Pouring and Siphoning a Gas

Density of Gases

Introduction

When most students hear the word "fluid" they automatically think of liquids—something that flows with no rigid shape of its own, something that can be poured from one container into another. Gases are fluids, too. Perhaps it is because of their invisible nature or their lack of any self-defined boundaries that gases maintain such a low profile. Thus it is certainly an eye-opener when students see (or rather, don't see) a dense gas such as carbon dioxide being poured or even siphoned from one container into another.

Concepts

- Density
- Properties of carbon dioxide
- Index of refraction
- Sublimation

Materials

Baking soda (sodium bicarbonate), $NaHCO_3$, 10 g

Dry ice, 2–3 thumb-sized chunks (optional)

Vinegar (5% acetic acid), 200 mL

Butane safety lighter

Plastic soda bottles, preferably colorless, 2-L, 2

Plastic tubing, $^1/_2$″ ID, 70 cm*

Scissors, heavy duty

Stiff baling wire (or coat hanger wire), 40 cm

Tape

Tea light candle

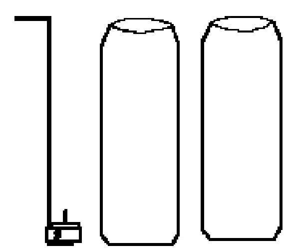
**Use a new piece of tubing, not one that may have been used in the chemistry lab!*

Safety Precautions

Care should be taken when dealing with open flames. Remove all combustible materials from the vicinity of the demonstration. Dry ice is cold enough to cause frost bite—wear well insulated gloves when handling it. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Follow all laboratory safety guidelines. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Preparation

1. Rinse out the 2-L soda bottles and remove labels.
2. Cut off the very tops of the soda bottles to make two tall plastic "beakers," leaving the sides curved in a bit at the top.
3. Bend the wire into a "Z" a few centimeters taller than the beakers, and then tape the candle to the bottom (see Figure 1).

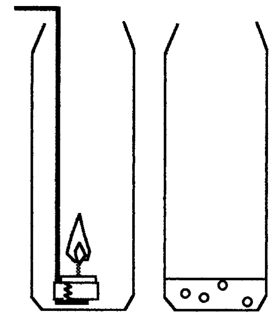


Procedure

Pouring a Gas

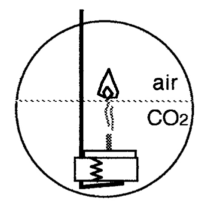
1. Gesturing to the two "beakers," ask the students what they think is in them (consensus: Air).
2. Place two rounded teaspoons (approx 10 g) of baking soda in the bottom of one of the beakers and add 1 cup (250 mL) of vinegar. Ask the students what they observe (chemical reaction, bubbles, gas being produced). Ask if they know what that gas is; you may give them the hint that baking soda is called "sodium bicarbonate" (CO_2). Ask if they know what carbon dioxide is often used for (fire extinguishers). And finally ask them where they think the CO_2 product is now (many think that it has probably diffused out into the room). "Well let's see . . ."
3. Light the candle and ask what enables it to burn (air—or more precisely, the oxygen in the air). Have them predict

what will happen when the candle is lowered into the unused beaker, still presumably filled with air.



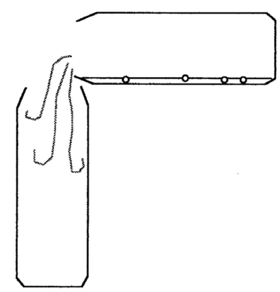
4. Carefully lower the burning candle into the beaker and observe that it continues to burn. This serves as a control to confirm that the act of lowering the candle into a beaker does not put out a flame (see Figure 2).

5. Remove the candle from the control beaker and lower the candle into the beaker in which the baking soda and vinegar reacted (but not into the liquid at the bottom), and observe that it is quickly extinguished. *Note:* If the candle is lowered slowly enough, and then held just below the CO_2–air interface, sometimes a very unusual phenomenon occurs—the bottom of the flame is extinguished, but the very top continues to burn. That is, the flame gets lifted off the wick—sometimes by as much as a centimeter or two—with a thin trail of smoke extending from the tip of the wick to the bottom of the flame, as shown in Figure 3. If the candle is then lifted back up, the flame reattaches itself to the wick! Achieving this effect is not easy, and it helps to have the CO_2 freshly generated, so that the CO_2–air interface is very sharp and well-defined, not "blurred" as it becomes due to mixing and diffusion if too much time has passed between the generating of the CO_2 and the introduction of the flame.



