20.2 Oxidation Numbers

Connecting to Your World The bursts of bright white light produced by fireworks are the result of metals being burned. Powdered or flaked aluminum or magnesium are included in the fireworks. When these metals are heated to high temperatures in the explosion, they burn with an intense white light. The fireworks shown here are called stars. As elements burn, their oxidation numbers change. In this section, you will learn about how oxidation and reduction are defined in terms of a change in oxidation number.

Guide for Reading

Key Concepts

- What is the general rule for assigning oxidation numbers?
- How are oxidation and reduction defined in terms of a change in oxidation number?

Vocabulary

oxidation number

Reading Strategy

Using Prior Knowledge Before you read, predict how oxidation and reduction affect the charge on an ion. Check your predictions as you read.

Assigning Oxidation Numbers

An **oxidation number** is a positive or negative number assigned to an atom to indicate its degree of oxidation or reduction. **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.** 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, and in compounds with the more electronegative fluorine, where it is positive.
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.
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.

Section Resources

Print

- **Guided Reading and Study Workbook,** Section 20.2
- **Core Teaching Resources,** Section 20.2 Review
- **Transparencies,** T234–T237

Technology

- **Interactive Textbook with ChemASAP,** Problem-Solving 20.9, Assessment 20.2

20.2

1 FOCUS

Objectives

20.2.1 Determine the oxidation number of an atom of any element in a pure substance.

20.2.2 Define oxidation and reduction in terms of a change in oxidation number and **identify** atoms being oxidized or reduced in redox reactions.

Guide for Reading

Build Vocabulary L2

Word Forms Students can relate the vocabulary term, *oxidation number*, with other similar terms used in the chapter.

Reading Strategy L2

Outline Have students preview the section's content by making an outline of the content.

2 INSTRUCT

Connecting to Your World

Have students read the opening paragraph. Ask, **What other colors can be seen in fireworks?** *(yellow, green, blue-green, red)* **Which elements are responsible for these colors?** *(yellow — sodium; green — barium; blue-green — copper; crimson — strontium)*

Assigning Oxidation Numbers

Relate L2

Explain that determining oxidation numbers of elements in compounds is a way for chemists to keep track of electron transfer during redox reactions. Ask, **What are other examples where items are numbered to keep track of movement?** *(The numbers on sports players' jerseys and the area codes assigned to telephone numbers in different regions are two examples.)*

TEACHER Demo

Redox Reactions L2

Purpose Students will observe a redox reaction and write a balanced equation describing it.

Materials 100 mL of 0.1*M* lead(II) acetate [$Pb(CH_3COO)_2$], 150-mL beaker, 1-cm × 3-cm zinc strip, sodium sulfide, 3*M* NaOH, 1*M* iron(III) chloride, stirring rod, plastic container

Safety Lead(II) acetate is a suspected carcinogen and mutagen. Perform activity in a well-ventilated area. Wear safety goggles, aprons, and disposable gloves.

Procedure Pour 100 mL of 0.1*M* lead(II) acetate [$Pb(CH_3COO)_2$] into a 150-mL beaker and add a 1-cm × 3-cm zinc strip. **CAUTION!** Lead(II) acetate is a suspected carcinogen and mutagen.

Expected Outcome A "tree" of lead that forms over the course of a day.

Ask students to write a balanced equation for the redox reaction that occurred. Have them assign oxidation numbers for Zn and Pb. Based on changes in oxidation number, have students identify the elements that were oxidized and reduced.

Answers $Zn(s) + Pb(CH_3COO)_2(aq) \rightarrow Pb(s) + Zn(CH_3COO)_2(aq)$ Reactants: Zn^0, Pb^{+2}; Products: Zn^{+2}, Pb^0
Zn was oxidized. Pb was reduced.

Disposal Dispose of the reaction mixture by adding a threefold molar excess of sodium sulfide to the contents of the beaker and stirring occasionally for one hour. Adjust the pH to neutral with 3*M* NaOH. Filter or decant the mixture. Allow the residue to dry. Add a threefold excess of 1*M* iron(III) chloride to the filtrate, stirring slowly. Allow the precipitate to settle, then filter or decant. Put the dry precipitate and filtered residue in a plastic container and bury in an approved landfill site. Flush the neutral filtrate down the drain with excess water.

Figure 20.11 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 and vapor). The potassium and phosphorus are stored under a liquid to prevent them from reacting with oxygen in the air.

