VERIFICATION OF GAS LAWS

Experiment Title:

Verification of Gas Laws

Date Experiment Performed: September 13th, 2019

Date Experiment Submitted: September 20th, 2019

Attach here (if required, indicate the appropriate document(s)):

Medical or other Acceptable Document: **Not applicable**

Change of Lab Day Form: **Not applicable**

Change of Lab Section Form: **Not applicable**

Late Pass: **Not applicable**

Procedure – Charles' Law: As outlined in the lab manual ("What in the World ISN'T Chemistry", General Chemistry Laboratory Manual, Dr. Rashmi Venkateswaran, 2019.).

Procedure – Boyle's Law:

1. Obtain a clean, dry 20 mL syringe.
2. Attach the syringe to the Gas Pressure Sensor.
3. Trap a 10 mL volume of gas in the syringe.
4. Hold the plunger in place until LabQuest 2 displays a stable number for pressure.
5. Record the pressure.
6. Repeat steps 3 and 4 for volumetric measurements of 4 mL, 6 mL, 8 mL, 12 mL, 14 mL and 16 mL of gas.

Discussion:

Charles' Law

Table 1: Data obtained from the lab which includes initial and final temperatures in both degrees celsius and kelvin, volume from the erlenmeyer flask (Volume$_{cw}$), the total amount of water in the erlenmeyer flask (Volume$_1$), and the calculated volume (Volume$_2$) for Trial 2.

Trial	**Initial Temperature of Boiling Water (℃)**	**Initial Temperature of Boiling Water (K)**	**Initial ice bath temperature (℃)**	**Final ice bath temperature (℃)**	**Final ice bath temperature (K)**	**Volume$_{cw}$ (mL)**	**Volume$_1$ (mL)**	**Volume$_2$ (mL)**
2	100	373.15	10.6	13	286.15	32	156.4	124.4

Table 2: Table of results which indicates the constant of [V1/T2] and [V2/T2], the percent error, and the averages for Trial 2.

Trial	$\frac{V1\ mL}{T1\ K}$	$\frac{V2\ mL}{T2\ K}$	**Percent Error (%)**
2	0.42	0.43	-3.7

According to Maeng & Bell (2013), "kinetic molecular theory explains Charles's law, the direct relationship between the temperature and volume of a gas." (p.39) Charles' Law describes the proportional relationship between the temperature and, at a constant pressure, the volume of gas. This proportional relationship can be demonstrated through relating initial temperature over initial volume and final temperature over final volume ($\frac{V1}{T1} = \frac{V2}{T2}$) and is observed in *Table 2*. During the lab, the procedure involved the heating of a dry flask and transferring it to an ice bath. In this case, T1 and T2 are the temperatures that correspond to the volumes V1 and V2. T1 relates to the temperature of the boiling water whereas T2 corresponds to the ice bath temperature. V2 was then calculated from the volume in the erlenmeyer flask and V1 (Volume$_2$

= Volume_1 - $\text{Volume}_{cw)}$, as noted in the *appendix sample calculations for Charle's law.* The flask after being heated and transferred to the ice bath quickly began to fill with water due to atmospheric pressure as the flask cooled.

Although Trial 1 and Trial 2 are similar, Trial 2 will be discussed due to its higher accuracy, as noted in the *appendix additional data for Charles' law Table 5*. Trial 2 clearly illustrates the close relationship between 0.42 from V1/T2 and 0.43 from V2/T2, as noted from Table 2. This similarness is indicative of Charle's law, that gases expand and contract at a given temperature, as the constant from (V1/T1) 0.42 is very close to 0.43 from V2/T2. In Charles' Law the theoretical value would have $\frac{V1}{T1} = \frac{V2}{T2}$ exactly equate to each other. In theory, V2 should have been 119.9 mL rather than the acquired 124.4 mL, see *Table 1*. However, during the lab, the results being rendered can be due to the source of error that occurred. Furthermore, the percent error is -3.2%, see *Table 2*. This negative percent error indicates that yield of the experiment was lower than anticipated. All of which can be attributed to sources of error that occurred.

