CHAPTER 15 SULFUR DIOXIDE

SULFUR DIOXIDE HAS BEEN KNOWN SINCE ANCIENT TIMES. In the Odyssey, Homer describes the burning of sulfur (formation of SO_2) as a way to fumigate homes.

Sulfur dioxide is a colorless gas with a suffocating, choking odor. It is toxic to humans and concentrations as low as 8 ppm will produce coughing. Volcanic activity is the primary source of $SO_2(g)$ in nature. Human activity, specifically the combustion of sulfur-rich coal and petroleum, accounts for much more of the $SO_2(g)$ in nature and is the main culprit in acid rain.

Sulfur dioxide has a melting point of -72.7 ^oC. Its boiling point is -10 ^oC, thus, it can easily be condensed to a liquid. Sulfur dioxide is over twice as dense as air.

Sulfur dioxide is produced by the combustion of sulfur or hydrogen sulfide. Sulfur dioxide also is produced by ***roasting*** (heating in the presence of oxygen) metal sulfides such as $FeS_2(s)$. The vast majority of the sulfur produced in the world is converted to $SO_2(g)$ as the first step in the production of sulfuric acid.

In addition to its use as a fungicide, sulfur dioxide has widespread toxic effects on microorganisms and is thus used to disinfect and preserve food and wine. Dried fruits are preserved with $SO_2(g)$ which prevents discoloration (browning) as well as preserves the fruit. Traces of sulfur dioxide can be noticed when smelling or eating dried fruit. Sulfur dioxide is also extensively used in the brewing industry. It is used as a bleaching agent for paper, textiles, oils, etc.

Sulfur dioxide is highly soluble in water. At room temperature, the dissolving process is exothermic:

$$SO_2(g) \rightarrow SO_2(aq) \qquad \Delta H^o = -95.6 \text{ kJ}$$

At room temperature, the solubility of $SO_2(g)$ is approximately 200 g SO_2/L. Thus, 1 mL of water could dissolve 76 mL $SO_2(g)$. The solubility of $SO_2(g)$ in water is highly temperature dependent and is about 400 times more soluble at 0 ^oC (228 g/L) than it is at 90 ^oC (5.8 g/L).

Sulfur dioxide is often thought of as the "anhydride of sulfurous acid", H_2SO_3. However, it is questionable if sulfurous acid exists at all. If it does, the equilibrium constant is so small that fewer than 1 molecule per billion is in the form of sulfurous acid:

$$SO_2(aq) + H_2O(l) \rightleftarrows H_2SO_3(aq) \qquad K << 1 \times 10^{-9}$$

Suitability

All of these experiments are suited for use as classroom demonstrations. These experiments are not advised for use as laboratory experiments conducted by first year high school chemistry students due to the toxicity of SO_2. All of these experiments are suitable for university-level students. The following experiments are included in this chapter.

Experiment 1. Sulfur dioxide reacts with water
Experiment 2. Sulfur dioxide reacts quickly with sodium hydroxide
Experiment 3. Sulfur dioxide and potassium permanganate react
Experiment 4. Sulfur dioxide discolors many natural colors
Experiment 5. Acid-rain microchemistry
Experiment 6. Sulfur dioxide reacts with aqueous bromine

Generally, the production of sulfur dioxide and the experiments that go with this gas should be conducted by individuals familiar with and experienced with gas production using the syringe method introduced in Part 1. Sulfur dioxide has properties that make its proper use and handling more important than was the case for carbon dioxide, hydrogen and oxygen.

All of these experiments serve to demonstrate several basic concepts of sulfur dioxide chemistry including (a) sulfur dioxide is extremely soluble in water (all of the experiments, but especially Experiments 1 and 2); (b) sulfur dioxide is an acid anhydride (Experiment 1, 2, 4 and 5); and (c) sulfur dioxide functions as a reducing agent (Experiments 3 and 6).

Background skills required

Students should be able to:

- generate a gas as learned in Chapter 1.
- know how to prevent accidental/unintentional discharge of gas.
- understand fundamental concepts of high school chemistry so that observations can be interpreted.

