GAS LAWS

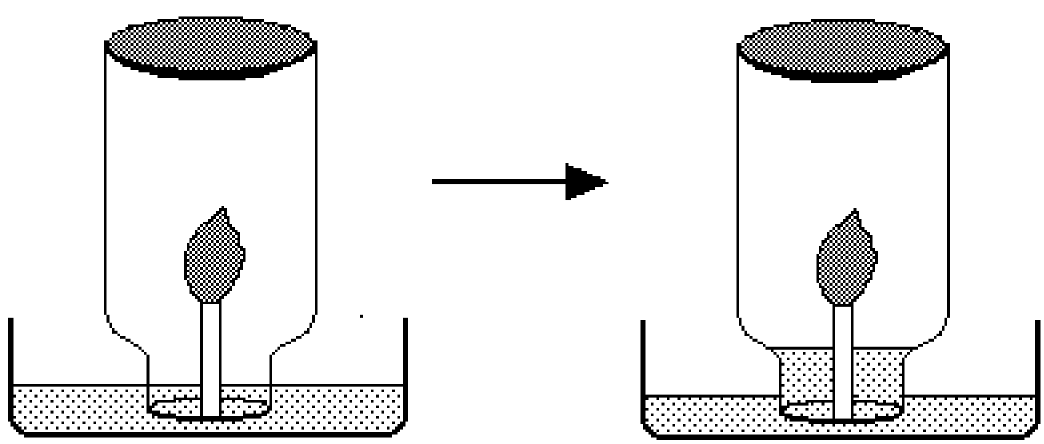
PROBLEM PRESENTATION / EXPLORATION

A. The Collapsing Soft Drink Can

1. The materials needed for this demonstration include:
   - a.) 2 empty soft drink cans (Diet Coke® and Classic Coke®)
   - b.) a plastic dish pan filled with ice water
   - c.) tongs
   - d.) hot plate
2. Fill the pan half full with ice water.
3. Before the students enter the class pour 5 mL of water into an empty Classic Coke® can; do not put any water into the Diet Coke® can.
4. Place both cans on the hot plate and heat them for about 10 minutes.
5. Using tongs, quickly invert the Classic Coke® can into the pan of ice water.
6. Observe what happens.
7. Now invert the Diet Coke® can into the pan of ice water.
8. Have the students try and explain what happened. [When the Classic Coke® can is inverted and placed into the ice water the can immediately is crushed. When the Diet Coke® can is inverted and placed into the ice water very little collapsing takes place.]

B. Rising Tide

1. The materials needed for each student station in this activity include:
   - a.) birthday candles (a different number at each station)
   - b.) 1 dish (the bottom of a large petri dish or pie plate works well)
   - c.) 1 quart jar
   - d.) food coloring
   - e.) water
   - f.) matches
2. Secure the candle to the bottom of the dish or pie pan using hot wax or clay.
3. Half fill the dish with colored water.
4. Light the candle.
5. Invert the jar and place over the candle.
6. Record observations.



CLASS RESPONSE / CONCEPT INVENTION

A. Why did the Classic Coke® can collapse but the Diet Coke® can did not? In the Classic Coke® can the water was changed to vapor and filled the can with water vapor. When the temperature was lowered the water vapor changed back into liquid water. This takes up only about 1/70 of the volume that the vapor occupied. (Remember, the molecules are much closer together in the liquid state than in the gaseous state.) The air pressure on the outside of the can is now so much larger than the tiny pressure on the inside of the can that it collapses. In the Diet Coke® can the original air was heated up by being heated on the hot plate. But when this can was inverted and lowered into the ice water there was no drastic change in volume because the volume that the hot air took up was only reduced a small amount by being cooled down. (There was no change of state in this case.) Thus, even though the pressure on the inside of the Diet Coke® can was a little lower than on the outside, it wasn't different enough to cause the collapsing that we saw in the Classic Coke® can. You will probably see a little collapsing of the Diet Coke® can. Make note of it since this can be used to reinforce the example of Charles' Law that will be taken up later. (Charles' Law says that volume and temperature are directly proportional to each other. In other words, as the temperature of the air decreased, the volume of the air also decreased.)

B. Why did the water rise inside the jar as the candles went out?

1. In many science books an erroneous explanation has been advanced for this demonstration. It has, in the past, been used to show that the air is composed of 20% oxygen because when the candle goes out (due to there not being enough oxygen to support combustion) about 20% of the volume of the bottle fills with water. However, **this is wrong !** The amount of water entering the jar depends on a number of factors and will not rise 20% of the volume every time. Some of these factors are the number of candles, the temperature of the water, and how fast the jar is placed over the candle.
2. The first major thing wrong with this explanation is that there is just about as much carbon dioxide produced in the burning of the candle as there was oxygen used up. Therefore, the overall volume of the gas should not change much at all if this explanation were correct.
3. The second major thing wrong with this explanation is that the oxygen is not actually all used up when the candle goes out. There is a classic film loop from the old ESS curriculum that shows this by putting a mouse in the bottle in which the candle is burning. The mouse has no trouble breathing after the candle goes out since there is quite a bit of oxygen remaining.

