THE MEANING OF OXIDATION AND REDUCTION

section 22.1

objectives

- Define oxidation and reduction in terms of the loss or gain of oxygen or hydrogen and the loss or gain of electrons
- State the characteristics of a redox reaction, and identify the oxidizing agent and reducing agent

key terms

- oxidation
- reduction
- oxidation–reduction reactions
- redox reactions
- reducing agent
- oxidizing agent

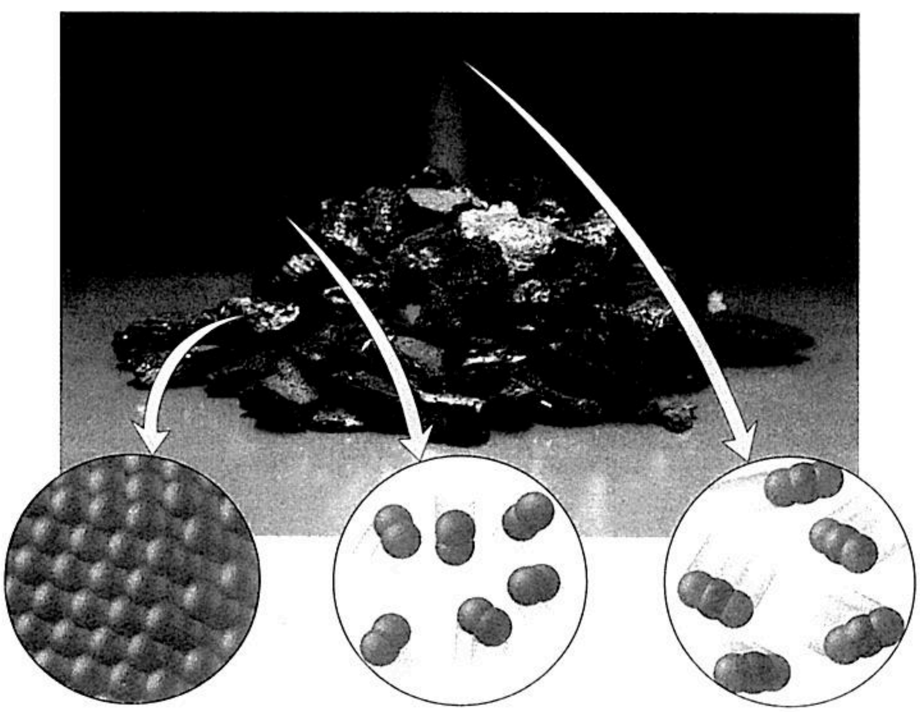
During winter in cold climates, salt is often used on roads to prevent the buildup of slippery ice. Although salt may make driving safer, it can cause the metallic parts of cars to corrode or rust relatively quickly. This problem is often so severe that people will not drive their newer cars in winter because they fear that their car will end up looking as rusty as this ship. **What property causes metal to corrode?**

Oxygen in Redox Reactions

The combustion of gasoline in an automobile engine and the burning of wood in a fireplace produce energy. So does the metabolism of food by your body. Such reactions are among the principal sources of energy on Earth, and all involve a process called oxidation. What do you think of when you hear that term? If you answered "oxygen," you are thinking along the same lines as most early chemists.

Oxidation originally meant the combination of an element with oxygen to produce oxides. As you will soon learn, the term also has a more modern, and wider, meaning. When gasoline or wood burns in air, it oxidizes and produces carbon dioxide. So does coal, as shown in **Figure 22.1.** Methane (CH_4), a component of natural gas, also burns in air. Methane oxidizes to form oxides of carbon and hydrogen.

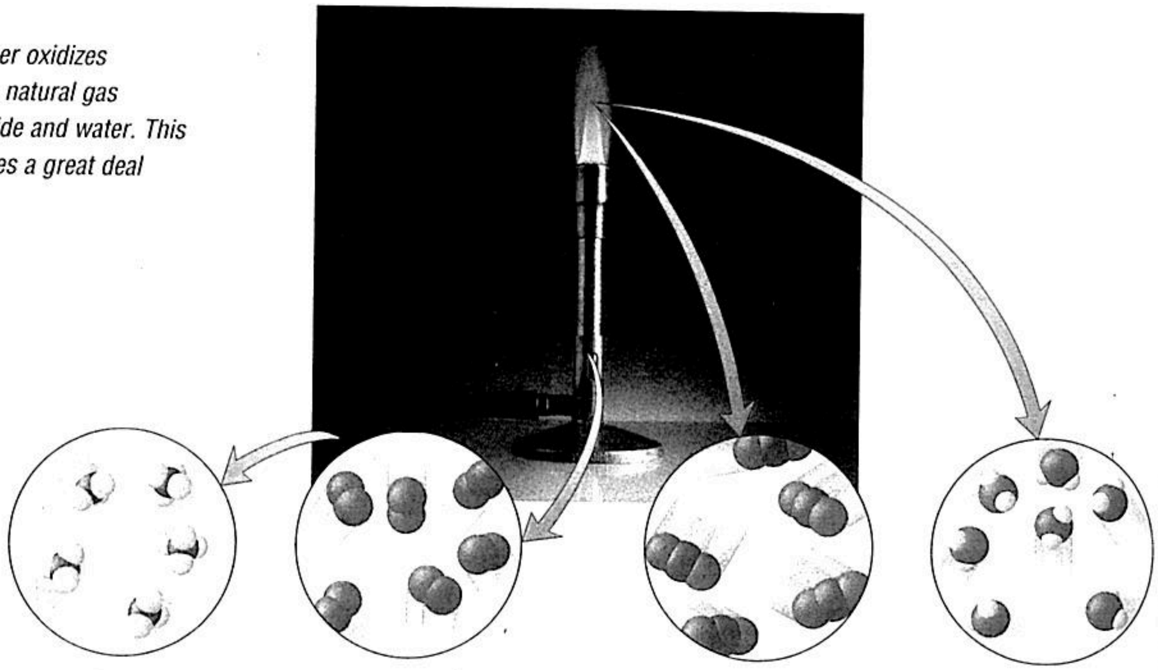
Not all oxidation processes involve burning, however. Bleaching is an example of oxidation that does not involve burning. Bleaches are substances that remove stains or unwanted color from fabrics and other



$C(s) \quad + \quad O_2(g) \longrightarrow CO_2(g)$

Figure 22.1
When coal, which is mostly carbon, is burned in air, carbon dioxide and heat are produced.

Figure 22.2
A Bunsen burner oxidizes the methane in natural gas to carbon dioxide and water. This reaction releases a great deal of heat.



