The Copper Cycle

Most of the background material for this laboratory will be covered in greater detail in the lecture course later in the semester. Here is some background information so you will understand the chemistry behind the reactions you will perform.

Many aspects of our lives involve chemical reactions—from the batteries that power our cars and cell phones to the thousands of processes occurring within our bodies. Most of these reactions can be classified into one of **three main types** of chemical reactions: **precipitation** reactions, **acid-base neutralization** reactions, and **oxidation-reduction** (also called "**redox**") reactions.

Aqueous Solutions(aq)

Many reactions occur in an **aqueous** environment (i.e., in a solution where ions and compounds are dissolved in water). When we indicate that a reactant or product has the physical state *(aq)*, we mean the **substance is dissolved in water**. When an ionic compound is in aqueous solution, the individual ions are present in solution; for example, NaCl*(aq)* exists as Na^+ and Cl^- ions moving around in water.

Solubility Rules

Many ionic compounds are ***soluble***—i.e., they dissolve in water. Others generally do not dissolve in water and are considered ***insoluble***. To determine if an ionic compound is soluble—i.e., will dissolve—in water, we use the Solubility Rules:

Solubility Rules for Ionic Compounds in Water

The **compound is SOLUBLE** if it has:	The **compound is INSOLUBLE** if it has:
1. Li^+, Na^+, K^+, or NH_4^+ ion (**ALWAYS!**)	5. CO_3^{2-}, CrO_4^{2-}, PO_4^{3-}, **except compounds with** Li^+, Na^+, K^+, NH_4^+ **are soluble**
2. $C_2H_3O_2^-$, NO_3^-, ClO_4^-	6. S^{2-}, **except compounds with** Li^+, Na^+, K^+, NH_4^+, Ca^{+2}, Sr^{+2}, Ba^{+2} **are soluble**
3. Cl^-, Br^-, or I^-, **except compounds with** Ag^+, Pb^{+2}, and Hg_2^{+2} are **insoluble**	7. Hydroxide ion, OH^-, **except compounds with** Li^+, Na^+, K^+, NH_4^+ **are soluble**
4. SO_4^{2-} **except compounds with** Ag_2SO_4, $CaSO_4$, $SrSO_4$, $BaSO_4$, $PbSO_4$, and Hg_2SO_4 **are insoluble**	

The Solubility Rules indicate which compounds are soluble, and thus are represented as aqueous: e.g., KI*(aq)*, $BaCl_2$*(aq)*, NaOH*(aq)*, etc. The Solubility Rules also indicate which compounds are **insoluble**—i.e., do not dissolve in water and remain as solids: e.g. $BaSO_4$*(s)*, AgCl*(s)*, $CaCO_3$*(s)*, etc.

Double Replacement/Precipitation Reaction

For example, consider the reaction between aqueous lead(II) nitrate with aqueous potassium bromide, as shown below:

$$Pb(NO_3)_2(aq) + KBr(aq) \rightarrow PbBr_2 + KNO_3$$

Note that the chemical formulas for the products formed are based on their charges, not how they appear on the reactant side of the chemical equation.

Based on Solubility Rules #4 and #1, we find that $PbBr_2$ is insoluble and KNO_3 is soluble. Thus, the ***complete, balanced*** equation is:

$$Pb(NO_3)_2(aq) + 2\ KBr(aq) \rightarrow PbBr_2(s) + 2\ KNO_3(aq)$$

We can cancel the spectator ions from the ionic equation and write the net ionic equation:

$$Pb^{2+}(aq) + 2\ Br^-(aq) \rightarrow PbBr_2(s)$$

This reaction produces a cloudy mixture with small particles of the solid suspended in the solution. When enough solid has formed, it will begin to settle at the bottom of the beaker. Thus, a clear solution becoming cloudy when another solution is added is often taken as experimental evidence of a solid or precipitate forming.

