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Article

Impact of aluminum surface area on the rate of reaction with aqueous copper (II) chloride solutions

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SUMMARY

The tests in this study sought to improve current understanding regarding the efficient use of aluminum and $CuCl_2$ in production chemistry. We tested whether more concentrated solutions of $CuCl_2$, when mixed with pure aluminum, would yield more rapid temperature increases; whether aluminum with a smaller surface area would produce a lower final temperature than the mixtures with a bigger surface area of aluminum; and whether the temperature increase over an elapsed time for all mixtures would be linear. Throughout this experiment, we investigated the change in reaction temperature over the course of 120 seconds, with 30 second timepoints, following the introduction of either strips or grains of pure aluminum to variable concentrations of aqueous $CuCl_2$ to gain insight into the role that metal surface area plays in the rate of this single replacement reaction. We found that, for the aluminum strips, the experiments demonstrated an expected increase in temperature; however, granular aluminum did not follow this same trend. We demonstrated that when $CuCl_2$ solutions were mixed with aluminum strips, the change in temperature followed a linear relationship, however when $CuCl_2$ was mixed with granular aluminum, the 0.33 M mixture created a linear trend, but the 0.66 M mixture and the 1.00 M mixture did not show a trend. Therefore, we concluded that by collision theory, the greater surface area inherent to equal masses of granular aluminum allowed a faster reaction rate compared to the aluminum strips.

INTRODUCTION

Through this study, we hoped to gather new information pertinent to production chemistry. By exploring the impact of different concentrations of $CuCl_2$ solution and two different surface areas of 3003 grade aluminum on reaction rate, production methods involving aluminum and various solutions could be improved. Therefore, the large-scale manufacturing of chemicals like $AlCl_3$ could be done more efficiently. The significance of such a chemical could extend to medications and food production.

The stoichiometric equation for the reaction of $CuCl_2$ with aluminum is $2Al + 3CuCl_2 \rightarrow 2AlCl_3 + 3Cu$, which is a classic single replacement reaction. Furthermore, when $CuCl_2$ is mixed with aluminum, the reaction is exothermic and heat energy is generated. This indicates that as the processes of bond breaking and bond formation take place, the reactants proceed from a state of high potential energy to products possessing lower potential energy. This discrepancy of energy is ultimately released into the surrounding aqueous solution as heat and the resulting solution temperature increases accordingly (1).

Surface area is one of the multitude of factors that influence reaction rate. In this specific case, as the surface area of aluminum increases, an increase in reaction rate should follow (2). The greater the surface area of a solid-phase reactant, such as a metal, the more reactant sites are exposed, translating to a greater chance of the necessary chemical entities colliding with each other. A greater chance of collisions, according to collision theory, directly results in a greater chance of a productive reaction, thus leading to an increase in the overall rate of reaction (3).

Through this experiment, we aimed to demonstrate how the temperature changed when we mixed 1.00 gram of aluminum in the form of strips or granules into different concentrations of $CuCl_2$ solution. We used pure aluminum, not aluminum alloy, in this experiment to make sure that there were the same number of moles of aluminum in the strips and granules. Furthermore, this data should provide direct insight into the reaction rate. We hypothesized that mixing aluminum with more concentrated $CuCl_2$ would produce more rapid increases in temperature based on an increase in reactant collisions. We also hypothesized that the mixtures with aluminum strips would produce a lower final solution temperature than the mixtures with granular aluminum due to the fact that the strips possessed a surface area three times lower than that of the granular aluminum (approximately 2.54 cm^2 vs. 7.53 cm^2, respectively, for 1.00 g of material). Furthermore, with respect to the mathematical fit of individual reaction rate for each specific mixture, we hypothesized that the temperature increase over the course of 120 seconds for all mixtures would be linear, representing a constant turnover of products, and therefore heat energy, per unit of time.

