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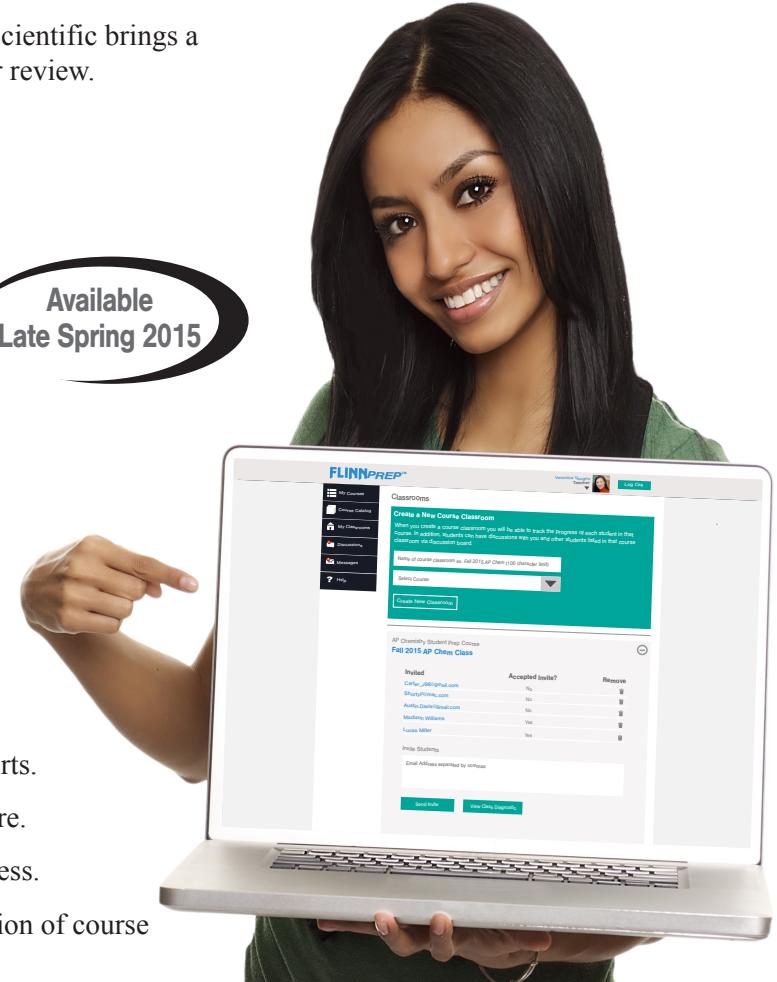
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Teachers may purchase a one-year licence for students in their class by ordering registration key codes through Flinn. Use a purchase order or credit card and specify Catalog# EL1000. A registration key code will be provided for each license purchased. One license/key code must be purchased per student. Key codes will be sent to the ordering teacher.

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One Year Access

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  - Nomenclature
  - Bonding Basics
  - Molecular Structure and Properties
  - Chemical Reactions
  - Gases
  - Stoichiometry
  - Solutions
  - Acids and Bases
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#### Print GHS-Compliant Chemical Labels

GHS pictograms and hazard statements can also be printed on chemical labels through the Online Chemventory™ program.

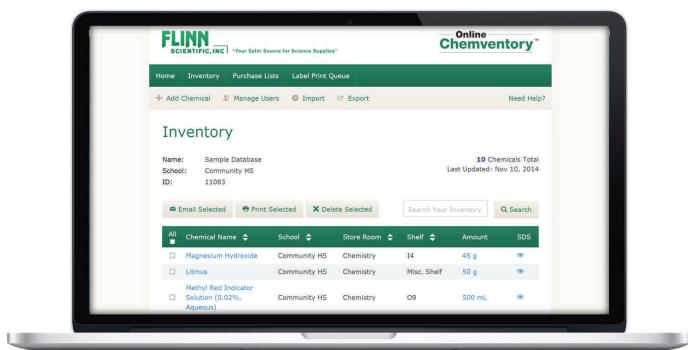
### 5-Year License only \$189.95

Track your entire chemical inventory for less than \$4 a month! For \$189.95 you get 5 years of access to build, view, and update your chemical inventory database. You may order the program online using Flinn Catalog No. SE3000 and we will e-mail your unique Registration Key Code. You may also order SE3000 by mail, fax, or phone and we will contact you with your unique registration code.



#### Flinn Online Chemventory™ enables you to:

- Print, Email, and Export Your Inventory List
- Import Data from Previous Chemventory™ Versions
- Print GHS-Compliant Chemical Labels
- Create a Purchase List
- Easily Add All Your Flinn Chemicals



For more information go to [flinnsci.com](http://flinnsci.com)

Flinn Online Chemventory™ allows you to access your inventory information from multiple locations on multiple devices. Manage your chemical inventory safely and conveniently with Flinn's affordable Online Chemventory™ program!

Catalog No.	Description	Price/Database License	Price/Database License 5 or More
SE3000*	Flinn Online Chemventory™ Program	\$189.95	\$159.95

\*A minimum of one 5-year license must be purchased per participating school.

# **FLINN SCIENTIFIC**

# **Morning of Chemistry 2015**

## **The Best of Chem West**

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# 2015 Morning of Chemistry Presenters



## Alex Aschoff

Alexander Aschoff has been teaching chemistry and physics at Proviso East High School since 2004. He learned to teach from two great chemistry minds, Dr. Willy Hunter at Illinois State University and Glen Lid at Proviso East High School. During his teaching career Alex has been awarded Outstanding Science Teacher at Proviso East, teacher of the month/year, recognized by the Board of Education for revitalizing the school golf program, and has been awarded a teaching scholarship award from the Common Roots Initiative. Alex has also been asked to serve on the National Honors Society Committee, and coached both golf and wrestling for over 6 years.



## Sue Bober

Sue started teaching in 1990 (23 years including “mom years”). She teaches chemistry and physical science at Schaumburg High School. Sue is on the Board of Directors of the National Mole Day Foundation and received the Davidson Award from CICI in 2011.



## Bill Grosser

Bill Grosser has taught chemistry in the Chicago suburbs for the past 25 years. He currently teaches two different levels of chemistry at Oak Park and River Forest High School. During the summer Bill serves as the coordinator of the Golden Apple STEM Institute helping bring inquiry and the Next Generation Science Standards into K-12 classrooms across Chicago. Bill has been named to the National Honor Roll of Teachers by the Association of Science Centers, is a National Board Certified teacher, has been awarded with a RadioShack National Technology Award and is a Golden Apple fellow.



## Mike Heinz

Mike has been teaching in the suburbs of Chicago for 23 years. He is currently serving in his 4th year as science department chairman at Downers Grove North High School. Mike has a Bachelor's of Biology from Quincy University and Masters in the Art of Teaching from Lewis University. He represents DuPage County on the Board of Directors for the Northern Illinois Science Educators. In the past Mike has worked with Flinn Scientific as a Finn Foundation presenting teacher from 2004-2007, and has also been featured in Flinn's Best Practices for Teaching Chemistry video series as well as presented at previous Morning of Chemistry shows. Mike has won the 2005 Chemical Industry Council of Illinois' Davidson Award, the 2009 American Chemical Society Great Lakes Region Award for Excellence in High School Teaching and was a 2010 Golden Apple Finalist.

# 2015 Morning of Chemistry Presenters



## Tanya Katovich

Tanya Katovich has two passions: chemistry and teaching. Over the past twenty years, she has taught high school chemistry in the northwest suburbs of Chicago and enriched her knowledge by participating in a variety of partnerships offered at the university level including at Duke and Northwestern. She has worked as Teacher Fellow for Northwestern, co-taught workshops for high school teachers, and helps coordinate an annual GEMS conference (STEM conference for 5th/6th grade girls). She currently teaches in a 1:1 classroom and incorporates flipped learning. She sits on the Board of Directors of Northern Illinois Science Educators (NISE).



## Kevin Kopack

Kevin Kopack, National Board Certified Teacher, is the chair of sciences and teaches Honors chemistry and Advanced Placement chemistry at Lane Tech College Prep High School in Chicago. His engaging and exciting approach to teaching chemistry has developed over the past 15 years. He participates in local and national chemistry conferences. Kevin also sponsors his school's chemistry club where he and his students are involved in such events as American Chemistry Society (ACS) Day and Science Saturdays to introduce students to the amazing world of chemistry.



## Karl Kraddock

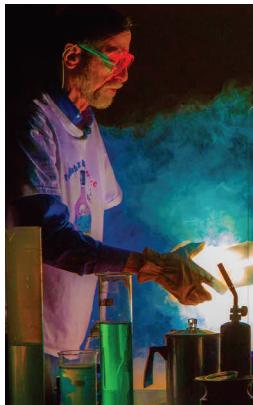
Karl Craddock has been teaching Chemistry, Physical Science and AP Chemistry at William Fremd High School in Palatine, IL for the past 15 years. He earned his B.S. in Chemistry from Illinois State University and his M.A. in Educational Leadership from Roosevelt University. He is currently the President of the ChemWest Teacher Network and on the Board of Directors for Fermilab Friends for Science Education. He has been awarded his National Board Certification. Karl earned the Davidson Award for Illinois Chemistry Teacher of the Year from the Chemistry Industry Council of Illinois.



## Glen Lid

Currently retired. Taught 35 years at Proviso East High School: Regular, Honors, Biology, Physical Science, General Science, and AP Chemistry. Head baseball coach, assistant wrestling and golf coach. Currently adjunct professor at Elmhurst College teaching chemistry for nurses, also still coaches wrestling and golf at Proviso East. Glen is the two time runner-up for the presidential award in science and math, chemical industry counsel of Illinois Davidson award winner, Golden Apple teacher of distinction, High School Teacher of the year: Disney Hand Teacher Awards, National Teachers Hall of Fame, and Illinois Baseball Coaches Hall of Fame.

# 2015 Morning of Chemistry Presenters



## Lee Marek

Lee Marek taught chemistry and AP chemistry at Naperville North High School in Naperville, IL. He taught 12 years at University of Illinois at Chicago. Lee received his B.S. in Chemical Engineering from the University of Illinois and M.S. degrees in both Physics and Chemistry from Roosevelt University. He has more than 40 years of teaching experience. Lee is one of the authors on Flinn's 23 volume set of ChemTopic Lab Manuals and is a lead presenter in Flinn's Best Practices for Teaching Chemistry Video Series. Together with two other ChemWest members, Lee also founded Weird Science and has presented 600 demonstration and teaching workshops for more than 300,000 students and teachers across the country. Lee or his students have performed science demonstrations on the David Letterman Show 35 times. Lee has received the Presidential Award, the James Bryant Conant Award in High School Teaching and the Helen Free Award for Public Outreach from the American Chemical Society. Lee's full biography can be found on his website: <http://www2.chem.uic.edu/marek/>



## Mike Palazzolo

Mike Palazzolo is a 4<sup>th</sup> year Chemistry Teacher at Schaumburg High School. He participates in a variety of community programs that enriches teachers knowledge of Chemistry and science. Mike regularly attends ChemWest, a local group for the sharing of best practices in chemical education and he volunteers at a D211 annual STEM conference called GEMS. Mike is currently pursuing his masters degree in chemical education at Northeastern Illinois University in Chicago.



## Jill Serling

Jill Serling teaches chemistry and physics at Glenbrook South High School in Glenview, Illinois. During her seven years of teaching, she has incorporated many demonstrations that she originally learned from her peers at ChemWest meetings. Currently, she is pursuing a master's degree in chemistry at Northeastern Illinois University.



## JulieAnn Villa

JulieAnn Villa has been teaching science for 17 years. She currently teaches at chemistry and science research Niles West High School in Skokie, IL. JulieAnn began her teaching career in middle school science overseas and has taught in 3 countries. She holds a bachelor's degree from M.I.T. and a master's degree from DePaul University. Her greatest passion is connecting students with their own research ideas about science and molecular modeling based chemistry. She was honored as a 2007 Golden Apple Award Finalist. She serves as the co-director for IJAS Region 6 Science Fair and is active in AACT and NSTA.

## Leaf Blower Bernoulli

### Introduction

What causes the lift that takes a jet airplane off the runway? What makes a baseball curve? An 18th century Swiss scientist and mathematician Daniel Bernoulli (1700–1782) gave some insight into the answer. Bernoulli proposed that the faster a fluid such as air or water moves, the less pressure it exerts. Let's see!

### Science Concepts

- Pressure
- Bernoulli's principle

### Materials

Bernoulli Demonstrator

### Safety Precautions

Follow standard laboratory precautions. To avoid contamination and spread of germs, do not allow students to share demonstrators. The devices can be sanitized after use with a Lysol® solution (1½ oz to 1 gal water) or bleach solution. Dip the demonstrator in the sanitizing solution, rinse thoroughly with water, and allow to air dry.

### Procedure

1. With the ball in the funnel cup of the Bernoulli Demonstrator, blow gently into the mouthpiece end of the demonstrator.
2. Observe what happens to the ball—see Bernoulli's principle in action! With practice, the ball can be made to levitate above the funnel cup.

### Disposal

The Bernoulli Demonstrator may be sanitized and reused.

### Tips

- Care must be taken when blowing into the Bernoulli Demonstrator because too much force will cause the ball to fly out of the column of air. For Bernoulli's principle to work for this demonstration, an air column and streamline of air must be created around the ball, which can only be created with a slow, steady flow of air. If the air flow is too fast, the friction from the ball's surface will create too much turbulence at the base, and a streamline of air will not travel around the ball. The Venturi effect (see *Discussion*) will not occur and this will simply turn into a Newton's third law demonstration of action-reaction. The air molecules will strike the base of the ball with a force much greater than the force of gravity pulling down, and the ball will fly out of the Bernoulli Demonstrator.
- Make a giant Bernoulli Demonstrator using a leaf blower and a beach ball! Place the beach ball in a vertical column of air produced by the leaf blower approximately a foot from the end of the nozzle. Make sure the ball is centered in the air column. Then carefully let go of the ball. It should float and possibly rise up and remain suspended in the air column, just like the Bernoulli Demonstrator. This may take a steady hand and a few attempts. Be patient.