In binary ionic compounds, such as NaCl and $CaCl_2$, the oxidation numbers of the atoms equal their ionic charges (Rule 1). 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 the sign is put before the oxidation number.

Because water is a molecular compound, no ionic charges are associated with its atoms. However, oxygen is reduced in the formation of water. Oxygen is more electronegative than hydrogen. So, 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 two 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, and the oxidation number of each hydrogen is +1 (Rules 2 and 3). Oxidation numbers are often written above the chemical symbols in a formula. For example, water can be represented as

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

Many elements can have several different oxidation numbers. Use rules 5 and 6 to determine the oxidation number of atoms of these elements, plus other elements not covered in the first four rules. The substances in Figure 20.12 are compounds of chromium, but chromium has a different oxidation number in each compound.

Figure 20.12 Yellow potassium chromate (K_2CrO_4) and orange potassium dichromate ($K_2Cr_2O_7$) each have a chromium-containing polyatomic ion. **Inferring** ***What is the oxidation number of chromium in each compound?***

640 Chapter 20

Differentiated Instruction

Gifted and Talented L3

Have students do research and report to the class why carbon monoxide is extremely poisonous to living systems whereas carbon dioxide is relatively harmless. Point out that carbon in carbon monoxide has an oxidation number that is lower than that in carbon dioxide. Therefore, one difference is that carbon monoxide is a better reducing agent than carbon dioxide.

CONCEPTUAL PROBLEM 20.2

Assigning Oxidation Numbers to Atoms

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

a. SO_2 **b.** CO_3^{2-} **c.** Na_2SO_4 **d.** $(NH_4)_2S$

1 Analyze ***Identify the relevant concepts.***

Use the set of rules you just learned to assign and calculate oxidation numbers.

2 Solve ***Apply concepts to this situation.***

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

$$\overset{+4-2}{SO_2}$$

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

$$\overset{?-2}{CO_3^{2-}}$$

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 $+4 + (3 \times (-2)) = -2$.

$$\overset{+4-2}{CO_3^{2-}}$$

c. The oxidation number of each sodium ion, Na^+, is the same as its ionic charge, +1 (Rule 1). The oxidation number of oxygen is −2 (Rule 3).

$$\overset{+1\ ?-2}{Na_2SO_4}$$

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

$$\overset{+1+6-2}{Na_2SO_4}$$

d. Ammonium ions, NH_4^+, have an ionic charge of +1, so the sum of the oxidation numbers of the atoms in the ammonium ion must be +1. The oxidation number of hydrogen is +1 in this ion.

$$\overset{?\ +1}{NH_4^+}$$

$$? + 4(+1) = +1$$

So, the oxidation number of nitrogen must be −3 because $-3 + (+4) = +1$.

Two ammonium ions have a total charge of +2. Since the compound $(NH_4)_2S$ is neutral, sulfur must have a balancing oxidation number of −2.

$$\overset{-3+1\ \ -2}{(NH_4)_2S}$$

Practice Problems

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

a. S_2O_3 **b.** Na_2O_2 **c.** P_2O_5 **d.** NO_3^-

10. Determine the oxidation number of chlorine in each of the following substances.

a. $KClO_3$ **b.** Cl_2 **c.** $Ca(ClO_4)_2$ **d.** Cl_2O

Interactive Textbook

Problem-Solving 20.9 Solve Problem 9 with the help of an interactive guided tutorial.

with ChemASAP

Discuss L2

Lead a class discussion on determining oxidation numbers for elements in compounds by applying the rules for assigning oxidation numbers. Go over many examples until students feel comfortable with determining these numbers. Also explain that the same element can have several oxidation numbers, depending on the compound. Ask, **What are the oxidation numbers of iodine in HIO_4, HIO_3, HIO, I_2, and HI?** *(+7, +5, +1, 0, −1, respectively)*

CONCEPTUAL PROBLEM 20.2

Answers

9. a. S, +3; O, −2
b. Na, +1; O, −1
c. P, +5; O, −2
d. N, +5; O, −2

10. a. +5
b. 0
c. +7
d. +1

CLASS Activity

Assigning Oxidation Numbers L2

Purpose Students will work cooperatively to assign oxidation numbers for various elements.

Procedure Write formulas for ionic and molecular compounds on the chalkboard. Divide the class into groups of two and have students work together to assign oxidation numbers to all of the elements in each compound. Students can use the problem-solving strategy outlined in Conceptual Problem 20.2.

