Lab 5
Oxidation-Reduction Reactions

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INTRODUCTION

Oxidation-reduction reactions, often called **redox reactions**, are common among pairs of chemical compounds. Similar to acid-base reactions, redox reactions always involve two reactions: an oxidation reaction and a reduction reaction. The electrons lost in the oxidation reaction are the same electrons gained in the reduction reaction. In fact, they are so intricately connected that each oxidation and reduction reaction is often considered a half-reaction. Two half-reactions makes up one complete redox reaction.

Figure 1: Cast iron metal turns rusty when oxidized. Different chemicals can cause this, such as hydrogen, oxygen, salt, or other acids and corrosives.

OXIDATION

Oxidation is the loss of electrons by a molecule, atom, or ion. When a molecule or atom undergoes oxidation, it is said to have been oxidized. One simple example of oxidation, shown below, is the formation of the sodium ion from solid sodium. In this equation, solid sodium loses an electron to become the sodium ion:

$$Na\,(s) \rightarrow Na^+ + e^-$$
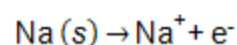
The resulting electrical charge on an atom after an oxidation reaction is referred to as the **oxidation state**. Transition metals, for example, are often oxidized to varying degrees. Copper can be oxidized to Copper (I) (also known as cuprous, or Cu^+) or Copper (II) (also know as cupric, or Cu^{2+}). It is important to keep track of the number of electrons lost during an oxidation reaction because of this variability. The number of electrons lost by one atom or molecule is equal to the number of electrons gained by another. Table 1 lists some of the common metals that have multiple oxidation states.

COLOR CHANGES WITH OXIDATION STATE

Different energy levels create different wavelengths, and thus produce different colors. The colors that the human eye can see are the wavelengths that are not absorbed by a material and bounce back toward the eyes. Electron transitions require and create a specific amount of energy. A change in the oxidation state (the number of electrons) consequently changes the amount of energy present in a given ion. If energy is lost or gained through a reaction, the resulting amount of energy present may alter the color of the product. Therefore, some oxidation reactions can be viewed through the color change that occurs.

Color changes can be easily viewed in liquid solutions. However, many minerals and precious stones contain transition metals that may exist at different oxidation states, which changes the color of the stone. For example, the mineral, beryl (Figure 2) contains several different metals that be present at different oxidation states. Beryl with iron (II) (Fe^{2+}) will create a pale blue or aquamarine color, but beryl with iron (III) (Fe^{3+}) creates a gold or yellow color. Further, if the mineral contains both Fe^{2+} and Fe^{3+} the resulting color will be dark blue. Beryl can also change colors if heated or irradiated (i.e., if energy is added).

Figure 2: The mineral beryl appears in different colors based on the oxidation states of its chemical makeup.

Oxidation-Reduction Reactions

REDUCTION

Reduction is the gain of electrons. When a molecule or atom undergoes reduction, it is said to have been reduced. For example, the formation of iron (II) from iron (III) is a reduction reaction:

$$Fe^{3+} + e^- \rightarrow Fe^{2+}$$
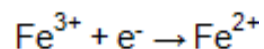
Once again, it is important to keep track of the number of electrons gained by the molecule or atom being reduced. As mentioned earlier, the number of electrons gained by the molecule or atom being reduced is equal to the number of electrons lost by the molecule or atom undergoing reduction. Thus, the molecule or atom undergoing reduction is called the **oxidizing reagent**, while the molecule or atom undergoing oxidation is referred to as the **reducing agent.**

Table 1: Common Metals With More Than One Potential Oxidation State

Element (Metal)	Cation Name	Cation Symbol
Chromium	Chromium (II) (*Chromous*)	Cr^{2+}
	Chromium (III) (*Chromic*)	Cr^{3+}
Manganese	Manganese (II) (*Manganous*)	Mn^{2+}
	Manganese (III) (*Manganic*)	Mn^{3+}
Iron	Iron (II) (*Ferrous*)	Fe^{2+}
	Iron (III) (*Ferric*)	Fe^{3+}
Copper	Copper (I) (*Cuprous*)	Cu^+
	Copper (II) (*Cupric*)	Cu^{2+}

OXIDATION NUMBERS

Oxidation numbers allow scientists to keep track of the total number of electrons transferred in a redox reaction. They express the charge of the central atom in a coordination compound (a chemical compound with a central transition metal at the core, and neutral or negatively charged molecules bound to that core) if all ligands and electron pairs were removed from the compound. Often, the numerical value of an oxidation number is the same value as the oxidation state, although this is not always the case. Importantly, oxidation numbers are not the remaining charge of an ion; they can be positive or negative numbers, and are assigned to elements using the rules defined in Table 2.

