Fantastic Four-Color Oscillator

FLINN SCIENTIFIC
CHEM FAX!

Introduction

You won't believe your eyes as you watch this amazing oscillating reaction! This four-color oscillator will get your students' undivided attention as they observe a solution flash from green to blue to purple to red. And that's not all—this four-color cycle will repeat itself for well over an hour!

Concepts

- Oscillating reactions
- Chemical equilibrium
- Oxidation–reduction reactions
- Kinetics/catalysts
- Reaction mechanisms

Materials

1,10-phenanthroline, $C_{12}H_8N_2 \cdot H_2O$, 0.23 g

Cerium(IV) ammonium nitrate, $Ce(NH_4)_2(NO_3)_6$, 2.7 g

Iron(II) sulfate, $FeSO_4 \cdot 7H_2O$, 0.12 g

Malonic acid, $CH_2(CO_2H)_2$, 8 g

Potassium bromate, $KBrO_3$, 9.5 g

Potassium bromide, KBr, 1.8 g

Sulfuric acid solution, H_2SO_4, 3.0 M, 225 mL

Distilled or deionized water, approximately 1 L

Beaker, 1-L

Erlenmeyer flasks, 250-mL, 3

Graduated cylinder, 50-mL

Magnetic stirring bar

Magnetic stirring plate

Spatula

Safety Precautions

A small amount of elemental bromine gas is released from the reactions in this demonstration; adequate ventilation is necessary. Potassium bromate is a strong oxidizing agent and poses a fire risk in contact with organic material; it is a strong irritant and moderately toxic. Malonic acid is a strong irritant, slightly toxic, and corrosive to eyes, skin, and respiratory tract. Potassium bromide is slightly toxic by ingestion and a severe body tissue irritant. Cerium(IV) ammonium nitrate is a strong oxidizer and a skin irritant. Iron(II) sulfate is slightly toxic by ingestion and 1,10-phenanthroline is highly toxic by ingestion. Sulfuric acid solution is corrosive to eyes, skin, mucous membrane, and other body tissue. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Preparation

Solution A (0.23 M $KBrO_3$) — In a 250-mL Erlenmeyer flask, dissolve 9.5 grams of potassium bromate in 250 mL of distilled or deionized water. Label this solution as "Solution A."

Solution B (0.31 M $CH_2(CO_2H)_2$ and 0.059 M KBr) — In a second 250-mL Erlenmeyer flask, dissolve 8 grams of malonic acid and 1.8 grams of potassium bromide in 250 mL of distilled or deionized water. Label this solution as "Solution B."

Solution C (0.019 M $Ce(NH_4)_2(NO_3)_6$ and 2.7 M H_2SO_4) — In a third 250-mL Erlenmeyer flask, dissolve 2.7 grams of cerium(IV) ammonium nitrate in 25 mL of distilled or deionized water and add 225 mL of 3.0 M sulfuric acid. Label this solution as "Solution C."

Solution D (0.50% ferroin solution) — To prepare 50 mL of stock solution, dissolve 0.12 grams of iron(II) sulfate in 50 mL of distilled or deionized water. In the resulting solution, dissolve 0.23 grams of 1,10-phenanthroline. Label this solution as "Solution D."

Procedure

1. Place a 1-L beaker on the magnetic stirring plate and place the magnetic stirring bar in the beaker.

2. Pour all of Solution A and Solution B into the 1-L beaker.

3. Adjust the stirrer to produce a vortex in the solution. The solution may become amber, and will turn colorless after about one minute.

4. Once the solution is colorless, add all of Solution C and only 15 mL of Solution D. (Note: The solution composition is now 0.077 M BrO_3^-, 0.10 M malonic acid, 0.020 M Br^-, 0.0063 M Ce^{4+}, 0.90 M H_2SO_4, and 0.17 mM ferroin.)

5. Keep stirring the green cloudy mixture and it will become a green solution. Over a period of about a minute, the color of the solution will change from green to blue, then to violet, and finally to red-brown.

6. The color of the solution will suddenly return to green, and the cycle will repeat itself more than 20 times, lasting over an hour.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. The reaction mixture should be neutralized with sodium carbonate and flushed down the drain with excess water according to Flinn Suggested Disposal Method #24a.

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
Content Standards: Grades 9–12
　　Content Standard B: Physical Science, structure and properties of matter, chemical reactions

Tips

- Solutions A–D may be prepared in advance, stoppered, and labeled until needed for the presentation. Solution shelf life is good. The demonstration can easily be scaled up or down to suit your needs.

- The discussion and explanation of this oscillating reaction can be quite complicated for beginning chemistry or general science students. However, it is not necessary to fully understand the reaction mechanism in order to appreciate the spectacular chemistry that occurs in this demonstration. Thus, this demonstration can be performed at any level of science with the explanation suited to the level of the class. A complete discussion is included in this handout; however, further information can be found by reviewing the original reference (see References section).

Discussion

This oscillating reaction demonstrates the classic Belousov-Zhabotinsky (BZ) reaction which is a cerium-catalyzed bromate–malonic acid reaction.

The overall reaction occurring in this demonstration is the cerium-catalyzed oxidation of malonic acid by bromate ions in dilute sulfuric acid. The bromate ions are reduced to bromide ions, while the malonic acid is oxidized to carbon dioxide and water. The overall reaction can be represented by Equation 1:

$$3CH_2(CO_2H)_2 + 4BrO_3^- \rightarrow 4Br^- + 9CO_2 + 6H_2O \hspace{2cm} \textit{Equation 1}$$
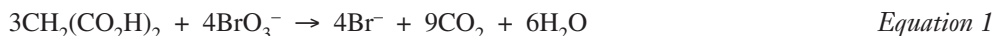
In order to gain some understanding and appreciation for how this overall reaction can produce the amazing, repetitive color changes observed in the demonstration, it is necessary to look at the reaction mechanism or, in other words, how the reactants are transformed into products.

The mechanism involves two different competing processes—Process A involves ions and two-electron transfers; Process B involves radicals and one-electron transfers. The dominant process at any particular time is dependent on the bromide ion concentration. Process A (see Equation 2a) occurs when the bromide ion concentration rises above a certain critical level, while Process B (see Equation 3a) is dominant when the bromide ion concentration falls below a certain critical level. Oscillations occur because Process A consumes bromide ions, leading to conditions which favor Process B. Process B (indirectly) produces bromide ions, which leads to conditions which favor Process A.

Process A

$$BrO_3^- + 5Br^- + 6H^+ \rightarrow 3Br_2 + 3H_2O \qquad \textit{Equation 2a}$$
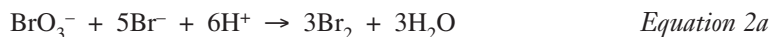
Bromate ions are reduced by bromide ions through a series of oxygen transfers (two-electron reductions) as shown in Equation 2a. This reaction occurs when Solutions A and B are mixed. The amber color which may develop is caused by the production of elemental bromine. This color soon disappears as the bromine reacts with malonic acid as shown in Equation 2b.

$$Br_2 + CH_2(CO_2H)_2 \rightarrow BrCH(CO_2H)_2 + Br^- + H^+ \qquad \textit{Equation 2b}$$
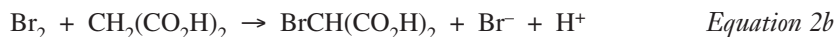
Process A results in an overall decline in the bromide ion concentration and, once the necessary intermediates are generated and most of the bromide ions are consumed, the rate becomes negligible and Process B takes over.

Process B

$$2BrO_3^- + 12H^+ + 10Ce^{3+} \rightarrow Br_2 + 10Ce^{4+} + 6H_2O \qquad \textit{Equation 3a}$$
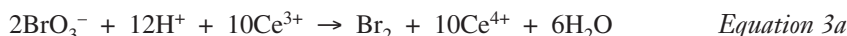
Bromate ions are reduced by cerium(III) ions to produce bromine through a simple redox reaction as shown in Equation 3a. Process B produces Ce(IV) ions and Br_2. Both of these species react at least in part to oxidize the malonic acid (see Equation 2b) and the bromomalonic acid (see Equation 3b) to form additional bromide ions. As the concentration of bromide ions increases, the rate of Equation 2a increases until eventually Process A once again dominates.

$$BrCH(CO_2H)_2 + 4Ce^{4+} + 2H_2O \rightarrow Br^- + 4Ce^{3+} + HCO_2H + 2CO_2 + 5H^+$$

$$\textit{Equation 3b}$$

As the reaction oscillates between Process A and Process B, triggered by changes in the bromide ion concentration, concentrations of other species in solution oscillate as well—these concentration changes will explain the color changes observed. While Process A occurs, the cerium ions are in their reduced state, Ce(III). During Process B, some cerium ions are oxidized to Ce(IV) and thus the ratio of Ce(III) to Ce(IV) oscillates as well.

The indicator used in this demonstration is ferroin, which is tris(1,10-phenanthroline) ferrous sulfate. As the concentration of Ce(IV) increases, the Ce(IV) oxidizes the iron in ferroin from Fe(II) to Fe(III). The Fe(II) complex is red while the Fe(III) complex is blue; thus the color of the solution changes as the iron is oxidized. As the concentration of Ce(III) increases, the Fe(III) is reduced back to Fe(II) and the color of the solution changes accordingly.

The color changes in this demonstration, however, are more complex than simple red-blue oscillations from the ferroin. There are also changes in color due to the cerium ions in solution—Ce(III) is colorless while Ce(IV) is yellow. A simplified equation to help explain the color changes is shown below, Equation 4:

$$Ce(IV) + Fe(II) \leftrightarrows Fe(III) + Ce(III) \qquad \textit{Equation 4}$$

yellow   red   blue   colorless
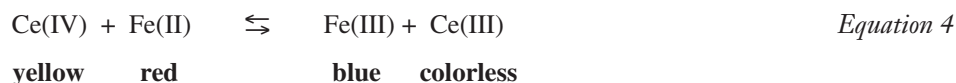
A possible explanation for the appearance of the oscillating colors in solution is provided; however, a more complete understanding of the color changes might be gained by reviewing the original references.

Green = The yellow Ce(IV) is oxidizing Fe(II) to blue Fe(III); a small amount of Fe(II) has been oxidized to the blue Fe(III) complex; thus, the mixture of yellow and blue forms a green solution.

Blue = All Ce(IV) is reduced to colorless Ce(III); all Fe(II) is oxidized to the blue Fe(III) complex; thus, the solution is blue.