Facts and Figures

How a Breathalyzer Works

A breath analyzer, or Breathalyzer, is used to estimate the level of alcohol in a person's bloodstream. The chemical potassium dichromate is used in this process. If any ethanol is present in the person's breath when breathing into the apparatus, the Cr^{+6} in the potassium dichromate (orange) is reduced to one of its lower states, such as Cr^{3+} (green). The color change can be related to the amount of alcohol in the bloodstream.

Answers to...

Figure 20.12 K_2CrO_4, Cr +6; $K_2Cr_2O_7$, Cr +6.

Oxidation-Number Changes in Chemical Reactions

CLASS Activity

Identifying Redox Reactions L2

Purpose Students will identify and write equations for redox reactions.

Materials 3 small, glass test tubes; zinc; copper; magesium; 0.1*M* HCl; medicine dropper or pipet

Procedure Have students work in pairs. Give each pair a set of three small, glass test tubes containing a small portion of zinc, copper, and magnesium. Have students add approximately 10 drops of 0.1*M* HCl to the metals and note signs of reaction. They should see bubbles form in the tubes containing zinc and magnesium, but not in the tube containing copper. Ask, **Can you guess the identity of the gas in the tubes containing zinc and magnesium?** (*H_2*(g)) Have students write equations for the reactions of zinc and magnesium with HCl. Have students explain why these reactions are redox reactions.

Use Visuals L1

Figure 20.14 Have students examine the figure 20.14 and ask, **What would happen if a copper nail were added to a solution of iron(II) sulfate instead?** *(Nothing would happen because copper is below iron in the activity series of metals.)*

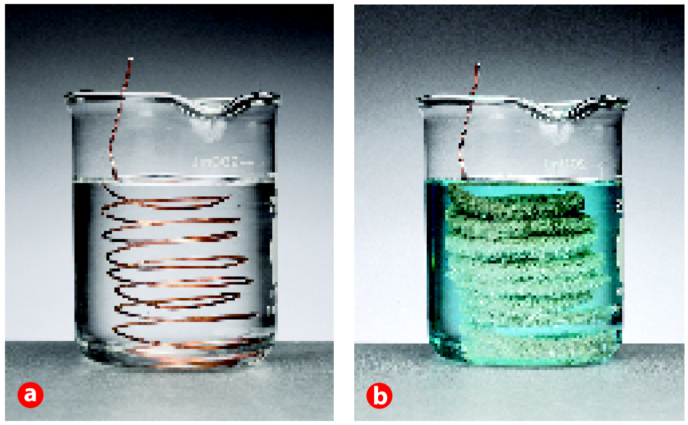


Figure 20.13 Copper reacts with silver nitrate. **a** When a copper wire is placed in a silver nitrate solution crystals of silver coat the wire. **b** The solution slowly turns blue as a result of the formation of copper(II) nitrate. **Drawing Conclusions** ***What change occurs in the oxidation number of silver? How does the oxidation number of copper change?***

Oxidation-Number Changes in Chemical Reactions

Figure 20.13 shows what happens when copper wire is placed in a solution of silver nitrate. In this reaction, the oxidation number of silver decreases from +1 to 0 as each silver ion (Ag^{1+}) gains an electron and is reduced to silver metal (Ag^0). Copper's oxidation number increases from 0 to +2 as each atom of copper metal (Cu^0) loses two electrons and is oxidized to a copper(II) ion (Cu^{2+}). Here is the equation with oxidation numbers added:

$$\overset{+1\,+5\,-2}{2AgNO_3}(aq) + \overset{0}{Cu}(s) \longrightarrow \overset{+2\,+5\,-2}{Cu(NO_3)_2}(aq) + \overset{0}{2Ag}(s)$$

Figure 20.14 illustrates a redox reaction that shows what occurs when a shiny iron nail is dipped into a solution of copper(II) sulfate.

