Electrochemical Reactions

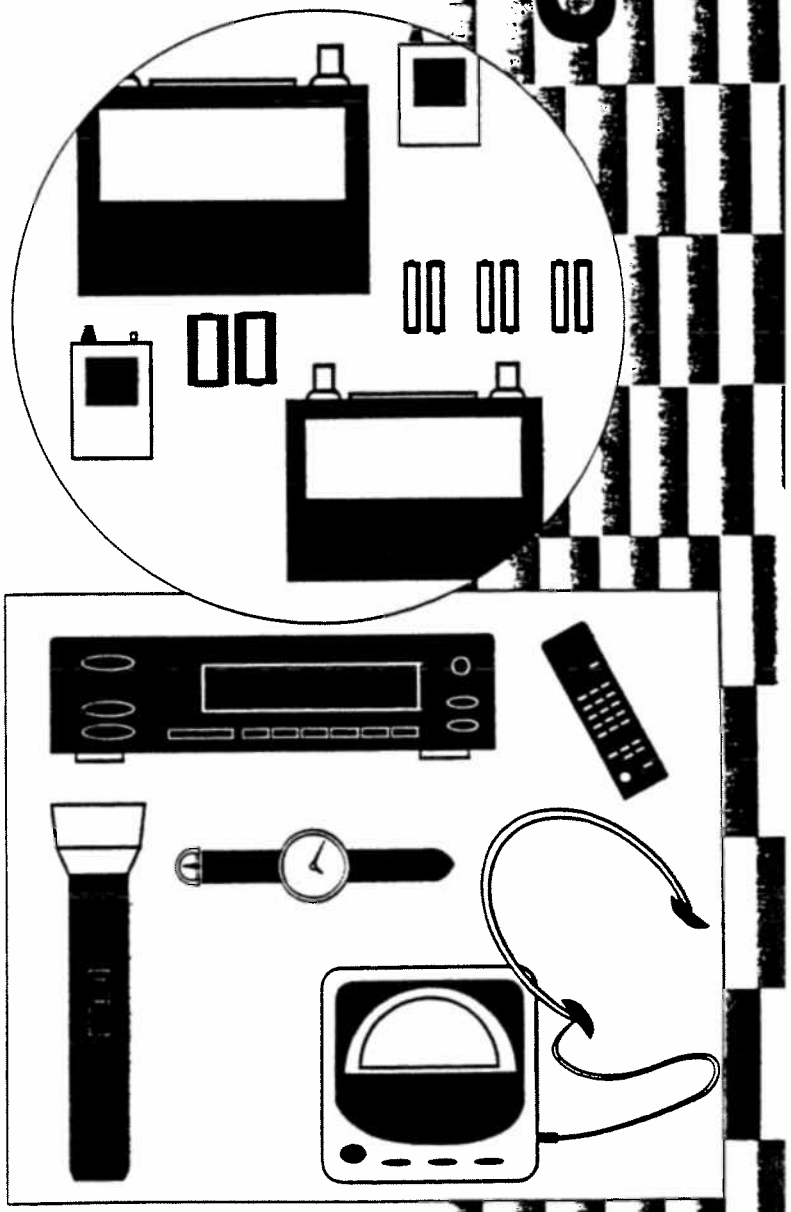
UNIT 8

Think about car engines powered by batteries instead of gasoline. Since electric cars do not produce exhaust fumes, how would this affect air pollution? Electric cars have been made for many years, but they are not widely used. A major reason for the limited use of electric cars is that the batteries used to power these cars are very heavy and don't last very long. In recent years, however, new batteries have been developed to overcome these problems.

What inside a battery produces electricity? The answer is chemical reactions that take place inside the battery. In these reactions, electrons move without any outside energy being added.

In other chemical reactions, energy must be added to make electrons move. Such reactions are used in the process of electroplating. In electroplating, metal ions such as gold and silver are plated onto less expensive metals to make jewelry and other products.

In this unit, you will learn about different kinds of electrochemical reactions and how they occur.



