Physics 12 IB   **The Gas Laws**

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1   Introduction

The lab activity investigates the behavior of gases by analyzing measurements of pressure, volume, temperature and mass. The experiments make use of the macroscopic properties of fixed masses of gas (air) to investigate a number of gas laws such as a) Boyle's law b) Gay-Lussac's Law and c) the Pressure Law.

2   Aims

1. To determine the relationship between the pressure of a fixed mass of gas and its volume at constant temperature.
2. To determine the relationship between the volume of a fixed mass of gas and its temperature at constant pressure.
3. To determine the relationship between the pressure of a fixed mass of gas and its temperature at constant volume.
4. To determine Absolute Zero by extrapolation.

3   Data Collection

Part I

Room Temperature = 22.4 ℃
Volume uncertainty = ± 0.5 cm^3
Pressure uncertainty = ± 0.25 psi

See Table 1

Part II

Initial pressure of gas = pressure of air = 1.01×10^5 Pa
Diameter of capillary tube   = 1.5 mm
Absolute uncertainty of ruler   = ± 0.5 mm
Absolute uncertainty of thermometer   = ± 0.5 ℃

Volume (cm^3)	Pressure (psi)
8	17.0
10	14.9
12	12.5
15	11.0
17	10.4
20	10.1
21	10.3
25	9.6

Table 1: Constant Temperature

T (℃)	Length (cm)
7.5	4.7
11.0	4.6
13.5	4.8
31.0	5.3
38.0	5.5
48.0	6.0
49.7	5.9
55.5	6.2
61.0	6.5

Table 2: Constant Pressure

T (℃)	Pressure (kPa)
17.0	96.11
23.0	97.63
30.0	99.80
30.0	99.80
48.0	105.78
53.0	107.32
60.0	109.21
65.0	110.97
72.0	112.56
74.0	112.94
75.5	114.00

Table 3: Constant Volume

See Table 2

Part III

Volume of Florence flask = 250 ml ± 5% ml
Temperature uncertainty = ± 0.5 ℃

See Table 3

4   Data Analysis

Attached

5   Evaluation

In this lab we were able to verify the 3 primary gas laws by determining the relationship between the pressure, volume and temperature of a fixed mass of gas:

1. Boyle's Law
2. Gay-Lussac's Law
3. The Pressure Law

Part I

By analyzing the graphs, there seems to be an inverse relationship between pressure and volume of a fixed mass of gas at constant temperature. This conclusion was reached due to the following reasons:

1. The graph of pressure versus volume looks like the graph of $\frac{1}{x}$.
2. The graph of volume versus the reciprocal of pressure is a straight line.

The graph of the product of pressure and volume versus pressure was too erroneous to deduce any useful results. According to the conclusion reached, this relationship should have been a linear one.

We thus conclude that

$$P \; \alpha \; \frac{1}{V} \tag{1}$$

Part II

The slope of the graph that shows the variation of volume with temperature is a constant and represents the ratio between some constant k and the pressure P.

The equation of the graph is:

$$V = 0.0024T + 0.3027$$

In order to determine the temperature (in ℃) for which the volume of the gas would be zero, we let $V = 0$ and solve for T. In doing so, we obtain a T of -126.13 ℃.

It is necessary to keep in mind that in trying to compute this temperature, we have let the volume of the gas equal to zero. However, the gas of mass m is matter, and according to definition, all matter occupy space and cannot therefore be volume-less. Since $V \neq 0$, it follows that a temperature of -126.13 ℃ is unattainable.

It thus seems that the numerical value -126.13 is the absolute minimum temperature for the gas. More accurate experiments in the past have shown that the absolute minimum temperature for any gas is -273.15 ℃.

Sources for error in this experiment could have arisen from the following sources:

1. The gas in use (air) wasn't of low density.
2. A constant pressure wasn't maintained during the experiment.
3. Inaccurate measurements of the length and diameter of the air column leading to an incorrect volume calculation.

We thus conclude that the temperature of a gas (in Kelvins) is directly proportional to its volume for a given fixed pressure, or mathematically

$$T \; \alpha \; V \tag{2}$$

Part III

The equation for the graph that shows the relationship between pressure and temperature for a fixed volume is

$$P = 0.3057T + 90.795$$

The slop of the above linear relationship is the coefficient of T or 0.3067. This value represents the ratio $\frac{P}{T}$ and will later be shown to be equal to a constant k divided by the fixed volume under which the experiment was performed.

The graph of a V vs. T graph is also a straight line, but as this line is horizontal, its slope is 0. This characteristic arises from the fact that for any given temperature, the volume of the gas remains the same as the gas was confined to a rigid container that could neither expand nor contract. On the contrary, the slope of a pressure versus temperature curve is non-zero, but remains constant because the relationship is linear.

We thus conclude that the relationship between pressure and temperature a gas of fixed volume is a linear one. That is to say that

$$P \; \alpha \; T \tag{3}$$

for a fixed volume of the gas.

Part IV

By extrapolating the trendline of the pressure vs. temperature graph to lower and lower temperatures, the line crosses the x-axis at $-297.0\,°C$. In reality, no gas can be cooled to this temperature because all gases liquefy before reaching it. It then seems that $-297.0\,°C$ is the *absolute zero point* for temperature measurement. This means that temperatures lower than this value cannot be reached by continually cooling a gas or any other substance. If lower temperatures could be reached, then further extrapolation of the straight line in the graph would suggest that negative absolute gas pressures could exist. Such a situation would be impossible, because a negative absolute gas pressure has no meaning.

Nevertheless, the graph has been redrawn such that there is no negative temperature. This is the Kelvin scale and has deliberately chosen so that its zero temperature point is the lowest temperature attainable.

There still seems to be an error in our calculation as past (and more accurate) experiments have determined the absolute temperature for a substance to be $-273.15\,°C$.

Percentage Discrepancy $= \frac{297.06-273.15}{273.15} = 8.75\%$

6 Conclusion

Although not experimented upon, it is quite obvious to expect that there exists a relation between the absolute pressure and the number of molecules in a gas. It is possible to increases the pressure of a gas by adding more molecules, and this is why pumping more air into a balloon or a tire inflates it. Past experiments have indicated that when the volume and temperature of a low-density gas are kept constant, doubling the number of molecules doubles the pressure. Thus, the absolute pressure of a gas is directly proportional to the number of mlecules or, equivalently, to the number of moles n of the gas.

$$P \; \alpha \; n \tag{4}$$

The four relationships discussed above (Equations (1), (2), (3) and (4)) between pressure, volume, temperature and number of moles of a low-density gas may be succinctly summarized as follows:

$$P \; \alpha \; \frac{nT}{V}$$

This proportionality can be written as an equation by inserting a proportionality constant R. The resulting equation is known as the **ideal gas law**.

$$PV = nRT \tag{5}$$

where T is the Kelvin temperature, and R is the constant of proportionality called the *universal gas constant*. The value R has been determined experimentally to be 8.31 J/(mol·K) for any real gas whose density is sufficiently low to ensure ideal gas behavior.