Acids and Bases

Acids can be defined as substances that produce ***hydronium ions*** (***H_3O^+***) when they are dissolved in water. A hydronium ion is the product of a hydrogen ion that reacts with a water molecule: $H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$. A hydrated hydrogen ion ($H^+(aq)$) is equivalent to an aqueous hydronium ion. The two equations below both represent the ionization of hydrochloric acid, $HCl(aq)$, but the second one shows a particular water molecule explicitly.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Acids are usually easy to recognize since their formulas start with H and contains nonmetal elements other than H—e.g. $HCl(aq)$, $HNO_3(aq)$, and $H_2SO_4(aq)$ are all acids. *Note that the physical state aqueous, (aq), must be included to distinguish a compound that is acting like an acid from other forms of a substance. For example, the formula "HCl" can also be used for hydrogen chloride gas, HCl(g), so to indicate aqueous hydrochloric acid, one must specify HCl(aq).*

One useful definition of bases is that ***bases*** are compounds that ***produce hydroxide ions*** (***OH^-***) when dissolved in water. The dissociation of sodium hydroxide, NaOH, is shown below. :

$$NaOH(s) \rightarrow NaOH(aq) \text{ which is equivalent to } Na^+(aq) + OH^-(aq)$$

Acid-Base Neutralization Reactions

In an acid-base neutralization reaction, a hydrogen ion-containing acid reacts with a hydroxide-containing base to produce water and a salt (an ionic compound):

$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{NaOH(aq)} \rightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{NaCl(aq)}$$

Acids can react with bases, regardless of whether the salt is soluble or insoluble. There are other types of acids and bases that can react without forming water.

If the reactants and products of an acid/base reaction are colorless and soluble, it is impossible to monitor the progress of an acid-base reaction based solely on the appearance of the solutions. To help us monitor acid-base reactions, we use **litmus paper** to determine if a solution is acidic or basic. Litmus paper changes color depending on the presence of H^+ or OH^- ions in the substance being tested. Blue litmus paper turns **red** in **acidic** solutions containing H^+ ions, and red litmus paper turns **blue** in **basic** solutions containing OH^- ions.

Oxidation/Reduction Reactions

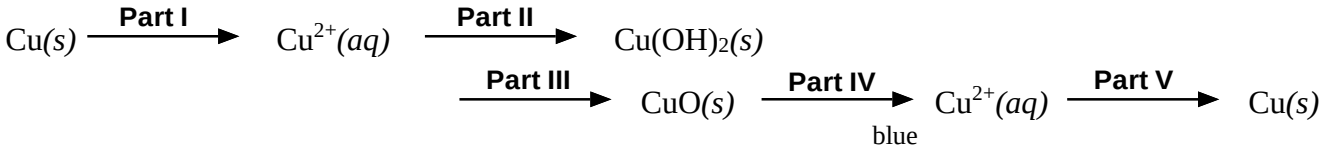
In an oxidation/reduction reaction, electrons are transferred from one reactant to the other. In the simplest form of these reactions, single-displacement reactions (also called single-replacement reactions), metal ions react with pure metals. If the reaction proceeds, the pure metal gives electrons to the metal cation. This causes the pure metal to become a cation and the cation to become a pure metal. The cation must always have an anion partner which is present either in an ionic solid or in a solution. For example:

$$\underset{\text{metal}}{Mg(s)} + \underset{\text{cation}}{2\ Ag^{+}(aq)} \rightarrow \underset{\text{metal}}{2\ Ag(s)} + \underset{\text{cation}}{Mg^{2+}(aq)}$$

If the charge of an element is changing, that is a good indication that an oxidation/reduction reaction is taking place. Later in the semester you will learn about oxidation numbers which are used to keep track of more complicated oxidation/reduction reactions.

Step I: Chemistry

The different copper species obtained in each part is shown in Equation 1 below:



I. Oxidizing Copper Metal with Concentrated Nitric Acid, $HNO_3(aq)$

The first step involves transforming Cu metal to copper(II) ions, Cu^{2+}, using concentrated nitric acid, $HNO_3(aq)$. At the same time, the nitrate ions (NO_3^-) undergo a series of reactions to form nitrogen monoxide, NO. This product rapidly reacts with oxygen in the air to form NO_2, a brown gas. The presence of $Cu^{2+}(aq)$ makes the solution blue.

