Vapor Pressure of Water

Introduction

Dry air in buildings and homes causes many health problems during the winter months. Symptoms of low relative humidity include dry skin and rashes, sore throats, and coughs. The amount of water vapor that the air can hold depends on temperature. Why does warm air hold more water than cold air?

Concepts

- Evaporation and condensation
- Kinetic-molecular theory
- Vapor pressure
- Ideal gas law and Dalton's law

Experiment Overview

The purpose of this experiment is to determine the vapor pressure of water at different temperatures. A small amount of air will be trapped in an inverted graduated cylinder and heated to 80 °C (see Figure 1). As the temperature increases, the air will expand and become saturated with water vapor. The volume of gas will be measured at 5 °C-intervals as the water bath is cooled to 50 °C. Finally, ice will be added to cool the gas to 0 °C and the volume of "dry air" will be recorded—at this temperature, the vapor pressure of water is low and we assume that the trapped gas contains only air.

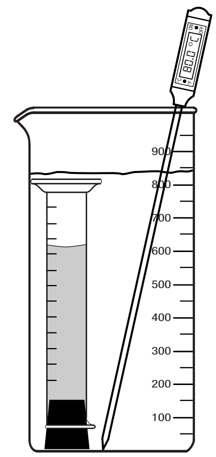


Figure 1.

Materials

- Barometer
- Beaker, tall-form, 500-mL or 1-L
- Beral-type pipet, jumbo
- Gloves, heat-protective
- Graduated cylinder, glass, 25-mL
- Hot plate or Bunsen burner set-up
- Ice
- Ring stand and ring (optional)
- Rubber stopper, 1-or 2-hole
- Thermometer, digital
- Stirring rod (plastic)
- Water

Safety Precautions

Exercise care when working with the hot water bath. To avoid scalding, wear heat-protective gloves or use a hot vessel gripping device when removing the beaker of hot water from the heat source. Turn the hot plate or burner off when not in use. Wear chemical splash goggles whenever working with chemicals, heat or glassware in the chemical laboratory.

Procedure

1. Fill a 500-mL or 1-L tall-form beaker about two-thirds full with *hot tap water* (45–50 °C).
2. Add 20–21 mL of distilled water to a 25-mL graduated cylinder and place a two-hole rubber stopper in the cylinder. Close the stopper holes with your finger and quickly invert and lower the graduated cylinder into the beaker filled with hot water. *Note:* There should be about 10 mL of air trapped in the inverted graduated cylinder.
3. Add more hot tap water as needed to the beaker so that the entire graduated cylinder is under water and the trapped air is also surrounded by the water bath (refer to Figure 1).
4. Place a digital thermometer in the hot water and carefully heat the beaker on a hot plate at a medium setting or with a very low Bunsen burner flame until the water temperature is about 80 °C. *Note:* Do NOT allow the volume of gas to expand beyond the scale marked on the graduated cylinder.
5. Turn off the hot plate or Bunsen burner and carefully remove the heat source. Stir the water in the beaker to ensure an even distribution of heat. Record the temperature of the hot water bath and the volume of gas in the graduated

cylinder.

6. Continue stirring the water and allow the beaker and contents to slowly cool. Measure and record the volume of gas and the temperature at 5 °C-intervals until the temperature is 50 °C. *Hint:* Don't cool too fast.
7. Cold water or crushed ice may be added to speed up the cooling process. *It is important, however, to keep the water level in the beaker about the same throughout the experiment.* Remove water as needed using a jumbo plastic pipet.
8. After the temperature has reached 50 °C, cool the water rapidly to about 0 °C by adding ice to the beaker. *Note:* Remove water as needed to keep the overall level of ice and water surrounding the graduated cylinder the same (see step 7).
9. When the temperature of the ice-water bath is between 0–5 °C, measure and record the volume of gas and the temperature. Record the barometric pressure. Note the units!

Disposal

None required.

Tips

- The water bath should be cooled slowly in order to achieve good heat transfer.
- It is not practical to raise the graduated cylinder out of the water bath to equalize the pressure of the gas inside the cylinder with atmospheric pressure. The temperature of the gas cools rapidly when the graduated cylinder is lifted out of the bath, giving rise to large errors in the resulting volume measurements. The height difference between the gas level in the graduated cylinder and the water level in the beaker is thus a source of error. The pressure inside the cylinder is slightly greater than atmospheric pressure because the trapped gas is supporting the extra height of the column of water (assuming the water level in the beaker is above the gas level in the graduated cylinder). The pressure difference may be accounted for by measuring the height (h) in millimeters of the water level above the gas level in the graduated cylinder. Sample results were obtained without applying this pressure correction.