Lesson 40 Oxidation and Reduction

Key Words

oxidation:	the loss of electrons
reduction:	the gain of electrons
oxidizing agent:	takes or gains electrons from another substance
reducing agent:	gives or loses electrons to another substance

KEY IDEAS

The electrons moving around the nuclei of atoms may also move from one atom to another. Such movement takes place during both oxidation reactions and reduction reactions. Both types of reactions happen at the same time. The loss of electrons is called oxidation. The gain of electrons is called reduction.

Firefighters need to be aware of the processes of oxidation and reduction. When materials burn, they combine with oxygen; thus, oxidation takes place. To prevent or stop fires, firefighters must know how to keep oxygen away from materials that can burn.

Oxidation has two meanings. One meaning is "the combining of a substance with oxygen." For example, calcium combines with oxygen in the oxidation reaction shown below.

$$2\,Ca + O_2 \longrightarrow 2\,CaO$$

Loss and Gain of Electrons. Another meaning of **oxidation** (ahks-ih-DAY-shuhn) is the "loss of electrons." In the oxidation reaction shown above, the calcium loses electrons. See Fig. 40-1.

Fig. 40-1

$$Ca: \longrightarrow Ca^{2+} + 2e^-$$

Calcium atom — Calcium ion — 2 Electrons

Calcium can also react with sulfur in an oxidation reaction.

$$Ca + S \longrightarrow CaS$$

In this reaction, the calcium loses electrons just as it does in the calcium-oxygen reaction shown on page 198.

During oxidation, the electrons that are lost go to another substance. For example, electrons lost by calcium can go to sulfur. See Fig. 40-2.

Fig. 40-2

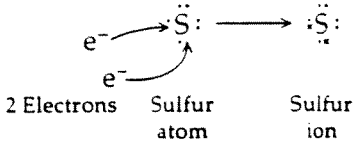
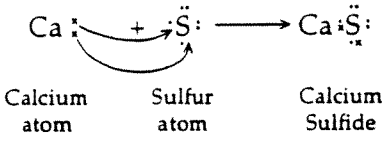


Fig. 40-3



In the calcium-sulfur reaction, shown in Fig. 40-3, the electrons that are lost by the calcium are gained by the sulfur. The gain of electrons by a substance in a reaction is called **reduction** (rih-DUK-shuhn).

Oxidizing and Reducing Agents. In the calcium-sulfur reaction, shown in Fig. 40-3, the sulfur is the **oxidizing agent** (AHKS-ih-dyz-ihng AY-juhnt) because it causes the loss of electrons. The oxidizing agent is reduced because it gains electrons.

The calcium is the **reducing agent** (rih-DOOS-ihng AY-juhnt) because it causes the gain of electrons. The reducing agent is oxidized because it loses electrons.

1. **What is oxidation?** _____

_____

2. **What is reduction?** _____

_____

Changing Charges. During oxidation reactions, charges on atoms change. Look again at the reaction of calcium and sulfur.

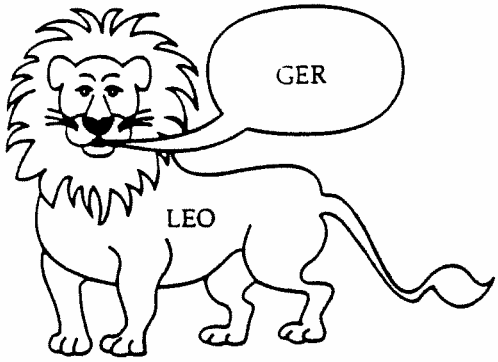
$$Ca^\circ + S^\circ \longrightarrow Ca^{2+}S^{2-}$$

Here is how the charges change: Calcium (Ca°), the reducing agent, becomes more positive (Ca^{2+}); sulfur (S°), the oxidizing agent, becomes more negative (S^{2-}).