When the reaction mixture is diluted with water, the Cu^{2+} ions are hydrated (surrounded by water) to form the octahedral complex ion, $[Cu(H_2O)_6]^{2+}$, as shown below. Six water molecules (shown as red O and white H atoms) are bonded to a Cu^{2+} ion (shown in gray as the central atom).

$$Cu^{2+}(aq) + 6\ H_2O(l) \rightarrow [Cu(H_2O)_6]^{2+}(aq)$$

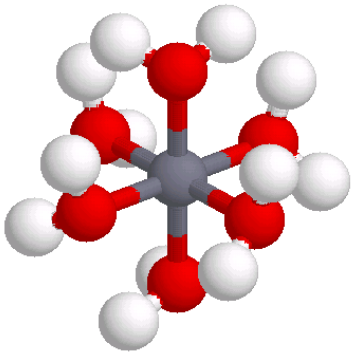


Figure 1

Step II: Chemistry

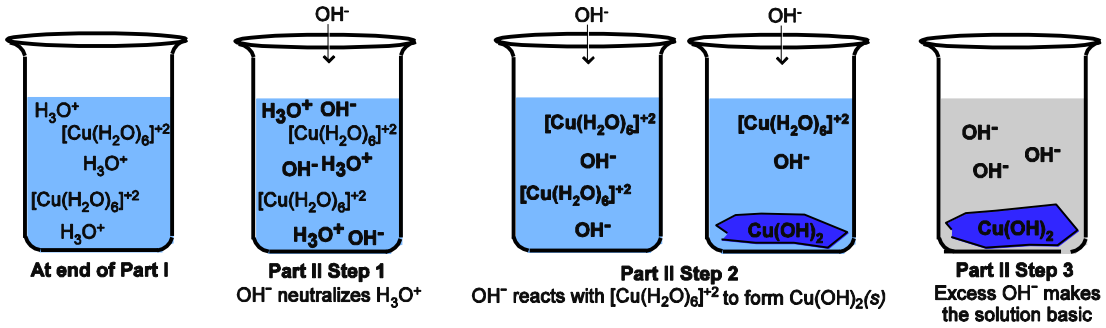
II. Precipitating $Cu(OH)_2$*(s)* with NaOH*(aq)*

In Part II, two reactions are carried out by adding NaOH*(aq)*. In the first reaction, the hydroxide ions (OH^-) from the NaOH*(aq)* neutralize the excess hydronium ions (H_3O^+) left over from the previous part:

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)$$

Once all the H_3O^+ ions are neutralized, additional OH^- ions react with the Cu^{2+} ion to form $Cu(OH)_2$ precipitate. Once all the Cu^{2+} ions have reacted, no more precipitate forms. Adding more OH^- ions makes the solution basic, so it can turn red litmus paper blue. Figure 2 on the next page shows the step-wise reaction of Cu^{2+} with NaOH.

Figure 2: Step-wise Illustration of the Precipitation of $Cu(OH)_2$ in Part II – Remember: $[Cu(H_2O)]^{2+}$ indicates the same substance as Cu^{2+}.



1st Beaker: At the end of Part I, hydrated copper complex, Cu^{2+} are present, making the solution blue, and excess hydronium ions (H_3O^+) remain from the nitric acid used.

2nd Beaker: Adding NaOH*(aq)* to the blue solution results in the OH^- ions neutralizing the H_3O^+ ions to form water: $H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)$. The Na^+ ions and resulting water molecules are not shown.

3rd and 4th Beakers: Once all the H_3O^+ are neutralized, adding more NaOH*(aq)* results in the OH^- ions reacting with the Cu^{2+} to form the blue $Cu(OH)_2(s)$ precipitate shown at the bottom of the beaker. Water molecules released from the complex ion are not shown.

5th Beaker: When all of the Cu^{2+} ions have been converted to $Cu(OH)_2(s)$ precipitate, adding more NaOH*(aq)* results in unreacted OH^- ions in solution, which makes the solution basic. Red litmus paper can be used to confirm the solution is basic. Note that the solution is no longer blue since no Cu^{2+} ions are present in the solution.

