A POURABLE Greenhouse Gas

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ILLUSTRATIONS BY CEASR CAMINERO

Carbon dioxide (CO_2) gas in our atmosphere has the capacity to absorb escaping infrared light and "reradiate" it back down toward the Earth's surface. The role of this atmospheric CO_2 is similar to that of the glass panes in greenhouses that keep indoor temperatures high enough for plants to thrive during winter months. During the long history of life on our planet, CO_2 did its part to keep Earth warm and inhabitable. But if serious measures are not taken soon to curb the increasing amounts of CO_2 that we release into the atmosphere, we could end up with a very different climate than what we have now.

Good ideas about controlling and containing CO_2 start with the knowledge of its chemical and physical properties. Besides trapping radiant energy, carbon dioxide has some other remarkable features. In this investigation, you can find out how carbon dioxide behaves when it is released into the atmosphere.

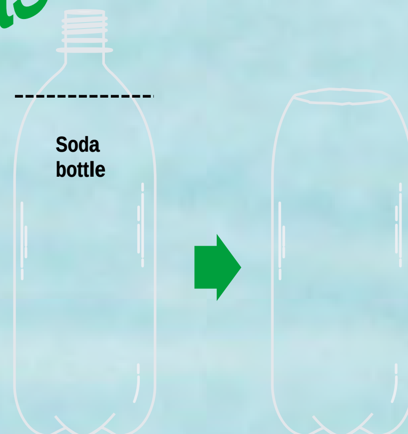


Figure 1

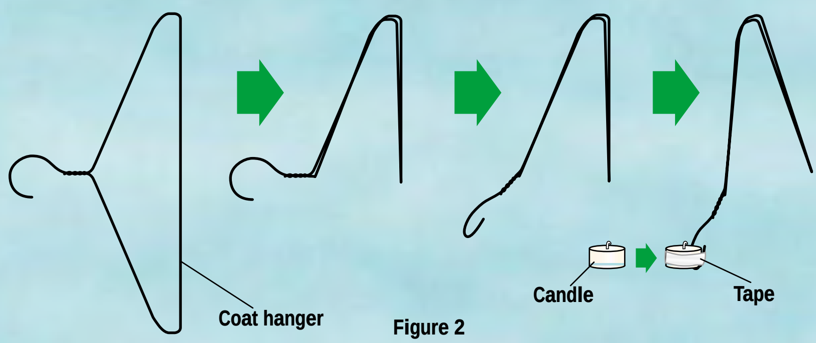


Figure 2

Materials and tools for each group

Safety goggles, apron, and disposable plastic gloves for each student
Two 2-L soda bottles (empty, rinsed out, labels removed)
scissors and utility knife
small, flat candle (sold as "tea light" or "votive")
wire coat hanger
tape
matches
baking soda
vinegar

For extension activities (optional)

0.5-inch-diameter plastic tubing, 60–70 cm long (available at hardware stores)
disposable alcohol pad in sealed package
dry ice (about 1 lb for entire class) in chips (available from ice cream stores and frozen food suppliers)
water

Safety notes: This activity is designed for a supervised safety-equipped chemistry classroom. Students should wear safety goggles, aprons, and disposable gloves throughout the procedure. All combustible materials (papers, books, etc.) should be removed from the area surrounding the open flame. Standard chemistry lab safety rules should be observed.

Preparation

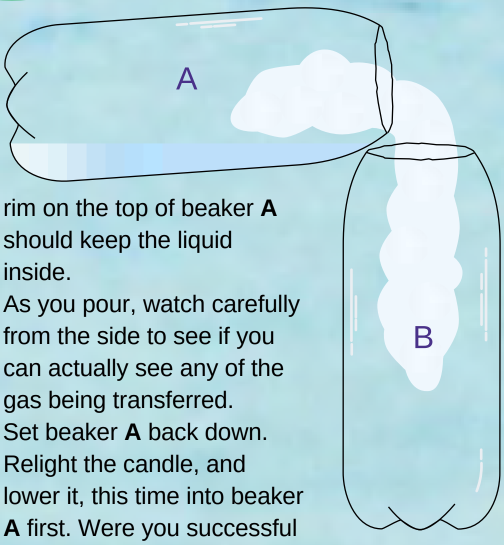
1. Cut the tops off both soda bottles as shown in Figure 1. Use a utility knife to start the cut, then a good sharp pair of scissors to continue it all the way around. When finished, you will have two tall plastic beakers with rims bent slightly inward at the top. Label them "A" and "B".
2. Bend the coat hanger in half, as shown in Figure 2. Bend the hook down and flatten it. Tape the candle to it securely.

Activity

Tip: For best results, this activity should be conducted in a place with very little draft or air movement.

1. Place about 120 mL (1/2 cup) of baking soda into beaker **A**. Pour in about 240 mL (1 cup) of vinegar and observe the reaction. Those are bubbles of CO_2 gas you are seeing!
2. Wait a minute or so to give the CO_2 plenty of time to diffuse out into the room.
3. Light the candle with a match, and slowly lower it into beaker **B**. This serves as a control to show that a beaker filled with regular air does not affect the candle flame.
4. Remove the candle from beaker **B** and, while it is still burning, lower it slowly into beaker **A**.
   What did you observe? What ***chemical property*** of carbon dioxide did this demonstrate? What ***physical property*** did it demonstrate? Why do you think the CO_2 is still in the beaker? How long do you think it would stay in there?
5. Pick up beaker **A** and try to pour the CO_2 slowly over into beaker **B.** Carefully pour only the gas and none of the liquid across. The curved

for you to try!



rim on the top of beaker **A** should keep the liquid inside.

