 2NH_3(g)}_{\text{oxidized (partial loss of } e^-)}$$

Hydrogen has an electronegativity of 2.20, and nitrogen's electronegativity is 3.04. For the purpose of studying oxidation-reduction reactions, the more-electronegative atom (in this case nitrogen) is treated as if it had been reduced by gaining electrons from the other atom (hydrogen). Conversely, the less-electronegative atom (hydrogen) is treated as if it had been oxidized by losing electrons to the other atom (nitrogen).

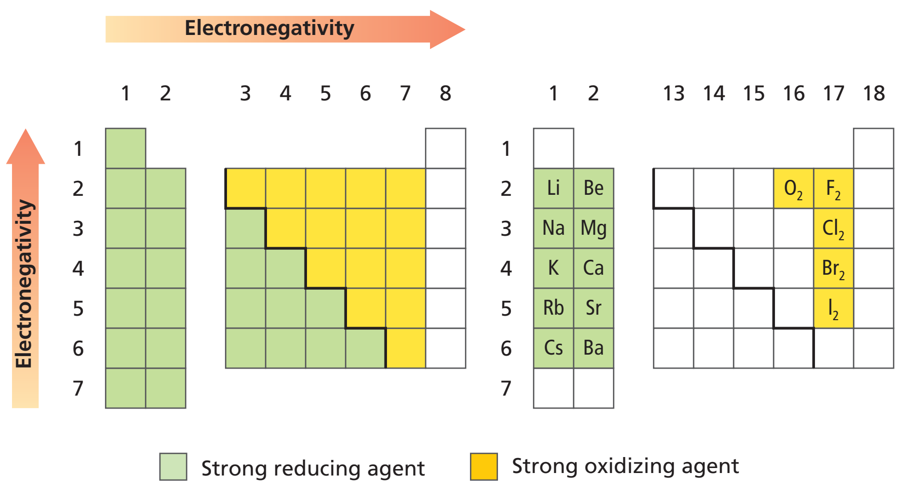


Figure 19.4 The electronegativity of elements increases from left to right across the periodic table, and it decreases going down a group. Elements with low electronegativity are strong reducing agents, and elements with high electronegativity are strong oxidizing agents.

Predict ***which element would be the strongest oxidizing agent. Which is the strongest reducing agent?***

EXAMPLE Problem 19.1

Identify Oxidation-Reduction Reactions The following equation represents the redox reaction of aluminum and iron.

$$2Al + 2Fe^{3+} + 3O^{2-} \rightarrow 2Fe + 2Al^{3+} + 3O^{2-}$$

Identify what is oxidized and what is reduced in this reaction. Identify the oxidizing agent and the reducing agent.

1 Analyze the Problem

You are given the reactants and products in the reaction. You must determine the electron transfers that occur. Then, you can apply the definitions of oxidizing agent and reducing agent to answer the question.

2 Solve for the Unknown

Identify the oxidation process and the reduction process.

$Al \rightarrow Al^{3+} + 3e^-$ (loss of e^- is oxidation) — **Aluminum loses three electrons and becomes an aluminum ion.**

$Fe^{3+} + 3e^- \rightarrow Fe$ (gain of e^- is reduction) — **The iron ion accepts the three electrons lost from aluminum.**

Aluminum is oxidized and is therefore the reducing agent.
Iron is reduced and is therefore the oxidizing agent.

3 Evaluate the Answer

In this process, aluminum lost electrons and was oxidized, whereas iron gained electrons and was reduced. The definitions of oxidation, reduction, oxidizing agent, and reducing agent apply. Note that the oxidation number of oxygen is unchanged in this reaction; therefore, oxygen is not a key factor in this problem.

PRACTICE Problems

Extra Practice Pages 989–990 and glencoe.com

1. Identify each of the following changes as either oxidation or reduction. Recall that e^- is the symbol for an electron.
   - **a.** $I_2 + 2e^- \rightarrow 2I^-$
   - **b.** $K \rightarrow K^+ + e^-$
   - **c.** $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   - **d.** $Ag^+ + e^- \rightarrow Ag$
2. Identify what is oxidized and what is reduced in the following processes.
   - **a.** $2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$
   - **b.** $2Ce + 3Cu^{2+} \rightarrow 3Cu + 2Ce^{3+}$
   - **c.** $2Zn + O_2 \rightarrow 2ZnO$
   - **d.** $2Na + 2H^+ \rightarrow 2Na^+ + H_2$
3. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

$$Fe(s) + Ag + (aq) \rightarrow Fe^{2+}(aq) + Ag(s)$$

4. **Challenge** Identify the oxidizing agent and the reducing agent in each reaction.
   - **a.** $Mg + I_2 \rightarrow MgI_2$
   - **b.** $H_2S + Cl_2 \rightarrow S + 2HCl$

Real-World Chemistry
Oxidation

Rust When moist air comes in contact with iron, the iron oxidizes. Iron oxide (Fe_2O_3), called rust, is common because iron combines readily with oxygen. Pure iron is uncommon in nature. Steel, a mixture that contains iron, is a commonly used form of iron. Several protective methods, such as plating, painting, and applying an enamel or plastic coating, can inhibit the production of iron oxide.

■ **Figure 19.5** Banded iron—shown in this cross-section of rock—is a result of different oxidation states of iron, which depends on which mineral is present.

Determining Oxidation Numbers

In order to understand all types of redox reactions, you must have a way to determine the oxidation number ($n_{element}$) of the atoms involved in the reaction. **Table 19.2** outlines the rules chemists use to make this determination easier.

Many elements other than those specified in the rules below, including most of the transition metals, metalloids, and nonmetals, can be found with different oxidation numbers in different compounds. For example, iron has different oxidation numbers, indicated by the different colors as shown in **Figure 19.5**, depending on which mineral is also present.

Table 19.2 Rules for Determining Oxidation Numbers

Rule	Example	$n_{element}$
1. The oxidation number of an uncombined atom is zero.	Na, O_2, Cl_2, H_2	0
2. The oxidation number of a monatomic ion is equal to the charge of the ion.	Ca^{2+}	+2
	Br^-	−1
3. The oxidation number of the more-electronegative atom in a molecule or a complex ion is the same as the charge it would have if it were an ion.	N in NH_3	−3
	O in NO	−2
4. The oxidation number of the most-electronegative element, fluorine, is always −1 when it is bonded to another element.	F in LiF	−1
5. The oxidation number of oxygen in compounds is always −2 except in peroxides, such as hydrogen peroxide (H_2O_2), where it is −1. When it is bonded to fluorine, the only element more electronegative than oxygen, the oxidation number of oxygen is positive.	O in NO_2	−2
	O in H_2O_2	−1
6. The oxidation number of hydrogen in most of its compounds is +1, except in metal hydrides; then, the oxidation number is −1.	H in NaH	−1
7. The oxidation numbers of group 1 and 2 metals and aluminum are positive and equal to their number of valence electrons.	K	+1
	Ca	+2
	Al	+3
8. The sum of the oxidation numbers in a neutral compound is zero.	$CaBr_2$	$(+2) + 2(-1) = 0$
9. The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to the charge of the ion.	SO_3^{2-}	$(+4) + 3(-2) = -2$

EXAMPLE Problem 19.2

Determine Oxidation Numbers Use the rules for determining oxidation numbers to find the oxidation number of each element in potassium chlorate ($KClO_3$) and in a sulfite ion (SO_3^{2-}).

1 Analyze the Problem

In the rules for determining oxidation numbers, you are given the oxidation numbers of oxygen and potassium. You are also given the overall charge of the compound or ion. Using this information and applying the rules, determine the oxidation numbers of chlorine and sulfur. (Let $n_{element}$ equal the oxidation number of the element in question.)

Known	Unknown
$KClO_3$	$n_{Cl} = ?$
SO_3^{2-}	$n_S = ?$
$n_O = -2$	
$n_K = +1$	

2 Solve for the Unknown

Assign the known oxidation numbers to their elements, set the sum of all oxidation numbers to zero or to the ion charge, and solve for the unknown oxidation number.

$(n_K) + (n_{Cl}) + 3\,(n_O) = 0$

$(+1) + (n_{Cl}) + 3(-2) = 0$

The sum of the oxidation numbers in a neutral compound is zero. For group 1 metals, $n_{element} = +1$. Substitute $n_K = +1$, $n_O = -2$.

$1 + n_{Cl} + (-6) = 0$

$n_{Cl} = +5$

Solve for n_{Cl}.

$(n_S) + 3\,(n_O) = -2$

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

The sum of the oxidation numbers in a polyatomic ion equals the charge on the ion. Substitute $n_O = -2$.

$n_S + (-6) = -2$

$n_S = +4$

Solve for n_S.

3 Evaluate the Answer

The rules for determining oxidation numbers have been correctly applied. All of the oxidation numbers in each substance add up to the proper value.

