

# Pressure vs. Volume and Boyle's Law

## Boyle's Law

### Introduction

In 1642 Evangelista Torricelli, who had worked as an assistant to Galileo, conducted a famous experiment demonstrating that the weight of air would support a column of mercury about 30 inches high in an inverted tube. Torricelli's experiment provided the first measurement of the invisible pressure of air. Robert Boyle, a "skeptical chemist" working in England, was inspired by Torricelli's experiment to measure the pressure of air when it was compressed or expanded. The results of Boyle's experiments were published in 1662 and became essentially the first gas law—a mathematical equation describing the relationship between the volume and pressure of air. What is Boyle's law and how can it be demonstrated?

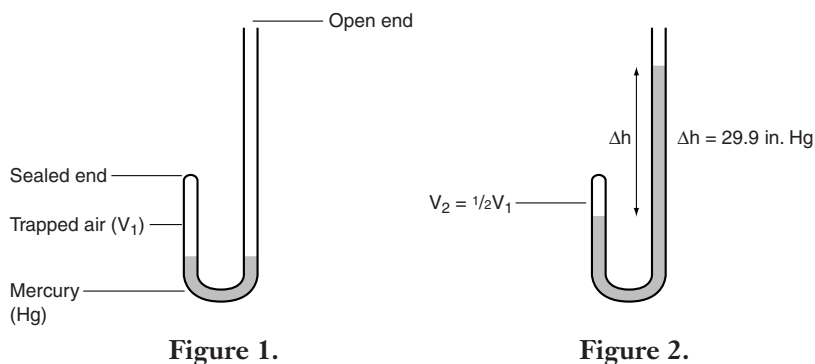
### Concepts

- Gas properties
- Boyle's law
- Pressure
- Kinetic-molecular theory

### Background

Robert Boyle built a simple apparatus to measure the relationship between the pressure and volume of air. The apparatus consisted of a J-shaped glass tube that was sealed at one end and open to the atmosphere at the other end. A sample of air was trapped in the sealed end by pouring mercury into the tube (see Figure 1). In the beginning of the experiment, the height of the mercury column was equal in the two sides of the tube.

The pressure of the air trapped in the sealed end was equal to that of the surrounding air and equivalent to 29.9 inches (760 mm) of mercury. When Boyle added more mercury to the open end of the tube, the air trapped in the sealed end was compressed into a smaller volume (see Figure 2). The difference in height of the two columns of mercury ( $\Delta h$  in Figure 2) was due to the additional pressure exerted by the compressed air compared to the surrounding air. Boyle found that when the volume of trapped air was reduced to one-half its original volume, the additional height of the column of mercury in the open end of the tube measured 29.9 inches. The pressure exerted by the compressed air was twice as great as atmospheric pressure. The mathematical relationship between the volume of the air and the pressure it exerts was confirmed through a series of measurements.



### Experiment Overview

The purpose of this experiment is to perform a modern version of Boyle's classic experiment without the hazards associated with the use of mercury. The experiment will be carried out using air trapped inside a sealed syringe within a "pressure bottle." The bottle will be pressurized by pumping in air to obtain a pressure several times greater than that of the surrounding air. As some of the excess pressure within the bottle is then released, the volume of the trapped air inside the syringe will change. Volume measurements will be made at several different pressures and the results will be analyzed by graphing to derive the mathematical relationship between pressure and volume.

## Pre-Lab Questions

1. According to our modern understanding of the gas laws, there are four measurable properties (variables) of a gas. These variables are P (pressure), V (volume), T (temperature), and n (number of moles). In Boyle's experiment, which two variables were held constant?
2. Fill in the blanks to summarize the relationship among the gas properties in Boyle's experiment: For a fixed \_\_\_\_\_ of gas at constant \_\_\_\_\_, the \_\_\_\_\_ of a gas increases as the \_\_\_\_\_ of its container decreases.
3. Pressure is defined in physics as force divided by area ( $P = \text{force/area}$ ). According to the kinetic-molecular theory, the particles in a gas are constantly moving and colliding with the walls of their container. The pressure of the gas is related to the total force exerted by the individual collisions. Use the kinetic theory to explain the results of Boyle's experiment.
4. The pressure scale on a tire gauge is marked in units of pounds per square inch (psi). The scale starts at zero when the gauge is exposed to the surrounding air. This means that the total pressure is equal to the gauge pressure *plus* the pressure of the surrounding air. Standard atmospheric pressure (1 atm) is equal to 14.7 psi. Assume that you have just inflated the tire on your bicycle to 82 psi using a bicycle pump. What is the total pressure of air in the tire in psi? In atmospheres?