Time required

If used as a demonstration, this chapter can be covered in one or two lecture periods. If students are doing these experiments, most, but probably not all of these experiments can be done in one laboratory period. Splitting the experiments between classroom demonstration and laboratory experiment is advisable. For example, during one 45 minute laboratory period, students could do:

Preparation of Sulfur dioxide
Experiment 1. Sulfur dioxide reacts with water
Experiment 2. Sulfur dioxide reacts quickly with aqueous sodium hydroxide
Experiment 3. Sulfur dioxide and potassium permanganate react

The remaining experiments could be performed as classroom demonstrations or selected experiments could be performed during a second laboratory period:

Experiment 4. Sulfur dioxide discolors many natural colors
Experiment 5. Acid-rain microchemistry
Experiment 6. Sulfur dioxide reacts with aqueous bromine

Website

This chapter is available on the web at website:

http://mattson.creighton.edu/Microscale_Gas_Chemistry.html

Instructions for your students

PREPARATION OF SULFUR DIOXIDE[1]

General Safety Precautions

Always wear safety glasses. Gases in syringes may be under pressure and could spray liquid chemicals. Follow the instructions and only use the quantities suggested.

Toxicity

Sulfur dioxide has an irritating odor and is a poisonous gas. Care must be taken when handling SO_2(g). Exposure to concentrations as low as 8 ppm will produce coughing. If you start coughing due to SO_2 inhalation, leave the laboratory to seek fresh air. Deadly concentrations for rats start at 1000 ppm. To put these concentrations in perspective, if 8 mL SO_2 were dispersed evenly into a volume of 1 m^3, the concentration of SO_2 would be 8 ppm. Exercise caution when working with poisonous gases and vacate areas that are contaminated with unintentional discharges of gas.

Because sulfur dioxide is extremely water soluble, wash out syringes with plenty of water to minimize the amount of the gas that dissipates into the room.

Use a fume hood if available

The gas-generation steps should be carried out inside a working fume hood if possible.

Equipment

- Microscale Gas Chemistry Kit (Chapter 1)
- extra-large vial cap (still must fin inside the syringe barrel, but should have greater capacity than the vial cap in the kit)
- hot plate or water heater
- 600 mL beaker
- thermometer

Chemicals

- 4 g sodium hydroxide, NaOH
- 1.5 g sodium bisulfite, $NaHSO_3$(s)
- 3 mL 6 M HCl(aq)
- 5 mL universal indicator solution
- concentrated ammonium hydroxide (only the fumes will be used)

[1] Content for this chapter first appeared as "Microscale Gas Chemistry, Part 9. Experiments with Sulfur Dioxide" Mattson, B.; Anderson, M.; Nguyen J; Lannan, J., ***Chem13 News***, **259**, September, 1997.

The following instructions will produce approximately 40 - 50 mL of $SO_2(g)$. The rate of gas-production of is slow and requires heat to drive the gas from the solution. It typically takes over a minute to collect 40 – 50 mL $SO_2(g)$. The reaction is:

$$NaHSO_3(s) + HCl(aq) \rightarrow SO_2(g) + NaCl(aq) + H_2O(l)$$

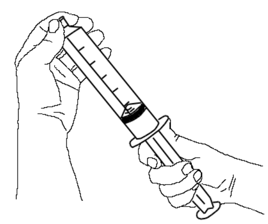
Preparation of Neutralization Solution

Prepare 100 mL of 1 M NaOH (4 g NaOH in H_2O to make 100 mL) in a 250 mL flask. Keep the flask stoppered when not in use. Label the flask "Neutralization Solution, 1 M NaOH". This solution will be used to neutralize excess reagents in the experiments.

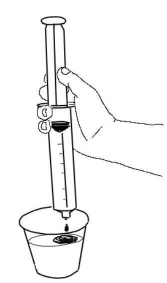
Preventing unwanted discharge of sulfur dioxide

Sulfur dioxide is a noxious gas and must not be discharged into breathable air. The use of syringes to generate such gas samples works exceptionally well and far better than any other method in preventing undesired discharges. There are two simple considerations to keep in mind whenever handling noxious gases:

(1) When opening the syringe (by removing the syringe cap), do so with the plunger slightly withdrawn so the contents are under reduced pressure. Use your thumb to maintain the plunger in this position as shown in the drawing. This will allow some air to enter the syringe but no noxious gas to escape.



