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20 Oxidation-Reduction Reactions

Big idea REACTIONS

20.1 The Meaning of Oxidation and Reduction

For students using the Foundation edition, assign problems 2–6, 8, 10–16, 18, 19.

Essential Understanding Oxidation and reduction are opposite chemical processes during which electrons are shifted between reactants.

Reading Strategy

Compare and Contrast Organizing information in a table helps you compare and contrast several topics at one time. Consider comparing and contrasting oxidation and reduction and how they relate to oxygen and electrons. As you read, ask yourself, "How are they similar? How are they different?"

As you read Lesson 20.1, use the compare and contrast table below. Complete the table, comparing how oxidation and reduction relate to electrons and oxygen.

Oxidation and Reduction		
	Compare	**Contrast**
Oxygen	***If oxygen is involved in the process, a change occurs in the number of oxygen atoms.***	***During oxidation, a substance gains oxygen. During reduction, a substance loses oxygen.***
Electrons	***A change occurs in the number of electrons.***	***During oxidation, a substance loses one or more electrons. During reduction, a substance gains one or more electrons.***

EXTENSION Use the information about electrons in the chart to explain why oxidation and reduction always occur together.

When electrons are gained by one substance, the electrons must come from another substance, which loses the electrons.

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Lesson Summary

What Are Oxidation and Reduction? Oxidation and reduction are chemical processes that occur together and are called "redox" reactions.

- Oxidation involves losing electrons and, usually, gaining oxygen.
- Reduction involves gaining electrons and, usually, losing oxygen.
- A reducing agent loses electrons and an oxidizing agent accepts electrons.
- In redox reactions involving covalent compounds, electrons shift but do not completely transfer.

Corrosion Corrosion of metals is caused by a redox reaction.

- The presence of salts and acids can speed up corrosion.
- Metals that are resistant to losing electrons corrode slowly, or not at all.
- Some corrosion, such as rust, damages metal objects; other corrosion, such as the formation of a patina on copper, does not have negative effects on the metal.

After reading Lesson 20.1, answer the following questions.

What are Oxidation and Reduction?

1. What was the original meaning of the term *oxidation*?

Oxidation originally meant the combination of an element with oxygen to produce an oxide.

2. Circle the letter of each sentence that is true about oxidation.

(a.) Gasoline, wood, and natural gas (methane) can all burn in air, producing oxides of carbon.

b. All oxidation processes involve burning.

(c.) Bleaching is an example of oxidation.

(d.) Rusting is an example of oxidation.

3. Look at Figures 20.1 and 20.2. Describe what is happening in each chemical reaction.

a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Methane and oxygen quickly combine to produce carbon dioxide and water.

b. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Iron atoms slowly combine with oxygen in moist air to produce compounds such as iron(III) oxide.

4. What is the name of the process that is the opposite of oxidation?

reduction

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5. Circle the letter of each sentence that is true about oxidation and reduction.

a. Oxidation never occurs without reduction, and reduction never occurs without oxidation.

b. You need to add heat in order to reduce iron ore to produce metallic iron.

c. When iron oxide is reduced to metallic iron, it gains oxygen.

d. Oxidation-reduction reactions are also known as redox reactions.

6. What substance is heated along with iron ore in order to reduce the metal oxide to metallic iron?

carbon

7. Look at the chemical equation for the reduction of iron ore on page 693. When iron ore is reduced to metallic iron, what oxidation reaction occurs at the same time?

Carbon combines with oxygen from the iron(III) oxide to form carbon dioxide.

8. Is the following sentence true or false? The concepts of oxidation and reduction have been extended to include many reactions that do not even involve oxygen.

true

9. What is understood about electrons in redox reactions?

Redox reactions involve a shift of electrons between reactants.

10. In the table below, fill in either "gain" or "loss" to correctly describe what happens to electrons and oxygen during oxidation and reduction.

	Oxidation	Reduction
Electrons	*loss*	*gain*
Oxygen	*gain*	*loss*

11. Look at Figure 20.3. Circle the letter of each sentence that is true about the reaction of magnesium and sulfur.

a. When magnesium and sulfur are heated together, they undergo a redox reaction to form magnesium sulfide.

b. Electrons are transferred from the metal atoms to the nonmetal atoms in this reaction.

c. When magnesium atoms lose electrons and sulfur atoms gain electrons, the atoms become less stable.

d. Magnesium is the oxidizing agent and sulfur is the reducing agent in this reaction.