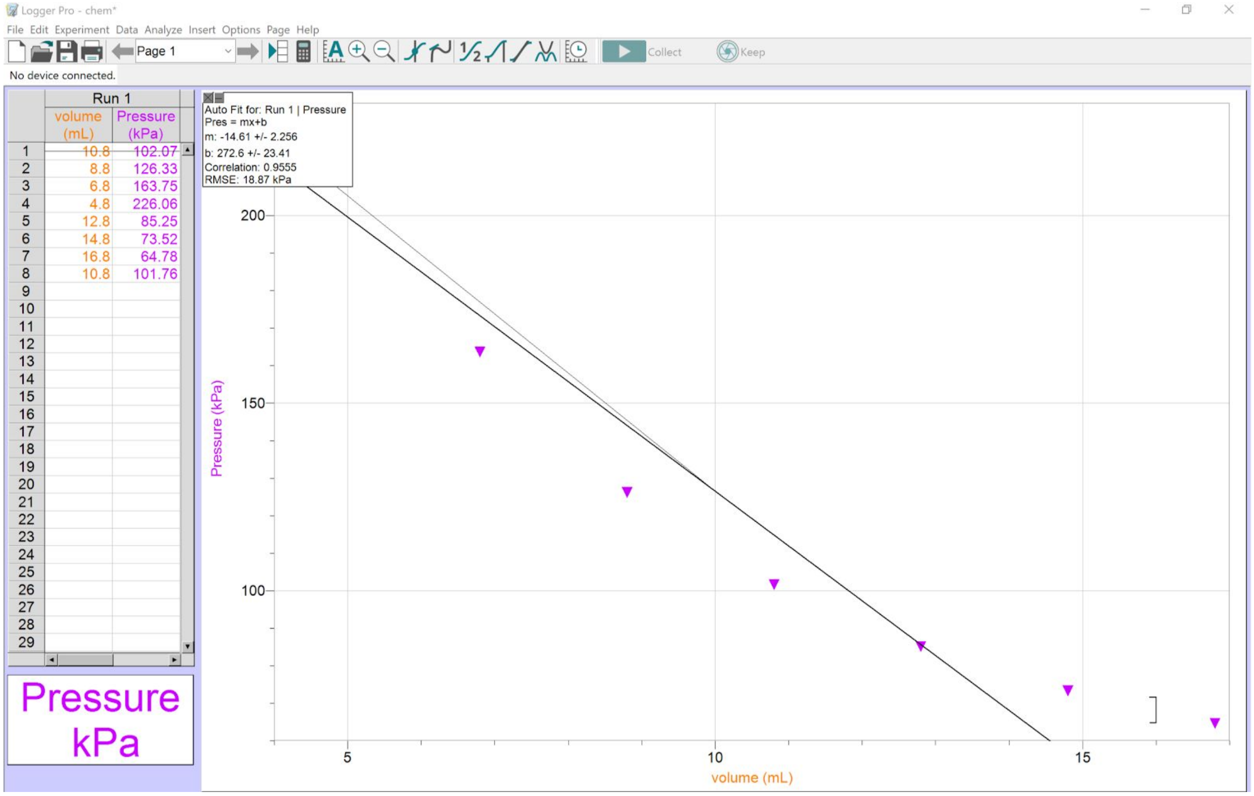
One of the main sources or error could be the possibility that the flask was not allowed to heat for the proper amount of time and thus resulted in a limited expansion of air in the flask. This then would go on to skew the results in regards to the calculation of Volume_2 as the proper amount of gas would be unable to be recorded. This further accounts for the negative percent error. Similarly, this also could be the result of the flask not being able to cool down appropriately. This once again would affect volume_2 as the proper amount of gas would not be noted. Referring back to Trial 1, Trial 2 is more precise based on the 0.01 difference between

$\frac{V1}{T1}$ and $\frac{V2}{T2}$. Whereas, Trial 1 results had a 0.05 difference between $\frac{V1}{T1}$ and $\frac{V2}{T2}$ as 0.40 is the constant from $\frac{V1}{T1}$ and 0.45 is from $\frac{V2}{T2}$. The larger difference in Trial 1, once again indicates the source of error that may have occurred. Regardless, the constant calculated from the acquired is similar in size and thus goes on to verify Charles' Law.