A hair dryer will also work if a leaf blower is not available or is not convenient to use. A blown-up balloon with a counter-weight works well. Tie a piece of string around a washer and then tie the free end of the string to the tied end of the blown-up balloon (see Figure 1). Place the balloon into the air current of the hair dryer. If the balloon is not heavy enough, add another washer to the string until the balloon balances in the air current of the hair dryer.

## Discussion

Bernoulli's principle states when the speed of a fluid increases, the pressure on the surrounding surface drops. This is also known as the Venturi effect, and explains why planes fly, why baseballs curve, and why a race car needs a spoiler.

*What causes the ball to levitate and remain suspended in the air stream above the funnel cup?* The reason for the ball's levitation can be explained by Bernoulli's principle, as well as Newton's laws of motion. When the ball is placed into the column of rapidly moving air, the friction from the surface of the ball creates turbulence and slows the air speed down at the surface. This creates high pressure on the bottom of the ball. The ball's surface also deflects the air column and the air is channeled around the ball in what is known as a streamline. However, as the streamline of air reaches the top of the ball, it wants to continue traveling in a straight line and does not want to follow the spherical surface of the ball. Therefore, small, turbulent air currents are produced at the top surface of the ball, which do not produce as much pressure as there is on the bottom, creating a pocket of low pressure (see Figure 2). Since there is more pressure (force) below the ball than above, the ball will rise up until the net upward force and the force of gravity pulling down are balanced. The ball will remain stable so long as the air velocity remains stable.

Bernoulli's principle is also responsible for keeping the ball in the column of air. For example, if the ball were to drift slightly to the right, the right side of the ball will move out of the air column, creating more turbulence that will slow the air down on that side. Meanwhile, the left side of the ball has moved into the air column and has become more streamlined with the air column, creating less turbulence, and the air will travel over the surface more quickly. So this situation creates more pressure on the right side (slower air speed) and less pressure on the left side (faster air speed), and the ball is forced back into the column of air until the pressure on both sides becomes equal. This is why the ball will remain in the center of the air column.

## 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 5–8***

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, understanding of motions and forces

### ***Content Standards: Grades 9–12***

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, motions and forces

## Acknowledgments

Special thanks to members of the Weird Science group—Lee Marek, Naperville North High School, Naperville, IL and Bob Lewis, Downers Grove High School, Downers Grove, IL—for bringing this demonstration to our attention.

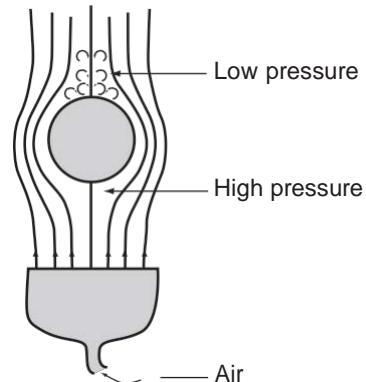
## Reference

Tipler, Paul A. *Physics For Scientists and Engineers*, Third Edition, Volume 1; Worth Publishers: New York, 1990; pp 346–351.

## The Bernoulli Demonstrator is available from Flinn Scientific, Inc.

Catalog No.	Description
AP5933	Bernoulli Demonstrator

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



**Figure 2**

## Iodine Clock Reaction

### Introduction

Use this dramatic iodine clock reaction to demonstrate the effect of concentration, temperature, and a catalyst on the rate of a chemical reaction.

### Concepts

- Kinetics/Catalysts
- Clock Reactions

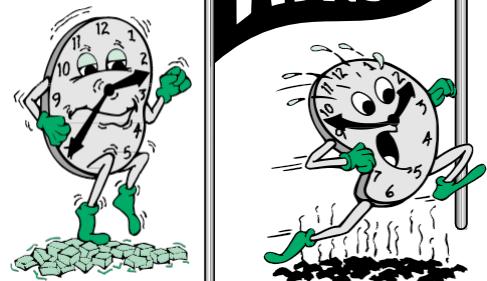
### Materials (for each demonstration)

Potassium iodate solution, 0.20 M,  $\text{KIO}_3$ , 325 mL\*  
Sodium meta-bisulfite solution, 0.20 M,  $\text{Na}_2\text{S}_2\text{O}_5$ , 60 mL\*<sup>†</sup>  
Starch solution, 2%, 180 mL\*<sup>†</sup>  
Sulfuric acid solution, 0.1 M,  $\text{H}_2\text{SO}_4$ , 10 mL\*  
Distilled or deionized water, 1105 mL  
Graduated cylinders, 10-mL, 2  
Graduated cylinders, 50-mL, 2  
Graduated cylinders, 100-mL, 2

\*Materials furnished in kit.

<sup>†</sup>This solution must be prepared fresh from solid material furnished in the kit.

Beakers, 250-mL, 6  
Beakers, 400-mL, 6  
Bucket or utility pan (for ice bath)  
Timer or stopwatch  
Thermometer  
Hot Plate  
Ice



### Safety Precautions

Potassium iodate is an oxidizer. It is moderately toxic by ingestion and a body tissue irritant. Sodium meta-bisulfite is a skin and tissue irritant. Sulfuric acid solution is corrosive to eyes, skin and other tissues. Wear chemical splash goggles, a chemical-resistant apron and chemical-resistant gloves. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

### Preparation

#### Preparation of "A" Solutions:

1. Label six 400-mL beakers as 1A, 2A, 3A, 4A, 5A, and 6A.
2. Pour 50 mL of the 0.2 M potassium iodate solution plus 150 mL of distilled or deionized water into beakers 1A, 4A, and 5A.
3. Pour 100 mL of the 0.2 M potassium iodate solution plus 100 mL of distilled or deionized water into beaker 2A.
4. Pour 25 mL of the 0.2 M potassium iodate solution plus 175 mL of distilled or deionized water into beaker 3A.
5. Place beaker 4A on a hot plate and warm it to about 45 °C. Place beaker 5A into an ice bath and cool it to about 10 °C.
6. Pour 50 mL of the 0.2 M potassium iodate solution, 140 mL of distilled or deionized water, and 10 mL of the 0.1 M sulfuric acid solution into beaker 6A.

7. Check to make sure each beaker contains 200 mL of solution. Use the table below as a guide.

	<b>Beaker 1A</b>	<b>Beaker 2A</b>	<b>Beaker 3A</b>	<b>Beaker 4A</b>	<b>Beaker 5A</b>	<b>Beaker 6A</b>
Potassium Iodate Solution, 0.2 M	50 mL	100 mL	25 mL	50 mL	50 mL	50 mL
Distilled or Deionized Water	150 mL	100 mL	175 mL	150 mL	150 mL	140 mL
Sulfuric Acid Solution, 0.1 M	—	—	—	—	—	10 mL
[KIO <sub>3</sub> ] before mixing with Solution B	0.05 M	0.1 M	0.025 M	0.05 M	0.05 M	0.05 M
[KIO <sub>3</sub> ] after mixing with Solution B	0.04 M	0.07 M	0.02 M	0.04 M	0.04 M	0.04 M

#### Preparation of “B” Solutions:

1. Add 420 mL of distilled or deionized water to the 500-mL bottle that contains (16 g) solid sodium meta-bisulfite. Put on the cap and shake well until the solid is dissolved. This is now a 0.20 M sodium meta-bisulfite solution. Prepare this solution fresh.
2. Prepare 1300 mL of 2% starch solution by making a smooth paste of 26 g soluble starch and 130 mL distilled or deionized water. Pour the paste into 1170 mL of boiling water while stirring. Cool to room temperature before using. Starch solution has a poor shelf life and will form mold if kept for too long. Prepare this solution fresh.
3. Label six 250-mL beakers as 1B, 2B, 3B, 4B, 5B, and 6B.
4. Pour 10 mL of the 0.20 M sodium meta-bisulfite solution, 30 mL of the starch solution, and 40 mL of distilled or deionized water into each of these beakers. Stir each solution.

	<b>Beaker 1A</b>	<b>Beaker 2A</b>	<b>Beaker 3A</b>	<b>Beaker 4A</b>	<b>Beaker 5A</b>	<b>Beaker 6A</b>
Sodium meta-Bisulfite Solution, 0.20 M	10 mL					
Starch Solution	30 mL					
Distilled or Deionized Water	40 mL					
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ] before mixing with Solution A	0.025 M					
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ] after mixing with Solution A	0.007 M					

## Procedure

1. Make a data table on the chalkboard or on an overhead like the following one.

	1 (Control)	2	3	4	5	6
[KIO <sub>3</sub> ]	0.04 M	0.07 M	0.02 M	0.04 M	0.04 M	0.04 M
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ]	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M
Temperature	Room Temp	Room Temp	Room Temp	Warm	Cool	Room Temp
Catalyst Added	No	No	No	No	No	Yes
Time Until the Appearance of the Blue Color (sec)						

### *Control Reaction:*

2. Place the 400-mL beakers containing the "A" Solutions on the demonstration table in order.
3. Pour Solution 1B into Solution 1A. Carefully time the reaction with a stopwatch or timer. Measure the time from when the two solutions are mixed until the appearance of the blue color. Record the time in column 1 of the data table. The rate of this reaction will be the rate against which the others are compared. This is the control.

### *Effect of Concentration on the Rate:*

4. Pour Solution 2B into Solution 2A. Measure the time until the appearance of the blue color. Solution 2A is twice as concentrated as Solution 1A, so this rate should be about twice that of the rate in the control reaction (the time will be half that of the control). Record the time in column 2 of the data table.
5. Pour Solution 3B into Solution 3A. Measure the time until the appearance of the blue color. Solution 3A is half as concentrated as Solution 1A, so this rate should be about half that of the rate in the control reaction (the time should be double that of the control). Record the time in column 3 of the data table.

### *Effect of Temperature on the Rate:*

6. Carefully remove Solution 4A from the hot plate. Measure the temperature of solution 4A. Immediately pour Solution 4B into Solution 4A. Measure the time until the appearance of the blue color. Because Solution 4A was warmed, the rate of reaction should be faster than that of the control reaction (the time should be less than that of the control). Record the temperature of Solution 4A and the time in column 4 of the data table.
7. Remove Solution 5A from the ice bath. Measure the temperature of Solution 5A. Immediately pour Solution 5B into Solution 5A. Measure the time until the appearance of the blue color. Because Solution 5A was cooled, the rate of reaction should be slower than that of the control reaction (the time should be more than that of the control). Record the temperature of Solution 5A and the time in column 5 of the data table.

### *Effect of a Catalyst on the Rate:*

8. Pour Solution 6B into Solution 6A. Measure the time until the appearance of the blue color. The sulfuric acid solution added to Solution 6B serves as a catalyst in the reaction, so the rate of this reaction should be faster than that of the control reaction (the time should be less than that of the control). Record the time in column 6 of the data table.

## Results

Beaker	1	2	3	4	5	6
Time Until the Appearance of the Blue Color (sec)	6 sec	3 sec	12 sec	4 sec	8 sec	2 sec

## 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 final solutions may be reduced with sodium thiosulfate solution according to Flinn Suggested Disposal Method #12a. Add just enough reducing agent to decolorize the blue color of the starch–iodine complex.

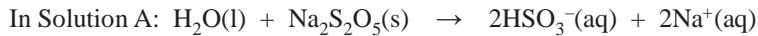
## Tips

- This demonstration may be more dramatic if six student volunteers are recruited to help. They can each be given prepared solutions and asked to mix the solutions simultaneously in front of the class. This twist requires the use of six stopwatches or timers.
- All concentrations/dilutions in the *Preparation* section were calculated using the equation  $M_1V_1 = M_2V_2$ . Concentrations before and after mixing are different because combining the solutions dilutes each reactant.

## Discussion

In this reaction, potassium iodate and sodium meta-bisulfite react to form iodine. The starch solution serves as an indicator of the end of the reaction, forming a dark-blue colored starch–iodine complex in the presence of iodine. The chemical pathway for the formation of iodine is complicated and not completely understood, but the following mechanism serves as an outline.

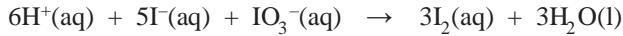
Step 1: Solution A and Solution B contribute hydrogen sulfite ions,  $\text{HSO}_3^-$ (aq), and iodate ions,  $\text{IO}_3^-$ (aq), to the solution.



Step 2: The iodate ions react with the hydrogen sulfite ions to produce iodide ions,  $\text{I}^-$ (aq).



Step 3: In the presence of hydrogen ions,  $\text{H}^+$ , the iodide ions react with excess iodate ions to produce iodine,  $\text{I}_2$ (aq).



Step 4: Before the iodine can react with the starch to produce a dark-blue colored complex, it immediately reacts with any hydrogen sulfite ions still present to form iodide ions.



Step 5: Once all of the hydrogen sulfite ions have reacted, the iodine is then free to react with the starch to form the familiar dark-blue colored complex.

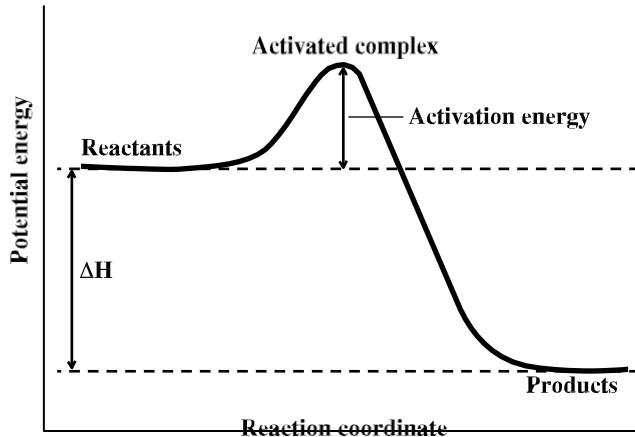


The dark-blue color of the complex is due to the presence of the pentaiodide anion,  $\text{I}_5^-(\text{aq})$ . By itself, the pentaiodide ion is unstable; however, it is stabilized by forming a complex with the starch.

