Instructor's Guide

to accompany the

Shakhashiri Chemical Demonstrations Videos

Dear Fellow Teacher

This Guide contains brief descriptions of each of the demonstrations in the *Shakhashiri Chemical Demonstrations Videos*. Also included are background information about the chemical reactions, the course topics with which the demonstrations would be appropriate, sources for further information, and sample questions for students to consider while watching the demonstration or afterward.

The demonstrations are in clusters as indicated in the Table of Contents. Each of the eight topical headings includes demonstrations that might also belong under another heading. You may choose to create your own clusters by using the demonstrations in any sequence you desire.

The suggested contexts and sample questions are just that, suggestions and samples. They are intended to assist you in using these demonstrations. All of these demonstrations can be used in situations other than those suggested and also to communicate concepts that are not articulated in the sample exercises. You are invited to adapt these and other demonstrations into your course design. They can be used as a focus for your presentations and discussions. Our lecture presentations usually include one or more demonstrations.

We urge you to consider the following, as appropriate:

1. Give the students the question handout and play back the video *twice*. During the first playback ask the students to watch. Then, ask the students to read and answer as many questions as they can. Finally, show the tape the second time so the students can complete their answers.

2. Turn off the sound during playback and provide your own narration.

3. Use the still-frame feature of your DVD player frequently to show appropriate details.

4. Take advantage of the slow-motion playback features we have incorporated in some of the demonstrations.

5. Refer students who may need additional assistance to:

   Workbook for General Chemistry, 3rd edition, Bassam Z. Shakhashiri and Rodney Schreiner: Stipes Publishing, 2005.

6. Do the demonstrations yourself!

We sincerely hope that these videos will stimulate you to include more chemical demonstrations in your courses to increase your students' experience of the phenomena of chemistry.

Bassam Z. Shakhashiri
Rodney Schreiner

March 2008
Department of Chemistry
University of Wisconsin-Madison

Contents

Acid-Base Indicator from Red Cabbage

Juice is prepared from a red cabbage. This juice is mixed with an assortment of household products. The mixtures turn a variety of colors. The colors correspond to the acidity of the household products.

Running time = 5:44

The pigments in red cabbage juice can be extracted into water. The extract functions as a pH indicator, undergoing several color changes over the range of 1 to 10 pH units. In highly acidic media, the extract is red, and in very alkaline media it is green. Between these extremes, the red cabbage juice turns purple, blue, and blue-green. The initial color of the extracted juice depends on the pH of the water used for the extraction. If distilled water is used, the initial color is purple.

This demonstration is suitable for discussions of acids and bases and of acid-base indicators.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 50-57, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

HANDOUT FOR
ACID-BASE INDICATOR FROM RED CABBAGE

In this demonstration, red cabbage juice is mixed with a variety of household products. The color of the cabbage juice changes in each mixture. In the chart below, record your observation of the color in each mixture. Then, indicate whether the household product is acidic or basic.

Product	Color of cabbage juice	Acidic or basic
vinegar		acidic
ammonia		basic
Sprite		
Tide		
Vanish		
milk of magnesia		
Drāno		
baking soda		

Reactions of Carbon Dioxide in Aqueous Solutions

Six pairs of tall cylinders are arranged in a row. Each pair of cylinders contains a different colored liquid. Pieces of solid carbon dioxide (dry ice) are dropped into one cylinder of each pair. In these cylinders, bubbles rise through the liquid, and fog forms at the top. Gradually, the colors of the liquids change in the cylinders to which dry ice was added.

Running time = 3:28

This demonstration involves a series of attractive color changes in the cylinders. These color changes are produced by the pH indicators in the cylinders. As bubbles of carbon dioxide gas from the subliming dry ice rise through the solutions, some carbon dioxide dissolves in the solutions. This decreases the pH of the solutions. The carbon dioxide gas bubbling through the solutions is very cold. This cold gas causes water vapor to condense, forming a fog at the tops of the cylinders.

The demonstration may be used in discussions of the properties of carbon dioxide and its aqueous solutions, acid-base indicators, and water vapor and fog.

Students should be prepared to note any color changes that occur in the cylinders, and the time it takes for the color changes to occur in the different cylinders.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 114-120, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 3-26, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Why was the dry ice added to every other cylinder, rather than to all of them?

2. Why did the solution in some cylinders take longer to change color than others?

3. How is the time it takes for the color to change related to the amount of base in the cylinder?

4. How is the time it takes for the color to change related to the particular indicator in the cylinder?

5. Why does the gas that escapes from the cylinders drift downward?

6. What household materials might have been added to the cylinders to change the color of the solutions in them?

Ammonia Fountain

An inverted round-bottomed flask rests on a ring stand. The flask is sealed with a 2-holed stopper through which two glass tubes are inserted. One of these is short and has a rubber bulb attached to the end. The other is long and extends to the bottom of an upright round-bottomed flask filled with a colorless liquid. When the rubber bulb is squeezed, a small amount of water is injected into the inverted flask. Immediately the liquid in the upright flask rushes up the tube and into the inverted flask. The rushing liquid turns red and forms a fountain as it enters the inverted flask.

Running time = 1:10

The inverted flask is filled with ammonia gas. When a small amount of water is injected into the flask, some of the ammonia dissolves in the water. This reduces the pressure of the ammonia in the flask, allowing the atmosphere to push water from the lower flask into the ammonia-filled flask. As the water enters the flask, more ammonia dissolves, further reducing the pressure and increasing the rate at which water flows up the tube. The water contains phenolphthalein indicator. This indicator is colorless in neutral solutions and red in basic solutions. The solution becomes basic as ammonia dissolves in it.

The demonstration is effective in discussions of ammonia and its properties, acid-base indicators, gas solubility, and the effects of atmospheric pressure.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 205-210, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the purpose of squirting a little water into the inverted flask?

2. Why does the liquid move into the upper flask?

3. Why isn't the lower flask sealed with a stopper?

4. Why does the liquid change color when it enters the upper flask?

Demonstration 4

Multiple Ammonia Fountain

Four inverted flasks are arranged one above another and connected in series with tubing. A tube extends from the lowest flask into a beaker of colored liquid. A small amount of water is injected into the top flask. The colored liquid immediately rushes into the lowest flask and changes color. The liquid flows sequentially through the first three flasks and goes through a series of color changes. At the end, each flask is filled with a different colored liquid.

Running time = 1:45

Striking color changes occur as the liquid rushes through the flasks. The color changes result from pH changes and their effect on universal indicator as ammonia dissolves in the water. This demonstration is effective in discussions of the properties of ammonia, acid-base indicators, and solubility.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 92-95, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

HANDOUT FOR MULTIPLE AMMONIA FOUNTAIN

1. Answer the following questions while watching the demonstration:

   a. What is the color of the liquid in the beaker?

   b. What is in the four flasks?

   c. How is the process started?

   d. What color does the liquid turn just as it enters the bottom flask?

   e. What are the colors of the four flasks when the process is complete?

   top flask:

   second flask:

   third flask:

   bottom flask:

2. The liquid is water dyed with universal pH indicator. The colors of the indicator at various pH values are:

pH:	2	4	6	8	10
color:	red	orange	yellow	green	purple

   What is the pH of the liquid in each flask?

   top flask:

   second flask:

   third flask:

   bottom flask:

Oxidation of Luminol

A clear, colorless liquid and a clear, blue liquid are poured into a funnel at the top of a coil of glass tubing. With the lights off, the mixture of the two liquids glows bright blue as it flows through the funnel.

Running time = 0:59

This is an example of a chemiluminescent reaction, the oxidation of luminol by hydrogen peroxide. A product of the reaction is formed in an excited state, and this excited state decays to the ground state by emitting light. The oxidation is catalyzed by copper sulfate, which is responsible for the blue color of one of the solutions.

This demonstration can be used in presentations of chemiluminescence, excited states, energy transfers, and molecular orbital theory.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 156-167, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What forms of energy are released by this reaction?

2. Why is the reaction carried out in the dark?

3. Why does the glow disappear when the lights are turned on?

4. Describe some other forms of chemiluminescence.

Chemiluminescent Ammonia Fountain

An inverted round-bottomed flask sealed with a two-holed stopper rests on a ring stand. Through one hole of the stopper is a short piece of glass tubing connected to a rubber bulb. Through the other hole of the stopper, a long piece of glass tubing is connected to a Y connector. From the Y connector, clear plastic tubing extends into two flasks, one containing a colorless liquid, the other a blue liquid. When the rubber bulb is squeezed, a small amount of water enters the round-bottomed flask. Immediately the liquids from the flasks are driven into the round bottom flask. When the lights are turned off a blue glow is visible where the two liquids mix and inside the inverted round-bottomed flask.

Running time = 1:27

Light is emitted when the two liquids mix. The reaction is the chemiluminescent oxidation of luminol by hydrogen peroxide in a basic mixture. The demonstration is an adaptation of the ammonia fountain. The inverted round-bottomed flask is filled with ammonia gas. When a small amount of water is injected into the flask, some of the ammonia dissolves, reducing the pressure of the gas in the flask. The reduced pressure allows the atmosphere to push the liquids into the round-bottomed flask. As the liquids enter the flask, more ammonia dissolves, further reducing the pressure. This demonstration may be used in discussions of chemiluminescent reactions, or to show the solubility of ammonia.

If one is emphasizing the chemiluminescent nature of the reaction, have students review the solubility of ammonia in water, and what effect that has on the pressure outside the flask where the ammonia dissolves.

FURTHER INFORMATION

Nicholas C. Thomas, *Journal of Chemical Education*, Volume 67, page 339 (1990).

