

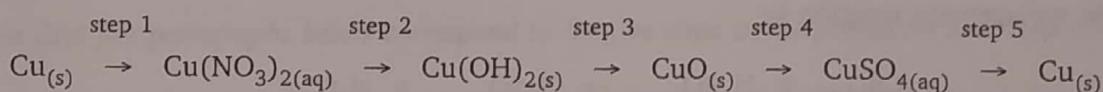
## Experiment 7. Chemical Reactions of Copper

### I. Purpose:

1. To observe some of the chemical reactions of copper.
2. To determine the percent yield of a series of chemical reactions.

### II. Theory:

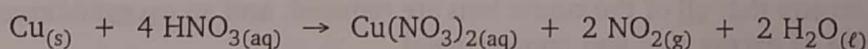
One interesting study of a group of chemical systems involves starting with a substance and transforming it into another substance, then repeating this process several times, ending eventually with the starting material. The efficiency of the overall reactions can be evaluated by calculating the percent yield of the starting material. In this experiment, copper metal is transformed into various copper-containing compounds by a series of sequential chemical reactions. The chemical transformations are shown below:



The meaning of the letters in parentheses is as follows: (s) = solid, (l) = liquid, (g) = gas, and (aq) = substance is dissolved in water to yield an aqueous solution.

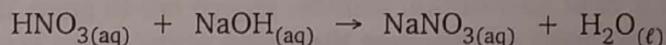
#### Step 1

Step 1 involves an oxidation-reduction reaction where copper metal is oxidized by nitric acid to form cupric nitrate,  $\text{Cu(NO}_3)_2$ , which exists in solution in the form of blue cupric ions ( $\text{Cu}^{2+}$ ) and colorless nitrate ions ( $\text{NO}_3^-$ ). As the copper dissolves, a brown gas, nitrogen dioxide (also called nitrogen (IV) oxide), is produced. The balanced equation for the reaction is:<sup>1</sup>



#### Step 2

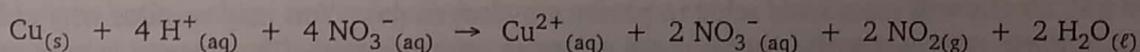
In the process of dissolving the copper metal, an excess of nitric acid was added in step 1 to ensure that all of the copper metal would go into solution. In step 2, sodium hydroxide (NaOH) solution is added to the cupric nitrate solution. At first, the sodium hydroxide neutralizes the excess nitric acid left over:



When the nitric acid is completely neutralized, the remaining sodium hydroxide will react with the cupric

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<sup>1</sup> - In this "molecular" equation, all substances are written in molecular form. In Chem. 2, you will learn that "strong electrolytes" dissociate completely into their ions in an aqueous solution. It is possible to convert a "molecular" equation into a "total ionic" equation by splitting up any strong electrolytes which may be present in the solution. If this is done, the first equation would then look like this:



For simplicity, only "molecular" equations have been used in describing the reactions which take place in this experiment. However, you should remember that soluble ionic substances such as cupric nitrate,  $\text{Cu(NO}_3)_2$ , dissociate completely into their constituent ions in the solution.

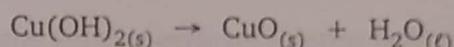
nitrate to form an insoluble precipitate of cupric hydroxide,  $\text{Cu}(\text{OH})_2$ :



This reaction is a double replacement reaction which essentially goes to completion because of the formation of the insoluble precipitate (this is indicated by the (s) and downward arrow).

### Step 3

The blue cupric hydroxide precipitate formed in step 2 is bulky and very difficult to filter. To make it easier for you to filter the solid in the next step, the solution containing this precipitate is heated to cause it to decompose partially into the more dense, black, crystalline cupric oxide.



### Step 4

After the cupric oxide formed above is filtered to remove it from the rest of the solution, sulfuric acid is added. The solid cupric oxide goes into solution because soluble cupric sulfate is formed. The equation for this double replacement reaction is:



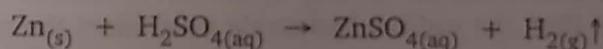
### Step 5

In order to recover the copper in this final step, zinc metal is added to the cupric sulfate solution. The zinc reduces the cupric ions, forming solid copper metal and zinc ions. The equation for this oxidation-reduction reaction is:



When this reaction is complete, the blue color of the cupric ion disappears (because the zinc ions are colorless.) In order to ensure that all of the cupric ions are reduced, and an excess of zinc metal is added, leaving behind a mixture of excess zinc and copper metal.

Fortunately the solution is still acidic (from the excess sulfuric acid added in step 4 above). The left-over zinc reacts with this sulfuric acid to produce a solution of zinc sulfate and hydrogen gas:



### Percent Yield

The only solid which remains is the copper metal, which is collected, dried, and weighed. From the mass of copper used in Step 1 and the mass of copper recovered in Step 5, the percent yield of copper in this set of reactions will be calculated.

$$\text{Percent yield} = \frac{\text{Mass of copper recovered}}{\text{Initial mass of copper}} \times 100$$

Note: In general, metals will react with acids to produce hydrogen gas. You may wonder why zinc will react with the sulfuric acid, while the copper does not react with the acid. The answer is that **copper is lower on activity series than hydrogen**. You will study the activity series in greater detail in Chem. 2 next semester.