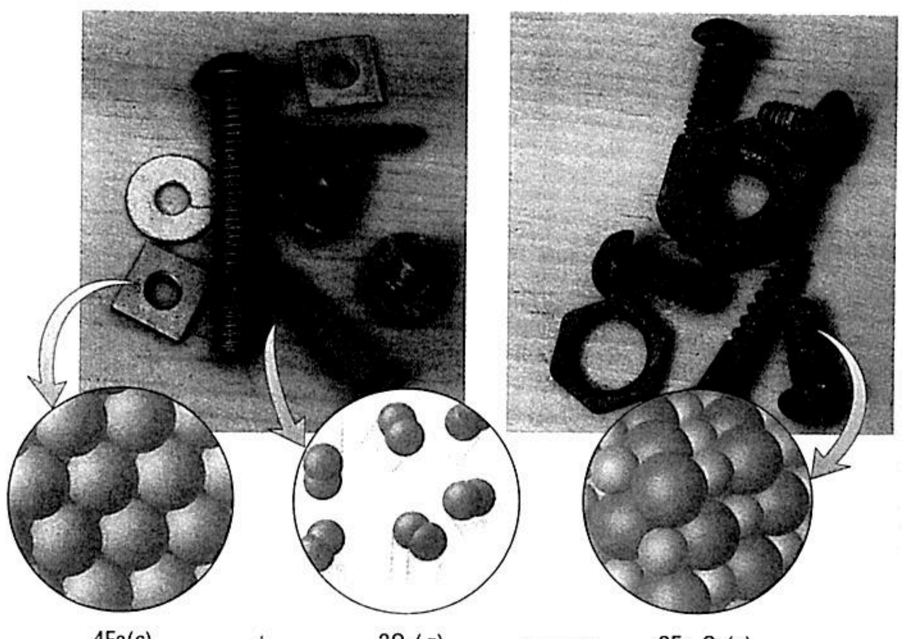
$CH_4(g)$ + $2O_2(g)$ $\longrightarrow$ $CO_2(g)$ + $2H_2O(g)$

materials. Common liquid household bleaches contain sodium hypochlorite (NaClO). Powder bleaches may contain calcium hypochlorite ($Ca(ClO)_2$) or sodium perborate ($NaBO_3$).

Hydrogen peroxide (H_2O_2) is another good oxidizing agent. Common household peroxide is both a bleach and a mild antiseptic that kills bacteria by oxidizing them. Notice that, in keeping with the original definition of oxidation, all these substances have one thing in common: They all contain oxygen.

Another familiar example of an oxidation process that does not involve burning is rusting. When elemental iron turns to rust, it slowly oxidizes to compounds such as iron(III) oxide (Fe_2O_3).

Figure 22.3
When items made of iron are exposed to moist air, the Fe atoms react with O_2 molecules. The iron rusts; it is oxidized to compounds such as iron(III) oxide (Fe_2O_3).



$4Fe(s)$ + $3O_2(g)$ $\longrightarrow$ $2Fe_2O_3(s)$

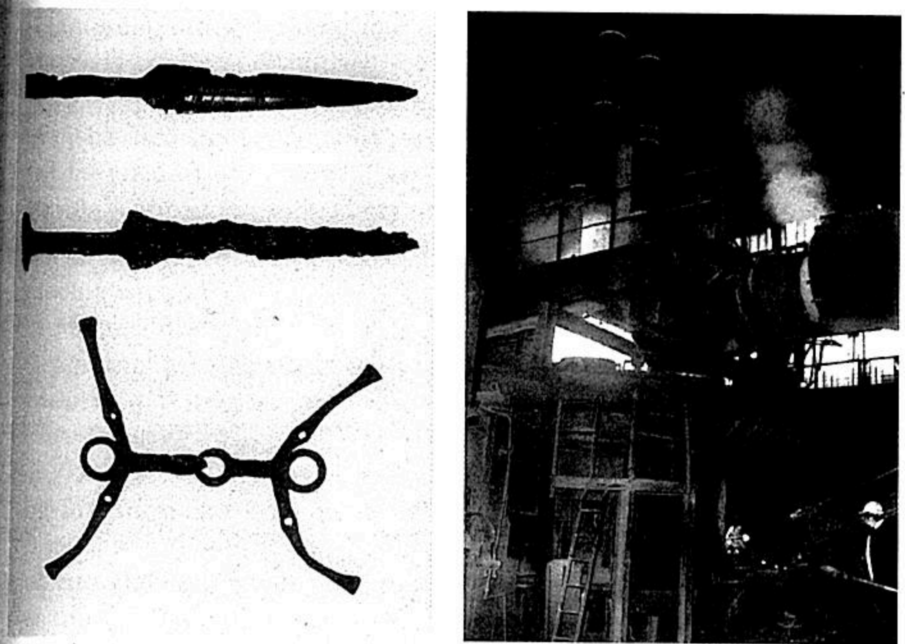


Figure 22.4
These iron objects (left) were made in ancient times. The iron was obtained by reduction of iron ore with charcoal. A similar process is carried out in a modern-day blast furnace (right).

A process called reduction is the opposite of oxidation. Originally, **reduction** meant the loss of oxygen from a compound. The reduction of iron ore to metallic iron involves the removal of oxygen from iron(III) oxide. The reduction is accomplished by heating the ore with charcoal. A large decrease in volume occurs during the reduction of a metal oxide to metal. Can you explain how reduction got its name? The equation for the reduction of iron ore is

$$\underset{\text{Iron(III) oxide}}{2Fe_2O_3(s)} + \underset{\text{Carbon}}{3C(s)} \longrightarrow \underset{\text{Iron}}{4Fe(s)} + \underset{\text{Carbon dioxide}}{3CO_2(g)}$$

The artifacts in **Figure 22.4** show that ancient people reduced iron ore to iron in the early Iron Age, more than 2500 years ago!

The equation for the reduction of iron also includes an oxidation process. Oxidation and reduction always occur simultaneously. As iron oxide is reduced to iron by losing oxygen, carbon is oxidized to carbon dioxide by gaining oxygen. No oxidation occurs without reduction, and no reduction occurs without oxidation. Reactions that involve these processes are therefore called **oxidation–reduction reactions.** Oxidation–reduction reactions are also known as **redox reactions.**

Electron Transfer in Redox Reactions

Today, the concepts of oxidation and reduction have been extended to include many reactions that do not even involve oxygen. Redox reactions are understood to involve a shift of electrons between reactants. **Oxidation** is redefined to mean complete or partial loss of electrons or gain of oxygen. **Reduction** is complete or partial gain of electrons or loss of oxygen.

Oxidation	**Reduction**
Loss of electrons	Gain of electrons
Gain of oxygen	Loss of oxygen

"*LEO* the lion goes *GER*" may help you remember the definitions of oxidation and reduction. *LEO* stands for *L*osing *E*lectrons is *O*xidation; *GER* stands for *G*aining *E*lectrons is *R*eduction.

