EXPERIMENT 26

Introduction to Oxidation–Reduction Reactions

Performance Goal

26–1 Determine experimentally the relative strengths of a selected group of oxidizing agents.

CHEMICAL OVERVIEW

Oxidation is defined as the process in which a loss of electrons occurs; **reduction** is a gain of electrons. From a broader viewpoint, in oxidation the oxidation number of an element increases (becomes more positive, as $+3 \rightarrow +5$, or $-3 \rightarrow -1$); whereas in reduction, the oxidation number decreases (becomes more negative, as $0 \rightarrow -1$, or $+7 \rightarrow +2$).

When a metal combines chemically with a halogen to form an ionic compound, an oxidation-reduction (redox) reaction occurs. Electrons are lost by the metal and gained by the halogen. Redox reactions may be thought of as electron transfer reactions, much as acid–base reactions may be viewed as proton transfer reactions. Each redox reaction may be considered the sum of two ''half-reactions'' or half-cell reactions:

$M \rightarrow M^{2+} + 2\ e^-$ Oxidation of a metal to M^{2+} ion by losing 2 electrons,

and

$X_2 + 2\ e^- \rightarrow 2\ X^-$ Reduction of a halogen to X^- by gaining 1 electron per atom.

Addition produces the net ionic redox equation:

$$\begin{array}{r} M \rightarrow M^{2+} + 2\ e^- \\ X_2 + 2\ e^- \rightarrow 2\ X^- \\ \hline M + X_2 \rightarrow M^{2+} + 2\ X^- \end{array}$$

Observe that the number of electrons lost by the metal exactly equals the number of electrons gained by the halogen. This balance is essential in any redox reaction; there can never be a deficiency or excess of electrons. Just as an ordinary chemical equation has to be balanced, so does a redox equation. Balancing is achieved by adjusting one or both half-reactions in order to equate the number of electrons lost and gained.

Example 1

$$\begin{array}{rl} A \rightarrow A^{2+} + 2\ e^- & \text{Oxidation} \\ B + e^- \rightarrow B^- & \text{Reduction} \end{array}$$

Multiply the reduction half-reaction by 2 and add:

$$\begin{array}{rl} A \rightarrow A^{2+} + 2\ e^- & \text{Oxidation} \\ 2\ B + 2\ e^- \rightarrow 2\ B^- & \text{Reduction} \\ \hline A + 2\ B \rightarrow A^{2+} + 2\ B^- & \text{Net ionic redox equation} \end{array}$$

Example 2

$$\begin{array}{rl} A \rightarrow A^{3+} + 3\ e^- & \text{Oxidation} \\ B + 2\ e^- \rightarrow B^{2-} & \text{Reduction} \end{array}$$

Multiply the oxidation reaction by 2 and the reduction reaction by 3 to equate the electrons lost in oxidation to the electrons gained in reduction:

$$\begin{array}{rl} 2\ A \rightarrow 2\ A^{3+} + 6\ e^- & \text{Oxidation} \\ 3\ B + 6\ e^- \rightarrow 3\ B^{2-} & \text{Reduction} \\ \hline 2\ A + 3\ B \rightarrow 2\ A^{3+} + 3\ B^{2-} & \text{Net ionic redox equation} \end{array}$$

In each of the preceding examples, the first element loses electrons to the second element; that is, the first element provides the electrons that reduce the second. Thus, the first element is referred to as a **reducing agent.** By accepting electrons, the second element causes the oxidation of the first element. Hence it is called an **oxidizing agent.** A summary of these terms is presented below:

If a species	*the species undergoes*	*and is called the*
Gains electrons	Reduction	Oxidizing agent (oxidizer)
Loses electrons	Oxidation	Reducing agent (reducer)

Just as acids vary in their strength (the ease with which they release protons), so reducers vary in their strength (the ease with which they release electrons). Similarly, oxidizers have different tendencies to capture electrons, just as bases vary in their attraction for protons. A strong oxidizer (oxidizing agent) has a great affinity for electrons.

In this experiment, you will investigate the relative ease with which certain metals and halides release electrons and will thereby build a partial qualitative chart of oxidizer strengths.