Step III: Chemistry

III. Converting solid $Cu(OH)_2$ to solid CuO

In Part III of the sequence, the reaction mixture is heated. This transforms the $Cu(OH)_2$ precipitate to CuO precipitate.

The CuO precipitate is separated from the solution, called the **supernatant liquid**, using a method called **gravity filtration**. The mixture is filtered using a filter funnel, and the solid is collected on filter paper. The supernatant liquid runs through the filter paper and collects in a beaker. This resulting filtered solution is called the **filtrate**.

Step IV: Chemistry

IV. Dissolving CuO*(s)* with sulfuric acid, H_2SO_4*(aq)*

In Part IV, the CuO precipitate is dissolved using sulfuric acid, H_2SO_4*(aq).* This redox reaction returns copper to its aqueous phase.

Step V: Chemistry

V. Reducing Cu^{2+} ions with Zinc Metal

In Part V, zinc metal (Zn) is added to the copper solution to convert the copper ions back to copper metal, Cu*(s)*. The resulting solution will contain colorless zinc ions, Zn^{2+}*(aq)* and copper solid. Visible evidence of this reaction is observed as bubbles of gas being released from the solution. (Since the H_3O^+ ions do not dissolve the Cu metal, the amount of copper yielded is not affected by excess acid.) Identify the gas displaced from the acid in this reaction.

When the solution becomes colorless, all of the Cu^{2+} ions have been converted to Cu metal.

All of the excess Zn metal is also converted to Zn^{2+} ion by the excess H_3O^+ ions from the sulfuric acid, H_2SO_4*(aq),*used to dissolve the CuO precipitate in Part IV.

Once all the Zn metal is dissolved, the Cu metal can be isolated by **decanting**, or **pouring off**, the supernatant liquid. The Cu will then be rinsed, dried, and weighed as described in the procedure.

The Copper Cycle

In this experiment, you will carry out a series of reactions starting with copper metal. This will give you practice handling chemical reagents and making observations. It is typical for scientists to observe materials before they react, what happens during a reaction and how it looks when the reaction has come to completion. The product of the final reaction will be copper metal and the percent copper that is recovered will be calculated.

****Lab Notebook****

You should include one table that contains the mass of copper at the beginning and end of the experiment along with % of copper recovered. This table should include:

- Mass of copper at the start of experiment (in Part I)
- Mass of copper + evaporating dish (from Part V)
- Mass of empty evaporating dish (from Part V)
- Mass of copper recovered (from Part V)
- Percent of copper recovered

Record observations for each of the steps (I-V) of the copper cycle in your lab book. Be sure to label each step (I-V). The observations for each step should include:

- the appearance of the reactants before the reaction
- the appearance of the reactants ***during*** the reaction (for example, bubbles, flames, etc.)
- the appearance of the products after the reaction.

Your observations should include state(s) of matter, color, texture, smell, etc. where applicable. If your observations are not detailed, you may not receive full credit.

One step also requires a specific chemical test using litmus paper to check for acidity. Be sure to also record the results of these tests in your lab notebook.

**You will turn in worksheet pages 11-12 along with the duplicate pages from your lab notebook.

Step I: Procedure - Oxidizing Cu with concentrated nitric acid, HNO_3*(aq)*

1. Place a sample of weighing paper in the balance. Tare the balance, so it reads 0.0000 g. Use forceps to transfer about 0.35-0.40 g of Cu strips onto the weighing paper. Record the mass of the Cu strips. Transfer the Cu strips into a clean 250-mL beaker labeled with one of your group member's initials. Record the appearance of the copper metal in your lab report.

   CAUTION: Concentrated nitric acid is highly corrosive, so it can cause severe chemical burns and damage clothing. Handle with care and avoid breathing the fumes. Any nitric acid spilled on skin must be rinsed immediately with water for 15 minutes. Any acid spilled on your work area must be neutralized then the entire area should be washed and dried.

   CAUTION: Concentrated nitric acid reacts with copper metal to form brown *toxic* NO_2 gas. Leave the reaction beaker in the fume hood until all of the brown gas is vented in the hood.