PRACTICE Problems

Extra Practice Pages 989–990 and glencoe.com

5. Determine the oxidation number of the boldface element in the following formulas for compounds.

a. Na**Cl**O_4 **b.** Al**P**O_4 **c.** H**N**O_2

6. Determine the oxidation number of the boldface element in the following formulas for ions.

a. **N**H_4^+ **b.** **As**O_4^{3-} **c.** **Cr**O_4^{2-}

7. Determine the oxidation number of nitrogen in each of these molecules or ions.

a. NH_3 **b.** KCN **c.** N_2H_4

8. **Challenge** Determine the net change of oxidation number of each of the elements in these redox equations.

a. $C + O_2 \longrightarrow CO_2$

b. $Cl_2 + ZnI_2 \longrightarrow ZnI_2 + I_2$

c. $CdO + CO \longrightarrow Cd + CO_2$

Table 19.3 Various Oxidation Numbers

Oxidation Number	+1	+2	+3	−1	−2
Aluminum			X		
Barium		X			
Bromine				X	
Cadmium		X			
Calcium		X			
Cesium	X				
Chlorine				X	
Fluoride				X	
Hydrogen	X			X	
Iodine				X	
Lithium	X				
Magnesium		X			
Oxygen					X
Potassium	X				
Sodium	X				
Silver	X				
Strontium		X			

Oxidation Numbers in Redox Reactions

Having studied oxidation numbers, you should be able to relate oxidation-reduction reactions to changes in oxidation number. Refer to the equation for a reaction that you saw at the beginning of this section—the replacement of bromine in aqueous potassium bromide (KBr) by chlorine (Cl_2).

$$2KBr(aq) + Cl_2(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$$

To learn how oxidation numbers change, start by assigning numbers, using **Table 19.3,** to all elements in the balanced equation. Then, review the changes, as shown in the equation below.

change: −1 reduction

change: +1 oxidation

$$2K\overset{+1}{}Br\overset{-1}{}(aq) + \overset{0}{Cl_2}(aq) \longrightarrow 2K\overset{+1}{}Cl\overset{-1}{}(aq) + \overset{0}{Br_2}(aq)$$

no change in oxidation number

You should notice that the oxidation number of bromine changed from −1 to 0, an increase of 1. At the same time, the oxidation number of chlorine changed from 0 to −1, a decrease of 1. Therefore, chlorine is reduced and bromine is oxidized. All redox reactions follow the same pattern. When an atom is oxidized, its oxidation number increases. When an atom is reduced, its oxidation number decreases. Note that there is no change in the oxidation number of potassium. The potassium ion takes no part in the reaction and is therefore a spectator ion.

Section 19.1 Assessment

Section Summary

- Oxidation-reduction reactions involve the transfer of electrons from one atom to another.
- When an atom or ion is reduced, its oxidation number is lowered. When an atom or ion is oxidized, its oxidation number is raised.
- In oxidation-reduction reactions involving molecular compounds (and polyatomic ions with covalent bonds), the more-electronegative atoms are treated as if they are reduced. The less-electronegative atoms are treated as if they are oxidized.

9. MAIN Idea **Explain** why oxidation and reduction must always occur together.
10. **Describe** the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?
11. **Write** the equation for the reaction of iron metal with hydrobromic acid to form iron(III) bromide and hydrogen gas. Determine the net change in oxidation for the element that is reduced and the element that is oxidized.
12. **Determine** the oxidation number of the boldface element in these compounds.
    a. HNO_3 c. Sb_2O_5
    b. CaN_2 d. $CuWO_4$
13. **Determine** the oxidation number of the boldface element in these ions.
    a. IO_4^- c. $B_4O_7^{2-}$
    b. MnO_4^- d. NH_2^-
14. **Make and Use Graphs** Alkali metals are strong reducing agents. Make a graph showing how the reducing abilities of the alkali metals would increase or decrease as you move down the family from sodium to francium.

Chemistry Online Self-Check Quiz glencoe.com

Objectives

- **Relate** changes in oxidation number to the transfer of electrons.
- **Use** changes in oxidation number to balance redox equations.
- **Balance** net ionic redox equations using the oxidation-number method.

Review Vocabulary

net ionic equation: an ionic equation that includes only the particles that participate in the reaction

New Vocabulary

oxidation-number method
species
half-reaction

Balancing Redox Equations

MAIN Idea Redox equations are balanced when the total increase in oxidation numbers equals the total decrease in oxidation numbers of the atoms involved in the reaction.

Real-World Reading Link When fatty substances in foods spoil, they are referred to as rancid. Large molecules are broken down through redox reactions that result in foul-smelling products. The equation for this process is complicated but can be balanced using the same rules for simpler equations.

The Oxidation-Number Method

Chemical equations must be balanced to show the correct quantities of reactants and products. Study the following unbalanced equation for the reaction that occurs when copper metal is placed in concentrated nitric acid. This reaction is shown in **Figure 19.6.** The brown gas that is produced is nitrogen dioxide (NO_2), from the reduction of nitrate ions (NO_3^-), and the blue solution is the result of the oxidation of copper (Cu) to copper(II) ions (Cu^{2+}).

$$Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$$

Note that oxygen appears in only one reactant, HNO_3, but in all three products. Nitrogen appears in HNO_3 and in two of the products. Redox equations such as this one, in which the same element appears in several reactants and products, can be difficult to balance. As you have read, when an atom loses electrons, its oxidation number increases; when an atom gains electrons, its oxidation number decreases. The number of electrons transferred from atoms must equal the number of electrons accepted by other atoms. Therefore, the total increase in oxidation numbers (oxidation) must equal the total decrease in oxidation numbers (reduction) of the atoms involved in the reaction. The balancing technique called the **oxidation-number method** is based on these principles, and is described in **Table 19.4.**

■ **Figure 19.6** Some chemical equations for redox reactions, such as the reaction between copper and nitric acid, can be difficult to balance because elements might appear more than once on each side of the equation.

Table 19.4 The Oxidation-Number Method

1. Assign oxidation numbers to all atoms in the equation.
2. Identify the atoms that are oxidized and the atoms that are reduced.
3. Determine the change in oxidation number for the atoms that are oxidized and for the atoms that are reduced.
4. Make the change in oxidation numbers equal in magnitude by adjusting coefficients in the equation.
5. If necessary, use the conventional method to balance the remainder of the equation.

EXAMPLE Problem 19.3

The Oxidation-Number Method Balance the following redox equation.

$$Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO_2 + H_2O$$

1 Analyze the Problem

Use the rules for determining oxidation number. The increase in oxidation number of the oxidized atoms must equal the decrease in oxidation number of the reduced atoms. Adjust the coefficients to balance the equation.

Chemistry Online
Personal Tutor For an online tutorial on balancing redox equations, visit glencoe.com.

2 Solve for the Unknown

Assign oxidation numbers to all atoms in the equation.

$$\overset{0}{Cu} + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3} \longrightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O_3})_2 + \overset{+4}{N}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O}$$

The oxidation number of copper increases from 0 to +2. The oxidation number of nitrogen decreases from +5 to +4.

Identify which atoms are oxidized, which are reduced, and which do not change.

Cu is oxidized.
N is reduced.
H does not change.
O does not change.
N does not change in the nitrate ion (NO_3^-).

Determine the change in oxidation number for the atoms that are oxidized and for the atoms that are reduced.

Change in oxidation number:
Oxidized: Cu +2
Reduced: N −1

Copper loses electrons. It is oxidized. Nitrogen gains electrons. It is reduced.

Make the change in oxidation numbers equal in magnitude by adjusting coefficients in the equation.

$$Cu + 2HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$
$$2(-1) = -2$$

Because the change in oxidation number for N is −1, you must add a coefficient of 2 to balance. This coefficient applies to both HNO_3 and NO_2.

Use the conventional method to balance the remainder of the equation.

$$Cu + 2HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$

$$Cu + 4HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$

The coefficient of HNO_3 must be increased from 2 to 4 to balance the four nitrogen atoms in the products.

$$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Add a coefficient of 2 to H_2O to balance the four hydrogen atoms on the left.

3 Evaluate the Answer

The number of atoms of each element is equal on both sides of the equation.
No subscripts have been changed.

PRACTICE Problems

Extra Practice Pages 989–990 and glencoe.com

Use the oxidation-number method to balance these redox equations.