Boyle's Law

Table 3: Values of pressure for each volume of gas obtained from the second trial, and the calculated constant

TRIAL 2		
Volume of gas (mL)	**Pressure exerted (kPa)**	**Boyle's Law constant**
4.8	226.06	1.085088
6.8	163.75	1.1135
8.8	126.33	1.111704
10.8	101.76	1.099008
12.8	85.25	1.0912
14.8	73.52	1.088096
16.8	64.78	1.088304



	Run 1	
	volume (mL)	Pressure (kPa)
1	10.8	102.07
2	8.8	126.33
3	6.8	163.75
4	4.8	226.06
5	12.8	85.25
6	14.8	73.52
7	16.8	64.78
8	10.8	101.76
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Auto Fit for: Run 1 | Pressure
Pres = mx+b
m: -14.61 +/- 2.256
b: 272.6 +/- 23.41
Correlation: 0.9555
RMSE: 18.87 kPa

Figure 1: Plotting the second trial's values for the independent variable, volume, and the dependent variable, pressure, depicts the inversely proportional relationship, as shown by the line of best fit.

The two variables needed to verify Boyle's law are volume and pressure, with the former being our independent variable and the latter, our dependent variable. That is because it is much easier to record and control the volume with the way the experiment has been set up. Under constant temperature, gas and pressure are inversely proportional--as one variable increases the other decreases. This relationship was consistently exhibited in both Figure 1 and Figure 2. The ratio between each volume of gas multiplied by its corresponding pressure value should, theoretically, equal to the same number. As shown in Table 3, the products for each of the volume measurements multiplied by the corresponding pressure were all very similar. The mean of each of the calculated constant values will be calculated to determine a final Boyle's law constant, which which will be used to derive a formula.

Boyle's law constant: 1.0967

<u>Boyle's law derivation</u>

$$P \propto 1/V, \text{ therefore}$$

$$P = c(1/V), \text{ therefore}$$

$$PV = c, \text{ constant and } P_1V_1 = P_2V_2 = \ldots = c \text{ [under constant temperature]}$$

Upon reviewing our procedures, we realized that we assumed that there would be no drastic changes in temperature that would affect the trapped gas's behavior during the experiment. This was an oversight. It would have been best to wait at least a minute after altering the volume before reading the pressure, as the change in volume most likely produced an amount of heat that would have affected the gas's behavior. This would allow the gas to have time to return back to room temperature. This degrades the accuracy and validity of our results, as now the pressure was measured at varying temperatures. Heat increases particle mobility (heat → kinetic energy), so the volume of the gas would increase. However, gases in a fixed volume cannot expand and consequently, pressure increases. This means that our recorded pressure values would be higher than the theoretical values.

Our starting measurement was 10 mL of gas. This was because the syringe's total volume capacity was 20 mL, and choosing a volume value in the middle allowed various number values below and above 10 mL to be taken. 10 mL was also our final measurement.

When taking repeated measurements for the same variable, it is highly unlikely to attain the same exact result twice. It is almost certain that there will be some variation/inconsistencies in these measurements. This is due to the fact that these variations occur randomly, which implies that there is virtually no way to be as precise by only taking the values once and/or control them. Taking repeated measurement ensures a more accurate value is obtained and decreases the potential amount of sources of error. The chosen repetition of the first and final mark is once again due to taking more precise measurements. The initial trial is almost always less precise especially when new equipment is being utilized. Thus, the last measurement on the same mark decreases the amount of random error that may potentially occur. This transcends in regards to taking multiple readings of the dependent variable for the same value of the independent variable, as the dependent variable may constantly change as the experimenter does not have complete control of this variable. In contrast, the independent variable is completely controlled by the experimentar. We chose to keep our second value for 10 mL, as we felt that we were more familiar with how to properly use the equipment by the end of our experiment than our first time doing it.