Violet = The colorless Ce(III) is reducing the blue Fe(III) complex to the red Fe(II); the mixture of blue and red appears violet.

Red = All of the blue Fe(III) is reduced to the red Fe(II) complex; colorless Ce(III) is present; the solution appears red.

References

Shakhashiri, B. Z. Chemical Demonstrations: A Handbook for Teachers of Chemistry; University of Wisconsin Press: Madison; 1985; Vol. 2, pp 257–261.

Materials for the *Fantastic Four-Color Oscillator* are available from Flinn Scientific, Inc.

Catalog No.	Description
P0136	Potassium Bromate, 100 g
M0091	Malonic Acid, 25 g
P0137	Potassium Bromide, 100 g
C0287	Cerium(IV) Ammonium Nitrate, 25 g
S0143	Sulfuric Acid, 18 M, 473 mL
F0016	Iron(II) Sulfate, 500 g
P0155	1,10-phenanthroline, 5 g

The above demonstration was developed by Flinn Scientific, Inc. into a demonstration kit. Each kit contains enough chemicals to perform the experiment at least seven times. Excite, energize, and bring back the "magic" of chemistry using Flinn Chemical Demonstration Kits.

Catalog No.	Description
AP4833	Fantastic Four-Color Oscillator—Chemical Demonstration Kit

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

Stoichiometry Balloon Races
Limiting and Excess Reactants

Introduction

Most stoichiometry calculations in the classroom are performed using exact (stoichiometric) mole ratios of reactants and products. In real life, however, many commercial processes for preparing compounds are carried out using an excess amount of one reactant (and thus a limiting amount of another reactant). This demonstration uses the well-known reaction of sodium bicarbonate and acetic acid to illustrate the concepts of limiting and excess reactants. By comparing the amount of carbon dioxide generated when varying amounts of sodium bicarbonate react with a given amount of acetic acid, students will be able to identify immediately the limiting and excess reactant in each case.

Concepts

- Stoichiometry
- Mole ratio
- Limiting reactant
- Excess reactant

Materials

Acetic acid, CH_3COOH, 1 M, 150 mL

Sodium bicarbonate, $NaHCO_3$, 10.5 g

Balance, centigram (0.01-g) precision

Balloons, 6

Erlenmeyer flasks, 125-mL, 6

Powder

Funnel

Graduated cylinder, 25- or 50-mL

Permanent marker

Spatula

Weighing dishes, 6

Safety Precautions

Acetic acid is a skin and eye irritant. Avoid contact with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please consult current Safety Data Sheets for additional safety information.

Preparation

1. Label six Erlenmeyer flasks 1–6. Using a graduated cylinder, add 25 mL of 1 M acetic acid to each flask.

2. Obtain six weighing dishes and label them 1–6.

3. Measure the appropriate amount of sodium bicarbonate into each weighing dish, according to the following table.

Table 1.

Sample	1	2	3	4	5	6
Mass $NaHCO_3$	0.50 g	1.00 g	1.50 g	2.00 g	2.50 g	3.00 g

4. Obtain six balloons. Stretch the balloons and blow them up at least once, then let out as much air is possible.

5. Use a powder funnel to add the first sodium bicarbonate sample (#1) to one of the balloons.

6. Flatten out the balloon to remove any extra air and then carefully stretch the neck of the balloon over the mouth of Erlenmeyer flask #1. Do not allow the solid to drop into the flask at this time.

7. Repeat steps 5 and 6 with the other sodium bicarbonate samples #2–6.

Procedure

1. Introduce the concepts of limiting and excess reagents by asking students how many automobiles can be assembled if the following parts are available: 140 car bodies, 520 tires, and 270 headlights.

2. Which automobile part is present in a quantity that "limits" the total number of automobiles that may be assembled? Is the limiting part the same as the part that is present in the least number? Explain. *Note:* The tires are the limiting parts in this example, even though there are more tires than anything else.

3. Show students the balloon/flask assemblies and ask them to predict what will happen when the sodium bicarbonate is added to the acetic acid in the flask. Write the reaction equation on the board.

4. Line up flasks 1–6 from right to left on the lecture desk. Lift each balloon in turn and shake it to allow the solid to fall into the solution. Make sure the neck of the balloon stays firmly attached to the flask.

5. The reactions will be immediate and vigorous. The white solids will dissolve, the solutions will start to bubble and fizz, and the balloons will become inflated.

6. Allow the reactions to proceed until the bubbling stops. Compare the size of the inflated balloons and whether all the solid has dissolved in each case. *(The balloon size should increase fairly uniformly for flasks 1–5, and then stay constant. It may be hard to tell the difference between flasks 4, 5, and 6.)*

7. Discuss the observations and carry out the necessary calculations to explain the results (See Table 2). Identify the limiting reactant and the excess reactant in each case.

Table 2.

Flask	Acetic Acid		Sodium Bicarbonate		Moles CO_2 Produced (Theoretical)
	Volume	**Moles**	**Mass**	**Moles**	
1	25 mL	0.025	0.50 g	0.0060	0.0060
2	25 mL	0.025	1.00 g	0.0119	0.0119
3	25 mL	0.025	1.50 g	0.0179	0.0179
4	25 mL	0.025	2.00 g	0.0238	0.0238
5	25 mL	0.025	2.50 g	0.0298	0.0250
6	25 mL	0.025	3.00 g	0.0357	0.0250

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. All of the leftover solutions may be disposed of down the drain with excess water according to Flinn Suggested Disposal Method #26b.

Tips

- Use new balloons each time the demonstration is performed. Once used, the balloons will have been stretched, and their inflated sizes may vary considerably from the expected size.

- Unreacted (undissolved) sodium bicarbonate will be visible in flasks 5 and 6. The solid, although normally water soluble, does not dissolve in the saturated carbon dioxide solutions.

- Add a few drops of universal indicator to each flask and observe the color changes as the pH of the solution changes over the course of the reaction.

- Excess reactants are used commercially in cases where reactions are reversible and thermodynamically unfavorable. An example of a gas-phase inorganic reaction is the synthesis of ammonia from nitrogen and hydrogen. Excess nitrogen is used to drive the reaction to completion. The synthesis of organic esters from organic alcohols and acids is another example of commercial processes that are normally carried out in the presence of excess reactants.

- Carry out the demonstration using sodium carbonate instead of sodium bicarbonate. The maximum amount of CO_2 evolution will be observed at a lower mass of sodium carbonate, due to the 2:1 mole ratio for reaction of acetic acid with sodium carbonate.

Discussion

Equation 1 summarizes the reaction between sodium bicarbonate and acetic acid.

$$NaHCO_3(s) \ + \ CH_3CO_2H(aq) \ \rightarrow \ NaCH_3CO_2(aq) \ + \ CO_2(g) \ + \ H_2O(l) \qquad \textit{Equation 1}$$

Sodium bicarbonate is the limiting reactant in Flasks 1–4. In Flasks 5 and 6, acetic acid is the limiting reactant.

Carbon Snake
Chemical Demonstration Kit

Introduction

The following experiment demonstrates an example of an exothermic reaction. It can also be used as an example of a chemical change.

Concepts

- Decomposition reaction
- Catalyst
- Combustion (energy)

Activity Overview

The purpose of this experiment is to demonstrate decomposition of chemicals by combustion.

Materials *(for each demonstration)*

Isopropyl Alcohol, 30 mL*

Sand*

Sodium Bicarbonate, 1 g*

Sucrose, 6 g*

Included in Kit

Butane Safety Lighter

Container Vial, small*

Evaporating Dish, 80 mL*

Safety Precautions

Isopropyl alcohol is a flammable liquid and a fire hazard; keep away from all flames, sparks, and heat sources. Isopropyl Alcohol is also slightly toxic by ingestion and inhalation [LD_{50}: 5045 mg/kg]. Wear chemical splash goggles, chemical-resistant gloves and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Follow all laboratory safety guidelines. Please review current Safety Data Sheets for additional safety, handling and disposal information.

Procedure

1. Mass 6 g of sucrose and pour into a small container.
2. Mass 1 g of sodium bicarbonate and also pour into the small container.
3. Mix dry chemicals in the container.
4. Pour sand, approximately 170 g, into the evaporating dish.
5. Measure 30 mL of isopropyl alcohol into a 50 mL graduated cylinder.
6. Carefully add the isopropyl alcohol over the sand making sure to cover all of the sand.
7. Add the mixture of solid chemicals toward the center of sand.
8. With a lighter, ignite the alcohol in the sand.
9. Set a timer and watch a carbon snake form out of the sand.
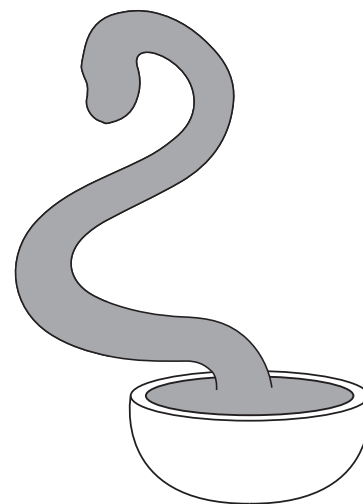
Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. The solid products of this reaction may be disposed via Flinn disposal method #26a. The sand can be reused or disposed via Flinn disposal method #26a.

Tips

- Replace sucrose with dextrose and compare reaction speed or density of the carbon snake.

- Add copper chloride or strontium chloride to the solid mixture to add color to the flame.
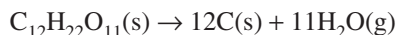
Discussion

The "snake" consists of mainly carbon that comes from the heated sugar, but which was not volatilized in the flame. The carbon is what makes the snake black. There is also Na_2CO_3 in the snake, which results from the decomposition of the baking soda when heated. The sugar and baking soda snake proceeds according to the following chemical reactions, where sodium bicarbonate breaks down into sodium carbonate, water vapor and carbon dioxide gas; while burning the sugar in oxygen produces water vapor and carbon dioxides gas. The pressure created from the release of the carbon dioxide gas causes the snake to grow. The snake gets its black appearance due to the sugar being caramelized by the heat.

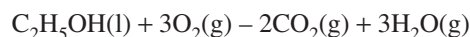- Combustion of sugar to give carbon dioxide and water vapor:

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(g)$$
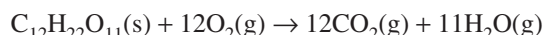
- Thermal decomposition of sugar to give carbon and water vapor:

$$C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(g)$$

- Thermal decomposition of baking soda to give sodium carbonate, carbon dioxide, and water vapor:

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

- Ignition of isopropyl alcohol:

$$C_2H_5OH(l) + 3O_2(g) - 2CO_2(g) + 3H_2O(g)$$

The *Carbon Snake Chemical Demonstration Kit* is available from Flinn Scientific, Inc.