2. **In a fume hood**, use a 10-mL graduated cylinder to carefully measure about 3 mL of concentrated nitric acid, HNO_3*(aq)*. Slowly pour the nitric acid onto the Cu strips in the beaker, swirling the beaker to maximize contact between the Cu and nitric acid until all of the solid Cu has dissolved and the NO_2 gas has escaped. ***Keep the reaction beaker in the hood until all the toxic brown NO_2 gas is gone, and keep your face away from the hood to avoid inhaling nitric acid fumes and NO_2 gas.*** Describe the reaction between HNO_3 and the Cu metal in your lab report.

3. Dilute the resulting solution with about 10 mL of deionized water. Describe the appearance of the resulting solution containing Cu^{2+} in your data table.

Step II: Chemistry - Precipitating $Cu(OH)_2$*(s)* with NaOH*(aq)*

In Part II, two reactions are carried out by adding NaOH*(aq)*. In the first reaction, the hydroxide ions (OH^-) from the NaOH*(aq)* neutralize the excess hydronium ions (H_3O^+) left over from the previous part.

Once all the H_3O^+ ions are neutralized, additional OH^- ions react with the Cu^{2+} complex ion to form a gelatinous blue $Cu(OH)_2$ precipitate.

Once all the Cu^{2+} ions have reacted, no more precipitate forms. Adding more OH^- ions makes the solution basic, so it can turn red litmus paper blue. The picture sequence on the next page outlines the step-by-step process that occurs during this step.

Step II: Procedure - Precipitating $Cu(OH)_2$ with NaOH solution

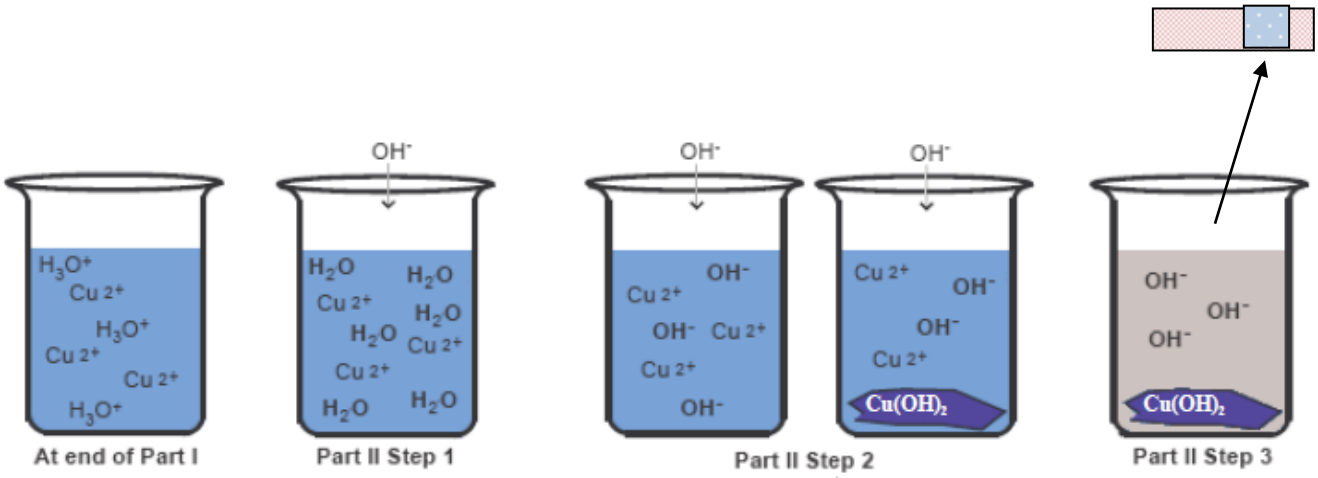
CAUTION: Sodium hydroxide (NaOH) can easily damage eyes. It is corrosive and can cause chemical burns and damage clothing. Any NaOH splashed into eyes or spilled on skin must be rinsed immediately with water for 15 minutes. Any base spilled on your work area must be neutralized then the entire area should be washed and dried.