(2) After the gas sample has been generated, discharge the used reagents into a cup containing the ***Neutralization Solution***.



Step-by-step instructions for preparing sulfur dioxide

Before starting, have your microscale gas kit, the two reagents needed and the Neutralization Solution ready.

1. Wear your safety glasses!
2. Make sure the syringe plunger and barrel are a good combination for each other and that the plunger moves with reasonable ease in the barrel without binding or sticking.

3. Fill a 600 mL beaker with 400 mL water and start to heat to 70 – 80 ^{o}C. Use a thermometer to assure that the temperature does not exceed 80 ^{o}C.

4. Measure out 0.35 g solid $NaHSO_3(s)$. Place the solid directly into the vial cap to prevent loss.

5. Fill the syringe barrel with water. Place your finger over the hole to form a seal.

6. Float the vial cap containing the solid reagent on the water surface. Lower the cap by flotation. Release the seal made by finger to lower the cap into the syringe barrel without spilling its contents.

7. Install the plunger while maintaining the syringe in a vertical position, supported by the wide-mouth beverage bottle or flask.

8. Fill the weighing dish with 6 M HCl(aq). Draw 3 mL of this solution into the syringe.

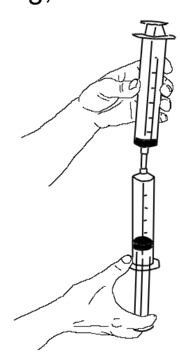
9. Push the syringe fitting into the syringe cap.

10. Place the syringe in the hot water bath. Shake the syringe side-to-side in order to mix the reagents, but do not splash the hot water. Assist the movement of the plunger move up the barrel. Continue this step until 40 – 50 mL gas have been collected. Remove the syringe from the hot water bath.

11. With the plunger slightly withdrawn to assure reduced pressure inside the syringe, and with the syringe held "cap-up", remove the syringe cap.

12. Discharge the liquid reagent into the Neutralization Solution. Immediately cap the syringe to prevent loss of gas.

13. Syringe-to-Syringe Transfer (instead of washing). Do NOT wash the SO_2-filled syringes. Sulfur dioxide is extremely soluble in water. Another simple technique is used to accomplish the same objective. Using a 3 cm piece of tubing, connect the SO_2-filled syringe to a clean dry syringe. Hold the two syringes in a vertical position with the clean, dry syringe on top as in the figure. Transfer the sulfur dioxide to the clean dry syringe by simultaneously pushing and pulling on the two plungers in 10 mL increments. Do not transfer any of the liquid reagent. After transfer is complete, pull the plungers outward by 3- 5 mL to assure reduced pressure in the syringes. Remove the connector tubing and cap the syringes.



Preparation of Sulfur Dioxide in the Microwave Oven

Samples of $SO_2(g)$ also can be prepared conveniently in a microwave oven. See Chapter 24 for details.

Disposal of sulfur dioxide samples

Unwanted SO_2(g) can be converted to water-soluble sulfites by bubbling the gas through the Neutralization Solution.

Teaching tips

1. The rate at which sulfur dioxide can be generated is greatly enhanced by pulling the plunger outward while tapping the syringe.
2. The solubility of a gas in water increases as the temperature decreases.

Questions

1. Write the chemical reaction that took place inside the syringe.
2. Why is sulfur dioxide so soluble in water?
3. The unit "parts per million" (ppm) has a different meaning when it is used to describe a gas mixture rather than a solution or solid mixture. With solids and solutions, the unit refers to the mass relationship and is conveniently thought of in terms of mg solute per kg solution, etc. With gases, however, the unit refers to moles of solute per million moles of total gas. In equation form, it looks like:

$$Conc(ppm) = \frac{10^6 \times n_{solute}}{n_{total}}$$

Calculate the concentration of sulfur dioxide in units of ppm if 10 mL sulfur dioxide were accidentally discharged into 1 m^3 of air.