Some examples will illustrate these redefined concepts. Consider reactions between a metal and a nonmetal. Electrons are transferred from atoms of the metal to atoms of the nonmetal. For example, the ionic compound magnesium sulfide is produced when magnesium metal is heated with the nonmetal sulfur. See **Figure 22.5.**

$$\underset{\text{Magnesium atom}}{\mathrm{Mg}:} + \underset{\text{Sulfur atom}}{\cdot\ddot{\mathrm{S}}:} \longrightarrow \underset{\text{Magnesium ion}}{\mathrm{Mg}^{2+}} + \underset{\text{Sulfide ion}}{:\ddot{\mathrm{S}}:^{2-}}$$

The result of this reaction is the transfer of two electrons from a magnesium atom to a sulfur atom. Because it loses electrons, the magnesium atom is said to be oxidized to a magnesium ion. How many electrons does the magnesium atom lose? Simultaneously, the sulfur atom gains two

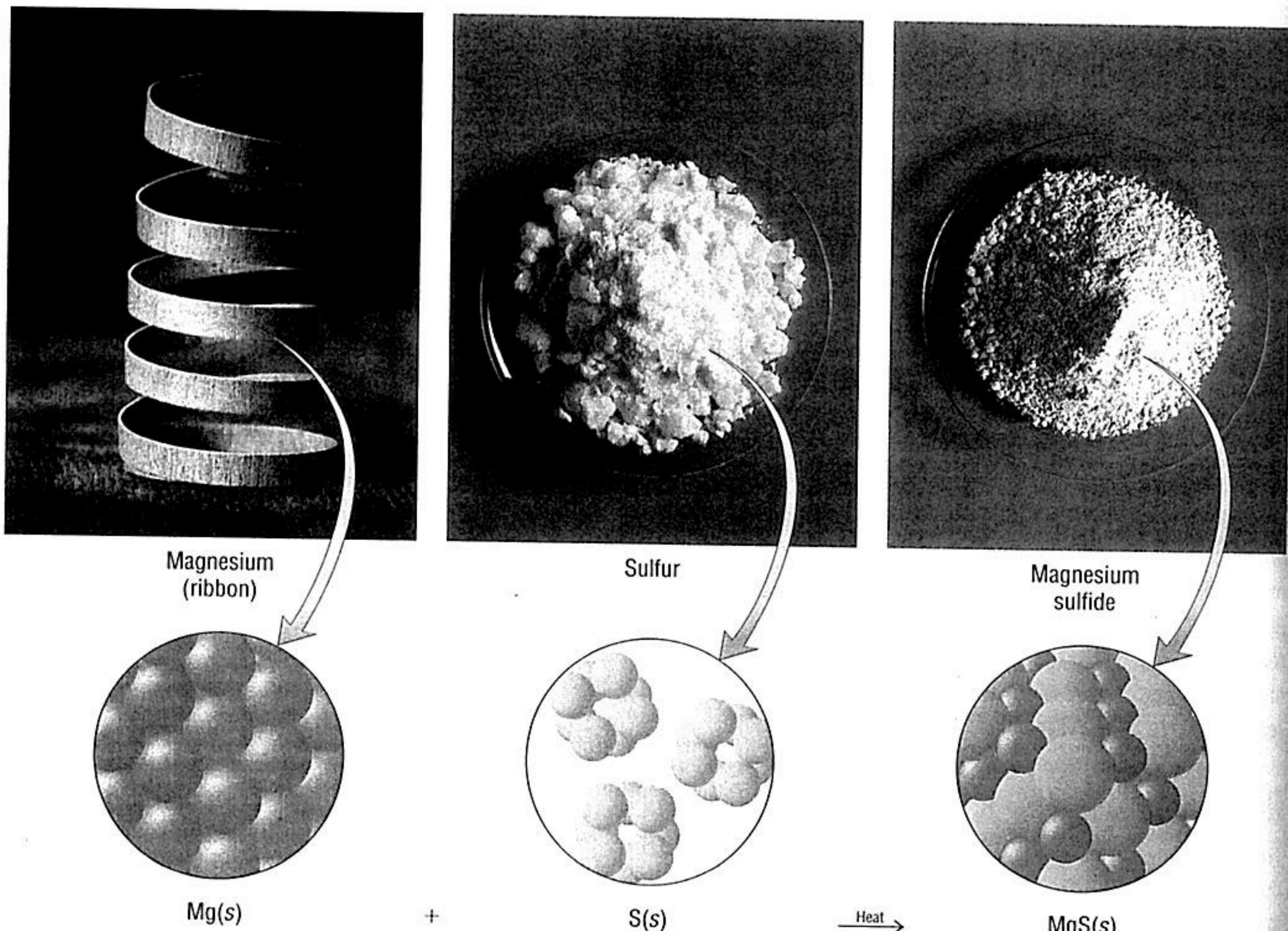


Figure 22.5
When magnesium and sulfur are heated together, they undergo an oxidation–reduction reaction to form magnesium sulfide. The magnesium atoms become more stable by the loss of electrons (oxidation). The sulfur atoms become more stable by the gain of electrons (reduction).

electrons and is reduced to a sulfide ion. The overall process is represented as the two component processes below.

$$\text{Oxidation:}\quad \cdot\text{Mg}\cdot \longrightarrow \text{Mg}^{2+} + 2e^- \quad \text{(loss of electrons)}$$

$$\text{Reduction:}\quad \cdot\ddot{\underset{\cdot\cdot}{\text{S}}}\colon + 2e^- \longrightarrow \colon\ddot{\underset{\cdot\cdot}{\text{S}}}\colon^{2-} \quad \text{(gain of electrons)}$$

The substance in a redox reaction that loses electrons is called the **reducing agent.** By losing electrons to sulfur, magnesium reduces the sulfur. Magnesium is thus the reducing agent. The substance in a redox reaction that accepts electrons is called the **oxidizing agent.** By accepting electrons from magnesium, sulfur oxidizes the magnesium. Sulfur is thus the oxidizing agent.

oxidized

$$\text{Mg}(s) + \text{S}(s) \longrightarrow \text{MgS}(s)$$

reduced

Magnesium (reducing agent)	Sulfur (oxidizing agent)	Magnesium sulfide

Sample Problem 22-1

What is oxidized and what is reduced in this single-replacement reaction? What is the oxidizing agent? The reducing agent?

$$2\text{AgNO}_3(aq) + \text{Cu}(s) \longrightarrow \text{Cu(NO}_3)_2(aq) + 2\text{Ag}(s)$$

1. *ANALYZE* *Plan a problem-solving strategy.*
Begin by rewriting the equation and showing the ions. Determine what loses electrons (is oxidized) and what gains electrons (is reduced).

