Amazing Student

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Chemistry

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Theoretical Versus Actual Yield Lab

Background: The purpose of this experiment was to demonstrate how human error can result in an actual yield that differs from the theoretical yield. Whereas the theoretical yield is "the maximum amount of products that can be produced from the reactants," the actual yield is "the amount of product that is actually formed when a reaction is carried out" (CK-12 Foundation). Both theoretical and actual yield are related to the Law of Conservation of Mass which states that matter can be neither created nor destroyed (Sterner et al.). The mass of the reactants must equal the mass of the products, whether the actual yield is close to the theoretical yield or not.

Question: How does the experimental mole ratio compare with the theoretical mole ratio?

Hypothesis: The actual yield after the reaction between Iron and copper(II) sulfate will be less than the theoretical yield.

Materials:

- Copper (II) sulfate pentahydrate ($CuSO4.5H_2O$)
- Iron metal filings (20 mesh)
- Distilled water
- 150-mL beaker
- 100-mL graduated cylinder
- Bunsen burner
- Beaker
- Tongs
- Balance
- Stirring rod
- 400-mL beaker
- Weighing paper

Procedures:

1. Measure the mass of a clean, dry 150-mL beaker. Record all measurements in a data table.
2. Place approximately 6 g $CuSO4.5H_2O$ into the 100-mL beaker, and measure the combined mass.
3. Add 25-mL of distilled water to the $CuSO4.5H_2O$.
4. Place the mixture above the flame, and stir until all of the solid dissolves (do not boil). Using tongs, remove the beaker from the hot plate.
5. Measure about 1 g of iron filings onto a piece of weighing paper. Measure the mass of the fillings.
6. While stirring, slowly add the iron filings to the hot copper(II) sulfate solution. Be careful not to splash the hot solution,
7. Allow the reaction mixture to sit for 5 minutes.
8. Use the stirring rod to decant (pour off) the liquid into a 400-mL beaker. Be careful to decant only the liquid- leave the solid copper metal behind.
9. Add 15 mL of distilled water to the copper solid, and carefully swirl the beaker to wash the copper. Decant the liquid into the 400-mL beaker.
10. Repeat Step 9 two more times.
11. Place the beaker containing the wet copper on the hot plate. Use low heat to dry the copper.
12. After the copper is dry, use the tongs to remove the beaker from the hot plate, and allow it to cool.
13. Measure the mass of the beaker and the copper.

14. The dry copper can be placed in a waste container. Moisten any residue that sticks to the beaker and wipe it out using a paper towel.
15. Pour the unreacted copper (II) sulfate and iron (II) sulfate solutions into a large beaker. Return all lab equipment to its proper place.

Data:

Table 1: Trials 1, 2, and 3

	Beaker (g)	Beaker and Copper (II) Sulfate (g)	Iron (g)	Iron (mol)	Beaker and Copper (g)	Copper (g)	Copper (mol)	Iron to Copper Mole Ratio	Percent Error (%)
Trial 1	80.12	86.15	1.00	0.0179	81.15	1.03	0.0162	1.00:0.905	9.65
Trial 2	80.14	86.15	1.00	0.0179	81.09	0.95	0.0149	1.00:0.832	17.00
Trial 3	80.13	86.11	1.00	0.0179	81.09	0.95	0.0149	1.00:0.832	17.00

Quantitative Data:

The chemical equation for the reaction was CuSO4 + Fe → Cu + FeSO4. For each trial we used approximately 6.00g copper (II) sulfate and 1.00g iron. Given these amounts of reactants, the actual yield of copper should have been 1.14g which would have resulted in a theoretical mole ratio of 1 mol Fe: 1 mol Cu. Due to human error, we had an average actual yield of 0.98g of copper with an average percent error of 14.55% (See Table 1). The calculated moles of iron and copper resulted in an average actual mole ratio of 1 mol Fe: 0.850 mol Cu. This ratio makes sense because our actual yield of copper was less than the theoretical yield, and therefore, our mole ratios were lower as well.

Qualitative Data:

After dissolving the copper (II) sulfate in 25 mL of distilled water, we added 1.00g of iron to the solution. After five minutes, a layer of solid, dark brown rust formed on top of the bluish iron (II) sulfate solution, and the reddish-brown Copper sunk to the bottom.