Catalog No.	Description
AP9584	The Carbon Snake – Chemical Demonstration kit

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

Safe Swimming with Sodium

FLINN SCIENTIFIC CHEM FAX!

Introduction

No chemistry class is complete without the spectacular demonstration of alkali metals reacting with water. *Safe Swimming with Sodium* is a novel variation that is much safer to perform than the standard demonstration of simply dropping a small piece of sodium metal into a beaker of water.

Concepts

- Alkali metals
- Single replacement reactions
- Density

Materials

Lithium metal, Li, 1 small piece (optional)

Mineral oil, 200 mL

Phenolphthalein, 0.5% solution, a few drops

Sodium metal, Na, 1 small piece

Water, 200 mL

Forceps

Glass cylinder, approximately 500-mL

Ring stand and clamp

Safety Precautions

Sodium metal is a flammable, corrosive solid and is dangerous when exposed to heat or flame. Sodium reacts vigorously with moist air, water, or any oxidizer. Purchasing pre-cut pieces for performing this demonstration greatly reduces the potential hazard of the material. Sodium reacts with water to produce flammable hydrogen gas and a solution of sodium hydroxide. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Procedure

1. Clamp a hydrometer cylinder or large graduated cylinder to a ring stand for support.

2. Add about 200 mL of water to the cylinder followed by a few drops of phenolphthalein solution.

3. Add 200 mL of mineral oil, forming a layer above the water. Tilt the cylinder to reduce mixing at the interface.

4. Using forceps, drop a piece of sodium, about the size of a kernel of corn, into the cylinder and observe the reaction.
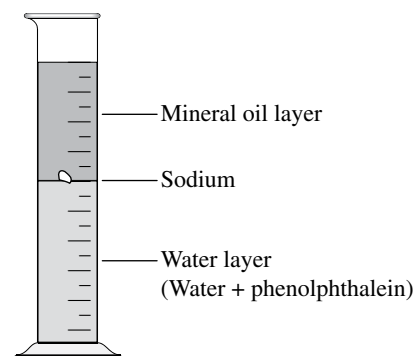
Mineral oil layer

Sodium

Water layer
(Water + phenolphthalein)

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. Do not dispose of anything until the sodium has completely reacted. The mineral oil can be stored and reused for future demonstrations and labs. The aqueous solution can be flushed down the drain with excess water according to Flinn Suggested Disposal Method #26a

Tips

- The colorless water–phenolphthalein layer can be regenerated by the addition of a small amount of dilute acid, such as
  1 M HCl. The setup can be used several times during the day.

- Sometimes during the first few reactions, the sodium metal may react very vigorously and briefly melt. If this occurs, the sodium becomes porous and "too light" to sink in the mineral oil. This piece of sodium will no longer swim—try another piece. This sometimes occurs because the mineral oil is wet or becomes wet during the setup.

- A video of this demonstration, *Safe Swimming with Sodium*, presented by Irene Cesa, is available for viewing as part

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of the Flinn Scientific "Teaching Chemistry" eLearning Video Series. Please visit the eLearning Web site at http://elearningflinnsci.com for viewing information. The video is part of the *Single Replacement Reaction* video package

Discussion

When added to the cylinder, sodium will sink in the mineral oil until it reaches the interface between the oil and water layers, at which time it reacts with water, forming hydrogen gas and sodium hydroxide, a strong base. The reactions of sodium and other alkali metals with water are classified as single replacement reactions or oxidation–reduction reactions. Sodium metal is oxidized to sodium ions, which "replace" the formal H^+ ions in water molecules. The hydrogens in water are reduced to elemental hydrogen gas (Equation 1).

$$2Na(s) \; + \; 2H_2O(l) \; \rightarrow \; H_2(g) \; + \; 2NaOH(aq) \hspace{3cm} \textit{Equation 1}$$
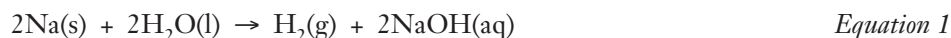
The evolution of hydrogen gas is evident, and hydrogen bubbles adhering to the sodium will carry it into the hydrocarbon layer, temporarily stopping the reaction. The amount of hydrogen and heat evolved is kept under control by this "swimming" behavior, making this demonstration quite safe. The piece of sodium repeatedly dives down to the water–hydrocarbon interface, reacts, then "swims" back up into the hydrocarbon layer until the reaction is complete. During the reaction, the piece of sodium is largely devoid of corrosion, allowing the students to view its gray, metallic appearance. The aqueous layer contains phenolphthalein and turns pink due to the production of a base, sodium hydroxide.

Density is an important physical property that can be used to separate materials or control reactions. Sodium has a density of 0.97 g/mL and sits at the interface of the water and oil layers. Lithium, in contrast, has a density of 0.54 g/cm^3, and will float on top of the hydrocarbon layer. (Try it!) The interface between two immiscible solvents is an effective site for controlling chemical reactions. Many industrial processes use this concept to react aqueous salts with nonpolar hydrocarbons.

NGSS Alignment

This laboratory activity relates to the following Next Generation Science Standards (2013):

Disciplinary Core Ideas: Middle School	**Science and Engineering Practices**	**Crosscutting Concepts**

MS-PS1 Matter and Its Interactions
    PS1.A: Structure and Properties of Matter
    PS1.B: Chemical Reactions
Disciplinary Core Ideas: High School
HS-PS1 Matter and Its Interactions
    PS1.A: Structure and Properties of Matter
    PS1.B: Chemical Reactions
	Developing and using models
	Constructing explanations and designing solutions
		Cause and effect
		Structure and function
		Stability and change

Acknowledgment

Special thanks to Ken Lyle, Duke University, Durham NC, for bringing this demonstration to our attention.

Reference

Alexander, M. D., J. Chem. Ed. 1992, 69, 418.

Materials for *Safe Swimming with Sodium* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP8916	Safe Swimming with Sodium—Chemical Demonstration Kit
S0329	Sodium, Bottle of 5 small pieces for demonstration
M0064	Mineral Oil, Light, 500 mL
P0115	Phenolphthalein Indicator Solution, 0.5%, 100 mL
AP8599	Hydrometer Cylinder, 600-mL
L0057	Lithium, 2.5 g

Consult the Flinn Scientific website for current prices.

Magic Genie
Sodium Iodide

Introduction

When sodium iodide is dropped into a flask containing 30% hydrogen peroxide, a "magical" genie appears in the form of water vapor and oxygen.

Concepts

- Exothermic reaction
- Catalysts
- Decomposition reactions

Materials

Hydrogen peroxide, 30%, H_2O_2, 50 mL

Sodium iodide, NaI, 4 g

Graduated cylinder, 50-mL or 100-mL

Volumetric flask, Pyrex®, 1000-mL

Filter paper

Safety Precautions

Hydrogen peroxide, 30%, will act as an oxidizing agent with practically any substance. It deserves the science teacher's special handling and storage attention. This substance is severely corrosive to the skin, eyes and respiratory tract; a very strong oxidant; and a dangerous fire and explosion risk. Do not heat this substance. Sodium iodide is mildly toxic, LD_{50}: 4340 mg/kg. The reaction flask will get extremely hot; use only a Pyrex flask and hold with a towel around it to prevent burns. Do not point the mouth of the flask towards yourself or anyone else. Never tightly close a vessel containing hydrogen peroxide—it may explode. Wear chemical splash goggles, chemical-resistant gloves and chemical-resistant apron. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Procedure

1. Wrap 4 g of sodium iodide in a small piece of filter paper or tissue. Staple the filter paper so that no sodium iodide leaks out.

2. Add 50 mL of the 30% hydrogen peroxide solution to a 1000-mL Pyrex volumetric flask. ***Caution:*** Wear rubber gloves when handling 30% H_2O_2. Contact with skin may cause burns.

3. Set the flask on a counter and hold the flask with a thick cloth towel. Drop in one packet of the sodium iodide solid. Point the flask up and in a safe direction away from yourself and your students as the magic genie (water–vapor) emerges from the flask. The flask will get extremely hot. The towel will hide the flask contents as well as protect your hand from the heat produced.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. Immediately clean up any liquid that may have splattered on the floor. Pour any liquid remaining in the flask down the drain with excess water. Rinse the flask thoroughly with water. Remove any remains of the sodium iodide packet and place it in the trash.

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 of atoms, structure and properties of matter, chemical reactions

Tips

- It is very important that this demonstration be done in a borosilicate (i.e., Pyrex) flask. A flask that is not borosilicate glass can crack from the evolution of heat.

- A large flask (1000-mL) is necessary because a brownish liquid can spurt out at the end of the reaction. A large flask will help prevent this from happening. The brown liquid results from the presence of free iodine produced from the extreme oxidizing ability of the 30% hydrogen peroxide.

- A thick cloth towel will prevent your students from seeing what is happening in the flask, as well as protect you from the heat evolved in the reaction. Another option is to wrap the flask in aluminum foil and decorate it like a "genie bottle". A more colorful option is to add food coloring to the flask.

- Manganese(IV) oxide can be substituted for sodium iodide in the demonstration. Both chemicals catalyze the reaction and will cause the release of oxygen from hydrogen peroxide.

- The sodium iodide packet can also be attached to a piece of thread and hung inside the flask. Attach the thread to the ou side of the flask with tape or a stopper. *Warning:* Do not use a solid stopper or cap. If the reaction starts prematurely, the pressure buildup may explode the flask. Use a one- or two-holed stopper and place it loosely on the flask.

Discussion

The Magic Genie demonstrates the decomposition of hydrogen peroxide into oxygen gas and water vapor. The decomposition is catalyzed by iodide (I^-), which is not changed during the reaction. It is an exothermic reaction and will evolve a lot of heat. The reaction is:

$$2H_2O_2(aq) \xrightarrow{\ I^-(aq)\ } 2H_2O(g) \ + \ O_2(g) \ + \ \text{Heat Energy}$$
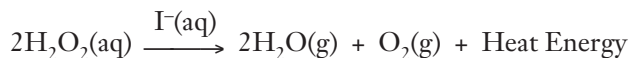
Acknowledgment

Special thanks to Jim and Julie Ealy, The Peddie School, Hightstown, NJ, who provided us with the instructions for this activity.

Reference

Stone, Charles, H. *J. Chem. Ed.*, **1944,** *21*, 300.