2. *SOLVE* *Apply the problem-solving strategy.*
The rewritten equation showing the ions is:

$$2\text{Ag}^+ + 2\text{NO}_3^- + \text{Cu} \longrightarrow \text{Cu}^{2+} + 2\text{NO}_3^- + 2\text{Ag}$$

In this reaction, two electrons have been lost from a copper atom (Cu) because it becomes a Cu^{2+} ion. These electrons are gained by two silver ions (Ag^+), which become neutral silver atoms.

$$\text{Oxidation:}\quad \text{Cu} \longrightarrow \text{Cu}^{2+} + 2e^- \quad \text{(loss of electrons)}$$

$$\text{Reduction:}\quad 2\text{Ag}^+ + 2e^- \longrightarrow 2\text{Ag} \quad \text{(gain of electrons)}$$

The Cu is oxidized and is therefore the reducing agent. The Ag^+ is reduced and is therefore the oxidizing agent.

3. *EVALUATE* *Does the result make sense?*
Ag^+, the oxidizing agent, gained electrons and is thus reduced. Cu, the reducing agent, lost electrons and is thus oxidized. The definitions of oxidation and reduction have been correctly applied.

Practice Problems

1. Determine what is oxidized and what is reduced in each reaction. Identify the oxidizing agent and reducing agent in each case.
   a. $2\text{Na}(s) + \text{S}(s) \longrightarrow \text{Na}_2\text{S}(s)$
   b. $4\text{Al}(s) + 3\text{O}_2(g) \longrightarrow 2\text{Al}_2\text{O}_3(s)$
2. Identify these processes as either oxidation or reduction.
   a. $\text{Li} \longrightarrow \text{Li}^+ + e^-$
   b. $2\text{I}^- \longrightarrow \text{I}_2 + 2e^-$
   c. $\text{Zn}^{2+} + 2e^- \longrightarrow \text{Zn}$
   d. $\text{Br}_2 + 2e^- \longrightarrow 2\text{Br}^-$

Chem ASAP!

Problem-Solving 1
Solve Problem 1 with the help of an interactive guided tutorial.

Link to

PHOTOGRAPHY

Redox in Photography

Black-and-white photography involves oxidation–reduction reactions. Exposing black-and-white film to light activates very fine grains of silver bromide in the film. The film is developed by placing it in a developing solution that is actually a reducing agent. The developer, usually an organic chemical such as hydroquinone ($C_6H_4(OH)_2$), reduces the activated silver bromide to finely divided, black metallic silver. Any silver bromide that remains unactivated is removed from the film by using a solvent called a fixer. Sodium thiosulfate, commonly called hypo, is used for this purpose. The areas of the film exposed to the most light appear darkest because they have the highest concentration of metallic silver. The reversed image, called a negative, is used to produce a positive print, the black-and-white photograph, of the image.

It is easy to identify complete transfers of electrons in ionic reactions such as those just examined. But what about reactions that produce covalent compounds; that is, compounds in which complete electron transfer does not occur? Consider the reaction of hydrogen and oxygen.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

According to the older definition of oxidation as a combination with oxygen, it is clear that the hydrogen is oxidized to water. The newer definition involving electron transfer will also provide the same answer. Consider what happens to the bonding electrons in the formation of a water molecule. The bonding electrons in each reactant hydrogen molecule are shared equally between the hydrogen atoms. In water, however, the bonding electrons are pulled toward oxygen because it is much more electronegative than hydrogen. The result is a shift of bonding electrons away from hydrogen, even though there is not a complete transfer.

$$\mathrm{H{-}H} \quad \text{electrons shared equally} \qquad \begin{array}{l}\overset{+\!\longrightarrow}{\mathrm{H{-}O}}\\ \ \ \mathrm{|}\\ \ \ \mathrm{H}\end{array} \uparrow \quad \text{shift of bonding electrons away from hydrogen}$$

Hydrogen is oxidized because it undergoes a partial loss of electrons. Thus the old gain-of-oxygen and the newer loss-of-electron definitions of oxidation agree that it is the hydrogen that is oxidized when the water forms.

What about oxygen, the other reactant? The bonding electrons are shared equally between oxygen atoms in the reactant oxygen molecule, but there is a shift of electrons toward oxygen in water. Oxygen is thus reduced because it undergoes a partial gain of electrons.

$$\mathrm{O{-}O} \quad \text{electrons shared equally} \qquad \begin{array}{l}\overset{+\!\longrightarrow}{\mathrm{H{-}O}}\\ \ \ \mathrm{|}\\ \ \ \mathrm{H}\end{array} \uparrow \quad \text{shift of bonding electrons toward oxygen}$$

In every redox reaction, including those that produce covalent products, there is an oxidizing agent and a reducing agent. In the reaction of hydrogen and oxygen to produce water, hydrogen is the reducing agent because it is oxidized. Oxygen is the oxidizing agent because it is reduced. This redox reaction is highly exothermic—that is, it releases a great deal of energy, as shown in **Figure 22.6.**

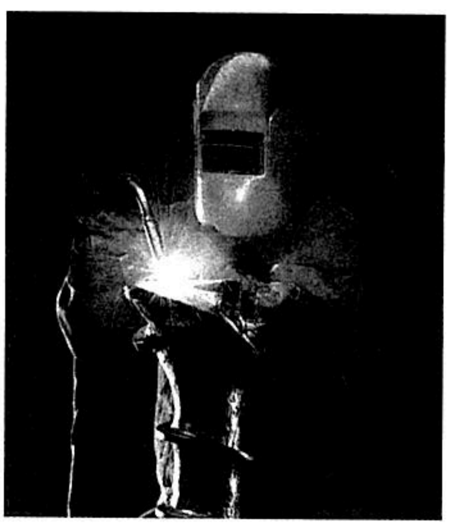


Figure 22.6

This artist uses an oxyhydrogen torch to cut and weld steel to make a sculpture. When hydrogen burns in oxygen, the redox reaction generates temperatures of about 2600 °C.

In some reactions involving covalent products, the partial electron shifts are less obvious. Some general guidelines are thus helpful. For example, for carbon compounds, the addition of oxygen or the removal of hydrogen is always oxidation. **Table 22.1** lists processes that constitute oxidation and reduction. The last entry in the table identifies oxidation numbers, which are another way to describe oxidation and reduction. You will learn more about oxidation numbers in Section 22.2.