## Materials

Bicycle pump with pressure gauge, or electric air pump

Graph paper, 2 sheets

Petroleum jelly, small bead

Pressure bottle, 1-L, with tire valve

Syringe, 10-mL, with syringe tip cap

Barometer (optional)

Tire gauge (optional)

## Design of the Pressure Bottle

The "pressure bottle" is a 1-L PETE (polyethylene terephthalate) soda bottle. The bottle cap has been fitted with a tire valve to give an airtight seal (see Figure 3). Pumping air into the bottle using an ordinary bicycle pump makes it possible to pressurize the bottle above atmospheric pressure. The bottle retains its volume when it is pressurized—any expansion is negligible. The plastic used to make these bottles will withstand pressures up to about 100 psi.

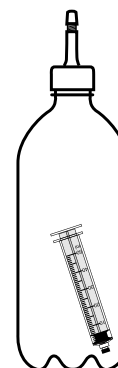


Figure 3.

## Safety Precautions

*The pressure bottle is safe if used properly. The bottle should not be inflated above 100 psi. Even if the bottle should "pop," the plastic construction will only result in a quick release of air, an accompanying loud noise, and a hole in the bottle. The bottle will split but will not shatter. Wear eye protection (safety glasses or chemical splash goggles) when working with the pressure bottle.*

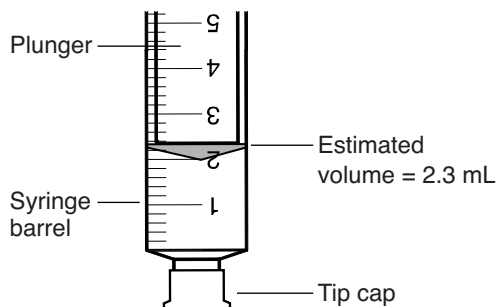
## Procedure

1. Using a barometer, measure the value of the local air pressure. *Note:* If a barometer is not available, consult an Internet site such as the national weather service site (<http://weather.gov>) to obtain a current pressure reading for your area. Record the barometric pressure in the Data Table.
2. Obtain a 1-L pressure bottle and a 10-mL syringe with a rubber tip cap.
3. Remove the tip cap from the syringe and pull on the plunger to draw about 9 mL of air into the syringe. Replace the tip cap to seal the air inside the syringe.
4. Place the sealed syringe inside the 1-L pressure bottle.
5. Run a small bead of petroleum jelly around the rim of the bottle.
6. Close the bottle with the special cap fitted with a tire valve. Tighten the cap securely.
7. Connect the tire valve to a bicycle pump or an electric air pump. *Note:* Exercise caution if using an electric air pump. Do

## Pressure vs. Volume and Boyle's Law *continued*

not exceed the maximum suggested pressure of 50–60 psi. 8. Pump air into the pressure bottle to obtain a pressure reading of 50–60 psi on the tire gauge. **Do NOT exceed 100 psi.** *Note:* Using a manual pump provides its own safety feature—it is very difficult to pump more than about 70 psi into the pressure bottle by hand.

- Loosen the connection between the pressure bottle–tire valve and the pump to release a small amount of pressure. As soon as you see the syringe plunger start to move, immediately retighten the tire valve to the pump.
- Using the pump gauge, measure and record the pressure to within  $\pm 1$  psi.
- Measure and record the volume of air trapped in the syringe at this bottle pressure. *Note:* Measure the volume at the black rubber seal, not at the inverted V-shaped projection (see Figure 4). The syringe barrel has major scale divisions marked every milliliter, and minor scale divisions every 0.2 mL. The volume should be estimated to within  $\pm 0.1$  mL.
- Loosen the connection between the pressure bottle–tire valve and the bicycle pump to release a small amount of pressure from the pressure bottle. Try to reduce the pressure by no more than about 10 psi. Immediately retighten the tire valve to the pump.
- Measure both the new pressure on the pump gauge and the new volume of the air trapped inside the syringe. Record all data in the Data Table. *Note:* If you are using a tire gauge to measure pressure, press lightly on the brass pin in the tire valve to release some air pressure. It may be necessary to bleed off enough air initially to get the first pressure reading below 50 psi, which is the scale maximum on many tire gauges.
- Repeat steps 10 and 11 to measure the volume of gas trapped in the syringe at several different pressures down to about 5 psi. It should be possible to obtain at least 5–6 pressure and volume measurements in this range.
- When the pressure on the tire gauge measures close to zero, remove the tire valve from the pump. Press down on the brass pin inside the tire valve to release all of the excess pressure within the pressure bottle. Record the final volume of air in the syringe at atmospheric pressure.
- If time permits, repeat steps 6–13 to obtain a second, independent set of pressure–volume data. Record this data as Trial 2 in the Data Table.



**Figure 4.**

# Pressure vs. Volume and Boyle's Law

## Data and Results Table

| Barometric Pressure |                          |                 |      |                |                          |                 |      |
|---------------------|--------------------------|-----------------|------|----------------|--------------------------|-----------------|------|
| Trial 1             |                          |                 |      | Trial 2        |                          |                 |      |
| Gauge Pressure      | Volume of Air in Syringe | Total Pressure* | 1/V† | Gauge Pressure | Volume of Air in Syringe | Total Pressure* | 1/V† |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |
|                     |                          |                 |      |                |                          |                 |      |

\*See Post-Lab Question #2. †See Post-Lab Question #5.