Conclusion:

Our results appear to generally follow the theoretical trends of Charles' Law and Boyle's Law and thus were able to verify Charles' Law and Boyle's Law by examining the proportional and inversely proportional relationships from the acquired results.

Reference(s):

"What in the World ISN'T Chemistry", General Chemistry Laboratory Manual, Dr. Rashmi Venkateswaran, 2019.

Maeng, J., & Bell, R. (2013). THEORIES LAWS, and HYPOTHESES. *The Science Teacher, 80*(7), 38-43. Retrieved from http://www.jstor.org/stable/43557281

Appendix:

Raw Data:

Additional Data, Charles' Law

Table 4: The obtained data from the lab which includes initial and final temperatures in both degrees celsius and kelvin, volume from the erlenmeyer flask (Volume$_{cw}$), Volume$_1$ and Volume$_2$ for Trial 1 and Trial 2.

Trial	**Initial Temperature of Boiling Water (℃)**	**Initial Temperature of Boiling Water (K)**	**Initial ice bath temperature (℃)**	**Final ice bath temperature (℃)**	**Final ice bath temperature (K)**	**Volume$_{cw}$ (mL)**	**Volume$_1$ (mL)**	**Volume$_2$ (mL)**
1	100	373.15	3	4	277.15	25.1	149.1	124
2	100	373.15	10.6	13	286.15	32	156.4	124.4

Table 5: Table of results which indicates the constant of [V1/T2] and [V2/T2], the percent error, and the averages for Trial 1 and Trial 2.

Trial	$\frac{V1\ mL}{T1\ K}$	$\frac{V2\ mL}{T2\ K}$	**Percent Error (%)**	**Theoretical Value of Volume$_2$ (mL)**
1	0.40	0.45	-12	119.9
2	0.42	0.43	-3.7	110.7

Average of Trial 1 and Trial 2	0.41	0.44	-7.9	115.3

Additional Graphs, Charles' Law

Not Applicable

Sample Calculations, Charles' Law

Volume$_2$ (the volume of the gas at Temperature$_2$) Calculation Trial 1:

Volume$_2$ = Volume$_1$ - Volume$_{cw}$

Volume$_2$ = 149.1mL - 25.1mL

Volume$_2$ = 124 mL

Charles' Law Verification from Trial 1

$$\frac{V1}{T1} = \frac{V2}{T2}$$

$$\frac{V1}{T1} = \frac{149.1mL}{373.15K} = 0.3993 \approx 0.40 \ \frac{mL}{K}$$

$$\frac{V2}{T2} = \frac{124mL}{277.15K} = 0.4474 \approx 0.45 \ \frac{mL}{K}$$

Average Charles' Law Constant Calculation for Trial 1 and Trial 2

$$\text{Average } \frac{V1}{T1} = \frac{(Trial\ 1\ \frac{V1}{T1} + Trial\ 2\ \frac{V1}{T1})}{2}$$

$$\text{Average } \frac{V1}{T1} = \frac{(0.40 + 0.42)}{2}$$

$$\text{Average } \frac{V1}{T1} = 0.41$$

Percent Error Calculation for Trial 1

$$\%\ \text{Error} = \frac{\frac{V1}{T1} - \frac{V2}{T2}}{\frac{Vi}{T1}} \times 100\%$$

$$\%\ \text{Error} = \frac{\frac{149.1\ mL}{373.15\ K} - \frac{124\ mL}{277.15\ K}}{\frac{149\ mL}{373.15\ K}} \times 100\%$$

$$\%\ \text{Error} = \frac{-0.048107}{0.399303} \times 100\%$$

$$\%\ \text{Error} = -12\ \%$$

Theoretical Value of Volume$_2$ for Trial 1

$\frac{V1}{T1} = \frac{V2}{T2}$ which can be rearranged as $\frac{V1xT2}{T1} = V_2$ which then calculates Volume_2