Table 22.1

Processes Leading to Oxidation and Reduction	
Oxidation	**Reduction**
Complete loss of electrons (ionic reactions)	Complete gain of electrons (ionic reactions)
Shift of electrons *away* from an atom in a covalent bond	Shift of electrons *toward* an atom in a covalent bond
Gain of oxygen	Loss of oxygen
Loss of hydrogen by a covalent compound	Gain of hydrogen by a covalent compound
Increase in oxidation number	Decrease in oxidation number

Corrosion

Billions of dollars are spent yearly to prevent and to repair damage caused by the corrosion of metals. Iron, a common construction metal often used in the form of the alloy steel, corrodes by being oxidized to ions of iron by oxygen. Water in the environment accelerates the rate of corrosion. Oxygen, the oxidizing agent, is reduced to oxide ions (in compounds such as Fe_2O_3) or to hydroxide ions. The following equations describe the corrosion of iron to iron hydroxides at moist conditions.

$$2Fe(s) + O_2(g) + 2H_2O(l) \longrightarrow 2Fe(OH)_2(s)$$

$$4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \longrightarrow 4Fe(OH)_3(s)$$

Corrosion occurs more rapidly in the presence of salts and acids. These substances produce conducting solutions that make electron transfer easier. The corrosion of some metals can be a desirable feature, as **Figure 22.7** shows.

Not all metals corrode easily. Gold and platinum are called noble metals because they are very resistant to losing their electrons by corrosion. Other metals lose electrons easily but are protected from extensive corrosion by the oxide coating formed on their surface. For example, aluminum oxidizes quickly in air to form a coating of very tightly packed aluminum oxide particles. This coating protects the aluminum object from further corrosion, as

Figure 22.7
Oxidation–reduction reactions cause corrosion. The copper on this roof reacted with water vapor, carbon dioxide, and other substances in the air to form a patina. This patina consists of a pale-green film of basic copper(II) carbonate. Because patinas enhance the surface appearance of objects, they are valued by architects and artists.

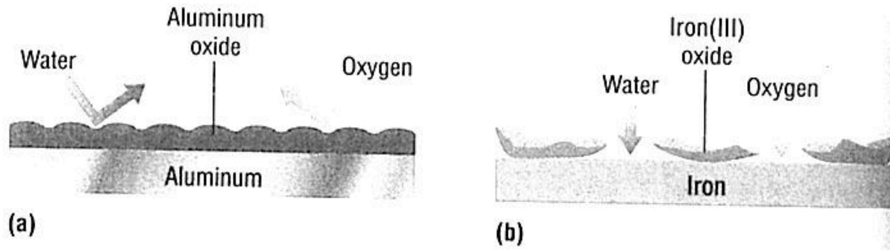


Figure 22.8
*Oxidation causes the complete corrosion of some metals. Aluminum, however, resists such corrosion because it forms a protective coating of aluminum oxide. How does the aluminum oxide on aluminum **(a)** differ from the iron(III) oxide formed on corroding iron **(b)**?*

shown in **Figure 22.8.** Iron also forms a coating when it corrodes, but the coating of iron oxide that forms is not tightly packed. Water in the air can penetrate the coating and attack the iron metal below it. The corrosion continues until the iron object becomes only a pile of rust.

The corrosion of objects such as shovels or knives is a common problem but not usually a serious one. In contrast, the corrosion of a steel support pillar of a bridge or the hull of an oil tanker is much more serious and costly to repair! To prevent corrosion in such cases, the metal surface may be coated with oil, paint, plastic, or another metal, as shown in **Figure 22.9.** These coatings exclude air and water from the surface, thus preventing corrosion. If the coating is scratched or worn away, however, the exposed metal will begin to corrode.

In another method of corrosion control, one metal is used to save a second metal. For example, to protect an iron object, a piece of magnesium may be placed in electrical contact with the iron. When oxygen and water attack the iron object, the iron atoms lose electrons as the iron begins to be oxidized. However, because magnesium is a better reducing agent than iron and is more easily oxidized, the magnesium immediately transfers electrons to the iron ions, reducing them back to neutral iron atoms.

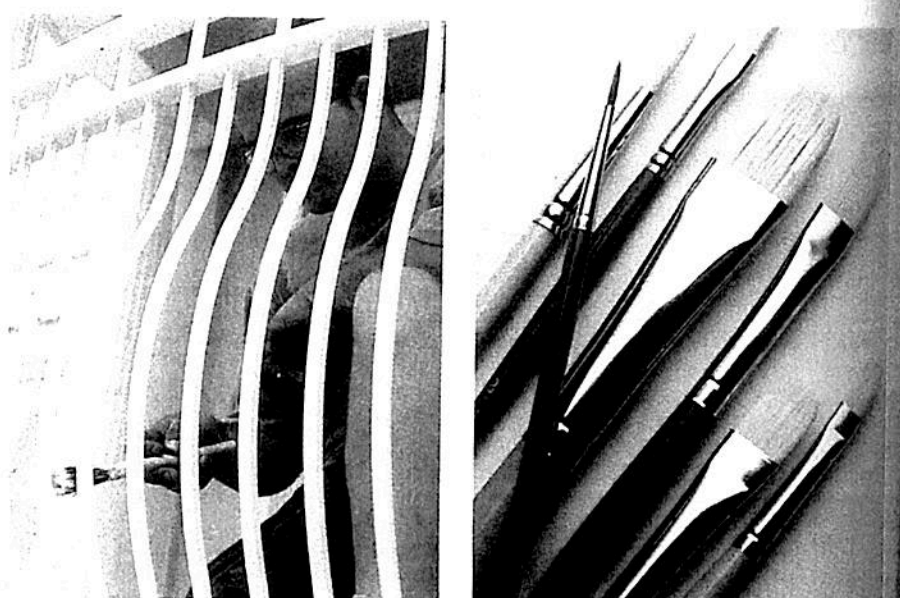


Figure 22.9
Painting a surface (left) protects it from the effects of the environment. Chromium metal also serves as a protective coating and imparts an attractive, mirrorlike finish (right). Like aluminum, chromium forms a corrosion-resistant oxide film on its surface.

Sacrificial zinc and magnesium blocks are sometimes attached to piers and ship hulls to prevent corrosion damage in areas submerged in water, as shown in **Figure 22.10.** Underground pipelines and storage tanks may be connected to magnesium blocks for protection. Obviously, it is easier and cheaper to replace a block of magnesium or zinc than to replace a bridge or a pipeline.

