

Tested Demonstrations

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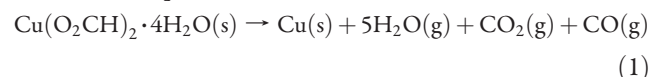
Metals in Metal Salts: A Copper Mirror Demonstration

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The presence of metal within an ionic metal salt is usually taken as an assumption in high school and college chemical education. Several laboratory experiments that demonstrate or even quantify the content of metal within metal compounds have been published in this *Journal* (1–6). In addition, there are a variety of high school- and college-level experiments involving the redox chemistry of copper in pennies (7, 8). However, none of these experiments is readily amenable to a simple lecture demonstration that can accompany initial discussions of ionic compounds. Consideration of ionic substances is important in both high school- and college-level introductory chemistry courses. Cations in all common ionic compounds, except ammonium salts, are metal ions. Demonstration of the latent presence of metal within a salt via reduction of the metal cation affords a visual touchstone for this important concept.

A convenient method for the reduction of metal salts would involve an anion that easily undergoes reductive elimination, involve a readily reducible metal ion, and produce a readily identifiable metallic residue. All of these characteristics are fulfilled by the use of copper(II) formate tetrahydrate. The thermal decomposition of this salt is



The formation of a copper mirror from copper(II) formate as described in eq 1 was first reported in 1947 (9); copper mirrors generated under inert atmosphere have been used as high-surface-area reducing agents for organic reactions (10). A related demonstration has been reported in a relatively obscure German-language journal (11).

Materials

Copper(II) formate tetrahydrate
250 mL Borosilicate glass round-bottom flask
Three-prong clamp (without vinyl or rubber sleeves)
Meeker burner and gas supply
Spatula
Wire gauze
Cork ring
Safety goggles or face shield
Disposable gloves
Nitric acid (for cleanup)
A fire extinguisher should be readily available during any demonstration involving flames

Preparation and Presentation

Approximately 3 g of $\text{Cu}(\text{O}_2\text{CH})_2 \cdot 4\text{H}_2\text{O}$ are placed in a very clean 250 mL borosilicate round-bottom flask that is

held with a metal clamp. The flask is heated with swirling over a Meeker burner or propane torch flame for about 2 min. It should be cooled appropriately (e.g., on a wire gauze) before handling. Once the flask is cool, a cork ring is useful for displaying the flask.

The bright copper mirror can be removed with nitric acid. The nitric acid waste should be neutralized (e.g., with NaHCO_3) before disposal, which should be carried out in accordance with local regulations for metal-containing wastes.

Hazards

Safety goggles or a face shield, latex gloves and a lab coat should be worn. Heated glass surfaces can cause burns and should be cooled for at least 10 min before handling. Glassware should be inspected carefully prior to heating it to avoid possible stress fracturing. The liberation of CO is potentially dangerous but the CO is spontaneously combusted during the experiment. Although a fume hood is not necessary, care should be taken not to breathe the gases liberated during the experiment. A small amount of metal oxide fume is evident during the reaction; however, these fumes are relatively heavy and do not pose a significant inhalation hazard, so long as the flask is kept well away from the demonstrator's face. A fire extinguisher should be readily available for this demonstration. Goggles, gloves, and a fume hood should be used when handling nitric acid during cleanup, because nitric acid is corrosive and its reduction by copper produces hazardous NO_2 gas.

Discussion

In the lecture demonstration, copper(II) formate is decomposed in a round-bottom flask using the heat from a Meeker burner. During the reaction, which takes about 2 min with continuous swirling, the blue–green copper compound first becomes a royal blue color, then liquefies and effervesces with the release of the hydration water. Suddenly, the residue becomes dark as the metal is liberated. At about the same time, the gas evolved will usually support a small green flame above the neck of the flask. The flame is presumably due to the combustion of carbon monoxide, but the green color is associated with copper(II) ions (begging a discussion of fireworks colorants). The flame can be blown out and will spontaneously reignite, amusingly. Soon thereafter, a high-quality copper mirror develops on the inside surface of the flask. Given the unique color of copper metal, the highly reflective mirror is readily identifiable as copper. The mirror that is produced will behave as would any fresh preparation of copper. For example, it can be converted to a blue solution of $\text{Cu}(\text{NO}_3)_2$ by treatment with HNO_3 , and then to $\text{Cu}(\text{OH})_2$, CuO , CuSO_4 solution, and back to Cu in the well-known copper cycle laboratory sequence (12). Depending

on the makeup of the audience, it might be appropriate to discuss the fact that copper, along with its congeners silver and gold, is a relatively noble metal. As such, copper(II) is reduced to the metal with relative ease



Copper is sometimes mined in pure ("native") form. The world's largest deposits of native copper were mined from upper Michigan in the 19th and early 20th centuries.

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Literature Cited

1. Sheeran, D. J. *Chem. Educ.* **1998**, *75*, 453–456.

2. O'Klatner, B. L.; Rabinovich, D. J. *Chem. Educ.* **2000**, *77*, 251–252.
3. Osella, D.; Ravera, M.; Soave, C.; Scorza, S. J. *Chem. Educ.* **2002**, *79*, 343–344.
4. Sanger, M. J.; Geer, K. J. *Chem. Educ.* **2002**, *79*, 994–996.
5. Sago, S. G. J. *Chem. Educ.* **2004**, *81*, 530–531.
6. Yee, G. T.; Eddleton, J. E.; Johnson, C. E. J. *Chem. Educ.* **2004**, *81*, 1777–1779.
7. Rosenhein, L. D. J. *Chem. Educ.* **2001**, *78*, 514–515.
8. Thomas, N. C.; Faulk, S. J. *Chem. Educ.* **2008**, *85*, 817–818.
9. Korosy, F. *Nature* **1947**, *160*, 21.
10. Pike, R. D.; Starnes, W. H., Jr.; Jeng, J. P.; Bryant, W. S.; Kourtesis, P.; Adams, C. W.; Bunge, S. D.; Kang, Y. M.; Kim, A. S.; Kim, J. H.; Macko, J. A.; O'Brien, C. P. *Macromolecules* **1997**, *30*, 6957–6965.
11. Weissenhorn, R. G. *Prax. Naturwiss., Chem.* **1990**, *39*, 38.
12. Postma, J. M.; Roberts, J. L.; Hollenberg, J. L. *Chemistry in the Laboratory*, 6th ed.; W. H. Freeman: New York, 2004; Expt. 5.

Supporting Information Available

A video of the demonstration. This material is available via the Internet at <http://pubs.acs.org>.