EXPERIMENTS WITH SULFUR DIOXIDE

Universal Indicator/pH 8 Solution

Experiment 1 and 5 require a slightly basic universal indicator solution. Prepare a solution by mixing 200 mL distilled water plus 20 mL universal indicator solution. Raise the pH to 8 by bubbling through the solution a pipetful of gaseous ammonia taken from the vapors above a solution of concentrated ammonium hydroxide solution.

The following chart provides indicator colors vs. the corresponding pH.

Indicator Colors		
pH	**Universal**	**Red Cabbage**
4.0	Red	Red
5.0	Orange Red	Purple
6.0	Yellow Orange	Purple
7.0	Dark Green	Purple
8.0	Light Green	Blue
9.0	Blue	Blue-Green
10.0	Reddish Violet	Green
11.0	Violet	Green
12.0	Violet	Green
13.0	Violet	Green-Yellow
14.0	Violet	Yellow

EXPERIMENT 1. SULFUR DIOXIDE REACTS WITH WATER

Equipment

Microscale Gas Chemistry Kit

Chemicals

SO_2(g), 20 - 30 mL

universal indicator pH 8 Solution (or cabbage juice)

Suitability

advanced high school chemistry lab, university lab and classroom demonstration

Applications, Topics, Purpose

physical and chemical changes, properties of sulfur dioxide, polar molecules, solutions, the dissolving process, solution equilibrium, acids and bases

Instructions

Generate SO_2 as described. Transfer 50 mL universal indicator pH 8 solution to a plastic cup or beaker. Remove the syringe cap and attach a 15 cm length of tubing to the syringe. Slowly dispense 10 mL of the SO_2 near the surface of the water and notice the production of an acidic solution at the surface. The solution process taking place is:

$$SO_2(g) \rightarrow SO_2(aq)$$

It may be necessary to dispense an additional 10 or 20 mL of the SO_2 to achieve the desired effect.

The ammonia present in the pH 8 solution reacts with sulfurous acid to produce aqueous ammonium bisulfite:

$$SO_2(aq) + H_2O + NH_3(aq) \rightarrow NH_4HSO_3(aq)$$

Teaching tips

1. The universal indicator solution, adjusted to a high pH with ammonia, becomes acidic when exposed to gaseous sulfur dioxide.
2. Sulfur dioxide is a Lewis acid in water. The water acts as a Lewis base.

$$SO_2(aq) + H_2O \rightleftarrows H^+(aq) + HSO_3^-(aq) \text{ or}$$

$$SO_2(aq) + 2\ H_2O \rightleftarrows H_3O^+(aq) + HSO_3^-(aq)$$

3. In general, most oxides of nonmetals such as sulfur are acidic in water solution.
4. Refer to the chart of indicator colors at the beginning of this section.
5. Most texts show the reaction of $SO_2(aq)$ and water to form sulfurous acid, $H_2SO_3(aq)$. This has an uncertain existence. The correct form is simply $SO_2(aq)$.

Questions

1. Based on your observations during the demonstration, does $SO_2(g)$ form an acidic or basic solution?
2. Why was it unnecessary to bubble the sulfur dioxide below the surface of the water? (Recall that the gas was discharged near the surface of the water.)
3. Sulfur dioxide, like many non-metal oxides, forms acidic solutions when dissolved in or reacted with water. Many non-metal oxides such as $Na_2O(s)$, form bases when reacted with water. In this example, sodium hydroxide would be formed. Write the reactions that takes place when: (a) sodium oxide and water react; (b) diphosphorus pentoxide and water react.

EXPERIMENT 2. SULFUR DIOXIDE REACTS QUICKLY WITH AQUEOUS SODIUM HYDROXIDE

Equipment

Microscale Gas Chemistry Kit

Chemicals

SO_2(g), 60 mL, two syringes full

50 mL 3 M NaOH

Suitability

advanced high school chemistry lab, university lab and classroom demonstration

Applications, Topics, Purpose

physical and chemical changes, properties of sulfur dioxide, chemical reactivity of sulfur dioxide, solutions, acids and bases

Instructions

Add 50 mL of water to a beaker. Add 50 mL 3 M NaOH to another beaker. Generate two syringes full of SO_2. Replace the syringe caps from a SO_2-filled syringes with 15 cm lengths of tubing. Hold the syringe by the barrel and not by the plunger for this next part! First, draw 2-3 mL of water into the fist syringe of SO_2(g) and then pinch the tubing closed with your fingers. Repeat the experiment using the second syringe of SO_2(g) and the 3 M NaOH solution. The reaction in the second part is:

$$SO_2(g) + NaOH(aq) \rightarrow NaHSO_3(aq)$$

Teaching tips

1. The plunger will rapidly be pulled inward as the SO_2 reacts.
2. Sulfur dioxide is very soluble in water because both molecules are polar covalent and form hydrogen bonding attractive forces with each other.

