Lab #11 Gas Laws Laboratory

Purpose: a. Part I: To examine the relationship between the temperature and pressure of a constant volume of gas (air) in order to determine experimentally the value of absolute zero. (Gen. & Univ. Phys I)

b. Part II: To examine the relationship between pressure & kinetic energy of an enclosed gas. (Univ. Phys I)

Equipment: Absolute zero demonstrator device (hollow sphere attached to a pressure gauge)
Several water baths at different temperatures from boiling to freezing
Thermometers

Discussion Part One:

The following is known as the ideal gas law:

$$P \cdot V = n \cdot R \cdot T \qquad (1)$$

P stands for pressure, and is defined as force per unit area, with units of newtons per square meter (N/m^2), or pascals (Pa), in the metric system (pounds per square inch (lb/in^2) in the British system). V represents volume (m^3), and T represents temperature in kelvins.

The actual quantity (mass) of gas is given by n, and represents the number of moles of the gas. Just as distance can be measure in feet or meters, mass of a substance can be measured in grams or moles. A mole is the amount of a substance that has a mass in grams numerically equal to the molecular weight of the substance. For example, one mole of water (H_2O) has a mass of 18 grams, while one mole of oxygen gas (O_2) has a mass of 32 grams. The thing that ANY mole of a substance has in common with a mole of some other substance is that they always contain the SAME NUMBER OF MOLECULES, Avogadro's number ($6.022 \cdot 10^{23}$). It can easily be shown that equal volumes of gases at the same temperature and pressure contain an equal number of molecules. This last statement is known as Avogadro's Principle.

R in equation (1) above is the universal gas constant, and when pressure is measured in Pa, volume in m^3, n in moles, and temperature in kelvins, it has a value of 8.31 Joule/mole • K.

Consider a process or problem in which the temperature of a confined (constant number of moles) gas is to remain constant. Then equation (1) becomes:

$$P\ V = \text{constant}, \qquad \text{or } P_1\ V_1 = P_2\ V_2$$

This is known as Boyle's Law: For a confined gas at constant temperature, pressure is INVERSELY proportional to volume.

Next, consider a process in which the pressure of a confined gas is to remain constant. (Perhaps the gas is in a cylinder with a freely-moving piston, so that we can observe the relation between temperature and volume.) Past experience tells us that when something is heated, it expands; this is confirmed by equation (1), which now becomes:

‘$V/T = (n \cdot R) / P = \text{constant}$ or $V_1 / T_1 = V_2 / T_2$ or $V_1 / V_2 = T_1 / T_2$

This is known as the Law of Charles and Gay-Lussac: For a confined gas at constant pressure, volume is DIRECTLY proportional to absolute temperature.

The third and final variable to hold constant in equation (1) is volume, which brings us to today’s laboratory. The hollow metal sphere of the demonstrator is filled with air whose pressure is measure with an absolute pressure gauge. As the temperature of that constant volume is change (by immersion in a liquid bath of known temperature), then the pressure changes can be read from the gauge. The temperature-pressure relationship of the gas does not depend on the quantity of gas used, nor does it depend on the initial pressure of gas so long as the quantity of gas (number of molecules) does not change during the experiment. Now equation (1) becomes:

$P / T = (n \cdot R) / V = \text{constant}$ or $P_1 / T_1 = P_2 / T_2$ or $P_1 / P_2 = T_1 / T_2$

In other words, for a confined gas at constant volume, pressure is DIRECTLY proportional to absolute temperature.

We can combine these three relations as follows:

(2)

The gas we will be using during this lab exercise is air, while air is a real (not ideal) gas, its behavior, like that of most real gases, is very closely approximated by equation (1). Remember: Whenever dealing with gas laws, the temperature numbers in a problem must be in kelvins.

Procedure Part I:

1. Using the absolute zero demonstrator device, measure and record five different temperatures and pressures for the air contained in your sphere. Begin with freezing OR boiling, but increase or decrease bath temperatures IN ORDER (Don’t go from hot to cold to warm to cool, etc.). During each immersion, make sure the sphere is completely submerged and leave it there long enough to obtain an accurate final reading. Tap the gauge lightly to ensure no sticking of the needle. Read to the nearest kPa.
2. Change the amount of gas of your sphere by activating the pin valve with ballpoint pen or car key to equalize the pressure between the room and the sphere. Do so just after recording the fifth and final temperature / pressure change of your

first run (when the sphere will be very hot or very cold). Now repeat step 1 with this different number of moles of air in the sphere.

3. Plot the pressure (kPa) as a function of temperature (°C) for both runs. By extrapolation, determine for each run the temperature at which the pressure goes to zero. Average these two values to arrive at your experimental answer for absolute zero, which defines the starting point of the Kelvin absolute temperature scale. Compare your result to the actual value of –273.15 °C. Determine percent error.

Data Part I:

Run 1:	Temperature (°C)	Pressure (kPa)	KE (J)
	______________	____________	______
	______________	____________	______
	______________	____________	______
	______________	____________	______
	______________	____________	______
Run 2:			
	______________	____________	______
	______________	____________	______
	______________	____________	______
	______________	____________	______
	______________	____________	______

Discussion Part Two – (University Physics I Lab):
You have seen that as the temperature of a gas increases, so does the pressure of the gas. The increase in pressure is an indication of the increase of kinetic energy of the air molecules themselves. The kinetic energy of the molecules of an enclosed gas is given by:

KE = 3/2 k • T • N   for gases having one atom per molecule

or   KE = 5/2 k • T • N   for gases having two atoms per molecule

Since air is mainly composed of nitrogen and oxygen, both of which exist as diatomic molecules, we will use the latter equation today. In the equations above, k is Boltzman's constant, and equals $1.38 \cdot 10^{-23}$ J/K. T is the temperature of the gas in kelvins, and N is the number of molecules.

To determine the number of molecules in your sphere, we can use equation (2) as follows:

$$\frac{P_o \cdot V_o}{T_o} = \frac{P_A \cdot V_A}{T_A}$$

where P_o and V_o are the standard pressure and volume respectively at STANDARD temperature T_o (STP), and P_A, V_A, and T_A are the actual pressure, volume, and temperature of the gas in your sphere. Standard pressure is $1.013 \cdot 10^5$ Pa, and standard temperature is 0 °C (273 kelvins). With this information and knowing that one mole of gas ($6.022 \cdot 10^{23}$ molecules) occupies a volume of $22.4 \cdot 10^3$ cm^3 at standard temperature and pressure, the number of molecules in the sphere can be determined. (Use a proportion.) The ACTUAL volume of your sphere has been found to be 552 cubic centimeters.

Procedure Part II:

1. Before you begin Part I, run 1, the pressure in your sphere was equalized at room pressure while at room temperature, so we will use run 1 data already gathered. Find the pressure and temperature of the classroom and determine the STANDARD volume for the gas in your sphere.

2. From the standard volume, determine the number of molecules you had in the sphere during run 1.

3. Convert Celsius degrees to kelvins for the run 1 temperatures, and calculate the kinetic energy of the gas for each temperature.

4. Plot the pressure as a function of kinetic energy. What kind of relation do you see?