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 156-167, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the purpose of the squirting the liquid into the inverted flask?

2. Why do the liquids move from the lower flasks up the tube and into the inverted flask?

3. Why aren't the tubes leading into the lower flask sealed with a one-hole stopper?

4. What forms of energy did you observe being produced in this demonstration?

5. Why is the reaction carried out in the dark?

Reaction of Carbon Disulfide and Nitrogen(II) Oxide

A small amount of liquid carbon disulfide is injected into a 1¼-meter glass tube containing nitrogen(II) oxide gas. The lights are turned off, a match is dropped into the tube, and a blue flash rushes through the tube. The rapid reaction is also shown in slow motion. When the lights are turned on, a yellow deposit is visible inside the tube.

Running time = 1:57

The reaction between CS_2 vapor and NO gas is explosive. It produces a bright blue flash and barking sound. The visual effect is particularly attractive in slow motion, where the flash can be followed down the tube. The demonstration is suitable for discussions of chemical change and its effects, reactions of gases, exothermic reactions, and chemiluminescent reactions.

Both physical and chemical properties of the carbon disulfide and nitrogen(II) oxide reactants are displayed in this demonstration. Students should look for these properties.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 117-120, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What evidence reveals that the glass tube contains more than air?

2. Why did the stopper pop off from the tube?

3. What evidence is there that a chemical change has taken place after the match is dropped into the tube?

4. What forms of energy do you observe being released when the mixture is ignited?

5. What is the identity of the yellow substance inside the tube after the reaction?

Demonstration 8

Photochemical Bleaching of Thionine

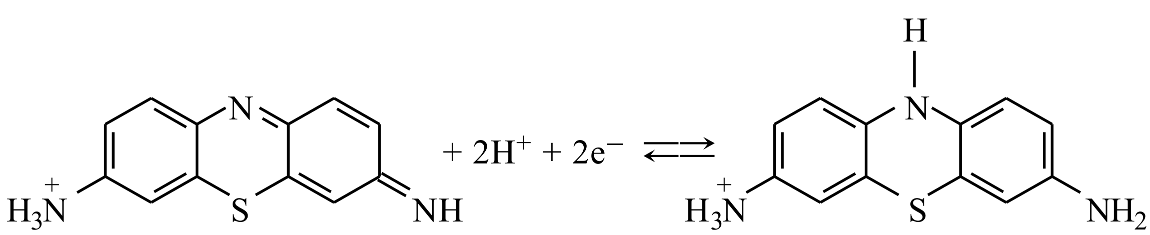
The beam of a slide projector irradiates a tube containing a blue solution. Where the light strikes the solution, the solution first appears magenta, and then becomes colorless. When the light is extinguished, the bleached solution becomes blue again.

Running time = 1:11

The blue solution contains a mixture of thionine, iron(II) sulfate, and sulfuric acid. Thionine is blue. Thionine is also an oxidizing agent, but it not sufficiently powerful to oxidize iron(II) ions. However, when it absorbs light, it becomes a more potent oxidizing agent, and it oxidizes iron(II) to iron(III). Simultaneously, thionine is reduced to a colorless form. When the light is extinguished, iron(III) oxidizes the colorless, reduced form of thionine back to the blue form. The magenta that appears when the light beam first strikes the solution results from a red fluorescence by thionine. The reduced form of thionine does not emit red light.

This demonstration is suitable for discussions of photochemical processes, oxidation-reduction reactions, and factors affecting equilibria.

Thionine is a blue dye related to the common pH indicator, methylene blue. It is a salt, usually chloride or acetate, of 7-amino-3-imino-3H-2-phenothiazine. Thionine can undergo a 2-electron reduction to a colorless form.



oxidized
(blue)

reduced
(colorless)

The reduction potential of thionine is about 0.4 volts at a pH of 2. The standard reduction potential of Fe^{3+} to Fe^{2+} is 0.77 volts. Therefore, the reduction potential of thionine in its ground state is too small for thionine to be reduced by Fe^{2+} to a significant extent. The reduction potential of Fe^{3+} is great enough to oxidize the reduced form of thionine virtually completely. Therefore, in low light intensities, the position of the thionine reduction equilibrium lies to the left, and the solution is blue. However, when thionine is excited by the absorption of visible light, the reduction potential of thionine is increased, and the excited molecules can be reduced by Fe^{2+}. The absorption of light shifts the equilibrium to the right, and, if the intensity of the light is great enough, the blue color of thionine disappears altogether.

FURTHER INFORMATION

L. H. Heidt, *Journal of Chemical Education*, Volume 26, pages 525-526 (1949).

M. D. Archer, *Journal of Applied Electrochemistry*, Volume 5, pages 26-27 (1975).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Does thionine absorb visible light? How can you tell?

2. Does thionine react with iron(II) sulfate at low light levels?

3. In bright light, thionine reacts with iron(II) ions to form reduced thionine and iron(III) ions. This is a reversible equilibrium reaction. Addition of light energy shifts the equilibrium in which direction? When the light energy is removed, which way does the equilibrium shift? Under what conditions would the color of the mixture be the deepest?

4. How would adding iron(III) sulfate to the mixture affect the rate at which the solution returns to blue after the light is extinguished?

Demonstration 9

Reduction of Copper(II) Oxide with Hydrogen

A piece of copper metal is heated strongly in a burner flame, and its surface turns black. After the burner is turned off, but while the copper piece is still hot, a funnel carrying hydrogen gas from a high-pressure cylinder is placed over the blackened copper. The black coating disappears, revealing the original coppery color. When the funnel is removed, the black coating reappears. This cycle is repeated several times.

Running time = 2:19

The demonstration shows the oxidation of hot copper by elemental oxygen from the air forming black copper(II) oxide, and the reduction of hot copper(II) oxide by elemental hydrogen, forming copper and water vapor. Students should realize that appearance and disappearance of the black copper(II) oxide coating is the result of two different chemical reactions, not a single, reversible chemical reaction.

This demonstration can be used in presentations of oxidation-reduction reactions, the atmosphere, oxygen, and hydrogen.

FURTHER INFORMATION

J. H. Walton, *Journal of Chemical Education*, Volume 19, pages 453-454 (1942).

***CAUTION**: Contrary to the instructions in this reference, the gas burner should be extinguished before the hydrogen cylinder is opened, to avoid the risk of a hydrogen explosion.*

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What evidence of a chemical reaction do you see when copper metal is heated in the flame?

2. What is the likely other reactant in the reaction besides copper?

3. Write a balanced chemical equation for the reaction.

4. What substance is oxidized, what substance is reduced, what is the oxidizing agent, what is the reducing agent in this reaction?

5. Why must the burner be turned off before the funnel containing hydrogen gas is placed over the copper?

6. What evidence of a chemical reaction is there when the darkened copper is bathed in gaseous hydrogen?

7. What product of this reaction is visible? What is the other product that cannot be seen?

8. Write a balanced equation for the reaction that occurs when the hydrogen-filled funnel is placed over the blackened copper.

9. In this second reaction, what substance is oxidized, what substance is reduced, what is the oxidizing agent, what is the reducing agent?

Growing Metal Crystals

A solution of tin(II) chloride is placed in a shallow dish containing two electrodes. The electrodes are connected through a switch to four D cells. When the switch is closed, needle shaped crystals form at one of the electrodes. When the polarity of the switch is reversed, the first crystals shrink, and new ones grow on the other electrode.

Running time = 1:53

When a direct current is passed through a solution of tin(II) chloride, at one electrode, the cathode, tin(II) is reduced to tin metal. The reduced metal forms into long needles. At the anode, chlorine gas is produced. When the polarity of the electrodes is reversed, the tin already formed is oxidized back to tin(II), while tin metal is reduced at the new cathode.

This demonstration is suitable for discussions of electrochemistry, electrolysis, and oxidation-reduction reactions.

FURTHER INFORMATION

> *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Volume 4, Demonstration 11.17, Bassam Z. Shakhashiri, University of Wisconsin Press (in press).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Write balanced equations for the reactions that occur at each electrode.

2. What is the total voltage applied across the electrodes?

3. What is the minimum voltage that must be applied across the electrodes to produce the reaction?

Copper to Silver to Gold

A copper coin is placed on top of granular zinc in a beaker of boiling sodium hydroxide solution. After the coin is removed and rinsed, it seems to have changed to a silvery metal. The coin is heated in the flame of a burner for several seconds and cooled in water. The coin now appears golden.

Running time = 2:26

Although this demonstration appears to show the transmutation of copper to silver to gold, this is not the answer to the alchemists' dream. The silvery coating on the copper coin is zinc metal. The golden coating is brass, which forms when the zinc coating is alloyed with the underlying copper. The demonstration shows how brass, an alloy of zinc and copper, is formed. The coin used in this demonstration must be a pre-1983 cent, because newer cents are copper-coated zinc, which would melt in the burner flame.

This demonstration is suitable for discussions of alloys, reactions of metals, physical and chemical changes, and diffusion processes.

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What happens to the pennies when they are placed on the powdered zinc in the sodium hydroxide solution?

2. Why are the pennies placed in the beaker of water after they are removed from the sodium hydroxide and zinc mixture?

3. What metal coats the pennies after they are rinsed?

4. How do we know that some type of diffusion process takes place while the silver colored penny is held in the Bunsen burner flame?

5. What is the term given to a mixture of metals?

Precipitates and Complexes of Nickel(II)

A variety of ligand solutions is added to green aqueous nickel(II) sulfate solutions. The ligands are ammonia, ethylenediamine, dimethylglyoxime, and cyanide ions. Upon addition of each ligand solution to the nickel(II) sulfate solutions, the color changes to a color characteristic of the nickel complex formed between nickel(II) ions and the added ligand.