1. While constantly stirring the Cu solution, slowly add 6M NaOH*(aq)* from the dropper bottles. First, the OH^- from the NaOH added will neutralize the excess acid left over from Part I.
2. Once all the acid is neutralized, additional OH^- ions react with the Cu^{2+} to form $Cu(OH)_2$*(s)*, a blue precipitate. Record what you observe in your lab report.

When adding more NaOH does not produce more precipitate, the solution can be tested to determine if all the Cu^{2+} has been precipitated and additional OH^- has made the solution basic. Use red litmus paper to test if the solution is basic as follows. *Without disturbing any precipitate*, use a glass stir rod to place a drop of **solution** (NOT the precipitate) on a piece of red litmus paper. If it turns blue, the solution is basic. Stop adding NaOH when the solution turns red litmus paper blue. Describe your litmus test in your lab report.

Step-wise Illustration of the Precipitation of $Cu(OH)_2$ in Part II

Check solution using red litmus paper (refer to background handout). Continue adding base until solution is basic.



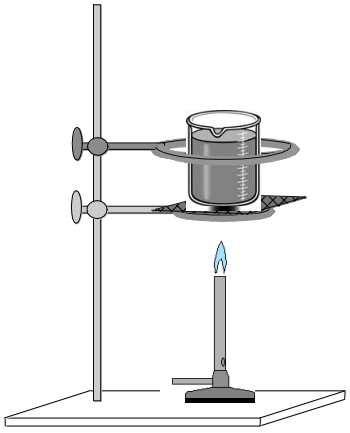
1st Beaker: At the end of Part I Cu^{2+} ions are present, making the solution blue, and excess hydronium ions (H_3O^+) remain from the nitric acid used.

2nd Beaker: Adding NaOH*(aq)* to the blue solution results in the OH^- ions neutralizing the H_3O^+ ions to form water: $H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)$. The Na^+ ions are not shown.

3rd and 4th Beakers: Once all the H_3O^+ are neutralized, adding more NaOH*(aq)* results in the OH^- ions reacting with the Cu^{2+} to form the blue $Cu(OH)_2(s)$ precipitate shown at the bottom of the beaker.

5th Beaker: When all of the Cu2+ ions have been converted to Cu(OH)2*(s)* precipitate, adding more NaOH*(aq)* results in unreacted OH– ions in solution, which makes the solution basic. Red litmus paper can be used to confirm the solution is basic. Note that the solution is no longer blue since no Cu^{2+} ions are present in the solution. In reality, your solution may still appear blue because of the dispersion of the $Cu(OH)_2$ in the solution by mixing.

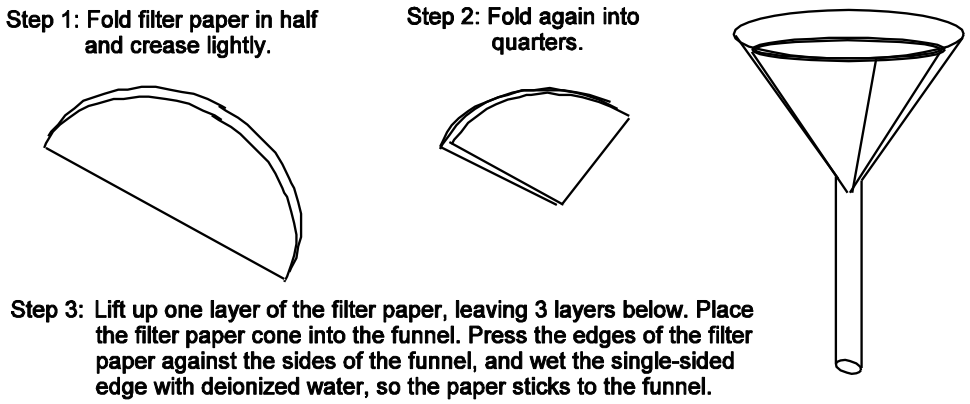
Step III: Procedure - Converting $Cu(OH)_2(s)$ to $CuO(s)$



1. Set up a ring stand as shown in the figure at the right. Set up a ring clamp, and put a wire gauze on top of it. Above it, attach another ring clamp with a diameter large enough to go around a 250-mL beaker. You are going to set your 250 mL beaker on the lower ring and gauze. The upper clamp will hold the beaker in place so it does not fall.