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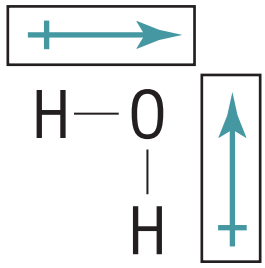
12. Is the following sentence true or false? In any redox reaction, complete electron transfer must occur. ***false***

13. Is the following sentence true or false? A redox reaction might produce covalent compounds. ***true***

14. Draw arrows showing the shift of bonding electrons during formation of a water molecule. Then complete the table listing the characteristics of this reaction.



Formation of Water by Reaction of Hydrogen and Oxygen	
Chemical equation	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
Shift of bonding electrons	***away from hydrogen and toward oxygen***
Reduced element	***oxygen***
Oxidized element	***hydrogen***
Reducing agent	***hydrogen***
Oxidizing agent	***oxygen***
Is heat released or absorbed?	***Heat is released.***

15. For each process described below, label it *O* if it is an oxidation or *R* if it is a reduction.

___***O***___ **a.** addition of oxygen to carbon or carbon compounds

___***R***___ **b.** removal of a metal from its ore

___***R***___ **c.** complete gain of electrons in an ionic reaction

___***O***___ **d.** shift of electrons away from an atom in a covalent bond

___***R***___ **e.** gain of hydrogen by a covalent compound

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Corrosion

16. Circle the letter of each sentence that is true about corrosion.

a. Preventing and repairing damage from corrosion of metals requires billions of dollars every year.

b. Iron corrodes by being oxidized to ions of iron by oxygen.

c. Water in the environment slows down the rate of corrosion.

d. The presence of salts and acids increases the rate of corrosion by producing conducting solutions that make the transfer of electrons easier.

17. Why are gold and platinum called noble metals?

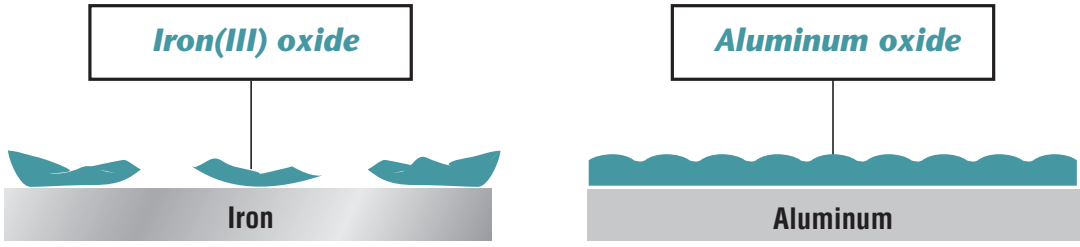
because gold and platinum are very resistant to losing their electrons through corrosion

18. Look at Figure 20.5. Why is corrosion desirable in the situation shown?

In this case, corrosion forms a patina that enhances the appearance of the buildings.

19. Look at Figure 20.6. Complete the sketch below to show how oxides form on the surface of each metal. Explain how differences between the oxides affect further corrosion of the metals.

Aluminum oxide forms a tightly packed protective layer that does not admit water. Because iron(III) oxide is not tightly packed, it allows water to penetrate and attack the metal below.



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20.2 Oxidation Numbers

For students using the Foundation edition, assign problems 1–5.

Essential Understanding An oxidation number is a positive or a negative number assigned to an atom to indicate how oxidized or reduced it is.

Lesson Summary

Assigning Oxidation Numbers Oxidation numbers are assigned according to established rules.

- If an atom is bonded to another atom, its oxidation number is the charge it would have if all the electrons in the bond were assigned to the more electronegative atom.
- The oxidation number of a monatomic ion is the same as its charge, in sign and magnitude.
- In binary ionic compounds, the oxidation number of each atom is its ionic charge.
- The oxidation number of an uncombined atom is 0.
- Some elements have only one oxidation number, but some elements have multiple oxidation numbers.
- The sum of the oxidation numbers of the atoms in a neutral compound is 0.
- The sum of the oxidation numbers in a polyatomic ion equals the ionic charge of the ion.

Oxidation-Number Changes in Chemical Reactions Oxidation numbers change any time an atom is oxidized or reduced.

- If the oxidation number increases, the atom or ion is oxidized.
- If the oxidation number decreases, the atom or ion is reduced.

After reading Lesson 20.2, answer the following questions.

Assigning Oxidation Numbers

1. Is the following sentence true or false? 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 more electronegative element. ***true***

2. For each binary ionic compound listed in the table, write the symbols for both ions, their ionic charges, and their oxidation numbers.