Running time = 4:47

This demonstration shows a range of dramatic color changes when colorless solutions are mixed with a clear, green solution. This demonstration can be used in discussions of the properties of transition metal complexes. Among the concepts involved are colors, stoichiometry, formation constants, and nomenclature of complex ions.

Students should be encouraged to note the initial color of the nickel(II) sulfate solution and the identities of the substances added to the nickel(II) sulfate solutions and to record the changes that result upon each addition. This may be accomplished by distributing the sample handout on the next page before showing the demonstration and instructing students to record their observations in the second column of the chart. After the demonstration, students may compare their observations to the information provided in the table at the bottom of the page. From this information they can determine the identity of the complex that was formed at each addition. They may record their conclusions in the third column of the chart.

The ligands used are ammonia, NH_3, ethylenediamine, $NH_2CH_2CH_2NH_2$ (abbreviated “en”), dimethylglyoxime, $HONC(CH_3)C(CH_3)NOH$ (abbreviated “Hdmg”), and cyanide ion, CN^-.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 299-306, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

Workbook for General Chemistry Audio-Tape Lessons, 2nd edition, pages 329-377, Bassam Z. Shakhashiri, Rodney Schreiner, and Phyllis Anderson Meyer, Saunders College Publishing (1981).

Bassam Z. Shakhashiri, Glen E. Dirreen, and Fred Juergens, *Journal of Chemical Education*, Volume 57, pages 900-901 (1980).

SAMPLE STUDENT EXERCISES AFTER THE DEMONSTRATION

1. Write a net ionic equation for the equilibrium that forms $[Ni(H_2O)_6]^{2+}$ from its constituents. Write a similar equation for each complex formed in this demonstration.

2. Write the K_f expressions for the equilibrium reactions of question 1.

3. Is the value of K_f for $[Ni(H_2O)_6]^{2+}$ larger or smaller than K_f for $[Ni(NH_3)_6]^{2+}$?

4. Methylamine has a molecular formula of CH_3NH_2. Write the net ionic equation for the reaction that occurs when an excess of methylamine is added to a solution containing $[Ni(H_2O)_6]^{2+}$. What is the most likely color of the resulting solution?

HANDOUT FOR PRECIPITATES AND COMPLEXES OF NICKEL(II)

All beakers initially contain 320 mL of aqueous 0.06M $NiSO_4$ solution. This solution is pale green, the characteristic color of the hexaaquanickel(II) ion, $[Ni(H_2O)_6]^{2+}$.

Ligand solutions are added to the $NiSO_4$ solutions. The ligands are ammonia (NH_3), ethylenediamine (abbreviated "en"), dimethylglyoxime (abbreviated "Hdmg"), and cyanide ions (CN^-). During the demonstration, record your observations in the second column of the chart. After the demonstration, use your observations and the information in the table at the bottom of the page to identify the nickel complex formed in each beaker. Write the formula of the complex in the third column of the chart.

Solution added	Observations	Nickel complex
40 mL 5M NH_3(aq)		
5 mL 25% en		
10 mL 25% en		
15 mL 25% en		
25 mL 1% Hdmg		
80 mL 1M KCN		

Colors, formation constants, and geometries of Nickel(II) complexes.

Formula	Color	Formation constant	Geometry
$[Ni(H_2O)_6]^{2+}$	green	—	octahedral
$[Ni(NH_3)_6]^{2+}$	blue	6.5×10^8	octahedral
$[Ni(H_2O)_4(en)]^{2+}$	light blue	3.6×10^7	octahedral
$[Ni(H_2O)_2(en)_2]^{2+}$	blue	5.6×10^{13}	octahedral
$[Ni(en)_3]^{2+}$	purple	3.3×10^{18}	octahedral
$[Ni(dmg)_2]$	red	4.2×10^{17}	square planar
$[Ni(CN)_4]^{2-}$	yellow	3.1×10^{30}	square planar

Precipitates and Complexes of Silver(I)

A sequence of clear colorless solutions is added one by one to a stirred clear colorless solution in a beaker. Precipitates of various colors appear, change color, disappear and reappear upon each addition.

Running time = 3:57

The initial solution contains silver ions. Various precipitating and complexing agents of silver ion are added in a sequence designed to cause precipitates to appear and dissolve. Each precipitate and complex is more stable than the previous one. The most insoluble is the last, silver sulfide.

This sequence is applicable to discussions of precipitation reactions, complex-ion formation, chemical equilibria, solubility products, and formation constants. The demonstration and its discussion and explanation may easily occupy an entire class period.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 307-313, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

Workbook for General Chemistry Audio-Tape Lessons, 2nd edition, pages 329-377, Bassam Z. Shakhashiri, Rodney Schreiner, and Phyllis Anderson Meyer, Saunders College Publishing (1981).

HANDOUT FOR
REACTIONS OF THE SILVER ION

In this experiment, a series of solutions is added, one by one, to a beaker containing distilled water. After each addition, record your observations in the chart below. Use the information in the tables on the next page to determine the identity of the substance formed after each addition.

Solution added	Observations	Substance formed
1. $AgNO_3$		
2. $NaHCO_3$		
3. NaOH		$Ag_2O \cdot H_2O$
4. NaCl		
5. NH_3		
6. NaBr		
7. $Na_2S_2O_3$		$Ag(S_2O_3)_2^{3-}$
8. KI		
9. KCN		
10. Na_2S		

Solubility of Silver Salts in Aqueous Solution

Substance	Color	K_{sp} at 25°C	Solubility, mol Ag^+/L
Ag_2CO_3	white	8.1×10^{-12}	2.5×10^{-4}
$Ag_2O{\bullet}H_2O$	brown	1.8×10^{-15} *	2.1×10^{-4}
AgCl	white	1.8×10^{-10}	1.3×10^{-5}
AgBr	pale yellow	8.0×10^{-13}	8.9×10^{-7}
AgI	yellow	4.5×10^{-17}	6.7×10^{-9}
Ag_2S	black	2.0×10^{-50}	3.4×10^{-17}

*This K_{sp} is for the reaction $Ag_2O \cdot H_2O \rightleftharpoons 2Ag^+ + 2OH^-$.

Formation Constants of Silver Complexes

Complex	Formation Constant
$Ag(NH_3)_2^+$	1.6×10^7
$Ag(S_2O_3)_2^{3-}$	4.3×10^{12}
$Ag(CN)_2^-$	1.0×10^{20}

EXERCISES

1. Write a net ionic equation for each reaction that occurs in steps 2 through 10.

2. Use the information in the tables on the other side of the sheet to calculate the value of the equilibrium constants for the reactions of steps 3, 5, 7, and 9.

3. What is the minimum concentration of CN^- required to dissolve 0.15 mole of AgI per liter?

4. How many moles of AgBr will dissolve in a 2.0M solution of $Na_2S_2O_3$?

5. How many moles of AgBr will dissolve in a 2.0M solution of NH_3?

Orange Tornado

Aqueous potassium iodide and aqueous mercury(II) nitrate are combined in a large beaker of water on a magnetic stirrer. An orange precipitate of mercury(II) iodide forms in the vortex, forming an "orange tornado." The orange precipitate dissolves and reforms as more potassium iodide and mercury(II) nitrate solutions are added.

Running time = 3:51

When potassium iodide solution is added to mercury(II) nitrate solution, orange mercury(II) iodide precipitates. If an excess of potassium iodide is added, the orange precipitate dissolves. The precipitate dissolves because soluble complex ions form between mercury(II) and iodide. If more mercury(II) is added, the precipitate reappears because the ratio of iodide ions to mercury(II) ions has been reduced.

This demonstration is suitable for discussions of solubility and complex ion formation.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 271-279, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

Workbook for General Chemistry Audio-Tape Lessons, 2nd edition, pages 329-377, Bassam Z. Shakhashiri, Rodney Schreiner, and Phyllis Anderson Meyer, Saunders College Publishing (1981).

SUGGESTED QUESTIONS AFTER THE DEMONSTRATION

1. What two solutions are combined in the beaker?
2. What is responsible for the orange color in the vortex?
3. The addition of which solution causes the orange color to disappear?
4. The addition of which solution causes the orange color to reappear?

Formation of Nylon

A colorless liquid is poured onto another colorless liquid in a small dish. The liquids do not mix. With a pair of forceps, a solid white strand of material is pulled from the interface between the two liquids. The strand is drawn up and attached to a horizontal shaft. As the shaft is rotated, a continuous strand is drawn from the interface. This process continues for a substantial period of time.

Running time = 2:16

The white strand is nylon 6-10, made from sebacoyl chloride dissolved in hexane in the top layer, and 1,6-diaminohexane dissolved in aqueous sodium hydroxide in the bottom layer. The reaction of the acid chloride with the amine produces an amide linkage, as well as hydrochloric acid, which is neutralized by the base. Because both sebacoyl chloride and 1,6-diaminohexane are difunctional, a polymeric amide is formed. The strand will continue to form at the interface as long as both solutions contain reactants.

This demonstration can be used to show the formation of an amide from an acid chloride and an amine, the formation of a condensation polymer, or the production of a polymer strand.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 213-215, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Why is it important that the two liquids do not mix? What would happen if they did?