## Post-Lab Questions

- Convert the local barometric pressure to psi units and enter the value to the nearest psi in the Data and Results Table. Some appropriate conversion factors are shown below.

$$1 \text{ atm} = 760 \text{ mm Hg} = 29.92 \text{ in Hg} = 14.7 \text{ psi}$$

- The pressure gauge measures the relative pressure in psi above atmospheric pressure. For each pressure reading in the Data and Results Table, add the local barometric pressure to the gauge pressure to determine the total pressure of air inside the pressure bottle. Enter the total pressure to the nearest psi in the table.
- Plot a graph of volume on the  $y$ -axis versus total pressure on the  $x$ -axis. *Note:* The origin of the graph should be (0,0). Choose a suitable scale for each axis so that the data points fill the graph as completely as possible. Remember to label each axis and give the graph a title.

- Describe the shape of the graph. Draw a best-fit straight or curved line, whichever seems appropriate, to illustrate how the volume of a gas changes as the pressure changes.
- The relationship between pressure and volume is called an “inverse” relationship—as the pressure increases the volume of air trapped in the syringe decreases. This inverse relationship may be expressed mathematically as  $P \propto 1/V$ . Calculate the value of  $1/V$  for each volume measurement and enter the results in the Data and Results Table.
- Plot a graph of pressure on the y-axis versus  $1/V$  on the x-axis and draw a best-fit straight line through the data. *Note:* The origin of the graph should be (0,0). Choose a suitable scale for each axis so that the data points fill the graph as completely as possible.
- Another way of expressing an inverse relationship between two variables ( $P \propto 1/V$ ) is to say that the product of the two variables is a constant ( $P \times V = \text{constant}$ ). Multiply the total pressure ( $P$ ) times the volume ( $V$ ) for each set of data points. Construct a Results Table to summarize the  $P \times V$  values.
- Calculate the average value of the  $P \times V$  “constant” and the average deviation. What is the relative percent error (uncertainty) in this constant?  
$$\text{Relative percent error} = (\text{Average deviation}/\text{Average value}) \times 100\%$$
- At constant temperature, the pressure of a gas is proportional to the concentration of gas particles in the container. When some of the pressure was released from the bottle, the syringe plunger moved up. Why did this happen? Use diagrams and explain in words what happens to the gas particles moving around both inside and outside the syringe before and after the pressure is released.
- (Optional)* Research the properties of PETE on the Internet. What characteristics of PETE make it an ideal plastic for use in soda bottles?

# Teacher's Notes

## Pressure vs. Volume and Boyle's Law

### Disposal

The pressure bottles and syringes are reusable—no disposal required.

### Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

**Unifying Concepts and Processes: Grades K–12**

Constancy, change, and measurement

**Content Standards: Grades 9–12**

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, structure and properties of matter, motions and forces

### Lab Hints

- The laboratory work for this experiment can reasonably be completed in about 20 minutes. This should allow ample time in a normal 50-minute class period for students to share equipment, if student groups of two are preferred.
- Pressure bottles may be prepared using readily available 2-L soda bottles. See the *Supplementary Information* section for instructions on how to prepare home-made pressure bottles. Cut or saw off the flanges on the top of the syringes. Otherwise, the syringes will not fit inside the pressure bottles. Flinn Scientific sells a pre-made pressure bottle (Catalog No. AP5930) equipped with the attached tire valve and syringe with tip cap.
- If the bottle leaks air around the cap when pressurized, remove the cap assembly and put additional petroleum jelly around the inside seal of the cap. Vaseline® may also be used in place of the petroleum jelly.
- Best results are obtained using a bicycle pump with an attached pressure gauge. If enough bicycle pumps cannot be obtained with student help, the experiment may be done using common automobile tire gauges. The teacher should prepare the syringe/bottle assemblies and pressurize the bottles ahead of time. Since the scale maximum on many tire gauges is 50 psi, students may have to bleed off enough air initially to get the pressure below 50 psi. The units shown on some pressure gauges may be psig (gauge pressure per square inch). The correct units for the “total pressures” calculated in the Data and Results Table are psia (pounds per square inch absolute) rather than psi.
- For best results, use a barometer to measure the local barometric pressure. The national weather service site reports corrected, sea-level air pressures. Note that these are *not* actual barometric pressure readings. Meteorologists convert station pressure values to what they would be if they had been taken at sea level. The following equation can be used to recalculate the barometric pressure (in inches Hg) from the reported sea-level pressure (in inches Hg). Elevation must be in meters.

$$\text{barometric pressure} = \text{sea-level pressure} - (\text{elevation}/312 \text{ m})$$

On many syringes, the black rubber seals have two “seal lines.” Make sure students are consistent in where they measure the volume! A small amount of pressure is released in step 7 before volume measurements are made—this is to overcome the friction between the rubber seal and the syringe barrel.

