

CHEM-FAX!

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Foiled Again — Aluminum Loses to Copper

A Redox Chemical Demonstration Kit

Introduction

Watch aluminum foil disappear as it is added to a green-blue solution of copper(II) chloride. Observe color changes, production of a gas, formation of solid metallic copper, and a drastic change in temperature. Learn the crucial role of a catalyst as the single-replacement, oxidation-reduction reaction proceeds.

Chemical Concepts

- Oxidation-reduction
- Catalysis
- Single replacement reaction

Materials (for each demonstration)

Aluminum foil, 6" × 12", 2 pieces*	Graduated cylinder, 250-mL
Copper(II) chloride solution, CuCl_2 , 1.0 M, 140 mL*	Spatula
Copper(II) sulfate solution, CuSO_4 , 1.0 M, 140 mL*	Stirring rod
Sodium chloride solution, NaCl , 1.0 M, 140 mL*	Thermometer (optional)
Water, distilled or deionized, 420 mL	Wood splint and matches (optional)
Beakers, Pyrex®, 600-mL, 3 or graduated cylinders, Pyrex®, 500-mL, 3	

*Materials included in kit.

Safety Precautions

Copper(II) chloride solution is toxic by ingestion and copper(II) sulfate solution is slightly toxic by ingestion. Small volumes of hydrogen gas are produced from the reaction. Hydrogen is a highly flammable gas; keep flammable materials away from the reaction mixture. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review Material Safety Data Sheet for further safety information. This activity requires the use of hazardous components and/or has the potential for hazardous reactions. Please review the Safety Precautions section and relevant Material Safety Data Sheets before beginning this activity.

Procedure

Part 1. Aluminum and Copper(II) Chloride

1. Place a 600-mL Pyrex® beaker (or a 500-mL graduated cylinder) on the demonstration table.
2. Use a graduated cylinder to measure 140 mL of 1.0 M CuCl_2 solution. Pour this into the beaker.
3. Measure and add 140 mL of distilled or deionized water to the beaker. The solution is now 0.5 M CuCl_2 .
4. Cut a piece of aluminum foil approximately 6" × 12". Loosely crumple the foil enough to fit into the beaker. (Note: Do not wad up the foil tightly in a ball as this decreases the surface area and slows any reaction that may occur.)
5. If desired, measure the temperature of the solution before adding the foil.
6. Place the piece of crumpled foil into the beaker, using a stirring rod to push it down completely into the solution. Measure the temperature of the reaction mixture again. Notice the great increase in temperature, indicating a very exothermic reaction.

7. Have students make detailed observations of the reaction occurring in the beaker. Have students generate hypotheses as to the reaction(s) that are occurring and have them write an equation for the reaction(s) they observe. Discuss which substances are reacting species and which are spectators in the reaction, if any. Students may hypothesize that aluminum reacts with copper(II) ions to form solid copper and aluminum ions. Test this hypothesis by performing Part 2, in which aluminum is again mixed with copper(II) ions, but this time from a different source, a copper(II) sulfate solution. *Note:* Set the beaker from Part 1 aside for comparison.

Part 2. Aluminum and Copper(II) Sulfate

8. Repeat the procedure (steps 1–6) from Part 1 in a different 600-mL beaker, except this time using 70 mL of 1.0 M CuSO_4 solution and 70 mL of distilled or deionized water. The solution is now 0.5 M CuSO_4 .

9. Have students make detailed observations. Did a reaction occur? Students will observe that no reaction occurs between aluminum and copper(II) sulfate. Why not? Have students make modifications to the original hypothesis and generate a new hypothesis. Students at this point may propose trying various experiments to test their hypotheses, so additional materials may be needed. Students may now hypothesize that in Part 1, the aluminum is reacting with the chloride ions rather than the copper(II) ions (since the copper(II) ions were present in both Parts 1 and 2 and no reaction occurred in Part 2). Test this hypothesis by performing Part 3, in which aluminum is mixed with chloride ions from sodium chloride. *Note:* Set the beaker from Part 2 aside for use in Part 4.

