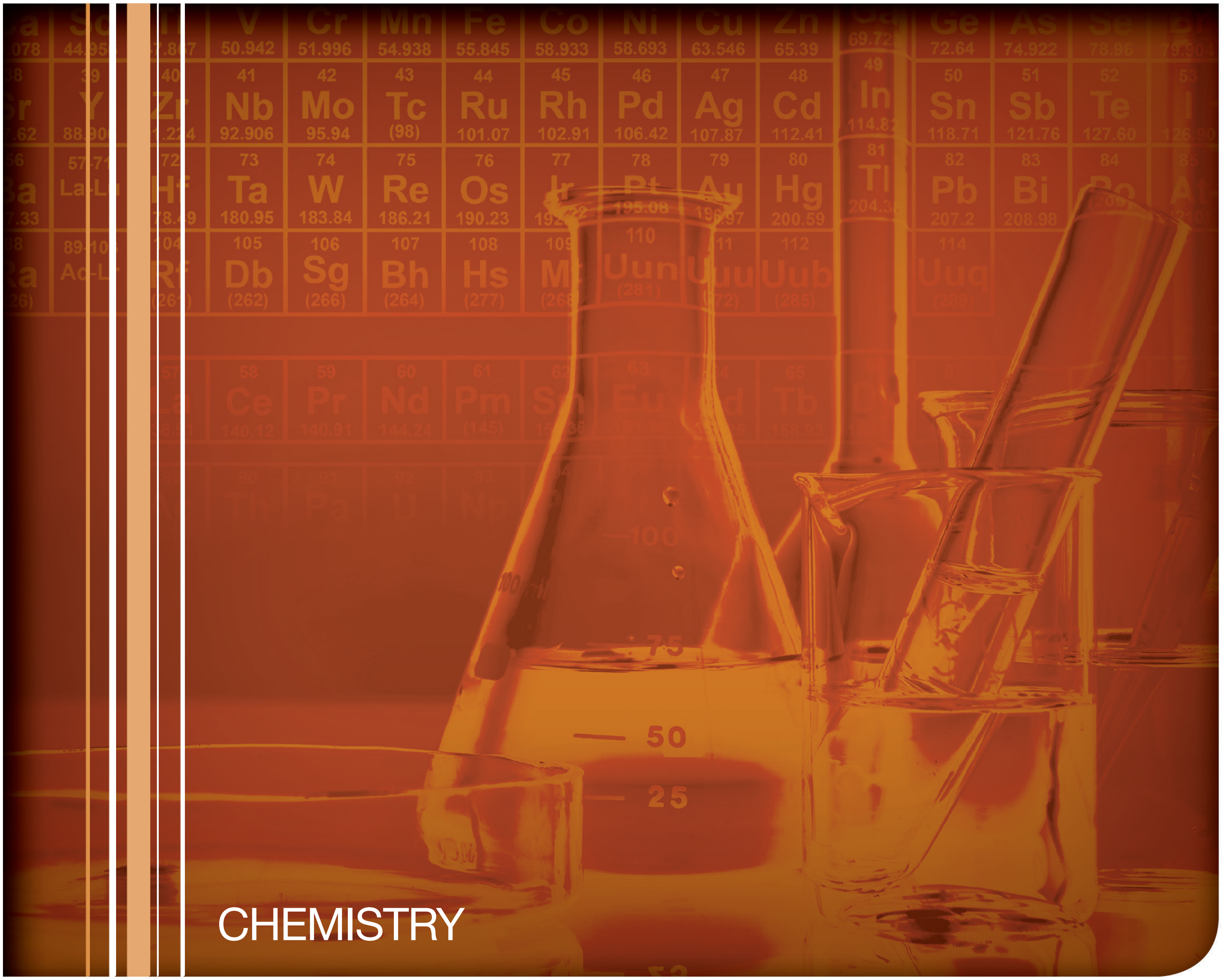

Single Replacement Reaction Stoichiometry and Percent Yield

Investigation Manual

CAROLINA
DISTANCE LEARNING

SINGLE REPLACEMENT REACTION STOICHIOMETRY AND PERCENT YIELD

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Overview

In this experiment, aluminum metal is reacted with copper sulfate to produce copper metal. The first step will determine the limiting reactant of the reaction and the theoretical yield produced from the reactant starting quantities. The theoretical yield is then compared with the actual yield obtained and the percent yield calculated.