In the reaction between copper (II) sulfate and iron, copper (II) sulfate was the excess reactant. This was demonstrated both in the analysis of the chemical equation as well as the observation of the product that was yielded after the reaction. After the copper (II) sulfate and iron reacted, the liquid remained blue because not all of the copper (II) sulfate was used during the reaction.

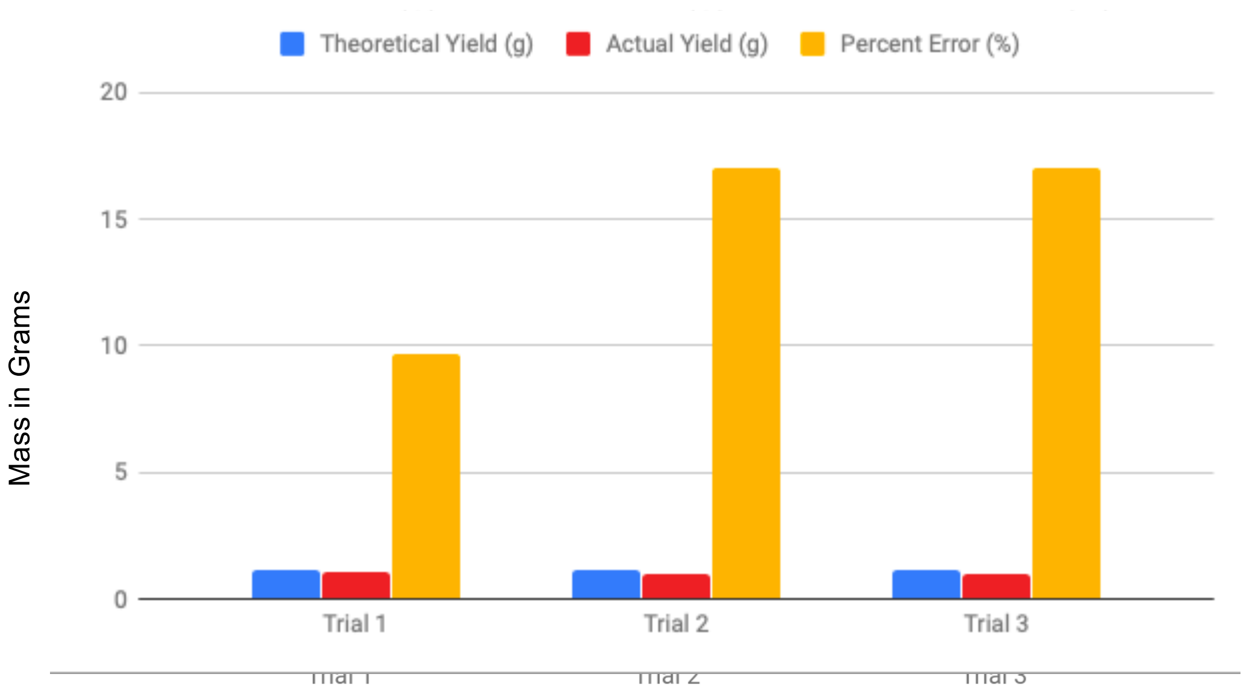
Discussion:

We began by dissolving approximately 6.00g copper (II) sulfate in water. When we added 1.00g of iron to the solution, a layer of rust formed on top. This was caused by the oxidation of the iron. A layer of copper accumulated at the bottom of the solution. Copper was formed because the reaction ($CuSO4 + Fe \rightarrow Cu + FeSO4$) is a single replacement reaction. Originally, the copper was attached to the sulfate molecule. The copper was then replaced by the iron to form iron (II) sulfate. The copper, now no longer attached to the sulfate molecule, sank to the bottom of the iron (II) sulfate solution. The copper weighed 1.03 g in trial one and 0.95 g in trials 2 and 3 (See Graph 1).