Compound	Ions	Ionic charges	Oxidation numbers
NaCl	Na^+ Cl^-	1+ 1−	+1 −1
CaF_2	Ca^{2+} F^-	2+ 1−	+2 −1

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3. Is the following sentence true or false? Even though water is a molecular compound, you can still obtain oxidation numbers for the bonded elements by imagining that the electrons contributed by the hydrogen atoms are completely transferred to oxygen.
true

4. Write the oxidation number, or the sum of the oxidation numbers, for the given atoms, ions, or compounds.

+2 **a.** Cu(II) ion

+1 **b.** Hydrogen in water

−1 **c.** Hydrogen in sodium hydride (NaH)

0 **d.** Potassium sulfate (K_2SO_4)

Oxidation-Number Changes in Chemical Reactions

5. Label each change *O* if it describes oxidation or *R* if it describes reduction.

R **a.** Decrease in the oxidation number of an element

O **b.** Increase in the oxidation number of an element

20.3 Describing Redox Equations

For students using the Foundation edition, assign problems 2–4, 5, 7.

Essential Understanding As with all chemical reactions, redox reactions can be described by balanced chemical equations.

Lesson Summary

Identifying Redox Reactions

Redox reactions are those during which an electron transfer occurs.

- Single-replacement, combination, decomposition, and combustion reactions are redox reactions, and double-replacement and acid-base reactions are not redox reactions.
- Redox reactions can be identified by a change in the oxidation number of elements during the reaction.
- A color change sometimes indicates a redox reaction.

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Balancing Redox Equations As with other chemical equations, the chemical equations that describe redox reactions must be balanced.

- Redox equations can be balanced using the oxidation-number-change method or the half-reaction method.
- In the oxidation-number-change method, an equation is balanced by comparing the increases and decreases in oxidation number.
- A half-reaction is an equation showing either an oxidation or reduction that takes place in a redox reaction.
- In the half-reaction method, the half-reactions are balanced and then combined into a balanced redox equation.

After reading Lesson 20.3, answer the following questions.

Identifying Redox Reactions

1. Name two kinds of reactions that are not redox reactions.

Double-replacement reactions and acid-base reactions are not redox reactions.

2. Look at Figure 20.11b. Write the oxidation numbers of all the elements in the reactants and products. Then answer the questions about the reaction.

	Reactants			Products			
	Zinc		Hydrochloric acid	Zinc chloride		Hydrogen	
Oxidation numbers	***0***		***+1 −1***	***+2 −1***		***0***	
Chemical equation	$Zn(s)$	+	$2HCl(aq)$	$\rightarrow$	$ZnCl_2(aq)$	+	$H_2(g)$

a. Is this a redox reaction? ***yes***

b. Which element is oxidized? How do you know?

Zinc; the oxidation number for zinc increases from 0 to +2.

c. Which element is reduced? How do you know?

Hydrogen; the oxidation number for hydrogen decreases from +1 to 0.

3. When a solution changes color during a reaction, what can you conclude about that reaction?

The reaction could be a redox reaction.

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Balancing Redox Equations

4. Answer these questions to help you balance the following equation using the oxidation-number-change method.

$$H_2(g) + O_2(g) \rightarrow H_2O(l) \text{ (unbalanced)}$$

a. What are the oxidation numbers for each atom in the equation?

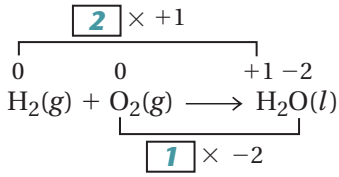
0 0 +1 −2

$$H_2(g) + O_2(g) \rightarrow H_2O$$

b. Which element is oxidized in this reaction? Which is reduced?

Hydrogen is oxidized; oxygen is reduced.

c. Use your answers to question 4*a* above to balance the equation. Write the coefficients needed to make the total change in oxidation number equal to 0.



d. What is the final balanced equation?

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

5. The equations for which reactions are balanced separately when using the half-reaction method?

The oxidation and reduction equations are balanced separately.

6. For what kind of reaction is the half-reaction method particularly useful?

ionic reactions in solution

7. When would you choose to use oxidation-number changes to balance an equation?

when the oxidized and reduced species appear only once on each side of the equation and no acids or bases are involved

8. What method would you choose to balance an equation for a reaction that takes place in an acidic or alkaline solution?

the half-reaction method

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Guided Practice Problems

Answer the following questions about Practice Problem 10.

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^-

S_2O_3

Step 1. What is the oxidation number for oxygen? Use Rule 3. *−2*

Step 2. What is the oxidation number for all of the oxygen atoms? *−2 × 3 = −6*

Step 3. What is the oxidation number for all of the sulfur atoms? *+6*

Step 4. What is the oxidation number for each sulfur atom? $\frac{+6}{2} = +3$

Step 5. How do you know your answers are correct?

The total of the oxidation numbers for all atoms must equal zero.

Na_2O_2

Step 1. What is the oxidation number of oxygen? Use Rule 3. (Hint: This compound is a peroxide.) *−1*

Step 2. What is the oxidation number of sodium? *+1*

Step 3. How do you know your answers are correct?