Questions

1. Identify the acid and base in the reaction between SO_2(g) and NaOH(aq).
2. In which syringe did the sulfur dioxide dissolve and in which syringe did it react?
3. Sketch the Lewis structure of one water molecule and one sulfur dioxide molecule with the hydrogen atom of water close to a oxygen atom of sulfur dioxide. Draw a dotted line from the water's hydrogen atom (d+) to an oxygen atom (d-) on sulfur dioxide. This picture will illustrate one possible attractive force between these two molecules.

EXPERIMENT 3. SULFUR DIOXIDE AND POTASSIUM PERMANGANATE REACT

Equipment

Microscale Gas Chemistry Kit

Chemicals

SO_2(g), 60 mL

dilute aqueous solution of potassium permanganate (See Appendix D)

1 M HCl(aq), 25 mL

Suitability

advanced high school chemistry lab, university lab and classroom demonstration

Applications, Topics, Purpose

oxidation-reduction reactions, chemical formulas, chemical reactions, writing balanced chemical equations, classifying chemical changes

Instructions

Prepare a very dilute aqueous solution of potassium permanganate. Prepare a syringe full of SO_2(g) as described above and transfer to a clean, dry syringe. Pour 5 mL of the permanganate solution into the weighing dish and draw the solution into the SO_2-filled syringe. Cap with the syringe cap and shake the solution. The pink color of permanganate will disappear. The balanced equation for the oxidation by permanganate under neutral or slightly acidic conditions is:

$$5\ SO_2(g) + 2\ MnO_4^-(aq) + 2\ H_2O(l) \rightarrow 2\ Mn^{+2}(aq) + 5\ SO_4^{-2}(aq) + 4\ H^+(aq)$$

Teaching tips

1. This demonstration can be performed without worrying about balancing the oxidation-reduction reaction that takes place. On the other hand, this is a good demonstration of a oxidation-reduction reaction.
2. To clean-up the brown stains left from permanganate solutions, soak the syringes in 1 M HCl(aq).

Questions

1. What are some of the indications that a reaction has taken place?
2. Balance the oxidation-reduction reaction that takes place if the manganese product were MnO_2(s) rather than Mn^{+2}(aq).
3. Given that Mn^{+2}(aq) is colorless solution, how could you be sure whether or not you produced Mn^{+2}(aq) or MnO_2(s)? Both are plausible and each is produced under certain conditions.

EXPERIMENT 4. SULFUR DIOXIDE DISCOLORS MANY NATURAL COLORS.

Equipment

Microscale Gas Chemistry Kit

Chemicals

SO_2(g), 60 mL, several syringes needed

Materials

flowers (impatiens, African violets, red roses work well)
radish
fruit drinks (optional)