2. What is the evidence that a chemical reaction is occurring?

3. Write a structural formula for each reactant.

4. Write an equation showing the reaction of an acid chloride with an amine.

5. Write a structural formula for several repeating units of the polymer.

5. Why is a continuous strand of material formed rather than small particles?

Demonstration 16

Polyurethane and Polystyrene Foams

Two viscous brown liquids are mixed in a beaker, and after a short time, the volume increases to many times the original size, overflowing the beaker. Acetone is poured into a foamed polystyrene cup held over another beaker, and the bottom falls out of the cup, dumping its contents into the beaker. The cup is dropped into the beaker and it collapses. Packing “peanuts” are dumped into this beaker, and they, too, collapse.

Running time = 3:03

The foam produced in the first beaker is polyurethane foam. Polyurethane is formed by the polymerization of a diisocyanate and a dialcohol. One of the components also contains a liquid with a low boiling point. The polymerization reaction is exothermic, and the heat it releases causes the liquid to vaporize. The vapor produces bubbles in the polymer, resulting in a foam. The volatile liquid (blowing agent) in this demonstration is a chlorfluorocarbon. When polyurethane foam is produced commercially, other blowing agents, such as carbon dioxide or nitrogen gases, are used to produce the bubbles.

Polystyrene foam is produced commercially in a similar fashion. Acetone is poured into a cup made of foamed polystyrene, the foam collapses. The polystyrene does not dissolve in acetone but instead absorbs acetone. The absorbed acetone softens the polymer, and the bubbles in the foam break, allowing the gas trapped in the bubbles to escape.

Both physical and chemical properties of polymers and foams are displayed in this demonstration. It is suitable in discussions of polymers and their properties.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 348-350, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SUGGESTED QUESTIONS DURING THE DEMONSTRATION

1. How many liquids are poured into the first beaker? Describe the appearance of these liquids. Estimate the volume of each liquid.

2. Estimate the volume of the foam produced in the reaction.

4. Do the cup and packing "peanuts" dissolve in acetone?

3. Estimate the volume of packing "peanuts" added to the second beaker. What is the approximate volume of the residue in the beaker?

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is a foam? How is a foam produced?

2. What produces the increase in volume after the two liquids are mixed?

3. Why does the volume of the packing peanuts decrease when they are mixed with acetone?

Demonstration 17

Rod-Climbing Polymer

A paddle stirrer is operated in colored water. The liquid moves away from the stirrer, and a vortex forms. The paddle is operated in a polymer solution. The liquid moves toward the stirrer and climbs the shaft of the stirrer.

Running time = 1:37

Most common liquids behave as the water in this demonstration: they move in the same direction as an applied stress. The paddle pushes out the water and the water moves away. Some liquids, however, move in a direction perpendicular to the stress. The paddle pushes out and the liquid moves up. This kind of behavior in a liquid indicates that there is relatively long-range order in the liquid. This order is provided by very large polymer molecules. In this demonstration, the liquid is poly(acrylamide) dissolved in glycerine. This behavior is also exhibited by bread dough in an electric mixer, where the dough climbs the shafts of the beaters. In bread dough, the polymer is starch (gluten) from flour.

This demonstration is appropriate for discussions of liquid properties and polymers.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 335-336, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Which liquid is more viscous, the one that moves away from the stirrer, or the one that climbs the stirrer?

2. When you stir a beverage, such as coffee, does the liquid move to the center or to the edge of the container?

3. What liquids, besides bread dough, move to the center of the container when they are stirred?

A Superabsorbent Polymer

One beaker of a pair contains a green liquid. A small amount of white powder is poured into the other beaker. The colored liquid is poured into the beaker with the powder. The mixture cannot be poured back into the first beaker, because a gel has formed. Some table salt is sprinkled onto the gel and the mixture is stirred. As the salt mixes with the gel, the gel liquefies and can be poured into the first beaker.

Running time = 1:40

The white powder is a polyacrylamide polymer. This polymer has a structure similar to polyethylene, but with extensive branching and with a hydrogen on every other carbon replaced with an amide, $–CONH_2$. This polymer has a strong affinity for water and can absorb many times its volume in water, as this demonstration shows. The water can be displaced from the polymer by adding ions to the gel. In the demonstration, sodium chloride is the source of ions.

This polymer is used in situations where water must be retained. Common applications include use in agriculture as seed coatings and in consumer products such as superabsorbent diapers. It is also used to dry diesel and aviation fuels.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 368-371, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. How quickly does the gel form after the polymer and water are mixed?

2. How does the volume of the gel compare to that of the polymer?

3. What breaks down the structure of the polymeric gel?

Plastics between Polarizers

A piece of polyethylene film is held between crossed polarizing filters and stretched. Brightly colored stretch marks appear in the film. A piece of acrylic sheet in a C clamp is placed between the filters. When the C clamp is tightened, bands of various colors appear in the acrylic sheet.

Running time = 3:21

When light passes through a polarizing filter, the light is polarized into a single plane. When the polarized light strikes a second filter with its polarizing plane perpendicular to that of the first, none of the light passes through. When a substance that rotates the plane polarized light is placed between the filters, light can now pass through the second filter. When the piece of polyethylene film is stretched, it rotates some colors of polarized light by a different degree than others. When it is stretched between crossed polarizers, light of various colors passes through the second polarizer. A similar effect is produced by the acrylic sheet, when it is stressed between the crossed polarizers.

This demonstration is suitable for discussions of properties of polymers and polarized light.

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What happens to the intensity of the light that passes through the polarizers, when one of them is rotated? Does the light change color when there is only air between them?

2. What happens when the polyethylene film is rotated between the crossed polarizers? Do colors appear?

3. What must be done to the polyethylene film and to the piece of acrylic to produce colors?

Explosive Reaction of Hydrogen and Oxygen

Four buoyant balloons are tethered to the bench, two yellow and two red. Each balloon is ignited with a torch, the first of each pair in room light and the second in the dark. All produce a loud explosion and a fire ball. However, the red balloons produce smaller fireballs and louder explosions. The explosions are shown in slow motion.

Running time = 4:42

The yellow balloons contain hydrogen gas, and the red balloons contain a mixture of hydrogen and oxygen. Hydrogen and oxygen mixtures are explosive over a wide range of compositions. When the hydrogen-filled balloon is ignited, the balloon breaks, and the hydrogen mixes with oxygen in the air forming an explosive mixture. Because the hydrogen and oxygen must mix before an explosion can occur, the explosion is relatively slow and diffuse. When the balloon filled with a mixture of hydrogen and oxygen is ignited, the explosion occurs rapidly in a compact area, producing a louder, shorter bang.

This demonstration is suitable for discussions of the factors affecting the rate of reactions, the properties of hydrogen and oxygen, and combustion reactions.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 106-116, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Which produced the louder bang, the red balloons or the yellow balloons? Which balloons produce the smaller fireball? Which produced the faster explosion?

2. What are the contents of each balloon?

3. Why was the size of the fireball smaller and the reaction faster with the mixture of gases than with pure hydrogen?

Pop-bottle Combustion of Hydrogen

The mouth of a carbonated-beverage ("pop") bottle covered with mesh is held near the flame of a burner. A short, sharp pop is produced and a flame darts from the mouth of the bottle. The demonstration is repeated with a second bottle in the dark, when the flame is much more apparent. In slow motion, the path of the flame is easily followed.

Running time = 1:29

This demonstration shows the explosive reaction between hydrogen and oxygen. The blue flash is most easily seen in the dark and in slow motion. Students should be alerted to watch the bottle and note the color of the flame and the path it follows as it burns.

The demonstration is suitable for discussions of combustion reactions in particular, chemical changes in general, and rates of reactions occurring in the gas phase.

FURTHER INFORMATION

> *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Volume 1, pages 106-112, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Why is the bottle stoppered before starting the reaction?

2. Why is the bottle covered with the plastic mesh?

3. What forms of energy are released when the mixture is ignited?

4. What ratio of hydrogen to oxygen would give the fastest reaction and loudest explosion?

Ignition of Alcohol Vapor

Two nails are inserted though the sides of a plastic bottle so their tips are separated by a small gap. A little ethyl alcohol is squirted into the plastic bottle, which is then corked and shaken. A Tesla coil, which produces an electric spark, is held near the head of one of the nails. The cork is ejected explosively from the mouth of the bottle. Slow motion shows a blue flame leaping from the mouth of the bottle after the cork is ejected.

Running time = 2:50

The Tesla coil produces a high-voltage electric spark between the nail points inside the plastic bottle. Inside the bottle is a mixture of alcohol vapor and air. The mixture of ethanol vapor and air is flammable, and the spark initiates the combustion reaction. The reaction is exothermic, and it produces both carbon dioxide and water vapor at high temperature. The production of hot gases causes the pressure inside the bottle to increase suddenly. This pressure increase expels the cork from the neck of the bottle.

This demonstration is suitable when discussing combustion reactions in general; when discussing properties of organic compounds, to contrast the slow, quiet burning of ethanol in a burner to the rapid, noisy explosion of ethanol vapor in air; or to illustrate the conversion of chemical energy to mechanical energy.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 216-219, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What evidence reveals that a chemical reaction has taken place?

2. What forms of energy are released when the mixture is ignited?

3. In what kind of motor is this type of energy release harnessed to provide useful work?

4. Write the balanced equation for this chemical reaction.

Reaction of Zinc and Sulfur

A small pile of sulfur and zinc powder is touched with a red-hot metal rod, and the mixture immediately ignites, producing sparks and a cloud of smoke.

Running time = 0:44

The hot metal triggers the reaction between zinc and sulfur. The heat produced by the reaction ignites the rest of the mixture. Some of the sulfur combines with oxygen of the air, producing sulfur dioxide. Some of the zinc, too, reacts with atmospheric oxygen, producing zinc oxide. The sparks are pieces of burning zinc, and the smoke contains sulfur dioxide and particles of zinc oxide and zinc sulfide.

