

Gases and Absolute Zero

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Purpose: To investigate the way in which the volume occupied by a gas sample at constant pressure changes with pressure and to determine the location of absolute zero.

Equipment: 600 mL beaker, 145 mL flask, meter stick, barometer, glass tubing, leveling bulb, rubber connectors, gas measuring tube, pinch clamp, ring stand, utility clamp, iron ring, Bunsen burner, wire gauze, rubber hose, thermometer

Materials: distilled water

Introduction: Gases have several unique properties that separate them from the other phases of matter. The volume of a gas is affected by changes in pressure, unlike solids and liquids, which have constant volume regardless of pressure. The relationship between pressure and volume in a gas is inversely proportional; that is, the pressure times the volume equals some constant. This was stipulated by Robert Boyle in Boyle's Law, which states $PV = \text{constant}$. This experiment will attempt to verify Boyle's Law experimentally.

In this experiment, pressure will be measured using a barometer. This device uses liquid mercury to measure atmospheric pressure. The pressures exerted by gases can be derived from the partial pressures exerted by the gases. For example, when a gas is collected over water, the total pressure includes the pressure exerted by the water vapor over the water. This pressure can be deducted from the total pressure by looking up the pressure exerted by water vapor at a certain temperature, as the pressure varies with temperature. This process is an important part of this experiment, as the partial pressure of the gas being collected is required, not the total pressure.

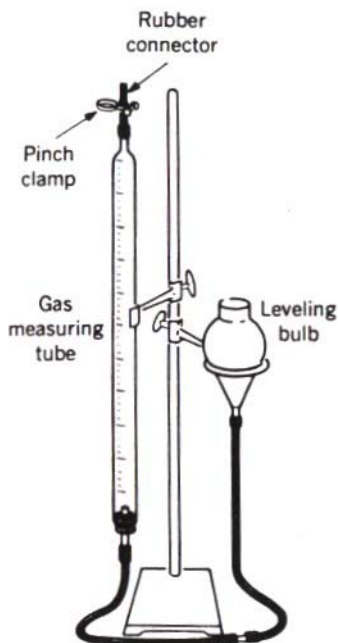
This experiment also aims to measure absolute zero, the lowest attainable temperature. In order to determine absolute zero, the pressures exerted by gases will be

measured at different temperatures, and then a graph will be constructed based on that data. A best-fit line will be used to estimate absolute zero.

Procedure:

Part A

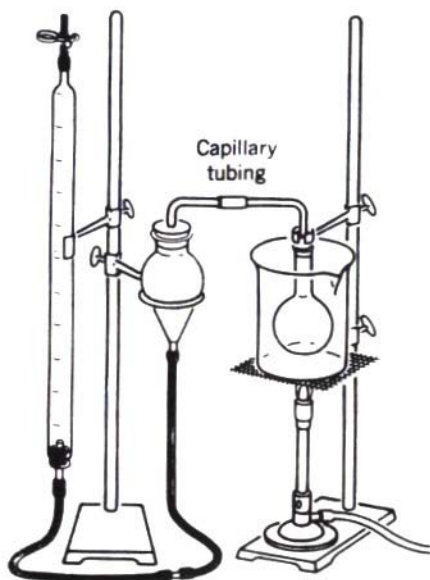
1. The following apparatus was assembled from the labeled parts:



2. Approximately 40 mL of air was trapped in the gas-measuring tube. The clamp, indicated above, was set on the rubber connector. The leveling bulb was raised and lowered to chase air bubbles out of the rubber tubes.
3. The height of the leveling bulb was adjusted to that the water in it was at the same level as the water in the gas-measuring tube. The volume of the gas sample was recorded.
4. The leveling bulb was raised one meter. The gas volume was recorded.
5. The leveling bulb was lowered one meter. The gas volume was recorded.
6. The water temperature and barometric pressure were measured and recorded.

Part B

1. The following apparatus was assembled from the labeled parts:



2. With the pinch clamp open and the bottom of the neck of the leveling bulb opposite the 40-mL mark, room temperature water was poured into the leveling bulb until it was filled to the neck. The temperature of the water was measured and recorded.
3. Water was poured into the beaker to the level of the bottom of the rubber stopper. The temperature of the water in the beaker was recorded.
4. The pinch clamp was closed and the gas volume was recorded. The water was heated. Every time the temperature of the water had increased 10 degrees, the gas measuring tube was lowered until the meniscus was level with the leveling bulb. The temperature of the water in the beaker was measured and recorded, and the level of the gas in the tube was measured and recorded. This step was repeated until the temperature reached 100 degrees.

Observations:

Part A.

Note: All observations were calibrated in the following manner. Original readings from the gas measuring tube (which are not reported) were calibrated by adding 4.6 mL to each reading. The measurements reported have already been calibrated according to this factor. Temperature readings are calibrated by interpolating the measured temperature (which is not reported) on a calibration curve obtained during a previous experiment.

