1. *Boyle's Law and the Empty Space in Air*

Introduction

Unlike liquids and solids, gases are easily compressed. This property was first studied quantitatively by Robert Boyle (1627–1691), who formulated the relationship that was later to be known as *Boyle's law* (Ebbing/Gammon, Section 5.2): the volume of a given amount of gas at a given temperature is inversely proportional to the applied pressure, or

$$PV = k \text{ (constant)}$$

(at a fixed temperature for a fixed amount of gas).

The ready compression of gases was later explained by one of the postulates of the *kinetic-molecular theory of gases* (Ebbing/Gammon, Section 5.6). This postulate states that gases are composed of molecules whose size is negligible in comparison with the distance between them. Gases are easily compressed because most of their volume is empty space!

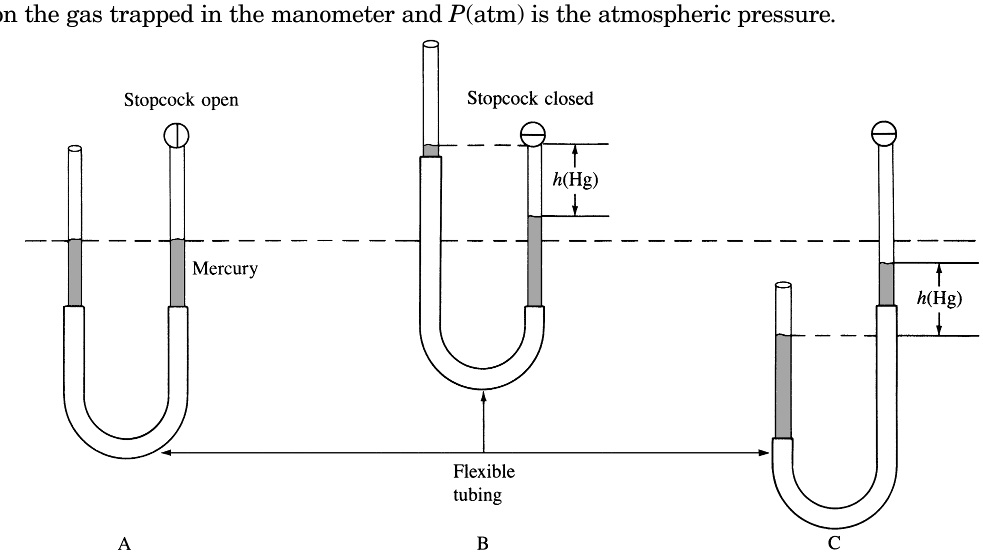
Purpose

This experiment has two objectives. First, you will examine the validity of Boyle's law, using air as the gas. Your laboratory instructor will make most of the required measurements for you. However, you will have one opportunity to use the apparatus that is needed in this part of the experiment. Second, you will determine the amount of empty space in air.

Manometers

A flexible manometer is required in this experiment. This device, which allows the pressure of an enclosed gas to be compared with that of the atmosphere, is U-shaped and contains mercury. The shape is achieved by attaching a piece of flexible tubing to two glass tubes, as shown in Figure 1.1A. One of these glass tubes is open-ended and the other is equipped with a stopcock.

FIGURE 1.1
A flexible tube manometer with (*A*) $P = P(\text{atm})$, (B) $P > P(\text{atm})$, and (C) $P < P(\text{atm})$, where P is the pressure exerted on the gas trapped in the manometer and $P(\text{atm})$ is the atmospheric pressure.



When the stopcock is open, both columns of mercury have identical heights, as in Figure 1.1A, because only atmospheric pressure is exerted on each of them. This stopcock, however, is closed more often than it is open. It is used during an experiment to prevent air from escaping from that end of the manometer and to allow pressures less than or greater than atmospheric pressure.

The flexible tubing allows the open-ended side of the manometer to be raised or lowered. Raising or lowering the open side affects the pressure exerted on the trapped air according to the relationship

$$P = P(\text{atm}) + h(\text{Hg})$$

In this equation, P is the pressure exerted on the trapped air, and P(atm) is the pressure of the atmosphere. A barometer (Ebbing/Gammon, Section 5.1) is used to measure P(atm). To obtain the term h(Hg), we subtract the height of the mercury in the closed end from the height of mercury in the open end, or

$$h(\text{Hg}) = h(\text{open}) - h(\text{closed})$$

When the open end is raised, $h(\text{open}) > h(\text{closed})$, so h(Hg) is positive (see Figure 1.1B). Thus the pressure exerted on the gas will be greater than atmospheric pressure by h(Hg). When the open end is lowered, $h(\text{closed}) > h(\text{open})$, so h(Hg) is negative (see Figure 1.1C). In this case, the pressure exerted on the trapped air will be less than atmospheric pressure.

Another way to state Boyle's law

The manometer will not allow you to measure directly the volume of the air trapped in the glass tube between the stopcock and the mercury. For this reason, it is convenient to redefine Boyle's law in terms of the distance between the stopcock and the surface of the mercury.

The glass tube containing the air will be treated as a uniform cylinder whose volume is given by

$$V = \pi r^2 L$$

where r is the radius of the cylinder and L is its length. Boyle's law becomes

$$PV = P\pi r^2 L = \text{k}$$

When all the constants have been collected on the right-hand side of the equation, the result is

$$PL = \frac{\text{k}}{\pi r^2} = \text{c (another constant)}$$

Boyle's law can now be restated in terms of this equation: if a fixed quantity of a gas is enclosed in a cylinder with a variable length, the observed length will be inversely proportional to the applied pressure.

