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A Colorful Catalysis Demonstration

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Abstract: A simple, yet colorful, homogeneous catalysis demonstration is reported. When concentrated H_2O_2 is added to a yellow-orange aqueous solution of Fe^{3+}, the solution turns black in color. After a brief induction period, exothermic decomposition of H_2O_2 ensues with the evolution of oxygen gas. As the reaction subsides, the solution returns to the original yellow-orange color. These color changes demonstrate one of the most important (and often most misunderstood by first-year students) concepts of catalysis—although a catalyst may be altered as it participates in a reaction, it is not consumed.

Introduction

Catalysis is one of the central themes of modern chemistry [1]. The concept of a catalyst to provide a new reaction pathway, normally with a lower activation barrier and enhanced rate, is an important topic in introductory chemistry courses for majors and nonmajors alike [2]. Often, demonstrations are used to portray this theme. For example, the catalytic oxidation of ammonia by Cr_2O_3 results in spectacular visual effects [3], and the reaction of hydrogen and oxygen gases with a 5% Pd/C catalyst provides students with audible evidence that a reaction has occurred [4].

One of the most commonly employed molecules for the investigation of catalysis is hydrogen peroxide. The thermodynamic instability of hydrogen peroxide with respect to disproportionation to oxygen and water (eq 1), coupled with its sluggish rate in the absence of a catalyst, make it an ideal system for investigating catalysis.

$$H_2O_2\,(aq) \rightleftharpoons H_2O\,(1) + \tfrac{1}{2}\,O_2\,(g) \qquad \Delta H^0 = -94.6\ \text{kJ mol}^{-1} \quad (1)$$

In a homogeneous system, a cobalt catalyst undergoes color changes during the course of the oxidation of tartrate by H_2O_2 [5]; however, this demonstration is difficult to perform and requires preheating of solutions to 70 °C. Photographs showing the color changes associated with the bromide-ion-catalyzed decomposition of H_2O_2 appear in a recent text [6]. Decomposition of H_2O_2 by heterogeneous catalysts such as $KMnO_4$ provides a simple means of generating oxygen [7], and a microscale investigation of the kinetics of decomposition using a pyrolusite rock has been reported [8]. Prompted by a clever calorimetric experiment using aqueous Fe^{3+} to catalyze the decomposition of H_2O_2 [9], we have developed a demonstration that uses color changes to show that, although a catalyst may change as it participates in a chemical reaction, it is ultimately regenerated and may be used again.

Procedure

Caution: Concentrated hydrogen peroxide solutions are strong oxidants. Gloves should be worn whenever these solutions are handled. It is also recommended that the demonstrator and audience remain a minimum of 5 ft away from the reaction vessel. This is because the possible ejection of “suds” during the boiling phase.

If an overhead projector is to be used: A 400-mL beaker containing 10 mL of 0.50 M aqueous $Fe(NO_3)_3$ is placed on a projector. The addition of 5 mL of 10% H_2O_2 (diluted from 30%) then causes the solution to turn black. After a brief induction period, the solution bubbles as oxygen gas is produced and eventually returns to the original yellow-orange color when the H_2O_2 is consumed. A beaker of catalyst solution diluted with 5 mL of H_2O may serve as a control. The use of more concentrated H_2O_2 makes the reaction harder to visualize as condensed water vapor produced by the reaction blocks the optics of the projector. A second aliquot of 10% H_2O_2 may be added to demonstrate that the catalyst remains active.

Using a demonstration-sized test tube: A 45 by 400-mm test tube is filled with 5 mL of aqueous 0.25 M $Fe(NO_3)_3$ and 10 mL of 30% H_2O_2. Again, the solution immediately turns black, and after a short induction period it begins to boil, eventually returning to the original color. The results from this demonstration are shown in Figure 1.

Using a super-sized test tube: A 75 by 800 mm test tube (fabricated by a professional glass blower) is filled with 20 mL of aqueous 25 M $Fe(NO_3)_3$ and 50 mL of 30% H_2O_2. The results of this demonstration are similar to those of the demonstration-sized test tube, but the larger scale makes this method ideal for bigger lecture halls.

Addition of a second (smaller) aliquot of H_2O_2 to the test tubes results in a more rapid reaction because the temperature of the test tubes is now above room temperature, and this demonstrates that the catalyst is still active.

If 30% hydrogen peroxide is unavailable, it is possible to use over-the-counter 3% H_2O_2 from a drugstore. In this case, the smaller demonstration-sized test tube, 20 mL of 0.50 M $Fe(NO_3)_3$, and 50 mL of 3% H_2O_2 are recommended. Although the color changes are still clearly evident, this leads to a slower reaction that has a markedly less intense boiling phase.