You can define oxidation and reduction in terms of a change in oxidation number. **An increase in the oxidation number of an atom or ion indicates oxidation. A decrease in the oxidation number of an atom or ion indicates reduction.**

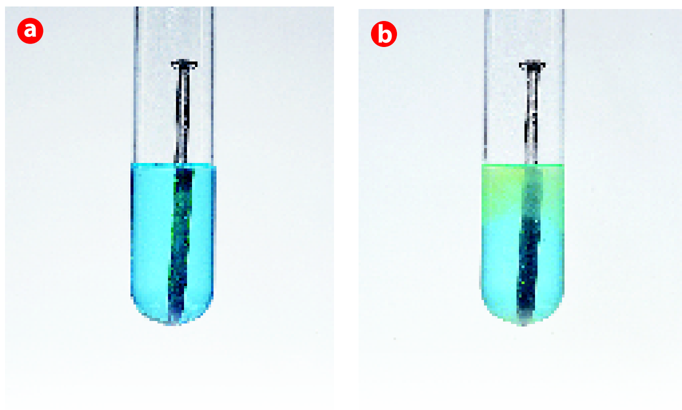
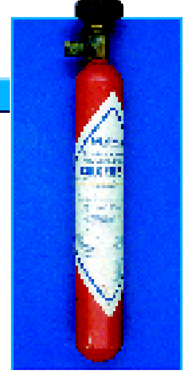


Figure 20.14 A redox reaction occurs between iron and copper. **a** An iron nail dipped in a copper(II) sulfate solution becomes coated with metallic copper. **b** The iron reduces Cu^{2+} ions in solution and is simultaneously oxidized to Fe^{2+}.

642 Chapter 20

CONCEPTUAL PROBLEM 20.3



Identifying Oxidized and Reduced Atoms

Use changes in oxidation number to identify which atoms are oxidized and which are reduced in the following reaction. Also identify the oxidizing agent and the reducing agent.

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

1 Analyze ***Identify the relevant concepts.***

Use the rules to assign oxidation numbers to each atom in the equation. An increase in oxidation number indicates oxidation. A decrease in oxidation number indicates reduction. The substance that is oxidized in a redox reaction is the reducing agent. The substance that is reduced is the oxidizing agent.

2 Solve ***Apply concepts to this situation.***

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

The element chlorine is reduced because its oxidation number decreases (0 to −1). The bromide ion from HBr(*aq*) is oxidized because its oxidation number increases (−1 to 0). Chlorine is reduced, so Cl_2 is the oxidizing agent. The bromide ion from HBr(*aq*) is oxidized, so Br^- is the reducing agent.

Practice Problems

11. Identify which atoms are oxidized and which are reduced in each reaction.

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

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

12. Identify the oxidizing agent and the reducing agent in each equation in Practice Problem 11.

Interactive Textbook

Problem-Solving 20.11 Solve Problem 11 with the help of an interactive guided tutorial.

with **ChemASAP**

20.2 Section Assessment

13. **Key Concept** What is the general rule for assigning oxidation numbers?

14. **Key Concept** How is a change in oxidation number related to the process of oxidation and reduction?

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

a. $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

b. $2HNO_3(aq) + 6HI(aq) \longrightarrow 2NO(g) + 3I_2(s) + 4H_2O(l)$

16. Identify the oxidizing agent and the reducing agent in each reaction in Problem 15.

Connecting Concepts

Types of Reactions Reread Section 11.2 and identify which of the five types of reactions are most likely redox reactions.

Interactive Textbook

Assessment 20.2 Test yourself on the concepts in Section 20.2.

with **ChemASAP**

CONCEPTUAL PROBLEM 20.3

Answers

11. a. H_2 oxidized, O_2 reduced **b.** N reduced, O oxidized

12. a. H_2 reducing agent, O_2 oxidizing agent **b.** N oxidizing agent, O reducing agent

3 ASSESS

Evaluate Understanding L2

Ask, **What are the oxidation numbers of the elements in the following compounds? KNO_3**(*K +1, N +5, O −2);* **Na_2SO_3**(*Na +1, S +4, O −2);* **P_2O_5** *(P +5, O −2);* **MgH_2**(*Mg +2, H −1);* **ClF_3** *(Cl +3, F −1);* **HIO** *(H +1, I +1, O −2)*

What does the sign and magnitude of the oxidation number of an element in a compound indicate? *(The sign indicates loss (+) or gain (−) of electrons compared to the isolated element; the magnitude indicates the degree of the gain or loss.)* **If fluorine is converted to F^- from F_2, what is the change in oxidation number? Is this oxidation or reduction?** *(change is 0 to −1; reduction)*

Reteach L1

Stress that oxidation numbers refer to the combining capacity of single atoms. Thus, a change in oxidation number must be calculated on a per-atom basis. Use examples such as Mn_2O_3 and $Cr_2O_7^{2-}$ to explain to students that they must consider the number of atoms involved.