15. $HCl + HNO_3 \longrightarrow HOCl + NO + H_2O$

16. $SnCl_4 + Fe \longrightarrow SnCl_2 + FeCl_3$

17. $NH_3(g) + NO_2(g) \longrightarrow N_2(g) + H_2O(l)$

18. Challenge $SO_2 + Br_2 + H_2O \longrightarrow HBr + H_2SO_4$

Balancing Net Ionic Redox Equations

Sometimes, chemists prefer to express redox reactions in the simplest possible terms—as an equation showing only the oxidation and reduction processes. Refer again to the balanced equation for the oxidation of copper by nitric acid.

$$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Note that the reaction takes place in aqueous solution, so HNO_3, which is a strong acid, will be ionized. Likewise, copper(II) nitrate ($Cu(NO_3)_2$) will be dissociated into ions. Therefore, the equation can also be written in ionic form.

$$Cu(s) + 4H^+(aq) + 4NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + 2NO_3^-(aq) + 2NO_2(g) + 2H_2O(l)$$

There are four nitrate ions among the reactants, but only two of them undergo change to form two nitrogen dioxide molecules. The other two nitrate ions are only spectator ions and can be eliminated from the equation. To simplify things when writing redox equations in ionic form, chemists usually indicate hydrogen ions by H^+ with the understanding that they exist in hydrated form as hydronium ions (H_3O^+). The equation can then be rewritten showing only the substances that undergo change.

$$Cu(s) + 4H^+(aq) + 2NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

Now look at the equation in unbalanced form.

$$Cu(s) + H^+(aq) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g) + H_2O(l)$$

You might also see this same reaction expressed in a way that shows only the substances that are oxidized and reduced.

$$Cu(s) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g) \text{ (in acid solution)}$$

In this case, the hydrogen ion and the water molecule are eliminated because neither is oxidized nor reduced. In acid solution, hydrogen ions (H^+) and water molecules are abundant and free to participate in redox reactions as either reactants or products. Some redox reactions can occur only in basic solution. When you balance equations for these reactions, you can add hydroxide ions (OH^-) and water molecules to either side of the equation.

DATA ANALYSIS LAB

Based on Real Data*

Analyze and Conclude

How does redox lift a space shuttle? The space shuttle gains nearly 72% of its lift from its solid rocket boosters (SRBs) during the first two minutes of launch. The two pencil-shaped SRB tanks are attached to both sides of the liquid hydrogen and oxygen fuel tank. Each SRB contains approximately 499,000 kg of propellent mixture.

Data and Observations

SRB Propellent Mixture

Component	Percent Composition
Ammonium perchlorate	69.6
Aluminum	16
Catalyst	0.4
Binder	12.04
Curing agent	1.96

*Data obtained from: Dumoulin, Jim. "Solid Rockt Boosters." *NSTS Shuttle Reference Manual.* 1988

Think Critically

1. **Balance an equation** Use the oxidation-number method to balance the chemical equation for the SRB reaction.
   $$NH_4ClO_4(s) + Al(s) \longrightarrow Al_2O_3(g) + HCl(g) + N_2(g) + H_2O(g)$$
2. **State** Which elements are reduced and which are oxidized?
3. **Infer** What are the benefits of using SRBs for the first two minutes of launch?
4. **Calculate** How many moles of water vapor are produced by one SRB?

EXAMPLE Problem 19.4

Balance a Net Ionic Redox Equation Balance the following redox equation.

$ClO_4^-(aq) + Br^-(aq) \longrightarrow Cl^-(aq) + Br_2(g)$ (in acid solution)

1 Analyze the Problem

Use the rules for determining oxidation number. The increase in oxidation number of the oxidized atoms must equal the decrease in oxidation number of the reduced atoms. The reaction takes place under acidic conditions. Adjust the coefficients to balance the equation.

2 Solve for the Unknown

Assign oxidation numbers to all atoms in the equation.

$$\overset{+7\,-2}{ClO_4^-}(aq) + \overset{-1}{Br^-}(aq) \longrightarrow \overset{-1}{Cl^-}(aq) + \overset{0}{Br_2}(g) \text{ (in acid solution)}$$

Use the rules in Table 19.2.

Identify which atoms are oxidized and which are reduced.

Br is oxidized.
Cl is reduced.

The oxidation number of bromine increases from −1 to 0. The oxidation number of chlorine decreases from +7 to −1.

Determine the change in oxidation number for the atoms that are oxidized and for the atoms that are reduced.

Change in oxidation number:

Br +1
Cl −8

Bromine loses electrons. It is oxidized.
Chlorine gains electrons. It is reduced.

Make the changes in oxidation number equal in magnitude by adjusting the coefficients in the equation.

$ClO_4^-(aq) + 8Br^-(aq) \longrightarrow Cl^-(aq) + 4Br_2(g)$ (in acid solution)

Because the oxidation number of Br is +1, you must add the coefficient 8 to balance the equation. $4Br_2$ represents 8 Br atoms to balance the $8Br^-$ on the left side.

Add enough hydrogen ions and water molecules to the equation to balance the oxygen atoms on both sides.

$ClO_4^-(aq) + 8Br^-(aq) + 8H^+(aq) \longrightarrow Cl^-(aq) + 4Br_2(g) + 4H_2O(l)$

Because you know the reaction takes place in acid solution, you can add H^+ ions on both sides of the equation.

3 Evaluate the Answer

The number of atoms of each element is equal on both sides of the equation. As with any ionic equation, the net charge on the right equals the net charge on the left. No subscripts have been changed.

PRACTICE Problems

Extra Practice Pages 989–990 and glencoe.com

Use the oxidation-number method to balance the following net ionic redox equations.

19. $H_2S(g) + NO_3^-(aq) \longrightarrow S(s) + NO(g)$ (in acid solution)

20. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)

21. $Zn + NO_3^- \longrightarrow Zn^{2+} + NO_2$ (in acid solution)

22. Challenge $I^-(aq) + MnO_4^-(aq) \longrightarrow I_2(s) + MnO_2(s)$ (in basic solution)

Connection to Biology What do many deep-sea fishes and fireflies have in common the bacterium, *Xenorhabdus luminescens*? These and other organisms emit light. Bioluminescence is the conversion of potential energy in chemical bonds into light during a redox reaction. Depending on the species, bioluminescence is produced by different chemicals and by different means. In fireflies, shown in **Figure 19.7,** light results from the oxidation of the molecule luciferin.

Scientists are still unraveling the mystery of bioluminescence. Some luminescent organisms emit light constantly, whereas others emit light when they are disturbed. Deep-sea fishes and some jellyfish appear to be able to control the light they emit, and one species of mushroom is known to emit light of two different colors. Zoologists have also determined that some light-emitting organisms do not produce light themselves; they produce light by harboring bioluminescent bacteria.

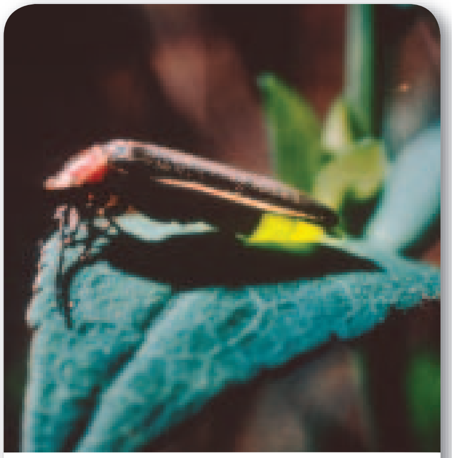


Figure 19.7 Organisms appear to use bioluminescence for different purposes. Some purposes might include attracting a mate and defense against prey. In the ocean depths, bioluminescence probably aids vision and recognition.

Balancing Redox Equations Using Half-Reactions

In chemistry, a **species** is any kind of chemical unit involved in a process. In the equilibrium equation $NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$, there are four species: the two molecules NH_3 and H_2O and the two ions NH_4^+ and OH^-. Oxidation-reduction reactions occur whenever a species that can give up electrons (reducing agent) comes in contact with another species that can accept them (oxidizing agent). For example, iron can reduce many species that are oxidizing agents, including chlorine.

$$2Fe + 3Cl_2 \longrightarrow 2FeCl_3$$

In this reaction, each iron atom is oxidized by losing three electrons to become an Fe^{3+} ion. At the same time, each chlorine atom in Cl_2 is reduced by accepting one electron to become a Cl^- ion.

$$\text{Oxidation: } Fe \longrightarrow Fe^{3+} + 3e^-$$
$$\text{Reduction: } Cl_2 + 2e^- \longrightarrow 2Cl^-$$

Equations such as these represent half-reactions. A **half-reaction** is one of the two parts of a redox reaction—the oxidation half or the reduction half. **Table 19.5** shows a variety of reduction half-reactions that involve the oxidation of Fe to Fe^{3+}.