Materials for *Magic Genie* are available from Flinn Scientific, Inc.

Catalog No.	Description
H0037	Hydrogen Peroxide, 30%, 100 mL
H0008	Hydrogen Peroxide, 30%, 500 mL
S0083	Sodium Iodide, 25 g
GP4045	Pyrex Volumetric Flask, 1000-mL
AP2092	Magic Genie Demonstration Kit

Consult the Flinn Scientific website for current prices.

Foiled Again!
Single Replacement Reaction

Introduction

Watch aluminum foil disappear as it is added to a solution of copper(II) chloride. Observe color changes, production of a gas, formation of a solid metal, and a drastic change in temperature. Learn about the unexpected role of a catalyst in this single-replacement reaction at a metal surface.

Concepts

- Single replacement reaction
- Oxidation–reduction
- Metal activity
- Catalysis

Materials

Aluminum foil, 6″ × 12″, 2 pieces

Copper(II) chloride solution, $CuCl_2$, 1 M, 140 mL

Copper(II) sulfate solution, $CuSO_4$, 1 M, 140 mL

Sodium chloride solution, NaCl, 1 M, 140 mL

Water, distilled or deionized

Beakers, Pyrex®, 600-mL, 3

Graduated cylinder, 500-mL

Spatula

Stirring rod

Thermometer

Wood splint and matches (optional)

Safety Precautions

Copper(II) chloride solution is toxic and copper(II) sulfate solution is slightly toxic by ingestion. Hydrogen gas, a highly flammable gas, is produced in the reaction. Keep flammable materials away from the demonstration area. Wear chemical splash goggles and chemical-resistant gloves and apron. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

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 M copper(II) chloride 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 roll the foil into a cylinder that will fit into the beaker. *Note:* Do not wad up the foil tightly into a ball—this will decrease the surface area and slow down the reaction.

5. If desired, measure the temperature of the solution before adding the foil.

6. Place the aluminum foil cylinder 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—the reaction is highly exothermic.

7. Have students make detailed observations of the reaction and ask them to generate a hypothesis for the reaction(s) in the beaker. Write an equation for the reaction(s) they observe. Discuss which substances are reacting species and which are spectators in the reaction, if any.

8. 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

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

10. Have students again make detailed observations. Students will observe that no reaction occurs between aluminum and copper(II) sulfate. Why not? Have students make modifications to this 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. *Note:* Set the beaker from Part 2 aside for use in Part 3.

Part 3 — Aluminum and Copper(II) Sulfate with Sodium Chloride

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

12. Add 70 mL of 1 M NaCl solution to the beaker.

13. Have students make detailed observations. Notice that a reaction now occurs between the aluminum and the copper(II) ions, as in Part 1. 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.
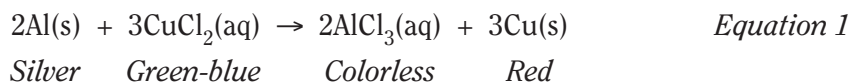
Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. 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 according to Flinn Suggested Disposal Method #26a.

Tips

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

- 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 catalyzes the reaction between aluminum and copper(II) sulfate. Observations of the reaction with the catalyst are the same as those observed when aluminum and copper(II) chloride are allowed to react. Other catalysts that have been found to catalyze this reaction include sodium bromide, potassium chloride, magnesium chloride, and hydrochloric acid. 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.

- Heating the solution of copper(II) sulfate and aluminum to 80 °C does not cause any reaction to occur. The chloride catalyst is necessary for the reaction to occur.

- A more dramatic demonstration display can be achieved by performing the demonstration in large 500-mL Pyrex® graduated cylinders or hydrometer cylinders. Loosely coil the foil into a tube and drop it into the cylinder of $CuCl_2$. The foil will slowly rise in the cylinder as gas bubbles attach to the foil surface. 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.

- Try using tap water in place of distilled water in the dilution of the $CuSO_4$ solution. Are there enough dissolved chloride ions in tap water to catalyze the reaction?

- The "Leftover Aluminum Wire Stoichiometry Lab" is available as a student laboratory kit from Flinn Scientific (Catalog No. AP4678). This experiment uses the same reaction to teach students about moles, limiting reactants, and stoichiometry.

Discussion

Aluminum foil reacts with an aqueous solution of copper(II) chloride according to Equation 1. The reaction may be classified as a single replacement, oxidation–reduction reaction.

$$2Al(s) \ + \ 3CuCl_2(aq) \ \rightarrow \ 2AlCl_3(aq) \ + \ 3Cu(s) \qquad \textit{Equation 1}$$

$$\textit{Silver} \qquad \textit{Green-blue} \qquad \textit{Colorless} \qquad \textit{Red}$$

The oxidation of aluminum metal to aluminum(III) ions (Al^0 to Al^{3+}) is inferred from the dissolving of the aluminum foil and is represented by the oxidation half-reaction below. The simultaneous reduction of copper(II) ions to copper metal (Cu^{2+} to Cu^0) results in the formation of solid copper metal according to the reduction half-reaction below. As copper(II) ions are reduced to copper, the green-blue color of the solution fades until it is colorless—the indication that the reaction is complete and all of the copper(II) ions have been reduced.

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$$ Oxidation Half-Reaction
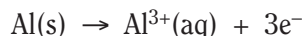
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$ Reduction Half-Reaction
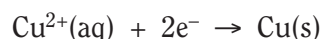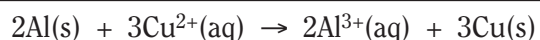
$$2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$$ Overall Balanced Equation

It is 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 consumes only a small amount of the aluminum.

$$2Al(s) + 6 H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g) \qquad \textit{Equation 2}$$
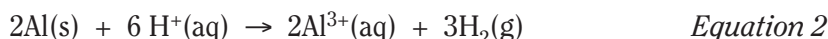
Why does aluminum react with copper(II) chloride and not with copper(II) sulfate or copper(II) nitrate? And why does it then react with copper(II) sulfate when chloride ions are added?

This interesting phenomenon was discovered serendipitously by a teacher in the classroom. The teacher wanted to perform the aluminum plus copper(II) chloride reaction with the class, but had run out of copper(II) chloride. Thinking that the net ionic equation should be the same with copper(II) sulfate, the teacher mixed aluminum with cupric sulfate. Nothing happened! After a few grams of salt were added, however, the reaction took off. This was unpredicted and raised a series of questions.

The first possible explanation is that the chloride ions catalyze the reaction at the metal surface. Other halide salts were tried and it was found that bromide ions also catalyzed the reaction. The second possible explanation has to do with the aluminum oxide coating. Normally, a chemically inert coating of aluminum oxide on the surface of aluminum metal protects the metal from oxidation. In the presence of chloride ions, however, the coating is breached and the underlying aluminum reacts. Because chloride is a relatively small ion, it is able to diffuse into and through the 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 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.

Colorful Iron Complexes
Chemical Demonstration Kit

Introduction

Easily distinguish between solutions of iron(II) and iron(III) ions by performing re-dox reactions between iron's two oxidation states. Simply add various complex ions to solutions of iron(II) or iron(III)—observe formation of the beautifully-colored Prussian blue precipitate or of the deep blood-red complex, the confirming test for iron(III).

Concepts

- Complex ions
- Oxidation–reduction
- Oxidation numbers
- Transition metals

Materials (for each demonstration)

Iron(III) chloride solution, 0.02 M, $FeCl_3 \cdot 6H_2O$, 60 mL*

Iron(II) sulfate solution, 0.02 M, $FeSO_4 \cdot 7H_2O$, 60 mL*

Potassium ferricyanide solution, 0.1 M, $K_3Fe(CN)_6$, 10 drops*

Potassium ferrocyanide solution, 0.1 M, $K_4Fe(CN)_6 \cdot 3H_2O$, 10 drops*

Potassium thiocyanate solution, 0.1 M, KSCN, 10 drops*

Water, distilled or deionized, 120 mL

Colorful Iron Complexes Worksheet

Graduated cylinder, 25-mL

Stoppers to fit tubes, 6*

Test tubes, 25 × 150 mm, 6*

Test tube rack

Materials included in kit.

Safety Precautions

This activity requires the use of hazardous components and/or has the potential for hazardous reactions. Potassium ferricyanide, potassium ferrocyanide, and potassium thiocyanate are dangerous if heated or in contact with concentrated acids since toxic hydrogen cyanide gas may be liberated. Potassium thiocyanate is moderately toxic by ingestion. Potassium ferricyanide, potassium ferrocyanide, and ferrous sulfate are slightly toxic by ingestion. Iron(II) sulfate is corrosive to skin, eyes, and mucous membranes. Iron(III) chloride may be a skin and tissue irritant. Avoid body contact with all chemicals. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information. Wash hands thoroughly with soap and water before leaving the laboratory.

Pre-Demonstration Activity

1. Copy enough worksheets for the class using the provided Colorful Iron Complexes Worksheet master.

2. Before performing the demonstration, write the chemical formulas for each of the five solutions on the chalkboard.

3. Have students determine the oxidation number of the iron atom in each compound using rules for assigning oxidation numbers in any standard chemistry text.

Procedure

1. Place six test tubes in a test tube rack. Label the tubes 1–6 and then label tubes 1–3 as Fe^{2+} and tubes 4–6 as Fe^{3+}.

2. Add approximately 20 mL of 0.02 M iron(II) sulfate solution and 20 mL of distilled or deionized water to test tubes 1–3. Stopper the tubes and invert to mix.

CHEM-FAX™...*makes science teaching easier.*

4. Add approximately 20 mL of 0.02 M iron(III) chloride solution and 20 mL of distilled or deionized water to test tubes 4–6. Stopper the tubes and invert to mix.