TAKE ANOTHER LOOK

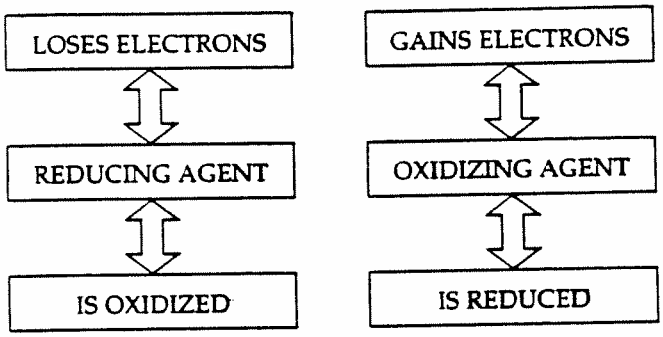
To remember what reactions involve losses and gains of electrons, think about this story. There once was a lion named **LEO**, which means Loss of Electrons is Oxidation. This king of the beasts was a little different from others. Instead of saying "roar," he said "**GER**", which means Gain of Electrons is Reduction. See Fig. 40-4.

Fig. 40-4



Look at Fig. 40-5. It shows the gain and loss of electrons by oxidizing agents and by reducing agents.

Fig. 40-5



Check Your Understanding

3. Use the terms *reducing agent*, *reduction*, and *substance oxidized* to complete the table comparing oxidation and reduction.

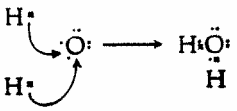
Loss of Electrons	Gain of Electrons
(a) oxidation	(a) _____
(b) _____	(b) oxidizing agent
(c) _____	(c) substance reduced

Fill in the blanks with the proper terms.

Oxidation is the (4)_____ of electrons. Reduction is the (5)_____ of electrons. The oxidizing agent (6)_____ electrons. The reducing agent (7)_____ electrons. The charge of the oxidizing agent becomes more (8)_____. The charge of the reducing agent becomes more (9)_____.

Two hydrogen atoms and one oxygen atom combine to form water as shown in Fig. 40-6.

Fig. 40-6



10. Which element is the oxidizing agent? _____

11. Which element is the reducing agent? _____

12. When magnesium burns, it combines quickly with oxygen. When iron rusts, the iron reacts slowly with oxygen. Write two equations to show that each of these reactions are examples of oxidation.

_____

_____

13. Iron ore is reduced to iron in the following reaction. Explain why carbon (C) is the reducing agent.

$$2FeO + C \longrightarrow CO_2 + 2Fe$$

14. Hydrogen will react with chlorine as shown in this equation.

$$H_2 + Cl_2 \longrightarrow 2HCl$$

Write the formula for the oxidizing agent (a)_____, the reducing agent (b)_____, the substance oxidized (c)_____, and the substance reduced (d)_____.

15. Write a paragraph to explain why oxidation and reduction both take place at the same time.

_____

_____

_____

_____

Lesson 41 Redox Reactions

Key Words

oxidation number: number given to each atom in a chemical formula to show the number of electrons that might be gained, lost, or shared during bond formation.

redox reaction: short term for an oxidation-reduction reaction

KEY IDEAS

In a redox reaction, oxidation numbers change. These numbers are used to show the direction of electron movement in the reactions. When an atom loses electrons, its oxidation number increases. When an atom gains electrons, its oxidation number decreases.

Redox reactions that take place in the body can lead to disease and aging. Antioxidants can stop or slow down harmful redox reactions. For this reason, nurses and other health care workers need to know about antioxidants present in foods and medicines.

Finding Oxidation Numbers. Electrons are gained, lost, or shared when atoms bond together. Oxidation numbers are used to keep track of electrons during bonding. It is easy to find the **oxidation number** of an atom by using the following set of rules:

The oxidation number of a one-atom ion is equal to its charge. For example, the oxidation number of calcium in Ca^{2+} is +2. The oxidation number of sulfur in S^{2-} is –2.

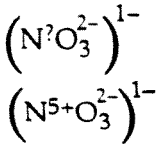
The oxidation number of an element is zero. An uncombined atom such as K or P has an oxidation number of zero. When atoms of the same element bond together, each atom also has an oxidation number of zero. Thus the oxygen atoms in O_2 and the oxygen atoms in ozone O_3 both have oxidation numbers of zero.