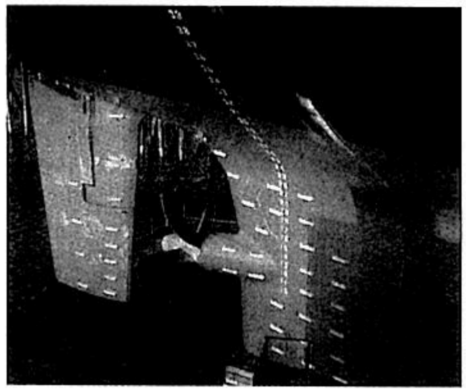


Figure 22.10
Zinc blocks, the white strips in the photograph, are attached to the steel hull of this ship to protect the steel from corrosion. The zinc blocks oxidize (corrode) and release electrons. The steel hull consumes the electrons supplied by the zinc, preventing it from corroding.

section review 22.1

3. Define oxidation and reduction in terms of

a. gain or loss of electrons.

b. gain or loss of hydrogen (in a covalent bond).

c. gain or loss of oxygen.

d. shift of electrons (in a covalent bond).

4. State the characteristics of a redox reaction, and explain how to identify the oxidizing agent and the reducing agent.

5. Which of the following would most likely be oxidizing agents and which would most likely be reducing agents? (*Hint:* Think in terms of tendencies to lose or gain electrons.)

a. Cl_2 **b.** K **c.** Ag^+

6. Refer to the electronegativity values in **Table 14.2** on page 405 to determine which reactant is oxidized and which reactant is reduced in each reaction. Also determine which reactant is the reducing agent and which is the oxidizing agent.

a. $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

b. $S(s) + Cl_2(g) \longrightarrow SCl_2(g)$

c. $N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g)$

d. $2Li(s) + F_2(g) \longrightarrow 2LiF(s)$

e. $H_2(g) + S(s) \longrightarrow H_2S(g)$

7. Use electron transfer or electron shift to identify what is oxidized and what is reduced in each reaction. Make use of electronegativity values, as needed, for molecular compounds.

a. $2Na(s) + Br_2(l) \longrightarrow 2NaBr(s)$

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

c. $S(s) + O_2(g) \longrightarrow SO_2(g)$

d. $Mg(s) + Cu(NO_3)_2(aq) \longrightarrow Mg(NO_3)_2(aq) + Cu(s)$

8. Why would a metal corrode more quickly in salt water than in distilled water?

Chem ASAP! **Assessment 22.1** Check your understanding of the important ideas and concepts in Section 22.1.

portfolio project

Shipwrecks of Spanish galleons often contain gold or silver treasures. Why is the recovered gold hardly changed while the silver has turned black? Research and write a report on how the thick layers of tarnish are removed from silver artifacts.

section 22.2

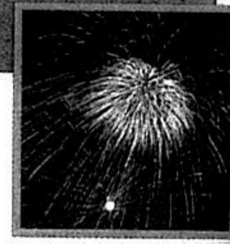
OXIDATION NUMBERS

objectives

- Determine the oxidation number of an atom of any element in a pure substance
- Define oxidation and reduction in terms of a change in oxidation number, and identify atoms being oxidized or reduced in redox reactions

key term

- oxidation number



Did you know that the different colors produced by fireworks are the result of various elements and compounds being burned? Burning sodium produces yellow light. Burning barium and copper compounds produces green light and blue-green light, respectively. The fireworks shown here are called red stars—their beautiful crimson glow coming from the burning of strontium compounds. As elements burn, their oxidation state often changes. **How are oxidation numbers assigned, and how are they used to analyze redox reactions?**

Assigning Oxidation Numbers

An **oxidation number** is a positive or negative number assigned to an atom according to a set of rules. Oxidation numbers can be thought of as a chemical bookkeeping device. As you will learn in Section 22.3, complex redox equations can be balanced by the use of oxidation-number changes. As a general rule, a bonded atom's oxidation number is the charge that it would have if the electrons in the bond were assigned to the atom of the more electronegative element. In binary ionic compounds, such as NaCl and $CaCl_2$, the oxidation numbers of the atoms equal their ionic charges. The compound sodium chloride is composed of sodium ions (Na^{1+}) and chloride ions (Cl^{1-}). Thus the oxidation number of sodium is +1, and that of chlorine is −1. Notice that when oxidation numbers are written, the sign is put before the number. Sodium in NaCl has an ionic charge of 1+ and an oxidation number of +1. What are the oxidation numbers of calcium and of fluorine in calcium fluoride (CaF_2)?

Because water is a molecular compound, no ionic charges are associated with its atoms. As you learned in Section 22.1, however, oxygen is reduced in the formation of water. Oxygen is more electronegative than

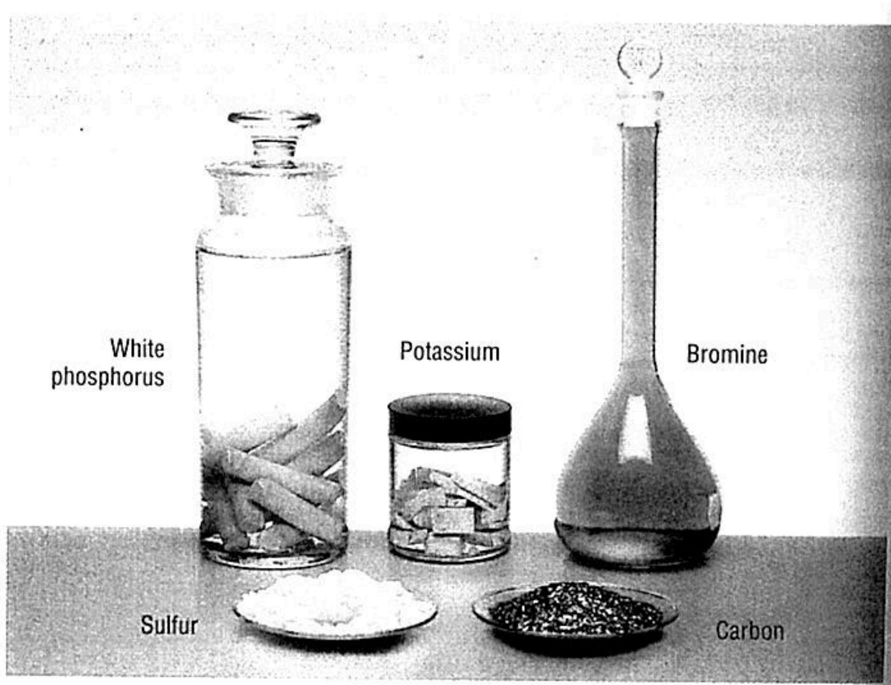


Figure 22.11

The oxidation number of any element in the free or uncombined state is zero. The elements shown here (left to right) are white phosphorus (stored under water), sulfur, potassium (stored under oil), carbon, and bromine liquid. The potassium and phosphorus are stored under a liquid to prevent them from reacting with oxygen in the air.