$$V_{2\text{ Theoretical}} = \frac{V1xT2}{T1}$$

$$V_{2\text{ Theoretical}} = \frac{149.1mL\ x\ 277.15K}{373.15\ K}$$

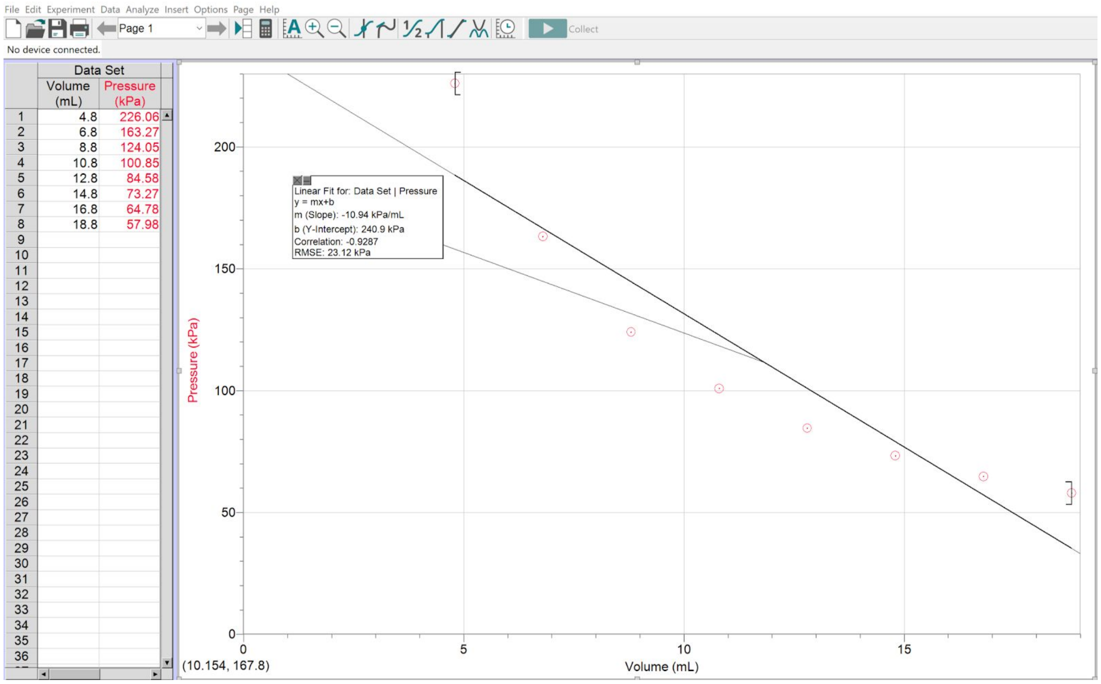
$$V_{2\text{ Theoretical}} = 110.7\text{ mL}$$

Additional Data, Boyle's Law

Table 6: Values of pressure for each volume of gas obtained from the first trial.

TRIAL 1	
Volume (mL)	**Pressure (kPa)**
4.8	226.06
6.8	163.27
8.8	124.05
10.8	100.85
12.8	84.58
14.8	73.27
16.8	64.78
18.8	57.98

Additional Graphs, Boyle's Law



	Data Set	
	Volume (mL)	Pressure (kPa)
1	4.8	226.06
2	6.8	163.27
3	8.8	124.05
4	10.8	100.85
5	12.8	84.58
6	14.8	73.27
7	16.8	64.78
8	18.8	57.98

Linear Fit for: Data Set | Pressure
y = mx+b
m (Slope): -10.94 kPa/mL
b (Y-Intercept): 240.9 kPa
Correlation: -0.9287
RMSE: 23.12 kPa

Figure 2. Plotting the first trial's values for the independent variable, volume, and the dependent variable, pressure, depicts the inversely proportional relationship, as shown by the line of best fit.

Sample Calculations, Boyle's Law

Calculating the constant for each volume

$pV = c$

→ 126.33 kPa(0.0088 L) = 1.111704

Average Boyle’s Law Constant

(1.085088+1.1135+1.111704+1.099008+1.0912+1.088096+1.088304)/7 = 1.0967

Assessment Criteria Sheet