Section Review

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.

Vocabulary

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

Part A Completion

Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

Oxidation–reduction, or 1, reactions are an important category of chemical reactions. Oxidation is considered to be any shift of electrons 2 from an atom. Reduction includes any shift of electrons 3 an atom. An oxidation reaction is always accompanied by a 4 reaction. The substance that does the oxidizing (the 5 agent) is 6. The substance that does the reducing (the 7 agent) is 8.

Part B True-False

Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

9. Reduction is the complete or partial gain of electrons by a substance. __________

10. In the reaction 2Na + Cl₂ → 2NaCl, sodium is the reducing agent. __________

11. In the reaction 2Na + Cl₂ → 2NaCl, sodium is being reduced. __________
12. To protect an iron ship hull, you should attach a metal that is easily reduced.

**Part C Matching**

*Match each description in Column B to the correct term in Column A.*

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>13. combustion</td>
<td>a. a metal that loses electrons easily</td>
</tr>
<tr>
<td>14. oxidation</td>
<td>b. complete or partial loss of electrons or gain of oxygen</td>
</tr>
<tr>
<td>15. oxidizing agent</td>
<td>c. oxidation of metals to metallic ions by oxygen and water in the environment</td>
</tr>
<tr>
<td>16. corrosion</td>
<td>d. a metal that resists corrosion</td>
</tr>
<tr>
<td>17. zinc</td>
<td>e. a chemical change in which oxygen reacts with another substance, often producing energy in the form of heat and light</td>
</tr>
<tr>
<td>18. gold</td>
<td>f. a substance that accepts electrons in a redox reaction</td>
</tr>
</tbody>
</table>

**Part D Questions and Problems**

*Answer the following in the space provided.*

19. Define *oxidation* and *reduction* in terms of the loss or gain of electrons.

20. In the equation given, identify the substance oxidized, the substance reduced, the oxidizing agent, and the reducing agent.

\[ \text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu} \]

21. Explain how putting a block of zinc or aluminum on the iron hull of a large ship will protect the ship from corrosion.
Section Review

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

Vocabulary
• oxidation number

Part A Completion
Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

The oxidation number of an element in an uncombined state is 1. The oxidation number of a monatomic ion is the same in magnitude and 2. as its ionic 3. The sum of the oxidation numbers of the elements in a neutral compound is 4. In a polyatomic ion, however, the sum is equal to the 5. Oxidation numbers help you keep track of 6. transfer in redox reactions. An oxidation number increase is 7. , while a 8. is reduction.

Part B True-False
Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

9. Oxygen is more electronegative than chlorine.
10. The oxidation number of each oxygen atom in most compounds is −2.
11. The oxidation number of Cl in KClO₃ is −1.
12. The oxidation number of each hydrogen atom in most compounds is −1.
13. The oxidation number for copper in a copper penny is +2.
14. In the reaction $\text{C} + \text{H}_2\text{O} \rightarrow \text{CO} + \text{H}_2$, the oxidation number of the hydrogen doesn't change.

15. In the reaction $\text{C} + \text{H}_2\text{O} \rightarrow \text{CO} + \text{H}_2$, the oxidation number of the carbon increases.

16. An increase in the oxidation number of an atom indicates oxidation.

**Part C Matching**

Match the oxidation number of nitrogen in each formula in Column B to the correct oxidation number in Column A.

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>17. -3</td>
<td>a. N$_2$</td>
</tr>
<tr>
<td>18. -2</td>
<td>b. HNO$_3$</td>
</tr>
<tr>
<td>19. -1</td>
<td>c. NO</td>
</tr>
<tr>
<td>20. 0</td>
<td>d. NH$_2$OH</td>
</tr>
<tr>
<td>21. +1</td>
<td>e. NH$_3$</td>
</tr>
<tr>
<td>22. +2</td>
<td>f. N$_2$O$_3$</td>
</tr>
<tr>
<td>23. +3</td>
<td>g. N$_2$O</td>
</tr>
<tr>
<td>24. +4</td>
<td>h. N$_2$H$_4$</td>
</tr>
<tr>
<td>25. +5</td>
<td>i. NO$_2$</td>
</tr>
</tbody>
</table>

**Part D Questions and Problems**

Answer the following in the space provided.