Part A. Ferrocyanide ions, Fe(CN)$_6^{4-}$ [Iron in the +2 oxidation state]

5. Add 5 drops of 0.1 M potassium ferrocyanide solution to Tube 1. Since both sources of iron are in the +2 state, the notable deep-blue precipitate does not form. Instead a light blue precipitate of potassium iron(II) hexacyanoferrate(II), K$_2$Fe[Fe(CN)$_6$], forms according to Equation 1.

$$Fe^{2+} + 2K^+ + [Fe(CN)_6]^{4-} \rightarrow K_2Fe[Fe(CN)_6](s) \qquad\qquad \textit{Equation 1}$$
Light blue ppt
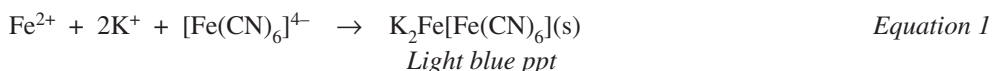
6. Add 5 drops of 0.1 M potassium ferrocyanide solution to Tube 4. A deep-blue precipitate will form according to equation 2, due to the presence of both iron(II) and iron(III) ions. This resulting deep blue precipitate is iron(III) hexacyanoferrate(II), Fe$_4$[Fe(CN)$_6$]$_3$, or Prussian blue. Have students determine the oxidation number of each iron atom in the complexes in Equations 1 and 2.

$$4Fe^{3+} + 3[Fe(CN)_6]^{4-} \rightarrow Fe_4[Fe(CN)_6]_3 \qquad\qquad \textit{Equation 2}$$
Prussian blue
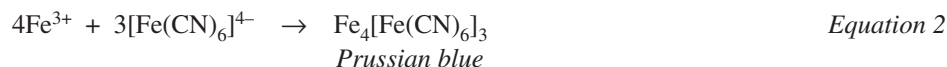
Upon standing for 5–10 minutes, the solution in Tube 1 will turn darker blue as the iron(II) is slowly oxidized to iron(III) by atmospheric oxygen to form the same Prussian blue precipitate as in equation 2.

Part B. Ferricyanide ions, Fe(CN)$_6^{3-}$ [Iron in the +3 oxidation state]

7. Add 5 drops of 0.1 M potassium ferricyanide solution to Tube 2. A deep-blue precipitate will form with the iron(III) sulfate. In this reaction, the ferricyanide ions, [Fe(CN)$_6$]$^{3-}$, oxidize iron(II) to iron(III) forming ferrocyanide ions, [Fe(CN)$_6$]$^{4-}$, according to Equation 3:

$$Fe^{2+} + [Fe(CN)_6]^{3-} \rightarrow Fe^{3+} + [Fe(CN)_6]^{4-} \qquad\qquad \textit{Equation 3}$$

The products of equation 3, the iron(III) ions and ferrocyanide ions, then combine to form iron(III) hexacyanoferrate(II) or Prussian blue, according to equation 2 above.

8. Add 5 drops of 0.1 M potassium ferricyanide solution to Tube 5. A brown solution is observed, indicating no reaction since both iron sources are in the +3 oxidation state, as shown below:

$$Fe^{3+} + [Fe(CN)_6]^{3-} \rightarrow \textit{No Reaction}$$
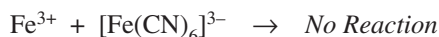
Part C. Thiocyanate ions, SCN$^-$

9. Add 5 drops of 0.1 M potassium thiocyanate solution to Tube 3. Some light red-brown coloring may appear due to slight oxidation of Fe^{2+} to Fe^{3+}, but no visible reaction is observed.

$$Fe^{2+} + 3SCN^- \rightarrow \textit{No Reaction}$$

10. Add 5 drops of 0.1 M potassium thiocyanate solution to Tube 6. A deep-red complex will form with the iron(III) sample according to Equation 4. This is a positive indicator test for the iron(III) ion.

$$Fe^{3+} + 3SCN^- \rightarrow Fe(SCN)_3 \qquad\qquad \textit{Equation 4}$$
Blood-red complex

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. Dispose of the solutions in the test tubes down the drain with excess water according to Flinn Suggested Disposal Method #26b. Flush any excess potassium thiocyanate, iron(II) sulfate, or iron(III) chloride solution down the drain according to Flinn Suggested Disposal Method #26b. Dispose of excess ferricyanide and ferrocyanide solution according to Flinn Suggested Disposal Method #14.

IN5946

Tips

- This kit contains enough materials to perform the demonstration seven times: six reusable test tubes and stoppers, 500 mL of the ferrous sulfate and iron(III) chloride solutions, and 10 mL each of the potassium ferrocyanide, potassium ferricyanide, and potassium thiocyanate solutions.

- It is very helpful to write Fe^{2+} and Fe^{3+} on the appropriate test tubes and to have the students fill in the provided worksheet as the demonstration is performed. Even the most skilled students (and teachers) may find it hard to keep track of which form of iron is in the tube and which is being added. The demo will have the greatest effect if this confusion is removed.

- Use this kit in the oxidation–reduction unit when teaching students about oxidation numbers. Have students perform the pre-demonstration activity to assure that they can determine the oxidation number for iron in each compound.

- After performing the demonstration, you may wish to use the tests to determine if an "unknown" contains iron(II) or iron(III) ions. You can use a few milliliters of either the iron(II) sulfate solution or the iron(III) chloride solution as the unknown. Or you can test any iron solution in your chemical stockroom. Enough of the testing reagents are provided for testing unknowns in each class.

Discussion

Many transition metals exhibit the ability to exist as relatively stable ions in different oxidation states. Iron can be found as the Fe^{2+} ion [iron(II)] in some compounds. Iron is also found as the Fe^{3+} ion [iron(III)] in other compounds.

The variable valence states can be explained by looking at the electron configuration of iron, which is $[Ar]4s^23d^6$. When transition metal atoms form positive ions, the outer s electrons are lost first because the inner d sublevels are lower in energy (more stable) than the outer s sublevels. In the iron(II) ion, the two 4s electrons have been lost, leaving $[Ar]3d^6$. In the iron(III) ion, the two 4s electrons and one 3d electron have been removed, leaving $[Ar]3d^5$. The iron(III) ion is more stable than the iron(II) ion since its d orbital is half-filled, containing five electrons, while that of iron(II) is one more than half-filled. Half-filled orbitals (and filled orbitals) have been shown to have greater stability. Therefore, a compound or a solution containing the iron(II) ion will slowly oxidize to the iron(III) state on exposure to air due the greater stability of the Fe^{3+} ion.

In order to distinguish between iron(II) and iron(III) ions, potassium ferrocyanide $[K_4Fe(CN)_6 \cdot 3H_2O]$ and potassium ferricyanide $[K_3Fe(CN)_6]$ complexes are used in this experiment. The cyano group in each complex has a charge of –1 and potassium has a charge of +1. Thus, the complex ferrocyanide, $Fe(CN)_6^{4-}$, contains iron in the +2 oxidation state while the complex ferricyanide, $Fe(CN)_6^{3-}$, contains iron in the +3 oxidation state.

A deep-blue (Prussian blue) precipitate results when either complex ion combines with iron in a different oxidation state from that present in the complex. The deep-blue color of the precipitate is due to the presence of iron in both oxidation states in the cyano complex. This provides a means of identifying either iron ion. Thus when a solution of iron(II) is mixed with ferricyanide [iron(III)], a deep-blue precipitate is formed; likewise, when a solution of iron(III) is mixed with ferrocyanide [iron(II)], a deep-blue precipitate is formed.

The deep-blue precipitate, Prussian blue, has the composition of $Fe_4[Fe(CN)_6]_3$. Prussian blue has been used as a pigment in printing inks, paints, cosmetics (eye shadow), artist colors, carbon paper, and typewriter ribbons.

The thiocyanate ion, SCN^-, provides an excellent confirming test for the Fe^{3+} ion. The soluble, blood-red $Fe(SCN)_3$ complex is formed from the Fe^{3+} ion, while no complex is formed with the Fe^{2+} ion.

Connecting to the National Standards

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

Unifying Concepts and Processes: Grades K–12
    Systems, order, and organization
    Evidence, models, and explanation
Content Standards: Grades 9–12
    Content Standard B: Physical Science, structure and properties of matter, chemical reactions

     IN5946

References

Bilash, B.; Gross, G.; Koob, J. *A Demo A Day: Another Year of Chemical Demonstrations, Vol 2;* Flinn Scientific: Batavia, IL, 1998; pp 244–246.

Tzimopoulos, N. D.; Metcalfe, H. C.; Williams, J. E.; Castka, J. F. *Modern Chemistry Laboratory Experiments;* Holt, Rinehart and Winston: New York, 1990; p 63.

The *Colorful Iron Complexes—Chemical Demonstration Kit* is available from Flinn Scientific, Inc.

Catalog No.	Description
AP5946	Colorful Iron Complexes—Chemical Demonstration Kit

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

IN5946

Simple Electrolysis

Decomposition of Water

Introduction

Demonstrate the decomposition of water in simple but colorful fashion on an overhead projector using just a 9-volt battery and alligator clips. No need for expensive power supplies or fancy glassware. Color changes produced in universal indicator solution reveal the pH changes that occur at the positive and negative electrodes as the reaction proceeds.

Concepts

- Decomposition reaction
- Oxidation–reduction
- Electrolysis
- pH

Materials

Sodium sulfate solution, Na_2SO_4, 1 M, 40 mL

Universal indicator solution, 2–3 mL

Water, distilled

Battery cap with alligator clip leads

Battery, 9-V

Beaker

Overhead projector

Pencil lead, 0.9 mm

Petri dish

Stirring rod

Safety Precautions

Universal indicator is an alcohol-based solution and is flammable; do not use near an open flame. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Procedure

1. Place the two halves of a Petri dish on the projection stage of an overhead projector. Pour enough sodium sulfate solution into each half of the Petri dish to just cover the bottom of each half dish.

2. Add 2–3 mL of universal indicator to each solution to give each a rich, transparent green color.

3. Break a pencil lead in half. Attach the leads to opposite sides of the Petri dish bottom with the alligator clips. Make sure the tip of each lead is submerged in the green solution and the alligator clips remain out of the solution. *Note:* The solution in the Petri dish top will serve as a control.

Figure 1.

4. To start the demonstration, clip the 9-volt battery into the snaps on the battery cap (Figure 1).

5. Let the demonstration run for 5–10 minutes and note the changing colors over time and the production of gases (bubbling) at each electrode. A deep purple color forms immediately at the negative electrode, while at the positive electrode a red-orange color appears more slowly. Over time, a wider range of universal indicator colors—purple, blue, green, yellow, and red—may develop.
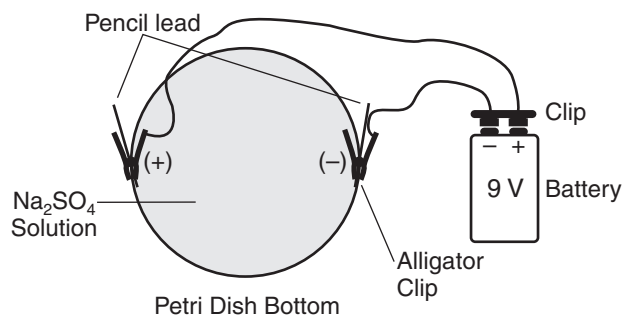
Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. The waste solutions may be disposed of down the drain with excess water according to Flinn Suggested Disposal Method #26b.