Outcomes

- Balance chemical equations.
- Use stoichiometry as a tool to determine limiting reactants and theoretical yields.
- Quantitatively analyze the chemical reaction product(s).
- Calculate percent yield.

Time Requirements

Preparation .... 5 minutes
Activity 1: Single Replacement Reaction
Stoichiometry and Percent Yield .... 90 minutes

Key

Personal protective equipment (PPE)

goggles gloves apron

follow link to video

photograph results and submit

stopwatch required

warning corrosion flammable toxic environment health hazard

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Background

Stoichiometry enables chemists to calculate quantitative relationships between reactants and products in a chemical reaction. A balanced chemical equation indicates the reactant proportions required to generate products. Balanced chemical equations allow you to derive mole ratios of the reactants and products from the coefficients in the reaction. Stoichiometry calculations are used to convert these mole ratios into mass or volume and vice versa.

Stoichiometry calculations are used to calculate the amounts of substances that react and how much they produce in a chemical reaction. The word “stoichiometry” comes from the Greek *stoikheion* “element” and *metriā* “measure.” Using a balanced chemical equation and the molar mass or concentration of the reactants, the amount of product formed can be calculated. Stoichiometry is also used to determine theoretical yield, percent yield, or how much reactant is needed to prepare a specific amount of product.

If the reactants are not present in the correct molar ratio based on the balanced chemical equation, then one of the reactants will be consumed first. The amount of product generated will be limited by the fully consumed reactant. This is called the limiting reactant or limiting reagent. The reactant that is left over is called the excess reactant or reagent.

Once the limiting reactant has been determined, the amount of product generated can be calculated. This is called the theoretical yield. After the experiment has been completed, the actual yield can be determined and the percent yield calculated by dividing the actual yield by the theoretical yield and multiplying by one hundred.

The first step in stoichiometry is to balance the equation and determine the mole ratios of all the reactants and products. The second step is to take the mass or volume data provided and convert them into moles. For mass, use the molar mass, also known as formula weight, to convert to moles. For solutions use the molarity and volume to calculate moles. If the reactant is a gas, remember that one mole of any gas at standard temperature and pressure (STP) equals 22.4 L. After the molar quantities of each reactant have been calculated, the theoretical yield of the products can be determined using the mole ratio from the balanced equation.

When performing stoichiometric calculations the units are very important. Always include the units and cancel them out during the calculations. This is a good double check to ensure that you are using the correct data, units, and equations.

Solving Stoichiometric Calculations

To determine product yield from stoichiometric calculations, use the following six-step process.

Step 1: Write a balanced chemical equation for the reaction.

Step 2: Balance the electric charges in the chemical equations.

Step 3: Calculate the number of moles of each reactant.

Step 4: Determine if one of the reactants is a limiting reactant.

Step 5: Calculate the number of moles of product produced based on the number of moles of the limiting reactant and the balanced chemical equation.

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SINGLE REPLACEMENT REACTION STOICHIOMETRY AND PERCENT YIELD

Background continued

Step 6: Convert the answer in Step 5 to the appropriate units, such as grams, volume, or concentration.

Step 7: Calculate percent yield after determining actual yield experimentally.

When working stoichiometry problems, always do the following:

- **a)** Write the units on all numbers.
- **b)** Check that the units cancel properly.
- **c)** Give the answer in correct units.
- **d)** Avoid rounding numbers too much during the calculation, but round the final answer to the correct number of significant figures.

Step 1: Balancing Chemical Equations

The first and most important step in stoichiometry is writing a correctly balanced chemical equation. The balanced chemical equation should also include the physical state of each species. This will dictate how to convert between the number of moles and measurement units. For example, if magnesium metal and hydrochloric acid react to produce magnesium chloride and hydrogen gas, the balanced chemical equation is:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

The balanced chemical equation provides two important pieces of information.