- Tire gauges come in many shapes and sizes. The best gauges for this experiment are those with an attached pressure dial or digital readout rather than a “pop-out” sliding scale.
- Students may need to be instructed in the use of a tire gauge. Students usually are too cautious—they tend to press the tire gauge softly against the valve, which allows air to escape and does not give an accurate pressure reading. The best advice is to work quickly and deliberately. There is an obvious value for students in learning to use a tire gauge so that they can inflate their automobile tires to the proper pressure. Maintaining proper tire pressure improves safety, tire wear, and gas mileage.

- Try not to exceed the maximum pressure recommended in the *Procedure* section. It was found that the graph of P versus 1/V became nonlinear as the pressure bottle was pressurized above about 70 psi (total pressure 85 psi). This can be used as a teaching point—deviations from ideal gas behavior are more important at higher pressures. Many textbooks show graphs of real versus ideal gas behavior as a function of pressure. For many gases, deviations from ideal behavior become significant at pressures greater than about 200 atm. Even at modest pressures, however, small deviations are common in the  $P \times V$  “constant.” Some of this deviation may also be due to a change in temperature. Compressing the gas will increase the temperature of the gas.
- This lab provides excellent data and is a great way to introduce the use of computer spreadsheet or graphical analysis programs. Using these programs, it is possible with just a “click of the mouse” to draw best-fit, straight, or curved lines (trendlines) through data and obtain regression equations. High school students are not expected to do the math involved in generating a best-fit straight line (linear regression). It is worthwhile, however, for students to learn how this important statistical tool is used to evaluate the reliability of results.
- See the *Supplementary Information* section for an example of how the data from this experiment can be analyzed to extrapolate the value of the atmospheric pressure in units of psi.

## Teaching Tips

- What does a pressure of 14.7 psi feel like? The “Atmosphere Bar” available from Flinn Scientific (Catalog No. AP5882) is a 52-inch steel bar that weighs 14.7 lbs. The base of the bar is one inch square. The resulting pressure of 14.7 pounds per square inch certainly feels impressive!
- If necessary, review the definition of the average deviation with your students. The average deviation is obtained by finding the difference between each individual value ( $X_i$ ) and the average value ( $\bar{X}$ ), taking the sum ( $\Sigma$ ) of their absolute values, and dividing by the number of measurements.

$$\text{average deviation} = \frac{\sum_{i=1}^n |x_i - \bar{x}|}{n}$$

- Does Boyle's law depend on the nature of the gas? The experiment can be extended to test the behavior of gases other than air by filling the syringes with pure gases such as helium, hydrogen, and nitrogen. Any source of gas may be used. Contact Flinn Scientific to obtain a complimentary copy of our ChemFax No. 10527, “Construction of a Simple Gas Delivery Apparatus.”
- The approaches shown here may be conceptually challenging for students but they lead to some interesting discussions. Are negative pressures possible? Why or why not?
- Boyle's law can also be demonstrated with every breath you take! When you inhale, your diaphragm moves down, and the volume of the body cavity surrounding the lungs expands. As the volume increases, the pressure decreases, and the outside air, which is at a higher pressure, rushes into the lungs. The opposite occurs when you exhale—the diaphragm moves upward, the lung cavity contracts, and the pressure increases relative to the outside air, pushing air out of the lungs. See the “Functioning Lung Model” (Catalog No. FB1110) available from Flinn Scientific for a large, demonstration-size model of the pressure–volume relationship involved in breathing.
- The great 17th century physicist Isaac Newton was a contemporary of Robert Boyle, and Boyle's results prompted Newton to propose an explanation for the pressure of a gas. Newton assumed that the particles in a gas were motionless (static) and that the pressure of a gas was due to the mutual repulsion of gas particles. That no less a person than the “father of physics” and the discoverer of the calculus proposed this model might serve as a cautionary tale for all teachers—our modern, obvious explanations for gas behavior may not always be obvious to students.
- The foundation of the modern kinetic-molecular theory of gases is usually attributed to Daniel Bernoulli. In 1734 Bernoulli proposed that a gas was composed of “hard” (inelastic) particles that were in constant and random motion. The force exerted by these particles as they collided with the walls of their container was the source of the gas pressure. Using this model, Bernoulli derived a mathematical equation for the pressure of a gas. His mathematical treatment correctly “predicted” the inverse pressure–volume relationship that had been observed empirically by Robert Boyle. The development of the kinetic theory culminated in the mid-1800s with the definition of temperature and the relationship between temperature and kinetic energy.

**Answers to Pre-Lab Questions** *(Student answers will vary.)*

1. According to our modern understanding of the gas laws, there are four measurable properties (variables) of a gas. These variables are P (pressure), V (volume), T (temperature), and n (number of moles). In Boyle's experiment, which two variables were held constant?

