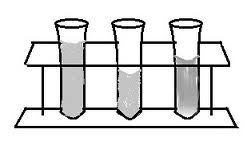
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**Instructions for**

**“*Using Chemical Demonstrations to Demonstrate Concepts in Physical Science*”**

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**PS-21**

**October 3, 2014**

**Demonstration 1 – Reaction of Baking Powder**

The gas forming reaction involves the reaction of sodium bicarbonate (baking soda) with potassium hydrogen tartrate (cream of tartar). Cream of tartar can be used as the acid in baking powder.Baking powder is a chemical [leavening agent](http://en.wikipedia.org/wiki/Leavening_agent); it a dry mixture of a [carbonate](http://en.wikipedia.org/wiki/Carbonate) or bicarbonate and a weak [acid](http://en.wikipedia.org/wiki/Acid) or acids. Sodium bicarbonate is most commonly used as the carbonate or bicarbonate component. The weak acid can vary. The reaction between the bicarbonate and acid when the dry mixture is added to water or an aqueous solution generates carbonic acid, which rapidly decomposes to water and carbon dioxide gas. The bubbles of CO2 gas formed in the wet mixture expand and thus [leaven](http://en.wikipedia.org/wiki/Leavening_agent) the mixture (Equations 1 and 2).

NaHCO3(s) + HX(s) 🡪 NaX(aq) + H2CO3(aq) (Eqn. 1)

H2CO3(aq) 🡪 H2O(l) + CO2(g) (Eqn. 2)

Cream of tartar (potassium hydrogen tartrate or potassium bitartrate) (Figure 1) has been used as the acid component of baking powder.



Materials

Weigh boats

50 mL beaker (optional)

25 mL or 50 mL graduated cylinder

Sodium bicarbonate

Potassium hydrogen tartrate

Stirring rod

Procedure

Weigh 0.42 g NaHCO3 and 0.94 g potassium hydrogen tartrate into plastic weigh boats. Combine the two solid in one of the weigh boats or in a 50 mL beaker. Note that nothing happens when the two solids are combined, as in solid baking powder. Add 20 mL H2O and watch for the formation of a gas. The mixture can be stirred with stirring rod to accelerate the rate of bubble production.

**Safety Notes:** Safety goggles are required.

As both reagents are commonly used in cooking, the resulting solution can be poured down the drain diluted with running water.

**Note:** Potassium hydrogen tartrate has a solubility of only 1 g in 162 mL of water, although the solubility increases appreciable as the temperature is increased (1g in 16 mL boiling water). This limits the use of potassium hydrogen tartrate for uses where rapid foam formation in a demonstration is desired (e.g., elephant toothpaste or a “volcano”).

**Demonstration 2 – Gas Formation**

The gas forming reaction involves the reaction of copper metal with concentrated nitric acid to produce nitrogen dioxide and copper(2+) ions (if the acid is diluted, another reaction occurs) (Equation 1). The resulting solution is initially green from the formation of cupric nitrate while a red-brown gas, NO2, is released. The gas should not be breathed so that the reaction must be performed outdoors or in a well-functioning fume hood.

Cu(s)  +  4 HNO3(aq) 🡪 Cu(NO3)2(aq)  +  2 NO2(g)  +  2H2O(l) (Eqn. 1)

The addition of water to the solution, diluting the nitrate ions present, will turn the solution the blue color of hydrated Cu2+ ions as the nitrate bound to the cupric ions is displaced by water.

Most notably, the observation of this reaction was an inspiration for the career of the well-known chemist Ira Remsen (1846-1927), who was the second president of John Hopkins University and the co-discoverer of the artificial sweetener saccharin.