Part 3. Aluminum and Sodium Chloride

10. Repeat the procedure (steps 1–6) from Part 1 in a different 600-mL beaker, except this time using 70 mL of 1.0 M NaCl solution and 70 mL of distilled or deionized water. The solution is now 0.5 M NaCl .

11. Have students make detailed observations. Did a reaction occur? Students will observe that no reaction occurs between aluminum and sodium chloride. Why not? Have students again make modifications to their hypotheses and generate new hypotheses. Students may now hypothesize that both copper(II) ions and chloride ions are necessary for the reaction to occur, since both were present in Part 1. Test this hypothesis by performing Part 4, in which aluminum is mixed with both copper(II) ions and chloride ions. *Note:* Set the beaker from Part 3 aside for use in Part 4.

Part 4. Aluminum and Copper(II) Sulfate and Sodium Chloride

12. Place the beaker from Part 2 on the demonstration table.

13. Add 70 mL of 1.0 M NaCl solution and 70 mL of distilled or deionized water to the beaker.

14. Have students make detailed observations. Notice that a reaction now occurs between the aluminum and the copper(II) ions, as in Part 1. Were the hypotheses correct? Discuss what is occurring in the beaker and write the chemical equation for the reaction. If chloride ions are not in the equation, what is the purpose of the chloride ions? Discuss the role of a catalyst in a reaction. Predict what would happen if CuSO_4 were added to the beaker from Part 3. Try this experiment.

15. Place the beaker from Part 3 on the demonstration table. This beaker contains the unreacted aluminum foil in the 0.5 M NaCl solution.

16. Add 70 mL of 1.0 M CuSO_4 solution and 70 mL of distilled or deionized water to the beaker.

17. Have students make observations. Generate hypotheses about what other materials might act as catalysts in this reaction.

Extensions

1. *Test for hydrogen gas with a burning splint* — Light a wood splint and hold it in the beaker over the bubbles that are released from the reaction. A positive test is indicated if a pop or a barking sound is heard.

Test for oxygen gas with a glowing splint — Light a wood splint and blow it out so that it is glowing. Hold the glowing splint over the bubbles that are released from the reaction. A positive test is indicated if the glowing splint re-ignites.

Notice that hydrogen gas is released from the reaction. Discuss the origin of the gas. Write the chemical equation for the production of hydrogen gas.

2. Consider extending this demonstration to an open-ended student laboratory experiment. Have students form research groups to brainstorm hypotheses and discuss variables to be tested. Have them discuss possible experiments that might test each hypothesis, being sure to discuss necessary “controls.” Have students pick the hypothesis that seems most viable and interesting and write up the experimental design in detail. Have each group conduct their experiment and write a complete lab report including hypothesis, experimental design, results of experimentation, interpretation of the results, and conclusions. Share the results with the class.

Disposal

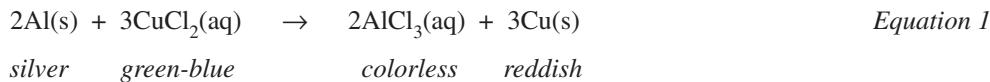
Allow the solid material in the beakers to settle. Decant the copper(II) chloride and copper(II) sulfate solutions down the drain according to Flinn Suggested Disposal Method #26b. Dispose of the solid copper and leftover aluminum foil in the solid waste according to Flinn Suggested Disposal Method #26a. Please consult your *Flinn Chemical & Biological Catalog/Reference Manual* for proper disposal procedures.

