**Stoichiometry of Iron-Copper (II) Sulfate Redox Reaction**

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**ABSTRACT**

Stoichiometry was used to determine the correct reaction equation between iron metal and a solution of copper (II) sulfate. The volume of copper (II) sulfate was calculated by finding the amount of copper (II) sulfate used to react with 1.0 gram of iron. The measured copper (II) sulfate solution was heated on a hotplate and then mixed with iron. Once the reaction has occurred, the solid copper was washed with distilled water and acetone and then dried on the hotplate. The mass of the copper metal was weighed and used to calculate the ratio of moles of copper produced to moles of iron reacted. The ratios for the three trials ended up being close to 1 which meant the correct reaction equation was Fe (s) + Cu2+ (aq) 🡪 Fe2+ (aq) + Cu (s).

1. **INTRODUCTION**

The purpose of the experiment is to find out which of the two balanced reactions between iron and copper (II) sulfate is correct:

Fe (s) + Cu2+ (aq) 🡪 Fe2+ (aq) + Cu (s) [1]

2 Fe (s) + 3 Cu2+ (aq) 🡪 2 Fe3+ (aq) + 3 Cu (s) [2]

In other words, we are trying to find out whether the iron ion is iron (II) or iron (III). This reaction is an oxidation-reduction reaction which means copper (II) gains electrons from the iron and is reduced. The iron loses electrons to the copper and is oxidized. We know that the redox reaction will create copper metal as a product so we can easily calculate the ratio of how much copper metal is produced to how much iron metal is reacted to find out which of the two reactions above is correct. The final result will be a ratio number close to 1 or 1.5 and that number will tell us which reaction is correct.

1. **EXPERIMENTAL**

For this experiment, the chemicals we used were iron (brown sand) and copper (II) sulfate (blue aqueous solution). I calculated the volume of copper (II) sulfate that would react with 1.0 gram of iron which is 4.73 grams. I weighed 1 gram of iron powder by using an electronic balance and placing a piece of paper on it and taring it. I poured the measured iron powder into a 150 mL beaker. I measured 4.7 mL of CuSO4 in a 10 mL graduated cylinder and poured the CuSO4 into a 250 mL Erhlenmeyer flask and began heating it on a hot plate to 85 degrees Celsius. After the CuSO4 was heated to a dark blue color, I poured it into the 150 mL beaker containing the iron powder and decanted the liquid from the beaker using a dropper after the reaction stopped. The metal that remained was a bright orange color. I poured 10 mL of distilled water twice into the copper metal remaining in the beaker and decanted that water. Then, I poured 5 mL of acetone into the beaker twice and decanted the liquid again. The copper metal was then dried on the hot plate and weighed with a piece of paper on an electronic balance.

The materials used during the course of this experiment were: electronic balance, weighing paper, beaker, graduated cylinders, Erhlenmeyer flask, hot plate and dropper.

1. **RESULTS AND DISCUSSION**

In this experiment, the goal is to find which of the two reactions given is correct. The results I achieved at the end of the experiment were three ratios from the three trials of moles of copper produced to moles of iron reacted. The trial one ratio was exactly 1.0, the trial two ratio was 1.1 and the trial three ratio was 0.9- (found in the calculations on the yellow pages). According to the balanced equations, the ratio you could obtain could either be 1.0 or 1.5. My answers are closest to 1.0 which indicates that the first equation or Fe (s) + Cu2+ (aq) 🡪 Fe2+ (aq) + Cu (s) is the correct one since there is a one to one ratio of iron and copper on reactant side and the product side. As with the second equation, 2 Fe (s) + 3 Cu2+ (aq) 🡪 2 Fe3+ (aq) + 3 Cu (s), there is 1 iron for every 1.5 copper.

The results that I achieved were from using a volume of 4.7 mL copper (II) sulfate. I am aware that there were other individuals who used volumes of 20 mL and more and thus their results could either be similar or different. Before the lab started, this volume was calculated because we knew how much grams of iron we were going to use. We could convert that number to moles of iron used and compare with one of the equations to find out how much copper (II) sulfate is used.

If a 5% loss of copper produced was lost, the observed ratios would all decrease 0.1 less. So instead of the ratios being 1.0, 1.1 and 0.9, they would become 0.9, 1.0 and 0.8. However, the results would still tell us that the correct equation is equation number one. A 10% loss of copper produced would just change the ratios much more than a 5% loss would. Perhaps if my ratios were in between 1.0 and 1.5, the assumptions that a percentage of copper was lost would definitely change the answer and determining which equation was correct would be a lot harder. Nonetheless, my ratios are very close to 1.0 and therefore, the correct equation is Fe (s) + Cu2+ (aq) 🡪 Fe2+ (aq) + Cu (s).

1. **CONCLUSION**

The results obtained from this experiment provide us with an idea of how to use stoichiometry to find ratios of substances in a reaction. In this particular experiment, there were two possible answers that could be obtained. We would only be able to find these answers if we understood what a redox reaction was and knew how to use ratios to find an unknown value from a known value. These concepts are especially useful in stoichiometry and allowed us to determine how much copper was produced from using a certain amount of iron. Moreover, these concepts would be useful in many fields of science and allow anyone to calculate the appropriate ratios of certain reactions and to create successful reactions from those ratios.

Table I. For three trials, the mass of iron powder was weighed in grams and the moles of iron and volume of copper (II) sulfate to use were calculated based on the mass of the iron powder. The mixture of iron and copper (II) sulfate yielded the mass of copper produced as well as the moles of copper produced.

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|  | Trial #1 | Trial #2 | Trial #3 |
| Mass of Iron Powder | 1.000 g | 1.000 g | 1.069 g |
| Moles of Iron | 1.8 \* 10-2 | 1.8 \* 10-2 | 1.9 \* 10-2 |
| Volume of 1.0 M Copper (II) Sulfate Used | 4.7 mL | 4.7 mL | 4.7 mL |
| Mass of Copper Produced | 1.228 g | 1.120 g | 1.168 g |
| Moles of Copper produced | 1.9 \* 10-2 | 1.8 \* 10-2 | 1.8 \* 10-2 |