Results and Discussion

Along with the low cost and toxicity and familiarity with the reagents, this demonstration is ideal for demonstrating

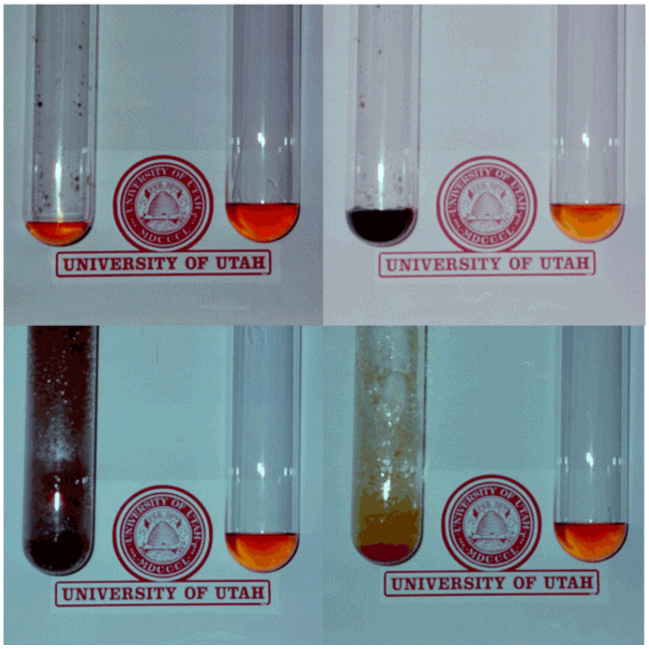

Figure 1. The reaction of H_2O_2 and aqueous Fe^{3+} in a demonstration-sized test tube. The sequence of photographs is arranged from top left to bottom right with the test tube on the right containing a catalyst control solution.

that a catalyst is not consumed during the course of a reaction. The black color of the solution illustrates that the iron catalyst is involved with the reaction, but the return to the original color confirms the idea that the catalyst is not consumed.

The yellow-orange color of the catalyst solution is due to the presence of $Fe(OH)(H_2O)_5^{2+}$ [9]. The black color generated upon addition of H_2O_2 to the aqueous Fe^{3+} catalyst is ascribed to an Fe–H_2O_2 complex that is believed to form during the reaction according to the following mechanism [10]:

$$\mathrm{Fe^{III}} + \mathrm{H_2O_2} \rightleftharpoons \mathrm{Fe^{III}(H_2O_2)}$$

$$\mathrm{Fe^{III}(H_2O_2)} \rightarrow \mathrm{Fe^{IV}OH} + \bullet\mathrm{OH}$$

$$\bullet\mathrm{OH} + \mathrm{H_2O_2} \rightarrow \mathrm{H_2O} + \mathrm{HO_2}\bullet$$

$$\mathrm{Fe^{IV}OH} + \mathrm{HO_2}\bullet \rightarrow \mathrm{Fe^{III}} + \mathrm{H_2O} + \mathrm{O_2}$$

As the H_2O_2 is broken down into O_2 and H_2O, the catalyst complex is destroyed, returning the solution to its original color as the reaction subsides.

Freshly prepared aqueous Fe^{3+} solutions give slightly shorter induction periods than aged solutions, making it easier to observe the black Fe–H_2O_2 complex. This difference is attributed to the formation of polynuclear species that are established in the complex equilibria of aqueous Fe^{3+} solutions [11]. This demonstration is also relevant to “Fenton” oxidations of organic substrates—the mechanistic details of these reactions remain a subject of much discussion [12].

References and Notes

1. Cusumano, J. A. *J. Chem. Educ.* **1995,** *72,* 959–964; Oyama, S. T.; Somorjai, G. A. *J. Chem. Educ.* **1988,** *66,* 765–769; Kolb, D. *J. Chem. Educ.* **1979,** *56,* 743–747.
2. Copper, C. L.; Koubek, E. *J. Chem. Educ.* **1999,** *76,* 1714–1715.
3. Volkovich, V. A.; Griffiths, T. R. *J. Chem. Educ.* **2000,** *77,* 177.
4. Owens, G.; Richmond, T. G. Demonstrating Catalysis with a Bang! Initiation of a Hydrogen/Oxygen Explosion with a Pd/C Catalyst. *Chem. Educator 1*(4), S 1430-4171(96)04045-9; DOI 10.1007/s00897960045a.
5. Summerlin, L. R.; Ealy, J. L., Jr. *Chemical Demonstrations: A Sourcebook for Teachers;* R. R. Donnelley and Sons: Chicago, IL, 1985, p 72.
6. Brown, T. L.; LeMay, H. E,. Jr.; Bursten, B. E. *Chemistry: The Central Science,* 8th ed.; Prentice Hall: Upper Saddle River, NJ, 2000, p 540.
7. Shakhashiri, B .Z. *Chemical Demonstrations: A Handbook for Teachers of Chemistry,* vol 2; The University of Wisconsin Press: Madison, WI, 1985, pp 137–141.
8. Ragsdale, R. O.; Vanderhooft, J. C. *J. Chem. Educ.* **1998,** *75,* 215–216.
9. Marzzacco, C. J. *J. Chem. Educ.* **1999,** *76,* 1517–1518.
10. Gopinath, E.; Bruice, T. C. *J. Am. Chem. Soc.* **1991,** *113,* 4657–4665.
11. Greenwood, N. N.; Earnshaw, A. *Chemistry of the Elements,* 2nd ed.; Butterworth-Heinemann: Oxford, England, 1997, p 1089.
12. Walling, C. *Acc. Chem. Res.* **1998,** *31,* 155–157; MacFaul, P. A.; Wayner, D. D. M.; Ingold, K. U. *Acc. Chem. Res.* **1998,** *31,* 159-162.