The demonstration is pertinent to discussions of oxidation, exothermic reactions, and factors affecting reaction rates.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 53-54, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Do zinc and sulfur react at room temperature?

2. What must be done to start the reaction?

3. Why do all of the zinc and sulfur react, even though only a small portion of the mixture is heated?

Demonstration 24

Reaction of Potassium Permanganate and Glycerine

Several milliliters of glycerine are poured onto a pile of potassium permanganate. After several seconds, smoke rises from the mixture. Shortly afterward violet flames erupt from the mixture.

Running time = 1:16

When mixed with potassium permanganate, glycerine undergoes spontaneous combustion. The slow oxidation of glycerine by potassium permanganate gradually speeds up at the reaction releases heat. Eventually, the mixture catches fire. The violet color of the flames is produced by potassium ions in the flame.

The demonstration is pertinent to discussions of oxidation, exothermic reactions, factors affecting reaction rates, and emission from excited ions (flame tests).

FURTHER INFORMATION

> *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Volume 1, pages 83-84, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What types of energy are released during the reaction?

2. What causes the reaction to accelerate?

3. How could the delay between mixing and the appearance of flames be shortened?

4. What color are the flames? What element produces this color in a flame?

Thermite Reaction

A clay flower pot is filled with a mixture of iron oxide and aluminum. A layer of potassium permanganate is placed on top of this. The pot is suspended over a bucket of sand, and a small amount of glycerine is poured into the pot. After several seconds, smoke begins to rise from the pot, followed by sparks, and after a brilliant flare, a stream of white hot liquid runs out of the pot into the bucket. After the reaction has subsided, a glowing hot ingot of iron is removed from the bucket.

Running time = 2:38

The reaction between iron oxide and aluminum is highly exothermic. The reaction releases so much heat that the iron produced is molten. The reaction is triggered by heating a small portion of the mixture very hot. This is accomplished with the reaction of potassium permanganate and glycerine (see Demonstration 24). The sparks emitted during the reaction are small pieces of burning aluminum, the same material used in sparklers.

This demonstration is suitable for discussions of exothermic reactions, the metallurgy of iron, and energy of activation.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 85-89, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What types of energy are released during the reaction?

2. What is the substance that ran out of the pot into the bucket?

3. Write a chemical equation for the reaction in the demonstration.

4. What is the purpose of the potassium permanganate placed on top of the iron oxide and aluminum mixture?

Explosion of Nitrogen Triiodide

A stand contains three rings, each holding a piece of filter paper covered with a brown powder. When the brown powder on the bottom paper is lightly touched with a feather on the end of a long pole, the material on all three pieces of paper explodes with a snap, forming a cloud of violet vapor. Slow motion shows that the powders on the three papers explode almost simultaneously.

Running time = 1:14

Nitrogen triiodide, the brown material on the filter paper, is extremely sensitive to mechanical shock when it is dry. This demonstration shows that a great deal of energy is released when NI_3 decomposes, and that the activation energy for this reaction is quite small.

This demonstration is appropriate when discussing the release of chemical energy, activation energy, explosions, and the properties of nitrogen or iodine.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 96-98, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SUGGESTED EXERCISES AFTER THE DEMONSTRATION

1. What evidence is there that a chemical reaction occurred?

2. Only one filter paper was touched. What caused the material on the other two pieces of filter paper to explode?

3. The violet cloud is a product of the reaction. What is this product?

Dehydration of Sugar by Sulfuric Acid

A small amount of concentrated sulfuric acid is poured into a beaker about one quarter full of powdered sugar. The mixture is stirred, and it gradually darkens. Suddenly a black, steaming column rises up and out of the beaker to a height about twice that of the beaker.

Running time = 1:18

Concentrated sulfuric acid has a very great affinity for water. This affinity is so great that it can dehydrate organic materials, removing the elements of water. Sugar is a carbohydrate. When sulfuric acid removes the elements of water, what remains is carbon. The carbon remnant is what darkens the mixture. When sulfuric acid combines with water, a great deal of heat is released. This heat is enough to vaporize some of the mixture, which causes it to foam.

This demonstration is suitable for discussions of the dehydrating properties of sulfuric acid, the properties of carbohydrates such as sugar, and exothermic processes.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 77-78, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is meant by a "carbohydrate?"

2. What remains after the elements of water are removed from a carbohydrate?

3. Why is concentrated sulfuric acid used, rather than an aqueous solution?

Demonstration 28

The Collapsing Can: Effect of Atmospheric Pressure

A small amount of water is poured into a five-gallon steel can, and the can is heated over a burner until the water inside has boiled for several minutes. The can is stoppered, removed from the heat, and set into a tray of ice. The can collapses. An aluminum soft-drink can containing a small amount of water is heated on a hot plate until the water boils. The can is quickly inverted and dipped into a beaker of water. The can immediately collapses.

Running time = 2:27

In the first segment, water is boiled in a 5-gallon steel can. The steam displaces some of the air in the can. When the can is filled with steam at atmospheric pressure, it is sealed and cooled. Cooling the can causes the steam to condense to liquid water. This reduces the pressure of the gases inside the can. The pressure outside the can is much greater than that inside the can, and the can is not strong enough to withstand the pressure difference. The atmosphere crushes the can.

In the second segment, water is boiled in an aluminum beverage can. After steam has displaced the air in the can, the can is inverted and its top is dipped into water. This cools the can and the steam within it condenses. The atmosphere crushes it immediately. The aluminum is so weak that the atmosphere crushes it before water can fill it.

This demonstration is useful when discussing the pressure of the atmosphere, changes of state of water, and the relative volumes occupied by the same amount of gas and liquid.

FURTHER INFORMATION

> *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Volume 2, pages 6-8, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SUGGESTED EXERCISES AFTER THE DEMONSTRATION

1. Does this demonstration involve chemical or physical changes?

2. What does the can contain before boiling commences? How do the contents of the can change as the water boils inside the can? What does the can contain immediately before it is sealed? What can be said about the temperature of the gas inside the can just before it is sealed?

3. Immediately before the can is sealed, what is the pressure inside the can? Outside the can?

4. When the can is cooled, what happens to the gases inside the can? What effect does this have on the pressure inside the can?

5. Estimate how many moles of steam (gaseous water) would fill the 5-gallon can at 105°C and atmospheric pressure. (Assume ideal gas behavior for the steam.) What would be the volume of this amount of water when it condenses to a liquid?

6. Suggest a way to restore the 5-gallon can to its original shape.

Demonstration 29

Effect of Temperature on the Properties of Gases

Dry ice is sealed into a balloon, and the balloon slowly inflates. A sealed flask containing chlorine gas is placed in a dish of an acetone-dry ice mixture, and the chlorine condenses to a yellow liquid. A sealed flask containing red-brown nitrogen(IV) oxide gas is immersed in liquid nitrogen, and a blue solid condenses inside the flask. The inflated balloon containing dry ice is immersed in the liquid nitrogen, and it collapses. As the balloon warms again, it slowly expands.

Running time = 4:45

Dry ice is solid carbon dioxide, and it sublimes to gaseous carbon dioxide at -78°C. When it is placed in a balloon at room temperature, the subliming gas inflates the balloon. When the balloon is chilled to below the sublimation temperature, the gaseous carbon dioxide solidifies, and the balloon collapses. Chlorine boils at -35°C, and the dry-ice acetone mixture has a temperature of -78°C. Therefore, gaseous chlorine liquefies when its container is placed in the bath. Nitrogen(IV) oxide is composed of a mixture of different molecules. At room temperature, the mixture contains red-brown NO_2 and colorless N_2O_4 in equilibrium. When nitrogen(IV) oxide is cooled in liquid nitrogen (-196°C), a disproportionation reaction occurs. One product is blue dinitrogen trioxide, N_2O_3.

This demonstration is suitable for discussions of the properties of gases and changes in state, and for discussions of the identity of the chemical species responsible for the color changes in the N_2O_4 system.

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Why did the balloon inflate after dry ice was sealed into it? Why did it shrink when it was immersed in liquid nitrogen?

2. Dry ice sublimes at -78°C and chlorine boils at -35°C. What can be said about the temperature of the dry ice and acetone mixture?

3. What happens to the intensity of the yellow color in the chlorine flask when the flask is cooled? Why does this happen?

4. What is the oxidation state of the nitrogen in N_2O_4 and in N_2O_3? Is another product formed when N_2O_3 forms from N_2O_4?

Combining Volumes of Gases

Two identical flasks, one containing nitrogen(II) oxide and the other oxygen, are connected by tubing. The nitrogen(II) oxide flask is also connected to a beaker of water. When clamps are removed from the tubing, a red-brown gas forms inside the flasks. Gradually, water is drawn into the flasks until one flask is completely filled and the other is half filled with water. The gas remaining is colorless. When a glowing splint is inserted into this gas, the splint bursts into flame.

Running time = 4:54

Nitrogen(II) oxide reacts with oxygen to produce red-brown nitrogen(IV) oxide. The product nitrogen(IV) oxide is soluble in water. The stoichiometry of the reaction is 2 moles of NO to 1 mole of O_2. The demonstration shows the stoichiometry in terms of volumes, 2 volumes of NO to 1 volume of O_2. The starting materials are equal volumes of NO and O_2. After the product is dissolved in water, one-half volume of a colorless gas remains. This gas is identified as oxygen, for it supports combustion. Therefore, the volume of oxygen consumed is half that of nitrogen(II) oxide.