1. Mole ratio: The equation coefficients are used to determine the molar ratio. One mole Mg reacts with two moles of HCl to produce one mole of $MgCl_2$ and one mole H_2.
2. Physical state: Mg is a solid, H_2 is a gas, and HCl and $MgCl_2$ are both aqueous solutions.

Step 2: Balancing Electrical Charges

Many chemicals dissolve into ions in water. They fall apart in atoms or elements containing an electric charge. In a chemical reaction, the electric charge on both sides of the equation needs to be the same. As an example, we are going to use water, H_2O, which has no electrical charge. The elements that make up water do contain an electrical charge when dissolved. Hydrogen, or H, has one electrical charge. This is indicated by H^+, where the plus sign indicates the positive electrical charge. When there is no number in front of the electrical charge, it means that there is only 1. Thus, in this case, hydrogen has an electrical charge of 1 plus. Oxygen has two negative electrical charges, indicated by O^{2-}. The minus sign indicates an electrical charge, and the 2 indicates two negative charges.

In order to balance the reaction of creating water, the net electrical charge needs to be zero. The negative and positive charges need to cancel each other out. An easy way to do this is to multiply the electrical charge of the positive molecule by the number of negative molecules reacting, and multiply the number of negative charges times the number of positive molecules available. In this case, 1 hydrogen times 2 negative charges gives us 2, which indicates how many hydrogen ions we need to cancel out the 2- charge of the oxygen.

We can do the same for the positive charge. The positive charge times 1 oxygen ion gives us 1. We only need one oxygen ion to cancel out the positive charge of the hydrogen.

This will give us $2H^+ + O^{2-} \longrightarrow H_2O$. On both sides of the equation the net electrical charge is zero. The equation is now balanced.

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Step 3: Calculate the Number of Moles of Reactants

Dimensional analysis is used in stoichiometric calculations to convert between units such as mL and L. When setting up a stoichiometric calculation, dimensional analysis ratios should be oriented so the same units cancel each other out.

> There are several different methods to perform dimensional analysis. Some instructors teach a method that uses grids to keep track of units. To keep track of units in the method shown here, the number always goes with the top unit and the bottom unit has an implicit '1' as the number. For example, 15 g/mol should be written as $\frac{15\text{ g}}{1\text{ mol}}$.

When setting up a stoichiometric calculation, begin with the units provided in the problem and convert the value to moles. If the reactant or product is a solid, begin with mass and divide by the molar mass calculated from the periodic table. If the reactant or product is a solution, start with molarity and multiply the volume to calculate moles. If it is a gas, use the ideal gas law (PV = nRT, where P is pressure, V is volume, n is number of moles, R is the universal gas constant, and T is temperature) to calculate the number of moles. To simplify this, if the gas is at standard temperature and pressure (STP), then one mole of gas occupies 22.4 L.

If 1.03 g of the magnesium metal and 100 mL of 0.5 M hydrochloric acid are used in the reaction above, the number of moles of the reactants are calculated as follows:

Mg(s)

$$\text{moles Mg} = \frac{\text{mass Mg}}{\text{molar mass Mg}}$$

$$\text{moles Mg} = \frac{1.03\text{ g}}{24.3\text{ g / mol}}$$

$$\text{moles Mg} = 0.0424\text{ moles}$$

HCl(aq)

$$\text{molarity} = \frac{\text{moles}}{\text{volume in Liters}}$$

$$\text{moles} = \text{molarity} \times \text{volume in Liters}$$

$$\text{moles HCl} = 0.5\text{ M} \times 100\text{ mL} \times \frac{1\text{ L}}{1{,}000\text{ mL}}$$

$$\text{moles HCl} = 0.05\text{ moles}$$

Step 4: Determine the Limiting Reactant

There are several different methods for determining the limiting reactant. One method is to determine the number of moles of product produced from each reactant. The reactant that generates the smaller amount of product will be the limiting reactant because it will be consumed first. Use the mole ratio of the reactant to product based on the coefficients in the balanced chemical equation. For the above example from step 3.