### III. Chemicals Required:

0.5-g piece of No. 16 or No. 18 copper wire; 20- or 30-mesh zinc metal; acetone; methanol; concentrated nitric acid,  $\text{HNO}_3$ ; 3.0 M sodium hydroxide,  $\text{NaOH}$ ; 6.0 M sulfuric acid,  $\text{H}_2\text{SO}_4$ ; boiling chips.

### IV. Special Equipment Needed:

250-mL beaker; Evaporating dish; hot plate

*~ 27 cm ~ 0.53 g*

### V. Procedure:

*use 23 - 28 cm Cu wire*

**Caution:** You must wear departmentally approved eye protection or safety glasses while doing this experiment. Concentrated nitric acid must be handled with extreme care. Use the acid only in the hood. Avoid breathing the fumes. Nitric acid will stain your hands yellow; be sure to rinse it off immediately with running water in case of skin contact.

*The first five paragraphs below correspond to the five steps in the procedure described above.*

- Place a label with your name(s) or initials on the 250-mL beaker which you have signed out from the technician. Go to the electronic balance and tare the balance with this beaker on it.
  - Your goal is to weigh a number of small pieces of copper wire in this beaker so that their total mass is between 0.45 g and 0.55 g.* To do this, measure a piece of No. 18 copper wire about 7 cm long and cut off the piece from the spool. Scrape off any cupric oxide coating which might be present on the wire using steel wool. Then hold the wire over the previously-tared 250-mL beaker which is already on the electronic balance, so that the pieces which you will cut off will fall into this beaker. Using a pair of scissors, cut this wire into small pieces, each about 1 cm long, and allow the pieces to fall into the beaker. Keep on adding more cut copper pieces until their total mass is 0.45 g to 0.55 g.
  - Record the exact mass of the wire pieces in the beaker to the nearest 0.001 g (line 1).
  - CAUTION: WEAR SAFETY GOGGLES!** *In the hood*, measure out about 4-5 mL of concentrated  $\text{HNO}_3$  using a graduate, and add it to the copper in the beaker. Describe the reaction as to color, evolution of gas, and change in temperature (exothermic or endothermic) on the report sheet. (line 4) If some of the copper remains undissolved after a few minutes, use a hot plate *in the hood* to heat the solution gently. It may also be necessary for you to add a little more nitric acid to dissolve the copper.
- After the reaction is complete, add 100 mL distilled water. Then add about 30 mL of 3.0 M  $\text{NaOH}$  (use graduate) to the solution in your beaker and describe the reaction. (line 5) Mix well with a stirring rod and test the solution using a piece of red litmus paper. (If the paper stays red, the solution is still acidic; if the paper turns blue, the solution is now basic.) If the solution is still acidic, add more  $\text{NaOH}$  solution until it becomes basic.
- Add three boiling chips and carefully heat the solution to the boiling point while stirring using a ring stand, wire gauze, and bunsen burner. Describe the reaction on your report sheet. (line 6) Allow the black  $\text{CuO}$  to settle; then decant the clear liquid at the top by pouring it into your 600 mL beaker which will be used to temporarily store wash solutions. Add about 200 mL of very hot distilled water, stir, and then allow the  $\text{CuO}$  to settle. Decant once more (again pour the clear liquid into the your 600 mL waste beaker). What are you removing by this washing and decantation? (line 7)