This demonstration is suitable for discussions of combining volumes of gases, stoichiometry, and properties of nitrogen(II) oxide.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 167-179, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the volume of each of the reactants, NO and O_2?
2. What is the red-brown gas produced in the reaction?
3. What happens to the red-brown gas?
4. What volume of gas remains after the reaction is complete?
5. What is the identity of this remaining gas?
6. What is the volume of the remaining gas?
7. What volume of NO and what volume of O_2 are consumed in the reaction?
8. What is the ratio of the volume of NO consumed to that of the O_2 consumed?
9. Write a chemical equation for the reaction of NO with O_2.

Red, White, and Blue Coin-Operated Reaction

Three flasks containing colorless liquids are arranged in a row and connected with glass tubing. The flask on the left is half filled, the middle flask completely filled, and the one on the right contains only a small amount of liquid. A piece of copper is dropped into the flask on the right, and a series of reactions begins. The flask on the right fills with red-brown gas. Liquid flows from the center flask to the one on the left, turning the contents of the left flask red. Then red liquid flows from the left flask back into the center flask, where it returns to colorless. Some of the colorless liquid from the center flask flows from the center flask into the flask on the right, where it turns blue. When the flask on the right is half filled with blue liquid, its weight causes it to sink into the pad on which it is resting and to trigger a mechanism that reveals an American flag.

Running time = 3:55

The process is initiated by dropping a piece of copper into concentrated nitric acid. The reaction between these produces red-brown nitrogen(IV) oxide gas, which pushes liquid from the center flask into the flask on the left. The left flask contains a solution of sodium hydroxide and the center flask a solution of phenolphthalein and dilute nitric acid. When these solutions mix, the phenolphthalein turns red. The red-brown gas dissolves in the solution in the center flask. This reduced the pressure in the flask, and liquid flows from the left to the center. The phenolphthalein returns to colorless when it mixes with the nitric acid in the center flask. As more nitrogen(IV) oxide gas dissolves, liquid flows from the center flask to the flask on the right, forming a blue solution of copper nitrate.

This demonstration is particularly useful for practicing observation skills. Students should be encouraged to speculate on the contents of the various flasks and what reactions account for their observations.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 83-91, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. The demonstration started when a piece of copper was dropped into a small amount of liquid in the flask on the right. What happened when this was done? What must the liquid have been to cause this reaction?

2. What is the red-brown gas?

3. Why does the liquid flow from the center flask to the left flask?

4. The red color is from phenolphthalein. Which solution originally contained the phenolphthalein, the one in the center flask or the one in the left flask?

5. Why does the liquid flow back into the center flask?

6. What produces the blue solution in the right flask at the end of the reaction?

7. Write an equation for the reaction between copper and concentrated nitric acid.

8. Write an equation for the reaction between aqueous nitric acid and aqueous sodium hydroxide.

Diffusion of Bromine Vapor

A small vial of liquid bromine is broken and dropped into a glass cylinder full of air. The red-brown vapor of bromine covers the bottom of the cylinder and slowly diffuses upward into the air. Subsequently, another similar vial of liquid bromine is broken in a long evacuated tube. The bromine vapor fills the tube immediately.

Running time = 3:05

In this demonstration, students should concentrate on the differences between the containers into which the bromine vapor is introduced. Specifically, the cylinder is open to the air, and therefore, filled with air, while the tube is sealed. Although it is not possible to see that the sealed tube is evacuated, the difference in the behavior of bromine vapor in the two containers indicates that the contents of the sealed tube are quite different than those of the open cylinder. The vapor of bromine, which as a vapor pressure of 100 torr at room temperature, rapidly occupies the entire tube, but it only slowly spreads upward in the cylinder.

The sealed tube was evacuated with a mechanical vacuum pump, and the pressure of the residual gas in the tube is about 0.01 torr. The pressure of the gas in the air filled cylinder is about 100,000 times that of the gas in the evacuated tube. This means that the concentration of gas molecules in the air-filled cylinder is about 1×10^5 times that of the molecules in the evacuated tube. Therefore, the molecules in the evaluated tube are about $(1\times10^5)^{1/3}$, or 50, times as far apart as those in the air-filled cylinder. Each molecule in the evacuated tube moves, on average, about 50 times farther between collisions than do those in the air-filled cylinder. Therefore, the diffusion of bromine vapor in the evacuated tube is much faster than in the air-filled cylinder.

FURTHER INFORMATION

> *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Volume 2, pages 63-68, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Why is the bromine stored in a sealed vial?

2. What happens to molecules of liquid bromine after it is poured into the open-top cylinder?

3. What evidence do you have for your answer to question 2?

4. As bromine molecules enter the gaseous state in the open cylinder, what are their immediate surroundings on a molecular scale?

5. What is in the evacuated apparatus before the bromine vial is broken?

6. When the sealed vial is broken in the evacuated apparatus, what are the immediate surroundings of the bromine molecules?

7. Compare your answers to questions 4 and 6 and explain the observed differences in speed at which the bromine vapor travels in the two containers.

Singlet Molecular Oxygen

A cylinder of chlorine is connected with tubing to a gas washing bottle containing a small amount of hydrogen peroxide solution, and this bottle is connected to a second bottle containing sodium hydroxide solution. When the cylinder valve is opened, bubbles form in the bottles, and a faint red glow appears in the bubbles in the hydrogen peroxide solution. A red dye is added to the hydrogen peroxide solution, and the glow becomes much brighter.

Running time = 1:55

The reaction between chlorine and hydrogen peroxide produces molecular oxygen in an excited state, a singlet state. Some of these excited molecules return to the ground state by the emission of red light. This process, in which light is emitted by a chemically generated species, is called chemiluminescence. The dye added to the mixture is violanthrone. Violanthrone receives energy from the singlet oxygen, producing an excited state of violanthrone. This excited state of violanthrone also returns to the ground state by emitting red light, but it does so more efficiently than singlet oxygen. The transfer of energy from singlet oxygen to violanthrone is also more efficient than the emission of light by singlet oxygen. Therefore, the emission of red light is greatly enhanced by the addition of violanthrone to the mixture.

The red emission results from the simultaneous return of *two* singlet excited-state oxygen molecules to their triplet ground state. The reverse of this process is responsible for the blue color of liquid oxygen, which can be observed in Demonstration 38.

This demonstration is appropriate in discussions of chemiluminescence, excited states, energy transfers, and molecular orbital theory.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 133-145, Bassam Z. Shakhashiri, University of Wisconsin Press (1983).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. How does the oxygen in the bubbles differ from oxygen in the air?

2. The oxygen in the bubbles is produced by a reaction between chlorine gas and hydrogen peroxide solution. Write a chemical equation for this reaction.

3. Describe some other examples of chemiluminescence.

The Blue Bottle Experiment

Three stoppered flasks are arranged in a row. Two are half filled with colorless liquids, and the third is half filled with a yellow liquid. The liquid in the first flask turns blue when it is shaken. The liquid in the second first turns red when it is shaken, then purple when it is shaken more vigorously. When the third flask is shaken, its contents change from yellow to orange, and, with more vigorous shaking, to green.

Running time = 1:28

All three flasks contain alkaline solutions of dextrose and methylene blue. The second flask also contains resazurin, and the third indigo carmine. All three flasks also contain air. When the flask is shaken, oxygen from the air dissolves in the solution. Methylene blue catalyzes the oxidation of dextrose by oxygen. Methylene blue is blue in its oxidized form and colorless in its reduced form. During the catalysis, methylene blue is oxidized by oxygen, becoming blue. The blue, oxidized form reacts with dextrose and is reduced back to the colorless form.

This demonstration is appropriate for discussions of catalysis, oxidation-reduction reactions, properties of carbohydrates, and of oxygen.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 142-146, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What does each flask contain?

2. How can the composition of a solution be changed when it is shaken?

3. Is there a limit to the number of times the solution can be turned blue by shaking it? Why?

Volume Changes upon Mixing of Liquids

Two 50-milliliter volumetric flasks are filled to their calibration marks with liquid, one with ethanol and the other with colored water. The contents of these flasks are combined in a 100-milliliter volumetric flask. The volume of the combined liquids is noticeably less than 100 milliliters. Two different 50-milliliter volumetric flasks contain carbon disulfide and colored ethyl acetate. These liquids are mixed in a 100-milliliter volumetric flask, and the volume of this mixture is greater than 100 milliliters.

Running time = 4:09

Our everyday experience with liquids leads us to expect that the volumes of liquids are conserved. In particular, when two liquids are mixed, the volume of the mixture is usually the sum of the volumes of the individual liquids. This is not always the case, as this demonstration shows. The volume of the mixture may be greater or less than the sum of the individual volumes. This happens, of course, only when the two liquids are different. The liquids used here have the largest volume changes upon mixing among common liquids.

The demonstration is suitable for discussions of preparing solutions, properties of liquids, and intermolecular attractive and repulsive forces.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 225-228, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. If the volumes of liquids were additive, what would be the volume of the mixtures?

2. What happens to the average distance between molecules when water and ethanol are mixed? When carbon disulfide and ethyl acetate are mixed?

3. Which intermolecular forces are more important when water and ethanol are mixed, attractive or repulsive? When carbon disulfide and ethyl acetate are mixed?

Miscibility and Density of Liquids

Four colorless liquids are poured into a glass cylinder. The four liquids settle into distinct layers. A dye is added to the cylinder. After the liquids are shaken together, they return to four separate layers, with the top and third liquids colored by the dye and the second and bottom layers colorless.