6. Once it has been established that the CO_2 gas is dense enough to "sit" down in the beaker, as though it were a liquid, pose the question, "Do you think that the carbon dioxide gas can be poured like a liquid from one container into another?"

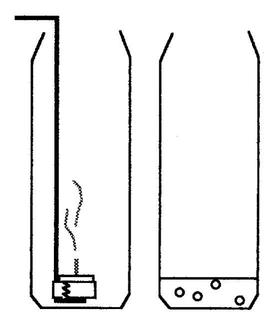
7. Place the candle aside and slowly pour the CO_2 from the reaction beaker into the control beaker, being careful to pour only gas, not any of the liquid (see Figure 4). Since there is virtually no visible evidence that the gas is being poured from one beaker to the other, the students will think you have lost your mind, especially if you say things like "I've got to be careful not to spill any. . . Oops, some just fell on the floor. . . . I'll clean it up later. . . . There, almost done . . ." *Note:* A very slight Schlieren effect may be observed as the CO_2 is poured through the air. This is similar to the wavy mirage-like distortions that can be seen in the air over asphalt or car tops on a hot day.



8. Relight the candle, then prove that the CO_2 has been transferred by lowering the lit candle into the original reaction beaker, and observe as it continues to burn. *Note:* If all the CO_2 was not poured out, the candle may go out if it is lowered too far into the beaker.

9. Students will demand better proof than that, so then lower the lit candle into the other beaker, the supposed recipient of the CO_2, and observe that the flame is extinguished (see Figure 5).

Siphoning a Gas



10. Empty the CO_2 out of the second beaker, by simply holding it upside-down for 5–10 seconds, and then place it upright on the table.

11. To verify that the CO_2 is gone, lower the lit candle back into the beaker and show that it continues to burn.

12. Use some additional vinegar and baking soda to generate some fresh CO_2 in the reaction beaker.

13. Insert one end of the plastic tubing down into the reaction beaker so that it rests 2–3 cm above the vinegar level.

14. Hold this beaker up high (level with your head), suck quickly on the other end of the hose, (You are likely to get a little vinegary taste of CO_2—like a combination of salad dressing and tonic water—but it won't last long!) and then immediately transfer the end of the hose into the empty beaker (see Figure 6). The reaction beaker does not have to remain as high as your head, but it does have to remain 20–30 cm higher than the empty beaker.

15. Hold the beaker and hose in this configuration for 20–30 seconds. (If you thought it was hard to tell when the CO_2 had finished pouring, try to tell when it's through being siphoned!)

16. Lower the beaker and remove the hose. Acknowledge that if there is any CO_2 in the lower beaker, it must have gotten there through the hose.

17. Test each beaker with the burning candle. The candle will be extinguished as it reaches the CO_2–air interface, about ½ to ⅔ of the way down into the reaction beaker (for it was not necessarily siphoned completely empty) and about ⅓ to ½ of the way down into the second beaker. The amount of CO_2 transferred depends, of course, on the height differential between the two beakers during the siphoning, the length of time the siphon was allowed to run, as well as the internal diameter of the hose.