1

Connecting to the National Standards

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

Unifying Concepts and Processes: Grades K–12
    Systems, order, and organization
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Content Standards: Grades 5–8
    Content Standard B: Physical Science, properties and changes of properties in matter, transfer of energy
Content Standards: Grades 9–12
    Content Standard B: Physical Science, structure and properties of matter, chemical reactions, interactions of energy and matter

Tips

- Sodium sulfate is used as a source of dissolved ions to increase the current flow through the solution. In the absence of added electrolyte, no reaction will occur when the battery is connected to the pencil leads—there are no ions to "carry" the current through the solution. The rate of electrolysis increases as the concentration of sodium sulfate increases.

- The sulfate ion is an extremely weak base. The initial indicator color for the electrolysis solution may be more blue-green rather than green. Test a small amount of sodium sulfate stock solution with universal indicator—the solution should turn green. If the solution is teal-colored, add one drop of 0.1 M hydrochloric acid to the stock solution.

Discussion

When an electric current is passed through a solution of water containing an ionic salt, such as sodium sulfate, the water molecules break apart or decompose into their constituent elements, hydrogen and oxygen. The overall reaction occurs as two separate, independent half-reactions. Reduction of hydrogen atoms in water to elemental hydrogen (H_2) occurs at the negative electrode (the cathode), while oxidation of the oxygen atoms in water to elemental oxygen (O_2) occurs at the positive electrode (the anode). Each half-reaction is accompanied by the production of OH^- or H^+ ions, as shown below:

$$4e^- + 4H_2O \rightarrow 2H_2(g) + 4OH^- \qquad \qquad \textit{Reduction}$$

$$2H_2O \rightarrow O_2(g) + 4H^+ + 4e^- \qquad \qquad \textit{Oxidation}$$

The overall reaction is the decomposition of water (Equation 1).

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g) \qquad \qquad \textit{Equation 1}$$

Universal indicator is an acid–base indicator that is different colors at different pH values (Table 1). Excess OH^- ions produced at the cathode cause the pH to increase, resulting in a color change of the universal indicator solution from green (neutral, pH 7) to purple (basic, pH $\geq$ 10). Excess H^+ ions produced at the anode cause the pH to decrease, resulting in a color change of the universal indicator solution from green to orange-red (acidic, pH $\leq$ 4). The electrolysis half-reactions can also be followed by observing the production of gas bubbles at each electrode.

Table 1. Universal Indicator Color Changes

pH	4	5	6	7	8	9	10
Color	Red	Orange	Yellow	Green	Blue/green	Dark blue	Purple

Materials for *Simple Electrolysis—Decomposition of Water* are available from Flinn Scientific, Inc.

Catalog No.	Description
A0352	Sodium Sulfate Solution, 1 M
U0001	Universal Indicator Solution, 100 mL
GP3019	Petri Dish

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

Cool Light
Chemiluminescence

Introduction

Chemiluminescence demonstrations are popular with students and teachers alike. When light is produced without heat, that's cool!

Concepts

- Chemiluminescence
- Oxidation–reduction
- Catalyst

Materials

Hydrogen peroxide, H_2O_2, 3%, 15 mL

Luminol, 0.1 g

Potassium ferricyanide, $K_3Fe(CN)_6$, 0.7 g

Sodium hydroxide solution, NaOH, 5%, 50 mL

Water, distilled or deionized, 2000 mL

Erlenmeyer flasks, 1-L, 2

Erlenmeyer flask, 2-L

Funnel, large

Graduated cylinder, 50-mL

Ring stand and ring

Safety Precautions

Hydrogen peroxide is an oxidizer and skin and eye irritant. Sodium hydroxide solution is corrosive, and especially dangerous to eyes; skin burns are possible. Much heat is released when sodium hydroxide is added to water. If heated to decomposition or in contact with concentrated acids, potassium ferricyanide may generate poisonous hydrogen cyanide. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Preparation

1. Prepare Solution A by adding 0.1 g of luminol and 50 mL of 5% sodium hydroxide solution to approximately 800 mL of distilled or deionized (DI) water. Stir to dissolve the luminol. Once dissolved, dilute this solution to a final volume of 1000 mL with DI water.

2. Prepare Solution B by adding 0.7 g of potassium ferricyanide and 15 mL of 3% hydrogen peroxide to approximately 800 mL of DI water. Stir to dissolve the potassium ferricyanide. Once dissolved, dilute this solution to a final volume of 1000 mL with DI water.

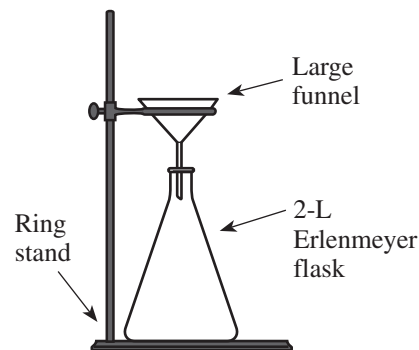3. Set up the demonstration equipment as shown in Figure 1.

Large funnel

2-L Erlenmeyer flask

Ring stand

Figure 1.

Procedure

1. Turn down the lights. The room should be as dark as possible.

2. Recruit a volunteer or second presenter. Together, pour Solution A and Solution B into the large funnel simultaneously. As the two solutions mix, chemiluminescence begins.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. The final solution may be disposed of according to Flinn Suggested Disposal Method #26b.

Tips

- This demonstration is especially appealing if it is set up so the students can see the mixture through spiraling clear plastic tubing. Simply follow the directions for setup at the end of these instructions. This type of apparatus gives a large surface area for light to be emitted and also provides a flowing effect along with the luminescence, thus increasing the overall visual impact.

Discussion

Chemiluminescence is defined as the production or emission of light that accompanies a chemical reaction. Light emission results from the conversion of chemical energy into light energy due to changes in the composition of a chemiluminescent material. The "flame test" colors observed when different metal salts are burned in a Bunsen burner flame are examples of a type of chemiluminescence known as pyroluminescence. The glow of solid phosphorus in air is another classic example of chemiluminescence—light, along with some heat, is produced when the phosphorus undergoes an oxidation reaction. The oxidation of luminol (3-aminophthalhydrazide) in this demonstration illustrates a type of "cool light" chemiluminescence in which little or no heat is produced.

The light-producing chemical reactions of luminol were discovered by H.O. Albrecht in 1928. Since that time numerous procedures have been developed to produce light using luminol. Experiments have shown that the following "ingredients" are necessary for luminol to exhibit chemiluminescence—a basic (alkaline) pH, an oxidizing agent, and a catalyst. In this demonstration, the oxidizing agent is hydrogen peroxide, the catalyst is the iron(III) cation in potassium ferricyanide, and sodium hydroxide is used to maintain the basic pH needed for the reaction to occur.

Oxidation of luminol and the resulting chemiluminscence occurs in the following sequence of reactions:

- Sodium hydroxide acts as a base and converts luminol (structure I) into a dianion.

- Hydrogen peroxide oxidizes the dianion form of luminol to the aminophthalate ion (structure II), which is produced in an excited electronic state. Nitrogen gas is also released in this reaction.

- The excited aminophthalate ion decays to a lower energy ground state and gives off light in the process. The emitted light has a wavelength of 425 nm, which is in the blue region of the visible spectrum.
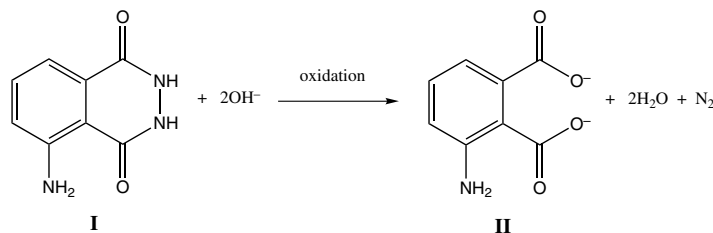
The "Cool Light" demonstration has value not only for the obvious reason of demonstrating chemiluminescence, but also to show the effects of a catalyst, pH, and temperature on a reaction, as well as how reaction rates are affected by concentration.

Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Cool Light* activity, presented by Irene Cesa, is available in *Chemiluminescence*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

Materials for *Cool Light* are available from Flinn Scientific, Inc.

Materials required to perform this activity are available in the *Cool Light—Demonstration Kit* available from Flinn Scientific. Materials may also be purchased separately.

| Catalog No.	Description
AP8627	Cool Light—Demonstration Kit
L0078	Luminol, 5 g
S0074	Sodium Hydroxide, 100 g
P0050	Potassium Ferricyanide, 100 g
H0009	Hydrogen Peroxide, 3%, 500 mL
GP9155	Erlenmeyer Flask, Borosilicate Glass, 2-L

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

Energy in Photons
Light Energy Demonstration

Introduction

Students often confuse the concepts of intensity of light and energy of light. This demonstration provides a clear way to demonstrate that the intensity, or brightness, of light is NOT the same as the amount of energy a particular color of light possesses.
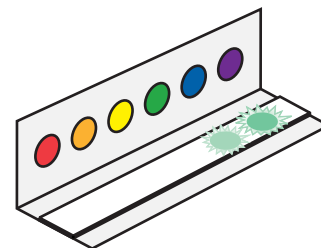
Concepts

- Phosphorescence
- Photoelectric effect
- Absorbance vs. transmittance
- Energy and wavelength of light

Materials

Energy in Photons Demonstrator Card—assembled and ready to use

Light source—classroom lights work well

Procedure

1. Open the Demonstrator Card and show the class the phosphorescent strip. Explain that the common term for phosphorescence is "glow-in-the-dark" and that this strip will glow in the dark.

2. Directly expose the entire phosphorescent strip to the classroom lights for about 15 seconds. Now turn off all the classroom lights and completely darken the room. The entire strip will glow brightly for several minutes and then begin to fade. Once students are satisfied with the glow, turn the classroom lights back on.

3. Show the six colored filters on the Demonstrator Card to the class. Hold the Demonstrator Card up to the light so that the color of light transmitted through each filter is clearly visible. Observe that the color of light transmitted through each filter is the same color as the filter.

4. Close the Demonstrator Card tightly making sure that no light can reach the phosphorescent strip from the sides. Paper clip the sides closed and place it top down on the desk for 2–3 minutes. Have the class predict what will happen if the Demonstrator Card is closed and only "filtered" light is allowed to shine upon the phosphorescent strip.