In compounds made up of only two elements, the more electronegative element has a negative oxidation number. The less electronegative element has a positive oxidation number. In PCl_3, chlorine is more electronegative than phosphorus. Chlorine therefore has an oxidation number of –1. Phosphorus in PCl_3 is less electronegative than chlorine. Thus, phosphorus has a charge of +3.

In compounds, hydrogen usually has an oxidation number of +1. Oxygen usually has an oxidation number of –2. In HCl, the oxidation number of hydrogen is +1. In CaO, the oxidation number of oxygen is –2.

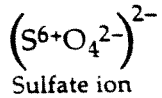
The sum of the oxidation numbers in an ion made up of many elements is equal to its charge. One example is the nitrate ion NO_3^-, shown in Fig. 41-1. In this ion, each oxygen atom has an oxidation number of –2. Three oxygen atoms have an oxidation number of –6, since 3(–2) = –6. The sum of the oxidation numbers is the charge on the ion, which is –1. That is, the oxidation number of nitrogen added to –6 should equal –1. So the oxidation number of the nitrogen must be +5.

Fig. 41-1



In the sulfate ion SO_4^{2-}, the oxidation numbers add up to –2. Look at Fig. 41-2. Each oxygen atom has an oxidation number of –2. The oxidation number of sulfur is +6 because (+6) + (4)(–2) = –2.

Fig. 41-2



Sulfate ion

The sum of the oxidation numbers in a compound is zero. In water, the oxidation number of the oxygen is –2. The oxidation number of each hydrogen is +1. The oxidation number of both hydrogens is 2(+1) = +2. The sum of –2 for the oxygen and +2 for the hydrogens is zero. In nitric acid HNO_3, the oxidation number of the hydrogen is +1, and the charge on the nitrate ion is -1.

1. **What is the oxidation number of a free element?** _____

2. **What is the usual oxidation number of oxygen?** _____

3. **What is the sum of the oxidation numbers in a compound?**

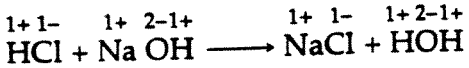
Oxidation Numbers in Reactions. A **redox reaction** is an oxidation-reduction reaction. Look at the equation shown in Fig. 41-3.

Fig. 41-3

In this reaction, the oxidation number of the bromine changes from –1 to 0. The oxidation number of the chlorine changes from 0 to –1. Each bromine atom loses an electron, which is oxidation. Each chlorine atom gains an electron, which is reduction. Thus, the reaction shown is a redox reaction.

Now look at the equation in Fig. 41-4. In this reaction, no change of oxidation numbers occurs. If none of the oxidation numbers change, no redox reaction takes place.

Fig. 41-4



TAKE ANOTHER LOOK

Look at the redox reaction between sodium (Na) and sulfur (S) shown in Figs. 41-5 and 41-6. The diagrams show the movement of electrons during the same reaction in different ways. Notice that sodium loses electrons, which is oxidation. Sulfur gains electrons, which is reduction. The oxidation number of each sodium atom increases from 0 to +1. The oxidation number of the sulfur atom decreases from 0 to –2.

Fig. 41-5

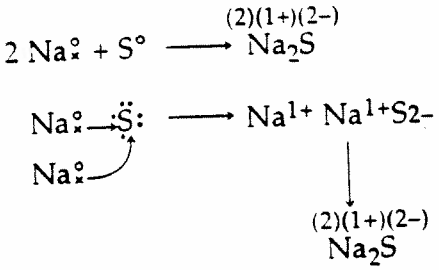
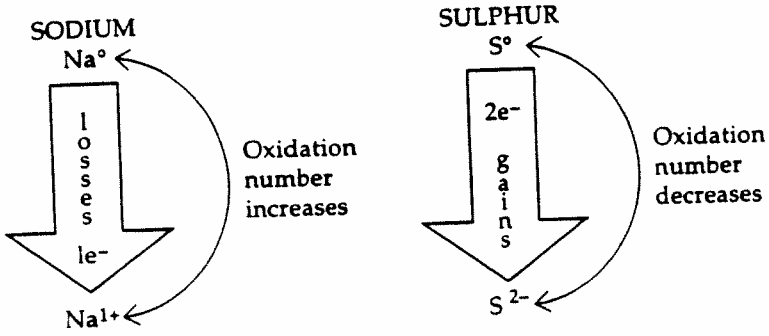


Fig. 41-6



Check Your Understanding

Fill in the blanks with the correct terms.