*Both the temperature (T) and the number of moles of gas (n) were held constant in Boyle's J-tube experiment.*

2. Fill in the blanks to summarize the relationship among the gas properties in Boyle's experiment:

*For a fixed **number of moles** of gas at constant **temperature**, the **pressure** of a gas increases as the **volume** of its container decreases.*

3. Pressure is defined in physics as force divided by area ( $P = \text{force/area}$ ). According to the kinetic-molecular theory, the particles in a gas are constantly moving and colliding with the walls of their container. The pressure of the gas is related to the total force exerted by the individual collisions. Use the kinetic theory to explain the results of Boyle's experiment.

*In Boyle's experiment, the pressure of air increased when it was compressed into a smaller volume container. According to the kinetic theory, confining the gas particles in a smaller volume will increase the number of collisions and hence the total force of the collisions with the container walls. (The distance the particles must travel between collisions decreases as the volume is reduced.)*

4. The pressure scale on a tire gauge is marked in units of pounds per square inch (psi). The scale starts at zero when the gauge is exposed to the surrounding air. This means that the total pressure is equal to the gauge pressure *plus* the pressure of the surrounding air. Standard atmospheric pressure (1 atm) is equal to 14.7 psi. Assume that you have just inflated the tire on your bicycle to 82 psi using a bicycle pump. What is the *total pressure* of air in the tire in psi? In atmospheres?

*Relative (gauge) pressure = Total pressure - Atmospheric pressure*

*82 psi = Total pressure - 14.7 psi*

*Total pressure = (82 + 14.7) psi = 97 psi (two significant figures)*

*Convert to atm:  $97 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} = 6.6 \text{ atm}$*

**Sample Data and Results Table** *(Student data will vary.)*

| Barometric Pressure |                          |                 |                       | 29.9 inches Hg = 15 psi |                          |                 |                       |
|---------------------|--------------------------|-----------------|-----------------------|-------------------------|--------------------------|-----------------|-----------------------|
| Trial 1             |                          |                 |                       | Trial 2                 |                          |                 |                       |
| Gauge Pressure      | Volume of Air in Syringe | Total Pressure* | 1/V†                  | Gauge Pressure          | Volume of Air in Syringe | Total Pressure* | 1/V†                  |
| 42 psi              | 2.0 mL                   | 57 psi          | 0.50 mL <sup>-1</sup> | 54 psi                  | 1.6 mL                   | 69 psi          | 0.63 mL <sup>-1</sup> |
| 35 psi              | 2.4 mL                   | 50 psi          | 0.42 mL <sup>-1</sup> | 35 psi                  | 2.4 mL                   | 50 psi          | 0.42 mL <sup>-1</sup> |
| 22 psi              | 3.3 mL                   | 37 psi          | 0.30 mL <sup>-1</sup> | 23 psi                  | 3.2 mL                   | 38 psi          | 0.31 mL <sup>-1</sup> |
| 17 psi              | 3.8 mL                   | 32 psi          | 0.26 mL <sup>-1</sup> | 14 psi                  | 4.4 mL                   | 29 psi          | 0.23 mL <sup>-1</sup> |
| 11 psi              | 5.0 mL                   | 26 psi          | 0.20 mL <sup>-1</sup> | 6 psi                   | 6.2 mL                   | 21 psi          | 0.16 mL <sup>-1</sup> |
| 6 psi               | 6.2 mL                   | 21 psi          | 0.16 mL <sup>-1</sup> | 3 psi                   | 7.3 mL                   | 18 psi          | 0.14 mL <sup>-1</sup> |
| 1 psi               | 8.4 mL                   | 16 psi          | 0.12 mL <sup>-1</sup> | 0 psi                   | 9.0 mL                   | 15 psi          | 0.11 mL <sup>-1</sup> |

\*See *Post-Lab Question #2*. †See *Post-Lab Question #5*.



**Answers to Post-Lab Questions** (*Student answers will vary.*)

- Convert the local barometric pressure to psi units and enter the value to the nearest psi in the Data and Results Table. Some appropriate conversion factors are shown below.

$$1 \text{ atm} = 760 \text{ mm Hg} = 29.92 \text{ in Hg} = 14.7 \text{ psi}$$

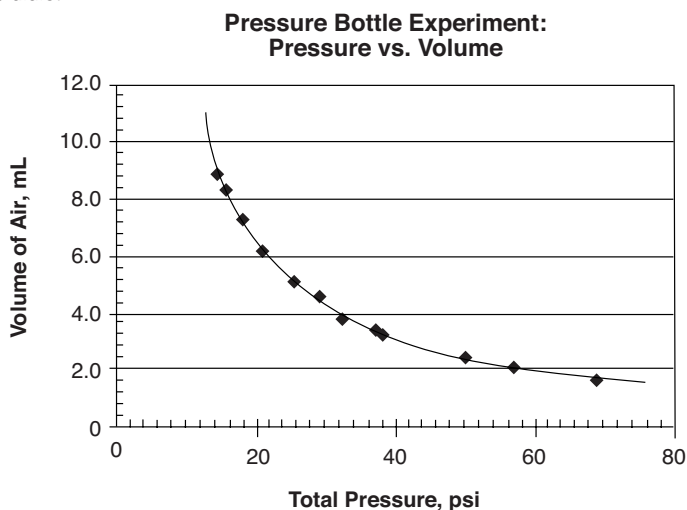
$$29.9 \text{ in Hg} \times \frac{14.7 \text{ psi}}{29.92 \text{ in Hg}} = 15 \text{ psi (rounded to the nearest psi)}$$