26. Define *oxidation* and *reduction* in terms of a change in oxidation number.

27. Use the change in oxidation number to determine which elements are oxidized and which are reduced in these reactions. (Note: It is not necessary to use balanced equations.)

   a. $\text{HNO}_3 + \text{HBr} \rightarrow \text{NO} + \text{Br}_2 + \text{H}_2\text{O}$
   b. $\text{KMnO}_4 + \text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + \text{H}_2\text{O} + \text{KCl}$
   c. $\text{Sb} + \text{HNO}_3 \rightarrow \text{Sb}_2\text{O}_5 + \text{NO} + \text{H}_2\text{O}$
   d. $\text{C} + \text{H}_2\text{SO}_4 \rightarrow \text{CO}_2 + \text{SO}_2 + \text{H}_2\text{O}$
**Section Review**

**Objectives**
- Balance a redox equation using the oxidation-number-change method
- Balance a redox equation by breaking a redox equation into oxidation and reduction half-reactions and then using the half-reaction method

**Vocabulary**
- oxidation-number-change method
- half-reaction
- half-reaction method

**Part A Completion**

Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

One method for balancing redox equations involves determining the change in oxidation number of the substances that are oxidized and reduced. Coefficients are then used to make the increase in oxidation number equal to the decrease.

The **2** method is another way to write a **3** equation for a redox reaction. In this method, the net **4** equation is divided into **5** half-reactions. Each half-reaction is balanced independently. Finally, the half-reactions are **6**.

The half-reaction method is particularly useful in balancing equations for **7** reactions.

**Part B True-False**

Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

- **8.** The reduction half-reaction in the reaction $\text{MnO}_4^- + Cl^- \rightarrow \text{Mn}^{2+} + Cl_2$ involves $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$
9. In an oxidation half-reaction, electrons occur on the right side of the equation.


11. \(2e^- + 2Cl^- \rightarrow Cl_2\) is a balanced half-reaction.

12. To balance the oxygen in a half reaction involving \(MnO_4^- \rightarrow Mn^{2+}\), \(2H_2O\) will be added to the product side of the equation.

13. In the equation \(2FeBr_2 + Br_2 \rightarrow 2FeBr_3\), the oxidation number of the iron doesn’t change.

**Part C Matching**

*Match each description in Column B to the correct term in Column A.*

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>14. half-reaction method</td>
<td>a. ions that are present but do not participate in or change during the reaction</td>
</tr>
<tr>
<td>15. spectator ions</td>
<td>b. (Fe^{2+} \rightarrow Fe^{3+} + e^-)</td>
</tr>
<tr>
<td>16. anions</td>
<td>c. balancing a redox equation by first balancing the oxidation and reduction half-reactions</td>
</tr>
<tr>
<td>17. oxidation half-reaction</td>
<td>d. balancing a redox equation by comparing the increase and decrease in oxidation numbers</td>
</tr>
<tr>
<td>18. half-reaction</td>
<td>e. equation showing either the reduction or the oxidation of a species in an oxidation-reduction reaction</td>
</tr>
<tr>
<td>19. oxidation-number-change method</td>
<td>f. ions that can serve as reducing agents</td>
</tr>
<tr>
<td>20. reduction half-reaction</td>
<td>g. (2e^- + Br_2 \rightarrow 2Br^-)</td>
</tr>
</tbody>
</table>

**Part D Questions and Problems**

*Answer the following in the space provided.*

   
   a. \(HNO_3(aq) + HI(g) \rightarrow NO(g) + I_2(s) + H_2O\)
   
   b. \(HNO_3(aq) + I_2(s) \rightarrow HIO_3(aq) + NO_2(g) + H_2O(l)\)

22. Balance these redox equations using the half-reaction method.
   
   a. \(H_2S(aq) + HNO_3(aq) \rightarrow S(s) + NO(g) + H_2O(l)\)
   
   b. \(Fe^{2+} + Cr_2O_7^{2-} \rightarrow Fe^{3+} + Cr^{3+}\)
Practice Problems

In your notebook, solve the following problems.