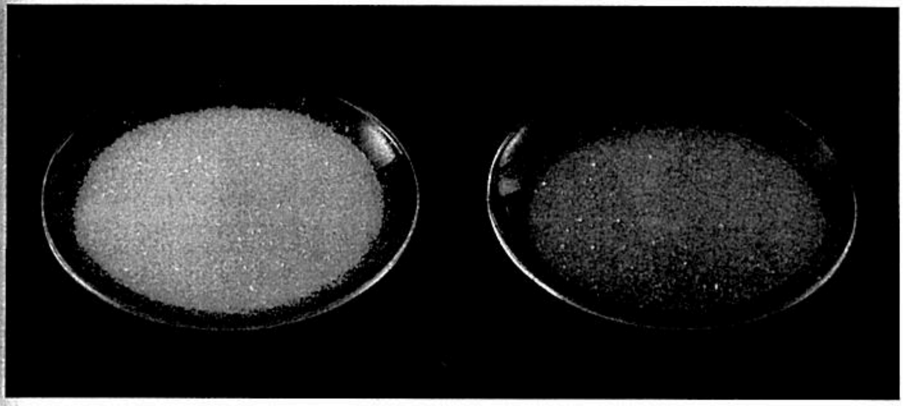


Figure 22.12
Yellow potassium chromate (K_2CrO_4) and orange potassium dichromate ($K_2Cr_2O_7$) each have a chromium-containing polyatomic ion. What are the formulas of the chromate ion and the dichromate ion?

hydrogen. In water, the two shared electrons in the H—O bond are shifted toward oxygen and away from hydrogen. Imagine that the electrons contributed by the hydrogen atoms are completely transferred to the oxygen. The charges that would result from this transfer are the oxidation numbers of the bonded elements. The oxidation number of oxygen is −2. The oxidation number of each hydrogen is +1. Oxidation numbers are often written above the chemical symbols in a formula. For example, water can be represented as

$$\overset{+1}{H_2}\overset{-2}{O}$$

The following set of rules should help you determine oxidation numbers.

Rules for Assigning Oxidation Numbers

1. The oxidation number of a monatomic ion is equal in magnitude and sign to its ionic charge. For example, the oxidation number of the bromide ion (Br^{1-}) is −1; that of the Fe^{3+} ion is +3.
2. The oxidation number of hydrogen in a compound is +1, except in metal hydrides, such as NaH, where it is −1.
3. The oxidation number of oxygen in a compound is −2, except in peroxides, such as H_2O_2, where it is −1.
4. The oxidation number of an atom in uncombined (elemental) form is 0. For example, the oxidation number of the potassium atoms in potassium metal (K) or of the nitrogen atoms in nitrogen gas (N_2) is 0. See **Figure 22.11.**
5. For any neutral compound, the sum of the oxidation numbers of the atoms in the compound must equal 0.
6. For a polyatomic ion, the sum of the oxidation numbers must equal the ionic charge of the ion.

The last two rules can be used together to determine the oxidation number of atoms not covered in the first four rules. The yellow crystals and the orange crystals in **Figure 22.12** are both compounds of chromium. What is the oxidation number of chromium in each compound?

Sample Problem 22-2

What is the oxidation number of each kind of atom in the following compounds?

a. SO_2 **b.** CO_3^{2-} **c.** K_2SO_4

1. ANALYZE ***Plan a problem-solving strategy.***

a.–c. Use the set of rules you just learned to assign oxidation numbers and to calculate unknown ones.

2. SOLVE ***Apply the problem-solving strategy.***

a. There are two oxygen atoms and the oxidation number of each oxygen is −2 (rule 3). You also know that the sum of the oxidation numbers for the neutral compound must be 0 (rule 5). Therefore the oxidation number of sulfur is +4, because $+4 + (2 \times (-2)) = 0$.

$$\overset{+4}{S}\overset{-2}{O_2}$$

b. The oxidation number of oxygen is −2 (rule 3).

$$\overset{?}{C}\overset{-2}{O_3}{}^{2-}$$

The sum of the oxidation numbers of the carbon and oxygen atoms must equal the ionic charge, 2− (rule 6). The oxidation number of carbon must be +4, because $+4 + (3 \times (-2)) = -2$.

$$\overset{+4}{C}\overset{-2}{O_3}{}^{2-}$$

c. The oxidation number of the potassium ion is the same as its ionic charge, +1 (rule 1). The oxidation number of oxygen is −2 (rule 3).

$$\overset{+1}{K_2}\overset{?}{S}\overset{-2}{O_4}$$

For the sum of the oxidation numbers in the compound to be 0 (rule 5), the oxidation number of sulfur must be +6, because $(2 \times (+1)) + (+6) + (4 \times (-2)) = 0$.

$$\overset{+1}{K_2}\overset{+6}{S}\overset{-2}{O_4}$$

3. EVALUATE ***Do the results make sense?***

The results are consistent with the rules for determining oxidation numbers. Rule 1 was used to find the oxidation number of potassium in K_2SO_4. Rule 3 was used to find the oxidation number of oxygen in all three compounds. Rule 5 was used to find the oxidation number of sulfur in SO_2 and K_2SO_4. Rule 6 was used to find the oxidation number of carbon in CO_3^{2-}. Also, addition of the oxidation numbers correctly gives the final overall charge for the ion and the two neutral compounds.

Practice Problems

9. Determine the oxidation number of each element in these substances.

a. S_2O_3 **b.** O_2 **c.** $Al_2(SO_4)_3$

d. Na_2O_2

10. Find the oxidation number of each kind of atom in the following.

a. P_2O_5 **b.** NH_4^+

c. $Na_2Cr_2O_7$ **d.** $Ca(OH)_2$

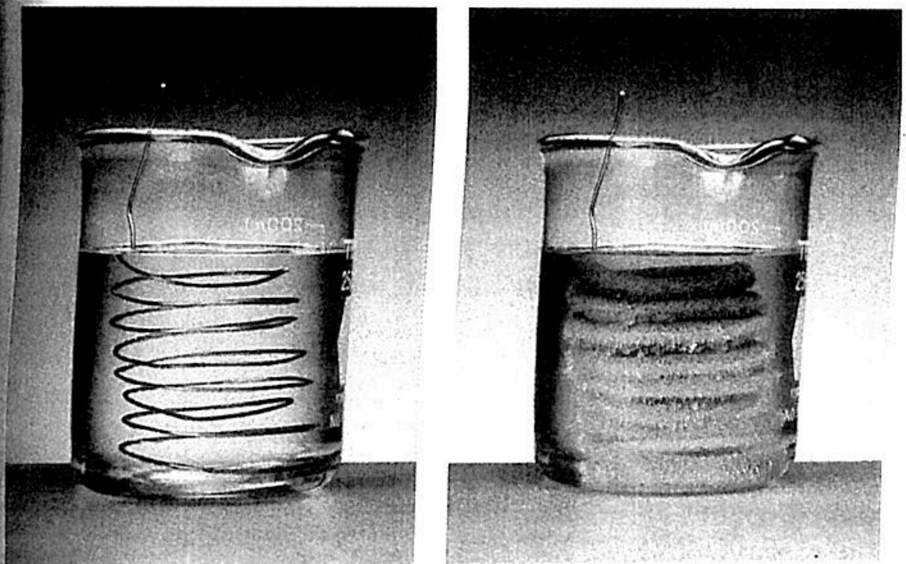


Figure 22.13
When a copper wire is placed in a colorless silver nitrate solution (left), crystals of silver coat the wire (right). The solution slowly turns blue as a result of the formation of copper(II) nitrate. What change occurs in the oxidation number of the silver? How does the oxidation number of the copper change?