Vocabulary

Science usage v. Common usage

Species

Science usage: in chemistry, any kind of representative particle involved in a process

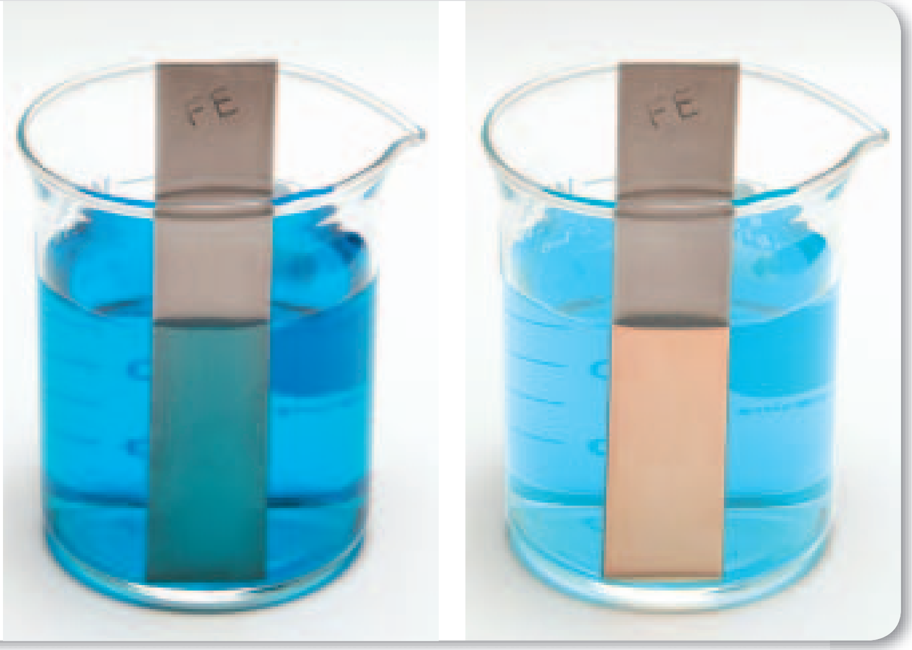
In a synthesis reaction, two distinct species combine to form a single product.

Common usage: a class of individuals having some common characteristics or qualities; a distinct sort or kind

Humans and chimpanzees are two different species.

Table 19.5 Redox Reactions that Oxidize Iron

Overall Reaction (unbalanced)	Oxidation Half-Reaction	Reduction Half-Reaction
$Fe + O_2 \longrightarrow Fe_2O_3$	$Fe \longrightarrow Fe^{3+} + 3e^-$	$O_2 + 4e^- \longrightarrow 2O^{2-}$
$Fe + F_2 \longrightarrow FeF_3$		$F_2 + 2e^- \longrightarrow 2F^-$
$Fe + HBr \longrightarrow H_2 + FeBr_3$		$2H^+ + 2e^- \longrightarrow H_2$
$Fe + AgNO_3 \longrightarrow Ag + Fe(NO_3)_3$		$Ag^+ + e^- \longrightarrow Ag$
$Fe + CuSO_4 \longrightarrow Cu + Fe_2(SO_4)_3$		$Cu^{2+} + 2e^- \longrightarrow Cu$



■ **Figure 19.8** As a result of this redox reaction between iron and copper sulfate solution, solid copper metal is deposited on the iron. To balance the chemical equation for this reaction, you could use half-reactions.

FOLDABLES Incorporate information from this section into your Foldable.

You will learn more about the importance of half-reactions when you study electrochemistry in Chapter 20. For now, however, you can learn to use half-reactions to balance a redox equation. For example, the following unbalanced equation represents the reaction that occurs when you put an iron nail into a solution of copper(II) sulfate, as shown in **Figure 19.8.**

$$Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + Fe_2(SO_4)_3(aq)$$

Iron atoms are oxidized as they lose electrons to the copper(II) ions. The steps for balancing redox equations by using half-reactions are shown in **Table 19.6.**

VOCABULARY

ACADEMIC VOCABULARY

Method:
a way of of doing something
Students study for an exam using different methods.

Table 19.6	**The Half-Reaction Method**
1. Write the net ionic equation for the reaction, omitting spectator ions. $Fe + Cu^{2+} + SO_4^{2-} \longrightarrow Cu + 2Fe^{3+} + 3SO_4^{2-}$; $Fe + Cu^{2+} \longrightarrow Cu + 2Fe^{3+}$	
2. Write the oxidation and reduction half-reactions for the net ionic equation. $Fe \longrightarrow 2Fe^{3+} + 6e^-$; $Cu^{2+} + 2e^- \longrightarrow Cu$	
3. Balance the atoms and charges in each half-reaction. $2Fe \longrightarrow 2Fe^{3+} + 6e^-$; $Cu^{2+} + 2e^- \longrightarrow Cu$	
4. Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction. $2Fe \longrightarrow 2Fe^{3+} + 6e^-$; $3Cu^{2+} + 6e^- \longrightarrow 3Cu$	
5. Add the balanced half-reactions and return spectator ions. $2Fe + 3Cu^{2+} \longrightarrow 3Cu + 2Fe^{3+}$; $2Fe(s) + 3CuSO_4(aq) \longrightarrow 3Cu(s) + Fe_2(SO_4)_3(aq)$	

EXAMPLE Problem 19.5

Balance a Redox Equation by Using Half-Reactions
Balance the redox equation for the reaction below using half-reactions.

$KMnO_4(aq) + SO_2(g) \longrightarrow MnSO_4(aq) + K_2SO_4(aq)$ (in acid solution)

1 Analyze the Problem

The reaction takes place in an acid solution. Use the rules for determining oxidation numbers and the steps for balancing by half-reactions to balance the equation for the reaction of permanganate and sulfur dioxide.

2 Solve for the Unknown

Write the net ionic equation for the reaction.

$MnO_4^- + SO_2 \longrightarrow Mn^{2+} + SO_4^{2-}$

Eliminate coefficients, spectator ions, and state symbols.

Write the oxidation and reduction half-reactions for the net ionic equation, including oxidation numbers.

$\overset{+4}{S}O_2 \longrightarrow \overset{+6}{S}O_4^{2-} + 2e^-$ (oxidation)

$\overset{+7}{Mn}O_4^- + 5e^- \longrightarrow \overset{+2}{Mn}^{2+}$ (reduction)

Use the rules in Table 19.2 and Table 19.6.

Balance the atoms and charges in the half-reactions.

$SO_2 + 2H_2O \longrightarrow SO_4^{2-} + 2e^- + 4H^+$ (oxidation)
$MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$ (reduction)

In an acid solution, H_2O molecules are available in abundance and can be used to balance oxygen atoms in the half-reactions; H^+ ions are readily available and can be used to balance the charge.

Adjust the coefficients so that the number of electrons lost in oxidation (2) equals the number of electrons gained in reduction (5).

$5SO_2 + 10H_2O \longrightarrow 5SO_4^{2-} + 20H^+ + 10e^-$ (oxidation)
$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$ (reduction)

The least common multiple of 2 and 5 is 10. Cross-multiplying gives the balanced oxidation and reduction half-reactions.

Add the balanced half-reactions and simplify by canceling or reducing like terms on both sides of the equation.

$5SO_2 + 10H_2O + 2MnO_4^- + 16H^+ + 10e^- \longrightarrow 5SO_4^{2-} + 20H^+ + 10e^- + 2Mn^{2+} + 8H_2O$

$5SO_2 + 2H_2O + 2MnO_4^- \longrightarrow 5SO_4^{2-} + 4H^+ + 2Mn^{2+}$

Return spectator ions (K^+), and restore the state descriptions.

$5SO_2(g) + 2H_2O(l) + 2KMnO_4(aq) \longrightarrow$
$K_2SO_4(aq) + 2H_2SO_4(aq) + 2MnSO_4(aq)$

Add the K^+ ions to the two MnO_4^- ions on the left and one of the SO_4^{2-} ions on the right. Split the remaining ions between the H^+ and Mn^+ ions.

3 Evaluate the Answer

A review of the balanced equation indicates that the number of atoms of each element is equal on both sides of the equation. No subscripts have been changed.

PRACTICE Problems

Extra Practice Pages 989–990 and glencoe.com

Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

23. $Cr_2O_7^-(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)

24. $Mn^{2+}(aq) + BiO_3^-(aq) \longrightarrow MnO_4^-(aq) + Bi^{2+}(aq)$ (in acid solution)

25. Challenge $N_2O(g) + ClO^-(aq) \longrightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution)

Problem-Solving Strategy

Balancing Redox Equations

Determine which species is oxidized, which species is reduced, which species is the oxidizing agent, and which species is the reducing agent.

Do the oxidized and reduced species appear more than once on either side of the equation, or does the reaction occurs in an acidic or basic solution?

No — **Oxidation-Number Method of Balancing Redox Equations**

1. Assign oxidation numbers to all of the elements.
2. Adjust the coefficients in the equation so that the oxidation numbers are equal in magnitude.
3. Balance the rest of the equation by the conventional method.

Yes — **Half-Reaction Method of Balancing Redox Equations**

1. Write the net ionic equation for the equation, omitting the spectator ions.
2. Determine the oxidation and the reduction half-reactions.
3. Balance the atoms and the charges in each half-reaction.
4. Adjust the coefficients so that the number of electrons lost and the number of electrons gained is equal.
5. Combine the balanced half-reactions and return spectator ions.

Apply the Strategy

Balance the following equation using this flowchart.

$P_4(s) \longrightarrow HPO_3^{2-}(aq) + PH_3(g)$ (in acid)

Section 19.2 Assessment

Section Summary

- Redox equations in which the same element appears in multiple reactants and products can be difficult to balance using the conventional method.
- The oxidation-number method is based on the number of electrons transferred from atoms equaling the number of electrons accepted by other atoms.
- To balance equations for reactions in an acid solution, add enough hydrogen ions and water molecules to balance the equation.
- To balance equations for reactions in a basic solution, add enough hydroxide ions and water molecules to balance the equation.
- A half-reaction is one of the two parts of a redox reaction.