Tips

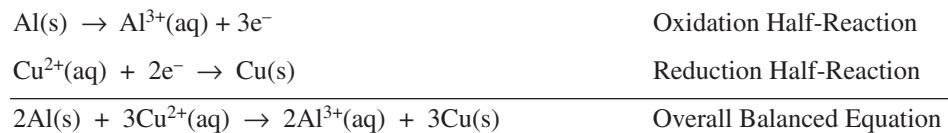
- This kit contains enough chemicals to perform the demonstration seven times using the amounts indicated in the procedure. You may opt to scale the reaction up or down proportionally, depending on class size and if you wish to try any extensions.
- The reaction temperature increases from room temperature (25 °C) to nearly 60 °C when aluminum is added to copper(II) chloride. Therefore, be sure to perform the demonstration in Pyrex® glassware.
- Sodium chloride is provided in this kit to catalyze the reaction between aluminum and copper(II) sulfate. Observations of the reaction with the catalyst are the same as those observed between aluminum and copper(II) chloride. Other catalysts that have been found to catalyze this reaction: sodium bromide, potassium chloride, magnesium chloride, and hydrochloric acid solution (1 M). It has also been found that copper(II) nitrate reacts in the same way as copper(II) sulfate. As an extension to this demonstration, consider trying other catalysts or other copper solutions.
- A more dramatic demonstration display can be achieved by performing the demonstration in large 500-mL Pyrex® graduated cylinders or hydrometer cylinders. In this case, loosely coil up the foil into a tube-like piece and drop it into the cylinder of CuCl₂. The foil will begin to slowly rise up in the cylinder from the gas bubbles attaching to the foil piece. Most of the reaction will occur at the top of the solution (which will turn gray or colorless) and there will be unreacted green-blue copper solution at the bottom of the cylinder. A long stirring rod will be useful in pushing the foil down to the bottom.
- It has been found that simply heating the solution of copper(II) sulfate and aluminum up to over 80 °C does not cause any reaction to occur. The chloride catalyst is necessary for the reaction to occur.
- Try using tap water in place of distilled water in the dilution of the CuSO₄ solution. Are there enough dissolved chloride ions in tap water to catalyze the reaction?
- After performing this demonstration to the class, have the class perform the *Leftover Aluminum Wire Stoichiometry Lab*, available from Flinn as a chemistry laboratory kit, Catalog No. AP4678. This lab uses the same reaction to teach concepts of moles, limiting reactants, and stoichiometry.

Discussion

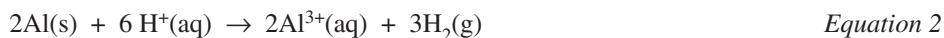
In the demonstration, aluminum foil is added to an aqueous solution of copper(II) chloride causing a single replacement oxidation-reduction reaction to take place (see *Equation 1*).



The oxidation of aluminum metal to aluminum(III) (Al^0 to Al^{3+}), which is visible by the dissolving of the aluminum foil to form aluminum chloride, is represented by the oxidation half-reaction below. The simultaneous reduction of copper(II) ions to copper metal (Cu^{2+} to Cu^0) will occur and solid copper metal precipitates from solution according to the reduction half-reaction below. As the copper(II) ions are reduced to copper, the green-blue solution color will fade until the solution is completely colorless—the indication that the reaction is complete and all of the copper(II) ions have been reduced.



It can be observed that hydrogen gas is simultaneously released from the reaction when aluminum metal foil is added to copper(II) chloride solution. If the pH of the copper(II) chloride solution is measured, it is found to be slightly acidic. Hence there are free hydrogen ions in solution, which cause the side reaction of hydrogen ions with the aluminum surface to form hydrogen gas and aluminum ions (See *Equation 2*). Due to the limited concentration of hydrogen ions, this reaction only consumes a small amount of the aluminum.



Why does aluminum react with copper(II) chloride and not with copper(II) sulfate or copper(II) nitrate? And why does it then react with copper(II) sulfate when chloride ions are added?

This interesting phenomenon was discovered serendipitously by Cliff Schrader. He had wanted to perform the aluminum plus copper(II) chloride reaction with his class, but had run out of copper(II) chloride. He thought that the net ionic equation would be the same with copper(II) sulfate, and he had lots of that on hand. Yet upon mixing aluminum with copper(II) sulfate, to his surprise, nothing happened. He tried adding a few grams of salt and the reaction went! This was quite unpredicted and opened up a whole new set of questions. While the reasons for Cliff's observations are still unclear, there are several possible explanations that can account for this interesting occurrence.

