CONSERVATION OF MASS

To a beginning student of chemistry one of the most fascinating aspects of the laboratory is the dazzling array of sights, sounds, odors, and textures that are encountered there. Among other things, we believe that this experiment will provide an interesting aesthetic experience. You will be asked to carry out a series of reactions involving the element copper and to carefully observe and record your observations. The sequence begins and ends with copper metal. Because no copper is added or removed between the initial and final steps, and because each reaction goes to completion, you should be able to quantitatively recover all of the copper you started with if you are careful and skillful. The following shows in an abbreviated form of the reactions of the cycle of copper:

$Cu \rightarrow Cu(NO_3)_2 \rightarrow Cu(OH)_2 \rightarrow CuO \rightarrow CuSO_4 \rightarrow Cu$

Like any good chemist, you will probably be curious to know the identity of each reaction product and the stoichiometry of the chemical reactions for each step of the cycle. They are listed below:

$$4HNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

$$Cu(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + 2NaNO_3(aq)$$

$$Cu(OH)_2(s) \xrightarrow{\text{Heat}} CuO(s) + H_2O(l)$$

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

$$CuSO_4(aq) + Zn(s) \rightarrow ZnSO_4(aq) + Cu(s)$$

These equations summarize the results of a large number of experiments but it is easy to lose sight of this if you just look at equations written on the paper. You can easily be overwhelmed by the vast amount of information found in an experiment and in chemistry textbooks. It is in fact a formidable task to attempt to learn or memorize isolated bits of information that are not reinforced by your personal experience. This is one reason why it is important to have a laboratory experience. Chemistry is preeminently an experimental science.

As you perform the experiment, watch closely and record what you see. It is easier to remember information that is organized by some conceptual framework. Observations and facts that have not been assimilated into some coherent scheme of interpretation are relatively useless.

Chemists look for relationships, trends, or patterns of regularity in organizing their observations of chemical reactions. The periodic table is a product of this kind of thinking. It groups the elements into chemical families. Each element bears a strong resemblance to other members of the same chemical family but also has its own unique identity and chemistry.

In a similar fashion, it is useful to classify reactions into different types. Several different kinds of classification schemes exist because no one scheme is able to accommodate all known reactions. A simple classification scheme we will use at the beginning is one based on ideas of

combination, decomposition, and replacement. We present an outline and some examples of each type of reaction.

A SCHEME FOR CLASSIFYING CHEMICAL REACTIONS

Combination Reactions

Reactions that involve the combination of two or more pure substances to form a single substance.

Examples: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

$CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$

Decomposition Reactions

A single substance decomposes (often promoted by heat or light) into two or more different pure substance.

Examples: $2HgO(s) \xrightarrow{heat} 2Hg(s) + O_2(g)$

$NO_2(g) \xrightarrow{light} NO(g) + O(g)$

Single Replacement Reaction

A metal replaces another metal in a binary (made up of only two different elements) compound, or a nonmetal replaces another nonmetal in a binary compound.

Examples: $2AgNO_3(aq) + Cu(s) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$

$MgCl_2(aq) + I_2(s) \rightarrow MgI_2(aq) + Cl_2(g)$

Double Replacement Reaction

Both partners in a binary compound exchange with a second binary compound.

Example: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

A second type of double replacement reaction is called a neutralization reaction. In this process an acid reacts with a base producing a salt and water.

Example: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

As you carry out each step of the cycle of copper reactions, think about what is happening in each reaction and try to fit it into the classification scheme.

CHEMICAL RESPONSIBILITY

SAFETY ALERTS

1. Nitric acid, HNO_3, is toxic, corrosive, and an oxidant. It can cause severe skin burns. It also turns skin yellow on contact.

2. Sodium hydroxide, NaOH, is toxic and corrosive. It can cause severe skin burns and can cause severe damage to eyes. Wear your safety glasses.

3. Sulfuric acid, H_2SO_4, is toxic and corrosive. It can cause severe skin burns and can cause severe damage to eyes. Wear your safety glasses.

Some of the intermediate products of this reaction also are hazardous.

4. Copper(II) sulfate, $CuSO_4$, is toxic and an irritant.

5. Zinc sulfate, $ZnSO_4$, is an irritant and could be toxic.

6. Copper(II) nitrate, $Cu(NO_3)_2$, is toxic, corrosive and an oxidant.

   If you spill any of the above chemicals on your skin, immediately flush the affected area with water for several minutes. Quick action can prevent serious damage even from acids and sodium hydroxide.