“While reading a textbook of chemistry I came upon the statement, "nitric acid acts upon copper." I was getting tired of reading such absurd stuff and I was determined to see what this meant. Copper was more or less familiar to me, for copper cents were then in use. I had seen a bottle marked nitric acid on a table in the doctor's office where I was then "doing time." I did not know its peculiarities, but the spirit of adventure was upon me. Having nitric acid and copper, I had only to learn what the words "act upon" meant. The statement "nitric acid acts upon copper" would be something more than mere words. All was still. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table, opened the bottle marked nitric acid, poured some of the liquid on the copper and prepared to make an observation. But what was this wonderful thing which I beheld? The cent was already changed and it was no small change either. A green-blue liquid foamed and fumed over the cent and over the table. The air in the neighborhood of the performance became colored dark red. A great colored cloud arose. This was disagreeable and suffocating. How should I stop this? I tried to get rid of the objectionable mess by picking it up and throwing it out of the window. I learned another fact. Nitric acid not only acts upon copper, but it acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and another fact was discovered. Nitric acid acts upon trousers. Taking everything into consideration, that was the most impressive experiment and relatively probably the most costly experiment I have ever performed. . . . It was a revelation to me. It resulted in a desire on my part to learn more about that remarkable kind of action. Plainly, the only way to learn about it was to see its results, to experiment, to work in a laboratory.”[1]

Materials

5 g Cu metal (fillings or wire)

40 mL of nitric acid (concentrated)

100 mL H2O

500-mL Erlenmeyer flask

100-mL graduated cylinder

Waste bottle

Procedure

Place the copper into the bottom of the Erlenmeyer flask. Outside or in a well-functioning hood, add the nitric acid. As the reaction slows with time, the container can be swirled. When the reaction is complete or close to completion, add the water slowly.

**Safety Notes:** Safety goggles are required.

Nitric acid is a corrosive chemical and should be handled with care. Wearing gloves is recommended when handling these chemicals. Spills should be cleaned with large volumes of water. If contact is made with your skin, rinse thoroughly with water. The reaction of copper metal and nitric acid produces nitrogen oxides that should not be inhaled; thus, this procedure should be performed in a fume hood or outside.

Unwanted chemicals are not to go down the drain but should be placed in appropriate waste containers.

**References**

1. from F. H. Getman, "The Life of Ira Remsen"; Journal of Chemical Education: Easton, Pennsylvania, 1940; pp 9-10; quoted in Richard W. Ramette, "Exocharmic Reactions" in Bassam Z. Shakhashiri, *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1.*Madison: The University of Wisconsin Press, 1983, p. xiv.

**Demonstration 3 – Solid Formation**

A precipitate is a solid, water-insoluble product formed in a reaction. A reaction that f9orms a precipitate from soluble reactants is a precipitation reaction. The formation of the precipitate lead chromate from the mixing of a solution of potassium chromate and a solution lead nitrate (Equation 2) is an excellent example as the precipitate is a bright yellow color. The color should be familiar to students as lead chromate was used previously as the pigment for yellow highway lines (before toxic concerns arose) as the solid has extremely limited solubility in addition to its bright color.

K2CrO4(aq) + Pb(NO3)2(aq) 🡪 PbCrO4(s) + 2 KNO3(aq)

Materials

200 mL of 0.1 M K2CrO4

200 mL of 0.1M Pb(NO3)2

1 L beaker

Stirring rod and Waste bottle

Procedure

The solutions are made by dissolving 6.62 g of Pb(NO3)2 in water to give a total volume of 200 and by dissolving 3.88 g K2CrO4 in water to give a total volume of 200 mL.Add one of the solutions to the 1-L beaker. Slowly pour the other solution into the beaker. Stir with a stirring rod to guarantee completion.

**Safety Notes:** Safety goggles are required for all experiments.

Lead(II) nitrate, potassium chromate, and the copper salts are toxic chemicals. Gloves should be worn when handling these chemicals.

Unwanted chemicals are not to go down the drain but should be placed in appropriate waste containers.

**Demonstration 4 – Heat Release: Lighting a Bunsen Burner**

Combustion reactions (burning a fuel in oxygen) give off energy as heat and light. A convenient demonstration of this is the flame from a Bunsen burner. The reaction for the burning of natural gas (methane, CH4) is given in equation 1. An alcohol burner, candle, lighter, or similar device would also work, although the reaction would need to be changed appropriately.

CH4(g) + 2 O2(g) 🡪 CO2(g) + 2 H2O(l) (Eqn. 1)

Materials

Bunsen burner

Spark lighter or other appropriate device to light burner

Procedure

Light the Bunsen burner and let students observe flame.

