Lab 1:
The Ideal Gas Law

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Intermediate Experimental Physics
Section 002

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The Objective of this week's experiment is to verify the widely-acknowledged law that the pressure of a gas decreases as its temperature increases. This law, now so commonplace that it is applicable to undergraduate physics courses, was pioneered by Robert Boyle & Jacques Charles. In this lab, we will use principles of scientific inquiry to continue their legacy.

Theoretical Background/ Abstract:

Gasses, when kept under isothermic conditions (in which temperature is kept constant) have an interesting property. The pressure of a gas will vary with the reciprocal of the gas's volume. These two parameters, pressure and volume, are said to be complementary. This becomes clear when presented with the following formula, known as Boyle's law.

$$P_1 V_1 = P_2 V_2.$$

It is also important that we consider the fact that the volume of a gas under isobaric conditions (constant temperature) is directly proportional to its temperature:

$$V \propto T$$

It is thanks to the brilliant mind of Émile Clapeyron that we are able to combine these notions of gasses—along with some concepts borrowed from stoichiometry—into one meaningful statement about how the temperature, pressure, and volume of a gas are related. This **Ideal Gas Law** (validated in its importance with the three-letter-acronym, IGL) states that the product of an ideal gas's pressure and volume is equal to its absolute temperature (in Kelvin) times it's number of moles and the ideal gas constant, equal to 8.31 J / K • mol.

This is most often shown with the formula:

$$PV = nRT = NK_BT$$

where *P* is pressure; *V* is Volume; *n* is the number of moles in the apposite gas; *R* is the Ideal Gas Constant, explicated above; and *T* is the temperature in absolute terms, Kelvin.

We now find ourselves scratching our heads, asking *what on earth is an ideal gas?* An ideal—or, more aptly put, idealized—gas is defined in the following terms:

- The molecules of the gas are indistinguishable, small, hard spheres (or points).
- All collisions are elastic and all motion is frictionless (no energy loss in motion or collision)

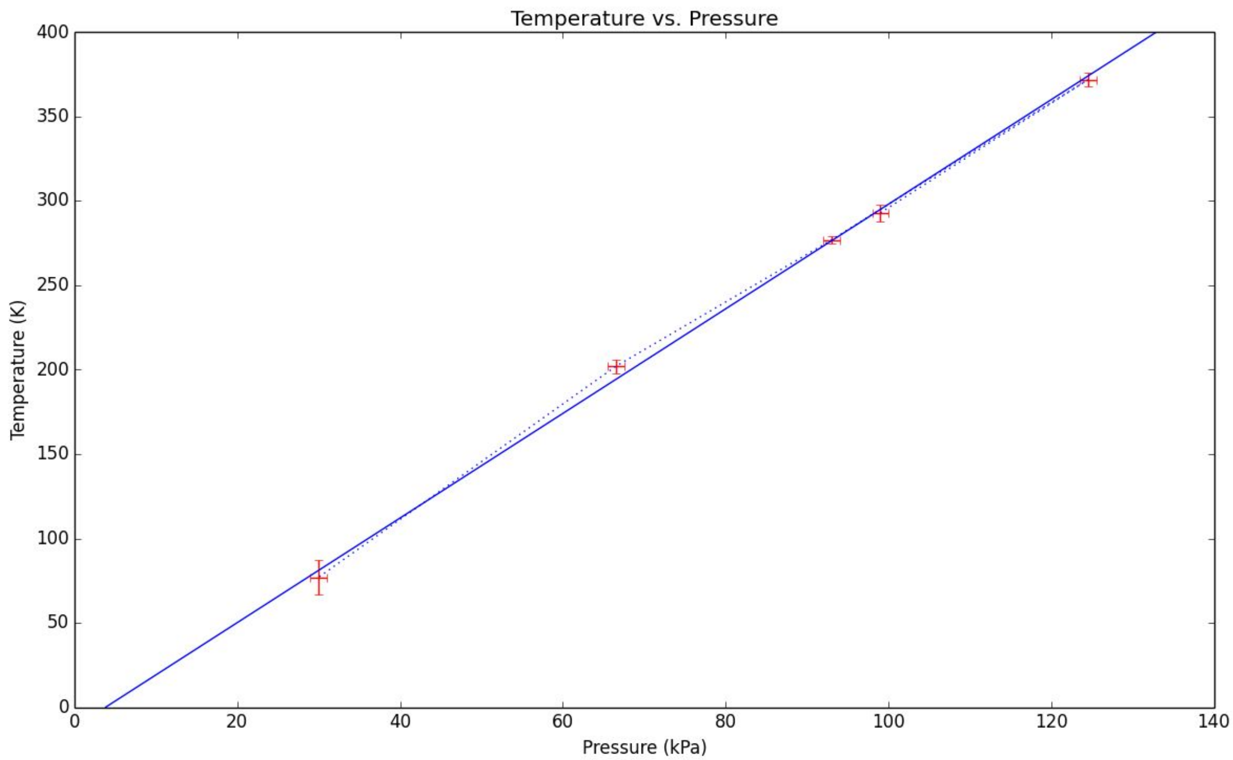
- Newton's laws apply
- The average distance between molecules is much larger than the size of the molecules
- The molecules are constantly moving in random directions with an even distribution of speeds
- There are no attractive or repulsive forces between the molecules apart from those that determine their point-like collisions
- The only forces between the gas molecules and the surroundings are those that determine the point-like collisions of the molecules with the walls
- In the simplest case, there are no long-range forces between the molecules of the gas and the surroundings.