Using the rules outlined in Table 2, we can determine the oxidation numbers for the components of permanganate ($MnO_4)^-$. Based on the rules, oxygen always has an oxidation number of –2. To determine the oxidation number of manganese, we must first determine the overall charge on oxygen. To do this, we multiply –2 by the total number of oxygens (4) to get a final charge of –8. Because the sum of the oxidation numbers in a polyatomic ion is always equal to the ion charge, the oxidation number of manganese must be +7, because –8 + 7 = -1.

Table 2: Oxidation Number Assignment Rules

| Rule Number	Rule
1	An element by itself (rather than in a compound) is always zero. This is also true for diatomic elements (e.g., H_2, O_2, Cl_2, etc.).
2	A monoatomic ion is the same as the charge on the ion.
3	The sum of all the oxidation numbers in a neutral compound always equals zero. Also, the oxidation number of a polyatomic ion is always equal to the ion charge.
4	Alkali metals in a compound are +1. Also, alkaline earth metals in a compound are +2.
5	Oxygen is almost always -2.
6	Hydrogen is almost always +1.
7	Fluorine is always -1. Also, the oxidation numbers of chlorine, bromine, and iodine are almost always -1.

Oxidation-Reduction Reactions

## REDOX REACTIONS: BALANCING AND HALF-REACTIONS

Oxidation and reduction can be used to explain **oxidation-reduction reactions (redox reactions)**. The electrons lost by the first molecule (oxidation) must be equivalent to the electrons gained by the second molecule (reduction). In order to determine how many electrons are lost or gained, the net ionic equations of the redox reaction must be balanced. Because this process is complex, the equations are often split into two **half-reactions**. Each reaction is balanced separately and then combined to create the full, balanced redox reaction. This is called the **ion-electron method** of balancing redox equations. Let's review this process with the example below.

ION-ELECTRON METHOD EXAMPLE

The redox reaction between Sn^{2+}, Hg^{2+}, and Cl^- is balanced using the ion-electron method. You need to balance the equation, but you aren't sure where to start. You know that the unbalanced equation for this reaction is:

$$Sn^{2+}(aq) + Hg^{2+}(aq) + Cl^-(aq) \rightarrow Hg_2Cl_2(s) + Sn^{4+}(aq)$$
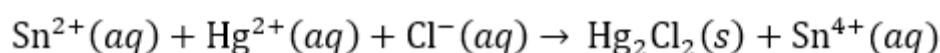
1. To begin balancing this equation, divide the total reaction into two half-reactions. Make sure that you do not change the charge or type of any atom referenced in the unbalanced equation.

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq)$$

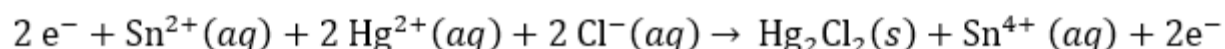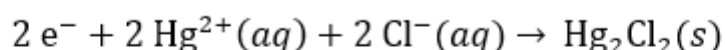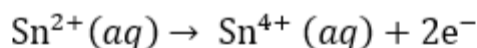$$Hg^{2+}(aq) + Cl^-(aq) \rightarrow Hg_2Cl_2(s)$$
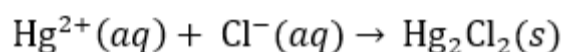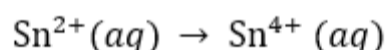
2. Next, determine which half-reaction is the oxidation reaction and which is the reduction reaction. In this case, Sn develops a more positive charge and loses electrons; therefore, it is oxidized. In contrast, Hg gains electrons, and is therefore reduced.

3. Balance each half-reaction by adjusting the coefficients where necessary.

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq)$$

$$2\,Hg^{2+}(aq) + 2\,Cl^-(aq) \rightarrow Hg_2Cl_2(s)$$
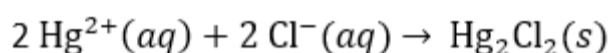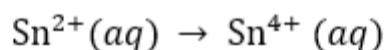
4. Balance the charges of each half-reaction by adding electrons to the more positive side of the equation (or, the less negative side of the equation).