SECTION 20.1 THE MEANING OF OXIDATION AND REDUCTION

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

1. $2\text{Sr} + \text{O}_2 \rightarrow 2\text{SrO}$
2. $2\text{Li} + \text{S} \rightarrow 2\text{Li}_2\text{S}$
3. $2\text{Cs} + \text{Br}_2 \rightarrow 2\text{CsBr}$
4. $3\text{Mg} + \text{N}_2 \rightarrow \text{Mg}_3\text{N}_2$
5. $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$
6. $\text{Cl}_2 + 2\text{NaBr} \rightarrow 2\text{NaCl} + \text{Br}_2$
7. $\text{Si} + 2\text{F}_2 \rightarrow \text{SiF}_4$
8. $2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO}$
9. $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$
10. $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$

SECTION 20.2 OXIDATION NUMBERS

1. Give the oxidation number of each kind of atom or ion.
   - a. Sn
   - b. K$^+$
   - c. S$^{2-}$
   - d. Fe$^{3+}$
   - e. Se
   - f. Mg$^{2+}$
   - g. Sn$^{4+}$
   - h. Br$^-$

2. Calculate the oxidation number of chromium in each of the following formulas.
   - a. $\text{Cr}_2\text{O}_3$
   - b. $\text{H}_2\text{Cr}_2\text{O}_7$
   - c. $\text{CrSO}_4$
   - d. $\text{CrO}_4^{2-}$

3. Use the changes in oxidation number to determine which elements are oxidized and which are reduced in these reactions. (Note: It is not necessary to use balanced reactions.)
   - a. $\text{C} + \text{H}_2\text{SO}_4 \rightarrow \text{CO}_2 + \text{SO}_2 + \text{H}_2\text{O}$
   - b. $\text{HNO}_3 + \text{HI} \rightarrow \text{NO} + \text{I}_2 + \text{H}_2\text{O}$
   - c. $\text{KMnO}_4 + \text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + \text{H}_2\text{O} + \text{KCl}$
   - d. $\text{Sb} + \text{HNO}_3 \rightarrow \text{Sb}_2\text{O}_5 + \text{NO} + \text{H}_2\text{O}$

4. For each reaction in problem 3 above, identify the oxidizing agent and reducing agent.
SECTION 20.3 BALANCING REDOX EQUATIONS

1. Balance these equations using the oxidation-number-change method.
   a. C + H₂SO₄ → CO₂ + SO₂ + H₂O
   b. H₂S + HNO₃ → S + NO + H₂O
   c. HNO₃ + HI → NO + I₂ + H₂O
   d. Sb + HNO₃ → Sb₂O₅ + NO + H₂O
   e. KMnO₄ + HCl → MnCl₂ + Cl₂ + H₂O + KCl
   f. KIO₄ + KI + HCl → KCl + I₂ + H₂O
   g. Zn + Cr₂O₇²⁻ + H⁺ → Zn²⁺ + Cr³⁺ + H₂O

2. Write half-reactions for the oxidation and reduction processes for each of the following reactions.
   a. Fe²⁺ + MnO₄⁻ → Fe³⁺ + Mn²⁺ (acidic solution)
   b. Sn²⁺ + IO₃⁻ → Sn⁴⁺ + I⁻ (acidic solution)
   c. S²⁻ + NO₃⁻ → S + NO (acidic solution)
   d. Mn²⁺ + H₂O₂ → MnO₂ + H₂O (basic solution)

3. Balance these reactions using the half-reaction method.
   a. Zn + HgO → ZnO₂²⁻ + Hg (basic solution)
   b. Fe²⁺ + MnO₄⁻ → Fe³⁺ + Mn²⁺ (acidic solution)
   c. Sn²⁺ + IO₃⁻ → Sn⁴⁺ + I⁻ (acidic solution)
   d. S²⁻ + NO₃⁻ → S + NO (acidic solution)
   e. Mn²⁺ + H₂O₂ → MnO₂ + H₂O (basic solution)
   f. CrO₂ + ClO⁻ → CrO₄²⁻ + Cl⁻ (basic solution)
To determine the relative amount of iron in a sample of iron ore, a chemist dissolved 2.938 g of the ore in 50.0 mL of dilute sulfuric acid (H₂SO₄) in a reaction flask. The colorless solution was then titrated to the end point with potassium permanganate. The half-reactions for the oxidation and reduction processes that occur during this titration are:

\[
\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} \\
\text{MnO}_4^- \rightarrow \text{Mn}^{2+}
\]

