Oxidation-Reduction Reactions
Prelab

1. What is the purpose of this experiment?

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

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

Write the two balanced half-reactions. In balancing the individual half-reactions, H₂O and/or H⁺ may be needed. Identify the substance oxidized and the substance reduced.
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. Zinc metal is observed to go into solution as zinc(II) ions (losing electrons), while silver metal deposits as a solid by reduction (gaining electrons) of silver(I) ions. This reaction can be written in terms of the two half-reactions

\[ \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2 \text{e}^- \quad (1) \]

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

A net ionic equation is obtained by adding the two reactions after multiplying equation (2) by two to make the electrons out of equation (1) equal to the electrons into equation (2). The number of electrons released in the oxidation half-reaction must always equal the number of electrons consumed in the corresponding reduction half-reaction.

\[ 2 \text{Ag}^+(aq) + \text{Zn}(s) \rightarrow 2 \text{Ag}(s) + \text{Zn}^{2+}(aq) \quad (3) \]

Procedure – Do steps 1 and 2 at the same time.

1. Reaction of Cu^{2+} with Fe

Add ~2 mL (40 drops) of 0.5 M CuSO_4 solution to a test tube. Add a small piece of steel wool (Fe) to the solution and mix until colorless. Decant the colorless solution into a clean test tube to be used in Reaction 3.

2. Reaction of Cu^{2+} with Zn

Add ~2 mL (40 drops) of 0.5 M CuSO_4 solution to a test tube. Add a small piece of zinc metal to the solution and mix.

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. Remove the zinc piece 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.
3. **Reaction of Fe$^{2+}$ with H$_2$O$_2$**

Add 4 drops of 3% hydrogen peroxide to the colorless solution from 1.

*To help characterize the product,* acidify the solution with 5 drops of 6 M HNO$_3$, then add 1 drop of 0.1 M KSCN.

Consult a table of standard reduction potentials for potential **half-reactions** of H$_2$O$_2$.

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$_2$**

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. Look at the tips of the tongs when the reaction is over.

5. **Reaction of Cu with O$_2$**

Place the crucible from Reaction 2 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. **Gently** heat the contents of the crucible until **dry**, then strongly for **five minutes**. Do **steps 6, 7, and 8** while it is heating. 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 and mix until dissolved. Add 10 drops of concentrated ammonia.

6. **Reaction of Zn with H$^+$**

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$^+$**

Place a piece of Mg ribbon in a test tube. Add about 2 mL of 10% HCl.

8. **Reaction of Cl$_2$ with I$^-$**

Place ~1 mL of 0.1 M KI solution in a test tube. Add several drops of aqueous Cl$_2$ to the test tube **just** until a color change is observed.

*In order to characterize the product of the reaction above,* add 1-2 drops of 1 M Na$_2$S$_2$O$_3$ solution. Note what happens **initially**.
Data Treatment

Write the (net ionic) half-reactions of each experiment by identifying the two products. Label 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 iron(II) sulfate solution was accidentally left open to the air. The solution has turned a pale yellow color. What happened (chemically)?