Oxidation-Number Changes in Chemical Reactions

An increase in the oxidation number of an atom indicates oxidation. A decrease in the oxidation number of an atom indicates reduction. Look again at the equation in Sample Problem 22-1 on page 649 and at **Figure 22.13.** Can you identify what is being oxidized and what is being reduced on the basis of oxidation-number changes? Here is the equation with oxidation numbers added:

$$2\overset{+1}{Ag}\overset{+5}{N}\overset{-2}{O_3}(aq) + \overset{0}{Cu}(s) \longrightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O_3})_2(aq) + 2\overset{0}{Ag}(s)$$

In this reaction, the oxidation number of silver decreases from +1 to 0, which indicates reduction: Silver ions (Ag^{1+}) reduce to silver metal (Ag^0). Copper is oxidized in this reaction. Its oxidation number increases from 0 to +2 as copper metal oxidizes from Cu^0 to Cu^{2+}. Note that these results agree with those obtained by analyzing the electron transfers that occur in the reaction. **Figure 22.14** illustrates a redox reaction that shows what occurs when a shiny iron nail is dipped into a solution of copper(II) sulfate.

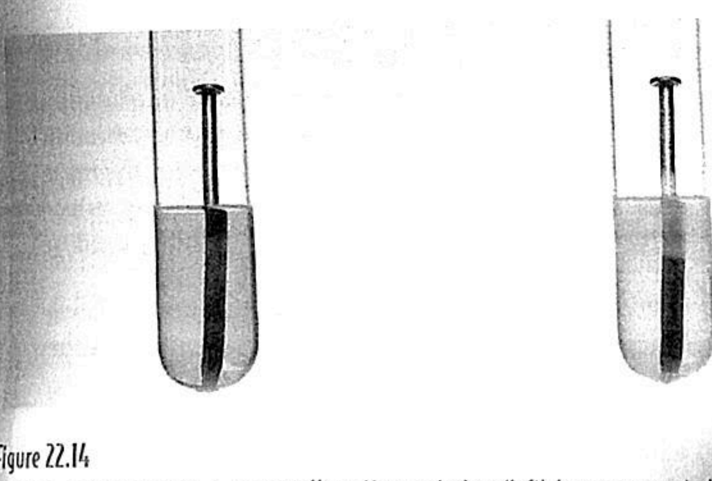


Figure 22.14
An iron nail dipped in a copper(II) sulfate solution (left) becomes coated with metallic copper (right). The iron reduces Cu^{2+} ions in solution and is simultaneously oxidized to Fe^{2+}. Write the balanced ionic equation for this redox reaction.

Sample Problem 22-3

Use the changes in oxidation number to identify which atoms are oxidized and which are reduced in each reaction.

a. $Cl_2(g) + 2HBr(aq) \longrightarrow 2HCl(aq) + Br_2(l)$

b. $C(s) + O_2(g) \longrightarrow CO_2(g)$

c. $Zn(s) + 2MnO_2(s) + 2NH_4Cl(aq) \longrightarrow$
$ZnCl_2(aq) + Mn_2O_3(s) + 2NH_3(g) + H_2O(l)$

1. ANALYZE ***Plan a problem-solving strategy.***

Use the rules to assign an oxidation number to each atom on both sides of the equation. Note the increases and decreases in oxidation numbers. On the basis of these changes, identify the atoms oxidized and those reduced. A decrease in oxidation number indicates reduction. An increase in oxidation number indicates oxidation.

2. SOLVE ***Apply the problem-solving strategy.***

a. $\overset{0}{Cl_2}(g) + 2\overset{+1-1}{HBr}(aq) \longrightarrow 2\overset{+1\,-1}{HCl}(aq) + \overset{0}{Br_2}(l)$

The element chlorine is reduced because its oxidation number decreases (0 to −1). The bromide ion is oxidized because its oxidation number increases (−1 to 0).

b. $\overset{0}{C}(s) + \overset{0}{O_2}(g) \longrightarrow \overset{+4-2}{CO_2}(g)$

The element carbon is oxidized (0 to +4). The element oxygen is reduced (0 to −2).

c. $\overset{0}{Zn}(s) + 2\overset{+4\,-2}{MnO_2}(s) + 2\overset{-3+1\,-1}{NH_4Cl}(aq) \longrightarrow$
$\overset{+2\,-1}{ZnCl_2}(aq) + \overset{+3\,-2}{Mn_2O_3}(s) + 2\overset{-3+1}{NH_3}(g) + \overset{+1\,-2}{H_2O}(l)$

The element zinc is oxidized (0 to +2). The manganese ion is reduced (+4 to +3).

3. EVALUATE ***Do the results make sense?***

Checking the results reveals that the assignments of oxidation numbers are correct; that is, the rules for assigning oxidation numbers have been correctly applied. In each case, a decrease in oxidation number has correctly been used to indicate reduction, and an increase in oxidation number has been used to indicate oxidation.

Recall that in every redox reaction there is an oxidizing agent and a reducing agent. The element that is oxidized is the reducing agent, and the element that is reduced is the oxidizing agent. Thus once you have identified the elements that are oxidized and the elements that are reduced in a reaction, you can use this information to identify the oxidizing agent and the reducing agent. This is illustrated in Sample Problem 22-4.

Practice Problem

11. Use the changes in oxidation numbers to identify which atoms are oxidized and which are reduced in each reaction.

a. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

b. $2KNO_3(s) \longrightarrow 2KNO_2(s) + O_2(g)$

c. $NH_4NO_2(s) \longrightarrow N_2(g) + 2H_2O(g)$
(*Hint:* Consider each N in NH_4NO_2 separately.)

d. $PbO_2(aq) + 4HI(aq) \longrightarrow I_2(aq) + PbI_2(s) + 2H_2O(l)$

Chem ASAP!

Problem-Solving 11

Solve Problem 11 with the help of an interactive guided tutorial.