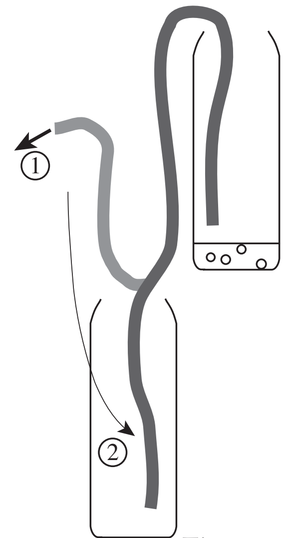


Figure 6

Disposal

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

Tips

- Students may wonder how long the CO_2 will remain in the beaker. If it is truly like a liquid, one would imagine it should remain in the beaker indefinitely. Pour a beaker of CO_2, but then wait two minutes before testing it with the burning candle. Try waiting five minutes . . . ten minutes . . .Generally, the CO_2 will remain in the beaker for 1–2 minutes, but not for more than 5–6 minutes. It eventually diffuses out into the room. This variation leads to a number of interesting experiments. Does a screen over the top affect how long the CO_2 will remain? How about a piece of paper or some plastic wrap? What effect does temperature have on this rate of diffusion?
- Certainly the impact of the demonstration hinges on the student not seeing a gas being poured, but then seeing evidence that indeed it must have been poured. As another possible follow-up demonstration, repeat the CO_2 pouring,but do so in front of a screen, with the overhead projector shining light on the beakers and the students watching the shadows. The aforementioned Schlieren effect can thus be accentuated. The students will find it interesting that although the CO_2 is virtually invisible, its shadow can easily be seen.
- Another way to make the CO_2 visible is to "color" it! Repeat the demonstration above, but instead of baking soda and vinegar, place 100 mL of water and two to three thumb-sized chunks of dry ice in the reaction beaker. The beaker will quickly fill with CO_2, colored whitish gray by a fine fog of water droplets. (These droplets form instantaneously as the vapor from the water condenses in the cold CO_2 gas, given off by the subliming dry ice.) Now when the CO_2 is poured or siphoned across into the second beaker, it is quite easy to see.
- Avoid breathing, talking, and definitely do not laugh directly toward the beakers as you are pouring the CO_2. With such turbulence, you might find that very little of the CO_2 makes it into the second beaker!
- Avoid pouring across any of the solution at the bottom of the reaction beaker, for if that gets poured across, students could argue that it was the transferred solution, still generating CO_2 that caused the flame to be extinguished.
- You may want to practice the pouring and/or the siphoning using dry ice CO_2 first, to enable you to see how slowly it pours across. *Note:* The slow pouring is not indicative of CO_2 having a high viscosity. Instead it reflects how lightweight CO_2 is—compared to something like water—and how air resistance therefore causes it to reach a very slow terminal velocity and fall very slowly.

- You may wish to use a vacuum pump (or a pipet bulb, or even a nasal aspirator) to initiate the siphon, especially if you are doubtful of the cleanliness of the hose, or if you are performing the dry ice variation and are concerned that a small piece of dry ice could get sucked up through the hose! This has never happened; in fact, it is virtually impossible to get it to happen intentionally. Still, it poses a potential risk and therefore warrants consideration.

Discussion

The nice thing about this demonstration of the invisible pourability of CO_2 is that it illustrates quite a few chemical and physical properties of the compound.

- It can be generated by the reaction of baking soda and vinegar (see Equation 1).

$$NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) + CO_2\,(g) \qquad \textit{Equation 1}$$

- It does not support combustion.
- It is a gas at room conditions.
- It is clear and colorless.
- Because it is a gas, it has low viscosity.
- Because of its high molar mass, it is denser than air.
- Its index of refraction is close to, but not quite equal to that of air (see second *Tip*).
- Due to Brownian movement, it eventually diffuses out in all directions to occupy the entire space permitted (see first *Tip*).

The CO_2 molecule is linear in shape (see Figure 7). The C=O double bonds are polar due to the very electronegative oxygen atoms "hogging" the bonding electron pairs. Because of its linear symmetry, however, the molecule as a whole is nonpolar; it has no permanent dipole moment. As a result, the attractive forces between the molecules are rather weak (only London forces), and easily overcome by the kinetic energy of the molecules even at relatively low temperatures. Thus CO_2 is a gas at room temperature and remains a gas all the way down to –78°C, at which point the intermolecular forces take over and pull the substance into a solid. *Note:* The liquid state of CO_2 is obtainable; it is just not stable at pressures as low as 1 atm.