Moles $MgCl_2$ produced based on Mg

$$0.0424\text{ moles Mg} \times \frac{1\text{ mole } MgCl_2}{1\text{ mole Mg}}$$

$$= 0.0424\text{ moles } MgCl_2$$

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Background continued

Moles $MgCl_2$ produced based on HCl

$$0.05 \text{ moles HCl} \times \frac{1 \text{ mole } MgCl_2}{2 \text{ moles HCl}}$$

$$= 0.025 \text{ moles } MgCl_2$$

HCl is the limiting reactant because 0.025 moles of $MgCl_2$ is less than 0.0424 moles of $MgCl_2$, therefore the hydrochloric acid will be consumed first. The Mg is an excess reactant because it still exists when all the HCl has been consumed.

Step 5: Calculate the Number of Moles of Product Produced

Use the limiting reactant to calculate the number of moles of each product. In most problems, the number of moles of one of the products will have already been calculated in Step 4. Remember to use the proper coefficients from the balanced chemical equation.

Moles $MgCl_2$ produced based on HCl

$$0.05 \text{ moles HCl} \times \frac{1 \text{ mole } MgCl_2}{2 \text{ moles HCl}}$$

$$= 0.025 \text{ moles } MgCl_2$$

Moles H_2 produced based on HCl

$$0.05 \text{ moles HCl} \times \frac{1 \text{ mole } H_2}{2 \text{ moles HCl}}$$

$$= 0.025 \text{ moles } H_2$$

Step 6: Convert Moles of Product to the Final Units

This step is similar to Step 3, but in reverse. Depending on the physical state of the product, convert moles to grams (solid), molarity (aqueous solution), or volume (gas). Remember that the solution volume will not change and that the gas volume is still subject to the ideal gas law (PV = nRT). In the example above, since the gas is produced at STP the temperature, pressure, and universal gas constant are the same on both sides of the equation and cancel out. This means that the gas volume can be calculated using a simple ratio.

Step 7: Calculate Percent Yield

$MgCl_2(aq)$

$$\text{molarity} = \frac{\text{moles}}{\text{volume in Liters}}$$

$$\text{molarity } MgCl_2 = \frac{0.025 \text{ moles } MgCl_2}{0.1 \text{ L}}$$

$$\text{molarity } MgCl_2 = 0.25 \text{ M } MgCl_2$$

$H_2(g)$

$$\frac{22.4 \text{ Liters of gas}}{1 \text{ mole of gas}} = \frac{\text{X Liters of } H_2}{0.025 \text{ moles of } H_2}$$

$$= 0.56 \text{ Liters of } H_2$$

The last step in some stoichiometry problems is to determine the percent yield of the reaction. Either product can be used to determine the percent yield. In general, it is easier to measure the mass or volume for solids and gases, respectively, than the molarity of a solution when calculating the yield.

If 0.49 Liters of H_2 gas is produced in the

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example, what is the percent yield?

$$\frac{\text{actual yield of } H_2}{\text{theoretical yield of } H_2} \times 100\% = \text{percent yield}$$

$$\frac{0.49 \text{ L of } H_2}{0.56 \text{ L of } H_2} \times 100\% = 87.5\%$$

The percent yield tells us how well the reaction worked in terms of forming the desired product. A percent yield below 100% is common. It could be due to the reaction not going to completion or that some product loss occurred when it was transferred from one container to another or filtered. Side reactions also lead to lower percent yields since they consume some of the reactants or products. It is also possible to have a percent yield greater than 100%. This could be due to weighing errors or an additional substance present with the product, making it weigh more. Incomplete drying may also inflate percent yield because moisture adds to the weight of a solid product.

Single Replacement Reactions

In this investigation, aluminum metal and a solution of copper(II) sulfate react to produce copper metal and an aluminum sulfate solution. A small amount of sodium chloride is added to the reaction as a catalyst. The chloride ion removes a very thin layer of aluminum oxide from the aluminum foil surface, allowing the copper to react with pure aluminum metal. Because the chloride ion is a catalyst and not involved in the overall reaction, it is not included in the balanced chemical equation.