The appearance of the dark-blue color in solution indicates that all of the reactants have been used up and the reaction has gone to completion. Therefore, the rate of reaction can be measured by recording the time to the appearance of the dark-blue color.

In general, the effect of concentration, temperature, and a catalyst on the reaction rate can be understood by looking at the energy profile for a given reaction.

In an energy profile diagram, the left side of the diagram represents reactants, while the right side represents products. In the diagram above, the products are lower in energy than the reactants. In terms of thermodynamics, this reaction is exothermic and should occur spontaneously. However, not all collisions between reactants will produce products. The collision energy for a particular collision must exceed a critical energy for products to be formed. This critical energy is called the *activation energy* and is represented by the hump in the energy profile diagram.



Why must reactant molecules overcome this activation energy, or get over the hump, to reach products? As the reactant molecules approach each other, their atoms interact causing distortion in the bonds of both molecules. This distortion reaches a maximum as the reactants form an *activated complex*, or transition state. The activated complex is a hybrid species formed as the reactant molecules come together and trade atoms to become products. Only those colliding molecules that have enough kinetic energy to reach this distorted intermediate will produce products. As is evident from the energy profile diagram above, the potential energy of this distorted transition state determines the activation energy, or height of the barrier, for a particular reaction. If the barrier is low, almost all colliding molecules will have sufficient energy to reach and overcome the barrier. These reactions will occur spontaneously. If the barrier is high, only a small percentage of collisions will occur with sufficient energy to reach and overcome the barrier and go on to form products. These reactions occur much more slowly than those with a low barrier. In general, as the height of the barrier increases, the rate of the reaction decreases. Therefore, the rate of a reaction depends on the height of the barrier, or the activation energy.

The above description of the energy profile assumes the reaction occurs in a single step. This theory can be applied to multi-step mechanisms, such as the one in this iodine clock reaction, by assuming that one of the steps in the mechanism is much slower than the other steps. This step then determines the rate of the reaction and is called the *rate-determining step*. It is generally a good approximation to say that the energy profile of a reaction describes the energy profile of the rate-determining step.

To increase the rate of a reaction, one of two things must occur: (1) more molecules with sufficient kinetic energy to overcome the barrier must be involved in the reaction to produce a higher number of successful collisions, or (2) the activation energy must be decreased.

One way to obtain a higher number of successful collisions is to increase the concentration of reactant molecules. In beaker #1, the ratio of potassium iodate molecules to sodium meta-bisulfite molecules is 5:1. This ratio climbs to 10:1 in beaker #2, while it drops to 2.5:1 in beaker #3. The number of collisions is proportional to the concentration so beaker #2 will have twice as many collisions as beaker #1, while beaker #3 will have half as many collisions as beaker #1. In each case, the same fraction of these collisions will possess sufficient energy to overcome the barrier as before. So, since twice as many collisions are occurring in beaker #2, the rate at which a given concentration of B is converted to products will double. Because half as many collisions are occurring in beaker #3, the rate at which a given concentration of B is converted to products is cut in half.

Another way to obtain more molecules with sufficient energy to overcome the barrier is to increase the temperature. The strong temperature dependence of reaction rates can be understood by looking at the relationship between temperature and energy. The average kinetic energy (KE) of a sample is directly proportional to the temperature (T) of the sample according to the following equation:

$$KE = \frac{3}{2} k T$$

where  $k$  is Boltzman's constant. As the temperature is increased, the average kinetic energy of the sample is increased providing a sample with more molecules that possess enough kinetic energy to reach and overcome the barrier.

To lower the activation energy, a catalyst may be added to the reaction mixture. A catalyst is a substance that, when added to the reaction mixture, participates in the reaction and speeds it up, but is not itself consumed in the reaction. In general, a catalyst provides a modified or new mechanism for the reaction that is faster than the original mechanism. The rate of the catalyzed reaction is faster because the activated complex in the catalyzed mechanism is of lower energy than the activated complex in the original mechanism. Hence the barrier to products is lower in the catalyzed reaction. A greater percentage of reactant molecules will possess the needed energy to successfully collide and overcome the barrier. Therefore, the rate of the reaction is increased.

Although thermodynamics determines whether reactants or products are energetically favored, it is kinetics that determines how fast even an exothermic reaction will occur. The speed at which different reactions occur varies widely, ranging from instantaneous to so slow that it appears as if no reaction occurs at all. The use of catalysts in industrial applications can turn costly (slow) reactions into more efficient (fast) processes. Every year more than a trillion dollars worth of goods are manufactured with the help of catalysts. Without them, fertilizers, pharmaceuticals, plastics, fuels, solvents, and surfactants would be in short supply.

## Connecting to the National Standards

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

### ***Unifying Concepts and Process: Grades K–12***

Systems, order, and organization

Evidence, models, and exploration

Constancy, change, and measurement

### ***Content Standards: Grades 5–8***

Content Standard A: Science as Inquiry

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

### ***Content Standards: Grades 9–12***

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, chemical reactions, conservation of energy and increase in disorder

## Answers to Worksheet

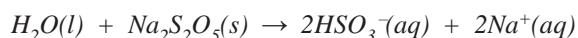
### Sample Results Table

	1 (Control)	2	3	4	5	6
<b>Contents of Beaker A</b>	50 mL 0.2 M $\text{KIO}_3$ , 150 mL water	100 mL $\text{KIO}_3$ , 100 mL water	25 mL $\text{KIO}_3$ , 175 mL water	50 mL $\text{KIO}_3$ , 150 mL water	50 mL $\text{KIO}_3$ , 150 mL water	50 mL $\text{KIO}_3$ , 140 mL water, 10 mL 0.1 M sulfuric acid
<b>Contents of Beaker B</b>	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water
[ $\text{KIO}_3$ ] after mixing solutions	0.04 M	0.07 M	0.02 M	0.04 M	0.04 M	0.04 M
[ $\text{Na}_2\text{S}_2\text{O}_5$ ] after mixing solutions	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M
Temperature	Room temperature	Room temperature	Room temperature	Warm	Cool	Room temperature
Catalyst present?	No	No	No	No	No	Yes
Reaction Time	6 sec	3 sec	12 sec	4 sec	8 sec	2 sec

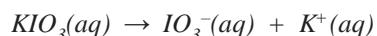
## Discussion Questions

1. Write the chemical equation for each of the steps in this reaction mechanism.

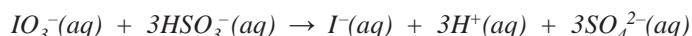
- a. Sodium meta-bisulfite produces hydrogen sulfite ions in water



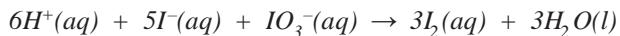
- b. Potassium iodate decomposes



- c. Iodate ions react with hydrogen sulfite ions



d. Iodide ions react with iodate ions in the presence of water



In order for a reaction to occur and products to be formed, reactant molecules need to reach a transition state, or activated complex. Only molecules with enough kinetic energy, or activation energy, to reach this state can produce products. The reaction rate depends on the activation energy necessary for molecules to form an activated complex. There are two ways to increase the rate of a reaction: more molecules with sufficient kinetic energy must be present, or the activation energy must be decreased.

2. How does the concentration of reactant molecules affect the reaction rate? Explain.

*The higher the concentration of reactant molecules, the faster the rate at which the reaction will occur. This is because an increase in concentration results in more successful collisions between different reactant molecules. Therefore, the likelihood of these collisions having enough kinetic energy to form an activated complex is also increased.*

3. How does temperature affect the reaction rate? Explain.

*Kinetic energy is directly proportion to temperature. Thus, when temperature is increased, the average kinetic energy of the sample also increases. This results in the presence of more molecules with enough kinetic energy to reach the transition state.*

4. What is a catalyst? What was the catalyst used in this demonstration?

*A catalyst is a substance that speeds up a reaction by lowering the activation energy, but is not consumed in the reaction. The catalyst in this demonstration was sulfuric acid.*

## Reference

Cardillo, C.; *Micro Action Chemistry*; Flinn Scientific: Batavia, IL, 1998; pp 85–87.

## Materials for the *Iodine Clock Reaction* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP4601	Iodine Clock Reaction—Effect of Concentration, Temperature, and a Catalyst on Reaction Rate
GP1047	Beaker, Pyrex®, 250-mL
GP1048	Beaker, Pyrex®, 400-mL
GP2040	Cylinder, Graduated, 10-mL
GP2044	Cylinder, Graduated, 50-mL
GP2046	Cylinder, Graduated, 100-mL
AP1452	Thermometer, Non-Mercury
AP8386	Hot Plate, Cimarec™, 7" x 7"
AP8874	Timer, Large Display
AP9176	Utility Pan, Plastic
W0001	Distilled Water, 1 gallon

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

# Iodine Clock Reaction Demonstration Worksheet

## Results Table

	1 (Control)	2	3	4	5	6
<b>Contents of Beaker A</b>						
<b>Contents of Beaker B</b>						
[KIO <sub>3</sub> ] after mixing solutions						
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ] after mixing solutions						
Temperature						
Catalyst present?						
Reaction Time						

## Discussion Questions

1. Write the chemical equation for each of the steps in this reaction mechanism.
  - a. Sodium meta-bisulfite produces hydrogen sulfite ions in water
  - b. Potassium iodate decomposes
  - c. Iodate ions react with hydrogen sulfite ions
  - d. Iodide ions react with iodate ions in the presence of water

Name: \_\_\_\_\_

In order for a reaction to occur and products to be formed, reactant molecules need to reach a transition state, or activated complex. Only molecules with enough kinetic energy, or activation energy, to reach this state can produce products. The reaction rate depends on the activation energy necessary for molecules to form an activated complex. There are two ways to increase the rate of a reaction: more molecules with sufficient kinetic energy must be present, or the activation energy must be decreased.

2. How does the concentration of reactant molecules affect the reaction rate? Explain.

3. How does temperature affect the reaction rate? Explain.

4. What is a catalyst? What was the catalyst used in this demonstration?

## Zombie Bubbles

### Overview

Two chemicals are added to a fish tank. The fish tank creates a layer of gas that blown bubbles float on.

### Purpose

This demo shows the density of different gases and how less dense materials float on more dense materials. It also shows an acid-base neutralization reaction and it's products.

### Materials

- A small (10 gallon) fish tank
- Baking Soda (sodium bicarbonate)
- Simple bubbles found at any drugstore
- 6M acetic acid (Vinegar works 'ok')
- Candle (optional)
- A ladle (optional)
- Music: "Zombie Nation" by Kernkraft 400 (optional)

### Procedure

Add baking soda to the bottom of the fish tank (amount depends on tank, but enough to fully dust the bottom).

Add approximately 100ml of the 6M acetic acid to the bottom of the tank. The neutralization reaction will bubble and produce carbon dioxide gas.

From approximately one foot away from the tank blow the bubbles so that they land into the tank (this takes some practice, but make sure you don't blow directly into the tank and blow the CO<sub>2</sub> out of the tank)

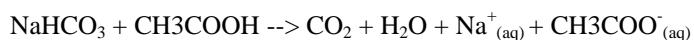
If sufficient CO<sub>2</sub> was created, the bubbles should float in the middle of the fish tank on the layer of CO<sub>2</sub>.

As an extension, you can light a candle and then place it into the tank. The candle will get suffocated and go out. When it goes out the smoke produced will stay in the CO<sub>2</sub> layer showing the fluid nature of the gas.

Or you can place the candle in a beaker and use the ladle to scoop out some CO<sub>2</sub> and "pour out" the candle. (This is tough to master. It is easier to produce some CO<sub>2</sub> in a beaker and then "pour it out" from the beaker directly.)

### Discussion

The neutralization reaction is as follows:



Carbon dioxide gas alone is much more dense (0.001977 g/mL) than the air mixture of N<sub>2</sub>, O<sub>2</sub>, CO<sub>2</sub> ...etc. (0.00128 g/mL) so it stays at the bottom of the tank. The bubbles seem to be more dense than the air but less dense than the CO<sub>2</sub>. The music adds an element of fun to the demonstration and shows the bubbles are "frozen like zombies."

### Safety

Acetic acid is corrosive; avoid contact with skin; wear eye protection. Also be aware of general precautions with the open flame of the candle.

## Smoke Rings

Years ago, toy manufacturers like Wham-O sold air blasters that sent bursts of air sailing across a room to the surprise and delight of any innocent victim. With a little practice, it was quite easy to shoot a cup off of someone's head from 20 feet away.

### Materials

- 5 gallon bucket or a large trash can
- Bungee cord
- Plastic shower curtain or thick plastic sheet
- Knife or keyhole saw

### EXPERIMENT

1. It's easy to make your own air blaster using materials that you can find at the hardware store. Start with a 5 gallon plastic bucket (about \$3 in any hardware store). Carefully cut a 2 to 3 inch hole in the bottom center of the bucket.
2. Stretch a membrane across the top of the bucket. Believe it or not, a piece of clear plastic shower curtain works great. Just stretch a piece of the shower curtain over the top of the bucket and secure it in place using a bungee-type cord.
3. Lightly hit the shower curtain with your hand or the end of a stick. An invisible blast of air shoots out of the hole. Just aim the air cannon at someone or something across the room and send a blast of air their way with a whack of the membrane.