5. After 2–3 minutes, expose the closed Demonstrator Card to the classroom lights for at least 30 seconds.

6. Turn off all the classroom lights and completely darken the room again. Open the Demonstrator Card and show the phosphorescent strip to the class. The strip will only glow under the blue and violet filters!

7. Compare the class' predictions with the actual results. Many students will be surprised that the brighter colors, like yellow and orange, do not let enough light through to cause the strip to glow. Explain that even though these colors may look brighter, or more intense, only the blue and violet filters let through light with enough energy to make the phosphorescent strip glow.

8. Have the class estimate the maximum wavelength needed to excite the phosphorescent strip (and cause it to glow) by using an approximate wavelength for each filter color. *(The maximum wavelength is about 480 nanometers.)*

9. Calculate the minimum energy a photon must have to cause the strip to phosphoresce, or glow:

$$E = hc/\lambda$$

where E = energy in joules, h = Planck's constant = 6.626×10^{-34} J·sec, c = speed of light = 2.998×10^8 m/sec, and λ = wavelength in meters.

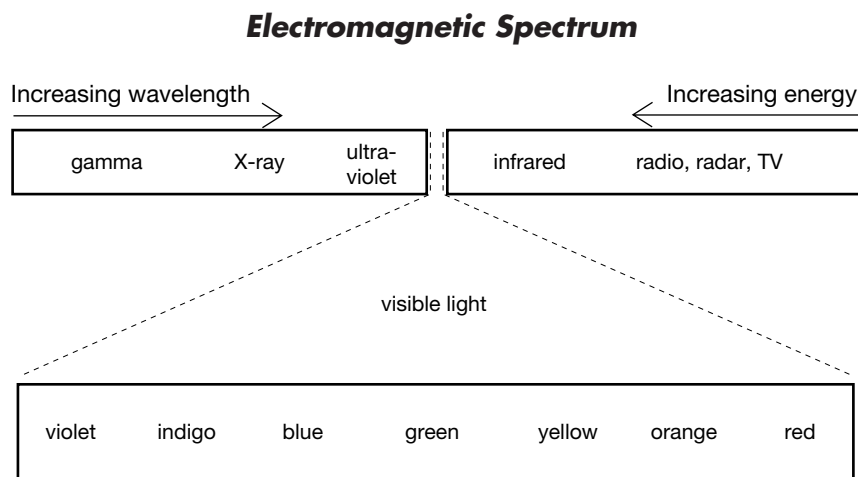The minimum photon energy required for phosphorescence is 4.1×10^{-19} J.

$$E = \frac{(6.626 \times 10^{-34} \text{ J·sec})(2.998 \times 10^8 \text{ m/sec})}{(480 \text{ nm})(1 \times 10^{-9} \text{ m/nm})} = 4.1 \times 10^{-19} \text{ J}$$

Tips

• The spots under the blue and violet filters should always glow brightly, but sometimes a small glow can be seen under some of the other filters. This is partly due to stray light creeping in around the filters. Try to keep the Demonstrator Card as tightly closed as possible to prevent any extra light from getting to the phosphorescent strip. The slight glow can also be explained by noting that the filters are not "perfect" filters in removing lights of other wavelengths.

• To store the Demonstrator Card, insert the solid black strip into the Card so that it is covering the phosphorescent strip. Close the Demonstrator Card and store it in the envelope. This will protect the phosphorescent strip from light and prolong its useful life.

• To reinforce the concept that it is the energy of the light that matters rather than the intensity, obtain two light sources of different intensity, such as a 40 W bulb and a 100 W bulb. Shine each lightbulb onto the closed Demonstration Card—in both cases only the violet and blue filters will cause the strip to phosphoresce.

Discussion

Visible light is a form of electromagnetic radiation. All forms of electromagnetic radiation consist of oscillating electric and magnetic fields traveling at a constant speed, the speed of light, 2.998×10^8 m/s. Other familiar forms of electromagnetic radiation include γ-rays, X-rays, ultraviolet radiation, infrared radiation, and radio waves.

Electromagnetic Spectrum

Increasing wavelength →　　　　　　　　　　　　　　← Increasing energy

gamma	X-ray	ultra-violet

infrared	radio, radar, TV

visible light

violet	indigo	blue	green	yellow	orange	red

The visible portion of the electromagnetic spectrum spans the wavelength region from about 400 to 700 nm. The human eye sees light of 400 nm as violet and 700 nm as red. Based on the quantum nature of light, the wavelength of light is inversely proportional to its energy (E) according to Planck's Law, $E = hc/\lambda$, where h is Planck's constant and c is the speed of light. Because it has a shorter wavelength than red light, violet light is greater in energy than red light. As the color of light changes, so does the amount of energy it possesses. White light, such as that from an incandescent or fluorescent light, contains all of the colors in the visible spectrum.

A second characteristic of light, in addition to its energy, is its intensity. Intensity corresponds to the brightness of light. According to the theories of classical physics, energy should be proportional to intensity, so that the more intense a light source, the more energy it gives off. Under this assumption, very bright (yellow) light should cause the phosphorescent strip in the Demonstrator Card to glow. However, this is not observed! Instead, the phosphorescent strip glows only when blue or violet light is shined on it. This phenomenon is analogous to the photoelectric effect, one of the classical paradoxes that led to the discovery of the quantum nature of light and Planck's Law.

The glowing of the phosporescent strip in the Demonstrator Card is due to two processes that occur in tandem. First, the phosphor must absorb light of the proper energy to excite an electron from its lowest energy ground state to a higher energy excited state. Once in the excited state, the electron has a natural tendency to want to return back to its more stable ground state. In doing so it releases energy in the form of light—this is the emission of photons that is observed as "glow-in-the-dark" phosphorescence. The phosphorescent material has a critical wavelength or energy of light. If a light is shined on the phosphorescent strip and it contains photons whose energy is greater than the energy needed to cause the strip to glow, it will glow. If the intensity of this source is increased, the strip will glow more brightly. If, however, a light source is shined on the phosphorescent strip that contains photons whose energy is less than the critical energy for the phosphorescent strip, no glowing will occur, no matter how bright the light source.

The cutoff wavelength for exciting the phosphorescent strip is about 480 nm. Photons (light energy) with a higher wavelength (less energy) will not cause the strip to glow, while photons with shorter wavelengths (more energy) will cause the phosphorescent glow. The cutoff wavelength of 480 nm is right on the border between blue and green light. Therefore, blue or violet photons which are transmitted through the filter will contain enough energy to excite the electrons in the phosphorescent material and cause it to glow. The green, yellow, orange and red filters, in contrast, absorb the blue and violet photons (their complementary colors) and do not allow them to be transmitted. Therefore, the light transmitted through these filters does not contain enough energy to excite the phosphorescent strip and no glow is observed.

The following table lists the wavelengths associated with each of the filter colors and their complements.

Representative Wavelength, nm	Wavelength Region, nm	Filter Color	Complementary Color
410	400–425	Violet	Yellow-green
470	425–480	Blue	Orange
490	480–500	Blue-green	Red
520	500–560	Green	Red-Violet
565	560–580	Yellow-green	Violet
580	580–585	Yellow	Violet
600	585–650	Orange	Blue
650	650–700	Red	Blue-green

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Catalog No. AP4848

Publication No. 4848

Fluorescent Dye Kit
Chemical Demonstration Kit

Introduction

Color is a result of the interaction of light with matter. The color that a solution appears to the human eye can change depending on the nature of the light source used to illuminate it. In this demonstration, four solutions that appear one color under visible light will "change" colors when exposed to an ultraviolet (black) light.

Chemical Concepts

- Fluorescence
- Transmittance
- Absorbance
- Emission

Materials

Eosin Y solution, 1%, 5 mL

Ethyl alcohol, 95%, 500 mL

Fluorescein solution, 1%, 15 mL

Rhodamine B solution, 1%, 1 mL

Water, distilled or deionized, 1000 mL

Water, tonic, 500 mL

Beakers, 600-mL, 4

Graduated cylinders, 10-mL, 3

Stirring rods, 3

Ultraviolet light source—black light

Visible light source—classroom lights work well

Safety Precautions

Please review current Material Safety Data Sheets for additional safety, handling, and disposal information. This activity requires the use of hazardous components and/or has the potential for hazardous reactions. Ethyl alcohol is flammable and a dangerous fire risk. Addition of denaturant makes ethyl alcohol poisonous—it cannot be made nonpoisonous. Dye solutions will easily stain hands and clothing; avoid all contact with skin and clothing. Do not look directly at the black light; its high-energy output can be damaging to eyes. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron.

Preparation

Beaker 1: Pour approximately 500 mL of tonic water into one of the 600-mL beakers.

Beaker 2: Add 15 mL of the 1% fluorescein solution to the second 600-mL beaker. Dilute the fluorescein solution by adding enough distilled or deionized water to reach the 500-mL mark on the beaker. Stir.

Beaker 3: Add 5 mL of the 1% eosin Y solution to the third 600-mL beaker. Dilute the eosin Y solution by adding enough ethyl alcohol to reach the 500-mL mark on the beaker. Stir.

Note: Eosin Y is soluble in water, but the fluorescence is not nearly as strong in water as it is in ethyl alcohol.

Beaker 4: Add 1 mL of the 1% rhodamine B solution to the fourth 600-mL beaker. Use a graduated cylinder or add 15 drops from the dropping bottle. Dilute the rhodamine B solution by adding enough distilled or deionized water to reach the 500-mL mark on the beaker. Stir.

Procedure

1. Place the four beakers in a row on the demonstration table in the following order: tonic water, fluorescein solution, eosin Y solution, and rhodamine B solution. Set the beakers on a light box or place a white background both below and behind the beakers. This will make it easier to clearly see the colors of each of the solutions.

2. Turn on the light box or just use the classroom lights to observe the colors of each of the solutions.

3. Turn off all of the lights and completely darken the room. The demonstration is most dramatic in a completely dark room.

4. Turn on the black light and place it on the demonstration table in front of the row of beakers. Do not look directly at the black light. Observe the fluorescence from each beaker. Note that the fluorescent color of each solution is different than the color observed under the normal classroom lights!

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. Flush all solutions down the drain according to Flinn Suggested Disposal Method #26b.

Tips

- The difference in the colors of the solutions under the classroom lights compared to the black light is most obvious in a completely darkened room. Try to extinguish all light sources.

- The solutions will last all day, even a whole week; however, some evaporation will occur. The solutions will keep for an extended period of time if the beakers are covered with Parafilm.®

- The tonic water does not have to be carbonated for fluorescence to occur—it will still fluoresce even if it goes flat.

