# Partial Pressure

**Strand** Phases of Matter and Kinetic Molecular Theory

**Topic** Investigating partial pressure and gas laws

**Primary SOL** CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include

a) pressure, temperature, and volume;
b) partial pressure and gas laws;
c) vapor pressure;
d) phase changes;
e) molar heats of fusion and vaporization;
f) specific heat capacity; and
g) colligative properties.

**Related SOL** CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

a) designated laboratory techniques;
b) safe use of chemicals and equipment;
c) proper response to emergency situations;
f) mathematical and procedural error analysis;
g) mathematical manipulations including SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, and dimensional analysis;
h) use of appropriate technology including computers, graphing calculators, and probeware for gathering data, communicating results, and using simulations to model concepts.

CH.3 The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include

a) nomenclature;
b) balancing chemical equations;
c) writing chemical formulas;
e) reaction types.

CH.4 The student will investigate and understand that chemical quantities are based on molar relationships. Key concepts include

a) Avogadro's principle and molar volume;
b) stoichiometric relationships.

## Background Information

Dalton's law of partial pressures states that the total pressure exerted by a mixture of nonreactive gases is equal to the sum of the partial pressures of the individual gases in the mixture.

 

By measuring the volume of air removed in a reaction of iron and the oxygen in a sample of air and then comparing this volume with the total volume of the air in the sample, we can determine the percentage of oxygen in the air.

## Materials

- Large test tubes
- Water
- Iron filings
- 250 mL beakers
- Ring stands and test tube clamps
- Markers
- Graduated cylinders
- Calculators
- Barometer

## Vocabulary

*atmospheric pressure, barometer, Charles' law, constant pressure, constant temperature, Dalton's law of partial pressures, scoopula*

## Student/Teacher Actions (what students and teachers should be doing to facilitate learning)

This activity should follow the one on Charles' law so that students have had previous experience with water filling space previously occupied by a gas.

This activity presents students with the problem of determining the percentage of oxygen in the air. Since air is largely nitrogen with most of the rest consisting of oxygen plus a small amount of other gases, one way to solve the problem would be to remove all the other gases besides oxygen from a sample of air and then determine oxygen's percentage of the total. Another way would be to remove the oxygen and then determine oxygen's percentage of the total, using the difference.

You may wish to present students with an array of laboratory equipment to use and a list of viable chemicals and have them research and then perform various experiments to solve the problem. Alternatively, you may have all students perform the following experiment.

1. Instruct each lab team to rinse a large, clean test tube with water in order to wet the inside surface of the tube. Excess water should be poured out.
2. Have each team use a scoopula to place a small pinch of iron filings in the test tube and shake to coat the inside surface with filings. Excess filings should be poured out.
3. Have each team slowly turn their test tube upside down and carefully lower it into a 250 mL beaker that is half full of water, making sure not to lose any of the air in the tube. (Placing a finger over the mouth of the tube until it is under water may help). Direct each team to use a ring stand and test tube clamp to support the tube in a vertical position so that it will not fall over.
4. The next day or next class period (at least 24 hours later), call students' attention to the fact that water has risen up into the test tube. Ask them why, and prompt them to respond that the volume of air in the tube must have decreased and that the water rose up into the tube to replace that volume. Discuss what might have caused this to happen. Lead students

to arrive at the conclusion that some component of the air reacted with the iron and was removed from the air. Display the balanced chemical equation for this reaction, $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$, and point out that it is oxygen that reacted with the iron to form iron oxide. Consequently, water replaced the volume of oxygen that was removed from the air by the reaction.