NOTE: Since metals and halides release electrons, they are the reducing agents, while metal ions and elemental halogens, Cl_2, Br_2, I_2, are the oxidizing agents.

Evidence of a metal–metal ion reaction is visible when a drop of solution containing the ion is placed on the metal. To detect a reaction between a halogen and a dissolved halide, however, a nonaqueous solvent that is immiscible with water must be used.

Halogens dissolved in trichloroethane have characteristic colors. Chlorine, Cl_2, in trichloroethane is colorless; bromine ranges from tan in dilute solutions to deep red or maroon when concentrated; and iodine is pale pink to deep purple, depending on concentration. You will use trichloroethane as a solvent to determine which halogen is present after a halogen and a halide ion have been combined. The trichloroethane is not involved in any chemical change; it serves simply as a solvent.

To interpret your observations correctly, consider the two examples below:

1. Suppose that the reaction between I_2 and Br^- goes to completion.

$$I_2 + 2\ Br^- \rightarrow 2\ I^- + Br_2$$

   Because in the reaction Br_2 is produced, the *bottom* layer will be yellow/brown (depending on the concentration).

2. Suppose the reaction in the preceding example did not occur. Because there is no reaction, the *bottom* layer will be purple, showing that I_2 is still present. Note that simply seeing a color does not mean that a reaction has taken place!

SAFETY PRECAUTIONS AND DISPOSAL METHODS

Bromine in elemental form and concentrated iodine solution cause severe burns on contact with the skin. Also, bromine and chlorine waters release vapors that are extremely harmful when inhaled. These liquids should be handled only in a fume hood—and with utmost care. Trichloroethane vapors are harmful and should not be inhaled.

Do not pour solutions containing trichloroethane down the drain. Dispose of these solutions in a stoppered bottle.

PROCEDURE

1. Metal–Metal Ion Reactions

A. Obtain one strip each of copper, zinc, and lead. Clean one side of each strip with emery paper. Lay the strips side by side on a paper towel on the desk, cleaned surface up.

B. Place on each strip 1 drop of each solution shown in Table 26.1. Record on your work page the combinations of metals and metal ions that showed evidence of a chemical change and those that did not. Wait about 5 minutes before you decide that no reaction has occurred. Include the silver ion, Ag^+, although the metal itself is not used

Table 26.1 *Solutions for Testing Metals*

Metal	*Test Solutions*
Copper, Cu	Zn^{2+}, Pb^{2+}, Ag^+
Zinc, Zn	Cu^{2+}, Pb^{2+}, Ag^+
Lead, Pb	Cu^{2+}, Zn^{2+}, Ag^+

because of its high cost. Assume that silver metal does not react with Zn^{2+}, Pb^{2+}, or Cu^{2+}.

C. On completing this part of the experiment, return or dispose of the metal strips as directed by your instructor.

Table 26.2 Solution Combinations

Test Tube	*Halogen Solution*	*Halide Solution*
1	Cl_2	KBr
2	Cl_2	KI
3	Br_2	NaCl
4	Br_2	KI
5	I_2	NaCl
6	I_2	KBr

2. Halogens

A. Select a small test tube and pour 10 drops of KBr into it. Then add about half a dropperful of trichloroethane (TCE). Shake the test tube to make sure the TCE, which is heavier than water, settles to the bottom.

B. Now add 3–4 drops of chlorine water. Shake the test tube well to make sure the free halogen is transferred to the bottom layer. Allow the two phases to separate and note the color of the *bottom layer.* By the color of this layer—and this layer only!—determine which elemental halogen (Cl_2, Br_2, or I_2) is present. Knowing this, you can then use your reasoning (see examples above) to determine if a reaction has taken place.

C. Repeat the procedure, using the same order: (1) halide solution first, (2) TCE next, and (3) halogen last, adding KI and Cl_2 (test tube 2 in Table 26.2).

D. Continue mixing the remaining pairs (test tubes 3–6), one at a time, and note the color of the *bottom layer.* Record your observations on your work page.