### HOW TO MAKE A SMOKE RING LAUNCHER

How can you make the invisible ring of air visible? With a little smoke, of course. The best smoke rings are made by filling the trash can with theatrical smoke. Smoke machines (foggers) are commonly used in stage productions. Check to see if your theater department has one, or see if you can borrow one from the disc jockey who plays for your school dances. The best smoke rings are also made by *gently* tapping the shower curtain. A hard smack results in a fast blast of air that is difficult to see. The flying vortices are best seen against a dark background with light coming from either side.

### HOW DOES IT WORK?

The proper name for the air cannon device is vortex generator. The "ball" of air that shoots out of the cannon is actually a flat vortex of air, similar to rings of smoke blown by a talented cigar smoker. A vortex is generated because the air exiting the bucket at the center of the hole is traveling faster than the air exiting around the edge of the hole. That swirling or vortex motion can be observed if a little smoke is blown into the bucket just before giving the rubber membrane a gentle push. This activity demonstrates that air occupies space... and the flying smoke rings are an added bonus.

## Monsters of the Midway Orange Blue Reaction

### Introduction

Fascinating reaction has it all—bubbling and effervescence, and reversible orange and blue color changes as a precipitate appears and then disappears. Many concepts come into play in this colorful demonstration of competing redox reactions, catalysis, and transition metal complex ions. A great demo all around, especially if your school colors are orange or gold and blue!

### Concepts

- Redox reactions
- Catalysis
- Complex ions

### Materials

Copper(II) sulfate solution, CuSO<sub>4</sub>, 1 M, 1 mL  
Hydrogen peroxide, H<sub>2</sub>O<sub>2</sub>, 3%, 240 mL  
Potassium sodium tartrate solution, KNaC<sub>4</sub>H<sub>4</sub>O<sub>6</sub>, 1 M, 60 mL  
Beaker, 1-L (a tall form beaker works nicely)

Graduated cylinders, 10-mL and 100-mL  
Hot plate/magnetic stirrer (or hot plate and stirring rod)  
Thermometer, -10 to 100 °C

### Safety Precautions

Dilute (3%) hydrogen peroxide solution is a weak oxidizing agent and a skin and eye irritant. Copper(II) sulfate is a skin and respiratory tract irritant and is toxic by ingestion. 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.

### Procedure

1. Using a 100-mL graduated cylinder, measure 60 mL of potassium sodium tartrate solution into a 1-L beaker.
2. Using a clean 100-mL graduated cylinder, measure 40 mL of 3% hydrogen peroxide. Add the hydrogen peroxide to the beaker while stirring the solution. Continue stirring throughout the rest of the demonstration.
3. Heat the beaker until the solution temperature reaches 50 °C. Turn off the heat. Very little reaction will be apparent at this point.
4. Using a 10-mL graduated cylinder, measure 1 mL of copper(II) sulfate solution and add it to the beaker. Observe the following:
  - a. Light blue color due to the copper(II) tartrate complex ion
  - b. The reaction starting, indicated by bubbling
  - c. The temperature rising to about 80 °C
  - d. Vigorous bubbling as additional gas is formed
  - e. The color suddenly changing to an opaque orange-gold due to the precipitation of copper(I) oxide (Cu<sub>2</sub>O).
5. Add another 40 mL of 3% hydrogen peroxide to the beaker and watch as the orange precipitate dissolves, the blue color returns, and then the reaction suddenly repeats itself (the orange precipitate returns).
6. The sequence of color changes can be repeated about six times by adding more hydrogen peroxide (step 5). The solution must be kept hot, at 70 °C or higher. The reaction can no longer be repeated when the solution becomes too dilute or the tartrate ion is depleted.

### Disposal

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

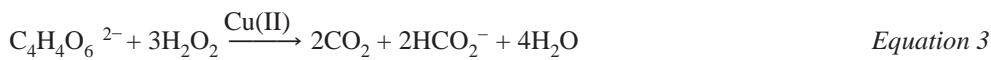
## Tips

- A 600-mL beaker or 500-mL tall form beaker may be used if a 1-L tall-form beaker is not available.
- A good discussion time is while the solution is heating. Heating the solution takes several minutes—this is a good time to initiate a discussion of the relevant chemistry concepts.

## Discussion

The reversible orange and blue reaction demonstrates several fundamental principles in chemistry. It also reveals that “typical” chemistry may be quite unusual! The catalytic decomposition of hydrogen peroxide to produce oxygen gas and water is well known, and many substances, including metals and metal ions, will catalyze the reaction. The decomposition reaction is an example of disproportionation in which hydrogen peroxide is both oxidized, to give oxygen gas, and reduced, to give water. Hydrogen peroxide can act therefore as either a reducing agent or an oxidizing agent, depending on reaction conditions and the substrates involved.

The intense blue color observed when Cu(II) ions are added to the solution containing hydrogen peroxide and tartrate ions signals the formation of copper(II)-tartrate complex ions (Equation 1). As the solution is heated these Cu(II) ions catalyze the decomposition of hydrogen peroxide, resulting in gas bubbles and the liberation of heat. The gas bubbles are a mixture of oxygen and carbon dioxide, suggesting that two competing redox reactions occur simultaneously. In one reaction, hydrogen peroxide is oxidized by copper(II), giving oxygen gas and a bright orange precipitate of copper(I) oxide (Equation 2). In an accompanying reaction, tartrate ions are oxidized by hydrogen peroxide to give carbon dioxide, formate ions, and water (Equation 3). Cu(II) ions catalyze this reaction as well—the oxidation of tartrate by hydrogen peroxide is slow in the absence of a transition metal catalyst. Adding more hydrogen peroxide when these initial reactions have subsided re-oxidizes copper(I) oxide, resulting in the disappearance of the orange precipitate and the return of the blue color of the copper(II) complex ions (Equation 4). The entire cycle then repeats itself until the tartrate ion concentration has been depleted. All of the redox equations given below are written in basic form since it is known that the pH of the reaction mixture increases over the course of the demonstration.



## 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 9–12*

Content Standard A: Science as Inquiry

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

## Acknowledgments

Special thanks to Marie C. Sherman, Ursuline Academy, St. Louis, MO for providing Flinn Scientific with the idea and instructions for this activity.

## Reference

Sherman, Marie C. and Deborah Weil, “A Reversible Blue-and-Gold Reaction.” *J. Chem. Educ.*, **1991**, 68, 1037.

**Chemicals for The Reversible Orange and Blue Reaction are available from Flinn Scientific, Inc.**

Catalog No.	Description
AP8684	The Reversible Orange and Blue Reaction
P0084	Potassium Sodium Tartrate, 100 g
P0085	Potassium Sodium Tartrate, 500 g
C0246	Cupric Sulfate Solution, 1 M, 500 mL
H0009	Hydrogen Peroxide, 3%, 473 mL

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

## Leaky Faucet and the CO<sub>2</sub> Crystal Ball

### Introduction

Four fun demonstrations using solid carbon dioxide, also known as dry ice. In the first demonstration, liquid detergent is added to a graduated cylinder containing some dry ice subliming under water, and foggy carbon dioxide bubbles spill out over the top. Once in contact with the air, however, these bubbles begin to shrink quite noticeably. Beautiful, mystical smoke rings are blown using dry ice and plastic cups in the next demonstration. In the third demonstration, dry ice is added to a bucket half-filled with water and an eerie heavy fog cascades over the rim. A soapy towel is used to create a soap film sheet. This film gradually inflates into a misty, colorful crystal ball that undulates gracefully to the air currents in the room! And finally, foggy CO<sub>2</sub> gas—from dry ice subliming under water—is passed through a faucetlike series of pipes. Dense CO<sub>2</sub> bubbles form, pinch off, and fall rapidly downward, resembling a huge leaky faucet!

### Concepts

- Density
- Solubility
- Sublimation

### Materials

Gloves, insulated (for all demonstrations)

#### For Shrinking Suds

Dish detergent, Joy® or Dawn® work best

Water, warm

Dry ice, 2–3 large chunks, 40–50 g each

Graduated cylinder, 1-L or larger

#### For Misty Smoke Rings

Dry ice, 100 g

Plastic cups, 9-oz, and 10-oz, 1 each\*

Water, warm

Test tube, 2-cm diameter\*

Bunsen burner\*

\*A tall plastic beverage cup with a domed top may be used instead.

#### For CO<sub>2</sub> Crystal Ball

Dry ice, 500–700 g, 1 or 2 fist-sized chunks

Bucket with a smooth, flat rim

Soap solution made with Joy or Dawn dish detergent

Flashlight, lantern-type (optional)

Water, warm

Paper towel or absorbent cloth

Beaker or plastic cup, 250-mL

Stirring rod

#### For CO<sub>2</sub> Leaky Faucet

Dry ice, 2–3 chunks, 100–200 g each

PVC pipe elbow joint connectors,  $\frac{3}{4}$ ", 2

Soap solution, 3–5%, 50 mL

Rubber band

Water, warm

Scissors

Clamps, buret, 2

Soda bottles, 2-L plastic, 2

Cup, plastic

Support stand, tall

Paper towel

T-shirt or piece of cloth, newly washed (optional)

PVC pipe,  $\frac{3}{4}$ ", 60", 15", and 4"

Tubing, plastic (Tygon), 1" OD, 6-cm

## Safety Precautions

Dry ice (solid carbon dioxide) is an extremely cold solid ( $-78.5^{\circ}\text{C}$ ) and will cause frostbite. Do not touch dry ice to bare skin; wear appropriate insulated gloves or use tongs whenever handling dry ice. Wear chemical splash goggles and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Follow all laboratory safety guidelines. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

## Preparation

### Misty Smoke Rings

1. If a dome-top beverage cup is not available, you will need to make your own using two plastic cups, one tall, 10-oz and one wide mouth 9-oz.
2. Heat the mouth of a test tube in a burner flame for about 1 minute, rotating constantly.
3. Press the hot mouth of the test tube in the center of the bottom of the wide mouth cup, until a hole is melted through. This will serve as the lid for the tall cup (see Figure 1).

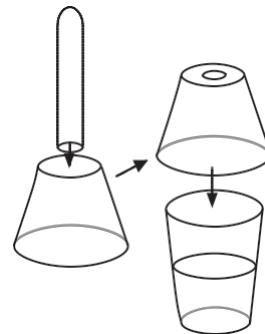


Figure 1.

### $\text{CO}_2$ Crystal Ball

1. Roll the paper towel or cloth into a strap, 3–4 cm in width and 10–20 cm longer than the diameter of the bucket.
2. Add 3–5 mL of dish soap to a 250-mL beaker or a plastic cup. Add water to the 100 mL mark. Stir gently.

### $\text{CO}_2$ Leaky Faucet

1. Cut the top off one 2-L soda bottle, leaving the sides curved in at the top. Save the bottom part.
2. Cut the top off a second 2-L bottle, just where the curve begins. Save the top part.
3. The saved bottom portion should fit inside the saved top portion and hold together with friction (see Figure 2).
4. Construct a large faucet head (like an upside-down “J”) with the PVC pipes and elbow joints (see Figure 3).
5. Roll up a paper towel and wrap it around the mouth of the faucet, securing the towel in place with a rubber band.
6. Use Tygon tubing to secure the bottom of the “faucet” over the 2-L bottle top (see Figure 3).

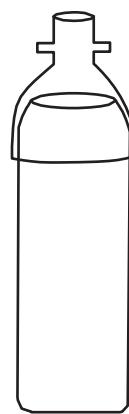


Figure 2.

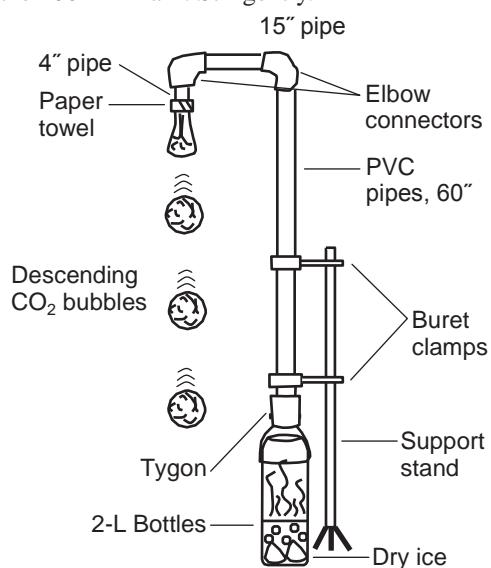


Figure 3.

## Procedure

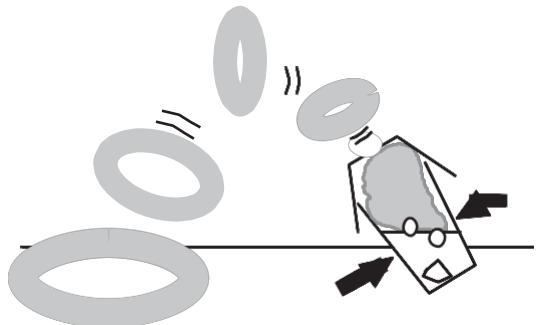
### Shrinking Suds

1. Fill the graduate cylinder roughly one-third of the way with warm water.
2. Wearing gloves, drop in 2–3 chunks of dry ice and allow the students to make observations. Notice the rapid sublimation and the “fog” produced.
3. Add a small squirt (2–3 mL) of detergent to the graduated cylinder. You will soon observe hundreds of fog-filled bubbles quickly accumulating on top of the water. As new ones form, the bubbles on top get pushed upward past the mouth of the container. Then a rather unusual phenomenon occurs—the bubbles on the outside layer begin to shrink.
4. To more easily observe the shrinking suds, wet one hand and scoop out a handful of the large, golf ball-sized bubbles. Watch as they shrink in a matter of seconds to the size of BB's.