The first possible explanation is that the chloride ions may catalyze the reaction. Cliff tried other halide catalysts and found that bromide ions also seemed to catalyze the reaction. The second possible explanation has to do with the aluminum oxide coating. Normally, the chemically inert coating of aluminum oxide on the surface of the aluminum metal protects the metal from further oxidation (by copper(II) ions). In the presence of chloride ions, however, the coating is breached and aluminum reduces copper(II) ions to copper metal. Because the chloride ion is a relatively small ion, it is able to diffuse into and through a protective metal oxide coating. Aluminum chloride, which is more soluble than aluminum oxide, can then form. This aluminum chloride salt leaches back through the oxide coating, and a path is now open for copper(II) ions to attack the underlying aluminum metal. The reaction is rapid and extremely exothermic. A third possible explanation for the observations in this demonstration involves electrochemical potentials. The copper–chloride electrode (with copper(II) chloride) has a greater half-cell potential than the copper–copper (II) electrode (with copper(II) sulfate). This supports the observation that the reaction with chloride ions is more favorable than without chloride ions. This experiment is an excellent open-ended demonstration which, like any good inquiry-based lesson, leads to more questions than answers.

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12

Evidence, models, and explanation

Constancy, change, and measurement

Content Standards: Grades 5–8

Content Standard B: Physical Science, properties and changes of properties in matter

Content Standards: Grades 9–12

Content Standard B: Physical Science, structure and properties of matter, chemical reactions

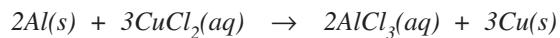
Answers to Worksheet Questions

Results Table

Part	Chemicals Used	Observations
Part 1	Aluminum foil, 0.5 M cupric chloride	<i>The solution around the aluminum fizzes and lets off a vapor. Copper coats the aluminum foil, and eventually the foil dissolves. The temperature of the solution rises, and the blue color fades.</i>
Part 2	Aluminum foil, 0.5 M cupric sulfate	<i>No evidence of a reaction.</i>
Part 3	Aluminum foil, 0.5 M sodium chloride	<i>No evidence of a reaction.</i>
Part 4	Aluminum foil, 0.5 M cupric sulfate, 0.5 M sodium chloride	<i>The solution around the aluminum fizzes and lets off a vapor. Copper coats the aluminum foil, and eventually the foil dissolves. The temperature of the solution rises, and the blue color fades.</i>

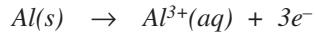
Questions

1. Write a balanced chemical equation for the reaction between aluminum foil and cupric chloride.

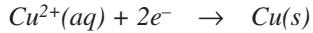


2. Aluminum metal is oxidized into aluminum(III), while the copper(II) is reduced. Write three balanced chemical equations:

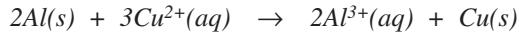
- a. For the oxidation half-reaction



- b. For the reduction half-reaction



- c. For the overall balanced equation



3. Propose an explanation for why there was no reaction with the aluminum by the cupric sulfate alone or by the sodium chloride alone, but there was when the cupric sulfate and the sodium chloride were mixed together.

The reaction only worked when cupric sulfate and sodium chloride were mixed together because even though the reaction is between aluminum and copper, it happens very slowly. Chloride ions then act as a catalyst, speeding up the reaction a great deal.

Acknowledgments

Special thanks to the late Cliff Schrader, Summit County ESC, Cuyahoga Falls, OH for providing Flinn with the idea for this activity. Andy Cherkas, Walter Rohr, and Pat Funk also provided insight and advice on this activity.

The Foiled Again—Aluminum Loses to Copper—Chemical Demonstration Kit is available from Flinn Scientific, Inc.

Catalog No.	Description
AP5936	Foiled Again—Aluminum Loses to Copper—Chemical Demonstration Kit
AP4678	Leftover Aluminum Wire—A Stoichiometry Lab Kit

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

Foiled Again — Aluminum Loses to Copper

Demonstration Worksheet

Results Table

Part	Chemicals Used	Observations
Part 1		
Part 2		
Part 3		
Part 4		

Discussion Questions

1. Write a balanced chemical equation for the reaction between aluminum foil and cupric chloride.
2. Aluminum metal is oxidized into aluminum(III), while the copper(II) is reduced. Write three balanced chemical equations:
 - a. For the oxidation half-reaction
 - b. For the reduction half-reaction
 - c. For the overall balanced equation
3. Propose an explanation for why there was no reaction with the aluminum by the cupric sulfate alone or by the sodium chloride alone, but there was when the cupric sulfate and the sodium chloride were mixed together.