- The pressure gauge measures the relative pressure in psi above atmospheric pressure. For each pressure reading in the Data and Results Table, add the local barometric pressure to the gauge pressure to determine the total pressure of air inside the pressure bottle. Enter the total pressure to the nearest psi in the table.

*Sample calculation for Trial 1: Total pressure = 42 psi + 15 psi = 57 psi*

*Refer to the Data and Results Table for the results of the other calculations.*

- Plot a graph of volume on the *y*-axis versus total pressure on the *x*-axis. *Note:* The origin of the graph should be (0,0). Choose a suitable scale for each axis so that the data points fill the graph as completely as possible. Remember to label each axis and give the graph a title.



- Describe the shape of the graph. Draw a best-fit straight or curved line, whichever seems appropriate, to illustrate how the volume of a gas changes as the pressure changes.

*The graph is curved. The volume decreases as the pressure increases. At first, there is a sharp reduction in the volume as the pressure increases. The decrease in volume then becomes more gradual and the volume appears to level off as the pressure increases further. Mathematically, the shape of the curve is described as hyperbolic. A hyperbolic curve of this type is obtained when there is an inverse relationship between two variables ( $y \propto 1/x$ ). See the graph for the best-fit curved line through the data.*

- The relationship between pressure and volume is called an “inverse” relationship—as the pressure increases the volume of air trapped in the syringe decreases. This inverse relationship may be expressed mathematically as  $P \propto 1/V$ . Calculate the value of  $1/V$  for each volume measurement and enter the results in the Data and Results Table.

*Sample calculation for Trial 1:*

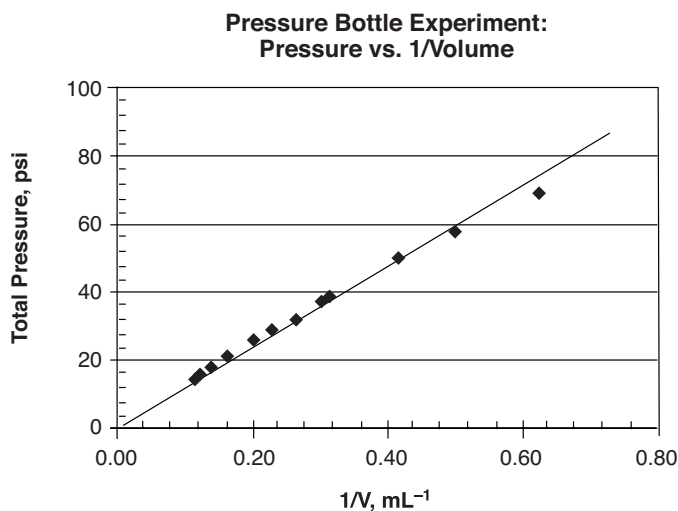
$$V = 2.0 \text{ mL}$$

$$1/V = 0.50 \text{ mL}^{-1}$$

*Refer to the Sample Data and Results Table for the results of the other calculations.*

## Pressure vs. Volume and Boyle's Law *continued*

6. Plot a graph of pressure on the y-axis versus  $1/V$  on the x-axis and draw a best-fit straight line through the data. *Note:* The origin of the graph should be (0,0). Choose a suitable scale for each axis so that the data points fill the graph as completely as possible.



7. Another way of expressing an inverse relationship between two variables ( $P \propto 1/V$ ) is to say that the product of the two variables is a constant ( $P \times V = \text{constant}$ ). Multiply the total pressure ( $P$ ) times the volume ( $V$ ) for each set of data points. Construct a Results Table to summarize the  $P \times V$  values.

### Sample Results Table

| Trial 1                  |        |                  | Trial 2 |        |                  |
|--------------------------|--------|------------------|---------|--------|------------------|
| P (psi)                  | V (mL) | P × V (psi · mL) | P (psi) | V (mL) | P × V (psi · mL) |
| 57                       | 2.0    | 110              | 69      | 1.6    | 110              |
| 50                       | 2.4    | 120              | 50      | 2.4    | 120              |
| 37                       | 3.3    | 120              | 38      | 3.2    | 120              |
| 32                       | 3.8    | 120              | 29      | 4.4    | 130              |
| 26                       | 5.0    | 130              | 21      | 6.2    | 130              |
| 21                       | 6.2    | 130              | 18      | 7.3    | 130              |
| 16                       | 8.4    | 130              | 15      | 9.0    | 140              |
| <b>Average Value</b>     |        |                  |         |        | 120              |
| <b>Average Deviation</b> |        |                  |         |        | 7                |