### Misty Smoke Rings

1. Fill the tall cup halfway with warm water.

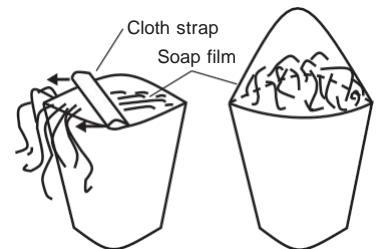
- Drop in one or two 20-g chunks of dry ice.
- Cap the tall cup with the domed lid or with the wide-mouth cup with the hole in the bottom (see Figure 1).
- Tilt the cup slightly, then give a few small, quick squeezes to the sides of the cup to generate CO<sub>2</sub> smoke rings (see Figure 4).
- If conditions are right (rate of sublimation and stillness of air in the room), the smoke rings will last for several seconds, following graceful projectile arcs. Then, as the rings approach the lab table top, they hover momentarily as they spread themselves out and dissipate.



**Figure 4.**

### CO<sub>2</sub> Crystal Ball

- Fill the bucket about halfway with warm tap water.
- Drop in one or two fist-sized chunks of dry ice. Allow students to observe as the fog eventually spills over the rim.
- Soak the rolled up paper towel strap in the soap solution, remove the strap from the beaker and allow the excess solution to drip off.
- Wet the rim of the bucket with the soapy strap.
- Draw the soapy strap slowly over the rim of the bucket to create a soap film "lid" (see Figure 5).
- As the film slowly inflates into a dome, blow gentle puffs of air at the dome and observe the way it resonates.



**Figure 5.**

### CO<sub>2</sub> Leaky Faucet

- Add 50 mL of a 3–5% soap solution made with liquid dish detergent to a plastic cup.
- Hold the cup of soap solution up to the mouth of the faucet to wet the paper towel and to establish a soap film across the opening.
- Fill the bottom portion of the 2-L bottle halfway with warm water and drop 2–3 chunks of dry ice into the water.
- Insert the 2-L bottle bottom into the top portion attached to the faucet until a good friction fit is made and the two portions hold together.
- Observe as a procession of misty CO<sub>2</sub> bubbles form and fall like huge, pearly water drops from a giant leaky faucet!
- Optional:* Try catching and bouncing the falling bubbles on a recently washed T-shirt or piece of cloth.

## 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 dry ice to sublime in a well-ventilated area. Soap solutions may be rinsed down the drain according to Flinn Suggested Disposal Method #26b.

## Tips

- Just for fun and to add to the effect, set up a few small ring toss targets for the Misty Smoke Rings, such as pencils held upright with balls of modeling clay.
- After the dry ice has been added to the bucket for the CO<sub>2</sub> Crystal Ball, place a lit lantern-type flashlight face-up in the bucket, then draw the soap film across. The crystal ball will now glow in the dark! Colored filters may be placed over the light to produce different colors!
- Try igniting the CO<sub>2</sub> Leaky Faucet "drips" to show that, unlike methane or hydrogen bubbles, carbon dioxide bubbles do not ignite; in fact, they will extinguish a flame! Thus, not only do different gases vary in their physical properties such as density, they also vary in their chemical properties such as flammability. Point out to the students, however, that these two

properties are not in any way linked—that is, there are some low density gases such as helium that do not burn, and plenty of high density gases such as propane and ether that do burn!

## Discussion

The stable state for CO<sub>2</sub> at normal room conditions is the gaseous state. Since barometric pressure (1 atm) is well below CO<sub>2</sub>'s triple point pressure (5.1 atm), the liquid state for CO<sub>2</sub> is not stable. Thus, solid CO<sub>2</sub> (dry ice) tends to change directly to a gas; that is, it sublimes.

Out in the air, it is hard to notice this sublimation, for the CO<sub>2</sub> gas produced is clear and colorless. When placed in water, however, this sublimation becomes much more obvious, for now the escaping CO<sub>2</sub> can be observed as large, rapidly evolving bubbles. Also, water serves as a better heat sink than air, and so the sublimation is sped up substantially.

It is important to point out that the thick fog that is observed is not the CO<sub>2</sub> gas; it is actually H<sub>2</sub>O in the liquid state—small, suspended droplets of condensation produced as warm vapor from the water comes in contact with the cold subliming CO<sub>2</sub>, much like the fog you see when your warm, humid breath is exhaled into cold winter air.

In the Shrinking Suds demo, the foggy CO<sub>2</sub> bubbles tend to shrink when exposed to air simply because the CO<sub>2</sub> gas, being relatively soluble in water, can diffuse readily across the soap film membrane. With air on the outside of the bubble, the higher CO<sub>2</sub> concentration is obviously on the inside, and the net diffusion would thus be outward. For the same reason, N<sub>2</sub> and O<sub>2</sub> from the air in the room would diffuse into the bubble. This should counteract the shrinking caused by the exiting CO<sub>2</sub>, but since CO<sub>2</sub>'s solubility in water is dramatically greater than air's, the net effect is one of shrinking. The inward diffusion of air could explain, however, why the bubbles reach a certain minimum size of about 1–2 mm in diameter and then stabilize. This minute volume very well might represent the air that has diffused in while the CO<sub>2</sub> was diffusing out. Perhaps this could even be used in a quantitative way to illustrate the relative solubility of the CO<sub>2</sub> and air, respectively. For example, if a CO<sub>2</sub>-filled bubble with an original diameter of 2.75 cm ( $V = 10.9 \text{ cm}^3$ ) shrank down to an air-filled bubble with a diameter of 0.12 cm ( $V = 0.00090 \text{ cm}^3$ ), this would imply that 10.9 cm<sup>3</sup> of CO<sub>2</sub> diffused out through the bubble membrane in the same time that only 0.00090 cm<sup>3</sup> of air diffused in. Does this indicate that CO<sub>2</sub> is 12,000 times more soluble than air?

It is also worth noting that the CO<sub>2</sub> bubbles do not appear to shrink as they are rising up inside the container, for there they are surrounded by other bubbles that presumably contain an equally high concentration of CO<sub>2</sub>. This is not to say that CO<sub>2</sub> diffusion between bubbles is not occurring. Rather, because the CO<sub>2</sub> concentrations are essentially the same on either side of the soap film, the diffusion rates into and out of the bubbles in the cylinder are more or less the same.

The Misty Smoke Rings are generated in much the same way as normal smoke rings. The quick pulse of air sends smoky mist out through the top hole of the cup lid. Backwards-spinning eddy currents occur as the outer edge experiences drag from the surrounding air. These seem to circulate completely around the ring (see Figure 6). Rather than follow the more-or-less straight-line path of normal smoke rings, however, the high density of the cool, moisture-laden carbon dioxide causes these misty smoke rings to follow a graceful parabolic path. As the rings are about to land on a flat horizontal surface, two more remarkable things happen. First, the rings hover as if on a cushion of air about 1–2 centimeters above the surface. Second, as they hover, the rings quickly spread out then disappear as the mist of tiny water droplets from which they are made evaporates into the warmer air.

In the CO<sub>2</sub> Crystal Ball demonstration, once the soap film lid is established, it quickly forms into a misty dome—dramatically resembling a fortune-teller's “crystal ball,” only considerably more fluid and pliable. Strategically blown puffs of air can set up any number of resonance frequencies. (This effect could perhaps be used in conjunction with the time-honored vibrating string analogy for modeling electron energy levels as harmonic frequencies!)

If left undisturbed, the film will quickly produce intense horizontal rainbows of color circling the dome. These are due to the interference patterns of the reflecting light. This happens in all soap films as a result of varying film thicknesses, but these colors seem especially pronounced in the crystal ball, perhaps because of the misty white backdrop.

After some time, another peculiar observation can be made. The dome stops growing, even though one can still hear the dry ice subliming vigorously in the bucket below. The explanation for this phenomenon centers on the relatively high solubility of CO<sub>2</sub> and its ability to diffuse readily through the soap film. First let us assume that the CO<sub>2</sub> is subliming at a more or less constant rate, and that this rate of sublimation is considerably greater than the rate at which the CO<sub>2</sub> can diffuse through the original soap

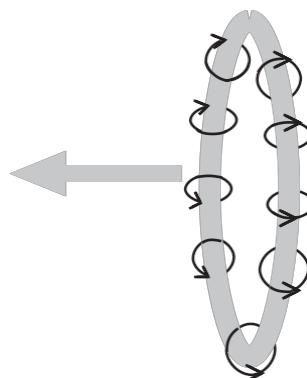


Figure 6.

film “lid” drawn across the bucket. The film is thus pushed outwards into a dome. As the film grows, however, and its surface area increases, so does the rate at which the CO<sub>2</sub> can diffuse through it. Hence, an equilibrium-like state is eventually reached where the rate of diffusion is equal to the rate of sublimation. In short, the dome reaches a point where it is “leaking” as fast as it is being filled, and so its size remains more or less constant. Actually, of course, the sublimation rate is decreasing as the dry ice gets used up. At first, this deceleration is too gradual to play much of a role in the demonstration, but as the chunks of dry ice begin to dwindle (and as the water temperature drops and layers of ice begin to form on the dry ice, insulating it from the rest of the water) the rate of sublimation can decrease quite substantially. When this happens, if the soap film has lasted long enough, one can see the dome start to shrink back down into the bucket.

The misty bubbles fall from the CO<sub>2</sub> Leaky Faucet remarkably fast due to the relatively high density of CO<sub>2</sub>. This is not to say that Galileo was wrong—that denser bodies have a greater gravitational acceleration. Rather, because of air resistance, all falling objects eventually reach terminal velocity—when the frictional force acting upward upon them matches the gravitational force acting downward. With all else being equal, objects that have a high density (and thus a high mass to surface area ratio) take longer to reach their terminal velocity; they therefore accelerate for a longer period of time and end up falling faster.

## 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 A: Science as Inquiry

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

### ***Content Standards: Grades 9–12***

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, structure and properties of matter, motions and forces

## Acknowledgments

David Brooks of the University of Nebraska, Lincoln, was the first to show the presenter the wonderful effects that can be achieved by incorporating soap films into dry ice demonstrations.

## References

Boys, C. V.; *Soap-Bubbles, Their Colours and the Forces Which Mold Them*; Dover, 1959.

Isenberg, C.; *Problem Solving with Soap Films: Part I*; Physics Education, September 1975, 10, 452–456.

Isenberg, C.; *Problem Solving with Soap Films: Part II*; Physics Education, November 1975, 10, 500–503.

Noddy, Tom; *Bubble Magic*; Running Press, 1988.

## Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Dry Ice Demonstrations* activity, presented by Bob Becker, is available in *Dry Ice Demonstrations*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

## Materials for Dry Ice Demonstrations are available from Flinn Scientific, Inc.

Catalog No.	Description
AP4416	Dry Ice Maker
C0241	Cleaner, Liquid, Detergent
SE1031	Gloves, Cotton and Canvas
GP9090	Cylinder, Borosilicate Glass, 1000-mL
AP6009	Bucket, Plastic Utility Pail
AP5998	Support Stand, 36" Rod

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

## Giant Chemical Twister

### Overview

Observe the color change of this twisting vortex as different amounts of acid and base are added.

### Purpose

This demo shows the color change of phenolphthalein from colorless to pink in the presence of base and a return to colorless as more HCl is added to solution.

### Materials

- 2000 mL Graduated Cylinder
- 2 M HCl
- 2 M NaOH
- Phenolphthalein
- Magnetic Stirrer
- Magnetic Stirrer Bar

### Procedure

Add a stirring bar to one liter of tap water in a graduated cylinder and place this on an electric stirrer. Turn on the electric stirrer. Add 50 mL of 2 M HCl to the graduated cylinder along with 3-5 drops of phenolphthalein. The solution should be colorless at this point. Then add 2 M NaOH drop wise into the vortex. The vortex will turn pink in the presence of a base. When the entire solution eventually turns dark pink, add more HCl to the graduated cylinder to make the solution colorless again. Repeat and have pHun!

### Safety

*Hydrochloric acid solution, although dilute, is severely corrosive to eyes, skin and other tissue. Sodium hydroxide solution, although dilute, is corrosive; skin burns are possible; very dangerous to eyes. The indicator solution contains ethyl alcohol, which is a flammable liquid and a fire risk; keep away from heat and 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.*

## Giant Balloon Blower Upper

### Introduction

Nitrogen makes up ~ 80% of the atmosphere. When it is cooled to -196 °C (-320°F) it condenses into a liquid. When the liquid is heated to room temperature, it rapidly turns to a gas and expands. Liquid nitrogen has a density of ~0.81 g/mL. As a gas, the density decreases to ~0.0013 g/mL. This translates to approximately a 600 X expansion in volume. The Giant Balloon Blower-Upper qualitatively shows this expansion by converting the liquid nitrogen into a gas, and trapping the gas in giant balloons attached to the “inflation” end.

### Supplies

- Giant Balloon Blower-Upper
- 20-24” latex balloons
- Dewar Flask (Flinn AP8562)
- Liquid nitrogen from local gas/welding supplier.

### Safety

- Liquid nitrogen can cause severe frostbite. Always wear a safety mask and/or goggles. Use appropriate thermal protection on your hands.
- Beware that if the large fill cap comes off while attempting to screw it on after liquid nitrogen has been added, some liquid can come back up out of the filling side of the pipe.
- Make sure the entire device is behind a blast shield, or at least 20 feet away from observers. In a classroom setting students should wear safety goggles to protect them from any small pieces of balloon.
- Check with students prior to the demonstration for latex allergies. Small fragments of the latex balloons can fly 10-15 feet when they explode.
- Warn students to cover their ears. The popping balloons can be loud and startling.
- Ear protection for the demonstrator is recommended.