The oxidation number for both oxygen atoms is *−2*. The sum of the oxidation numbers for all the atoms must be *zero*. Therefore, the oxidation number for both sodium atoms must equal *+2*.

P_2O_5

Step 1. What is the oxidation number for oxygen? Use Rule 3. *−2*

Step 2. What is the oxidation number for all of the oxygen atoms? $-2 \times 5 =$ *−10*

Step 3. What is the oxidation number for all of the phosphorous atoms? *+10*

Step 4. What is the oxidation number for each phosphorous atom? $\frac{+10}{2} =$ *+5*

Step 5. How do you know your answers are correct?

The total of the oxidation numbers for all of the atoms must be zero.

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NO_3^-

Step 1. What is the oxidation number for oxygen? Use Rule 3. *−2*

Step 2. What is the oxidation number for all of the oxygen atoms? $-2 \times 3 =$ *−6*

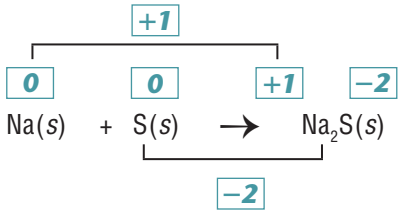
Step 3. What is the oxidation number for the nitrogen atom? *+5*

Step 4. How do you know your answers are correct?

The total of the oxidation numbers for all of the atoms must equal the net charge of the ion.

Extra Practice

Balance this redox equation using the oxidation-number-change method.



To balance the oxidation numbers, you must multiply the oxidation number of sodium by *2*. Add a coefficient of *2* in front of elemental sodium, but not in front of sodium sulfide because the sodium in sodium sulfide has a subscript of *2*.

The balanced equation is *$2Na(s) + S(s) \rightarrow Na_2S(s)$*.

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Apply the Big idea

Aluminum reacts with oxygen in the air to form aluminum oxide.

a. Write the unbalanced chemical equation for this reaction.

$Al(s) + O_2(g) \rightarrow Al_2O_3(s)$

b. What is oxidized?

$Alg(s)$

c. What is reduced?

$O_2(g)$

d. What is the oxidizing agent?

$O_2(g)$

e. What is the reducing agent?

$Al(s)$

f. Use the oxidation-number-change method to balance the equation. Show your work.

$+3 \times 2 = +6$

$0 \quad 0 \quad +3 \; -2$

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

$-2 \times 3 = -6$

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20 Self-Check Activity

For Questions 1–8, complete each statement by writing the correct word or words. If you need help, you can go online.

20.1 The Meaning of Oxidation and Reduction

1. Oxidation and ***reduction*** always occur at the same time.

2. Losing electrons is ***oxidation*** and gaining electrons is ***reduction***.

3. The metal iron ***corrodes*** when it reacts with oxygen to form rust.

20.2 Oxidation Numbers

4. The oxidation number of a bonded atom is the charge the atom would have if all the electrons in the bond went to the ***more electronegative*** element.

5. An increase in oxidation number indicates ***oxidation***, and a decrease in oxidation number indicates ***reduction***.

20.3 Describing Redox Equations

6. A reaction must be a(n) ***redox*** reaction if the oxidation number of any reactant species in the reaction changes when products form.

7. In a balanced chemical equation for a redox reaction, the total ***increase*** in oxidation number of one reactant species must equal the total ***decrease*** in oxidation number of another reactant species.

8. When balancing a redox reaction using ***half-reactions***, use separate equations for the oxidation and reduction parts of the reaction.

If You Have Trouble With…								
Question	1	2	3	4	5	6	7	8
See Page	693	694	697	701	704	708	710	712

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Review Vocabulary

Fill in each blank in the following paragraphs with a vocabulary term from this chapter.

When hydrogen burns, it combines with oxygen, and water forms. In this reaction, the (1) ***oxidation number*** of oxygen changes from 0 to −2, and that of hydrogen changes from 0 to +1. Because of this change, you know the reaction is a(n) (2) ***oxidation-reduction reaction***, also known as a(n) (3) ***redox reaction***. During the reaction, oxygen undergoes (4) ***reduction*** and also acts as a(n) (5) ***oxidizing agent***. At the same time, hydrogen undergoes (6) ***oxidation*** and also acts as a(n) (7) ***reducing agent***.

A chemical equation can be written for this reaction. The equation can be balanced by using the (8) ***half-reaction method***, in which increases and decreases in oxidation numbers are compared. It can also be balanced using the (9) ***oxidation-number-change method***, in which the (10) ***half-reactions*** for oxidation and for reduction are written and balanced separately.