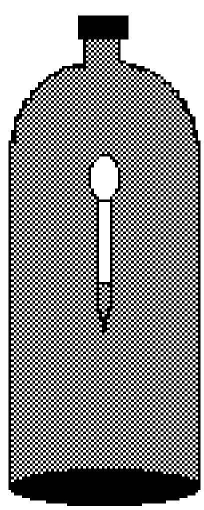
C. Why then did the water come up?

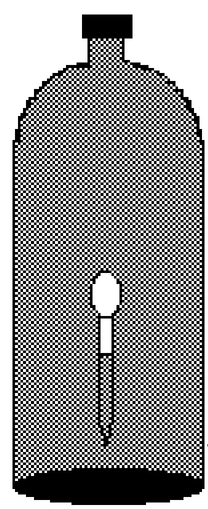
1. The foundation for explaining this can be provided by doing the following that will require these materials: a balloon, a ruler, an oven, and a freezer (This activity can be completed as a homework assignment if necessary.)
2. Inflate a balloon and tie it off. Allow enough room for the balloon to expand.
3. Using the ruler, or better yet using large calipers, measure and record the diameter of the balloon. If calipers are not available place the balloon between two shoe boxes on a table. Measure the span between the two boxes and this will be the diameter.

4. Place the balloon in an oven for 15 minutes at not more than 150°F (65°C). Remove the balloon and immediately measure and record the diameter.
5. Now, place the balloon in the freezer for 15 minutes. Remove and immediately measure and record the diameter.
6. What happened to the balloon in each case. What appears to be the relationship between volume and temperature?
7. It appears that as temperature goes up, volume goes up; and as temperature goes down, volume goes down. This is know as **Charles' Law: Temperature is directly proportional to volume, assuming that all other variables remain constant**.
8. A more correct reason why the water comes up into the bottle when the candle goes out can be seen in light of Charles' Law. While the candle was burning it heated up the air in the bottle. This caused the air to expand , pushing some of it out the bottom of the bottle as it was lowered over the candle. When the candle went out this now smaller mass of air cooled. As we have seen from above, when the air is cooled, the volume decreases. When this happens, the water rushes in to take up the space formerly occupied by the warmer air.
9. As has been mentioned above, there are other factors that contribute to the rising water. We will not investigate them at this time.

D. Boyle's Law - Cartesian Diver

1. The materials needed for each station in this activity include: a 2-L soft-drink bottle, a medicine dropper, and water
2. Remove the label from the soft-drink bottle and fill the bottle with water leaving no air in the bottle.
3. Add enough water to the dropper to cause the dropper to remain at the top of the bottle but completely submerged when placed into the 2L bottle.
4. Again make sure the bottle is filled to the top and screw the cap on the bottle and tighten.
5. Apply pressure to the sides of the bottle; release the pressure; what happens to the dropper?
6. How do you explained what happened? What is the relationship between pressure and volume of air trapped in the dropper?





BEFORE SQUEEZING **AFTER SQUEEZING**

7. Examine closely the air trapped in the medicine dropper. What happens to the amount of air when you press the sides of the bottle? What happens to

the amount of the air when you release the pressure? (According to Boyle's Law, increasing the pressure on a gas will decrease its volume. Applying pressure to the sides of the bottle increases the pressure on the water, forcing it into the dropper and reducing the volume of air in the dropper. The dropper becomes heavier and sinks. Releasing the pressure allows the volume to increase, forcing out some of the water. Since the medicine dropper now weighs less it will once again float.) Do you think the dropper would act the same way if the 2L bottle was only about two thirds full of water, in other words, the top third of the bottle was air? [You might have to squeeze a little harder, but it still works.]

8. The fancy way of stating **Boyle's Law is: The volume of a gas is inversely proportional to its pressure if all other variables are held constant.**
9. This set up is referred to as a Cartesian diver. Can you think of anything outside the laboratory that works on this principle? [Submarine]

E. If you have a vacuum pump, there are many activities that can show the relationship of volume and pressure. A vacuum pump simply pumps out air from an enclosed system. Since there are fewer molecules now present, the air pressure is lowered. When air is allowed to reenter the container the pressure will then increase.