4. **CAUTION: WEAR SAFETY GOGGLES!** *In the hood*, carefully add about 15 mL of 6.0 M  $\text{H}_2\text{SO}_4$ . Stir well with a stirring rod. What copper compound is present in the beaker now? (line 8)
5. Weigh out about 2.0 g of 30-mesh zinc metal using waxed paper, and bring the waxed paper and zinc to the hood. (The exact mass of the zinc need not be recorded.) *In the hood*, add the zinc metal all at once and **stir very well** until the blue color of the liquid has disappeared. If the solution remains light blue in color, add small amounts of zinc metal. After the blue color has disappeared, continue to stir well until no more bubbles of hydrogen gas have been produced. (If you perform this step too fast, you will have leftover zinc at the end along with the copper, and your results will be too high.) Describe the reaction on your report sheet. (line 9) What is present in solution? (line 10) **When gas evolution has become very slow, heat the solution gently for several minutes (but do not boil); you should observe some bubbles of a gas being produced.** Then allow it to cool. What gas is formed in this reaction? (line 11) How do you know? (line 12)
6. Weigh a clean, dry porcelain evaporating dish to the nearest 0.001 g. (line 2a)
7. When gas evolution has ceased, decant (pour off) the solution into your 600 mL waste beaker, and transfer the precipitate to a porcelain evaporating dish. Wash the precipitated copper with about 5 mL of distilled water. Allow the copper to settle, decant the solution, and repeat the process. What are you removing by washing? (line 13) Break up the lumps of copper metal formed using a glass rod. Wash the precipitate with about 5 mL of methanol (**KEEP THE METHANOL AWAY FROM FLAMES; IT IS FLAMMABLE!!**), settle, and decant *using the waste bottle in the hood.* (METHANOL IS ALSO EXTREMELY TOXIC: AVOID BREATHING THE VAPORS AS MUCH AS POSSIBLE.) Finally, wash with about 5 mL of acetone (**KEEP THE ACETONE AWAY FROM FLAMES; IT IS EXTREMELY FLAMMABLE!!**), allow to settle, and decant *using the waste bottle in the hood.*
8. Prepare a steam bath as illustrated in Figure 7.1 at the right. Be sure that the beaker is a convenient size to support the evaporating dish; a 250 mL beaker is often the correct size. Dry the copper metal on your steam bath for at least five minutes. Wipe the bottom of the evaporating dish with a towel, remove the boiling chips and allow the evaporating dish and contents (copper) to cool to room temperature. Then record the mass of the evaporating dish and copper. (line 2b) Calculate the mass of copper recovered (line 2c) and the percent yield. (line 3) What color is your copper sample? (line 14)
9. **Discard the copper metal as is directed by your instructor; DO NOT POUR IT INTO THE SINK because it is insoluble in water.**
10. **Waste Disposal:** At the conclusion of this experiment, transfer the solution in your 600 mL beaker into the waste container provided.

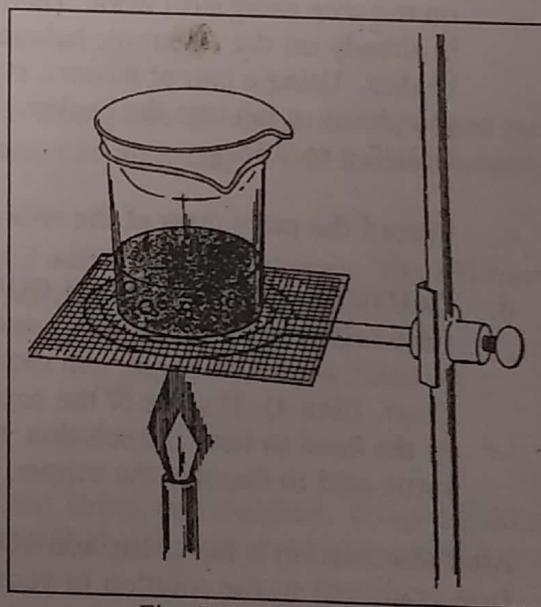


Fig. 7.1 - Steam Bath

## VI. Calculations:

Using the data on lines 1 and 2c, calculate the percent yield of copper in this sequence of reactions.

CHEM 110 - GENERAL CHEMISTRY I - LABORATORY REPORT

Experiment 7. Chemical Reactions of Copper

Name \_\_\_\_\_ Lab Partner's Name \_\_\_\_\_

Section \_\_\_\_\_ Date \_\_\_\_\_ Instructor \_\_\_\_\_

1. Initial mass of copper . . . . . \_\_\_\_\_

2a. Mass of evaporating dish . . . . . \_\_\_\_\_

2b. Mass of dish plus copper . . . . . \_\_\_\_\_

2c. Mass of copper recovered . . . . . \_\_\_\_\_

3. Percent yield of copper . . . . . \_\_\_\_\_

4. Describe the reaction  $\text{Cu} + 4 \text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + 2 \text{NO}_2 + 2 \text{H}_2\text{O}$  (Step 1).

5. Describe the reaction  $\text{Cu}(\text{NO}_3)_2 + 2 \text{NaOH} \rightarrow \text{Cu}(\text{OH})_2 + 2 \text{NaNO}_3$  (Step 2).

6. Describe the reaction  $\text{Cu}(\text{OH})_2 \rightarrow \text{CuO} + \text{H}_2\text{O}$  (Step 3).

7. What are you removing by the washing at the end of step 3? \_\_\_\_\_

8. What copper compound is present in the beaker at the end of step 4? \_\_\_\_\_

(Continued on the reverse side)

9. Describe the reaction  $\text{CuSO}_4 + \text{Zn} \rightarrow \text{ZnSO}_4 + \text{Cu}$  (Step 5).

10. After the above reaction (step 5) has taken place, what is present in the **solution**? \_\_\_\_\_

11. What is the **formula** of the gas produced during the heating process in step 5? \_\_\_\_\_

12. How do you know? (Hint: Look at your textbook where reactions of metals with acids are discussed.)

\_\_\_\_\_

\_\_\_\_\_

13. What are you removing by the washing in paragraph 7? \_\_\_\_\_

14. What color is your copper sample? \_\_\_\_\_