### Procedure

1. Seal the middle balloon opening so that only two balloons fill. This decreases the amount of nitrogen needed and generates questions after the demonstration.
2. Check three large balloons to make sure they have no leaks or defects by blowing them up slightly.
3. Attach three very large (~20-24”) latex balloons to each of the pipe ends. Each balloon should be secured with three rubber bands.
4. Unscrew the fill cap. Add approximately 2-Liters of liquid nitrogen. Quickly screw on the cap making sure that it forms a tight seal.
5. Move to a safe position approximately 10-15 feet away. Wait for the balloons to expand and burst\*.
6. Usually both balloons will explode at the same time. If too little nitrogen has been added the balloons will get very large, but may not burst. If this happens, approach the balloons and pop them with a pencil. If only one balloon bursts, the second balloon can be popped with a pencil. Sometimes a balloon may burst too soon due to a manufacturing flaw. If this happens, put on ear protection and place your gloved hand over the pipe with the burst balloon to seal off the leaking nitrogen. This should allow the second balloon to finish filling and burst.

**Credit:**

Lee Marek from Naperville North High School originally showed me this demo using a metal pipe about one foot in length. I asked my brother Ed the electrician to make me a slightly larger version. The Giant Balloon Blower-Upper is his creation. He enjoys scaling things up!

**Technical Notes:**

2 " ID galvanized heavy walled steel electrical conduit (Fill end and main pipe)

Threaded galvanized cap to secure fill end of the main pipe

1/2 " ID gallvanized heavy walled steel electrical conduit

Varoious reducing bushings to connect the 2" main pipe to 1/2" balloon pipes

- Total length, 41 inches center of fill pipe to center rise of balloon pipes
- Total height, 21 inches
- Width of balloon pipes, 11 inches center to center



## Pink Catalyst

### Introduction

Add a pink cobalt chloride solution to a colorless solution containing potassium sodium tartrate and hydrogen peroxide and watch as a very obvious green-colored complex forms. As the reaction ends, the solution will return to its original pink color—indicating that the cobalt chloride catalyst is not used up in the reaction.

### Concepts

- Catalysts
- Kinetics

### Materials

Cobalt chloride solution,  $\text{CoCl}_2$ , 0.1 M, 15 mL  
Hydrogen peroxide solution, 6%,  $\text{H}_2\text{O}_2$ , 40 mL  
Potassium sodium tartrate solution, 0.2 M, 100 mL  
Water, distilled, approximately 100 mL  
Beaker, 150-mL  
Beaker, 500-mL or 1-L

Graduated cylinder, 100-mL  
Hot plate  
Spatula or scoop  
Stirring rod or magnetic stirrer  
Thermometer

### Safety Precautions

*Cobalt chloride solution is moderately toxic by ingestion; the solid is a possible carcinogen as a fume or dust. Hydrogen peroxide is a strong oxidizer and a skin and eye irritant. Avoid contact of all chemicals with eyes and skin. 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

Prepare 0.2 M potassium sodium tartrate solution by dissolving 6 g of  $\text{KNaC}_4\text{H}_4\text{O}_6 \cdot 4\text{H}_2\text{O}$  in 100 mL of distilled water.

### Procedure

1. Using a graduated cylinder, measure out 100 mL of 0.2 M potassium sodium tartrate solution and pour it into a 500-mL or 1-L beaker.
2. Place the beaker on a hot plate at a medium setting and slowly warm the solution to 70 °C.
3. When the temperature of the potassium sodium tartrate solution reaches 70 °C, carefully add 40 mL of 6% hydrogen peroxide. No reaction is observed.
4. Note the initial pink color of the cobalt chloride solution, and then add the solution to the 500-mL beaker and stir.
5. Observe the rate and progress of the resulting chemical reaction. (*The solution turns bright green and vigorous bubbling ensues within one minute. The mixture begins to froth and foam, then just as suddenly subsides. When the rate of bubbling diminishes, the green color disappears and the pink color of the cobalt chloride solution returns.*)

### 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 final cobalt-containing solution may be treated according to Flinn Suggested Disposal Method #27d.

## Tips

- The use of a large beaker is important so that the reaction does not froth over.
- To measure the reaction rate, begin timing the reaction upon addition of the pink catalyst solution. Complete the timing after the original pink color of the cobalt chloride solution has returned. Using a hot plate–magnetic stirrer combination is strongly recommended to measure reaction times. The reaction rates may be measured at different temperatures (50, 60, and 70 °C)

## Discussion

The reaction of tartrate ions with hydrogen peroxide is an example of an oxidation–reduction reaction. Hydrogen peroxide is a strong oxidizing agent, resulting in the complete oxidation of tartrate ions to give carbon dioxide and water (Equation 1).



In the absence of a catalyst, the decomposition reaction, although thermodynamically favorable, is kinetically very slow. Thus, even at 70 °C, the reaction occurs at a barely noticeable rate.

To speed up the reaction, a catalyst must be used. Cobalt ions are known to catalyze the decomposition of hydrogen peroxide. The action of the cobalt catalyst can be followed by observing the color changes of the solution. The solution starts out pink, the color of the cobalt(II) aquo complex  $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ . The mixture then quickly turns green, indicating the formation of a cobalt(III) complex. The rapid production of gas bubbles due to oxidation of the tartrate ions occurs almost immediately after the green color has been observed. As the tartrate ions are consumed and the amount of gas production subsides, the color of the solution returns to the original pink color of the cobalt(II) catalyst.

The first step is the formation of a Co(II)-tartrate coordination compound. This is followed by oxidation to a green Co(III) complex which is thought to be the actual catalyst in the reaction. In the course of oxidation of the tartrate ions, the Co(III) complex is reduced back to Co(II). When all the tartrate has been consumed, the color of the solution reverts back to pink, indicating that only Co(II) ions are present.

Cobalt chloride is a catalyst because it is not consumed during the course of the reaction and it greatly speeds up the reaction.

## 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

## Acknowledgment

Special thanks to Jim and Julie Ealy (retired), The Peddie School, Hightstown, NJ for bringing this demonstration to our attention.

## References

Cesa, I. *Flinn ChemTopic™ Labs*, Vol. 14, *Kinetics*. Batavia, IL: Flinn Scientific, 2002.

## Materials for *The Pink Catalyst* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP2084	Pink Catalyst—Chemical Demonstration Kit
C0242	Cobalt Chloride Solution, 0.1 M, 500 mL
P0084	Potassium Sodium Tartrate, 100 g
H0028	Hydrogen Peroxide, 6%, 500 mL

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

## World's Largest Salt Bridge

**Description:** A galvanic cell is constructed using  $Zn^{2+}/Zn$  and  $Cu^{2+}/Cu$  couples with a “human” salt bridge.

**Materials:**

1 M $CuSO_4$	Voltmeter
1 M $ZnSO_4$	Connecting wires
Zn and Cu strips	

**Procedure:**

For large lecture halls, perform demonstration on base of document camera.

1. Pour 1 M  $CuSO_4$  and 1 M  $ZnSO_4$  solutions in separate containers (or use setup in Dabney 114). Set the Zn electrode in the  $ZnSO_4$  and the Cu electrode in the  $CuSO_4$  solution.
2. Connect the black wire (anode) to the Zn electrode and the red wire (cathode) to the Cu electrode. The voltmeter will not show any reading until the circuit is completed by touching each electrode with bare hands. The human body is used as the salt bridge

**Discussion:**

Due to their respective reduction potentials, a spontaneous redox process occurs when the zinc electrode is connected to the negative terminal and the copper electrode is connected to the positive terminal of the voltmeter. The spontaneous flow of electrons from anode and cathode generates a current with a voltage lower than the theoretical  $E^\circ$  cell for these couples (1.10 V). This demonstration should be performed in conjunction with the same experiment utilizing a NaCl salt bridge. Have students hypothesize why the two different types of salt bridges give different results. Compare dry versus “sweaty” hands by wiping NaCl solution on the student’s hands.

			$E^\circ_{red}(V)$
Anode	$Zn$	$Zn^{2+} + 2 e^{-}$	-0.76
Cathode	$Cu^2 + 2e^{-}$	$Cu$	+0.34

**Safety:** Wear proper protective equipment including gloves and safety glasses when preparing and performing this demonstration.

**Disposal:** Copper and zinc solutions can be reused as long as no contamination occurred. Electrodes should be rinsed, dried and cleaned by scrubbing with steel wool.

**References:**

Scharlin, P.; Battino, R. *J. Chem. Educ.* **1990**, 67, 156. (variation utilizing a wide variety of metals)

## Water to Wine Rap

### Materials:

6 wine glasses  
150 ml of distilled water  
10 drops of 1 m NaOH  
3 squirts of 1% phenolphthalein  
6 squirts of 6m HCl  
One small scoop of NaHCO<sub>3</sub> (s)  
3 drops of 1m AgNO<sub>3</sub> (aq)

### Procedure:

1. Set up glasses in a row at the front of the demonstration table. If using glasses, the beverages will be in the following order: wine, water, 7-Up, and milk.
2. Prepare the glasses/beakers as follows:  
Glass 1- Wine bottle 150 mL of distilled water and 10 drops 0.1 M NaOH solution.  
Glass 2 – Wine, 3 squirts 1% phenolphthalein (ADD JUST BEFORE START OF DEMO) Glass 3– Water, 6 squirts 6M HCl  
Glass 4 – 7-Up, small scoop of NaHCO<sub>3</sub> (s) - Glass 5 – Milk, 3 drops 1M AgNO<sub>3</sub> (aq)
3. Hold up the wine bottle and pour it into the wine glass. The phenolphthalein will turn pink in the basic environment (wine).
4. Pour contents of glass 1 into glass 2. The pink color will vanish, and the solution will return to colorless (water). Phenolphthalein is colorless in an acidic solution.
5. Pour the contents of glass 2 into glass 3. The acidic solution will react with the sodium bicarbonate to produce carbon dioxide bubbles in an acidic solution (7-Up).
6. Allow the bubbling to stop completely, then pour the contents of glass 3 into glass 4. The clear solution will develop a cloudy white precipitate (milk). The precipitate is formed when the barium ions in the barium nitrate solution react with the sulfate ions from the sulfuric acid to form barium sulfate, which is insoluble in water.

### **The Rap:**

*I've got a rappin' demo for you to see  
And it has to do with  
Chemistry, So all you  
people listen up,  
And cast your eyes upon this cup!*

*Chemistry is really fine,  
I can turn this water into wine,  
And if you aren't impressed  
my friend, I'll turn this wine  
into water once again.*

*And if you don't like what's  
in my cup, I'll turn this  
water into Seven-Up.*

*Or if you want a drink that's  
smooth as silk, I'll turn this  
Seven-Up into milk.*

*Now the thing that makes this  
demo great, Are the acids,  
bases, and precipitates,  
So now you've seen and you  
understand, I think this demo  
deserves a hand!*

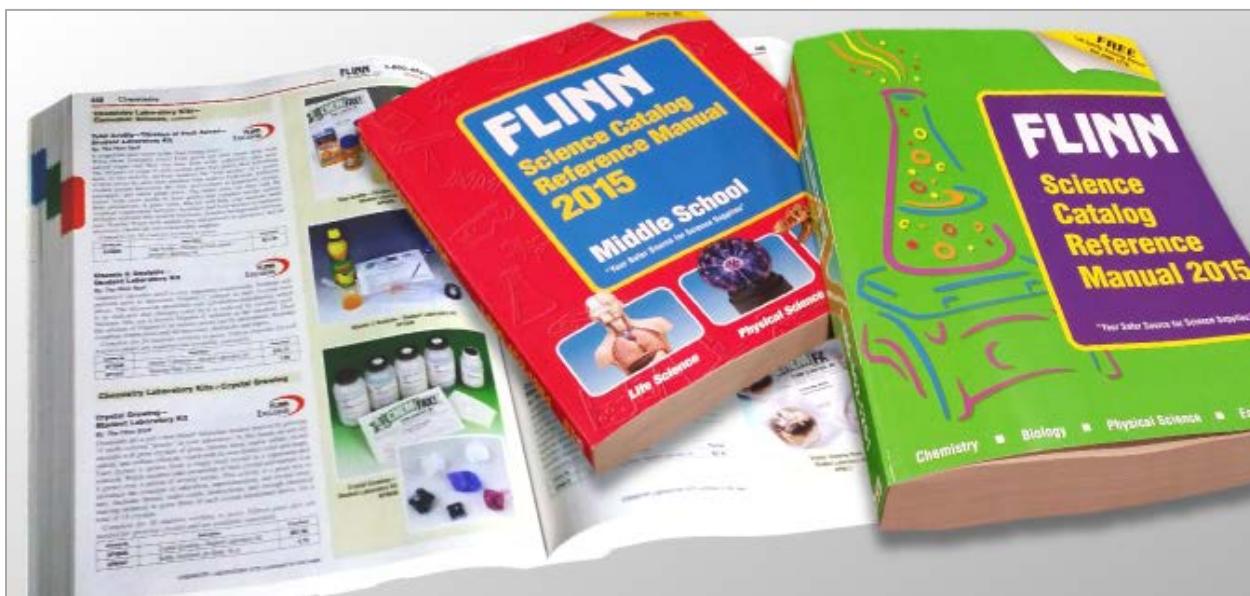
### **Reactions:**

- Glass 2 – NaOH + phenolphthalein --> pink solution The phenolphthalein turns pink in the basic environment (wine).
- Glass 3 - NaOH + H<sub>2</sub>SO<sub>4</sub> (in excess) --> H<sub>2</sub>O + Na<sub>2</sub>SO<sub>4</sub> (colorless solution) The pink color will disappear, and the solution will go back to colorless (water). Phenolphthalein is colorless in an acidic solution.
- Glass 4 - HCO<sub>3</sub><sup>-</sup> (aq) + H<sup>+</sup> (aq) --> H<sub>2</sub>CO<sub>3</sub>(aq) --> CO<sub>2</sub>(g) + H<sub>2</sub>O (l) (fizzing)  
$$2\text{NaHCO}_3(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + 2\text{CO}_2(\text{g})$$
 The acidic solution reacts with the sodium bicarbonate to produce carbon dioxide bubbles in an acidic solution (7-Up).
- Glass 5 - Ba<sup>2+</sup>(aq) + SO<sub>4</sub><sup>2-</sup>(aq) --> BaSO<sub>4</sub>(s) (white precipitate) The clear solution develops a cloudy white precipitate (milk). The precipitate is formed when the barium ions in the barium nitrate solution react with the sulfate ions from the sulfuric acid to form barium sulfate, which is insoluble in water.