5. Have each team raise or lower their test tube so that the level of water in the tube is the same at the level of water in the beaker, thus ensuring that the pressure of the gas in the tube is equal to the atmospheric pressure, which is constant. Ask them whether they can assume that the temperature of the gas in the tube is equal to the temperature of the air in the room, which is also constant. (Yes) Point out that, in other words, they are working with constant temperature and constant pressure.
6. Have each team use a marker to mark a line on the test tube indicating the level of the water in the tube. Then, direct each team to remove their test tube from the clamp and water bath, rinse out all remaining solid material, fill the tube with water up to the mark, and measure this volume of water with a graduated cylinder. Finally, have each team fill their test tube completely with water and measure this volume with a graduated cylinder.
7. Instruct each team to find the difference in the two volumes. Ask students what this difference equals. (The volume of oxygen that reacted with the iron filings and was removed from the air in the tube) Ask students how they can obtain the percentage of oxygen in the air. (By dividing the difference by the total volume)
8. Finally, have students determine the atmospheric pressure, using the class barometer. Then, have them multiply the atmospheric pressure by the percentage of oxygen in the air, which gives the partial pressure of oxygen in the air. Since air is primarily oxygen and nitrogen, the difference between the partial pressure of oxygen and the total atmospheric pressure is a good estimate of the partial pressure of nitrogen in the air.

## Assessment

- **Questions**
  - A sample of a gas contains 4.00 g of hydrogen, 8.00 g of oxygen, and 2.50 g of nitrogen. What's the mole fraction of each gas in this mixture?
  - The partial pressure of oxygen in a sample of air was observed to be 156 mm Hg on a day when air pressure was 743 mm Hg. What was the percent by volume of $O_2$ present?
  - Using the composition of air given above, is percent by volume the same as mole fraction?
  - If a mixture of 2.50 g each of $H_2$ and He exerts a total pressure of 0.75 atm, what is the partial pressure of each gas?
  - For a deep-sea dive, 46.0 L of $O_2$ and 12.0 L of He, both at STP, are pumped into a 5.00 L scuba diving tank. What is the total pressure in the tank and the partial pressure of each gas?
- **Journal/Writing Prompts**
  - Divers use compressed air when scuba diving. How are Dalton's law of partial pressures and Boyle's law important in scuba diving procedures? Why is Charles' law usually irrelevant to scuba diving?

  - It is fairly well documented that air contains about 20% oxygen. Explain what would happen to life as we know it today if there were a lower percentage of oxygen in the air. Explain how life would be affected if there were a greater percentage of oxygen in the air?
- **Other**
  Have students answer the following questions.
  - 5.00 L of $CO_2$ at 500 torr and 3.00 L of $CH_4$ at 400 torr are put into a 10.0 L container. What is the pressure exerted by the gases in the container?
  - At 27˚C and 750 torr pressure, what will be the volume of 2.3 mol of Ne?
  - A sample of propane gas, $CH_4$ , was collected over water at 25˚C and 745 torr. The volume collected was 1.25 L. What will be the volume of the dry propane at STP? (Vapor pressure of $H_2O$ at 25˚C is 23.8 torr.)
  - A gas occupies a volume of 410 mL at 27˚C and 740 mm Hg. Calculate the volume at STP.
  - What would be the partial pressure of $N_2$ gas collected over water at 25˚C and 705 torr pressure? (Vapor pressure of $H_2O$ at 25˚C is 23.8 torr.)
  - Normal scuba tanks are filled with compressed air. If the pressure of air in the tank is 200 atm, what is the partial pressure of the oxygen? Of the nitrogen?

## Extensions and Connections (for all students)

- Have students explain what happens to the percentage of oxygen in the air as altitude increases and as altitude decreases.
- Gases dissolve in water to a certain percentage. The concentration of oxygen in water is temperature dependent. As water temperature increases, the percentage of dissolved oxygen decreases. Have students explain the implications for marine life as the water temperature increases.

## Strategies for Differentiation

- Have students draw representations of Charles' law, Boyle's law, Avogadro's law, Dalton's law of partial pressures, and the ideal gas law.
- Have students develop a physical representation of each gas law other than a drawing.
- Ask a student who has experience with scuba diving explain to the class how the various gas laws apply to the diving experience. Alternatively, invite a diver to the class to talk about diving, and have students make the connection in small groups following the talk. If no one with scuba diving experience is available, show a video of scuba diving, and then discuss the application of the gas laws to what is seen.