Use the data in Table 1 and what you have learned about oxidation-reduction reactions to answer the following questions.
Table 1 Analysis of an Unknown Iron-Containing Ore

<table>
<thead>
<tr>
<th>Initial Volume of KMnO₄</th>
<th>48.65 mL</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final Volume of KMnO₄</td>
<td>23.35 mL</td>
</tr>
<tr>
<td>Volume of MnO₄⁻</td>
<td></td>
</tr>
<tr>
<td>Moles MnO₄⁻</td>
<td></td>
</tr>
<tr>
<td>Moles Iron(II), Fe²⁺</td>
<td></td>
</tr>
<tr>
<td>Mass of Iron</td>
<td></td>
</tr>
<tr>
<td>% of Iron in Ore</td>
<td></td>
</tr>
</tbody>
</table>

1. Match each component from the following list with the correct number shown in Figure 1. The same number may be used more than once.

   _____  a. oxidizing agent
   _____  b. reducing agent
   _____  c. standard solution of 0.0200M KMnO₄
   _____  d. acidic solution of iron(II) ion, Fe²⁺
   _____  e. reaction flask
   _____  f. buret

2. Use the half-reaction method to balance the equation for the redox reaction between permanganate ion and iron(II) ion. Write the net ionic equation only.

3. Explain what the end point of this particular titration means in terms of the reacting species in solution. How does the chemist recognize the end point when it occurs?

   _____________________________________________
   _____________________________________________

4. Use the stoichiometry of the balanced equation given in your answer to question 2 and the fact that the molar mass of Fe is 55.85 g to complete Table 1 above. Use the space below to show your work.
Vocabulary Review

Select the term from the following list that best matches each description.

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>half-reaction</td>
<td>oxidation-reduction reaction</td>
</tr>
<tr>
<td>half-reaction method</td>
<td>oxidizing agent</td>
</tr>
<tr>
<td>oxidation</td>
<td>redox reaction</td>
</tr>
<tr>
<td>oxidation number</td>
<td>reducing agent</td>
</tr>
<tr>
<td>oxidation-number-change method</td>
<td>reduction</td>
</tr>
<tr>
<td>1. the substance in a redox reaction that accepts electrons</td>
<td></td>
</tr>
<tr>
<td>2. a method of balancing a redox equation by comparing the increases and decreases in oxidation numbers</td>
<td></td>
</tr>
<tr>
<td>3. a process that involves a complete or partial gain of electrons or the loss of oxygen; it results in a decrease in the oxidation number of an atom</td>
<td></td>
</tr>
<tr>
<td>4. a method for balancing a redox equation by balancing the oxidation and reduction half-reactions</td>
<td></td>
</tr>
<tr>
<td>5. a positive or negative number assigned to a combined atom according to a set of arbitrary rules</td>
<td></td>
</tr>
<tr>
<td>6. a substance in a redox reaction that donates electrons</td>
<td></td>
</tr>
<tr>
<td>7. an equation showing either the reduction or the oxidation of a species in an oxidation-reduction reaction</td>
<td></td>
</tr>
<tr>
<td>8. a reaction that involves the transfer of electrons between reactants during a chemical change</td>
<td></td>
</tr>
<tr>
<td>9. a process that involves complete or partial loss of electrons or a gain of oxygen; it results in an increase in the oxidation number of an atom</td>
<td></td>
</tr>
<tr>
<td>10. another name for an oxidation-reduction reaction</td>
<td></td>
</tr>
</tbody>
</table>
Chapter Quiz

Choose the best answer and write its letter on the line.

1. The oxidation number of sulfur in each of the following is +6 except for 20.2
   a. SO₃²⁻  
   b. S₂O₃²⁻  
   c. SO₄²⁻  
   d. Na₂SO₄

2. Reduction is 20.1
   a. a gain of electrons.  
   b. a loss of electrons.  
   c. a gain of oxygen.  
   d. both a and c

3. Identify the oxidizing agent in the following reaction. 20.1
   2Na + S → Na₂S
   a. Na  
   b. S  
   c. Na₂S  
   d. Na⁺