Running time = 2:15

The liquids are decane (density 0.8 g/mL), water (1 g/mL), bis(2-chloroethyl)ether (1.2 g/mL), and perfluoro-1,3-dimethylcyclohexane (1.8 g/mL). Because the liquids are mutually immiscible, they settle into separate layers even after being shaken together. The densities of the liquids are quite different, and the liquids separate rather quickly.

This demonstration is suitable for discussions of the properties of liquids, miscibility, density, and solubility.

FURTHER INFORMATION

> *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Volume 3, pages 229-233, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Four colorless liquids were poured into the cylinder. Which one (first, second, third, or fourth) is the most dense? Which the least dense?

2. One of the liquids is water. Suggest a substance that could be added to the cylinder of colorless liquids to determine which layer is water.

3. No two of the liquids are miscible. What does this say about the forces between the molecules of these liquids?

Demonstration 37

Liquid Carbon Dioxide

Dry ice is crushed and sealed into a transparent tube attached to a pressure gauge. The gauge shows the pressure inside the tube rising. At a gauge reading of about 50 psi, a liquid appears in the tube. The pressure holds constant as the liquid forms. When the solid has disappeared, the liquid begins to boil, and the gauge shows the pressure increasing. When a valve on the tube is opened, the gauge reading returns to zero, and the liquid solidifies.

Running time = 4:56

At atmospheric pressure, dry ice (solid carbon dioxide) sublimes, that is, it changes directly from solid to gas. However, at pressures over about 5 atmospheres, solid carbon dioxide, as it warms, melts to a liquid before becoming a gas. The container used in this demonstration has walls of 1/2 inch thick acrylic, strong enough to resist the pressure that develops as solid dry ice sublimes. As the pressure increases, it eventually reaches 5 atmospheres, and the dry ice melts rather than subliming. The gauge attached to this apparatus reads the pressure differential, that is, the difference between the pressure inside the tube and outside. Atmospheric pressure (ca. 14 psi) is the outside pressure, so the total internal pressure is about 65 psi when the carbon dioxide melts.

This demonstration is particularly useful in discussions of phase changes and phase diagrams.

FURTHER INFORMATION

L. Andrews, *Journal of Chemical Education*, Volume 66, pages 597-598 (1989).

SUGGESTED QUESTIONS DURING THE DEMONSTRATION

1. What is the reading on the pressure gauge before the tube is sealed?

2. Does the pressure rise continuously, or does it pause at one value for a time?

3. What is the reading on the gauge as the solid liquefies?

4. What happens to the liquid when the valve is opened?

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the chemical name and formula of dry ice?

2. What is the temperature of dry ice?

3. Why is dry ice called *dry*?

4. What is the approximate value of atmospheric pressure in pounds per square inch (lb/in^2, psi)?

5. Convert to units of atmospheres the gauge reading when liquid forms.

6. The gauge reads the difference between the pressure outside the tube and the pressure inside the tube. What is the pressure inside the tube when the liquid forms?

Demonstration 38

Paramagnetism of Liquid Oxygen

Oxygen gas from a cylinder is passed through a copper coil immersed in liquid nitrogen. Blue liquid oxygen is collected at the outlet of the coil. When the liquid is poured into the gap between the poles of a magnet, it stays there until it evaporates.

Running time = 4:29

Liquid oxygen boils at -183°C. Therefore, gaseous oxygen can be condensed with liquid nitrogen, which has a temperature of -196°C. Oxygen is blue, and its color is easily seen when it is liquefied. It is also paramagnetic, meaning that it is attracted into a magnetic field. Its paramagnetism indicates that the molecules of oxygen contain unpaired electrons. These unpaired electrons are not predicted by the Lewis structure of oxygen, but they are accounted for by molecular orbital theory.

The blue color of liquid oxygen results from an absorption of light in the red portion of the visible spectrum. In this absorption *one* photon simultaneously promotes *two* ground-state oxygen molecules to their first excited state. The reverse of this process produces a red glow, which can be observed in Demonstration 33.

This demonstration is appropriate in discussions of paramagnetism, properties of oxygen, bonding, and molecular orbital theory.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 147-152, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Does oxygen have a higher or lower boiling point than nitrogen?

2. What color is liquid oxygen? What is the color of liquid nitrogen?

3. Is liquid oxygen paramagnetic or diamagnetic? Is liquid nitrogen paramagnetic or diamagnetic?

4. What is the minimum number of unpaired electrons in an oxygen molecule?

5. What is the minimum degeneracy of the highest occupied molecular orbitals in oxygen?

The Ice Bomb

A hollow cast iron sphere is filled with water, sealed tightly, immersed in an acetone-dry ice bath, and covered with a wooden box. After several seconds, the cast iron sphere explodes, throwing wooden slats around. The explosion is shown in slow motion. The process is repeated without the wooden box to show the pieces of the exploding sphere jumping out of the bath.

Running Time = 5:19

This demonstration strikingly illustrates the expansion of water as it freezes. Solid water is less dense than liquid water; ice floats in water. Therefore, a given mass of ice occupies a greater volume than the same mass of water. The pressure required to prevent the expansion of water as it freezes is immense. The cast-iron container is not strong enough to provide this pressure, and it ruptures explosively. A mixture of acetone and dry ice is used, because its low temperature (-77°C) causes the water to freeze quickly.

The demonstration is most useful for showing the properties of water and its changes in state. Concepts related to the demonstration include hydrogen bonding and its relationship to the arrangements of water molecules in the liquid and in the solid.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 310-312, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Does this demonstration involve a chemical change or a physical change?

2. What caused the water to freeze? How much colder than the freezing point of water is the acetone-dry ice mixture?

3. Why doesn't the cast iron container break immediately when the acetone-dry ice mixture is poured on it?

4. Did all of the water in the sphere freeze before the sphere broke?

5. What are some other phenomena you have seen that illustrate the expansion of water on freezing?

Demonstration 40

Fog: An Aerosol of Condensed Water Vapor

Chunks of dry ice, solid carbon dioxide, are placed in a pan of nearly boiling water. A dense white cloud of fog first rises above the pan. As more fog is produced, it stops rising and flows over the rim of the pan and down to the table top.

Running time = 1:35

Fog forms when water vapor in the air condenses into tiny suspended droplets. This condensation occurs when warm, humid air is cooled. The warm air over the hot water is nearly saturated with water vapor. This warm air is cooled by mixing it with the cold carbon dioxide gas that sublimes from dry ice. Initially the hot water heats the air above it, making it less dense, and causing the fog to rise. However, eventually, the cold carbon dioxide cools the air to the point that it becomes more dense than the air around the pan, and the fog sinks.

The demonstration is suitable for discussions of changes of state, the density of gases, and the meteorological formation of fog.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 114-120, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the white cloud that forms when the dry ice is dropped into the hot water?

2. Why does the cloud first rise in the air and then sink?

3. Very little fog forms when ice is dropped into hot water. What properties of dry ice are essential for the formation of fog?

4. What will happen to the temperature of the water in the dish pan? Why?

Crystallization of Supersaturated Sodium Acetate Solution

A few crystals of sodium acetate are placed in the center of a black board, and a clear, colorless solution is slowly poured onto the crystals. When the solution strikes the crystals, it solidifies. As more solution is poured onto the solid, a white column grows until it reaches nearly 30 centimeters in height.

Running time = 2:17

The clear, colorless solution is supersaturated with respect to sodium acetate trihydrate. The solution is so supersaturated, that when crystallization begins, all of the water is trapped inside the solid. The solution contains about 45% sodium acetate by weight. The crystallization process is exothermic, releasing about 20 kJ per mole of sodium acetate.

This demonstration is suitable for discussions of solutions, supersaturation, and crystallization.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1, pages 27-30, Bassam Z. Shakhashiri, University of Wisconsin Press (1983)

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the purpose of the crystals placed on the black board?

2. What would happen if one of these crystals were dropped into the flask containing the solution?

3. Heat is given off when the crystals form. What must be done to get the crystals to dissolve again?

Osmosis through a Copper Hexacyanoferrate(II) Membrane

A dropper injects a single drop blue solution near the bottom of a test tube containing a yellow solution. A rust-colored, transparent membrane forms at the interface between the two solutions, encapsulating a drop of the blue solution. The drop rises to the top of the yellow solution, stays there for a short time, then sinks slowly to the bottom of the tube.

Running time = 1:34

When the blue copper sulfate solution contacts the yellow potassium hexacyanoferrate(II) solution, a gelatinous membrane of copper hexacyanoferrate(II) forms. This membrane surrounds the drop of copper sulfate solution, keeping the liquids separate. The density of the copper sulfate solution is slightly less than that of the potassium hexacyanoferrate solution. For this reason, the blue drop slowly rises in the yellow solution.

The concentration of ions in the copper sulfate drop is less than that in the yellow potassium hexacyanoferrate solution. Therefore, the concentration of water is greater inside the drop than outside it. The membrane surrounding the drop is permeable to water but not to the ions in the solutions. As a result, osmosis drives water from the copper sulfate drop through the membrane into the potassium hexacyanoferrate(II) solution. This osmosis increases the density of the copper sulfate solution in the drop and, to a lesser extent, decreases the density of the potassium hexacyanoferrate solution. When the solution in the drop becomes more dense than the solution around it, the drop sinks.

This demonstration can be presented in discussions of colligative properties of solutions, specifically, osmosis.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 390-393, Bassam Z. Shakhashiri, University of Wisconsin Press (1989)

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the chemical identity of each of the solutes?