8. Calculate the average value of the  $P \times V$  "constant" and the average deviation. What is the relative percent error (uncertainty) in this constant?

$$\text{Relative percent error} = (\text{Average deviation} / \text{Average value}) \times 100\%$$

*Average value of the ( $P \times V$ ) constant = 120 psi · mL*

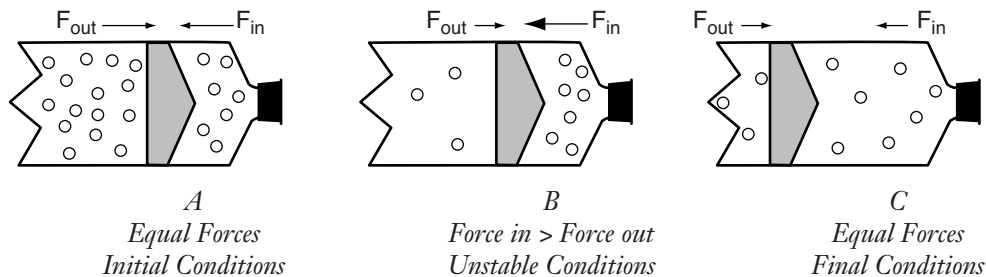
*Average deviation = 7 psi · mL*

*Relative percent error = (7 psi · mL / 120 psi · mL) × 100% = 6%*

*It appears that the mathematical product ( $P \times V$ ) is a constant within plus-or-minus 6%.*

9. At constant temperature, the pressure of a gas is proportional to the concentration of gas particles in the container. When some of the pressure was released from the pressure bottle, the syringe plunger moved up. Why did this happen? Use diagrams and explain in words what happens to the gas particles moving around both inside and outside the syringe before and after the pressure is released.

*Initially, the pressure was the same both inside and outside the syringe. This means there were equal concentrations of gas particles colliding with the inside and outside of the plunger, so the plunger stayed in place (A). When some of the pressure from the bottle was released, some air particles escaped, leaving fewer air particles to collide on the outside of the plunger (B). The concentration of particles inside the syringe overpowered those outside the syringe and pushed the plunger outward (the volume increased). As the volume increased, the particle concentration inside the syringe decreased. When the concentration of particles (pressure) inside the syringe was the same as that outside the syringe, the plunger stopped moving (C).*



10. (Optional) Research the properties of high density PETE on the Internet. What characteristics of PETE make it an ideal plastic for use in soda bottles?

*PETE is an ideal plastic for soda bottles for several reasons. It is transparent, crystal clear, pure, tough, and unbreakable. PETE also has very low permeability to oxygen, carbon dioxide, and water, and excellent chemical resistance to acids and bases. If the PETE does not contain additives, it is relatively easy to recycle.*

## Supplementary Information

### Construction of a Pressure Bottle

Cut off the top portion of a 2-L soda bottle. This will be used as a base or cap handle to hold the bottle caps that must be drilled. Screw one bottle cap securely onto the bottle top. Using a 1/2-inch drill and drill press, drill a hole through the center of the cap, as shown in Figure 5. (Notice you are drilling through the bottom of the cap.) Remove the cap from the bottle top, add rubber cement to the lip of the valve stem, and use pliers to pull the valve stem through the hole from the inside, as shown below. The tire valve will stick out of the top of the bottle cap when it is placed on the soda bottle. Clear away any burrs that may remain in the drilled hole. They will cause the seal to break.

