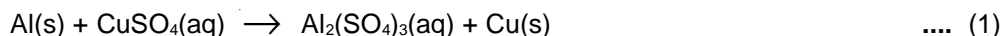


# The Stoichiometry of the Reaction Between Aluminum Metal and Copper (II) Sulphate

The object of this experiment is to find experimentally the mole ratio existing in the reaction between aluminum metal and copper (II) sulphate. The experiment has two distinct parts.

In the first part, you will react a piece of aluminum foil with an excess of copper (II) sulphate solution to produce copper metal. The unbalanced reaction equation is:



After filtering and washing the copper, you will react the copper metal with nitric acid, producing a solution containing copper ions.

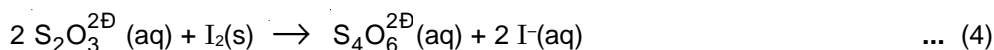


The colourless NO(g) immediately reacts with the air to form red-brown NO<sub>2</sub>(g). After neutralizing the excess nitric acid, you will dilute the copper ion solution to exactly 250.0 mL.

In the second part, you will find the concentration of copper ions in the diluted volume using a titration procedure. The titration procedure is as follows. First, a large excess of iodide ion is added to a sample having a known volume of dissolved copper ions, producing the following reaction.



The I<sub>2</sub> produced in this initial reaction is then titrated with a solution containing thiosulphate ions (S<sub>2</sub>O<sub>3</sub><sup>2-</sup>) having a known concentration, producing the following reaction.



Once the volume of thiosulphate solution is known, the moles of thiosulphate can be calculated and eventually you can find the moles of copper that originally reacted with the aluminum.

## PROCEDURE

1. Obtain a square of aluminum foil (about 10 cm x 10 cm) and record the mass. This mass should be in the range 0.37 g to 0.43 g; if the mass is outside this range, add or remove small bits of foil until the mass is in the range. **Record the mass of the aluminum you use.** Tear the aluminum foil into 25-50 small pieces and place them in a clean 250 mL beaker.
2. Obtain 60 mL of 0.5 M CuSO<sub>4</sub> solution and place it in a separate 250 mL beaker. Obtain about 1 g of solid sodium chloride, NaCl, and stir it into the CuSO<sub>4</sub> solution. When the NaCl has dissolved, pour the CuSO<sub>4</sub>/NaCl solution into the beaker containing the pieces of aluminum foil and stir slowly until all the aluminum has dissolved. (One or two tiny specks of aluminum left over is not a problem; many specks mean the reaction must continue for another minute or so.)
3. Filter the solid copper, taking care not to lose even a single speck of the red-brown powder. Use a wash bottle to get all the copper out of the beaker and into the filter. After the liquid from the reaction mixture has drained through the filter paper, wash the copper with two 50 mL portions of distilled water to ensure the copper is free from contamination. After the last of the water has drained through, wash the copper with about 50 mL of acetone to remove the last traces of water. (Washing with acetone lets the copper dry quickly and makes it easier to get the copper off the filter paper in the next step.)
4. Clean and dry the 250 mL beakers used above and scrape the copper off the filter paper into one of the beakers. Make sure to get as much as possible of the copper into the beaker – the more you leave behind, the more it will affect your results. Next, take the beaker to a fume hood and add 8.0 mL of 9 M nitric acid, HNO<sub>3</sub>. The reaction gives off red-brown NO<sub>2</sub> gas and leaves a dark green solution. When the reaction appears to be finished, carefully examine the bottom of the beaker to make sure all the copper has reacted. (CARE! Don't inhale any fumes coming out of the liquid.) A few small specks of unreacted solid are not a problem but larger amounts require you to wait another minute or two until the reaction finishes. When the reaction is complete, add enough distilled water to bring the volume up to about 50 mL.