2. Why is the copper sulfate solution injected near the bottom of the test tube, rather than at the top?

3. What is the evidence that a chemical reaction takes place between the solutions?

4. Why does the blue drop rise?

5. What evidence is there that the density of the blue drop changes during the demonstration? Does its density increase or decrease?

6. What is a semi-permeable membrane?

7. In which direction do water molecules move through the membrane during the demonstration? What evidence supports this inference?

Sugar Solution in Polarized Light

A beaker containing water is placed between crossed polarizing filters. One of the filters is rotated. The intensity of white light passing through the water is essentially the same as that passing through the filters. A beaker of sugar solution is placed between the polarizers. When one filter is rotated, the light passing through the sugar solution changes colors, showing all colors of the spectrum.

Running time = 2:00

When light passes through a polarizing filter, the light is polarized into a single plane. When the polarized light strikes a second filter with its polarizing plane perpendicular to that of the first, none of the light passes through. When a substance that rotates the plane polarized light is placed between the filters, light can now pass through the filters. Sugar is such a substance. In fact, sugar does not rotate all colors by the same amount. Therefore, as one filter is rotated, it is brought into alignment in turn with the planes of the various colors as they are rotated by the sugar solution.

This demonstration is suitable for discussions of stereoisomerism, properties of organic compounds, and polarized light.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3, pages 386-389, Bassam Z. Shakhashiri, University of Wisconsin Press (1989).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What happens to the intensity of the light that passes through the polarizers when one of them is rotated? Does the light change color with water between the filters?

2. What happens when one polarizer is rotated with sugar solution between the filters?

3. Why does the color of the sugar solution change as one filter is rotated? (Are all colors of polarized light rotated by the same amount by the sugar solution and by the polarizing filters?)

4. What does the fact that sugar solutions rotate a plane of polarized light indicate about the geometry of the sugar molecules?

The Landolt Iodine Clock Reaction

Ten beakers are arranged in two rows, one behind the other. Each beaker contains a colorless liquid. The contents of the end beaker in the back row is poured into the beaker in front of it. On the count of ten the contents of the second beaker in the back row are poured into the beaker in front. This counting and pouring continues down the rows. One by one, the colorless mixtures suddenly turn black.

Running time = 1:31

The beakers in the front row contain sodium bisulfite solutions. The beakers in the back row contain solutions of potassium iodate and starch. Sodium bisulfite and potassium iodate react with each other, producing iodide ions. When the bisulfite has been consumed, iodide ions react with iodate, producing iodine. The iodine and iodide ions form an intensely colored dark blue complex with starch.

This demonstration is suitable for discussions of stoichiometry of reactions, rates of reactions, and the chemistry of iodine and its compounds.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 4, Demonstration 10.1, Bassam Z. Shakhashiri, University of Wisconsin Press (in press).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Was there evidence of a reaction immediately upon the mixing of the solutions?

2. How can you tell a chemical reaction was occurring during the period that the mixtures remained colorless?

3. How could the amounts of solutions be varied to determine which one contains the limiting reagent in the reaction?

Color Variations of the Landolt Reactions

Six beakers containing colorless liquids are arranged in two rows, one behind the other. The contents of the each beaker in the back row are poured into the beaker in the front of it. The mixtures remain colorless for about 10 seconds, then the first suddenly turns orange, the second one red, and the third black.

Running time = 0:31 (This demonstration is done without narration.)

The reaction in these beakers is the same as the reaction in the Landolt reaction of Demonstration 44. The first mixture generates only a small amount of iodine, which colors the solution orange. The second generates more iodine, which turns the solution red. The third beaker contains starch, and the deep blue starch complex between iodine and iodide ions makes the solution appear black.

This demonstration is suitable for discussions of stoichiometry and rates of chemical reactions.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 4, Demonstration 10.2, Bassam Z. Shakhashiri, University of Wisconsin Press (in press).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Does the time between mixing and color change vary among the three reactions?

2. The last beaker contains starch, but the first two do not. What would happen if all beakers contained starch?

3. Could the colored substance produced in the first beaker be the same as that produced in the second beaker? If so, why do the colors look different?

Demonstration 46

The Briggs-Rauscher Oscillating Reaction

Three clear, colorless liquids are combined in a large beaker. The mixture first turns yellow, then black, then colorless. This cycle of color changes repeats several times.

Running time = 1:51

This is an example of an oscillating reaction, one that cycles through a series of changes repeatedly. The yellow color is produced by iodine in the mixture. A complex between iodine, iodide ions, and starch turns the mixture black.

The overall reaction that drives the color changes is the disproportionation of hydrogen peroxide to water and oxygen gas. This reaction is catalyzed by iodate ions. The catalysis occurs by two processes, one that generates iodine and another that regenerates iodate ions. Malonic acid in the mixture reacts with iodine to produce iodide ions. The oscillations stop when the amount of hydrogen peroxide in the solution is no longer sufficient to drive them. When the oscillations stop, the solution contains iodine, iodide ions, and starch (among many other species), and therefore, the mixture is deep blue.

This demonstration is appropriate for discussions of oscillating reactions and reactions involving competing processes. It is also useful as a system for detailed observation.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 232-256, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Does the rate at which the colors change remain constant?

2. What other changes occur in addition to the color changes?

3. Iodine produces the yellow color. A complex of iodine, iodide ions, and starch produces the black color. What must be changing in the mixture to account for the color changes?

4. What experiments could be performed to identify the gas in the bubbles in the solution?

Nitrogen Gas Oscillator

A solution of sodium nitrite and a solution of ammonium chloride are mixed in a test tube. Bubbles form in the mixture and rise to the surface. The bubbling repeatedly subsides and then increases.

Running time = 2:10

In aqueous solution, ammonium ions react with nitrite ions forming elemental nitrogen. When the solution becomes supersaturated in nitrogen, bubbles form and the gas escapes.

Oscillations occur in the formation of bubbles, because the rate at which nitrogen escapes from the solution varies, but the rate at which nitrogen is produced is essentially constant. The rate at which nitrogen escapes from the solution depends on the size of the bubbles. As the bubbles grow, their surface area increases, and the rate at which nitrogen leaves the solution and enters the bubbles increases. Eventually, nitrogen leaves the solution faster than it is produced, and the bubbling subsides. Bubbling increases after the solution again becomes supersaturated in dissolved nitrogen.

This demonstration is appropriate for discussions of oscillating reactions, supersaturated solutions, and bubble formation.

FURTHER INFORMATION

Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2, pages 301-304, Bassam Z. Shakhashiri, University of Wisconsin Press (1985).

SAMPLE QUESTIONS AFTER THE DEMONSTRATION

1. What are the two solutions combined in the test tube?

2. What is the product of the reaction between them?

3. Would you expect to find ammonium nitrite listed in a chemical catalog? Why?

Genie in a Bottle

The stopper is removed from a flask wrapped in aluminum foil. Almost immediately, a cloud of white smoke blasts from the mouth of the flask. A second flask that is not wrapped with foil shows a small packet suspended by a thread held by the stopper. When the stopper is removed, the packet falls into a colorless liquid in the flask, and the liquid quickly bubbles and froths, producing a blast of white smoke from the mouth of the flask.

Running time = 2:00

The flasks contain 30% hydrogen peroxide and a packet of manganese dioxide. Manganese dioxide catalyzes the decomposition of hydrogen peroxide into oxygen gas and water. This decomposition is very rapid and highly exothermic, forcing droplets of water and fog from the mouth of the flask.

This demonstration is suitable for discussions of catalysis, oxidation-reduction reactions, and exothermic reactions.

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. What is the liquid in the flask? What is in the small packet?

2. What it the white smoke that comes from the bottle?

3. What is the purpose of the material in the packet?

4. Write a balanced chemical equation for the reaction that occurs in this demonstration.

Simulation of a Nuclear Chain Reaction

Two rubber stoppers rest on each bail of 72 mousetraps in a rectangular array covered by a clear plastic box. A single stopper is dropped through a hole in the top of the box. When it strikes one of the mousetraps, it trips the mousetrap, which launches its two stoppers. This is the beginning of a chain reaction of tripping mousetraps launching stoppers. Slow motion clearly shows that the number of flying stoppers increases and then decreases during the process.

Running time = 1:51

This demonstration illustrates the essential feature of a chain reaction, nuclear or otherwise. In a chain reaction, each step produces the trigger that sets off a subsequent step. Therefore, one reaction can initiate a long sequence (chain) of similar reactions. Furthermore, if a reaction releases more than one trigger, then the rate at which these reactions occur will increase as long as sufficient reactant remains.

In this simulation, each tripped mousetrap releases two stoppers. The released stoppers are contained by the box, so they trigger other mousetraps. This generates a chain reaction whose rate at first accelerates, then decelerates as the mousetraps are triggered. This process is a mechanical analog of a nuclear fission chain reaction, such as that of uranium-235. In this reaction, a neutron triggers a U-235 nucleus to disintegrate into two smaller nuclei and several free neutrons. Each of the free neutrons can trigger another U-235 nucleus to undergo fission, causing the rate of fission to accelerate.

The demonstration is best suited to discussions of nuclear fission. It can also be used to illustrate other types of chain reactions.

SAMPLE EXERCISES AFTER THE DEMONSTRATION

1. Describe how the number of flying stoppers changes during the process.

2. What would happen to the overall length of the process if each mousetrap contained only one stopper?

3. What would happen if each mouse-trap had three rubber stoppers on it?

4. What purpose does the plastic box serve in this demonstration? (Hint: What would happen if the box were not there?) How is this related to the so-called “critical mass” of a nuclear fission reaction?