**Safety Notes:** Safety goggles are required.

Appropriate care should always be taken with the use of open flames.

**Demonstration 5 - Light**

A glow stick, also called a light stick, is generally composed of a transparent plastic case that holds chemicals in two compartments. One compartment is a glass or plastic tube that is sufficiently brittle to be broken when the outer plastic case is bent, releasing the components into the larger compartment. The mixing of the chemicals results in a chemiluminescence reaction. The reactions generally use hydrogen peroxide as an oxidant; luminol, oxalates, or oxalyl chloride as reductants; and possibly a sensitizer, a molecule that can absorb energy from a chemical reaction and then release the energy as photons.

A chemiluminescence reaction is a chemical reaction that produces a molecule in an excited electronic state and releases the excitation energy as a photon of light, rather than heat, to reach its ground electronic state. Chemiluminescence reactions are rather uncommon because they have several requirements. For example, this type of reaction must generate an atom, an ion, or a molecule capable of being excited electronically at a reasonable rate and produce a sufficient amount of energy to generate this species in an electronically excited state. Also, the excited species must be capable of returning to its ground state by releasing a photon. In addition, this process must be able to compete with other processes by which the excited molecule can relax, most notably by loss of the excess energy as heat or by transfer of the energy to another molecule, exciting it electronically. The latter process is called quenching. If the quenching molecule, the quencher, in turn releases the transferred energy as a photon, it is a sensitizer; this luminescence is called sensitized luminescence.

Materials

One glow stick

Procedure

Bend the glow stick to snap the glass or plastic container inside to start the chemiluminescence reaction.

**Safety Notes:** No safety equipment is required.

After use, the glow stick can be placed in a trash receptacle. Do not allow students to open the glow sticks as fluorescent molecules are often carcinogenic.

**Demonstration 6 – Heat Release: Decomposition of Hydrogen Peroxide**

Hydrogen peroxide is a thermodynamically unstable material that can readily be oxidized to produce dioxygen or reduced to produce water and releasing heat. Hydrogen peroxide actually spontaneously decomposes, disproportionating to give dioxygen and water. Normally, this

2 H2O2(aq) 🡪 2 H2O(l) + O2(g)

decomposition takes place slowly; thus, one can keep an aqueous solution of hydrogen peroxide in the dark in one’s bathroom for months. However, the addition of a catalyst can dramatically increase the rate of this decomposition and the rate of accompanying heat release.

2 Cu2+(aq) + H2O2(aq) 🡪 2 Cu1+(aq) + O2 + 2 H+(aq)

2Cu1+(aq) + H2O2(aq) + 2 H+(aq) 🡪 2 Cu2+(aq) + 2 H2O(l)

2 H2O2(aq) 🡪 2 H2O(l) + O2(g)

Materials

0.5 g CuCl2.2H2O

20 mL of 30 % H2O2

250-mL beaker

100-mL graduated cylinder

Waste bottle

Procedure

Add 20 mL of hydrogen peroxide to beaker. Add copper chloride. Solution will rapidly start to form bubbles and then boil. When the reaction is complete, a blue solution of hydrated cupric ions remains. Thus, the cupric ions are not consumed in the process.

**Safety Notes:** Safety goggles are required.

Hydrogen peroxide is an extremely reactive chemical and should be handled with care. Wearing gloves is recommended when handling this chemical. Spills should be cleaned with large volumes of water. If contact is made with your skin, rinse thoroughly with water. The reactions also are quite exothermic and become very hot; contact with the reactions and containers during the course of the peroxide decomposition should be avoided.

The copper salts are toxic chemicals. Gloves should be worn when handling these chemicals.

Unwanted chemicals are not to go down the drain but should be placed in appropriate waste containers.

**Demonstration 7 – Color Change**

Chlorine bleach (sodium hypochlorite, NaClO) is an oxidizing agent. The household chlorine bleach is a 3-8 % solution of sodium hypochlorite (with 0.01-0.05 % NaOH; the NaOH slows the decomposition of NaClO). Chlorine bleach will readily oxidize indigo (Figure 1), the blue dye used in blue jeans.