Suitability

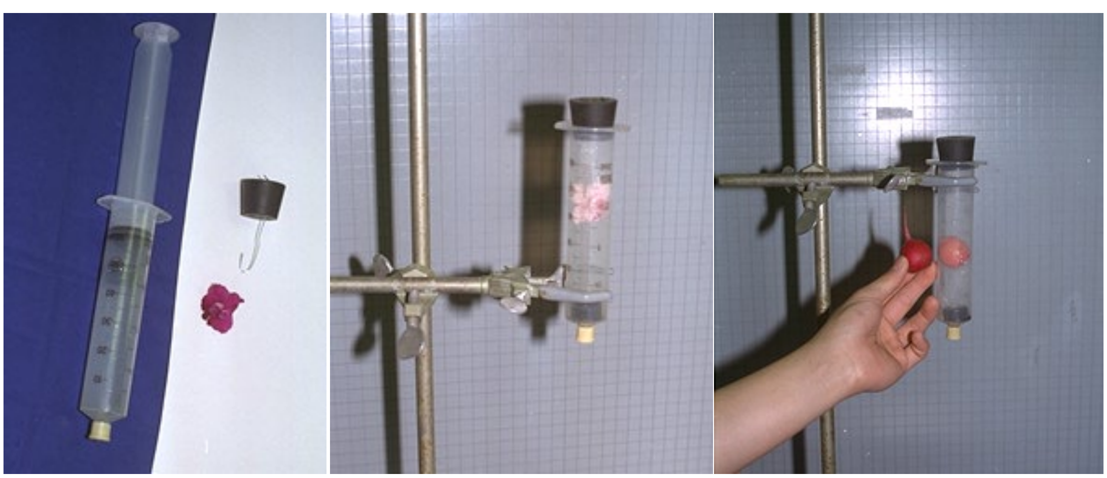
advanced high school chemistry lab, university lab and classroom demonstration

Applications, Topics, Purpose

household/familiar materials, bleaching, dyes, colored substances, physical and chemical changes, acids and bases, oxidation-reduction reactions

Instructions

Radishes and certain flowers including purple African violets and red roses are rapidly discolored by exposure to sulfur dioxide. Generate a several syringes full of SO_2. Place flowers or a radish in the syringe body of a clean syringe. Install the plunger as far as possible, but do not crush the flowers. With the short connector tubing, transfer 20 – 30 mL SO_2(g) to the syringe with the flowers or radish. As usual, it may be necessary to assist the movement of the plunger outward as the gas is being pushed inward. After 20 – 30 mL of gas has been transferred, pull back 5 mL on the plunger of the syringe containing the flower/radish to create a reduced pressure. Remove the connector tubing and cap both syringes. Observe the experiment over the next several minutes and again after one hour and again after overnight. The flower color change takes place within seconds. The radish becomes noticeably lighter in color within a few minutes and white overnight. Some fruit drinks also are affected by exposure to SO_2(g).



Teaching tips

1. Not all flowers and fruit drinks are discolored.
2. Use dry syringes! Sulfur dioxide will preferentially dissolve in drops of water rather than interact with the object to be discolored.
3. The explanation for the discoloration is based on one of two propertied of $SO_2(g)$. Sulfur dioxide is an acid anhydride and thus forms acidic solutions. If the pigment in the flowers is acid-base sensitive (an indicator), the color change could be thus explained. Sulfur dioxide is also a mild reducing agent and if the pigment can be reduced to a non-colored or differently colored compound that may occur. Generally, the acid-base explanation is more likely the correct one. For example, in Experiment 1, the acidic nature of sulfur dioxide was demonstrated. In Experiment 3, the reducing ability of sulfur dioxide was demonstrated.

Questions

1. When a color change occurs, some sort of chemical reaction must have taken place. Indicators change color due to a very simple, *reversible* reaction in which the form of the indicator in acid has an extra H^+ associated with it. As you have seen in Experiment 1, sulfur dioxide forms acidic solutions. The key word is *reversible* — if the pH were raised, the original color of the indicator would return. The other possible explanation for the color change caused by sulfur dioxide is that the compound reduced the natural pigment. Oxidation-reduction reactions are not generally as reversible as acid-base reactions. How might you determine whether an acid-base reaction or oxidation-reduction reaction took place with the natural pigment?
2. Classify these reactions as either an acid-base reaction or an oxidation-reduction reaction.

   (a) $3\ SO_2(g) + 2\ Cr^{+6}(aq) + 3\ H_2O(l) \rightarrow 3\ SO_4^{-2}(aq) + 2\ Cr^{+3}(aq) + 6\ H^+(aq)$

   (b) $SO_2(g) + NH_3(aq) + H_2O(l) \rightarrow HSO_3^-(aq) + NH_4^+(aq)$

EXPERIMENT 5. ACID RAIN MICROCHEMISTRY[1]

Equipment

Microscale Gas Chemistry Kit
24-well plate
gallon-sized (4 L) sealable plastic food storage bag