2. Add about 30-40 mL of deionized water to your reaction beaker from Part II. Carefully place the beaker on the ring stand inside the upper ring. **CAUTION:** Gently heat the beaker over a **medium** flame. (Set the inner cone of the Bunsen burner flame to a height of about 1.5 inch and the lower ring stand about 4 inches above the top of the Bunsen burner). Constantly stir the solution with the glass end of the stirring rod until all the blue precipitate turns black, and the solution is clear. If the solution starts to bump or boil, immediately remove the beaker from the heat and let the solution cool slightly. Describe what happens to the $Cu(OH)_2$ precipitate upon heating in your lab report.

3. Allow the beaker and contents to cool. While they are cooling, set up the gravity filtration apparatus. Obtain a second ring stand, and attach a ring clamp that is small enough to hold the plastic funnel. Prepare the filter paper as shown below:



Finally, place the plastic funnel in the small ring clamp, and place a 400-mL beaker beneath it to collect the filtrate (the liquid that goes through the filter paper). The funnel's stem should be just inside the beaker to prevent splashing.

4. Use the markings on a clean 150-mL beaker to measure out about 25 mL of deionized water. Boil the water on a hotplate to wash the precipitate in step 6.
5. When the 250-mL reaction beaker has cooled to room temperature, pour the CuO precipitate into the funnel to filter the contents. Transfer the last traces of the solid from the reaction beaker into the funnel, using a stream of deionized water.
6. Use a disposable pipet to wash the precipitate on the filter paper using the hot deionized water heated in the 150-mL beaker. Allow each portion of hot water to drain through the filter paper into the beaker below before adding the next portion. Use 15 mL of the hot deionized water to thoroughly wash the CuO precipitate.
7. Wash the 250-mL beaker, and rinse with deionized water. Replace the 400-mL beaker under the filter funnel with the clean 250-mL beaker. Discard the filtrate (solution) collected in the 400-mL beaker into the properly labeled waste container. Clean and dry the 400 mL beaker for use in Part V. ***Keep the CuO solid in the filter paper for use in Part IV.***

Step IV: Procedure - Dissolving CuO(s) with sulfuric acid, H_2SO_4*(aq)*

CAUTION: Sulfuric acid, H_2SO_4*(aq)*, is corrosive, so it can cause severe chemical burns and damage clothing. Handle with care and avoid breathing the fumes. Any sulfuric acid spilled on skin must be rinsed immediately with water for 15 minutes. Any acid spilled on your work area must be neutralized, and the entire area should be washed and dried.

1. Add about 10 mL of 3M sulfuric acid, H_2SO_4 (check the label before pouring), to the funnel to dissolve the CuO precipitate. Allow the solution to drain through the funnel to the rinsed 250-mL beaker. Repeat the procedure until all of the CuO solid dissolves to Cu^{2+} ions. Use as little of the sulfuric acid as possible in this step. Describe the reaction between the CuO precipitate and the H_2SO_4 in your lab report.

2. Use your water bottle to wash the last traces of solution from the empty funnel into the 250-mL beaker which now contains the acid solution and aqueous Cu^{2+} ions. **Keep this resulting solution for use in the Part V.**