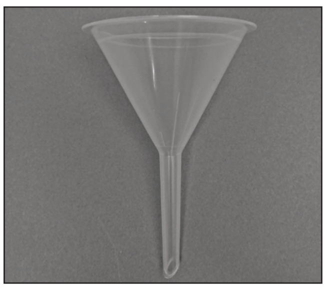
SINGLE REPLACEMENT REACTION STOICHIOMETRY AND PERCENT YIELD

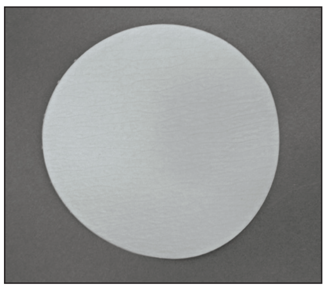
Materials

Included in the materials kit:

Copper(II) sulfate solution, 1 M, 100 mL


Aluminum foil


Filter funnel


Filter paper

Needed from the equipment kit:


Graduated cylinder, 50 mL


Electronic balance


Pipets


Spoons


Weighing boats


Plastic beaker, 250 mL

Reorder Information: A replacement kit for Single Replacement Reaction Stoichiometry and Percent Yield, item number 580338, can be ordered from Carolina Biological Supply Company.

Call 800-334-5551 to order.

Needed but not supplied:

- Ruler, 6"
- Paper towels
- Sodium chloride, NaCl, table salt, 1 g

Safety

Wear your goggles, lab apron, and gloves at all times while conducting this investigation.

Read all the instructions for this laboratory activity before beginning. Follow the instructions closely, and observe established laboratory safety practices, including the use of appropriate personal protective equipment (PPE) as described in the Safety and Procedure sections.


The copper sulfate solution is toxic when ingested and is a tissue irritant.


Copper sulfate is very toxic to aquatic life.

Safety continued

Wash your hands thoroughly after handling the copper sulfate or copper metal.

Do not eat, drink, or chew gum while performing this activity. Wash your hands with soap and water before and after performing this activity. Clean up the work area with soap and water after completing this investigation. Keep pets and children away from lab materials and equipment.

Preparation

1. Read through the procedure.
2. Obtain all the materials.
3. Clean and sanitize the work area.
4. If using the oven dry method, turn the oven on and set the temperature to 225 °F.

ACTIVITY 1

A Single Replacement Reaction Stoichiometry and Percent Yield

1. Use the 50-mL graduated cylinder, measure all the 1 M $CuSO_4$ in the bottle, and place it in the 250-mL beaker. Record the total volume in the data table.

> The bottle of copper sulfate will contain approximately 100 mL of solution. Use the 50-mL graduated cylinder. Measure 50 mL of solution and add it to the beaker. Measure out another 50 mL of solution and add that to the beaker. If there is less than 50 mL, record the volume before pouring. If some solution remains, measure that and add it to the beaker. The goal is to measure and record the starting volume of 1 M $CuSO_4$.

2. Add some of the NaCl to the $CuSO_4$ solution using a teaspoon.

> The sodium chloride is a catalyst so the **exact amount is not important**. One salt packet from a fast food restaurant will be sufficient.

3. Stir the solution with the plastic spoon to dissolve the NaCl.

4. Cut a 4 × 12" piece of aluminum foil.

5. Weigh and record the aluminum foil mass in the data table.

> For best results, use between 1.4 and 1.6 g of aluminum foil. If the aluminum mass is too high, tear a little off until it is within this range. If the mass is too low, add some more foil. The aluminum foil does not have to be one piece for this reaction. When weighing the aluminum foil, fold it up so it fits in/on the electronic balance pan.

6. Tear the aluminum foil into small pieces (approximately 1 × 1") and add them to the $CuSO_4$ – NaCl solution.

> **CAUTION:** The solution will get hot, but not hot enough to harm the plastic beaker.

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ACTIVITY

7. With a plastic spoon, stir to mix the aluminum foil in the solution.
8. Record all observations.
9. Let the reaction mixture sit for approximately 10 minutes. Stir occasionally with the plastic spoon to ensure all the aluminum reacts.