Volume of gas (levels equal) = $44.6 \pm .1$ mL

Volume of gas (bulb 1.00 m above) = $45.0 \pm .1$ mL

Volume of gas (bulb 1.00 m below) = $40.1 \pm .1$ mL

Temperature of water = $21.0 \pm .2$ °C (Note: this temperature was obtained with an alcohol thermometer, not the mercury thermometer used throughout the rest of this experiment. The alcohol thermometer provided a more accurate reading of water temperature. This is the only instance in the experiment where this thermometer was used.)

Barometric pressure = 769 mm Hg

Part B.

Temperature of water in leveling bulb = $42.0 \pm .2 \text{ }^\circ\text{C}$

Barometric pressure = 769 mm Hg

Volume of gas and temperature of flask:

Temperature of flask	Volume of gas
$40 \pm .2 \text{ }^\circ\text{C}$	$45.4 \pm .1 \text{ mL}$
$50 \pm .2 \text{ }^\circ\text{C}$	$44.2 \pm .1 \text{ mL}$
$60 \pm .2 \text{ }^\circ\text{C}$	$43.5 \pm .1 \text{ mL}$
$70 \pm .2 \text{ }^\circ\text{C}$	$42.6 \pm .1 \text{ mL}$
$80 \pm .2 \text{ }^\circ\text{C}$	$41.7 \pm .1 \text{ mL}$
$90 \pm .2 \text{ }^\circ\text{C}$	$41.2 \pm .1 \text{ mL}$
$100 \pm .2 \text{ }^\circ\text{C}$	$40.2 \pm .1 \text{ mL}$

Results:

Part A.

Observed pressure of wet gas. (barometric pressure) = **769 mm Hg**

Calculated pressure of dry gas (barometric pressure minus vapor pressure of H_2O) =
 $769 \text{ mm Hg} - 18.65 \text{ mm Hg} = \mathbf{750.35 \text{ mm Hg}}$

Volume of dry gas (volume of wet gas) = **44.6 mL**

PV product for dry gas (pressure of dry gas times volume of dry gas) =
 $750.35 \text{ mm Hg} \times 44.6 \text{ mL} = \mathbf{33466 \text{ mm Hg}\cdot\text{mL}}$

Bulb 1.00 m above:

Hg equivalent of 1.00-m H_2O column (height of column times reciprocal of weight of equivalent Hg column times 1000 conversion factor) =
 $1.00 \text{ m H}_2\text{O} \times (1 \text{ m Hg} / 13.5 \text{ m H}_2\text{O}) \times (1000 \text{ mm Hg} / 1 \text{ m Hg}) = \mathbf{74.1 \text{ mm Hg}}$

Observed pressure of wet gas (barometric pressure plus Hg equivalent of 1.00-m H_2O column) =
 $769 \text{ mm Hg} + 74.1 \text{ mm Hg} = \mathbf{843.1 \text{ mm Hg}}$

Calculated pressure of dry gas (observed pressure of wet gas minus vapor pressure of H_2O) =
 $843.1 \text{ mm Hg} - 18.65 \text{ mm Hg} = \mathbf{824.45 \text{ mm Hg}}$

PV product for dry gas (calculated pressure of dry gas times volume of gas) =
824.45 mm Hg x 40.4 mL = **33308 mm Hg-mL**

Bulb 1.00 m below:

Observed pressure of wet gas (barometric pressure minus Hg equivalent of 1.00-m H₂O column) =
769 mm Hg – 74.1 mm Hg = **694.9 mm Hg**

Calculated pressure of dry gas (observed pressure of wet gas minus vapor pressure of H₂O) =
694.9 mm Hg – 18.65 mm Hg = **676.25 mm Hg**

PV product for dry gas (calculated pressure of dry gas times volume of dry gas) =
676.25 mm Hg x 35.5 mL = **24007 mm Hg-mL**

Part B.

Temperature in flask (°C)	Pressure of dry gas in tube (mm Hg)	Pressure in flask (mm Hg)	Pressure in flask if volume constant (mm Hg)
40			769
50	733	788.4	794.9
60	744.8	800.2	810.7
70	760.5	815.9	831.7
80	776.9	832.3	853.4
90	786.4	841.7	866.1
100	805.9	861.3	892.2

Notes: The temperature in the flask was observed directly with a mercury thermometer.

The pressure of the dry gas in the tube was calculated by multiplying the original pressure (barometric pressure minus vapor pressure of H₂O) by the initial volume divided by the observed volume. The pressure in the flask was calculated by adding the vapor pressure of H₂O to the pressure of the dry gas in the tube. The pressure in the flask if the volume were constant was calculated by multiplying the pressure in the flask by (the sum of 145 and the net volume change in the measuring tube) divided by 145.