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$$

$$2\,e^- + 2\,Hg^{2+}(aq) + 2\,Cl^-(aq) \rightarrow Hg_2Cl_2(s)$$
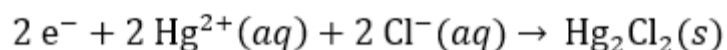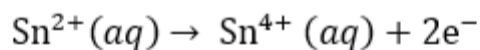
5. Add the two half-reactions together to create an equation in which the electrons gained equals the electrons lost, and add the two half-reactions together. Note that in our example, the electrons gained equals the electrons lost. This is not always the case!

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$$

$$2\,e^- + 2\,Hg^{2+}(aq) + 2\,Cl^-(aq) \rightarrow Hg_2Cl_2(s)$$

_____

$$2\,e^- + Sn^{2+}(aq) + 2\,Hg^{2+}(aq) + 2\,Cl^-(aq) \rightarrow Hg_2Cl_2(s) + Sn^{4+}(aq) + 2e^-$$
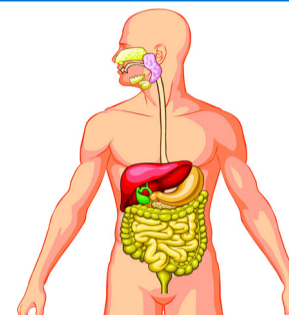
6. Cancel out any information that appears on both sides of the equation. **After completing this step, both the number of atoms and the total charge should be balanced**. At this point, your redox reaction is completely balanced, and you are left with:

$$Sn^{2+}(aq) + 2\,Hg^{2+}(aq) + 2\,Cl^-(aq) \rightarrow Hg_2Cl_2(s) + Sn^{4+}(aq)$$
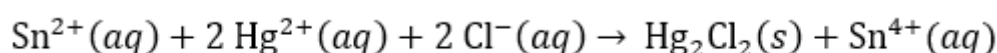
Oxidation-Reduction Reactions

Did you know that many of the chemical reactions that occur in your body are oxidation-reduction reactions? This process occurs when chemical energy is extracted from nutrients in a process called **catabolism** (the oxidative breakdown of molecules). The energy that is released from catabolism is then used to synthesize new molecules in a process called **anabolism** (the reductive synthesis of molecules).

Preparation of Iron (II) Acetate and Iron (III) Acetate

🧪 Experiment Inventory

Materials

150 mL 4.5% Acetic Acid ($C_2H_4O_2$)

150 mL 3% Hydrogen Peroxide (H_2O_2)

3 g Potassium Bitartrate ($KC_4H_5O_6$)

1 g Steel Wool

Matches

Sterno® Cooking Fuel

*Camera (camera phone is fine)

*Tap Water

*Paper Towel

*Pen or Pencil

Labware

(1) 100 mL Graduated Cylinder

(1) 250 mL Erlenmeyer Flask

(1) 500 mL Beaker

(1) Wire Mesh Stand

(2) Weigh Boats

(1) Funnel

(1) Filter Paper (Round)

(1) Glass Stir Rod

(1) Hot Pad

(1) Insulated Glove

(1) Metal Spatula

(1) Digital Scale

(1) Thermometer

*Timer (stopwatch, clock, internet, etc.)

Note: You must provide the materials listed in *red.

EXPERIMENT 1: PREPARATION OF IRON (II) ACETATE AND IRON (III) ACETATE

In this experiment, you will perform and observe an oxidation-reduction reaction.

LAB SAFETY: *Both the matches and the Sterno® can cause fire or burns to skin, clothing, or lab materials if the flame comes into contact with them. Do not stick any body part, loose clothing, or other flammable items in or near the flame. Be sure you have your safety glasses on, and never leave any burning fuel unattended!! Do this exercise in a well-ventilated area, and have a fire extinguisher nearby.*

PROCEDURE

Part 1: Preparation of Iron (II) Acetate

1. Use a 100 mL graduated cylinder to measure and pour 150 mL of acetic acid into a 500 mL beaker.

2. Place a weigh boat on the digital scale. Press the 0/T button on the right-hand side of the scale. Your scale should now read 0.0g. **Key Technique** ▶

3. Weigh out 1.0 g of steel wool. Then transfer the steel wool into the 500 mL beaker with the acetic acid.

☐ 4. Choose an area that is well-ventilated (yet not windy), and place your Sterno® cooking fuel can on a flat, heat-resistant surface.