RESULTS

To conduct our research, we prepared three different concentrations of $CuCl_2$ and mixed each with aluminum strips and granular aluminum, monitoring the temperature of each reaction using a Logger Pro probe over the time course of 120 seconds. From these experiments, we found

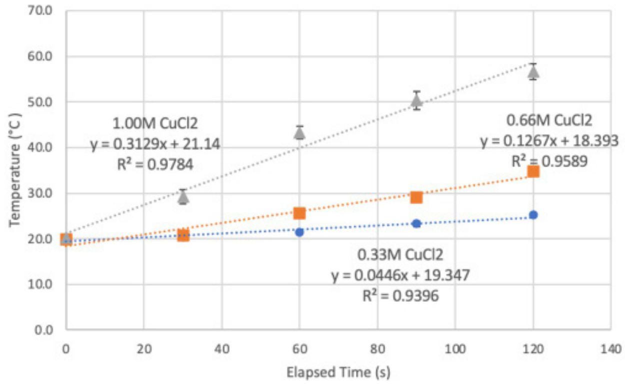

Figure 1: Elapsed time (s) vs. temperature (°C) for reaction with aluminum strips. We recorded the solution temperature for the reaction of aluminum strips with 0.33 M (blue), 0.66 M (orange), and 1.00 M (silver) concentrations of copper (II) chloride (CuCl2) (n=3, averaged). The temperature was taken at 30, 60, 90, and 120 seconds after the aluminum was mixed. The error bars represent average absolute deviation (AAD).

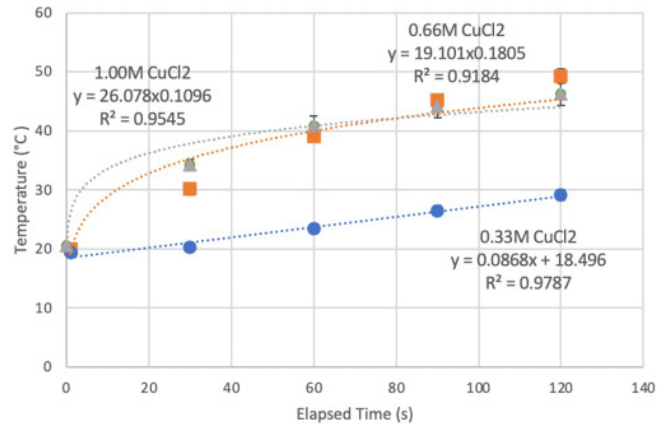

Figure 2: Elapsed time (s) vs. temperature (°C) when different concentrations of copper (II) chloride were mixed with 1.00 g of granular aluminum. Logger Pro Temperature Probe was used to measure the temperature of the mixture. Elapsed time was measured using a stopwatch. Temperature was taken 30, 60, 90, and 120 seconds after the aluminum was mixed. Data points come from the average of three replicates for each concentration. Error bars represent AAD.

that the temperature increased in a positive linear fashion when the different concentrations of $CuCl_2$ were mixed with the aluminum strips (**Figure 1**). The aluminum strips mixed with the 0.33 M $CuCl_2$ had a slope of 0.04 °C/s, the aluminum mixed with the 0.66 M CuCl2 had a slope of 0.13 °C/s, and the aluminum mixed with the 1.00 M $CuCl_2$ had a slope of 0.31 °C/s. The mixtures with a higher concentration of $CuCl_2$ had a greater slope and increased more rapidly than those with a smaller concentration (**Figure 1**). For all three different concentrations of $CuCl_2$, the R^2 values were close to 1 (0.94, 0.96, and 0.98 for the 0.33 M, 0.66 M, 1.00 M mixtures, respectively), showing that the data indeed demonstrated a reasonable linear trend.

The results when granular aluminum was mixed with 0.33 M $CuCl_2$ showed a linear trend with a slope of 0.09 °C/s. Though the granular aluminum mixed with 0.66 M and 1.00 M $CuCl_2$ did not show a well-fitted trendline, we used the first few data points to find a rough estimate of the slopes: for the granular aluminum mixed with 0.66 M $CuCl_2$, the slope was approximately 0.25 °C/s and for the granular aluminum mixed with 1.00 M $CuCl_2$, the slope was approximately 0.20 °C/s (**Figure 2**). The R^2 of the trendlines for the 0.33 M concentration of $CuCl_2$ with granular aluminum was 0.99. Also, the data for the granular aluminum shows the temperature somewhat plateauing for the 0.66 M and 1.00 M concentration of $CuCl_2$ (**Figure 2**).

DISCUSSION

Our experiment investigated how the surface area of aluminum would affect the reaction rate of aqueous copper (II) chloride solution. Our hypothesis was partially supported, as an increase in the molarity of $CuCl_2$ generally increased the reaction temperature across both types of aluminum. However, the relationship between the molarity of $CuCl_2$ and temperature was less linear in the set of experiments using granular aluminum.