Chemicals

SO_2(g), 60 mL
universal indicator/pH 8 solution, 150 mL

Suitability

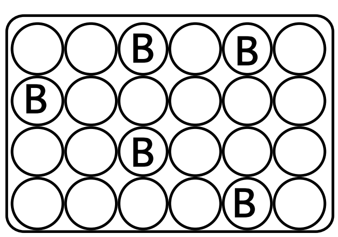
advanced high school chemistry lab, university lab and classroom demonstration

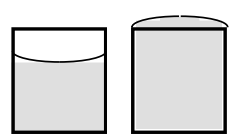
Applications, Topics, Purpose

environmental chemistry, acid-base reactions, diffusion, gas solubility, acid anhydrides

Instructions

Coal combustion produces sulfur dioxide that acts to produce acid rain. In this experiment, a 24-well plate is used to create a series of “lakes”, five or six of which are buffered. The 24-well plate is enclosed in a plastic bag to create an ecosystem. The layout of a typical ecosystem is shown in the left figure below. The "B" marks, randomly placed, indicate the six lakes that will be buffered. Fill all 18 of the unlabeled wells with this solution.



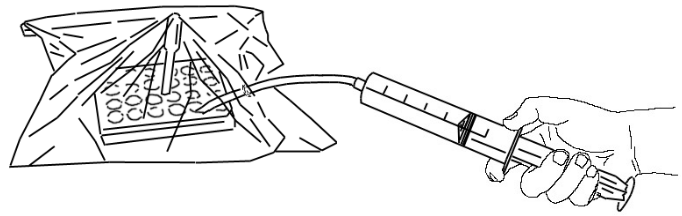


For use on an overhead projector: overfill the wells

Use a pipet to add the final drops to each well. To the remaining universal indicator solution, dissolve 0.1 g of sodium bicarbonate, $NaHCO_3$. Fill the remaining six lakes with this solution. Place a 6 cm length of a plastic pipet stem between the four middle wells in order to prop up the plastic bag above the surface of the filled wells. Pierce a small hole through the bag with a sharp pencil and work the tubing through the hole as shown in figure. (Moistening the tubing with soapy water helps to facilitate this process.) Next, slip the filled well plate into a plastic bag and zip shut.

[1] ***Chemistry Demonstration Aids That You Can Build!"*** Bruce Mattson, Mary Alice Kubovy, Jeff Hepburn, and Joe Lannan, Flinn Scientific Press. 1997.

Place the assembly on the overhead projector. Discharge the gas into the bag. As the gas drifts across the "landscape", the unbuffered lakes will become acidic. The buffered lakes will eventually become acidified as well. The entire acidification process takes 1-2 minutes for the unbuffered "lakes" and over 5 minutes for the buffered ones.



Teaching tips

1. Clean-up. Allow the bag to stand overnight – the $SO_2(g)$ will dissolve in the water.

2. Sulfur dioxide is produced when coal is burned. Coal naturally contains varying amounts of sulfur compounds, all of which leaves the smoke stack in the form of $SO_2(g)$. In the atmosphere, sulfur dioxide is usually converted to sulfur trioxide which reacts with water to form sulfuric acid.

$$SO_2(g) \rightarrow SO_2(aq)$$

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

   Both $SO_2(aq)$ and $H_2SO_4(aq)$ fall to earth with rain. In this acid rain demonstration, $SO_2(g)$ and aqueous sulfur dioxide are acidifying the lakes while in our environment, $SO_3(aq)$ and $H_2SO_4(aq)$ are the normal culprits.

3. The bicarbonate ion, $HCO_3^-(aq)$ of sodium bicarbonate is a base in water. The sodium ion is a neutral ion.

4. Most lakes are naturally slightly alkaline due to sedimentary materials, which contain carbonates. Acid rain causes the pH of lakes to drop to levels in which microorganisms cannot survive. Because many of the microorganisms affected form the foundation of the food chain, all life in the lake suffers.

5. Refer to the pH color chart at the beginning of this section.

Questions

1. How do the buffered lakes compare with the unbuffered lakes?

2. What was the pH of the buffered and unbuffered lakes before adding the $SO_2(g)$? How does the pH change for each kind of lake during the first ten minutes after addition of the $SO_2(g)$?