5. The reactions in steps 9 to 11 are spoiled in the presence of excess nitric acid, so you must neutralize the excess acid. Obtain about 10 mL of 1 M sodium carbonate solution,  $\text{Na}_2\text{CO}_3$ , and a dropping pipette. While constantly swirling the beaker, use the dropper to add  $\text{Na}_2\text{CO}_3$  in a relatively quick drop-by-drop fashion to the acidic copper solution. A precipitate will alternately form and then dissolve as the  $\text{Na}_2\text{CO}_3$  reacts with the excess acid. Continue adding  $\text{Na}_2\text{CO}_3$ , with swirling, until a few flecks of precipitate remain undissolved. (This will probably take about 6-8 mL and definitely less than 10 mL of  $\text{Na}_2\text{CO}_3$ .) Next, add a few drops of 1 M acetic acid,  $\text{CH}_3\text{COOH}$ , until the large flecks of precipitate have just dissolved – ignore any slight cloudiness or flecks of unreacted copper that were present before starting to add the  $\text{Na}_2\text{CO}_3$ .
6. Carefully pour the copper solution into a 250 mL volumetric flask, using a wash bottle containing distilled water to get ALL the copper solution out of the beaker and into the volumetric flask. Stopper the flask and shake well to ensure proper mixing.
7. Obtain about 100 mL of sodium thiosulphate solution,  $\text{Na}_2\text{S}_2\text{O}_3$ , in a clean and absolutely dry 250 mL beaker. **Record the molarity of the  $\text{Na}_2\text{S}_2\text{O}_3$** , which will be about 0.1 M. Wash out a burette with two 5 mL portions of  $\text{Na}_2\text{S}_2\text{O}_3$ , being certain to drain some of the wash liquid out the burette tip. Drain the last of the wash liquid and fill the burette with  $\text{Na}_2\text{S}_2\text{O}_3$  solution.
8. **Record the initial volume to the nearest 0.01 mL.**
9. Pipette exactly 25.00 mL of the diluted copper solution into a clean 250 mL beaker. Obtain 1 g of potassium iodide, KI, and dissolve the solid in about 10 mL of distilled water. After the solid is dissolved, pour the potassium iodide solution into the 25 mL portion of copper solution and swirl to complete the reaction. The mixture should be opaque chocolate brown.
10. Add  $\text{Na}_2\text{S}_2\text{O}_3$  solution from the burette into the beaker containing the chocolate-brown copper solution, with constant slow swirling, until the solution's colour becomes pale yellow-brown. CARE! The closer the solution gets to a pale yellow-brown colour the slower you must add the solution, eventually adding the solution in a drop-by-drop fashion. Before adding any more  $\text{Na}_2\text{S}_2\text{O}_3$ , add 2 mL of 1% starch solution to the mixture in the beaker and swirl to mix. (IMPORTANT – Do not add the starch solution until the mixture is pale yellow or your final result will not be as accurate as it should be.) After adding the starch, the mixture should become blue-purple. Now continue adding  $\text{Na}_2\text{S}_2\text{O}_3$  drop by drop, with swirling, stopping when the purple colour disappears and the mixture becomes pale brown in colour. **Record the final volume in the burette to the nearest 0.01 mL.**
11. Refill the burette and repeat steps 8 to 10. If the total volumes used in the 1st and 2nd titration are within 0.2 mL of each other, you are finished. If the volumes are more than 0.2 mL different, refill the burette and repeat steps 8 to 10 a third time and check if the 3rd volume is within 0.2 mL of the volume of either of the 1st or 2nd titration. If not, consult your teacher for advice.

## CALCULATIONS AND QUESTIONS

1. (a) Average the two closest volumes of  $\text{Na}_2\text{S}_2\text{O}_3$  used in the titrations to calculate the **average volume of  $\text{Na}_2\text{S}_2\text{O}_3$**  used. (Show your work.)
- (b) Use the molar concentration of the  $\text{Na}_2\text{S}_2\text{O}_3$  solution and the **average volume of  $\text{Na}_2\text{S}_2\text{O}_3$**  found in calculation 1(a) to calculate the moles of  $\text{Na}_2\text{S}_2\text{O}_3$  used. This result is the moles of  $\text{S}_2\text{O}_3^{2-}$  used in equation (4).
- (c) Use equation (4) and the moles of  $\text{Na}_2\text{S}_2\text{O}_3$  found in calculation 1(b) to calculate the moles of  $\text{I}_2$  reacted by the  $\text{Na}_2\text{S}_2\text{O}_3$ .
- (d) The moles of  $\text{I}_2$  reacted in equation (4) equals the moles of  $\text{I}_2$  produced in equation (3). Use this fact to calculate the moles of  $\text{Cu}^{2+}$  reacted in equation (3).
- (e) The moles of  $\text{Cu}^{2+}$  found in calculation 1(d) represents the moles of  $\text{Cu}^{2+}$  in each 25.00 mL portion of solution taken from the 250 mL volumetric flask. Use this information to calculate the total moles of  $\text{Cu}^{2+}$  contained in the 250 mL flask.

(f) The total moles of  $\text{Cu}^{2+}$  contained in the 250 mL volumetric flask equals the moles of  $\text{Cu(s)}$  reacted in equation (2) and also equals the moles of  $\text{Cu(s)}$  produced in equation (1) and the moles of  $\text{CuSO}_4$  reacted in equation (1). How many moles of  $\text{CuSO}_4$  reacted in your experiment? Note that you used an EXCESS of  $\text{CuSO}_4$  solution and that not all of the  $\text{CuSO}_4$  solution added in step 2 actually reacted; only the  $\text{CuSO}_4$  that reacted with the aluminum produced copper metal.

2. Use the mass of aluminum foil reacted to calculate the moles of aluminum used in the experiment.
3. Use the moles of  $\text{CuSO}_4$  that reacted and the moles of aluminum that reacted to calculate the value of the following ratio:

$$\text{mole ratio} = \frac{\text{moles of CuSO}_4 \text{ reacted}}{\text{moles of Al reacted}} .$$

4. Numbers such as 1, 2, 3, etc. are called "integers" and numbers such as 0.5, 1.5, 2.5, etc. are called "half-integers". Round the value for the mole ratio found in calculation 3 to the nearest integer or half-integer.
5. Balance equation (1). What type of reaction is represented by this equation?
6. Does the mole ratio you found in calculation 4 agree with the mole ratio predicted from your balanced equation (1)? Explain your answer.