26. MAIN Idea **Explain** how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction?

27. Describe why it is important to know the conditions under which an aqueous oxidation-reducation reaction takes place in order to balance the ionic equation for the reaction.

28. Explain the steps of the oxidation-number method of balancing equations.

29. State what an oxidation half-reaction shows. What does a reduction half-reaction show?

30. Write the oxidation and reduction half-reactions for the redox equation.

$Pb(s) + Pd(NO_3)_2(aq) \longrightarrow Pb(NO_3)_2(aq) + Pd(s)$

31. Determine The oxidation half-reaction of a redox reaction is $Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$, and the reduction half-reaction is $Au^{3+} + 3e^- \longrightarrow Au$. What minimum numbers of tin(II) ions and gold(III) ions would have to react in order to have zero electrons left over?

32. Apply Balance the following equations.

a. $HClO_3(aq) \longrightarrow ClO_2(g) + HClO_4(aq) + H_2O(l)$

b. $H_2SeO_3(aq) + HClO_3(aq) \longrightarrow H_2SeO_4(aq) + Cl_2(g) + H_2O(l)$

c. $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$ (in acid solution)

Chemistry Online Self-Check Quiz glencoe.com

In the Field

Career: Crime-Scene Investigator

Blood That Glows

In Shakespeare's play *MacBeth*, Lady MacBeth washes the blood of King Duncan from her hands but can still see the bloodstains. In modern forensics, a chemical called luminol gives investigators similar visual ability.

Blue-green whisper Luminol oxidizes when it comes in contact with iron, as shown in **Figure 1.** In the process, the molecules release energy in the form of distinctive blue-green light. In a dark room, the faint blue glow of luminol might reveal to investigators what their eyes alone could not see—hidden traces of blood. Red blood cells consist mainly of hemoglobin—a protein that contains iron.

To use luminol, investigators mix a white powder ($C_8H_7O_3N_3$) with hydrogen peroxide (H_2O_2) and other chemicals. This creates a liquid that can be sprayed onto areas suspected of holding hidden blood evidence. If blood is present—even in quantities to small to detect with the eye—the luminol will glow. Forensic photo-graphers then snap pictures with special cameras that can both capture the faint glow of the luminol and illuminate the surrounding area.

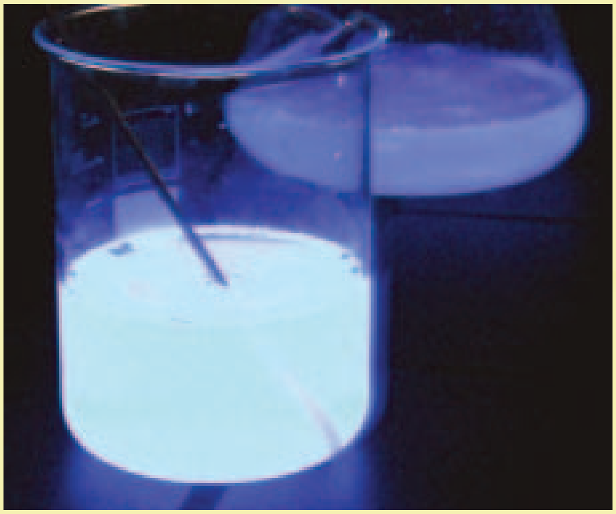


Figure 1 The luminol oxidizes within a beaker when an iron nail is added.

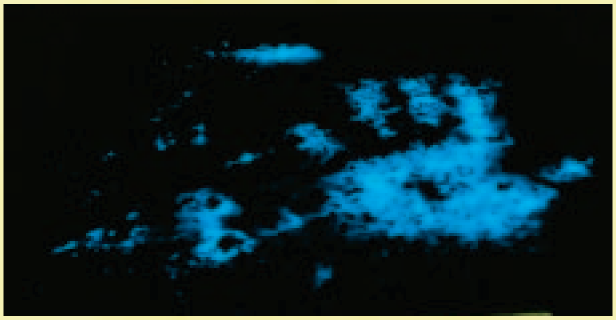


Figure 2 A luminol impression from a murder scene can be compared to a suspect's handprint.

Glowing evidence Bloodstains might reveal spatter patterns, giving clues about the type of weapon used to commit a crime. Faint luminol signals on carpet might lead investigators to much larger bloodstains. Bloody handprints, such as that in **Figure 2,** might even lead investigators to the assailant.

There are other uses for luminol besides murder investigations. In a car accident, luminol might reveal whether a victim was wearing a safety belt, even after the car has been subjected to rain, cold, or direct sunlight that can greatly alter bloodstains.

Spray of last resort Other iron-containing substances besides blood can cause luminol to glow, although experts can usually tell the difference. More importantly, luminol might interfere with other tests. For this reason, investigators normally do not use luminol until all their other investigations are complete.

WRITING in Chemistry

News Article Write a newspaper article that describes how luminol led investigators to a suspect. Describe the type of evidence that was used in the investigation. Visit glencoe.com to learn more about the use of luminol in crime-scene investigations.

CHEMLAB

SMALL SCALE

FORENSICS: IDENTIFY THE DAMAGING DUMPER

Background: Something is reacting with metals found on the hulls of many boats used on a nearby creek. The investigator has determined that there are three possible culprits, each with a different source. Your job is to test the three potential pollutants and compare them with a sample from the creek. The animals that rely on the creek as their primary water source are depending on you to solve this mystery of the damaging dumper.

Question: *How can a series of chemical reactions be used to determine what was dumped in a water supply?*

Materials

0.1*M* $AgNO_3$
0.1*M* HCl
0.1*M* $ZnSO_4$
unknown solution
Cu wire
Pb shot
Fe filings
Mg turnings
tongs or forceps
droppers (4)
24-well microscale reaction plate

Safety Precautions

WARNING: ***Silver nitrate*** ($AgNO_3$) ***is highly toxic and will stain skin and clothing.***

Procedure

1. Read and complete the lab safety form.
2. Create a table to record your data.
3. Place the well plate on a sheet of white paper.
4. Place a piece of copper wire in four wells in the first row.
5. Repeat Step 4, by adding a small sample of iron filings to wells in the second row.
6. Repeat Step 4, by adding a piece of lead shot to wells in the third row
7. Repeat Step 4, by adding a piece of magnesium ribbon to wells in the fourth row.
8. Count 20 drops of the silver nitrate solution ($AgNO_3$) into each well in the first column.
9. Repeat Step 8, adding hydrochloric acid (HCl) in the second column.
10. Repeat Step 8, adding zinc sulfate ($ZnSO_4$) in the third column.

Observations

	$AgNO_3$	HCl	$ZnSO_4$	Unknown
Cu				
Pb				
Fe				
Mg				

11. Repeat Step 8, adding the unknown solution in the fourth column.
12. Allow the reactions to proceed for 5 min, and then describe the reactions. Write *NR* for any wells that do not have evidence of a reaction.
13. **Cleanup and Disposal** Dispose of the solids and solutions as directed by your teacher. Wash and return all lab equipment to its designated location.

Analyze and Conclude

1. **Summarize** the results you observed in each well. How did you know a chemical reaction occurred?
2. **Model** Write a balanced reaction for each of the reactions you observed. In each one, identify the species being oxidized or reduced.
3. **Conclude** Based on your data, which solution was causing damage in the creek? Justify your answer.
4. **Use Variables, Constants, and Controls** Why was it important to compare the reactions of the unknown to more than one known solution?
5. **Research** Look up the MSDS for your chemical and report on what impact this chemical would have on the ecosystem.
6. **Extend** What would you expect if a solution of lead (II) nitrate ($Pb(NO_3)_2$) was one of the solutions?
7. **Error Analysis** Compare your results with those of other students in the laboratory. Explain any differences.

INQUIRY EXTENSION

Design an Experiment Hypothesize how you could remove this chemical from the creek without further damaging the ecology of the area. Design an experiment to test your hypothesis.

Study Guide

BIG Idea Oxidation-reduction reactions—among the most-common chemical processes in both nature and industry—involve the transfer of electrons.

Section 19.1 Oxidation and Reduction

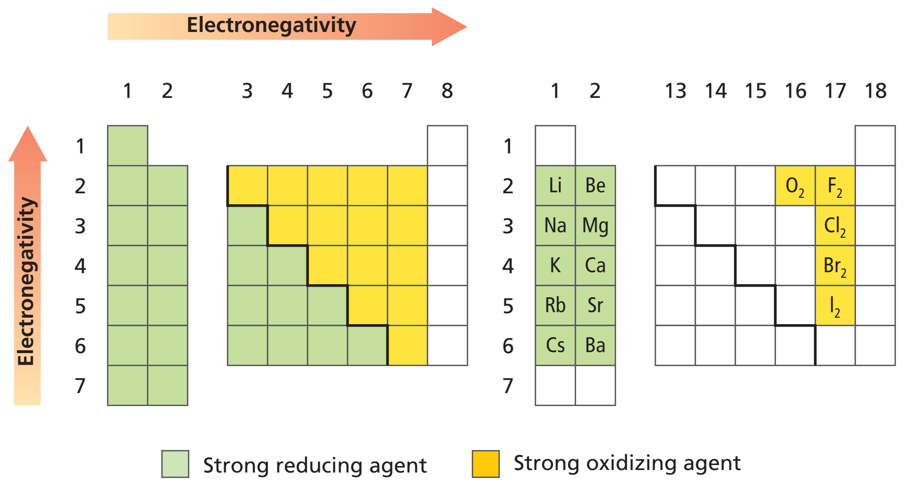
MAIN Idea Oxidation and reduction are complementary—as an atom is oxidized, another atom is reduced.