4. From the unbalanced equations below, identify the one that does not represent a redox reaction. 20.1
   a. HNO₃(aq) + H₃PO₃(aq) → NO(g) + H₃PO₄(aq) + H₂O(l)  
   b. H₂SO₄(aq) + NaOH(aq) → H₂O(l) + Na₂SO₄(aq)  
   c. C(s) + O₂(g) → CO₂(g)  
   d. H₂O₂(aq) + PbS(s) → PbSO₄(s) + H₂O(l)

5. Identify the oxidation half-reaction among the following. 20.3
   a. Fe²⁺ → Fe³⁺ + e⁻  
   b. Cl₂ + 2e⁻ → 2Cl⁻  
   c. O₂ + 4H⁺ + 4e⁻ → 2H₂O  
   d. Fe³⁺ + e⁻ → Fe²⁺

6. What will the coefficient of HNO₃ be when the following equation is completely balanced using the smallest whole-number coefficients? 20.3
   HNO₃ + MnCl₂ + HCl → NO + MnCl₄ + H₂O
   a. 2  
   b. 3  
   c. 6  
   d. 5

7. When the half-reactions I₂ + 2e⁻ → 2I⁻ and Na → Na⁺ + e⁻ are correctly combined, the balanced redox equation is 20.3
   a. Na + I⁻ → Na⁺ + 2I⁻  
   b. Na + I₂ → Na⁺ + 2I⁻  
   c. 2Na + I₂ → 2Na⁺ + 2I⁻  
   d. Na + I₂ + 2e⁻ → Na⁺ + 2I⁻ + e⁻

8. What is the reduction half-reaction for the following unbalanced redox equation? 20.3
   Cr₂O₇²⁻ + NH₄⁺ → Cr₂O₃ + N₂
   a. NH₄⁺ → N₂  
   b. N₂ → NH₄⁺  
   c. Cr₂O₃ → Cr₂O₇²⁻  
   d. Cr₂O₇²⁻ → Cr₂O₃

Name ___________________________ Date ___________________ Class ___________________
A. Matching

Match each term in Column A with the correct description in Column B.

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>a. half-reaction</td>
</tr>
<tr>
<td>2.</td>
<td>b. oxidation-number-change method</td>
</tr>
<tr>
<td>3.</td>
<td>c. oxidation</td>
</tr>
<tr>
<td>4.</td>
<td>d. oxidation number</td>
</tr>
<tr>
<td>5.</td>
<td>e. half-reaction method</td>
</tr>
<tr>
<td>6.</td>
<td>f. oxidation-reduction reaction</td>
</tr>
<tr>
<td>7.</td>
<td>g. spectator ion</td>
</tr>
<tr>
<td>8.</td>
<td>h. reducing agent</td>
</tr>
<tr>
<td>9.</td>
<td>i. reduction</td>
</tr>
<tr>
<td>10.</td>
<td>j. oxidizing agent</td>
</tr>
</tbody>
</table>

B. Multiple Choice

Choose the best answer and write its letter on the line.

11. Identify the oxidizing agent in the following reaction.

\[ 2Na + 2H_2O \rightarrow 2NaOH + H_2 \]

a. Na  
b. H_2O  
c. NaOH  
d. H_2

12. Identify the reducing agent in the following reaction.

\[ CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \]

a. H_2O  
b. O_2  
c. CO_2  
d. CH_4
13. Nitrogen has the same oxidation number in all of the following except
   a. $\text{NO}_3^-$.  
   b. $\text{N}_2\text{O}_5$.  
   c. $\text{NH}_4\text{Cl}$.  
   d. $\text{Ca(NO}_3)_2$.

14. Determine what happens in this reaction.
   \[ \text{S} + \text{Cl}_2 \rightarrow \text{SCl}_2 \]
   (Hint: Chlorine is the more electronegative element.)
   a. Sulfur is reduced.
   b. Chlorine is reduced.
   c. Chlorine is oxidized.
   d. Sulfur is the oxidizing agent.

15. $\text{Zn} \rightarrow \text{Zn}^{2+}$ represents
   a. oxidation.  
   b. reduction.  
   c. both a and b  
   d. neither a nor b

16. $\text{Sn}^{4+} \rightarrow \text{Sn}^{2+}$ represents
   a. oxidation.  
   b. reduction.  
   c. hydrolysis.  
   d. none of the above

17. What happens to chlorine (in $\text{ClO}_3^-$) in the following redox reaction?
   \[ \text{ClO}_3^- + \text{I}^- \rightarrow \text{Cl}^- + \text{I}_2 \]
   a. It is oxidized.
   b. Its oxidation number changes from +6 to -1.
   c. Its oxidation-number change is -6.
   d. Its oxidation-number change is +6.