The components of air

Air, the gas used in this experiment, consists of molecular nitrogen (N_2), approximately 80% by volume, and molecular oxygen (O_2), approximately 20% by volume (Ebbing/Gammon, Chapter 5, "A Chemist Looks At: Oxygen"). Although argon, carbon dioxide, and other substances are also present in the air, their presence is neglected in this experiment because their combined percentages amount to slightly less than 1%.

Because part of this experiment deals with the empty space in a sample of air whose total volume is known, the filled volume occupied by the molecules in the sample must be calculated. A molecule of N_2 and a molecule of O_2 have similar volumes. For purposes of this experiment, an identical volume of 3.8×10^{-24} cm^3/molecule will be assumed for each component.

Concept of the experiment

First, a series of measurements of the length of a cylinder containing trapped air and the corresponding pressure will provide the data for you to test Boyle's law. You should find that the product PL is a constant within the precision of the measurement.

Second, you will determine how much empty space there is in another sample of air. The pressure exerted on this sample and its volume and temperature will be obtained. By applying the ideal gas law (Ebbing/Gammon, Section 5.3) to these experimental data, you can calculate the combined number of moles of N_2 and O_2 in the sample. From the calculated number of moles, Avogadro's number, and the volume occupied by a single molecule, you can compute the filled volume used by the molecules in your sample. When the filled volume is subtracted from the actual volume of the sample, you will have determined the empty space.

Procedure

Getting started

1. Review the following definitions so that you understand the measurements that are required to test Boyle's law.

P	= pressure (mmHg) exerted on the trapped air
P(atm)	= atmospheric pressure (mmHg)
h(open)	= height (mm) of mercury in the open side of the manometer
h(closed)	= height (mm) of mercury in the closed side of the manometer
h(Hg)	= h(open) − h(closed)
L	= length of the gas column on the closed side, or, in other words, the result of subtracting h(closed) from the measured height of the bottom of the stopcock

Testing Boyle's law

1. Obtain and record the atmospheric pressure.
2. Obtain and record the height of the bottom of the stopcock.
3. Your laboratory instructor will close the stopcock and measure h(open) and h(closed) without moving the open end of the manometer. Record the results as Trial 1.

CAUTION: Mercury spills must be cleaned up immediately. Mercury vapor is very dangerous.

4. The open end will be raised to three successively higher positions. Record the measurements of h(open) and h(closed) at these positions as Trials 2, 3, and 4.
5. The open end will now be moved to three positions that are successively lower than the one recorded in Step 3. Record the measurements of h(open) and h(closed) at these positions as Trials 5, 6, and 7.
6. Your laboratory instructor will allow you to move the open end *cautiously* to any position that you desire. Measure h(open) and h(closed) at this position and record these measurements as Trial 8.
7. Calculate h(Hg), P, and L for each trial. Finally, calculate the product PL for each trial. Use scientific notation. Is PL a constant within the context of the use of significant figures (Appendix: Mistakes, Errors, Accuracy, and Precision)?

Determining the empty space in air

1. Measure and record the temperature of the laboratory. Convert that temperature to a Kelvin temperature.
2. Place a 100-mL graduated cylinder on the laboratory bench in front of you. Think about the air that occupies 50.0 mL of the volume of the graduated cylinder. What is the pressure exerted on this volume of air? Record your result. What is the temperature of this air? Record your result.

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Date:	....	Student name:	....
Course:	....	Team members:	....
Section:	....		....
Instructor:	....		....

Prelaboratory assignment

1. a. What is Boyle's law in its usual form?

   b. What is the form of Boyle's law that is used in this experiment?

   c. Explain in specific detail how you will know whether your data in this experiment demonstrate the validity of Boyle's law.

2. The atmospheric pressure on a particular day is 742 mmHg. Complete the following table for a gas trapped in the closed end of a manometer.

***h*(open)** **(mm)**	***h*(closed)** **(mm)**	***h*(Hg)** **(mmHg)**	***P*** **(mmHg)**
259	286		
348	322		

Calculations:

3. What special safety precaution must be observed during this experiment?

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Results

1. *Data for testing Boyle's law*

   Atmospheric pressure: .... mmHg

   Height of the bottom of the stopcock: .... mm

Trial	***h*(open)** **(mm)**	***h*(closed)** **(mm)**	***h*(Hg)** **(mmHg)**	***P*** **(mmHg)**	***L*** **(mm)**	***PL*** **(mmHg • mm)**
1	....	....	....	....	....	....
2	....	....	....	....	....	....
3	....	....	....	....	....	....
4	....	....	....	....	....	....
5	....	....	....	....	....	....
6	....	....	....	....	....	....
7	....	....	....	....	....	....
8	....	....	....	....	....	....

   Calculations:

2. *Data for determining the empty space in air*

   Laboratory temperature: .... °C; .... °K

   Pressure exerted on sample of air: .... mmHg

   Volume of sample of air: .... mL

Questions

1. Explain to what degree, if any, this experiment has supported Boyle's law within the precision of your results and the correct use of significant figures.

2. This question concerns the empty space in air.

   a. Calculate the number of moles of gas in your sample.

   b. Calculate the number of molecules of gas in your sample.

   c. Calculate the volume that is actually filled by these molecules.

   d. Calculate the empty space in your sample.