Connecting Concepts

combination, decomposition, single-replacement, combustion.

Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 20.2.

with **ChemASAP**

Answers to...

Figure 20.13 The oxidation number of silver changes from +1 to 0; the oxidation number of copper changes from 0 to +2.

Section 20.2 Assessment

13. The oxidation number is the charge a bonded atom would have if the electrons in the bond were assigned to the more electronegative element.

14. An increase in oxidation number indicates oxidation; a decrease in oxidation number indicates reduction.

15. a. Na, oxidized. Cl_2, reduced **b.** I, oxidized; N, reduced

16. a. Na reducing agent, Cl_2 oxidizing agent **b.** HNO_3 oxidizing agent, HI reducing agent

Technology & Society

Cold Water Makes a Hot Meal

Developing the FRH was not as simple as packaging a few grams of magnesium in a box. When exposed to air, magnesium oxidizes to magnesium oxide (MgO), which forms a thin film on the metal's surface. This film prevents further oxidation of the metal, either by air or by water. So the researchers had to find a way of letting water get through the MgO film.

They discovered that adding common table salt (NaCl) and powdered iron to the magnesium would do the trick. When water is added to this mixture, the chloride ions (Cl^-) react with the MgO film to form MgOHCl, which eats away at the remaining MgO.

Once the MgO film is gone, the water can react directly with the magnesium metal. The powdered iron seems to help this reaction take place more rapidly than it normally would. The researchers do not understand exactly how the iron helps, but they theorize that it assists in the transfer of electrons between the water and the magnesium. In any case, with the iron present, the reaction occurs rapidly and vigorously. As a result of putting corrosion to work, the energy stored in the chemical bonds of the reactants is released in just a few minutes, and a hot meal is ready.

After students have read the article, ask, **Why is water needed to activate the flameless ration heater?** *(It serves as an oxidizing agent.)*

Ask, **What is the reducing agent in this reaction?** *(Mg)*

Name other metals that might react with water to generate heat and hydrogen gas. *(Students answers will vary, but may include Ca, Sr, or Ba.)*

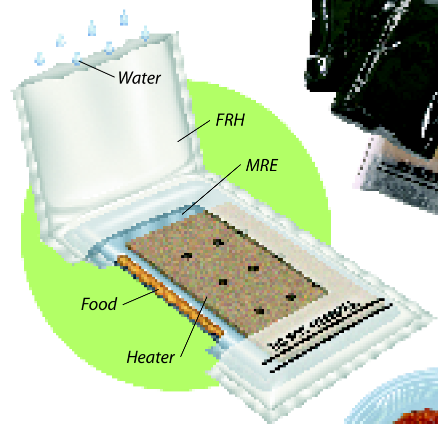
Applying Concepts

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2 + H_2(g)$

Technology & Society

Cold Water Makes a Hot Meal

Researchers have solved a long-standing problem: how to provide astronauts and military personnel with hot meals when there are no cooking facilities or time to cook. They developed the Flameless Ration Heater (FRH). An FRH is very small and weighs only one and a half ounces. It heats precooked, packaged meal rations with no flame or electricity. Simply place an MRE (Meal Ready to Eat) pouch in an FRH and add water. After about 15 minutes, the food is a warm 60°C. **Applying Concepts** ***What is the chemical equation for the reaction of magnesium and water?***



How it works An FRH uses corrosion to generate heat. Water oxidizes magnesium to form magnesium hydroxide, hydrogen, and heat. Just 24 grams of magnesium releases 355,000 joules of heat, enough to boil about 1 L of water.

Eating in space Astronauts N. Jan Davis and Stephen K. Robinson on the space shuttle Discovery eat Japanese rice. Astronauts use the FRH to heat packaged meals.

Facts and Figures

Preventing Metal Corrosion

People spend a considerable amount of time and energy trying to prevent or inhibit the simple redox reaction called corrosion. Many metals oxidize when exposed to the oxygen in air. To keep them from corroding or rusting, the metals can be painted, plated with chromium, or galvanized with zinc. But corrosion is not all bad. Because the oxidation of metal is generally an exothermic reaction, it can be used to generate useful heat—as long as the oxidation takes place quickly enough to release the heat energy in a short span of time.