Step V: Procedure - Reducing Cu^{2+} ions with Zn Metal

1. Use a weighing boat to measure and transfer about 1 g of Zn mesh to the Cu^{2+} solution in the 250-mL beaker. Constantly stir the mixture with a ***glass stirring rod***. **Do not use any metal object that will react with the acid to stir the solution.** Continue stirring until all the Cu^{2+} ions have been reduced to Cu metal as indicated by the solution becoming colorless. Dissolve any excess Zn by adding a few drops of 3M H_2SO_4*(aq)*. Describe your observations of the reduction of Cu^{2+} to Cu metal.
2. Allow the Cu metal to settle at the bottom of the beaker. Without losing any of the solid, carefully decant (pour off) as much of the supernatant liquid as possible into a 400-mL beaker. Some liquid will remain in the first beaker with the Cu metal. Wash the Cu metal 3 times using 20-mL portions of deionized water by stirring and then allowing the solid to re-settle. Again, decant the liquid into the 400-mL beaker each time.
3. Weigh a clean, dry evaporating dish. Transfer the Cu metal and any remaining water into the evaporating dish using a stream of deionized water. Decant most of the water from the evaporating dish. Use a disposable pipet to remove as much remaining water from the evaporating dish without losing any solid.
4. Write one of your group members' names on a folded piece of paper towel. Place your group's evaporating dish on the paper towel in the oven (between the hoods) to let the Cu completely dry. Check it after about 10 minutes. If the copper pieces are loose then it is dry. If it appears black in color, then the copper has been heated too much and has turned to copper (II) oxide.
5. When the Cu appears completely dry, let the evaporating dish cool to room temperature, and weigh the evaporating dish with the Cu. Record the final mass in your lab report.

Wash and dry all of your glassware, equipment, and your lab area to prevent chemical contamination and potential hazards.

Calculating Percent Copper Recovered

Theoretically, the mass of Cu recovered should be equal to the mass of the original Cu sample. The overall efficiency of the experiment is measured by calculating the percentage of copper recovered:

$$\textbf{percent recovered} = \frac{\textbf{mass of final product}}{\textbf{mass of initial sample}} \times 100\%$$

Ideally, the percent recovered should be close to 100%, which indicates that most (if not all) of the copper was successfully transformed through all five parts of the experiment.

The Copper Cycle

Name: _____

Partner: _____

Section Number: _____

*****Turn in pages 11-12 along with your lab notebook copies*****

For each equation shown below (1 for each step of the copper cycle), indicate

A. the products of the equation, include phases of all products,
B. balance the equation, and
C. tell what type of equation it is: combination, decomposition, single replacement, double replacement (precipitation), or acid/base neutralization (**Hint:** Many metal oxides can function as bases in aqueous solutions).

Step I. ___ $Cu_{(s)}$ + ____$HNO_{3\,(aq)}$ → ____$NO_{2\,(g)}$ + ____$H_2O_{(l)}$ + __ $Cu(NO_3)_{2\,(aq)}$

Type of reaction: This is a complicated reaction. NO^+ in the solution gets reduced to NO gas, which bubbles out of the solution and reacts with O_2 in the air to form NO_2. Balance the above equation even though it does not exactly describe the reaction you perform. The reaction has characteristics of decomposition, acid-base and oxidation-reduction reactions (single replacement).

Step II. ___$Cu(NO_3)_{2\,(aq)}$ + ____$NaOH_{(aq)}$ →

type of reaction: _____

Step III. ____$Cu(OH)_{2\,(s)}$ $\xrightarrow{\Delta}$ _____ + ____$H_2O_{(l)}$

type of reaction: _____

Step IV. ____$CuO_{(s)}$ +____ $H_2SO_{4\,(aq)}$ → _____ + _____

type of reaction: _____

Step V. ____$CuSO_{4\,(aq)}$ + ____$Zn_{(s)}$ → _____ + _____

type of reaction: _____

POST-LAB QUESTIONS:

1) Percent Copper Recovered

A student performing this experiment started with a 0.3769 g sample of copper turnings, which was dissolved in concentrated nitric acid. After completing the series of reactions, the student isolated 0.3492 g of copper. Calculate the percent copper recovered by the student. Show your work.

2) Indicate whether the following procedural errors would result in an **incorrectly high** or **incorrectly low** percent recovery. **Circle and explain your answer.**

a. The solution was not basic before being heated in Part III.
High Low

b. In Part III, the solution was poured into the funnel until it went above the top of the filter paper, and some black solid was disposed of with the filtrate.
High Low

c. The solution decanted in Part V was slightly blue in color.
High Low

d. After all the copper metal was obtained in Part V, it took too long for the excess Zn granules to dissolve, so a student added concentrated nitric acid to the solution, resulting in a brown gas.
High Low

e. In Part V, the copper metal was weighed while it was still damp.
High Low