## **Explanation:**

In this demonstration, it is clear that chemical changes have taken place. Indicators show that the chemical reaction has taken place by changing the color of the solution. The solutions in this experiment are acids and bases. When you change the amount of an acid or of a base in a solution, the color can change quickly, which makes it seem like different drinks are appearing. Another way that a solution can look different is when reactions make precipitates. The precipitates are actually solids that are made during the reaction that takes place.

In the first glass, I put 10 drops of sodium hydroxide and water. In the second glass, it was time for the indicator to be put in. The indicator that was used in this experiment was phenolphthalein. I put 3 squirts of this into a glass to make the “wine.” The indicator allowed the clear water to turn into pink. In the third glass, I put 6 squirts of Hydrochloric acid to get the clear color of water back. In the fourth glass I put a small scoop of baking soda, which is a base. The acid from the previous glass reacted with the baking soda to make it fizz, which looks like 7-up. Finally, for the last glass, I added in silver nitrate to make it turn into a cloudy, white precipitate. This gave the illusion of milk.



**Fill out the raffle card on the back cover of this handout to request a copy of the 2015 Flinn Catalog/Reference Manual.**

## Diving Tony Teaching with Toys

### Introduction

What goes down must come up? A variety of squeezable/sinkable cartesian divers can be made with the simplest of equipment and materials . . . and a little imagination.

### Concepts

- Density
- Boyle's Law

### Materials

Beaker, 600-mL

Pipets, Beral-type, disposable plastic, 100

Hex nuts, 100

Plastic soda bottle, 2-L

### Safety Precautions

The materials used in this activity are considered nonhazardous. Please follow all standard laboratory guidelines.

### Procedure

1. Fill the 600-mL beaker approximately  $\frac{4}{5}$  full with tap water.
2. Cut off all but 15 mm of the pipet stem (see Figure 1). Then screw the hex nut securely onto the truncated stem. The hex nut will make its own threads as it goes.

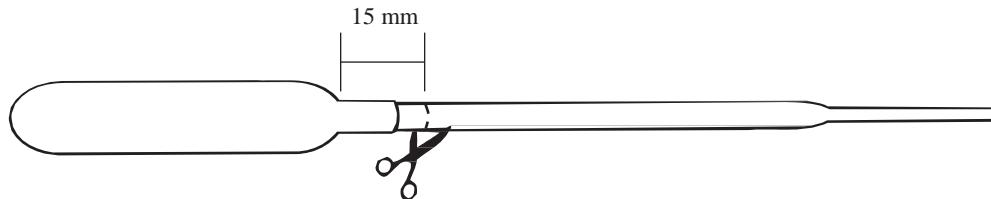


Figure 1. Cutting the Pipet

3. Place the pipet-nut diver assembly into the beaker of water and observe that it floats rather buoyantly in an upright position with the hex nut acting as ballast.
4. Squeeze out some of the air and draw some water up into the pipet. Now check the buoyancy. If you draw up too much water, the assembly will sink. If this happens, simply lift it out of the water, squeeze out a few drops of water and let air back in to replace the water.
5. Using this technique, adjust the amount of water in the assembly so that it just barely floats (in other words: fine tune the assembly's density to make it slightly less than that of water).
6. Place the diver assembly in a plastic 2-L bottle filled with water and screw on the cap securely (see Figure 2). Observe how the assembly dives to the bottom as you squeeze the bottle and how it rises to the surface as you release the squeeze.

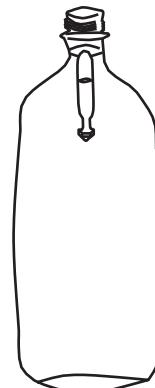


Figure 2.

## Discussion

Whether an object floats or sinks in a fluid depends on the object's density versus the density of the fluid. Density = mass/volume ( $D = m/v$ ). Thus, if the mass of an object increases while its volume remains constant, the object's density will increase. The density of an object will also increase if its volume is reduced while its mass remains constant. Boyle's Law states that as the pressure on a gas sample is increased, it gets compressed into a proportionately smaller volume. That is, there is an inverse relationship between the pressure exerted on a gas and its volume. While gases are compressible, liquids and solids are not.

In this activity, a Cartesian diver was constructed. Initially the diver was placed in the water containing gas and no water inside. Since the density of the air in the diver was less than the density of the water, the diver floated. As the diver was filled with water, its density increased and therefore sank in the beaker.

The Cartesian diver was placed in a 2-L bottle. In this case, the density of the diver remained the same but the density of the surrounding fluid changed. As the 2-L bottle was squeezed, the air pocket in the bottle was compressed and thus the total volume of the assembly decreases. Since the mass remains constant, the diver assembly's density increases. Conversely, as the 2-Liter bottle is released the air pocket is allowed to expand thus expanding the total volume of the apparatus. Since the volume is increasing, the diver's density decreases and therefore rises back to the top of the bottle.

## Tips

- It is considerably more convenient to adjust the density and to test for flotation in the 60-mL beaker or in a tub of water, rather than in the bottle itself. A cut-off 2-L bottle works well as a testing tank!
- It is advisable to fill the 2-L bottle completely with water. That way, when the bottle is squeezed, the work will go into compressing the air pocket in the diver and not into compressing a large air space at the top of the bottle.

## Extensions

Although the standard diver described above is amusing and educational, the real fun comes in trying some creative variations, such as those listed below:

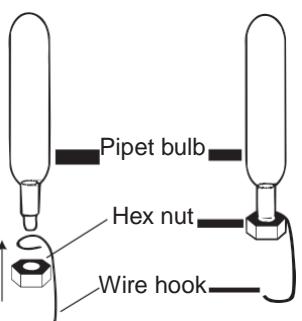


Figure 3.

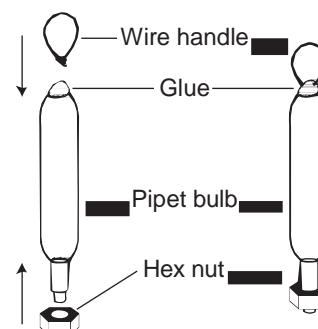


Figure 4.

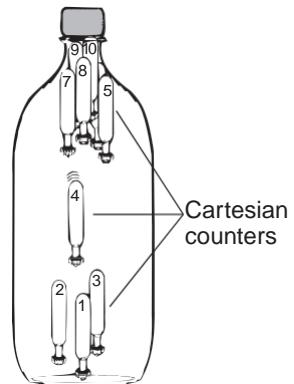


Figure 5.

1. **The Sunken Diver.** Adjust a diver's density so that it just barely sinks and then put it in a bottle of water. Try to find a way to make the diver ascend to the surface. Some ideas might include: taking the cap off and heating the bottle, or placing it in a vacuum jar, or perhaps adding a solute to increase the density of the surrounding water.

2. **Cartesian Retrievers.** Place two divers in the same bottle — one that barely floats and one that barely sinks, but with mechanisms or devices attached to them that will enable the floating one to dive down and retrieve the sunken one off the bottom. Use magnets, chewing gum, Velcro, a suction cup, a net, a hook and handle—whatever works! See Figures 3 and 4.

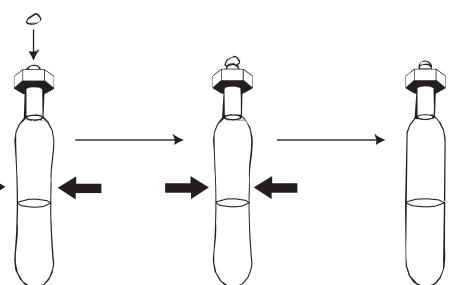


Figure 6.

3. **Cartesian Counters and Messages.** Place several numbered divers together in one bottle, but all with different densities, so they descend in order—1, 2, 3 . . . (or letter the divers to spell out a secret message!) See Figure 5.
4. **Diving Whirligigs.** Cut a small sheet of plastic into a pinwheel. Punch a hole in the center and fit it onto the stem of the pipet, just above the hex nut. Now the diver will spin gracefully as it sinks, and reverse its spin on the way up. Attach pipe-cleaner arms and legs to make an unusual diving ballerina!
5. **Closed-System Divers.** After the density has been adjusted, try sealing the mouth of the diver with a drop of hot glue. Now, when the bottle is squeezed, instead of water being forced up into the diver's mouth, the sides of the diver are forced noticeably inward (see Figure 6). This closed system allows the use of colored water inside the diver and results in divers

that can be stored and transported outside the bottle. What's more, the shape distortion may be used in several ways: for instance, wires may be attached to the sides of the diver and fashioned into "jaws" that hang downward. Then, when the middle gets pushed inward, the jaws spring open and a ferocious Cartesian shark dives downward with his mouth open to snatch a unsuspecting diver off the bottom!

6. **Density Column Divers.** Make a density column inside the bottle. This can consist of anything that forms layers (half oil/half water, for example). Use several divers—all adjusted to suspend themselves at different levels throughout the bottle.
7. **Remote-Controlled Divers.** Use airline tubing (aquarium tubing) or Tygon® tubing to connect two plastic bottle caps together. Screw one cap onto a water-filled bottle containing a standard diver (or any of the variations listed above) and screw the other cap onto a second bottle (the remote control) that just contains water. When the remote control is squeezed, the diver in the first bottle will descend, even from across the room through several meters of tubing! As mentioned above, it helps to have the bottles as full as possible and to have the tubing completely filled with water as well. Try replacing the water-filled bottle with a bottle of soda. Instead of squeezing, just shake! Or use the carbon dioxide-producing reaction between baking soda and vinegar to create the pressure in the remote control bottle.
8. **The Cartesian See-Saw.** Try to construct two Cartesian divers and attach them to the ends of a see-saw structure that alternately tips back and forth as you squeeze and release the bottle. At first, this might seem impossible, for both divers would increase in density as the bottle is squeezed. But, by varying the length of the lever arms or by making one diver more sensitive than the other (by using, for example, a regular diver on one end and a closed-system diver on the other), such an underwater see-saw is feasible!
9. **Concentric Divers.** Make a diver small enough to fit inside another one, so as the little one dives inside the bigger one, the bigger one dives inside the bottle!
10. **The Electric Diver.** Build a diver with a built-in circuit that causes a light to go on or a bell to ring when the diver descends.

*Note:* This list of variations has been presented with the intention of demonstrating the versatility and expandability of a concept that many may have considered very limited. From buoyancy to pressure to surface tension to density to chemical reactions and electrical conductivity, the Cartesian diver can act as a springboard (pun intended!) for a variety of fun and exciting educational activities.

## 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 A: Science as Inquiry

Content Standard B: Physical Science, structure and properties of matter, motions and forces

## Acknowledgments

Special thanks to Bob Becker for providing us with this activity. Cartesian Diver drawings provided by Susan Gertz.

## Further Reading

Sarquis, M.; Sarquis, J. L. *Fun with Chemistry: A Guidebook of K-12 Activities*; Institute for Chemical Education, University of Wisconsin: Madison, WI, in press; Vol. 3.

## Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Cartesian Divers* activity, presented by Jesse Bernstein, is available in *Teaching With Toys*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

## Materials for *Cartesian Divers* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP9082	Cartesian Diver Construction Kit

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

## Lota Bowl with Sodium Polyacrylate

### Introduction

A container is filled with water. The water is poured out of the container until it appears empty. In a few minutes, more water naturally appears in the container and it is also poured out. The lota bowl is a classic part of many magic shows. A version of it has been around since the time of the ancient Greeks. The apparent “magical” properties of the container make it an excellent device to teach problem-solving in many science classes.

### Concepts

- Problem-solving
- Air pressure

### Materials

Lota bowl	Large container to fill the lota bowl
Beaker, flask, or glass	Water

### Safety Precautions

*Although the materials used in this demonstration are considered nonhazardous, observe all normal laboratory safety rules.*

### Preparation

1. Fill a sink or container with enough water to submerge the entire lota bowl.
2. Place the lota bowl in the water and allow the bowl to completely fill. The bowl is filled when no more bubbles appear on the outside of the bowl. Pour out excess water until the water level lines up with outside hole.
3. Wipe all excess water from the outside of the lota bowl.
4. Place the lota bowl on a stand or countertop with the top hole facing away from students.

### Procedure

1. Casually show water in the lota bowl to a few students. Pour the water inside the lota bowl into a beaker or flask while keeping your finger over the *outside* hole (see Figure 1). Try to keep this outside hole away from student view.
2. Place the lota bowl back on a small stand or countertop and uncover the outside hole by removing your finger. The lota bowl will partially refill with water. Allow at least a minute for the bowl to refill.
3. Place your finger over the outside hole again and pour more water out of the lota bowl. Always keep your finger over the outside hole while pouring.
4. Replace the lota bowl on the stand or countertop, uncover the hole, and allow the inside bowl to again refill with water.
5. Repeat steps 3 and 4 for as long as you want to “milk” the presentation.

### Disposal

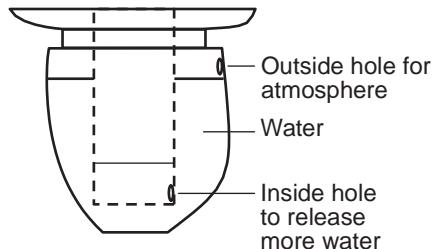
Empty any water remaining in the lota bowl before storing.

## Tips

- Practice handling and pouring water from the lota bowl so it looks natural covering the top hole.
- You can add more swirling and shaking of the bowl toward the end to give the impression that there are still large amounts of water being poured out of the container.
- This can also be done as a type of running gag throughout a class period. Pour the water once and then go onto something else. Return later in the same period and pour again. This can be done multiple times. It may take two or three pourings before students start questioning the device.