Procedure:

1. Record the atmospheric pressure in kiloPascals and the ambient temperature in degrees celsius. Release the valve on the pressure gauge to ensure that the gas in the bulb is at barometric equilibrium with the surrounding environment.
2. Submerse the bulb in a bucket of liberally-stirred ice-water, along with a thermometer with a suitable range of measurement. Wait for the needle to come to a standstill to indicate that thermal equilibrium has been achieved!
3. Record the temperature of the water (and the gas bulb), as well as the corresponding pressure.
4. Repeat steps 2 and 3 with the following mixtures:
    a. Boiling water
    b. Liquid Nitrogen (do not use a thermometer!)
    c. Alcohol + solid CO_2 (dry ice)

Experimental Data:
Data and results of calculations for the volume of air under various conditions

#	Pressure (kPa)	Temperature (K)	Estimated uncertainty in pressure (kPa)	Terror (K)
1	99.0	293.0	±0.5	±2.0
2	124.5	372.0	±0.5	±2.0
3	93.0	277.0	±0.5	±2.0
4	66.5	202.0	±0.5	±2.0
5	30.0	77.0	±0.5	±10.0



Answers to Questions:

1. What are the dimensions of temperature?

   Temperature is a fundamental physical dimension, often represented by Θ

2. List the units of temperature that you know.

   Planck Temperature, Delisle Scale, Kelvin, Rankine, Degrees Celsius and Fahrenheit

3. What are the dimensions of m and b?

   Equation of a line: $y = mx + b \Rightarrow$

   $$Temperature = m \cdot Pressure + Temperature \Rightarrow \begin{array}{l} b \sim Temperature \\ m \sim \dfrac{Temperature}{Pressure} \end{array}$$

4. What units of m and b are you using? Hint: They are not the same.

   $b \sim K$, $m \sim \dfrac{K}{Pa}$

Results and Error Analysis:

After much satisfying work at the ol' lab bench, my partner and I were able to make some important conclusions about the nature of gases-ideal. In plotting our data, we were pleased to see that the temperatures of our gas at various pressures lined up famously. That is to say that **our results were linear**, which is consistent with the current understanding of ideal gasses. The datapoints did not deviate from our best-fit line significantly. Most of our error bars were within a suitable range, and compensated for any discrepancy between values indicated by the best-fit line and the plotted points.

The data procured by the results of this experiment contain small quantities of error, or deviations from the expected results. This error, however, is not substantial enough to keep us from drawing broad conclusions about the relationship between pressure and temperature in an isochoric system at fixed mass.

We gave a 0.5 kiloPascal margin of error to all of our pressure calculations, because our instrument of measurement had subdivisions of 1 kPa. The uncertainty in temperature in the lowest measurement is a consequence of our not being able to stick a thermometer in the liquid nitrogen. Interestingly, this "measurement" is probably the most accurate of our plotted points, as derived it from the vaporization temperature of nitrogen, which is invariant at constant temperature and air pressure.

Our value for absolute zero is -280.5°C, which is about 5C° lower than we had expected. Possible reasons for this include possible swelling/contracting of the steel bulb at temperature extrema, and inflated inaccuracy of measurement by using three thermometers. Possible reasons for bulb inflation or contraction include the fact that steel experiences linear expansion of volume when exposed to an increase in temperature. This is given by the equations:

$$\Delta L = \alpha L_0 \Delta T \qquad \Delta V = \beta V_0 \Delta T$$

As we know, gasses also expand and contract with varying temperatures. Seeing as we reached 125 kPa of pressure in the bulb, it is not outrageous to imagine that the volume of the bulb was larger at 372 K than it was at 77 K, when the pressure was only 30 kPa. At this low-point, it's possible to assume that the atmospheric pressure in the lab has a compressing effect on the bulb.