RESULTS

Record the answers to the following questions in the space provided on the work page:

1. Which combination(s) yielded a redox reaction?
2. For each reaction, write the half-reaction equations for both oxidation and reduction.
3. If necessary, multiply either or both equations to equalize the electrons gained and lost in the half-reaction equations.
4. Add the half-reaction equations to get the net ionic redox equation and record them on the work page. If a reaction did not occur, write the equation and slash the arrow to show no reaction.
   Example: $Zn^{2+} + Cu \rightarrow Zn + Cu^{2+}$

5. From Part 1, list the oxidizers (oxidizing agents) in a column according to decreasing strength. Judge oxidizing strength by considering which species each oxidizer was capable of oxidizing and which species it could not oxidize. The last species will be the oxidizer that was incapable of oxidizing anything.
   Example: If metal A reacted with B^{2+} and C^{2+}, and metal B reacted with C^{2+} but not with A^{2+}, and metal C did not react with either A^{2+} or B^{2+}, then the strongest oxidizer is C^{2+} and the weakest oxidizer is A^{2+}. Remember, the oxidizers in this step are the metal ions.
6. Prepare a similar list for Part 2.

Name *Date* *Section*

Experiment 26

Advance Study Assignment

1. Define the following terms:

a. Reduction

b. Half-reaction

c. Oxidizer

2. Combine the following half-reactions to produce a balanced net ionic redox equation:

a.

$$A \rightarrow A^{+} + e^{-}$$
$$Z + 2\ e^{-} \rightarrow Z^{2-}$$

b.

$$B \rightarrow B^{2+} + 2\ e^{-}$$
$$Y + 3\ e^{-} \rightarrow Y^{3-}$$

c. Which specie in reaction a) is the reducing agent? Why?

Name | *Date* | *Section*

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Work Page

Part 1—Metal–Metal Ion Reactions

Metal	*Ion in Solution*	*Reaction*		*Half-Reactions*	*Net Ionic Equations*
		Yes	*No*		
Cu	Zn^{2+}				
	Pb^{2+}				
	Ag^{+}				
Zn	Cu^{2+}				
	Pb^{2+}				
	Ag^{+}				
Pb	Cu^{2+}				
	Zn^{2+}				
	Ag^{+}				
Ag	Cu^{2+}				
	Zn^{2+}				
	Pb^{2+}				

List of Oxidizers in Order of Decreasing Strength

Part 2—Halogens

Test Tube	*Halide Solution*	*Halide Ion*	*Halogen*	*Color (after mixing)*	*Reaction*	
					Yes	*No*
1	KBr	Br^-	Cl_2			
2	KI	I^-	Cl_2			
3	NaCl	Cl^-	Br_2			
4	KI	I^-	Br_2			
5	NaCl	Cl^-	I_2			
6	KBr	Br^-	I_2			

Half-Reactions and Net Ionic Redox Equations

List of Oxidizers in Order of Decreasing Strength

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Name ______________________ Date ____________ Section ____________

Experiment 26

Report Sheet

Part 1—Metal–Metal Ion Reactions

Metal	*Ion in Solution*	*Reaction*		*Half-Reactions*	*Net Ionic Equations*
		Yes	*No*		
Cu	Zn^{2+}				
	Pb^{2+}				
	Ag^{+}				
Zn	Cu^{2+}				
	Pb^{2+}				
	Ag^{+}				
Pb	Cu^{2+}				
	Zn^{2+}				
	Ag^{+}				
Ag	Cu^{2+}				
	Zn^{2+}				
	Pb^{2+}				

List of Oxidizers in Order of Decreasing Strength

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Part 2—Halogens

Test Tube	*Halide Solution*	*Halide Ion*	*Halogen*	*Color (after mixing)*	*Reaction*	
					Yes	*No*
1	KBr	Br^-	Cl_2			
2	KI	I^-	Cl_2			
3	NaCl	Cl^-	Br_2			
4	KI	I^-	Br_2			
5	NaCl	Cl^-	I_2			
6	KBr	Br^-	I_2			

Half-Reactions and Net Ionic Redox Equations

List of Oxidizers in Order of Decreasing Strength

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