18. Identify the atom that increases in oxidation number in the following redox reaction.
   \[ 2\text{MnO}_2 + 2\text{K}_2\text{CO}_3 + \text{O}_2 \rightarrow 2\text{KMnO}_4 + 2\text{CO}_2 \]
   a. C  
   b. K  
   c. Mn  
   d. O

19. Identify the reducing agent in this reaction.
   \[ \text{I}^- + \text{MnO}_4^- \rightarrow \text{I}_2 + \text{MnO}_2 \]
   a. $\text{I}^-$  
   b. $\text{MnO}_4^-$  
   c. $\text{I}_2$  
   d. $\text{MnO}_2$

20. What is the increase in oxidation number for the atom that is oxidized in the following balanced redox equation?
   \[ \text{Cr}_2\text{O}_7^{2-} + 8\text{H}^+ + 3\text{SO}_3^{2-} \rightarrow \text{Cr}^{3+} + 3\text{SO}_4^{2-} + 8\text{H}_2\text{O} \]
   a. +2  
   b. +6  
   c. -3  
   d. -6

21. To balance the oxygen and hydrogen for a redox reaction that takes place in basic solution, it is necessary to use
   a. $\text{H}_2\text{O}$ and $\text{H}^+$.  
   b. $\text{H}_2\text{O}$ only.  
   c. $\text{H}_2\text{O}$ and $\text{OH}^-$.  
   d. $\text{OH}^-$ only.

22. Which of the following is an oxidation half-reaction?
   a. $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}$  
   b. $\text{NO} + 2\text{H}_2\text{O} \rightarrow \text{NO}_3^- + 4\text{H}^+ + 3e^-$  
   c. $\text{Na}^+ + e^- \rightarrow \text{Na}$  
   d. $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$
23. What is the reduction half-reaction for the following unbalanced redox equation?

\[
\text{Cr}_2\text{O}_7^{2-} + \text{Fe}^{2+} \rightarrow \text{Cr}^{3+} + \text{Fe}^{3+}
\]

a. \(\text{Cr}^{3+} \rightarrow \text{Cr}_2\text{O}_7^{2-}\)

b. \(\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}\)

c. \(\text{Fe}^{3+} \rightarrow \text{Fe}^{2+}\)

d. \(\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}\)

24. Which atom is reduced in the following unbalanced redox equations?

\[
\text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{O} + \text{S} \rightarrow \text{KOH} + \text{Cr}_2\text{O}_3 + \text{SO}_2
\]

a. \(\text{S}\)

b. \(\text{O}\)

c. \(\text{Cr}\)

d. \(\text{K}\)

25. Identify a true statement about how to protect an iron object from corrosion.

a. Increase the amount of salt and/or acid in the water.

b. Place a gold or silver bar in contact with the iron.

c. Place a better reducing agent in contact with the iron.

d. Place a metal more easily reduced in contact with the iron.

26. Identify from the unbalanced equations below the one that does not represent a redox reaction.

a. \(\text{H}_2\text{O}_2(aq) + \text{MnO}_4^-(aq) \rightarrow \text{O}_2(g) + \text{Mn}^{2+}(aq)\)

b. \(\text{H}_2(g) + \text{N}_2(g) \rightarrow \text{NH}_3(g)\)

c. \(\text{NaCl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{NaNO}_3(aq) + \text{AgCl(s)}\)

d. \(\text{Cu}(s) + \text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag(s)}\)

C. Questions

Answer the following questions in the space provided.

27. Determine which substance is oxidized and which substance is reduced in each reaction. Identify the oxidizing agent and reducing agent in each case.

a. \(2\text{Na} + \text{Br}_2 \rightarrow 2\text{NaBr}\)

b. \(2\text{K} + \text{S} \rightarrow \text{K}_2\text{S}\)

28. Combine these two half-reactions to form a balanced redox equation.

\[
\text{Br}_2 + 2e^- \rightarrow 2\text{Br}^- \quad \text{and} \quad \text{Cr} \rightarrow \text{Cr}^{3+} + 3e^-
\]
29. Determine the oxidation number of each element in these substances.
   a. Li$_3$AlF$_6$
   b. Na$_2$O
   c. S$_8$

30. Balance the following redox equation, using either the oxidation-number-change method or the half-reaction method. Show all your work. (In using the half-reaction method, assume that the reaction occurs in aqueous acid solution.)