## Discussion

This lota bowl is an excellent device to use during discussions on air pressure or gas laws. It also has the essential properties to be a great problem-solving activity. The lota bowl is a container that has two shells. The inner shell holds water that is seen by the audience and poured out. There is also a hidden outer shell that contains a large amount of water between the outside wall and the inner shell that is not seen by the audience. When the inner shell is empty, some of the water from the outer shell will enter the inner shell once the outer hole is opened to the atmosphere. This process can be repeated many times although the amount of water being poured will gradually diminish. The water in the outer shell is able to enter the inner shell through another hole in the bottom of the inner shell (see Figure 1).



**Figure 1.**

An assignment to explain this device becomes interesting. As students begin to draw a sample model, they will usually leave out the outside hole. This outside hole is necessary to equalize pressure and prevent air lock. When the outside hole is sealed, most of the water stays in the outer shell and does not enter the inner shell. This is why there can be multiple pourings. As students attempt to build a model of the lota bowl, they will typically notice the necessity of the outside hole. Models can be built using different sizes of plastic cups or plastic soda pop bottles and a hole punch.

The lota bowl provides an excellent opportunity to discuss the effects of atmospheric pressure.

## 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 A: Science as Inquiry

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

### *Content Standards: Grades 9–12*

Content Standard A: Science as Inquiry

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

## Acknowledgment

Special thanks to Mr. Jeff Hepburn, Central Academy, Des Moines, IA, who provided us with the instructions for this activity.

## Reference

Sharpe, S. H., *Conjurors' Hydraulic and Pneumatic Secrets*; Hades: 1991.

**The Lota Bowl is available from Flinn Scientific, Inc.**

Catalog No.	Description
AP6212	Lota Bowl

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

## Rainbow Chemnection

### Introduction

A colorless solution is added to each of six beakers. A second colorless solution is added and each of the six resulting solutions turns a different color of the rainbow. Add more of the first solution and the rainbow of colors disappears.

### Concepts

- Acids and bases
- pH indicators

### Materials

Hydrochloric acid solution, 0.1 M, HCl, 1 L*	Beakers, 250-mL, 6
Sodium hydroxide solution, 0.1 M, NaOH, 1 L*	Beaker, 1000-mL
Six indicator solutions*: Violet, Blue, Green, Yellow, Orange, Red	Beaker, 2000-mL

\*Materials included in kit.

### Safety Precautions

Hydrochloric acid solution, although dilute, is severely corrosive to eyes, skin and other tissue. Sodium hydroxide solution, although dilute, is corrosive; skin burns are possible; very dangerous to eyes. The indicator solutions contain ethyl alcohol, which is a flammable liquid and a fire risk; keep away from heat and open flame. 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

1. Set up the six 250-mL beakers on an overhead projector or light box, or in front of the class.
2. Add 3 drops of “violet” indicator solution to the first beaker. Add 3 drops of “blue” indicator solution to the second beaker. Continue adding three drops of each of the other indicator solutions to the appropriate beakers.
3. Do not allow the solvent of each indicator solution to evaporate.
4. Dilute the 0.1 M sodium hydroxide 10 to 1 to make a 0.01 M solution. Do the same with the 0.1 M hydrochloric acid.

### Procedure

1. Add approximately 50 mL of the 0.01 M hydrochloric acid solution to each of the six beakers. All six resulting solutions should be clear.
2. Add approximately 75 mL of 0.01 M sodium hydroxide solution to each beaker. Each of the six solutions should change from clear to a color of the rainbow!
3. Add approximately 100 mL of the 0.01 M hydrochloric acid solution to each beaker. The solutions will once again be clear.

Precise amounts of acid and base solutions are not important. Each addition of acid or base solution must neutralize the solution in the beaker and drive the pH in the opposite direction. All solutions can be poured into the large two-liter beaker. The resulting solution will be acidic and clear.

## Disposal

The resulting solution may be flushed down the drain with excess water according to Flinn Suggested Disposal Method #26b; consult your current *Flinn Scientific Catalog/Reference Manual*.

## Tips

- The indicators are dissolved in 95% ethyl alcohol. The alcohol will readily evaporate, leaving the indicator powder in the beaker—unseen to the observers of the demonstration.
- You may use drops of a more concentrated acid or base to change the solutions from colored to clear, or a more concentrated base to change the solutions from clear to colored.

## Discussion

The three indicators used in this lab, phenolphthalein, thymolphthalein, and p-nitrophenol, are colorless in acidic solution. In a basic solution, phenolphthalein is red, thymolphthalein is blue and *p*-nitrophenol is yellow. Any color in the spectrum is possible using these primary colors.

The solvent of each indicator solution added to the beakers readily evaporates, leaving only a residue of the indicators on the bottom of each beaker. Students will not see this step of the procedure. They will only see the pouring of the acid and base solutions and the color changes. This demonstration can be done as a “magic” show. It is a great demonstration to get your students to look beyond what is happening and get them asking questions about how the disappearing rainbow occurs. Initially students might think the clear acid–base solutions contain a magical indicator. Knowing there is no such thing, have your students propose explanations for how these six colors could appear.

## 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 5–8*

Content Standard A: Science as Inquiry

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

### *Content Standards: Grades 9–12*

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, chemical reactions

## Answers to Worksheet Discussion Questions

1. Draw a diagram of the set-up. Include the chemicals that were added to the beakers. For each beaker, list the original color of the first solution, the color change after the second solution was added, and the final color after the third solution was added.

Beaker #1  
Colorless  
Violet  
Colorless

Beaker #2  
Colorless  
Blue  
Colorless

Beaker #3  
Colorless  
Green  
Colorless

Beaker #4  
Colorless  
Yellow  
Colorless

Beaker #5  
Colorless  
Orange  
Colorless

Beaker #6  
Colorless  
Red  
Colorless

*First solution: 0.01 M Hydrochloric Acid*

*Second solution: 0.01 M Sodium Hydroxide*

*Third solution: 0.01 M Hydrochloric Acid*

2. Given that the two chemicals added to the beakers were an acid and a base, what kind of chemical must have already been present in the beakers to produce the color changes?

*An acid-base indicator must have been present in each beaker, since the solutions were different colors when a base, NaOH, was added than when an acid, HCl, was added.*

3. Three indicators are used in this demonstration: phenolphthalein, thymolphthalein, and *p*-nitrophenol. Phenolphthalein is an indicator that is colorless in an acidic solution but pink-red in a basic solution. Thymolphthalein is also colorless in an acid, but blue in a base, and *p*-nitrophenol is colorless in an acid and yellow in a base. What indicator or combination of indicators was responsible for the color change in each beaker?

*Beaker #1 – thymolphthalein and phenolphthalein*

*Beaker #2 – thymolphthalein*

*Beaker #3 – Thymolphthalein and p-nitrophenol*

*Beaker #4 – p-nitrophenol*

*Beaker #5 – phenolphthalein and p-nitrophenol*

*Beaker #6 – phenolphthalein*

## Reference

Shakhshiri, B. Z. *Chemical Demonstrations*; University of Wisconsin: Madison, WI, 1989; Vol. 3, pp. 41–46.

**The Disappearing Rainbow—Chemical Demonstration Kit is available from Flinn Scientific, Inc.**

Catalog No.	Description
AP8979	Disappearing Rainbow—Chemical Demonstration Kit

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

# Disappearing Rainbow Worksheet

## Discussion Questions

1. Draw a diagram of the set-up. Include the chemicals that were added to the beakers. For each beaker, list the original color of the first solution, the color change after the second solution was added, and the final color after the third solution was added.
  2. Given that the two chemicals added to the beakers were an acid and a base, what kind of chemical must have already been present in the beakers to produce the color changes?
  3. Three indicators are used in this demonstration: phenolphthalein, thymolphthalein, and *p*-nitrophenol. Phenolphthalein is an indicator that is colorless in an acidic solution but pink-red in a basic solution. Thymolphthalein is also colorless in an acid, but blue in a base, and *p*-nitrophenol is colorless in an acid and yellow in a base. What indicator or combination of indicators was responsible for the color change in each beaker?

## Dynamite Soap

### Introduction:

Hydrogen and oxygen soap bubbles have always been avoided because they are difficult to prepare. In this demonstration the process has been made simple, something you could call "hands on!"

### Safety:

All present should wear goggles, ear protectors. Remind students to cover their ears prior to noise generation.

### Procedure:

- 1.Fill a plastic bottle with water. Invert it in a bucket of water. Using water displacement, fill the bottle with two volumes of hydrogen and one volume of oxygen. If hydrogen and oxygen cylinders are not available, use H<sub>2</sub>O<sub>2</sub> and MnO<sub>2</sub> for oxygen generation and zinc and hydrochloric acid for the hydrogen generation, see Flinn ChemFax *Gas Generator Bottle.*, publication #955.00
- 2.You should fill the bottle with the gases till there is about one inch of water left in the bottle. Put a few drops of liquid detergent in the screw cap of the bottle. With the bottle still inverted under water replace the bottle cap.
- 3.Shake the capped bottle and generate a bottle full of suds.
4. All present should wear goggles and ear protectors. Remind those in the room to cup their ears.
- 5.Squeeze out a small quantity of suds into the victims cupped hands. BE SURE TO MOVE THE BOTTLE OUT OF THE AREA BEFORE DOING WHAT COMES NEXT OR YOU WILL HAVE THE DYNAMITE ROCKET-- NOT PLEASANT! Ignite the bubbles with a match.
- 6.Variations include using only hydrogen, hydrogen and oxygen in a 1:1 ratio & in a 1:2 ratio, etc..

### Discussion:

If you seal the bottles with tape the gases can be kept in the bottles for several days. The correct reaction ratio for hydrogen or oxygen can be determined by the loudness of the explosion.

The reaction is : 2 H<sub>2</sub> + O<sub>2</sub> --> 2 H<sub>2</sub>O + 484 KJ / 2 MOLES H<sub>2</sub>O

### Credit:

A teacher from Oak Park River Forest High School, Ralph Lee, presented this demo at a CHEM WEST meeting.

## I've Got a Crush on You

You know we live at the bottom of an ocean of air. Air exerts pressure on us and on everything around us. This atmospheric pressure is really very large. At sea level the air pressure is 14.7 Lbs/in<sup>2</sup> [p.s.i.], or 760 mm Hg, or 1 atmosphere [ATM], or, I am sure more meaningful to you, 101.3 kilopascal [kPa]. To see this air pressure first hand we have the following demo that can be done any number of ways.

### Safety:

Instructor and students participating in this demonstration, or viewing the demonstration, should wear proper PPE such as gloves, goggles, and aprons.

### Procedure:

- a.) Put about 10 ml of water into a pop [soda if from the east coast] can and heat it until it boils vigorously for at least a minute. Using beaker tongs, grasp the can near the bottom and turn it upside down quickly into a trough containing a couple of centimeters of cold water. The can should collapse instantaneously.
- b.) Take a 5 gallon can and put a dent in the side. This is necessary to get some cans to collapse. Put 100 to 200 ml of water into the 5 gallon can and heat it on a hot plate until it boils vigorously for at least five minutes. A cloud of condensed water vapor should escape from the mouth of the can for at least two minutes. Turn off the hot plate and using gloves put the top on the can as tight as you can. Take the can off the hot plate right away and tighten the top with pliers. You can set the can on the desk top and wait. Or, wearing gloves, set the can on a tub of ice water. Hold the can down with a stick on its' top or better yet do not hold it at all, just let it sit in the tub. The can should collapse within a few minutes. {If you hold the can with your hand, you stand the chance of it collapsing on your hand- not a very good feeling}
- c.) Take a 55-gallon drum and put a dent in the side. Add about 1 gallon of water to it. Hot water will really speed up the process. Heat it on a hot plate or camping stove until the water boils vigorously for at least ten minutes. If the can is painted the paint may really stink if heated on an open flame so have adequate ventilation. A cloud of condensed water vapor should escape from the mouth of the can for at least ten minutes. Coat the threads of the screw top with silicon sealing material to ensure an airtight fit. Turn off the hot plate and, using gloves, put the top on the can as tight as you can. Take the can off the hot plate and tighten the top more with pliers or special tool for tightening drum tops. You can set the can on the desktop and wait. Or with gloves on set the can on a tub of ice water. Hold the can down with a stick on the top or, better yet, do not hold it at all, just let it sit in the tub. The can should collapse within a few minutes. Try hitting it with a CO<sub>2</sub> extinguisher to cool it off. {If you hold the can with your hand, you stand the chance of it collapsing on your hand and arm- even less of a good feeling- talk about being depressed!} Video clip of this demo on my web site below.

### d.) I am using a vacuum pump here because they did not want an LPGs in McCormick Place.

### So what is happening?

Some of the air is forced out of the can as it heat up with the top off. The steam condenses, and then turns back into water with the top on. There is a less air in the can to take the steam's place. So there is much less gas on the inside of the can thus there is less pressure inside. The air is pushing on the outside with about 15 pounds per every square inch. With a 5-gallon can this is over 18,000 pounds. The cans are not designed to withstand such pressure differentials so the cans implode.

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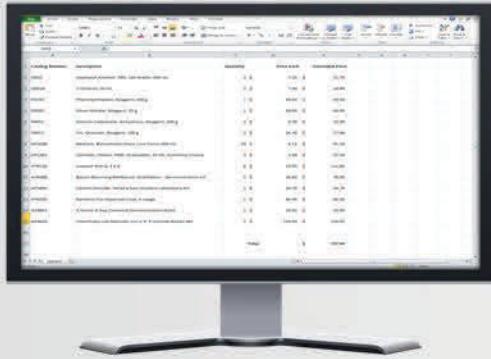
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