Materials

50-mL beaker

Stirring rod

Pasteur pipet and bulb

Dimethyl sulfoxide (DMSO)

Indigo

Commercial chlorine bleach

Procedure

Add 2 mg of indigo to 20 mL of DMSO in a 50-mL beaker. Stir until nearly all indigo dissolved; solution should be distinctly blue in color. Using a Pasteur pipet, add about 1 mL of bleach to the solution. Stir until blue color no longer becomes lighter. Continue to add bleach in ~1 mL aliquots until the blue color is gone.

**Safety Notes:** Safety goggles are required.

Commercial chlorine bleach is an extremely reactive chemical and should be handled with care. Wearing gloves is recommended when handling this chemical. Spills should be cleaned with large volumes of water. If contact is made with your skin, rinse thoroughly with water. The reactions also are quite exothermic and become hot, although the beaker only becomes warm to the touch on this scale. The waste cannot be poured down the drain.

**Demonstration 8 – Change of Properties**

A compound is a pure substance that is composed of two or more elements held together by chemical bonds. Compounds have their own unique properties, distinct from the elements that comprise them. However, compound can be broken down chemically into their component elements. An example of the start of such a process is the dehydration of sugar (sucrose, C12H22O11) to produce carbon. The reaction of sulfuric acid with sucrose is very exothermic producing water as steam. This steam carries the carbon with it as a black column of solid. This process is quite visually appealing to an audience; however, the steam also carries dissolved sulfuric acid, son that the demonstration must be performed in a well-functioning fume hood or outside.

C12H22O11(s) 🡪 12 C(s) + 11 H2O(l)

Materials

70 g granulated sugar (sucrose)

70 mL of sulfuric acid (concentrated)

400 mL beaker (tall form)

40 cm stirring rod

100 mL graduated cylinder

weigh boat

non-reactive surface

Procedure

Add the granulated sugar to the tall-form beaker. Place the beaker on a non-reactive surface that is easy to clean and does not melt easily. Add 70 mL of concentrated sulfuric acid to the sugar. Stir with the stirring rod until the column of black solid carbon begins to form and then stand back and allow the column to form. The carbon product should be well washed with water and disposed of in an appropriate container.

**Safety Notes:** Safety goggles are required.

Sulfuric acid is a corrosive chemical and should be handled with care. Wearing gloves is recommended when handling this chemical. Spills should be cleaned with large volumes of water, and the sulfuric acid should be neutralized with an appropriate agent such as sodium bicarbonate. If contact is made with your skin, rinse thoroughly with water. The reaction generates steam that contains dissolved sulfuric acid; thus, the reaction should only be performed in a well-functioning fume hood or outside. The carbon formed can be washed thoroughly to remove any sulfuric acid with water and disposed of in the trash

**Demonstration 9 – Breaking and Making Bonds**

Model kits are useful in demonstrating the structure of molecules and polyatomic ions and their bonding. A kit can be cheaply made from marshmallows and toothpicks. Student will demonstrate the change in connectivity of atoms associated with the recation of hydrogen nd chlorine (Equation 1). For this particular reaction green marshmallows representing chlorine atoms and white mini-marshmallows (representing smaller hydrogen atoms) are needed (Figure 1).

H2(g) + Cl2(g) 🡪 2 HCl(g) (Eqn. 1)

Numerous other reactions could be demonstrated with other colored marshmallows; marshmallows can also be painted to generate the desired colors. The common color scheme to depict atoms is as follows: black (carbon), white (hydrogen), red (oxygen), blue (nitrogen), green (chlorine), and yellow (sulfur).

Materials

White mini-marshmallows

Green marshmallows (often available in a bag of assorted color marshmallows)

Toothpicks

Procedure

Students construct a hydrogen (H2) molecule using two white mini-marshmallows and a toothpick and a chlorine (Cl2) molecule using two green marshmallows. Once made, the students disassemble the two molecule to make two molecules of hydrogen chloride (HCl), where each HCL molecule requires a toothpick, green marshmallow, and white mini-marshmallow.



**Safety Notes:** No safety precautions are necessary.