Experiment 12
Redox Reactions

OUTCOMES

After completing this experiment, the student should be able to:

- develop an activity series for different elements and ions.
- write balanced redox equations.
- describe the effect of pH on the degree of oxidation or reduction that takes place in a reaction.

DISCUSSION

Redox reactions, or **oxidation-reduction reactions**, are a family of reaction that are concerned with the transfer of electrons between species. **Oxidation** typically refers to the loss of electrons while **reduction** refers to the gain of electrons. More precisely, oxidation is defined as an increase in the oxidation state of an element (more positive oxidation number) and reduction is a decrease in the oxidation state of an element (more negative oxidation number). Redox reactions are always a matched set, meaning that there cannot be an oxidation reaction without a reduction reaction happening simultaneously. The oxidation and reduction portions of a particular reaction can be written separately and are called **half-reactions**. When writing half-reactions, the gained or lost electrons are usually explicitly included so that the half-reactions can be balanced with respect to electric charge.

While different species can either gain or lose electrons in a reaction, their willingness to be oxidized or reduced can vary depending on how strongly the electron is being held by the species. This can be shown quantitatively with **standard reduction potentials** such as those in your textbook. Half-reactions with more positive reduction potentials have reactants that gain electrons easily. Half-reactions with more negative reduction potentials have products that lose electrons easily. Using these reduction potentials allows one to determine whether or not a redox reaction between two species will occur.

Another way to determine whether or not a reaction will occur is to use an **activity series**. This is similar to the standard reduction potentials table, but in reverse order. The biggest difference between the two is that the activity series is based on qualitative observations instead of quantitative numbers such as those found with the reduction potentials. In the case of the activity series for metals, the activity series summarizes information about the reactions with acids and water, single displacement reactions, and the extraction of metals from their ores. The most reactive metals (most easily oxidized) appear at the top of the series, the least reactive at the bottom. Thus, a metal at the top of the series will prefer to exist as the ion and will donate electrons to the cations of the metals found below it.

In most cases, the order in which species occur in a standard reduction potential table and an activity series will be the same. However, the reduction potential table applies to species only in their standard state while the activity series looks at a variety of reactions and does NOT need to be in the standard state. Thus, there will be slight differences at times between the two, although it will usually just be a switch in two species that have similar reduction potentials.

PROCEDURE

⚠ ***Wear safety glasses or goggles at all times for this experiment.***
⚠ ***Avoid skin contact with chemicals in this experiment.***

Procedure 1: Halogens and Halides

This part should be performed under the hood.

1. Place 10 drops of the three different halogen solutions (X_2 water) into separate small test tubes. Add 10 drops of cyclohexane to each solution. Stopper the test tubes and shake vigorously for a couple of seconds. Record the color of the upper cyclohexane layer for each halogen. The color of this layer will tell you which halogen is present in solution in the subsequent reactions.

2. Into a new small test tube, combine 10 drops of a halogen solution with 10 drops of a 0.2 *M* halide ion solution (NaX). Add 10 drops of cyclohexane to the solution. Stopper the test tube and shake vigorously for a couple of seconds. Record the color of the cyclohexane layer.

3. Repeat step 2 until every possible combination of a halogen solution and halide ion solution has been made. You should have 6 combinations. *Note: Do not react the halogens with their corresponding halides.*

4. Discard the contents into the container labeled for halogenated organics. Clean your test tubes.

Procedure 2: Metals and Metal Ions

1. Obtain a 48-well reaction plate. Obtain seven pieces of each of six different metals: Cu, Zn, Pb, Fe, Mg, and an unknown metal. Place each piece of metal into a separate well.

2. React each metal with each of the following solutions: 0.1 *M* $CuSO_4$, 0.1 *M* $ZnSO_4$, 0.1 *M* $Pb(NO_3)_2$, 0.1 *M* $FeSO_4$, 0.1 *M* $MgSO_4$, 0.1 *M* $AgNO_3$, and 6 *M* HCl. Add enough of the solution to fill the reaction well half-full (about 15 to 20 drops).

3. Allow the reaction plate to sit for a few minutes. Determine the combinations that produced reactions and the ones that did not.

4. Empty the contents of the reaction plate into the designated container. Clean the reaction plate.

Procedure 3: Permanganate and Sulfite Ions

1. Obtain 3 small test tubes. Add 10 drops of 0.01 *M* $KMnO_4$ into each test tube.

2. Add 1 drop of 6 *M* HCl to the first test tube and 1 drop of 6 *M* NaOH to the third test tube.

3. Add 10 to 12 drops of 0.1 *M* Na_2SO_3 to each of the three test tubes. Record your observations, paying close attention to color, odor, and presence of gases

4. Empty the contents of the test tubes into the designated container. Clean the test tubes.

⚠ ***Dispose of all chemicals in the proper waste container.***

DATA ANALYSIS

1. Make a table of your experimental observations for each procedure.

2. For Procedure 1, compare the halogen added to the solution to the halogen observed after the reaction was completed to determine which combinations produced reactions.

3. Arrange the halogens by oxidizing strength from strongest to weakest. Arrange the halide ions by reducing strength from strongest to weakest.

4. Write balanced redox equations (net ionic) for each combination in Procedure 1 that produced a reaction.

5. For Procedure 2, determine the combinations that produced reactions and the ones that did not.

6. Arrange the metal ions by oxidizing strength from strongest to weakest. Arrange the metals by reducing strength from strongest to weakest.

7. Write balanced redox equations (net ionic) for any *five* of the combinations in Procedure 2 that produced a reaction.

8. For Procedure 3, identify the products of the three reactions based using the following guidelines:

   - The manganese(II) ion, Mn^{2+}, is faint pink in color although it will appear colorless at the concentrations present in this reaction. Manganese(IV) oxide, MnO_2, is a brownish-black solid that is insoluble in water . The manganate ion, MnO_4^{2-}, is a green ion in aqueous solution.

   - The sulfide ion, S^{2-}, is an ion that is colorless in aqueous solution but smells like rotten eggs in acidic solution. Sulfur trioxide, SO_3, is a white (not colorless) gas. The sulfate ion, SO_4^{2-}, is an ion that is colorless in aqueous solution and has no noticeable odor.

9. Write balanced redox equations (net ionic) for each of the combinations in Procedure 3 in the appropriate acidic or basic solution. In the second reaction, please note that while no acid or base was added, the reaction mixture was slightly basic. Also note that while the reactants for all three reactions are the same (MnO_4^{-} + SO_3^{2-}), one or more of the products will be different for the three reactions.

POSTLAB ACTIVITY

You will be turning in a worksheet for this experiment. It will be completed either individually or in pairs, according to your instructor's directions. The tables and equations that you generated in the data analysis will be incorporated into the worksheet.

Follow your instructor's directions for submitting the worksheet. If you are submitting electronically, please use the following convention for naming your worksheet: *Lastname1 Lastname2 Redox*. If you are emailing the worksheet, use a subject line of *Chem 1062: Redox Lab*.