As you pour, watch carefully from the side to see if you can actually see any of the gas being transferred.

6. Set beaker **A** back down. Relight the candle, and lower it, this time into beaker **A** first. Were you successful at pouring out all of the CO_2? What evidence supports your answer?
7. Relight the candle if necessary, and then lower it into beaker **B**. Did the CO_2 make it into beaker **B**? What evidence supports your answer? How many times do you think you could pour the CO_2 back and forth? Continue experimenting to test your prediction.

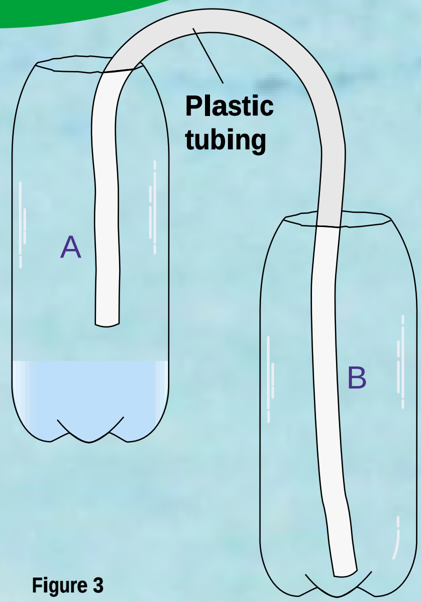


Figure 3

Extension Activities (optional)

Clean and dry beaker **A**, and use it to generate more carbon dioxide following the procedure in step 1. This time, rather than pouring the CO_2 from **A** to **B**, try *siphoning* it across. To do this, clean one end of the plastic tubing with a disposable alcohol pad. Place the other end of the plastic tubing down into beaker **A** so that it extends almost all the way down to the surface of the liquid on the bottom. Then hold the beaker above your head, and "prime the pump" by sucking quickly on the clean protruding end of the tubing, before quickly lowering that end down into beaker **B**. (See Figure 3)

Note: This is somewhat tricky to do, and you may want to practice first by trying to siphon some water from one cup to another. When you prime the siphon, try to avoid inhaling any of the CO_2. If you do, it may cause you to cough, but it will not harm you.

Try generating carbon dioxide gas by a different method. Into a clean and dry beaker **A**, place about 240 mL of warm water. Wearing gloves, place 2–3 thumb-sized pieces of dry ice (solid CO_2) into the water. Caution: **Dry ice is VERY cold. Avoid any direct contact with skin.** The CO_2 that fills the beaker will be "colored" gray—at least for a while—by the misty fog that forms when the cold CO_2 condenses some of the water vapor out of the air. Pour this "visible" carbon dioxide into beaker **B**. Describe and attempt to explain what you see.

Discussion

Baking soda is sodium bicarbonate ($NaHCO_3$). The bicarbonate ion (HCO_3^-) reacts with the hydrogen ion (H^+) ion from the vinegar, a solution of acetic acid ($HC_2H_3O_2$). The two ions produce the compound carbonic acid (H_2CO_3), which readily decomposes into $CO_2(g)$ (appearing as bubbles) and a byproduct you don't see—H_2O.

The two equations for these reactions are shown below.

$$NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow NaC_2H_3O_2(aq) + H_2CO_3(aq)$$

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l).$$

In this activity, you noticed an important chemical property of carbon dioxide: It does not allow things to burn the way oxygen does. This property makes CO_2 a good choice for use in fire extinguishers. It's surprising that CO_2 does not allow substances to burn, since carbon dioxide contains a lot of oxygen—about 73% by mass. In contrast, normal air contains only about 21% oxygen, yet it certainly allows many substances to burn. The low reactivity of CO_2 is explained by the fact that the oxygen in it is already bonded—to the carbon. And since the bonds in CO_2 are strong, the oxygen is not readily available to react with other elements.

But, if the oxygen in CO_2 is offered to an element with which it could form an even stronger bond than it makes with carbon, a reaction does take place. Magnesium metal is one such element. It burns quite vigorously in air, as well as in an environment of pure CO_2. This is why a CO_2 fire extinguisher should never be used on a magnesium fire!

Another physical property of CO_2 you observed in this activity is that it tends to diffuse upward very slowly. This is a result of its rather high density—1.96 g/L at standard temperature and pressure. In contrast, ordinary air has a density of only 1.29 g/L. Because CO_2 is so much denser than air, it tended to remain in the beaker, almost as though it were some kind of liquid. And like a liquid, CO_2 is dense enough to be poured, even siphoned, from one container to another. Pourability and siphonability are properties we do not usually associate with gases.

Here's a good question that may have already occurred to you. If CO_2 sinks to the bottom of the container of air, how does it get evenly mixed with other gases in the atmosphere? Think about that one as you learn more about the behavior of gases in your chemistry class.

Carbon dioxide, the greenhouse gas that both enables and threatens life on earth, has many fascinating chemical and physical properties. Any strategies for managing the levels of CO_2 in the atmosphere must result from a thorough understanding of this unique and abundant compound. Future articles in *ChemMatters* will explore more of the problems and some of the solutions challenging chemists who study this pourable greenhouse gas.