- The prescribed dilutions listed above are not strict. The same effects are noticeable over a range of dilutions; however, the fluorescence is most easily observed in dilute solutions. For best results, solutions more concentrated than those suggested above are not recommended.

Discussion

The Electromagnetic Spectrum

In 1865, J. C. Maxwell showed that visible light is a form of *electromagnetic radiation*. All forms of electromagnetic radiation consist of oscillating electric and magnetic fields travelling at a constant speed, the speed of light, 2.998×10^8 m/s. Other familiar forms of electromagnetic radiation include microwave radiation from a microwave oven, X-rays, the infrared radiation in heat from a fire, and radio waves. Together, all forms of electromagnetic radiation make up the electromagnetic spectrum.

Electromagnetic Spectrum
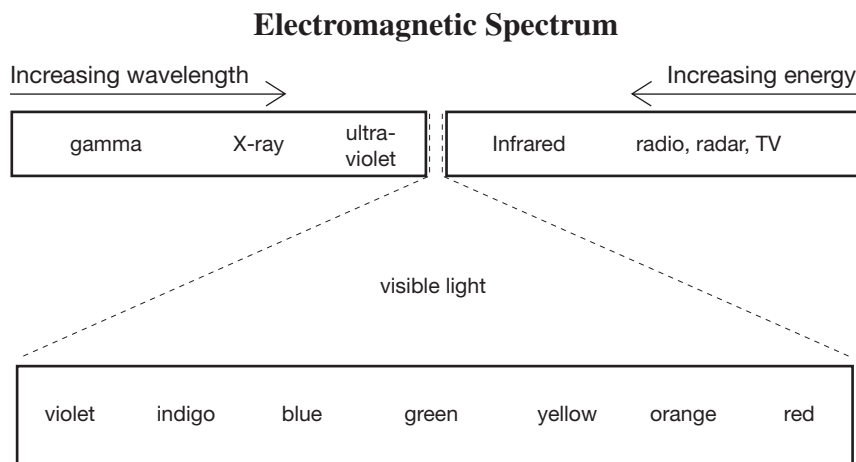
The visible portion of the electromagnetic spectrum is only a small part of the entire spectrum. It spans the wavelength region from about 400 to 700 nm. We see light of 400 nm as violet and 700 nm as red. Because wavelength is inversely proportional to energy according to the equation $E = hc/\lambda$, violet light is higher energy light than red light. The color of light we see with the

human eye varies from red to violet (low to high energy) according to the familiar phrase ROY G BIV: red, orange, yellow, green, blue, indigo, violet. As the color of the light changes, so does the amount of energy it possesses. White light, like that from normal classroom lights, contains all of the colors in the visible spectrum.

A typical black light (such as Flinn Catalog No. AP9030) gives off UVA light. UVA is ultraviolet light in the wavelength range from approximately 320 to 400 nm; therefore, it is higher energy light than visible light. The human eye cannot see ultraviolet light. Therefore, a substance that gives off ultraviolet light (and does not also give off visible light) will appear clear or colorless. The black light gives off higher energy light than the normal classroom lights. Because the light from a black light can be seen with the human eye, it clearly must give off some wavelengths of visible light in addition to the UVA wavelengths it gives off. These additional wavelengths are in the low 400's, so the black light appears purple to the human eye.

Absorption and Transmission of Visible Light

The four solutions appear different colors under the normal classroom lights. The tonic water is clear, the fluorescein solution is yellow-green, the eosin Y solution is yellow-orange, and the rhodamine B solution is pinkish-red. Why? They are each composed of different molecules—molecules that absorb different wavelengths of light. In general, a green solution looks green to the human eye because it is transmitting green light. When white light is shined through this solution, the molecules in the solution absorb some of the wavelengths of the light and transmit others. All non-green wavelengths of light will be absorbed by the green solution to some extent, although red light will be absorbed the most. The red photons hit the solution and are absorbed by the molecules in the solution. They do not make it through the solution, and hence, we do not see a red color from this solution. In contrast, green photons are not absorbed by the molecules in the green solution. So, they pass right through the solution, and we see a green color.

How do we know that the green solution absorbs the red wavelengths of light? Red and green are complementary colors—they are across from each other on the color wheel.

In general, colors opposite each other on the color wheel are complementary colors. For example, by looking at the wheel, it can be seen that violet and yellow are complementary colors. Therefore, it can be assumed that a yellow solution absorbs violet light and transmits yellow light. The color wheel and the idea of complementary colors can be used as a first estimation of the wavelengths that are absorbed by a substance based on its color.

The following table lists the wavelengths associated with each of the colors in the visible spectrum and their complements. The representative wavelength can be used as a benchmark for each color. For example, instead of referring to green as light in the wavelength range 500–560 nm, one could simply say that green light is 520 nm.
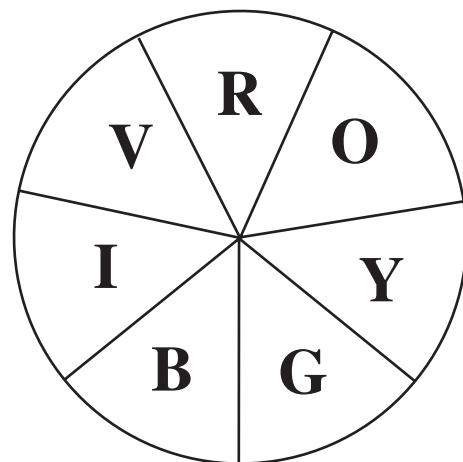
Representative Wavelength, nm	Wavelength Region, nm	Color	Complementary Color
410	400–425	Violet	Yellow-green
470	425–480	Blue	Yellow
490	480–500	Blue-green	Red
520	500–560	Green	Reddish-violet
565	560–580	Yellow-green	Violet
580	580–585	Yellow	Blue
600	585–650	Orange	Green-blue
650	650–700	Red	Blue-green

Fluorescence

Luminescence is the emission of radiation (light) by a substance as a result of absorption of energy from photons, charged particles, or chemical change. It is a general term that includes fluorescence, phosphorescence, and chemiluminescence, to name just a few special types. Fluorescence is different from other types of luminescence in that is it restricted to phenomena in which the time interval between absorption and emission of energy is extremely short. Therefore, fluorescence only occurs in the pres-

ence of the exciting source. This is different from phosphorescence, which continues after the exciting source has been removed. In this demonstration, the exciting source is the UV black light.

In fluorescence, when a light source is shined on a material, a photon is absorbed. The energy from the photon is transferred to an electron that makes a transition to an excited electronic state. From this excited electronic state, the electron naturally wants to relax back down to the ground state. When it relaxes back down to the ground state, it emits a photon (symbolized by the squiggly arrow in the diagram below). This relaxation may occur in a single step or in a series of steps. If it occurs in a single step, the emitted photon will be the same wavelength as the exciting photon. If the relaxation occurs in a series of steps emitting a photon along the way, the emitted photon will have a greater wavelength (lower energy) than the exciting photon.

Energy Level Diagram
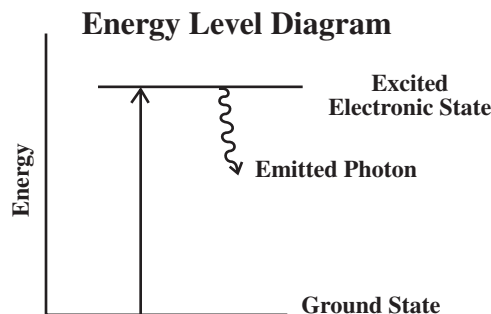
If the emitted photon's wavelength is in the visible portion of the spectrum, we observe a colorful, glowing effect. Emission of this form is termed fluorescence. This process is practically instantaneous so the fluorescence is observed as soon as the exciting source is present, and it disappears as soon as the exciting source is removed. The fluorescent glow is brighter than the color of the solution seen under normal fluorescent lights because light is being emitted from the solution, not just transmitted through it.

Absorption Curves and Color

Information about the absorption and emission curves of each of the solutions in this demonstration can be inferred from the observations made during the demonstration—that is, what wavelengths of light these solutions absorb and emit.

First consider the fluorescein, eosin Y, and rhodamine B solutions

- These solutions appear colored to the human eye under the normal classroom lights. Recall that normal classroom lights give off white light which is composed of all the visible wavelengths of light. Therefore, these solutions must absorb some wavelengths of visible light while transmitting others—the color of the solution is the transmitted color in each case. Each of these colored solutions has an absorption (and transmission) peak in the visible region of the electromagnetic spectrum (400–700 nm).

- When the normal classroom lights are turned off and the black light is shined on the solutions, they fluoresce. Under these conditions, the solutions are not being hit with visible light, but instead are being hit with UVA light (320–400 nm). In each case, when a molecule in the solution is hit with ultraviolet photons, the molecule absorbs an ultraviolet photon and promotes an electron up to an excited state. This electron then relaxes back down to the ground state in a series of steps emitting a visible photon along the way. It is evident that the photon is in the visible region of the spectrum because the fluorescence can be seen with the human eye. Therefore, the molecules in each of the solutions must have an absorption peak in the UVA portion of the electromagnetic spectrum with a corresponding emission peak in the visible portion of the spectrum.

  Clearly, each of these solutions has two absorption peaks—one in the visible and another in the UVA portion of the spectrum. If the transmitted wavelength of visible light is not the same wavelength as the emitted photon during fluorescence, the solution will appear to be two different colors under the two different light sources.

Now consider the tonic water solution. Tonic water appears clear to the human eye under the normal classroom lights. Therefore, it must not absorb any wavelengths of visible light. Consequently, in contrast to the three solutions discussed above, it does not have an absorption peak in the visible region of the spectrum. But, under the UVA black light, it is blue! When hit with ultraviolet light, one of the ingredients in tonic water, quinine, absorbs an ultraviolet photon and emits a visible photon in return. The human eye can see this visible photon, and therefore this solution appears to be colored when viewed under the black light.

It is evident from these examples that color is not an inherent quality of a substance, but instead, a result of the interaction of light with matter. If the wavelength of the light changes, the interaction, and hence the resulting color, may also change.

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
Content Standards: Grades 5–8
  Content Standard B: Physical Science, properties and changes of properties in matter, transfer of energy
Content Standards: Grades 9–12
  Content Standard B: Physical Science, structure and properties of matter, interactions of energy and matter

The *Fluorescent Dye Kit—Chemical Demonstration Kit* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP4848	Fluorescent Dye Kit—Chemical Demonstration Kit
AP9030	Ultraviolet Lamp, 18″

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

IN4848