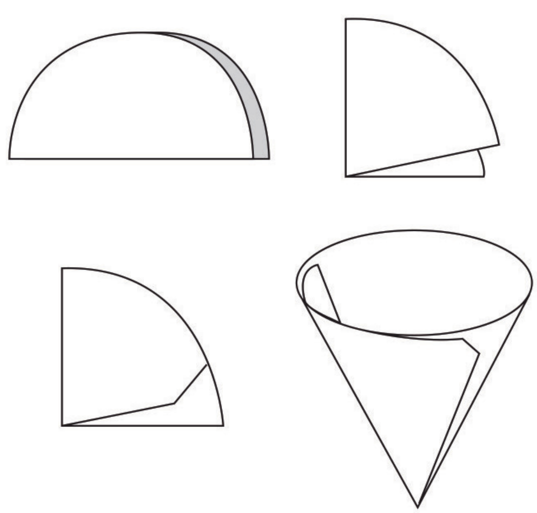
> The aluminum foil will float on top of the surface. Keep stirring to expose the aluminum on the foil until no more remains. Do not proceed to the filtration step if any aluminum is left.

10. After all the aluminum has reacted, let the solution sit for another 5 minutes, allowing the copper to settle to the bottom of the beaker.

Gravity Filtration

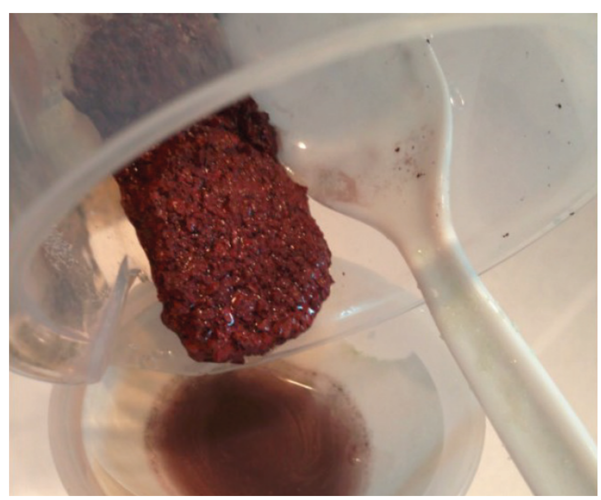
11. Prepare the filtration set-up.
    a. Fold the filter paper in half and then in half again. Tear a small piece from the corner of the filter paper. See Figure 1 below.

Figure 1.



    b. The folded filter paper has four curved edges, two interior and two exterior. Gently separate an exterior edge from the three remaining edges until it makes a cone. Place the filter paper cone into the funnel.
    c. Wet the filter paper with pure water and push the filter paper down into the funnel until it is flush with the walls.
    d. Place the filter funnel on top of the 50-mL graduated cylinder.
12. When no more aluminum foil is observed and the copper has settled to the bottom of the beaker, carefully decant the solution into a plastic cup. Use the spoon to keep all the solid copper in the beaker. Pour out as much of the solution as possible without losing any copper solid.
13. Add approximately 20 mL of pure water to the beaker containing the copper solids. Stir the copper suspension with the spoon and pour it in the filter funnel.
14. Transfer as much copper into the filter funnel as possible; use the spoon to help. See Figure 2.

Figure 2.



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Activity continued

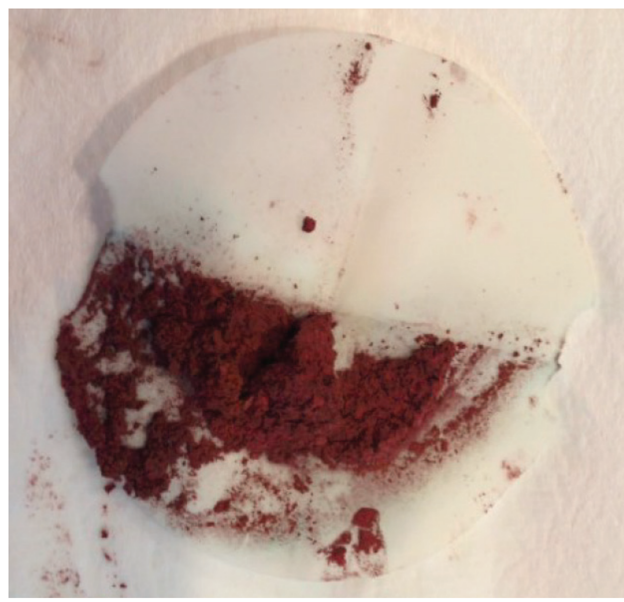
15. Add another 10 mL of water to the beaker, swirl, and pour the rest of the water and copper solids into the filter funnel. You may need to empty the graduated cylinder to be able to complete the rinsing.
16. It may take 10–15 minutes for the liquid to pass through the filter.
17. Use a plastic pipet and wash the copper solid with another 20 mL of pure water. Also add the water around the outside edges of the filter paper. This will wash away most of the remaining sulfate solution. Any remaining sulfate solution will crystallize during the drying step and add to the weight of the final product.
18. Wait another 10 minutes for the wash to pass through the filter. See Figure 3.