1. Put a partially inflated balloon that has been tied off under a bell jar.
2. Connect the vacuum pump to the bell jar. What will happen when the vacuum pump is turned on?
3. Before the pump is turned on there is equal pressure felt on both the inside of the balloon and on the outside of the balloon.
4. When the pump is turned on, air on the outside of the balloon is being pumped out and the pressure under the bell jar is suddenly less than the pressure on the inside of the balloon. Consequently the balloon expands. It is trying to get back to the condition where the pressure on the inside and on the outside are once again the same. Until this happens the volume of the balloon will expand.
5. When the pump is turned off, the volume stops expanding. The resulting pressure on both the inside and outside of the balloon is equal but lower than at the beginning. Note with this lower pressure, there is larger volume.
6. What will happen when air is let back into the bell jar? Since the outside pressure is now going up, the volume of the balloon is collapsing so that the pressure inside the balloon can also increase. When the pump is stopped again, the volume will also stop changing. As the pressure went up, the volume went down.
7. What do you think would happen if you dispensed some shaving cream into a beaker and put it under the bell jar? Try it. (Remember, there is gas trapped in the soap inside the can. This gas is the propellant that forces the shaving cream out when you press the button.)
8. Try putting marshmallows under the bell jar.
9. What other interesting things might you put into the bell jar to try to show the relationship between pressure and volume of a gas?

CONCEPT EXTENSION

A. Will the Cartesian Diver work with liquids other than water? [In most liquids, the answer is yes]

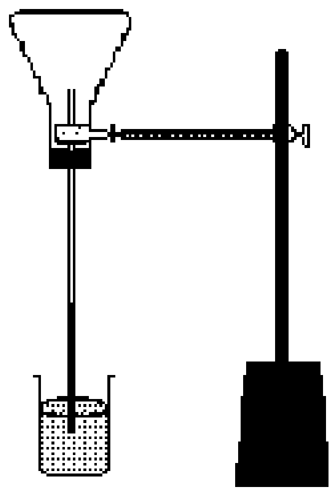
1. Try a liquid less dense than water and one that is more dense than water.
2. Because alcohol is less dense than water it is a good choice to test. You might want to use a 1L bottle or even one smaller such as a clear plastic dish washing soap bottle so that you don't have to use so much alcohol to fill it up. Rubbing alcohol is probably the cheapest alcohol. If you are careful, the alcohol will not be harmed by simply putting a dropper into it and can be used for other experiments.
3. Because the density of the alcohol is lower than water's density you might need a longer dropper to get it to work.
4. Another liquid that could be used that is less dense than water is cooking oil.

The dropper could be added right to the plastic bottle that many brands of cooking oil are packaged.

5. Many liquids that are more dense than water are not probably real safe to use. However, a 10% solution of salt is safe, cheap, and easy to make. (You need 10 grams of salt for every 90 grams of water.)
6. Another liquid that is more dense than water is 7-Up® because of all the sugar dissolved in the water. Ask the students if they think this will work? Let them try it. It won't work. Why? Let them figure this problem out. The carbon dioxide gas coming out of solution prohibits the Cartesian Diver from working properly. Once they have decided the bubbles are causing the problem, have them try and explain why they are interfering. Is there any way that 7-Up® could be made to work? Someone is bound to come up with the idea of heating the 7-Up® until all the carbon dioxide has been expelled and then using the "degassed" 7-Up®. Now it will work.

B. Since there appears to be a relationship between the volume of a gas and the temperature of the gas, could we make a thermometer made out of gas instead of alcohol or mercury?

1.



2. Constructing an "air thermometer"
    a. An "air thermometer" can be made by obtaining a Pyrex® flask with a one-holed rubber stopper fitted with a long piece of glass tubing (40-50 cm).
    b. Invert the flask so that the end of the glass tubing is well below a beaker of colored water, then clamp it to a ring stand.
    c. Heat the flask with a Bunsen burner. Note that the molecules of air inside the flask will expand and bubbles will be seen coming out of the end of the tube into the colored water.
    d. Now, place on the top of the flask a wet towel. The air will cool and contract allowing the colored water to move up the tube.
    e. When the system has equilibrated the tube can be calibrated. Mark on the tube with a grease pencil or other marking device the height of the water. From a mercury thermometer note the temperature of air in the room an assign this temperature to the mark just made.
    f. Continue to calibrate the "air thermometer" by changing the conditions to which the flask of air is subjected. Under each condition the mercury thermometer must be placed under the same conditions so that the "known" temperature can be assigned to the corresponding height of the colored water in the tube.
    g. Intermediary marks can be made between the known temperatures by proportionally drawing lines on the tube that correspond to the

difference in degrees of temperature between the known temperature marks.

3. This is a very sensitive thermometer as long as the glass tube is never allowed to stick out above the level of the colored water. If this happens it needs to be recalibrated.