   Fe$_2$O$_3$ + CO $\rightarrow$ Fe + CO$_2$ (acid solution)

D. Essay

31. How are oxidation numbers determined and used?
Chapter Test B

A. Matching

Match each term in Column B with the correct description in Column A. Write the letter of the correct term on the line.

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. the substance in a redox reaction that accepts electrons</td>
<td>a. oxidation-number-change method</td>
</tr>
<tr>
<td>2. the complete or partial gain of electrons or the loss of oxygen</td>
<td>b. reducing agent</td>
</tr>
<tr>
<td>3. those ions that do not change oxidation number or composition during a reaction</td>
<td>c. oxidation-reduction reactions</td>
</tr>
<tr>
<td>4. a positive or negative number assigned to an atom according to a set of arbitrary rules</td>
<td>d. spectator ions</td>
</tr>
<tr>
<td>5. the complete or partial loss of electrons or the gain of oxygen</td>
<td>e. oxidizing agent</td>
</tr>
<tr>
<td>6. the balancing of a redox reaction by comparing the increases and decreases in oxidation numbers</td>
<td>f. reduction</td>
</tr>
<tr>
<td>7. the chemical changes that occur when electrons are transferred between reactants</td>
<td>g. oxidation number</td>
</tr>
<tr>
<td>8. a method of balancing redox reactions by balancing the oxidation and reduction half-reactions</td>
<td>h. half-reaction method</td>
</tr>
<tr>
<td>9. the substance in a redox reaction that donates electrons</td>
<td>i. oxidation</td>
</tr>
<tr>
<td>10. another name for an oxidation-reduction reaction</td>
<td>j. redox reaction</td>
</tr>
</tbody>
</table>

B. Multiple Choice

Choose the best answer and write its letter on the line.

11. Which of the following is true about oxidation reactions?
   a. Oxidation reactions are the principal source of energy on Earth.
   b. All oxidation reactions are accompanied by reduction reactions.
   c. The burning of wood in a fireplace and the metabolization of food by your body are oxidation reactions.
   d. all of the above
12. What is the oxidized substance in the following reaction?
\[ \text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2 \]
   a. Fe  
   b. HCl  
   c. FeCl₂  
   d. H₂

13. The reducing agent in the reaction described in question 12 is
   a. Fe.  
   b. HCl.  
   c. FeCl₂.  
   d. H₂.

14. What is occurring in the following reaction?
   \[ \text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl} \]
   a. H₂ is being reduced.  
   b. Cl₂ is being oxidized.  
   c. H₂ is gaining two electrons.  
   d. Cl₂ is acting as an oxidizing agent.

15. What is the oxidation number of sulfur in H₂SO₃?
   a. +1  
   b. +2  
   c. +3  
   d. +4

16. What is the usual oxidation number of oxygen in a compound?
   a. −1  
   b. −2  
   c. +1  
   d. +2

17. In the unbalanced equation below, what element is being reduced?
   \[ \text{MnO}_2 + \text{HCl} \rightarrow \text{H}_2\text{O} + \text{MnCl}_2 + \text{Cl}_2 \]
   a. Mn  
   b. O  
   c. H  
   d. Cl

18. Which of the following is an oxidation reaction?
   a. \(\text{Co}^{3+} \rightarrow \text{Co}^{2+}\)  
   b. \(\text{Cl}_2 \rightarrow \text{ClO}_3^{-}\)  
   c. \(\text{AuCl}_4^- \rightarrow \text{AuCl}_2^-\)  
   d. \(\text{Mn}^{2+} \rightarrow \text{Mn}^{3+}\)

19. Among the following, which is an oxidation-reduction reaction?
   a. \(\text{Na}_2\text{S} + \text{CaCO}_3 \rightarrow \text{CaS} + \text{Na}_2\text{CO}_3\)  
   b. \(2\text{HNO}_3 + \text{Mg(OH)}_2 \rightarrow \text{Mg(NO}_3)_2 + 2\text{H}_2\text{O}\)  
   c. \(\text{H}_2 + \text{F}_2 \rightarrow 2\text{HF}\)  
   d. \(3\text{Ba(OH)}_2 + 2\text{H}_3\text{PO}_4 \rightarrow \text{Ba}_3(\text{PO}_4)_2 + 6\text{H}_2\text{O}\)

20. Which of the following is true concerning the reaction below?
   \[ \text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{S} + \text{NO} + \text{H}_2\text{O} \]
   a. S is reduced.  
   b. H is oxidized.  
   c. N is reduced.  
   d. O is oxidized.