$$P_{\text{cylinder}} = P_{\text{atm}} + h\ (\text{mm } H_2O) \times (1 \text{ mm Hg}/13.6 \text{ mm } H_2O)$$

- The effect of temperature on the vapor pressure of water is important in meteorology. Relative humidity is defined as the percentage of moisture that the air contains compared to the maximum that it can hold at a specific temperature (based on the saturation vapor pressure of water). The dew point is the temperature at which the amount of water vapor in the air would reach 100% relative humidity and start to condense.
- This activity was adapted from Solids and Liquids, Volume 11 in the *Flinn ChemTopic™ Labs* series; Cesa, I., Editor; Flinn Scientific, Batavia IL (2005).

Calculations and Analysis

1. Gas law calculations require that the temperature be in degrees kelvin (K). Convert temperature readings to kelvin.
2. There is a small error in the volume of gas caused by using the upside-down graduated cylinder (the meniscus is reversed). Correct volume measurements by subtracting 0.2 mL and convert to liters (L). Label this the *corrected volume* (V_{corr}).
3. Assume that "dry air" is the only gas in the inverted graduated cylinder at 0–5 °C. Use the ideal gas law to calculate the *moles of air* (n_{air}) in the graduated cylinder at 0–5 °C. *Hint:* The pressure is equal to the barometric pressure.
4. The moles of air (n_{air}) in the graduated cylinder is constant. Use the ideal gas law to calculate the *partial pressure of air* (P_{air}) in the graduated cylinder at each temperature between 50–80 °C. *Note:* Watch your units!
5. Between 50–80 °C, both air and water vapor are present. The total pressure is equal to the barometric pressure (P_{atm}). Use Dalton's law to calculate the *vapor pressure of water* (P_{water}) in mm Hg at each temperature.
6. Plot the vapor pressure of water in mm Hg versus temperature in °C, including points for 2 mm Hg at 0 °C and 760 mm Hg at 100 °C. Draw a *smooth-fit curve* through the data points.

Discussion

The rate of evaporation of a liquid depends on the nature of the substance and varies with temperature. Imagine a liquid placed in a closed container at a given temperature. As molecules of liquid evaporate from the liquid surface, the number of molecules in the gas phase, and hence the pressure, will increase. As the number of gas molecules increases, some of the molecules may collide with the surface of the liquid and condense. Eventually, the rate of evaporation of the liquid molecules will become equal to the rate of condensation of the gas molecules, and the pressure of the vapor will reach a constant or equilibrium value. The *vapor pressure* of a liquid is defined as the pressure exerted by a vapor in equilibrium with the pure liquid at a specific temperature. Liquids with high vapor pressures are considered volatile—they will evaporate readily from an open container.

The kinetic-molecular theory provides a model for explaining why the vapor pressure of a liquid increases as the temperature increases. Molecules in the liquid state are in constant motion. They are close enough together, however, that attractive forces between molecules keep them in a definite volume and influence their properties. At a given temperature, some of the molecules will be moving fast enough and have sufficient kinetic energy to break free from the "ties that bind them" and escape into the vapor phase. Increasing the temperature increases the average kinetic energy of the molecules. At higher temperatures, a larger number of molecules will be moving fast enough and have enough kinetic energy to overcome the forces of attraction in the liquid and enter the vapor phase. The number of molecules in the vapor phase above a liquid, and hence the vapor pressure, increases with increasing temperature.

Sample Data

Barometric Pressure *746 mm Hg*

Water Temperature, °C	*80.8*	*74.8*	*69.9*	*64.6*	*59.8*	*54.7*	*49.5*	*5.0*
Volume of Gas, mL	*18.4*	*15.2*	*13.2*	*12.3*	*11.2*	*10.6*	*10.2*	*7.8*

Sample Calculations and Analysis

1. Gas law calculations require that the temperature be in degrees kelvin (K). Convert temperature readings to kelvin.

   $K = °C + 273.2$
   Sample calculation: $5.0\ °C + 273.2 = 278.2\ K$

2. There is a small error in the volume of gas caused by using the upside-down graduated cylinder (the meniscus is reversed). Correct volume measurements by subtracting 0.2 mL and convert to liters (L). Label this the *corrected volume* (V_{corr}).