☐ 5. Thoroughly read all of the directions on your Sterno® cooking fuel can, and then use your metal spatula to pry open the lid. Put the lid aside, being careful not to touch any of the contents. If you accidently touch the contents, wash your hands before you proceed with the experiment.

☐ 6. Use your metal spatula to slide the Sterno® cooking fuel can under the wire mesh stand.

☐ 7. Carefully place your 500 mL beaker with the acetic acid and steel wool inside on top of the wire mesh stand.

☐ 8. Touch a lit match to the Sterno® contents, and drop the match into the contents. The contents will flame up immediately. Under bright lights you may not be able to see the flame, but know that it is there. **Key Technique** ▷

☐ 9. Bring the acetic acid plus steel wool mixture to a boil, and evaporate off 50 mL. This is the process to produce a solution of iron acetate and should take approximately 30–45 minutes.

☐ 10. When the acetic acid plus steel wool mixture is down to the 100 mL mark on the 500 mL beaker, use the insulated glove to place the hot flask on a hot pad to cool. **Do not place the hot flask directly on the counter because it can crack the glass!**

☐ 11. Extinguish your Sterno® cooking fuel can by placing the cover on top of the can. **Do not blow out the flame.** Once the can is cooled (look at the temperature indicator on the label), snugly fit the cover onto the can and store it in a safe place away from heat or flames. **Key Technique** ▷

☐ 12. Allow your solution of iron (II) acetate to cool down. This should take approximately 2–5 minutes.

☐ 13. Observe and record the color of the iron (II) acetate solution in Table 3 on the Experiment 1 Data Sheet.

☐ 14. Gather a piece of round filter paper, a funnel, and a clean, dry 250 mL Erlenmeyer flask.

☐ 15. Prepare a filtering funnel as shown in Figure 3. Fold the filter paper in half twice to make four quarters. Place the paper in the funnel so that three quarters are open on one side and one quarter is on the opposite side. Seat the filter paper into the funnel by moistening the paper with a small amount of tap water. **Key Technique** ▷

☐ 16. Insert the filter paper into the funnel, and place the funnel into the mouth of the 250 mL Erlenmeyer flask.

☐ 17. Filter the iron (II) acetate solution into the 250 mL Erlenmeyer flask, leaving the steel wool in the funnel. Place the 500 mL beaker aside.
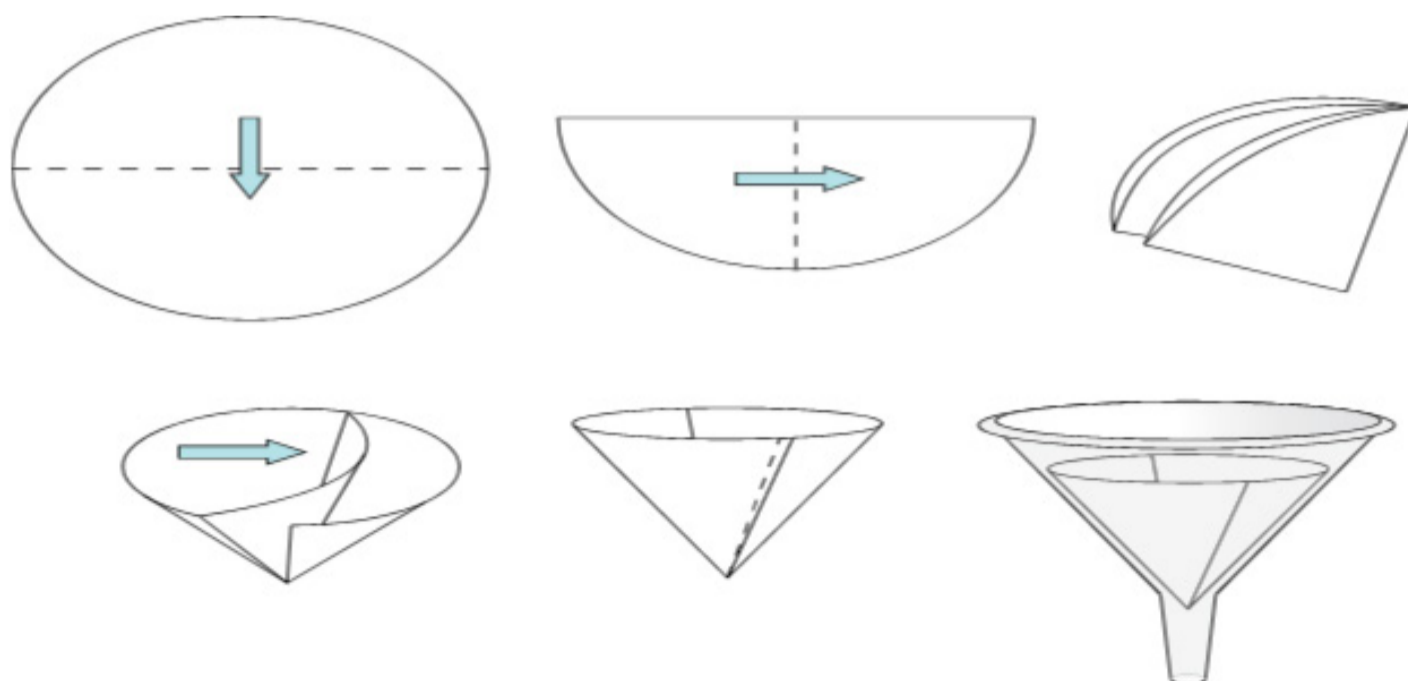
☐ 18. Clean and dry your 500 mL beaker.