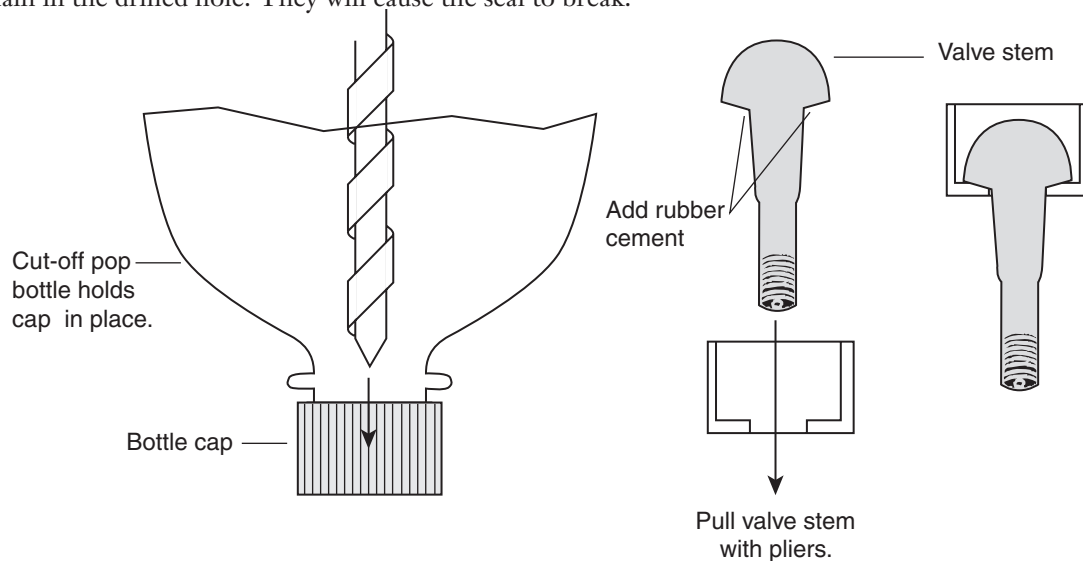
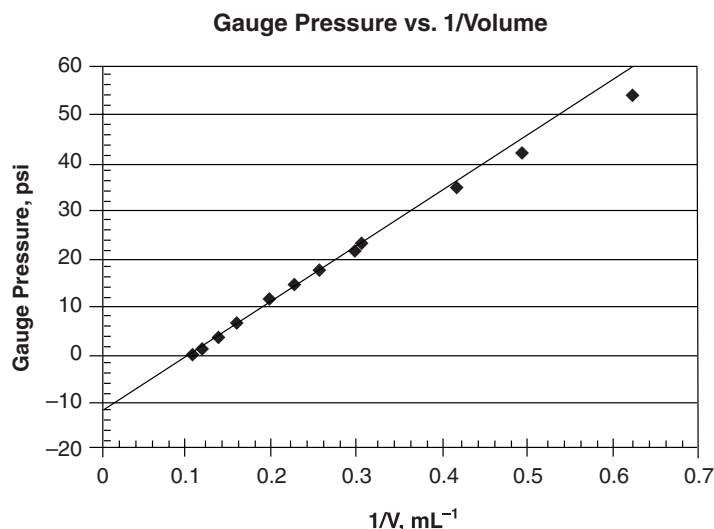


Figure 5.

## Graphical Analysis and the Value of the Atmospheric Pressure

As observed in Post-Lab Question #6, the graph of total pressure versus  $1/V$  is a straight line that passes through the origin (y-intercept equal to zero). This graph corresponds to an equation of the form  $P_{\text{total}} = m \times 1/V$ , where  $m$  is the slope of the line. If instead of the total pressure, students plot the gauge pressure ( $P_{\text{gauge}} = P_{\text{total}} - P_{\text{air}}$ ) versus  $1/V$ , they will obtain a straight line of the form  $P_{\text{gauge}} = (m \times 1/V) - P_{\text{air}}$  (where the y-intercept is equal to  $-P_{\text{air}}$ ). The value of  $P_{\text{air}}$  can be estimated by extending the best-fit straight line backwards to "negative" gauge pressures. The best estimate we were able to obtain using this method was about 12 psi (18% error).



## Algebraic Determination of the Value of the Atmospheric Pressure

It is also possible to use the gauge pressure–volume measurements to estimate a value for the atmospheric pressure in psi. This is a nice algebraic challenge for students. According to Boyle's law, multiplying any two sets of P–V data should give the same result ( $P_1V_1 = P_2V_2$ ). This is only true if the pressures are total pressures rather than relative gauge pressures ( $P_{\text{total}} = P_{\text{gauge}} + x$ , where  $x$  = value of atmospheric pressure in psi).

The following equation is obtained by combining the equations for Boyle's law and the total pressure:

$$(P_{\text{gauge}, 1} + x)(V_1) = (P_{\text{gauge}, 2} + x)(V_2)$$

Substituting the raw data for 17 psi and 1 psi (see Sample Data on page 11) into this equation gives:

$$(17 \text{ psi} + x)(3.8 \text{ mL}) = (1 \text{ psi} + x)(8.4 \text{ mL})$$

$$64.6 + 3.8x = 8.4 + 8.4x$$

$$56.2 = 4.6x$$

$$x = 12 \text{ psi}$$

## Reference

This experiment has been adapted from *Flinn ChemTopic™ Labs*, Volume 9, The Gas Laws, Cesa, I., ed., Flinn Scientific, Batavia, IL, 2003.

## Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Pressure vs. Volume and Boyle's Law* activity, presented by Bob Becker, is available in *Boyle's Law*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

## Materials for *Pressure vs. Volume and Boyle's Law* are available from Flinn Scientific, Inc.

Materials required to perform this activity are available in the *Boyle's Law in a Bottle—Student Laboratory Kit* available from Flinn Scientific. Materials may also be purchased separately.

| Catalog No. | Description                                    |
|-------------|--|
| AP6855      | Boyle's Law in a Bottle—Student Laboratory Kit |
| AP5930      | Pressure Bottle                                |
| AP6884      | Bicycle Pump, with Gauge and Release Valve     |
| AP1730      | Syringe, without Needle, 10 mL                 |
| AP8958      | Syringe Tip Cap, Pkg/10                        |

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.