Carbon Dioxide in Inhaled and Exhaled Air

Most people who have heard of John Dalton know him as the father of modern atomic theory, and so he was. However, Dalton's earliest scientific interests were in meteorology, and it was those interests that led him to study the gases of the air and their ultimate constituents, meaning atoms. Dalton also maintained records of the weather in Manchester, England for about 57 years. His studies of his own colorblindness are recognized in the use of "Daltonism" as a synonym for that condition.

One of Dalton's discoveries was that the gases of the atmosphere can physically combine and still maintain their individual pressures. In other words, if gases 1, 2, 3, ... have so-called partial pressures $P_1$, $P_2$, $P_3$, ... then

$$P_{\text{total}} = P_1 + P_2 + P_3 + ...$$

(1)

is the total pressure of their mixture. Equation (1) is known as Dalton's law of partial pressures. It implies that for any one gas, say "i", we will have

$$\% \text{ Gas } i = 100 \times \frac{P_i}{P_{\text{total}}}$$

(2)

to a good approximation, where $P_i / P_{\text{total}}$ is termed the mole fraction ($x_i$) of gas i in the mixture. In other words,

$$P_i = x_i P_{\text{total}}$$

(3)

a relationship with which you may be familiar from your chemistry textbook.

The purpose of this experiment is to give you experience using Dalton's law. You will apply equation (3) to determine the partial pressure of CO$_2$ ($P_{\text{CO}_2}$) in the atmosphere and in your lungs. The measurement will be made by using a concentrated NaOH solution to
remove CO₂ from air samples. Knowing the air pressure before and after CO₂ removal will let you quickly calculate $P_{CO₂}$, and thus $x_{CO₂}$.

**Chemical Reactions**

Although you are measuring a physical property, it is important not to lose sight of the chemistry. Sodium hydroxide is a strong base, and so it fully dissociates when dissolved in water:

$$\text{NaOH} \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)$$

If you combine CO₂ with a sodium hydroxide solution, the gas reacts with OH⁻(aq) to give the bicarbonate ion, as follows:

$$\text{OH}^-(aq) + \text{CO}_2 \rightarrow \text{HCO}_3^-(aq)$$

If NaOH is present in excess then

$$\text{OH}^-(aq) + \text{HCO}_3^-(aq) \rightarrow \text{H}_2\text{O} + \text{CO}_3^{2-}(aq)$$

will occur, with the carbonate ion precipitating as white Na₂CO₃. These are the reactions of interest in this experiment.

**Experimental Procedure - Inhaled Air**

The last page of this handout has a drawing and a photograph of the equipment needed for this experiment. Assemble the equipment if it has not already been put together for you. Note that it is imperative that the assembled apparatus be air tight.

To begin, pour about 5 mL of 10 $M$ NaOH solution into an empty litmus paper vial. *Be careful not to get the solution on you or your belongings since it is quite caustic.* Carefully put a few drops of mineral oil on top of the NaOH solution to keep it from prematurely reacting with atmospheric CO₂.

Now carefully lower the vial into the flask (or bottle) of your apparatus, being careful not to spill the NaOH solution. Forceps can be useful for this step. Insert the two-hole stopper snugly into the neck of the flask. Put the long glass tube into a 100-mL beaker, half-filled with water, as shown in the figure, and then add a few drops of food coloring. Check that the water levels inside and outside the glass
tube are the same. Assuming they are, seal the short open piece of rubber tubing with a small clamp.

Consult the last page again to be sure that you have everything set up correctly. Also consider the following:

- It would be a good idea to have the instructor or lab assistant check your equipment at this point.

- You'll need to know the barometric pressure to finish this experiment. Now is a good time to read the barometer and record the pressure in your lab notebook in mm Hg (torr).

Carefully tilt your flask, or bottle, so as to spill the NaOH solution and mineral oil. Gently swirl the flask and then use a ruler to measure any change in the water level in the long glass tube. Record the change, or lack thereof, in your lab notebook in units of millimeters. Any change in level is caused by the removal of CO₂ from the air in the flask. The change in the water level gives the atmosphere's partial pressure of CO₂, in mm H₂O units.

Now open the clamp, remove the glass tube from the beaker, and take the vial from the flask. Clean up the vial and flask and then repeat the measurement you just made so that you will have two PCO₂ values to average.

**Experimental Procedure - Exhaled Air**

Again assemble the equipment, but this time do not put the end of the long glass tube into the beaker. Take a breath of air, wait a second or so, then exhale into the short rubber hose. Now put the long glass tube into the beaker of water and proceed as before. This time the water level's rise will represent the pressure of CO₂ in your exhalation. Be sure to make two measurements for the exhaled air.

When you exhaled for the previous two measurements, some of the air you sent into the bottle was from dead space in your body and in the tubing. That air never came into contact with the capillary beds deep within your lungs. Now set up the equipment once more, but this time vent the first portion of your exhalation into the room and blow the rest into the rubber hose. This ensures that the air in the bottle actually comes from metabolized gases inside you. As before, measure the rise in the water level, which could well be a few hundred
millimeters. As always, repeat the measurement so you can average the two values.

Calculations

If you have not already done so, arrange the results of your six experimental runs in a table in your laboratory notebook. Multiply the barometric pressure by 13.6 to convert from mm Hg to mm H_2O units. The resulting value is \( P_{\text{Total}} \) for the air. Use Dalton’s law of partial pressures, from equation (3), to calculate the mole fraction of CO_2 and the % CO_2 in each of your six measurements as follows:

\[
\begin{align*}
    x_{\text{CO}_2} &= \frac{\text{water rise, in mm}}{P_{\text{total}}, \text{ in mm}} \quad (4) \\
    \% \text{ CO}_2 &= 100 \times \frac{\text{water rise, in mm}}{P_{\text{total}}, \text{ in mm}} \quad (5)
\end{align*}
\]

Finally, average the first two values (atmospheric CO_2), the middle two values (exhaled air), and the last two values (air from deep within lungs).

Lab Report

Write up the experiment in your laboratory notebook as usual. It would be a good idea to review the guidelines and to double-check with your instructor about any variations on them. As always, be sure to discuss your results.

References

This experiment was adapted from work in the University of Tennessee's chemistry department. An older version is in the April, 1974 issue of the Journal of Chemical Education (page 273).
Strange as it may seem, over 150 years ago Dalton's will stipulated that his eyes be left for medical examination. What can you learn of the results?

Dalton's original table of atomic and molecular symbols