Vocabulary

- oxidation (p. 681)
- oxidation-reduction reaction (p. 680)
- oxidizing agent (p. 683)
- redox reaction (p. 680)
- reducing agent (p. 683)
- reduction (p. 681)

Key Concepts

- Oxidation-reduction reactions involve the transfer of electrons from one atom to another.
- When an atom or ion is reduced, its oxidation number is lowered. When an atom or ion is oxidized, its oxidation number is raised.
- In oxidation-reduction reactions involving molecular compounds (and polyatomic ions with covalent bonds), the more-electronegative atoms are treated as if they are reduced. The less-electronegative atoms are treated as if they are oxidized.



Section 19.2 Balancing Redox Equations

MAIN Idea Redox equations are balanced when the total increase in oxidation numbers equals the total decrease in oxidation numbers of the atoms involved in the reaction.

Vocabulary

- half-reaction (p. 693)
- oxidation-number method (p. 689)
- species (p. 693)

Key Concepts

- Redox equations in which the same element appears in several reactants and products can be difficult to balance using the conventional method.
- The oxidation-number method is based on the number of electrons transferred from atoms equaling the number of electrons accepted by other atoms.
- To balance equations for reactions in an acid solution, add enough hydrogen ions and water molecules to balance the equation.
- To balance equations for reactions in a basic solution, add enough hydroxide ions and water molecules to balance the equation.
- A half-reaction is one of the two parts of a redox reaction.

CHAPTER 19 Assessment

Section 19.1

Mastering Concepts

33. What is the main characteristic of oxidation-reduction reactions?

34. Explain why not all oxidation reactions involve oxygen.

35. In terms of electrons, what happens when an atom is oxidized? When an atom is reduced?

36. Define *oxidation number.*

37. **Metals** What is the oxidation number of alkaline earth metals in their compounds? Of alkali metals?

38. How does the oxidation number in an oxidation process relate to the number of electrons lost? How does the change in oxidation number in a reduction process relate to the number of electrons gained?

a b

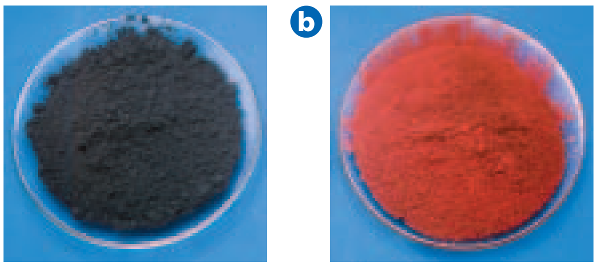


Figure 19.9

39. What is the oxidation number for copper in each of the compounds shown in **Figure 19.9?**

40. **Copper and air** Copper statues, such as the Statue of Liberty, begin to appear green after they have been exposed to air. In this redox process, copper metal reacts with oxygen to form solid copper oxide, which forms the green coating. Write the reaction for this redox process, and identify what is oxidized and what is reduced in the process.

Mastering Problems

41. Identify the species oxidized and the species reduced in each of these redox equations.
a. $3Br_2 + 2Ga \longrightarrow 2GaBr_3$
b. $HCl + Zn \longrightarrow ZnCl_2 + H_2$
c. $Mg + N_2 \longrightarrow Mg_3N_2$

42. Identify the oxidizing agent and the reducing agent in each of these redox equations.
a. $N_2 + 3H_2 \longrightarrow 2NH_3$
b. $2Na + I_2 \longrightarrow 2NaI$

43. What is the reducing agent in this balanced equation?

$$8H^+ + Sn + 6Cl^- + 4NO_3^{\,-1} \longrightarrow SnCl_6^{\,-2} + 4NO_2 + 4H_2O$$

44. What is the oxidation number of manganese in $KMnO_4$?

45. Determine the oxidation number of the boldface element in these substances and ions.
a. $Ca\mathbf{Cr}O_4$ **c.** $\mathbf{N}O_2^{\,-}$
b. $NaH\mathbf{S}O_4$ **d.** $\mathbf{Br}O_3^{\,-}$

46. Identify each of these half-reactions as either oxidation or reduction.
a. $Al \longrightarrow Al^{3+} + 3e^-$
b. $Cu^{2+} + e^- \longrightarrow Cu^+$

47. Which of these equations does not represent a redox reaction? Explain your answer.
a. $LiOH + HNO_3 \longrightarrow LiNO_3 + H_2O$
b. $MgI_2 + Br_2 \longrightarrow MgBr_2 + I_2$

48. Determine the oxidation number of nitrogen in each of these molecules or ions.
a. NO_3 **b.** N_2O **c.** NF_3

49. Determine the oxidation number of each element in these compounds or ions.
a. $Au_2(SeO_4)_3$ (gold (III) selenate)
b. $Ni(CN)_2$ (nickel (II) cyanide)

SO_3

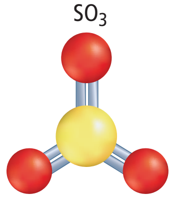


Figure 19.10

50. Explain how the sulfite ion ($SO_3^{\,2-}$) differs from sulfur trioxide (SO_3), shown in **Figure 19.10.**

Section 19.2

Mastering Concepts

51. Compare and contrast balancing redox equations in acidic and basic solutions.

52. Explain why writing hydrogen ions as H^+ in redox reactions represents a simplification and not how they exist.

53. Before you attempt to balance the equation for a redox reaction, why do you need to know whether the reaction takes place in acidic or basic solution?

54. Explain what a spectator ion is.

55. Define the term *species* in terms of redox reactions.

56. Is the following equation balanced? Explain.

$$Fe(s) + Ag^+(aq) \longrightarrow Fe^{2+}(aq) + Ag(s)$$

57. Does the following equation represent a reduction or an oxidation process? Explain your answer.

$$Zn^{2+} + 2e^- \longrightarrow Zn$$

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58. Describe what is happening to electrons in each half reaction of a redox process.

Mastering Problems

59. Use the oxidation-number method to balance these redox equations.

a. $Cl_2 + NaOH \longrightarrow NaCl + HOCl$

b. $HBrO_3 \longrightarrow Br_2 + H_2O + O_2$

60. Balance these net ionic equations for redox reactions.

a. $Au^{3+}(aq) + I^-(aq) \longrightarrow Au(s) + I_2(s)$

b. $Ce^{4+}(aq) + Sn^{2+}(aq) \longrightarrow Ce^{3+}(aq) + Sn^{4+}(aq)$

61. Use the oxidation-number method to balance the following ionic redox equations.

a. $Al + I_2 \longrightarrow Al^{3+} + I^-$

b. $MnO_2 + Br^- \longrightarrow Mn^{2+} + Br_2$ (in acid solution)

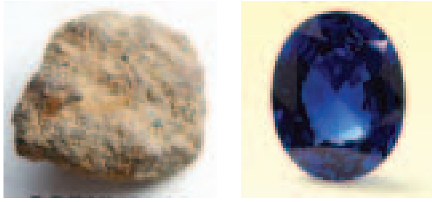
62. Use the oxidation-number method to balance these redox equations.

a. $PbS + O_2 \longrightarrow PbO + SO_2$

b. $NaWO_3 + NaOH + O_2 \longrightarrow Na_2WO_4 + H_2O$

c. $NH_3 + CuO \longrightarrow Cu + N_2 + H_2O$

d. $Al_2O_3 + C + Cl_2 \longrightarrow AlCl_3 + CO$



■ **Figure 19.11**

63. Sapphire The mineral corundum is comprised of aluminum oxide (Al_2O_3) and is colorless. Sapphire is mostly aluminum oxide, but it contains small amounts of Fe^{2+} and Ti^{4+}. The color of sapphire results from an electron transfer from Fe^{2+} to Ti^{4+}. Based on **Figure 19.11,** draw the reaction that occurs resulting in the mineral on the right. What are the oxidizing and reducing agents?