A graph was constructed, plotting the temperature in the flask on the x-axis against the pressure in the flask if the volume were constant on the y-axis. A best-fit line was drawn until an x-intercept was located. The graph is attached to this report. The best-fit line indicates that absolute zero, according to the experimental data, is approximately at -350°C .

Discussion: A percent error calculation can be performed on the absolute zero obtained experimentally.

$$\% \text{ error} = (|\text{observed}| - |\text{true}|) / |\text{true}| \times 100$$

$$350 - 273 / 273 \times 100 =$$

28.21% error.

This is a fair percent error considering the wide range of sources of error in this experiment. The procedure calls for the gas measuring tube to be moved in order to line up the leveling bulb with the meniscus in the tube. However, during the reported trial of this experiment, the tube was not raised or lowered to coincide with the leveling bulb. This led to incorrect readings of the meniscus, which in turn led to incorrect measurements of the amount of gas in the tube. Furthermore, the mercury thermometer used during the experiment was severely distorted. Water at 21.0 degrees Celsius was reportedly 43.0 degrees Celsius according to the mercury thermometer that was used for the reported trial. However, the discrepancy was not discovered until after the trial had taken place. The possibility for further, more egregious errors, coupled with the unavailability of a calibrated mercury thermometer, made another trial unfeasible. Another source of error was that the 145-mL flask used in the experiment may not have been 145 mL in volume. That figure is an average of several flasks available in the laboratory, since the actual volume of the flask was not measured or recorded at any point during the procedure.

Further sources of error include using an uncalibrated alcohol thermometer for one of the temperature readings. This produced inaccurate results, but results that were more accurate than the broken mercury thermometer. Another source of error is the failure to use capillary tubing to connect the entire gas-measuring apparatus. Instead, cross pieces with small volumes were used. This volume was not negligible; it detracted from the amount of gas in the gas measuring tube. However, there is no practical way to measure the amount of gas lost due to this equipment problem. Further sources of error include the practical failure to heat the flask uniformly. Since the water around the flask was being heated, and not the flask itself, it was impossible to ensure that the gas inside the flask was being heated uniformly. Furthermore, the neck of the flask had to stick out of the water slightly to allow a clamp to suspend the flask in the water-filled beaker. Finally, if there were any leaks in the system, whether detectable or not, the results of the experiment would be completely invalid. This plethora of sources of error, coupled with laboratory inexperience, produced enough error to distort the experimental value of absolute zero obtained in this experiment.

The theory associated with this experiment is the kinetic theory of gases. The kinetic theory of gases is responsible for the behavior of the gases in part A and explains why part B works. The kinetic theory states that gases consist of large numbers of molecules; that the volume of all of the molecules is negligible compared to the total volume in which the gas is contained; that attractive and repulsive forces between gas molecules are negligible; that energy can be transferred between molecules during collisions, but the average kinetic energy remains the same; and that the average kinetic energy of the mole-

cules is proportional to the temperature. Absolute zero is an average kinetic energy of zero. This means that there is no motion of particles, thus no volume.

There are many ramifications to this experiment. Most involve laboratory experience. New apparatuses were used in the laboratory during this experiment. New procedures were learned and practiced. This experiment served as troubleshooting practice, as many things went wrong due to the many sources of error. In terms of practical applications, the knowledge of the workings of gases could be applied to produce desired results in any industrial field. Pressure and volume can be adjusted under the assumption that they are inversely proportional to produce the necessary conditions for a gas.

Questions:

1. If there were a small air leak in the rubber connector at the top of the gas-measuring tube, when the water level in the bulb was 1.00 m above the level in the gas-measuring tube, the PV product would have been too low. The volume reading would have been lower than it should have been, causing a low product.
2. The effect on the results would have been negligible. Any temperature change between successive parts of a procedure when performed immediately after one another could not be great enough to change the results to any significant degree.
3. The decreasing solubility of air in water with increasing temperature would cause an error in the determination of absolute zero on the lower end of the temperature spectrum. Air dissolved in water increased the volume readings at low temperatures, but that air was not present at higher temperatures. Thus, the volume readings were more accurate at

higher temperatures. The best-fit line running through the plotted points would be shifted due to this error, causing the experimental value of absolute zero to be too low.

4. In order to make the calculations for alcohol instead of water, the vapor pressure for alcohol would be needed, as well as the ratio of millimeters alcohol to millimeters mercury, as provided in the introduction.

5. If the thermometer readings were always 2°C too low, the experimental value of absolute zero would be too low. The graph, and thus the best-fit line, would be shifted to the left, causing the x-intercept to move to the left and become more negative.

Conclusion: The experiment was completed to a reasonable degree of accuracy, considering the abundance of sources of error. The behavior of gases was investigated, and absolute zero was calculated with fair accuracy.