   $V_{corr} = V - 0.2\ mL$
   Sample calculation: $V_{corr} = 7.8\ mL - 0.2\ mL = 7.6\ mL\ (0.0076\ L)$ *at* $5.0\ °C\ (278\ K)$

3. Assume that "dry air" is the only gas in the inverted graduated cylinder at 0–5 °C. Use the ideal gas law to calculate the *moles of air* (n_{air}) in the graduated cylinder at 0–5 °C. *Hint:* The pressure is equal to the barometric pressure.

   $$n_{air} = \frac{P_{atm} \times V}{RT} = \frac{(0.982\ atm)(0.0076\ L)}{(0.0821\ L\ atm/mole\ K)(278\ K)} = 3.27 \times 10^{-4}\ moles$$

4. The moles of air (n_{air}) in the graduated cylinder is constant. Use the ideal gas law to calculate the *partial pressure of air* (P_{air}) in the graduated cylinder at each temperature between 50–80 °C. *Note:* Watch your units!

   Sample calculation for $T = 80.8\ °C\ (354\ K)$, $V_{corr} = 18.2\ mL\ (0.0182\ L)$:

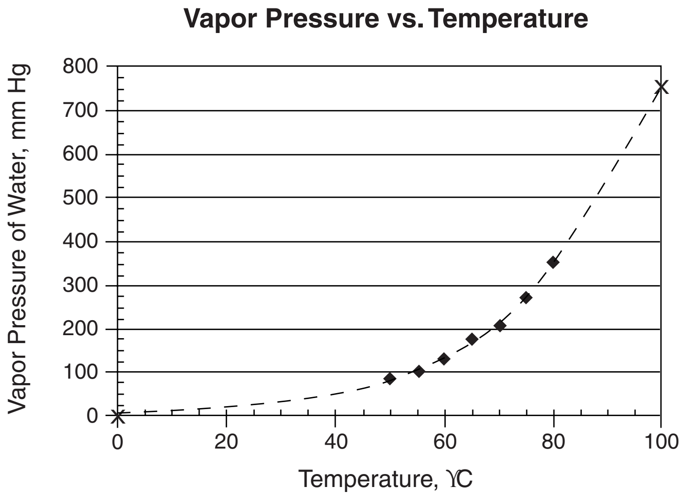
   $$P_{air} = \frac{(3.27 \times 10^{-4}\ moles)(0.0821\ L\ atm/mole\ K)(354\ K)}{(0.0182\ L)} = 0.522\ atm$$

5. Between 50–80 °C, both air and water vapor are present. The total pressure is equal to the barometric pressure (P_{atm}). Use Dalton's law to calculate the *vapor pressure of water* (P_{water}) in mm Hg at each temperature.

   Sample calculation for $T = 80.8\ °C\ (354\ K)$, $P_{atm} = 0.982\ atm$, $P_{air} = 0.522\ atm$:
   $P_{water} = 0.982\ atm - 0.522\ atm = 0.46\ atm$

$$P_{water} = 0.46\ atm \times \frac{760\ mm\ Hg}{1\ atm} = 350\ mm\ Hg$$

6. Plot the vapor pressure of water in mm Hg versus temperature in °C, including points for 2 mm Hg at 0 °C and 760 mm Hg at 100 °C. Draw a *smooth-fit curve* through the data points.



Sample Results

Temperature (°C)	Temperature (K)	V_{corr} (L)	P_{air} (atm)	P_{water} (atm)	P_{water} (mm Hg)
80.8 °C	*354 K*	*0.0182 L*	*0.521 atm*	*0.461 atm*	*350 mm Hg*
74.8 °C	*348 K*	*0.0150 L*	*0.622 atm*	*0.360 atm*	*274 mm Hg*
69.9 °C	*343.1 K*	*0.0130 L*	*0.707 atm*	*0.275 atm*	*209 mm Hg*
64.6 °C	*337.8 K*	*0.0121 L*	*0.748 atm*	*0.234 atm*	*178 mm Hg*
59.8 °C	*333 K*	*0.0110 L*	*0.811 atm*	*0.171 atm*	*130 mm Hg*
54.7 °C	*327.9 K*	*0.0194 L*	*0.845 atm*	*0.137 atm*	*104 mm Hg*
49.5 °C	*322.7 K*	*0.0100 L*	*0.865 atm*	*0.117 atm*	*89 mm Hg*
5.0 °C	*278.2 K*	*0.0076 L*	*0.981 atm*	*NA*	*NA*

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 B: Physical Science, structure and properties of matter, motions and forces, interactions of energy and matter

Materials for *Vapor Pressure of Water* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP4288	Graduated Cylinder, Polypropylene, 25 mL
GP1060	Beaker, Borosilicate Glass, Berzelius, 500 mL
AP8716	Flinn Digital Thermometer
AP5070	Barometer, Metric
AP8850	Beral-Type Pipets, Super Jumbo

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