64. Write the oxidation and reduction half-reactions represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$

b. $I_2(s) + Na_2S_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$

c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

65. Write the two half-reactions that make up the following balanced redox reaction.

$$3H_2C_2O_4 + 2HAsO_2 \longrightarrow 6CO_2 + 2As + 4H_2O$$

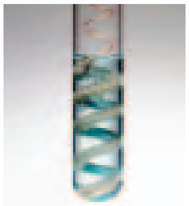
66. Label each half-reaction as reduction or oxidation.

a. $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

b. $MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$

c. $2H^+ + 2e^- \longrightarrow H_2$

d. $F_2 \longrightarrow 2F^- + 2e^-$



■ **Figure 19.12**

67. Copper When solid copper pieces are put into a solution of silver nitrate, as shown in **Figure 19.12,** silver metal appears and blue copper(II) nitrate forms. Write the corresponding chemical equation without balancing it. Next, determine the oxidation state of each element in the equation. Write the two half-reactions, labeling which is oxidation and which is reduction. Finally, write a balanced equation for the reaction.

68. Use the oxidation-number method to balance these ionic redox equations.

a. $MoCl_5 + S^{2-} \longrightarrow MoS_2 + Cl^- + S$

b. $TiCl_6^{2-} + Zn \longrightarrow Ti^{3+} + Cl^- + Zn^{2+}$

69. Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed.

a. $NH_3(g) + NO_2(g) \longrightarrow N_2(g) + H_2O(l)$

b. $Br_2 \longrightarrow Br^- + BrO_3^-$ (in basic solution)

70. Balance the following redox chemical equation. Rewrite the equation in full ionic form, then derive the net ionic equation and balance by the half-reaction method. Give the final answer as it is shown below but with the balancing coefficients.

$$KMnO_4(aq) + FeSO_4(aq) + H_2SO_4(aq) \longrightarrow Fe_2(SO_4)_3(aq) + MnSO_4(aq) + K_2SO_4(aq) + H_2O(l)$$

71. Write the oxidation and reduction half-reaction represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$

b. $I_2(s) + NaS_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$

c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

72. Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed. Keep balanced equations in net ionic form.

a. $Cl^-(aq) + NO_3^-(aq) \longrightarrow ClO^-(aq) + NO(g)$ (in acid solution)

b. $IO_3^-(aq) + Br^-(aq) \longrightarrow Br_2(l) + IBr(s)$ (in acid solution)

c. $I_2(s) + Na_2S_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$ (in acid solution)

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Chapter 19 Assessment

Mixed Review

73. Determine the oxidation number of the boldface element in each of the following.
a. $\mathbf{O}F_2$ **b.** $\mathbf{U}O_2^{2+}$ **c.** $\mathbf{Ru}O_4$ **d.** $\mathbf{Fe}_2O_3$

74. Identify each of the following changes as either oxidation or reduction.
a. $2Cl^- \longrightarrow Cl_2 + 2e^-$
b. $Na \longrightarrow Na^+ + e^-$
c. $Ca^{-2} + 2e^- \longrightarrow 2Ca$
d. $O_2 + 4e^- \longrightarrow 2O^{2-}$

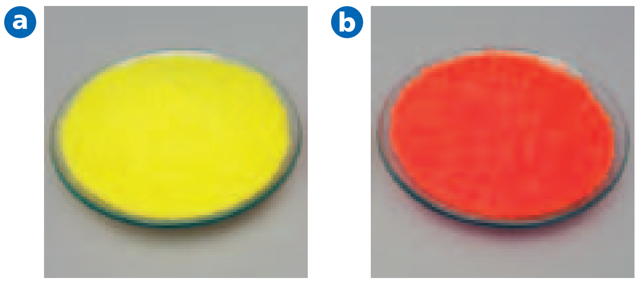
75. Use the rules for assigning oxidation numbers to complete **Table 19.7.**

Table 19.7 Oxidation Number Assignment

Element	Oxidation Number	Rule
K in KBr	+1	
Br in KBr		8
Cl in Cl_2		1
K in KCl		7
Cl in KCl	–1	
Br in Br_2	0	

76. Identify the reducing agents in these equations.
a. $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$
b. $Na_2SO_4 + 4C \longrightarrow Na_2S + 4CO$
c. $4IrF_5 + Ir \longrightarrow 5IrF4$

77. Write a balanced ionic redox equation using the following pairs of redox half-reactions.
a. $Fe \longrightarrow Fe^{2+} + 2e^-$
$Te^{2+} + 2e^- \longrightarrow Te$
b. $IO_4^- + 2e^- \longrightarrow IO_3^-$
$Al \longrightarrow Al^{3+} + 3e^-$ (in acid solution)
c. $I_2 + 2e^- \longrightarrow 2I^-$
$N_2O \longrightarrow NO_3^- + 4e^-$ (in acid solution)



■ **Figure 19.13**

78. What is the oxidation number of chromium in each of the compounds shown in **Figure 19.13?**

79. Balance these ionic redox equations by any method.
a. $Sb^{3+} + MnO_4^- \longrightarrow SbO_4^{3-} + Mn^{2+}$ (in acid solution)
b. $N_2O + ClO^- \longrightarrow Cl^- + NO_2^-$ (in basic solution)

80. Gemstones Rubies are gemstones made up mainly of aluminum oxide. Their red color comes from a small amount of chromium(III) ions replacing some of the aluminum ions. Draw the structure of aluminum oxide, and show the reaction in which an aluminum ion is replaced with a chromium ion. Is this a redox reaction?

81. Balance these ionic redox equations by any method.
a. $Mg + Fe^{3+} \longrightarrow Mg^{2+} + Fe$
b. $ClO_3^- + SO_2 \longrightarrow Cl^- + SO_4^{2-}$ (in acid solution)

82. Balance these redox equations by any method.
a. $P + H_2O + HNO_3 \longrightarrow H_3PO_4 + NO$
b. $KClO_3 + HCl \longrightarrow Cl_2 + ClO_2 + H_2O + KCl$

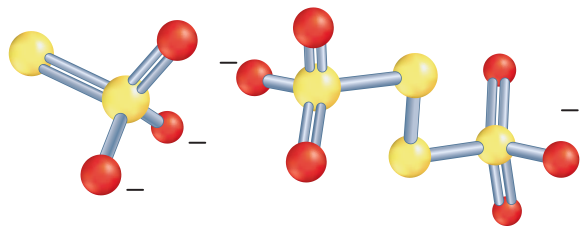
Think Critically

83. Apply The following equations show redox reactions that are sometimes used in the laboratory to generate pure nitrogen gas and pure dinitrogen monoxide gas (nitrous oxide, N_2O).

$$NH_4NO_2(s) \longrightarrow N_2(g) + 2H_2O(l)$$
$$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(l)$$

a. Determine the oxidation number of each element in the two equations, and then make diagrams showing the changes in oxidation numbers that occur in each reaction.
b. Identify the atom that is oxidized and the atom that is reduced in each of the two reactions.
c. Identify the oxidizing and reducing agents in each of the two reactions.
d. Write a sentence telling how the electron transfer taking place in these two reactions differs from that taking place here.

$$2AgNO_3 + Zn \longrightarrow Zn(NO_3)_2 + 2Ag$$



Thiosulfate ion ($S_2O_3^{2-}$) Tetrathionate ion ($S_4O_6^{2-}$)

■ **Figure 19.14**

84. Analyze Examine the net ionic equation below for the reaction that occurs when the thiosulfate ion ($S_2O_3^{2-}$) is oxidized to the tetrathionate ion ($S_4O_6^{2-}$). Balance the equation using the half-reaction method. **Figure 19.14** will help you to determine the oxidation numbers to use.

$$S_2O_3^{2-} + I_2 \longrightarrow I^- + S_4O_6^{2-} \text{ (in acid solution)}$$

85. Predict Consider the fact that all of the following are stable compounds. What can you infer about the oxidation state of phosphorus in its compounds?

$$PH_3,\ PCl_3,\ P_2H_4,\ PCl_5,\ H_3PO_4,\ Na_3PO_3$$

86. Solve Potassium permanganate oxidizes chloride ions to chlorine gas. Balance the equation for this redox reaction taking place in acid solution.

87. In the half-reaction $NO_3^- \longrightarrow NH_4^+$, on which side of the equation should electrons be added? Add the correct number of electrons to the side on which they are needed, and rewrite the equation.

■ **Figure 19.15**

88. The redox reaction between dichromate ion and iodide ion in acid solution is shown in **Figure 19.15.** Use the half-reaction method to balance the equation for this redox reaction.

Challenge Problem

89. For each reaction described, write the corresponding chemical equation without putting coefficients to balance it. Next, determine the oxidation state of each element in the equation. Then, write the two half-reactions, labeling which is oxidation and which is reduction. Finally, write a balanced equation for the reaction.

a. Solid mercuric oxide is put into a test tube and gently heated. Liquid mercury forms on the sides and in the bottom of the tube, and oxygen gas bubbles out from the test tube.

b. Solid copper pieces are put into a solution of silver nitrate. Silver metal appears and blue copper(II) nitrate forms in the solution.