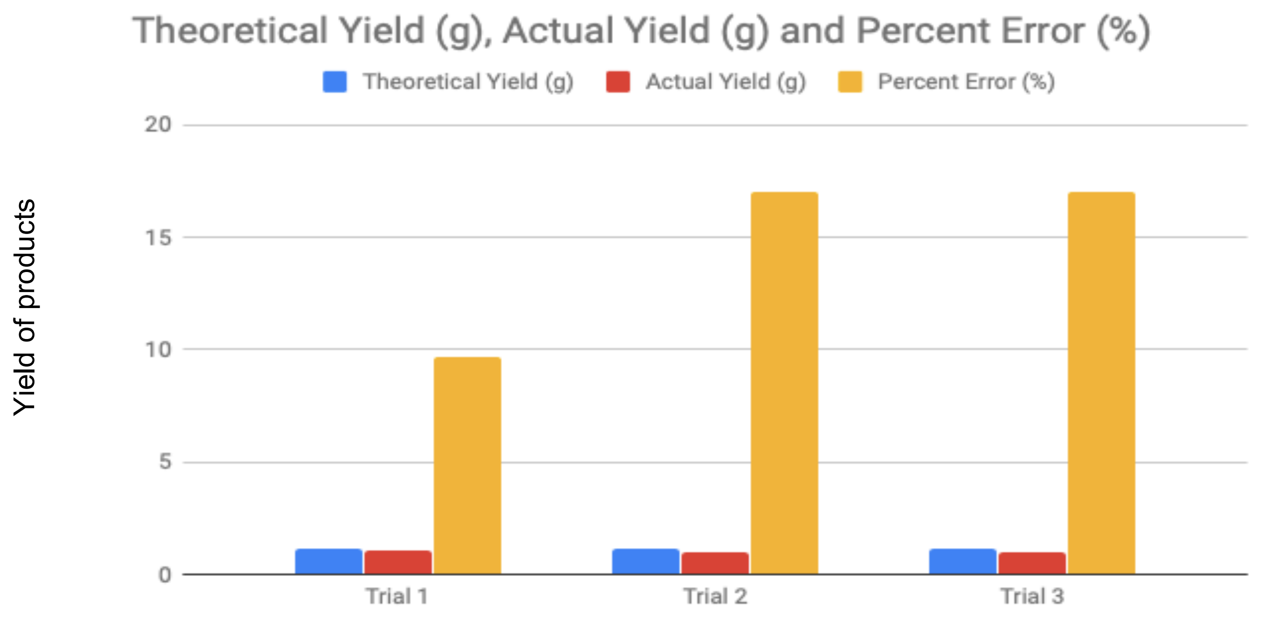
Given the equation of the reaction ($CuSO4 + Fe \rightarrow Cu + FeSO4$) and the amounts of copper (II) sulfate and iron which were used, the theoretical yield of the reaction was 1.14g of copper. Over the course of three trials, we had an average yield of 0.98g of copper with an average percent error of 14.55%. The average actual yield was 0.16g less than the theoretical yield due to human error during the experiment (See Graph 2).

In each trial, the actual yield of copper was less than the theoretical yield. In addition, the actual yield of the second and third trials was lower than the first trial., representing greater human error. This also meant that the percent error was greater in the second and third trials than the first (See Graph 2).

Mass of Reactants and Products



Graph 1: Mass of Reactants and Products



Graph 2: Theoretical Yield, Actual Yield, and Percent Error

Conclusion:

Based on the data acquired during the experiment, I learned that the actual yield is frequently less than the theoretical yield.

The information we gathered is applicable for factories that mass produce chemicals such as bleach or hydrogen peroxide. It would be helpful to know the average actual yield in addition to the theoretical yield so that the factory would be able to estimate the amount of product produced from a certain amount of reactants. In addition, the factory could analyze the chemical equation to determine the amount of reactants that is needed to produce the product and avoid using excess reactants.

I predict that in a similarly structured future experiment, our results would be the same. I predict that, once again, the actual yield would be less than the theoretical yield because of human error.

My hypothesis was correct. The actual yield was less than the theoretical yield which can be attributed to a number of potential causes stemming from human error. When we decanted the liquid, leaving behind the copper, we might have poured out some of the copper, reducing the yield. Some of the iron also might have stuck to the side of the beaker while we mixed it with the CuSO4 solution which would have reduced the amount of copper yielded.

Works Cited

CK-12 Foundation. “12 Foundation.” 11.3 Limiting Reactant and Percent Yield, www.ck12.org /user:dramartin/book/CLPS-Chemsitry-2017-2018/section/11.3/.

Sterner, Robert W, et al. “The Conservation of Mass.” *Knowledge Project*, 2011, www.nature.com /scitable/knowledge/library/the-conservation-of-mass-17395478.