The rules for finding oxidation numbers are the following: The oxidation number of an ion equals **(4)**_____. The oxidation number of an element equals **(5)**_____. In two-element compounds, the more electronegative element has a **(6)**_____ charge, and the less electronegative element has a **(7)**_____ charge. Hydrogen usually has an oxidation number of **(8)**_____, and the oxidation number of oxygen is usually **(9)**_____. The sum of the oxidation numbers of an ion equals **(10)**_____. The sum of the oxidation numbers of a compound equals **(11)**_____.

Assign oxidation numbers to each element in the following unbalanced equation.

$$NH_3 + O_2 \longrightarrow NO + H_2O$$

12. The oxidation number of the nitrogen changes from _____ to _____.

13. The oxidation number of the oxygen changes from _____ to _____.

14. What is the oxidation number of each element in the following compounds?

(a) HCl _____

(b) H_2SO_4 _____

(c) $KMnO_4$ _____

(d) NO _____

15. Which of the following reactions are redox reactions? _____

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

(b) $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

(c) $Zn + CuSO_4 \longrightarrow Cu + ZnSO_4$

(d) $2HgO \longrightarrow 2Hg + O_2$

Dinitrogen tetroxide (N_2O_4) and hydrazine (N_2H_2) are used as rocket fuels. The reaction between these two compounds produces nitrogen and water, as shown below.

$$N_2O_4 + 2N_2H_4 \longrightarrow 3N_2 + 4H_2O$$

16. What two changes of oxidation number does the nitrogen undergo?

_____

_____

17. Does the oxidation number of the oxygen change? _____

_____

Redox Reactions

LEO says GER

In each of the following balanced oxidation-reduction equations, identify those elements that undergo changes in oxidation number and indicate the magnitude of the change.

1. $I_2O_{5(s)} + 5\ CO_{(g)} \longrightarrow I_{2(s)} + 5\ CO_{2(g)}$

2. $2\ Hg^{2+}_{(aq)} + N_2H_{4(aq)} \longrightarrow 2\ Hg_{(l)} + N_{2(g)} + 4\ H^{+}_{(aq)}$

3. $3\ H_2S_{(aq)} + 2\ H^{+}_{(aq)} + 2\ NO_3^{-}{}_{(aq)} \longrightarrow 3\ S_{(s)} + 2\ NO_{(g)} + 4\ H_2O_{(l)}$

Indicate whether the following balanced equations involve oxidation-reduction. If they do, identify the elements that undergo changes in oxidation number.

4. $2\ OH^{-}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} \longrightarrow 2\ CrO_4^{2-}{}_{(aq)} + H_2O_{(l)}$

5. $3\ NO_{2(g)} + H_2O_{(l)} \longrightarrow 2\ HNO_{3(aq)} + NO_{(g)}$

6. $PBr_{3(l)} + 3\ H_2O_{(l)} \longrightarrow H_3PO_{4(aq)} + 3\ HBr_{(aq)}$

7. $2\ H_2SO_{4(aq)} + 2\ NaBr_{(s)} \longrightarrow Br_{2(s)} + SO_{2(g)} + Na_2SO_{4(aq)} + 2\ H_2O_{(l)}$

8. Solid lead (II) sulfide reacts at high temperatures with oxygen in the air to form lead (II) oxide and sulfur dioxide.

(a) Write a balanced equation for this reaction.

(b) Which substances are reductants, and which are oxidants?