Cumulative Review

90. A gaseous sample occupies 32.4 mL at −23°C and 0.75 atm. What volume will it occupy at STP? *(Chapter 13)*

91. When iron(III) chloride ($FeCl_3$) reacts in an atmosphere of pure oxygen, the following occurs:

$$4FeCl_3(s) + 3O_2(gv) \longrightarrow 2Fe_2O_3(s) + 6Cl_2(g)$$

If 45.0 g of $FeCl_3$ reacts and 20.5 g of iron(III) oxide is recovered, determine the percent yield. *(Chapter 11)*

Additional Assessment

WRITING in Chemistry

92. Steel Research the role of oxidation-reduction reactions in the manufacture of steel. Write a summary of your findings, including appropriate diagrams and equations representing the reactions.

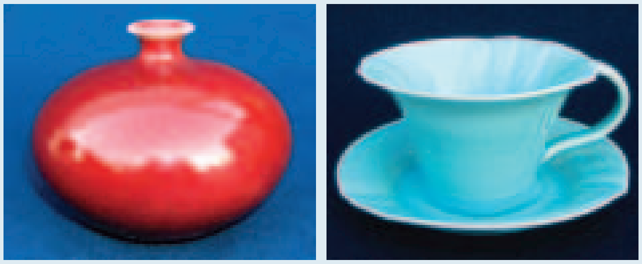
93. Silverware Practice your technical writing skills by writing a procedure for cleaning tarnished silverware by a redox chemical process. Be sure to include background information describing the process as well as logical steps that would enable anyone to accomplish the task.

94. Copper was a useful metal even before iron, silver, and gold metals were extracted and used from their ores and used as tools, utensils, jewelry, and artwork. Copper was smelted by heating copper ores with charcoal to high temperatures as early as 8000 years ago. Thousands of pieces of scrap copper have been unearthed in Virginia, where in the 1600s the colonists might have traded this material for food. Compare and contrast the processing and use of copper in those older civilizations with today.

DBQ Document-Based Questions

Glazes *The formation of color in ceramic glazes, such as in* **Figure 19.16,** *can be influenced by firing conditions. Metal ions such as copper that have more than one oxidation state can impart different colors to a glaze. In an oxidative firing, plenty of oxygen is allowed in the kiln, and copper ions present will make the glaze a green-to-blue color. Under reducing conditions, oxygen is limited and carbon dioxide is abundant. Copper ions in the glaze provide a reddish color.*

Data obtained from: Denio, Allen A. 2001. The joy of color in ceramic glazes with the help of redox chemistry. *Journal of Chemical Education.* 78 No 10.



■ **Figure 19.16**

95. Write the equation for what has occurred in the pottery shown in **Figure 19.16.**

96. Based on the color of the pottery, what is the oxidation state of the copper that is reduced? Oxidized?

Cumulative

Standardized Test Practice

Multiple Choice

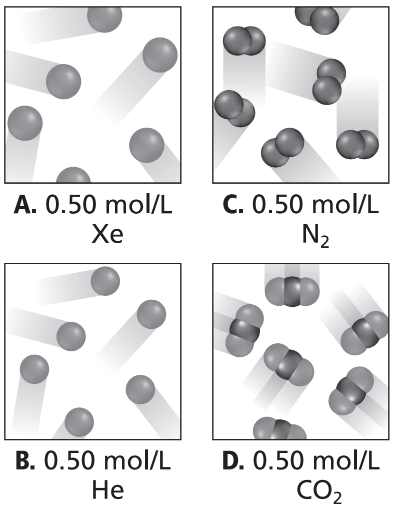
1. Which is NOT a reducing agent in a redox reaction?
A. the substance oxidized
B. the electron acceptor
C. the less-electronegative substance
D. the electron donor

2. The reaction between nickel and copper(II) chloride is shown below.

$$Ni(s) + CuCl_2(aq) \longrightarrow Cu(s) + NiCl_2(aq)$$

What are the half-reactions for this redox reaction?
A. $Ni \longrightarrow Ni^{2+} + 2e^-$, $Cl_2 \longrightarrow 2Cl^- + 2e^-$
B. $Ni \longrightarrow Ni^{2+} + e^-$, $Cu^+ + e^- \longrightarrow Cu$
C. $Ni \longrightarrow Ni^{2+} + 2e^-$, $Cu^{2+} + 2e^- \longrightarrow Cu$
D. $Ni \longrightarrow Ni^{2+} + 2e^-$, $2Cu^+ + 2e^- \longrightarrow Cu$

Use the diagram below to answer Questions 3 and 4. All four containers have a volume of 5.0 L and are at the same temperature.



3. Which container contains 110 g of its gas?
A. A
B. B
C. C
D. D

4. If a small hole is made in each container so that the gas can escape, which container will have the fastest rate of effusion?
A. A
B. B
C. C
D. D

5. The following system is in equilibrium:

$$2S(s) + 5F_2(g) \rightleftharpoons SF_4(g) + SF_6(g)$$

Which will cause the equilibrium to shift to the right?
A. increased concentration of SF_4
B. increased concentration of SF_6
C. increased pressure on the system
D. decreased pressure on the system

Use the table below to answer Question 6.

Data for the Formation of Cobalt(II) Sulfate at 25°C	
$Co(s) + S(s) + 2O_2(g) \longrightarrow CoSO_4(s)$	
ΔH_f°	−888.3 kJ/mol
ΔS_f°	118.0 J/mol • K
ΔG_f°	?

6. What is the ΔG_f° for the formation of cobalt(II) sulfate from its elements?
A. −853.1 kJ/mol
B. −885.4 kJ/mol
C. −891.3 kJ/mol
D. −923.5 kJ/mol

7. Which will be the result of increasing the temperature of a reaction in a system in equilibrium where the forward reaction is endothermic?
A. The equilibrium will shift to the left.
B. The equilibrium will shift to the right.
C. The rate of the forward reaction will be decreased.
D. The rate of the reverse reaction will be decreased.

8. The reaction between sodium iodide and chlorine is shown below.

$$2NaI(aq) + Cl_2(aq) \longrightarrow 2NaCl(aq) + I_2(aq)$$

The oxidation state of sodium remains unchanged for which reason?
A. Na^+ is a spectator ion.
B. Na^+ cannot be reduced.
C. Na is an uncombined element.
D. Na^+ is a monatomic ion.

Short Answer

Use the equation below to answer Questions 9 and 10.

The net ionic reaction between iodine and lead(IV) oxide is shown below.

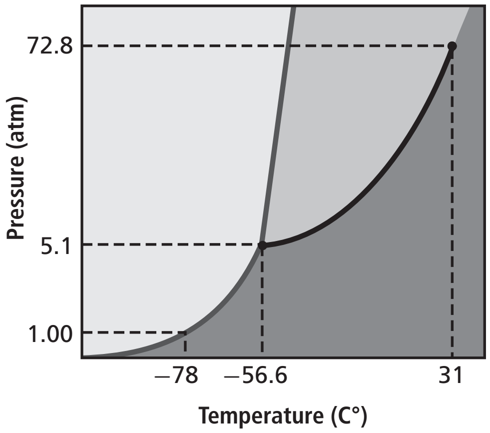
$$I_2(s) + PbO_2(s) \longrightarrow IO_3^-(aq) + Pb^{2+}(aq)$$

9. Identify the oxidation number in each participant in the reaction.

10. Explain how to identify which element is oxidized and which one is reduced.

Extended Response

Use the diagram below to answer Questions 11 to 13.



11. Explain what state or states of matter can exist at a temperature of −56.6°C and a pressure of 31.1 atm.

12. Suppose that you have a sample of CO_2 at 35°C and 83 atm. In what state of matter is the sample? Explain how you can predict this from the graph.

13. Is carbon dioxide denser in its liquid state or its solid state? Use the graph to explain.

SAT Subject Test: Chemistry

14. Which statement about the common ion effect is NOT true?
    - **A.** The effects of common ions on an equilibrium system can be explained by Le Châtelier's principle.
    - **B.** The decreased solubility of an ionic compound due to the presence of a common ion is called the common ion effect.
    - **C.** The addition of NaCl to a saturated solution of AgCl will produce the common ion effect.
    - **D.** The common ion effect is due to a shift in equilibrium toward the aqueous products of a system.
    - **E.** The addition of lead nitrate ($Pb(NO_3)$) to a saturated solution of lead chromate ($PbCrO_4$) will produce the common ion effect.

Use the list below to answer Questions 15 to 18.

Five flasks contain 500 mL of a 0.250*M* aqueous solution of the indicated chemical.

- **A.** KCl
- **B.** CH_3OH
- **C.** $Ba(OH)_2$
- **D.** CH_3COOH
- **E.** NaOH

15. Which chemical will dissociate into the greatest number of particles when in solution?

16. Which chemical has the greatest molar mass?

17. Which flask would contain 9.32 g of the labeled chemical?

18. Which flask's contents are composed of 18.6% oxygen?

NEED EXTRA HELP?																		
If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Review Section . . .	19.1	19.3	13.3	13.1	17.2	15.5	17.2	19.1	19.1	19.1	12.4	12.4	12.4	17.2	14.2	10.2	10.3	10.4