21. When the equation in question 20 is balanced, what is the coefficient for H₂O?
   a. 2  
   b. 4  
   c. 3  
   d. 6
22. In the equation \( \text{PbO}_2 + 4\text{HCl} \rightarrow 2\text{H}_2\text{O} + \text{PbCl}_2 + \text{Cl}_2 \), how many electrons are transferred?
   a. 1       c. 3
   b. 2       d. 4

23. The element oxidized in the reaction described in question 22 is
   a. Pb.       c. H.
   b. O.       d. Cl.

24. In the unbalanced equation given below, what is the element that is gaining electrons?
   \[ \text{HCl} + \text{MnO}_2 \rightarrow \text{MnCl}_2 + \text{H}_2\text{O} + \text{Cl}_2 \]
   a. H       c. Mn
   b. Cl       d. O

25. When the equation in question 24 is balanced, what is the coefficient for \( \text{HCl} \)?
   a. 1       c. 3
   b. 2       d. 4

26. Which of the following is true concerning redox reactions?
   a. Double-replacement reactions are always redox reactions.
   b. Single-replacement reactions may be redox reactions.
   c. Acid-base reactions are always redox reactions.
   d. all of the above

27. Identify a false statement about how to protect iron from corrosion.
   a. Coat the surface with oil, paint, or plastic.
   b. Attach a metal that is more easily reduced.
   c. Exclude air and water.
   d. Attach a metal that is a better reducing agent.

28. From the unbalanced equations below, identify the one that does not represent a redox reaction.
   a. \( \text{H}_2\text{CO}_3(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) \)
   b. \( \text{C}(s) + \text{H}_2\text{O}(g) \rightarrow \text{CO}(g) + \text{H}_2(g) \)
   c. \( \text{S}_2\text{O}_3^{2-}(aq) + \text{I}_2(s) \rightarrow \text{S}_4\text{O}_6^{2-}(aq) + \text{I}^-(aq) \)
   d. \( \text{FeBr}_2(aq) + \text{Br}_2(l) \rightarrow \text{FeBr}_3(aq) \)
C. Questions

Answer the following questions in the space provided.

29. For each of the following reactions, identify the element oxidized, the element reduced, the oxidizing agent, and the reducing agent.

<table>
<thead>
<tr>
<th></th>
<th>Oxidized</th>
<th>Reduced</th>
<th>Oxidizing Agent</th>
<th>Reducing Agent</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>K + I₂ → 2KI</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>b.</td>
<td>2Na + 2H₂O → 2NaOH + H₂</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>c.</td>
<td>H₂ + CuO → Cu + H₂O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>d.</td>
<td>Cu(NO₃)₂ + Mg → Mg(NO₃)₂ + Cu</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

30. Determine the oxidation number of each element in the following.
   a. K₂SO₄
   b. Cu(NO₃)₂
   c. HAsO₃
   d. MnO₄

31. Use the oxidation-number-change method to balance the equations given below. Show all your work.
   a. HNO₃ + Ag → AgNO₃ + NO + H₂O
   b. Br₂ + SO₂ + H₂O → H₂SO₄ + HBr
32. Use the half-reaction method to balance the equations given below. Show all your work.
   a. \( \text{HNO}_2 + \text{HI} \rightarrow \text{I}_2 + \text{NO} + \text{H}_2\text{O} \)

   b. \( \text{K}_2\text{Cr}_2\text{O}_7 + \text{FeCl}_2 + \text{HCl} \rightarrow \text{CrCl}_3 + \text{KCl} + \text{FeCl}_3 + \text{H}_2\text{O} \)

**D. Essay**

33. Explain why oxidation cannot occur without reduction, and vice versa.