For the mixture with aluminum strips and $CuCl_2$, the temperature increased more rapidly as the concentration of $CuCl_2$ increased (**Figure 1**). The reasoning behind the highly concentrated $CuCl_2$ reacting more actively comes from collision theory, as increasing the concentration of reactants will increase the frequency of collisions between the reactants, leading to a greater rate of reaction and more rapid rise in temperature (4). The linear model was appropriate because the increase in temperature was mostly uniform for the two minutes for all 0.33 M, 0.66 M, and 1.00 M $CuCl_2$ mixed with aluminum. The R^2 values of 0.94, 0.96, and 0.98, respectively, affirm the linear relationship.

For the mixture of granular aluminum and $CuCl_2$, the mixtures with higher concentrations of $CuCl_2$ generally had a greater change in temperature, but there was an exception for the mixture with 0.66 M $CuCl_2$ as the mixture had a higher temperature than the 1.00 M mixture at 120 seconds, possibly due to a plateauing effect seen in the 0.66 M and 1.00 M samples after 60 seconds. (**Figure 2**). One possible explanation for this result is that the mixture with the 0.66 M and 1.00 M had a faster reaction due to collision theory—caused by its higher concentration—within the first 30 seconds, but then had a decreased rate of reaction during the final 90 seconds due to limited remaining reaction sites. Thus, the increase in temperature was lower towards the end of the reaction and seemingly resulted in a plateau. We chose a linear fit for the 0.33 M $CuCl_2$ because the temperature increased at a constant rate. However, for the 0.66 M $CuCl_2$ and 1.00 M $CuCl_2$, we estimated a trend line using the first few data points as the reaction is very active in the first 30 seconds but slowed down substantially at longer times. For this experiment, the Average Absolute Deviation (AAD) values were low, ranging from 0.06 to 2.04 as the experiment didn't have much uncertainty. An alternative explanation for the similarities in the data with the mixture of 0.66 M $CuCl_2$ and the 1.00 M $CuCl_2$ with granular aluminum is that both mixtures of $CuCl_2$ had excess $CuCl_2$ which caused the 1.00 M $CuCl_2$ solution to react similarly to granular aluminum as the 0.66 M $CuCl_2$.

In this study, some sources of uncertainty were the difference in room temperature because we conducted this experiment over a couple of days, and the starting temperature was not always 20 °C. The largest difference in starting temperature was 0.5 °C. However, the difference in room temperature did not seem to markedly impact the final data. Another potential source of uncertainty was the speed of the stir bar. When we stirred the mixture with granular aluminum the speed seemed constant, but when the stir bar stirred the aluminum strip, we often heard clanking noises and the aluminum strip possibly slowed down the stirring speed. If we had used a larger beaker for our testing, we could have avoided this problem. However, the speed of the stirring didn't seem to have a large effect on the final data.

A potential alternative experiment would be mixing steel wool of different surface areas with different concentrations of acetic acid. This reaction would also create an exothermic reaction and the wool would rust, just like how aluminum did when we mixed it with $CuCl_2$ (5). We would use the same concentrations of 0.33M, 0.66M, and 1.00 M $CuCl_2$ and see if we get the same results as the aluminum with $CuCl_2$ experiment or varied results. Such a study would have implications for production chemistry as well, particularly with regard to other acids and materials, which would have different purposes in production chemistry ranging from commercial usage to environmental and agricultural products.

MATERIALS AND METHODS

To prepare the solutions, two hundred fifty mL beakers were filled with fifty mL of the corresponding concentration of $CuCl_2$ (0.33 M, 0.66 M, and 1.0 M). We set up three beakers for each concentration, and we used a total of 18 beakers: three beakers each with, 0.33 M, 0.66 M, and 1.0 M with aluminum strips and three beakers each with 0.33 M, 0.66 M, and 1.0 M with granular aluminum. Pure aluminum (Fisher Scientific) was used in this experiment, and granular aluminum particles had a radius of about 0.2mm while aluminum strips had lengths of about 2.54cm and a thickness of about 0.5mm. We measured the starting temperature using a Logger Pro Temperature Probe and then measured the temperature at 30 second increments for two minutes after we put the aluminum inside the beaker. To measure elapsed time, we used a stopwatch. Microsoft Excel was used to determine trendlines and R^2 values.

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