9. Hydrazine, N_2H_4, and dinitrogen tetroxide, form a self-igniting mixture that has been used as a rocket propellant. The reaction products are N_2 and H_2O

(a) Write a balanced equation for this reaction.

(b) Which substance serves as the reducing agent and which as the oxidizing agent?

10. Sulfuric acid functions as an oxidizing agent in (a) and as an acid in (b). How do you differentiate between these two functions?

(a) $4\ H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\ NaI_{(s)} \longrightarrow 2\ Na^{1+}_{(aq)} + I_{2(s)} + SO_{2(g)} + 2\ H_2O_{(l)}$

(b) $BaF_{2(s)} + 2\ H^+_{(aq)} + SO_4^{2-}{}_{(aq)} \longrightarrow 2\ HF_{(aq)} + BaSO_{4(s)}$

Identify the oxidizing agent and the reducing agent in each of the following balanced net ionic equations.

11. $5\ H_2C_2O_{4(aq)} + 2\ MnO_4^-{}_{(aq)} + 6\ H^+_{(aq)} \longrightarrow 2\ Mn^{2+}_{(aq)} + 10\ CO_{2(g)} + 8\ H_2O_{(l)}$

12. $3\ Cu_{(s)} + 8\ H^+_{(aq)} + 2\ NO_3^-{}_{(aq)} \longrightarrow 3\ Cu^{2+}_{(aq)} + 2\ NO_{(g)} + 4\ H_2O_{(l)}$

13. $Sn_{(s)} + 2\ H^+_{(aq)} \longrightarrow Sn^{2+}_{(aq)} + H_{2(g)}$

14. $8\ H^+_{(aq)} + 6\ Cl^-_{(aq)} + Sn_{(s)} + 4\ NO_3^-{}_{(aq)} \longrightarrow SnCl_6^{2-}{}_{(aq)} + 4\ NO_{2(g)} + 4\ H_2O_{(l)}$

15. $2\ MnO_4^-{}_{(aq)} + 10\ Cl^-_{(aq)} + 16\ H^+_{(aq)} \longrightarrow 5\ Cl_{2(g)} + 2\ Mn^{2+}_{(aq)} + 8\ H_2O_{(l)}$

16. $8\ H^+_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + 3\ SO_3^{2-}{}_{(aq)} \longrightarrow 2\ Cr^{3+}_{(aq)} + 3\ SO_4^{2-}{}_{(aq)} + 4\ H_2O_{(l)}$

17. $NO_3^-{}_{(aq)} + 4\ Zn_{(s)} + 7\ OH^-_{(aq)} + 6\ H_2O_{(l)} \longrightarrow 4\ Zn(OH)_4^{2-}{}_{(aq)} + NH_{3(aq)}$

Redox Reactions

Complete and balance the following equations.

1. $Pb(OH)_4^{2-}{}_{(aq)} + ClO^-{}_{(aq)} \longrightarrow PbO_{2(s)} + Cl^-{}_{(aq)}$ *basic solution*

2. $Tl_2O_{3(s)} + NH_2OH_{(aq)} \longrightarrow TlOH_{(s)} + N_{2(g)}$ *basic solution*

3. $Cr_2O_7^{2-}{}_{(aq)} + CH_3OH_{(aq)} \longrightarrow HCO_2H_{(aq)} + Cr^{3+}{}_{(aq)}$ *acidic solution*

4. $MnO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \longrightarrow Mn^{2+}{}_{(aq)} + Cl_{2(g)}$ *acidic solution*

5. $NO_2^-{}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} \longrightarrow Cr^{3+}{}_{(aq)} + NO_3^-{}_{(aq)}$ *acidic solution*

6. $Cr_2O_7^{2-}{}_{(aq)} + I^-{}_{(aq)} \longrightarrow Cr^{3+}{}_{(aq)} + IO_3^-{}_{(aq)}$ *acidic solution*

7. $MnO_4^-{}_{(aq)} + Br^-{}_{(aq)} \longrightarrow MnO_{2(s)} + BrO_3^-{}_{(aq)}$ (basic)