7. Hydrogen gas, H_2, is a by-product of the reaction between zinc metal and sulfuric acid. It is a highly flammable gas. Make sure there are no open flames in the laboratory when performing this reaction. It is required for you to do this reaction in the hood to minimize any chances of a fire or explosion.

Chemical Disposal

1. All solutions should be poured into the Conservation of Mass waste bottle or 5-gallon container.

PROCEDURE

Caution!! Wear safety glasses at all times!!

I. Conversion of Copper Metal to Copper(II) Nitrate

1. Weigh a 0.250 to 0.350 g sample of copper metal (to the nearest 0.001 g) in a 150 mL beaker.

2. **Do the following step in the hood because of the generation of toxic nitrogen dioxide fumes (NO_2).**
   Add 10 to 15 mL of 6M nitric acid (HNO_3) to the beaker.

3. If the copper does not completely react, gently heat the beaker on a hot plate **in the hood** until all the copper metal has completely reacted.

   Caution! Nitrogen dioxide gas is generated in this step. It is toxic! Keep the reaction in the hood!

4. If more than half the liquid evaporates, add more 6M nitric acid. Do not allow the solution to boil.

NOTE: The blue solution is copper(II) nitrate; the brown gas evolved by this reaction is nitrogen dioxide, NO_2.

5. Complete the reaction listed on the report sheet.

II. Conversion of Copper(II) Nitrate to Copper(II) Hydroxide

6. Allow the beaker containing the copper(II) nitrate solution to cool to room temperature. **You can now take your beaker back to your work station.**

7. After the beaker has cooled, add 25 mL of distilled water to the beaker.

8. Pour about 10 mL of 6M NaOH into a 50 mL beaker.

Add 6M NaOH to the beaker until pH paper test blue (that is, the solution has become basic). The clear blue solution will turn to a milky green and then to a milky blue color. The solution will not test basic until it turns the milky blue color. Using a dropper add the NaOH by the dropper full until the solution turns green and then drop-wise until it turns blue.

Caution: Considerable heat may be given off in this step if any unreacted nitric acid remains in your reaction mixture. Adding the NaOH solution drop-wise will limit the spattering that may occur. Keep the beaker covered with a watch glass as much as possible during the addition of the first portions of the NaOH solution. Add the NaOH through the opening caused by the spout of the beaker.

9. Test your solution with pH paper. It should be alkaline (basic) at this point. If the pH paper is green or blue, pH is 7 or greater. If it is not, add drop-wise additional NaOH until the solution tests basic.

NOTE: The pale blue precipitate, which is now present in your beaker is copper(II) hydroxide.

10. Complete the reaction on the report sheet.

III. Conversion of Copper(II) Hydroxide to Copper(II) Oxide

11. Add 50 mL of distilled water to your pale blue copper(II) hydroxide solution.

12. Boil the solution gentle with occasional stirring until the pale blue precipitate is converted to the black precipitate, copper(II) oxide (CuO). This will take at least 5 minutes of gentle boiling. Keep heating your solution until no blue color remains.

NOTE: This can be done on the hot plate in the hood or using your Bunsen burner at your lab desk.

13. Complete the reaction on the report sheet.

IV. Conversion of Copper(II) Oxide to Copper(II) Sulfate

14. Allow your solution to cool to near room temperature without disturbing the beaker. This will allow your precipitate to settle to the bottom of your beaker. While your solution is cooling, prepare to perform a gravity filtration. Collect the filtrate (the liquid) in a 600 mL beaker to be used for waste.

15. After the precipitate has settled, pour off as much of the liquid portion as you can into the gravity funnel. When most of this liquid has filtered through, pour the remainder of the solution into the funnel. Rinse the beaker that contained the copper (II) oxide with about 5 mL of water and pour this into the funnel. When filtration is complete (or nearly complete), wash the brown-black solid on your filter paper with two 5 mL portions of distilled water (add the 5mL of water to the gravity funnel) and when most of the liquid has filtered through add the remaining 5 mL of distilled water.