The particles in a gas are much farther apart than those in solids or liquids. Thus the volume of a gas sample depends almost entirely on the empty space between the particles. In turn, this spacing between gas particles is determined largely by the temperature and the pressure of the gas sample, and hardly at all by the particles themselves. What this means is that a collection of helium particles will have essentially the same volume as an equivalent number of neon, oxygen, or carbon dioxide particles, provided the gases are all at the same conditions of temperature and pressure. This is known as Avagadro's law, and it applies to all gases, but not to solids or liquids. One result of Avagadro's law is that one can simply use the molar masses of gaseous substances to predict their relative densities. We would expect carbon dioxide with a molar mass of 44.0 to be about 50% denser than air (average molar mass = 28.9), and it is.

One characteristic that liquids and gases do have in common is that their particles are free to move around one another. They are not locked into position the way they are in solids. This means neither liquids nor gases have defined shapes. Instead they flow to take on the shape of the container they are in. Substances that flow are referred to as *fluids*. The mobility of the particles, however, is not an all-or-none affair: The particles in some substances move, but with some degree of resistance, because of the attractive forces acting between them. This resistance to flow is known as *viscosity*. Gaseous substances like carbon dioxide, with their large spacings and weak attractions between the particles, are considered to have relatively low viscosities, and they can thus be poured quite easily. The pouring appears almost invisible in the demonstration because CO_2 is clear and colorless. But then again, water is clear and colorless, and yet it is quite easy to see when water is being poured from one container to another. A more complete explanation for why CO_2 pouring is nearly invisible would have to include mention of the index of refraction of carbon dioxide. *Index of refraction* is a physical property that pertains to how much light is slowed down as it travels through a substance. More specifically it is the ratio of light's velocity through a vacuum to light's velocity through the particular substance.

What enables us to see water, even though it's clear and colorless, is the fact that water's index of refraction (1.333) is substantially greater than that of air (1.00029), thus a beam of light slows down and gets bent quite significantly as it passes from air to water. Likewise, the light beam speeds up and gets bent again as it passes back from water to air. This phenomenon, called refraction, enables us to discern quite easily the interface between water and air. Reflection off the water's surface also helps. CO_2's index of refraction (1.00045) is so close to that of air that the interface is much harder to discern. Also, the

CO_2–air interface becomes considerably more blurred, especially over time, as the CO_2 starts to diffuse out and mix with the air. Nevertheless, the aforementioned Schlieren effect, although very faint, does attest to the fact that the CO_2–air interface is visible, and this Schlieren is a direct result of the slight refraction that occurs as light passes from air into CO_2 and then back out into air. The only way the Schlieren effect would not occur when one clear, colorless fluid is poured through another, would be if the two substances had identical indexes of refraction.

Diffusion is one other phenomenon that warrants discussion here. The random spreading out of the gas particles from an area of high concentration to an area of lower concentration, diffusion not only blurs the CO_2–air interface, it eventually leads to the complete evacuation of CO_2 out of the beaker (see *Tips* section). This might be seen as another difference between CO_2 and water—although they can both flow and can assume the shape of the containers they are placed in, water remains in its container, but CO_2 diffuses out. Yet, this also is only a difference by degree—the water would eventually diffuse completely out of the container as well, but since water's diffusion involves a phase change, it is called *evaporation*. And because the attractive forces are considerably greater between water molecules, the process takes considerably longer—months instead of minutes!

Connecting to the National Standards

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

Unifying Concepts and Processes: Grades K–12
Evidence, models, and explanation
Content Standards: Grades 5–8
Content Standard B: Physical Science, properties and changes of properties in matter
Content Standards: Grades 9–12
Content Standard B: Physical Science, structure and properties of matter, chemical reactions

Acknowledgments

The idea for siphoning the CO_2 was suggested by Ron Perkins of Greenwich High School, Greenwich, CT.

Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Pouring and Siphoning a Gas* activity, presented by Bob Becker, is available in *Density of Gases* and in *Properties of Carbon Dioxide*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

Materials for *Pouring and Siphoning a Gas* are available from Flinn Scientific, Inc.

Catalog No.	Description
S0043	Sodium Bicarbonate, 500 g
V0005	Vinegar, White, 3.78 L
AP8378	Plastic Tubing, $^{1}/_{2}$9 ID, 10-ft
AP8960	Butane Safety Lighter

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