3. What is the function of a buffer?

4. Is the acid rain in this experiment more likely to be caused by cars or by coal burning industries?

EXPERIMENT 6. SULFUR DIOXIDE REACTS WITH AQUEOUS BROMINE

Equipment

Microscale Gas Chemistry Kit

Chemicals

$SO_2(g)$, 60 mL

2 – 3 mL dilute aqueous solution of bromine (See Appendix D)

sodium bisulfite solution, $NaHSO_3(aq)$, 1 M

Suitability

advanced high school chemistry lab, university lab and classroom demonstration

Applications, Topics, Purpose

chemical changes, properties of sulfur dioxide, oxidation-reduction reactions

Instructions

Prepare a syringe full of $SO_2(g)$ and transfer to a clean, dry syringe. Pour 3 - 4 mL of water into weighing dish. Add several drops of the aqueous bromine solution until the solution is definitely colored. Draw the aqueous bromine solution into the syringe. Install the syringe cap and shake the solution. The red color of bromine will slowly disappear. The reaction under neutral or slightly acidic conditions is:

$$SO_2(aq) + Br_2(aq) + 2\ H_2O(l) \rightarrow 2\ Br^-(aq) + SO_4^{-2}(aq) + 4\ H^+(aq)$$

Excess aqueous bromine solution can be treated with 1 M $NaHSO_3(aq)$ until the color disappears and then safely discarded down the drain.

Teaching tips

1. Handle bromine water with extreme care.
2. Destroy all excess bromine water solutions as per the Instructions (last paragraph)

Questions

1. Which reactant was the oxidizing agent in this redox reaction? Which reactant was the reducing agent?
2. Write the two half-reactions that occurred in the syringe. Label one with oxidation and the other with reduction.
3. How might you confirm the presence of bromide ion?
4. Assign oxidation numbers to sulfur in SO_2 and SO_3^{-2}. What do you notice about the two oxidation numbers? Repeat for SO_3 and SO_4^{-2}.
5. From a chemical reactivity standpoint, the oxysulfur compounds with the same oxidation state will have similar oxidation-reduction properties. Make two pairs

from this list of four species with similar oxidation-reduction properties, SO_2, SO_3, SO_3^{-2} and SO_4^{-2}.

Clean-up and storage.

At the end of the experiments, clean the syringe parts, caps and tubing with water. Rinse all parts with distilled water if available. Be careful with the small parts because they can easily be lost down the drain. **Important:** Store plunger out of barrel unless both are completely dry.

Summary of Materials and Chemicals Needed for Chapter 15. Experiments with Sulfur dioxide

Equipment required

Item	Demo	For 5 pairs	For 10 pairs
Gas Chemistry Kit	1	5	10
24-well plate	1	2 - 3	4 – 5
extra-large vial cap	1	5	10
hot plate or water heater	1	5	10
600 mL beaker	1	5	10
thermometer	1	5	10

Materials required

Item	Demo	For 5 pairs	For 10 pairs
4 L (1 gallon) sealable plastic food storage bag	1	2 - 3	4 – 5
flowers*	a	a	a
radish	1	demo	demo
fruit drinks	a	a	a

* impatiens, African violets, red roses work well

a. brought in by students

Chemicals required

Item	Demo	For 5 pairs	For 10 pairs
sodium bisulfite, $NaHSO_3(s)$	2 g	10 g	50 g
sodium bisulfite, $NaHSO_3(aq)$, 1 M	5 mL	25 mL	125 mL
6 M HCl(aq)	10 mL	50 mL	100 mL
1 M HCl(aq)	25 mL	125 mL	250 mL
sodium hydroxide, NaOH	4 g	20 g	40 g
sodium hydroxide, 3 M NaOH	50 mL	250 mL	500 mL
universal indicator solution*	3 mL	15 mL	30 mL
concentrated ammonium hydroxide	< 1 mL	< 1 mL	< 1 mL
potassium permanganate	< 0.1 g	< 0.1 g	< 0.1 g
sodium bromide, NaBr	1.1 g	a	a
sodium hypochlorite **	8 mL	a	a

* or cabbage juice (you need to use more cabbage juice than universal indicator)

** household laundry bleach

a See Appendix D; recipe for “Demo” makes enough for entire group