16. Obtain about 15 mL of 3M sulfuric acid, H_2SO_4 (in a 50 mL beaker).

17. When most of the water has filtered, replace your waste beaker with the beaker that originally contained the copper(II) oxide. Pour the sulfuric acid onto the brown solid on the filter paper. When all the sulfuric acid has been added, gently stir the solution in the gravity funnel. Make sure not to put a hole in the filter paper. As the sulfuric acid reacts with the brown-black solid, it will turn to a clear blue solution. If any solid remains on the filter

paper, obtain a clean 150 or 250 mL beaker and replace the one catching the blue solution with this clean one. Pour the blue solution back through the filtration apparatus. If any brown solid still remains, repeat this procedure until the filter paper is clean. At the end of this process rinse the empty beaker with about 5 mL of distilled water and transfer it to the beaker containing the blue solution.

NOTE: The solution now contains blue copper(II) sulfate solution.

18. Complete the reaction on the report sheet.

V. Conversion of Copper(II) Sulfate to Copper Metal

19. Weigh out approximately 1 g of zinc metal (30 mesh).

20. Bring your beaker back to the hood. In the hood, slowly add the zinc to your copper(II) sulfate solution. If you add the zinc metal too quickly, it will react with excess sulfuric acid to produce a gas. This gas will bubble out of your solution, taking with it some of your liquid/solid mixture. This will result in a loss of copper metal. When the reaction has slowed down, you can bring the beaker back to your work station.

21. Allow the solution to stand until the blue color of copper(II) sulfate has completely disappeared. You should also notice the formation of copper metal at this point.

NOTE: The disappearance of the blue color is due to an electron transfer (that is, an oxidation-reduction) reaction taking place between metallic zinc and the copper(II) ions of copper(II) sulfate as follows:

$$Zn(s) + \underset{\text{blue}}{Cu^{2+}(aq)} \rightarrow \underset{\text{colorless}}{Zn^{2+}(aq)} + Cu(s)$$

22. Complete the reaction on the report sheet.

VI. Destruction of Excess Zinc Metal

NOTE: Once the blue color of the copper(II) ion has completely disappeared, it is safe to assume that all the dissolved copper (in the form of Cu^{2+}) has completely been reacted to form metallic copper. However, a closer look at the bottom of the beaker will also reveal the presence of unreacted zinc metal. This must be removed from our beaker before one can isolate the metallic copper. The metallic zinc is more reactive than metallic copper. The zinc metal can be converted to aqueous zinc sulfate solution by addition of sulfuric acid. H_2 gas is also a product of this reaction. The copper metal does not react with sulfuric acid.

23. Add about 10 to 15 mL of 3M sulfuric acid to your copper-zinc solution mixture.

24. Since it can take over an hour for all the zinc metal to react, you will stop at this point and finish this experiment next week.

25. From the supply table, obtain a marking pencil (red or blue, also called a china

marker) and write your name or initials on the side of the beaker. This will allow you to identify your beaker next week. When the bubbles (hydrogen gas) are forming at a slow rate, obtain a piece of parafilm and cover your beaker (make sure you leave a small opening so excess gas can still escape). Leave your beaker on your benchtop, the staff in the chemistry stockroom will store your beakers for you until next week.

26. Complete the reaction on the report sheet.

VII. Washing the Copper Metal

27. Allow the metallic copper to settle to the bottom of your beaker undisturbed.

28. Pour off as much of the liquid as possible into your waste copper solutions beaker. If you fill this beaker use your 400 mL beaker for your waste solutions and continue pouring off as much liquid as possible. Be careful not to lose any copper.

NOTE: Your copper metal solution is still contaminated with dissolved ions, which would affect the weight of your copper if not removed.

29. Weigh a clean, dry evaporating dish to the nearest 0.001 g.

30. Add 10 mL of distilled water to the copper metal solution. Transfer the solution containing the copper metal to the evaporating dish. Use your polywash bottle filled with distilled water to make sure that all the copper has been removed from the beaker.

31. Pour off as much of the liquid as possible into your waste copper solutions beaker without losing any of your copper metal.

32. Add about 10 mL of distilled water to the evaporating dish, stir thoroughly and then allow the copper to settle to the bottom of the evaporating dish. Pour off the liquid into your waste beaker.

33. Repeat this process twice more, discarding the washings into your waste copper solutions beaker and exercising care not to allow any of the copper to be lost.

VIII. Recovering the Copper Metal

34. Obtain a 250 mL beaker from your lab draw. Half fill it with tap water. Prepare to heat the beaker (that is, put a wire gauze on a ring and the beaker on the wire gauze).