Figure 3: Step-by-step process of folding the filter paper so that it will fit into the funnel.

Part 2: Oxidation of Fe²⁺ to Fe³⁺

☐   1.   Rinse your 100 mL graduated cylinder thoroughly with tap water. Dry your graduated cylinder with a paper towel.

☐   2.   While you are waiting for the iron (II) acetate to finish filtering, measure and pour 150 mL of 3% hydrogen peroxide into the clean, dry 500 mL beaker.

      Hint: First, use your clean, dry 100 mL graduated cylinder to measure 50 mL of 3% hydrogen peroxide, and pour it into the clean, dry 500 mL beaker. Then use the 100 mL graduated cylinder to measure 100 mL of 3% hydrogen peroxide, and pour it into the 500 mL beaker. Your beaker should now contain a total of 150 mL of 3% hydrogen peroxide.

☐   3.   Place a weigh boat on the scale. Press the 0/T button on the right-hand side of the scale. Your scale should now read 0.0g.

☐   4.   Use your metal spatula to weigh out 3 g of potassium bitartrate ($KC_4H_5O_6$) into your weigh boat.

☐   5.   Transfer your 3 g of potassium bitartrate into the 500 mL beaker that contains the 150 mL of 3% hydrogen peroxide solution.

☐   6.   Thoroughly stir the solution with the glass stir rod until the potassium bitartrate is evenly distributed throughout the hydrogen peroxide.

☐   7.   Remove the funnel from the Erlenmeyer flask, throw the filter paper with the steel wool into the trash, and rinse out your funnel.

☐   8.   Thoroughly rinse your 100 mL graduated cylinder with water. Dry your graduated cylinder with a paper towel.

☐   9.   Use your clean, dry 100 mL graduated cylinder to measure 15 mL of the iron acetate solution from the Erlenmeyer flask.

☐   10.   Place the thermometer into the beaker, and record the initial temperature (Time = 0) in Table 3 on the Data Sheet.

      Hint: Before proceeding to the next step, get your timer ready and have Table 3 and a pen or pencil in front of you.

☐   11. Add the 15 mL of the iron (II) acetate solution into the 500 mL beaker.

☐   12. Use Table 3 to record the temperature and color of the solution every 10 seconds for 2 minutes.

☐   13. When the reaction is complete, take a picture of your setup indicating the final color change, and submit it to your instructor along with your Post-Lab Questions.

Table 3: Temperature and Color Change Data

Time (seconds)	Temperature (°C)	Color
0		
10		
20		
30		
40		
50		
60		
70		
80		
90		
100		
110		
120		