Oxidation-Reduction Reactions

Introduction

Many chemical reactions proceed by an electron \((e^-)\) transfer from one reactant to another. These electron-transfer reactions are referred to as oxidation-reduction or redox reactions. Oxidation is defined as the part of a redox reaction in which a species loses electrons and increases in oxidation number. Reduction is the part of the redox reaction in which a species gains electrons and decreases in oxidation number. In any reaction in which oxidation occurs, reduction must also occur. An oxidizing agent is a species that oxidizes another species and is itself reduced. A reducing agent is a species that reduces another species and is itself oxidized. An example of a redox reaction is zinc metal reacting with silver nitrate solution. The molecular equation is

\[
2 \text{AgNO}_3(aq) + \text{Zn}(s) \rightarrow 2 \text{Ag}(s) + \text{Zn(NO}_3)_2(aq)
\]  

(1)

The net ionic equation is

\[
2 \text{Ag}^+(aq) + \text{Zn}(s) \rightarrow 2 \text{Ag}(s) + \text{Zn}^{2+}(aq)
\]  

(2)

Zinc metal loses two electrons to form the zinc(II) ions, while the two silver ions each gain one electron to form two silver metal atoms. Equation 2 can be written in terms of two half-reactions. A half-reaction is one of two parts of a redox reaction, one of which involves a loss of electrons and the other which involves the gain of electrons. The half-reactions for equation 2 are

\[
\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2 \text{e}^-
\]  

(3)

\[
2 \text{Ag}^+(aq) + 2 \text{e}^- \rightarrow 2 \text{Ag}(s)
\]  

(4)

The number of electrons released in the oxidation half-reaction must equal the number of electrons consumed in the corresponding reduction half-reaction. Since the number of electrons is exactly equal, they do not appear in the overall redox reaction equation (2).
Procedure

1. **Reaction of Cu\textsuperscript{2+} with Fe**
   a. Add about 2 mL of 0.5 M CuSO\textsubscript{4} solution to a test tube.
   b. Add a small piece of steel wool (Fe) to the solution and shake until colorless.
   c. Decant the **colorless** solution into a clean test tube to be used in Reaction 2.

2. **Reaction of Fe\textsuperscript{2+} with H\textsubscript{2}O\textsubscript{2}**
   a. Add 4 drops of 3% hydrogen peroxide to the colorless solution above.

   **To help characterize the product**, acidify the solution with 5 drops of 6 M HNO\textsubscript{3}, then add 1 drop of 0.1 M KSCN. Consult a table of standard reduction potentials for the redox half-reaction of H\textsubscript{2}O\textsubscript{2}.

3. **Reaction of Cu\textsuperscript{2+} with Zn**
   a. Add ~2 mL of 0.5 M CuSO\textsubscript{4} solution to a test tube.
   b. Add a small piece of zinc metal to the solution and shake.
   c. After a fair amount of reddish-brown solid has formed and flaked off the piece of zinc metal, pour all of the solution and solids into a small beaker; rinsing if necessary.
   d. Remove the zinc with tweezers and set it aside for use in Reaction 6. Transfer the reddish-brown solid into a clean crucible and remove as much water as possible by decanting. Use the crucible and solid in reaction 5.

   Light your Bunsen burner and adjust the airflow to achieve a hot flame. Use it in the following **two** reactions.

4. **Reaction of Mg with O\textsubscript{2}**
   a. Grasp a piece of Mg ribbon with **crucible tongs** and hold the Mg in the flame until it ignites. Do not look **directly** at the reaction until it is over.

5. **Reaction of Cu with O\textsubscript{2}**
   a. Place the crucible from Reaction 3 on a clay triangle supported by a ring and ring stand. Adjust the ring height to allow the hottest part of the Bunsen burner flame to just touch the bottom of the crucible.
   b. **Gently** heat the contents of the crucible until **dry**, then strongly for five minutes. Carefully stir the contents of the crucible with a stirring rod and heat again for two minutes. Allow to cool.

   **To help characterize the product**, transfer the solid from the crucible to a test tube, add ten drops of 10% HCl followed by 10 drops of concentrated ammonia.
6. **Reaction of Zn with H^+**
   a. Place the zinc saved from Reaction 3 in a test tube and add ~2 mL of 10% HCl. Observe the reaction, dilute with distilled water to slow the reaction, and discard.

7. **Reaction of Mg with H^+**
   a. Place a piece of Mg ribbon in a test tube. Add about 2 mL of 10% HCl.

8. **Reaction of Cl_2 with I^-**
   a. Place ~1 mL of 0.1 M KI solution in a test tube.
   b. Add aqueous Cl_2 to the test tube just until a distinct yellow solution is obtained.

   **In order to characterize the product of the reaction above,** add a drop or two of 1 M Na_2S_2O_3 solution. Note what happens initially.

**Question**

The voltage of an alkaline battery is supplied by the following (unbalanced) reaction:

\[
\text{Zn(s) + MnO}_2(\text{s}) \rightleftharpoons \text{ZnO(s) + Mn(OH)}_2(\text{s})
\]

Write the two balanced half-reactions. Identify the substance oxidized and the substance reduced.

**Data Treatment and Discussion**

Identify the product of each experiment and write the (net ionic) half-reactions. Identify each half-reaction as an oxidation or a reduction.

Combine the two half-reactions of each experiment to give the net ionic equations.

**Conclusion**

Address the following based on the experiments you have performed in this laboratory and your observations:

1. Which is a stronger reducing agent, Cu or Zn? Explain.
2. Which is a stronger oxidizing agent, Cl_2 or I_2? Explain.
3. Which is a stronger reducing agent, Fe or Cu? Explain.
4. A bottle of FeSO_4 solution was accidentally left open to the air. The solution has turned a pale yellow color. What happened (chemically)?