35. Place your evaporating dish on the top of the beaker.

36. Heat the apparatus with your Bunsen burner until the copper metal is dry. Occasionally remove the evaporating dish from the top of the beaker and tap the side of the dish on the desktop gently (**the evaporating dish will be hot: Use crucible tongs or paper towels to hold it**). If the copper moves freely, it is dry. If not, continue heating and repeat.

NOTE:If the solid starts to turn a dark reddish-brown color, turn off the Bunsen burner. This indicates the formation of copper(II) oxide, CuO.

37. Allow the evaporating dish to cool to about room temperature. Dry the bottom of the evaporating dish and then weigh it.

38. Determine the mass of copper metal recovered.

REMEMBER TO TURN IN THE YELLOW COPY OF YOUR LAB NOTEBOOK TO YOUR TA.

Conservation of Mass
REPORT SHEET
(40 POINTS)

Name _____ Lab Day_____

Lab Instructor _____ Date _____

EXPERIMENTAL

Weight of copper metal initially taken: _____ g

Weight of cleaned evaporating dish: _____ g

Weight of copper metal and evaporating dish: _____ g

Weight of copper metal recovered: _____ g

% copper metal recovered: _____ g

For each step of the cycle, write the products of the reaction and balance the chemical equation. Using the classification scheme presented in this experiment, write the reaction type (combination,

decomposition, single replacement or double replacement) for each reaction. Also record what you observe during each step in your lab notebook. Lastly answer any questions.

1. $Cu + HNO_3 + O_2 \rightarrow$

   What is in the solution after the reaction is complete?

   Observations

2. $Cu(NO_3)_2 + NaOH \rightarrow$

   Reaction type: _____

   $HNO_3 + NaOH \rightarrow$

   Reaction type: _____

   What is formed in the solution besides $Cu(OH)_2$? _____

   Observations:

3. $Cu(OH)_2 \xrightarrow{heat}$

   Reaction type: _____

   What is removed by the washing and decantation process at the end of step 12? (Consider the products of the reaction as well as reagents from previous steps.)

Observations:

4. $CuO + H_2SO_4 \rightarrow$

Reaction type: _____

What is in the solution at the end of this step? _____

Observations:

5. $Zn + CuSO_4 \rightarrow$

Reaction type: _____

$Zn + H_2SO_4 \rightarrow$

Reaction type: _____

What happens when the zinc is added? _____

What is removed by the washing and decantation at the end of this part of your experiment?

_____

Observations:

Conservation of Mass
Post-Lab QUESTIONS
(30 POINTS)

Name _____ Lab Day _____

Lab Instructor _____ Date _____

1. A student performed this experiment, but was not particularly careful in following the procedure. Explain how each of the following procedural changes would affect his/her mass of copper metal recovered (too high, too low, or not changed)

   A. Not all the copper metal had reacted with the nitric acid when the student started to add the sodium hydroxide.

   B. Not enough sodium hydroxide was added, so the solution was still acidic when the beaker was heated to convert the copper(II) hydroxide to copper(II) oxide.

   C. There was still some brown copper oxide present on the filter paper when the student moved on to the next step (adding zinc metal).

D. The copper metal was heated too long.

2. A 0.325 g sample of copper was weighed out by a student to start this experiment.

A. How many moles of copper were weighed out?

B. How many moles of Cu^{2+} ions should be produced when the nitric acid was added to the copper metal?

C. When the sodium hydroxide was added to the solution, how many moles of $Cu(OH)_2$ should have formed?

D. The directions require you to add 1.00 g of zinc. If you assume a 100 % yield of copper, how many grams of zinc were added in excess?

E. If magnesium metal were used instead of zinc metal, what is the minimum mass, in grams, of magnesium metal that should be used to ensure that all of the copper ions in the solution is converted back to copper metal?

Conservation of Mass
Preliminary Questions
(10 POINTS)

Name _____ Lab Day _____

Lab Instructor _____ Date _____

1. Why is it necessary to perform the reaction of copper metal with nitric acid in the hood?

2. State the Law of Conservation of Mass.

3. Describe the hazards associated with following chemicals.

   A. Sodium hydroxide, NaOH

   B. Sulfuric acid, H_2SO_4

   C. Nitric acid, HNO_3

4. A student weighed out 0.385 g of copper metal to start the experiment. After completing the last step, the student had recovered 0.327 g of copper metal. What was the student's percent recovery?