Figure 3.

CAUTION: The filter funnel gravity filtration set-up placed on top of the graduated cylinder is very top-heavy and prone to tipping over. Hold on to the cylinder base or weigh it down with heavy waterproof items to prevent it from tipping over.

19. Carefully remove the filter paper from the filter funnel.
20. Unfold the filter paper and lay it flat on several layers of paper towels to assist in the drying of the copper metal. See Figure 4.

Figure 4.



Drying the Product

The product must be thoroughly dried before the product is weighed. There are two recommended drying options. Using an oven or allowing it to air dry overnight. Do not use a microwave to dry the copper metal.

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ACTIVITY

A. Oven Method

1. Leave the wet filter paper on the paper towels for 10–15 minutes to wick away some of the moisture.
2. Preheat the oven to about 225 °F. Do not exceed 250 °F.
3. Place the filter paper on a sheet of foil. Place the foil on a cookie sheet or pie pan.
4. Spread the moist copper product out using the spoon. The greater the surface area exposed, the quicker the copper will dry. Be sure to remove any lingering copper from the spoon and onto the filter paper.
5. Heat the copper metal and filter paper for 25–30 minutes to remove all the water.
6. If the filter paper starts to char, remove it from the oven. The copper should break apart easily when it is thoroughly dry. See Figure 5.

Figure 5.

7. Remove the copper metal and filter paper from the foil and allow it to cool a few minutes.
8. Tare a weigh boat on the electronic balance.
9. Transfer the copper metal from the filter paper to the weigh boat.
10. Weigh the copper metal and record the mass in the Data Table.

B. Overnight Air-Dry Method

1. Keep the wet filter paper on the paper towels and place it in a well-ventilated but secure location for 24 hours.
2. Occasionally replace the paper towels to facilitate the drying process.
3. Spread the moist copper product out using the spoon. The greater the surface area exposed, the quicker the copper will dry.
4. When the copper metal breaks apart easily with the plastic spoon, it should be dry enough for weighing. This takes at least 24 hours, depending on the humidity and air flow around the sample.
5. Tare a weigh boat on the electronic balance.
6. Transfer the copper metal from the filter paper to the weigh boat.
7. Weigh the copper metal and record the mass in the Data Table.

Disposal and Cleanup

1. Dispose of all used sulfate solutions and filtrates in a sink with the water running. Allow the faucet to run a few minutes to dilute the solutions.
2. Clean and dry the beaker, graduated cylinder, filter funnel, pipet, and spoon before returning them to the equipment kit.
3. Sanitize the workspace.

Data Table 1: Single-Replacement Reaction of Aluminum and Copper Sulfate

	Answer	Show Your Work
Balanced chemical equation		
Volume of 1 M $CuSO_4$		
Mass of Al foil (from Step 5)		
Moles of $CuSO_4$ (using the method from Step 3)		
Moles of Al (using the method from Step 3)		
Moles of Cu product based on starting $CuSO_4$ (using the method from Step 4)		
Moles of Cu product based on starting Al (using the method from Step 4)		
Limiting reactant (using the method from Step 4)		
Theoretical yield of Cu in moles (using the method from Step 5)		
Theoretical yield of Cu in grams (using the method from Step 5)		
Actual yield of Cu in grams		
Percent yield of Cu (using the method from Step 7)		

NOTES

CHEMISTRY

Single Replacement Reaction Stoichiometry and Percent Yield Investigation Manual

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