Introduction:

Robert Boyle led an experiment which was able to determine a relationship between the volume and pressure of a gas. He had a U-shaped tube, with an open end on one side to which he poured mercury and observed that this was adding pressure to the trapped air, which also caused the volume of the gas to decrease. Boyle then concluded that when the temperature and amount of gas are kept constant, if you increase the pressure of the gas, the volume will decrease by the same amount. This is called Boyle's law and proves that $V \propto \frac{1}{p}$ (Clancy, C., Doram, T., Heimbecker, B. et al., 511).

Amadeo Avogadro proposed a hypothesis called Avogadro's law, which states that for ideal gases at constant temperature and pressure, that equal volumes contain the same number of molecules, regardless of their mass. Avogadro's law can be represented as $n \propto V$, meaning that an increase in the amount in moles of a gas at constant pressure and temperature will result in a greater volume (Clancy, C., Doram, T., Heimbecker, B. et al., 543-545).

John Dalton discovered the law of partial pressures when he observed that the pressure of dry air and water vapour mixed together is just the addition of both pressures. He realized that each gas has a partial pressure that is equal to the pressure the gas would exert by itself. This law can be represented by $p_{total} = p_1 + p_2 + p_3 + \ldots$ As well, the moles of each gas in a mixture can be represented as a mole fraction, X, which can be found by dividing the moles of a gas by the total amount of moles in a mixture. This can be related to partial pressure using this formula, $p_A = X_A p_{total}$ (Silberberg, M., Venkateswaran, R., Lavieri, S., et al., 167-168).

An ideal gas is a gas in which its particles are in constant motion and are considered to have no mass and no attractive or repulsive forces acting on each other. An ideal gas can be used to find the relationship between volume, pressure, temperature, and amount of moles using the ideal gas law. The ideal gas law states that pV = nRT, where R is the universal gas constant with a value of $0.08314 \frac{bar \cdot L}{mol \cdot K}$ or (Clancy, C., Doram, T., Heimbecker, B. et al., 505, 551).

Observations:

During the Boyle's law experiment, it was observed that at a constant temperature and amount of moles, when the pressure in the balloon was increased, the volume or the size of the balloon was noticeably reducing. As well, when the pressure of the balloon was raised, the volume of the balloon kept increasing until the balloon eventually popped. This observation was the same for every gas used in the experiment.

While performing the Avogadro's law experiment, it was observed that when the gas in the balloon was at a constant temperature and pressure, and the number of moles of gas in the balloon was increased, the volume of the balloon increased as well. When the moles of gas in the balloon was decreased, the size of the balloon decreased until it looked like the balloon was empty. Both gases used had the same observations.

In the Dalton's law of partial pressures experiment, when the three gases of different numbers of moles were added together in the same balloon, it was observed that the addition of the pressure of each individual gas measured out to be the total pressure in the balloon.

During the ideal vs. non-ideal gases procedure, it was observed that when the temperature was set to a low value and when the pressure was set to a high value, the volume in the balloon was very low compared to the other settings and the balloon size was small. On the other hand, when the temperature was high and the pressure was low, the volume in the balloon was very high, and the balloon had a very large size. When both the temperature and pressure were either both low or both high, the size of the balloon was in between the largest and smallest sizes.

Data:

Table 1. Boyle's Law Experiment Data and Results for Ideal Gas 1, MM = 4 g/mol

Volume (L)	**Pressure (bar)**	***pV***
7.431	1.000	7.431
3.716	2.000	7.432
2.477	3.000	7.431
1.858	4.000	7.432
1.487	5.000	7.435
1.239	6.000	7.434
1.062	7.000	7.434
0.9291	8.000	7.433
0.8259	9.000	7.433
0.7433	10.000	7.433

Table 2. Boyle's Law Experiment Data and Results for Ideal Gas 8, MM = 222 g/mol

Volume (L)	**Pressure (bar)**	**pV**
7.431	1.000	7.431
3.716	2.000	7.432
2.477	3.000	7.431
1.858	4.000	7.432
1.487	5.000	7.435
1.239	6.000	7.434
1.062	7.000	7.434
0.9291	8.000	7.433
0.8259	9.000	7.433

0.7433	10.000	7.433

Table 3. Boyle's Law Experiment Data and Results for Non-Ideal Gas N_2

Volume (L)	Pressure (bar)	pV
7.430	1.000	7.430
3.715	2.000	7.430
2.476	3.000	7.428
1.857	4.000	7.428
1.485	5.000	7.425
1.237	6.000	7.422
1.060	7.000	7.420
0.9277	8.000	7.422
0.8245	9.000	7.421
0.7420	10.000	7.420

Table 9. Avogadros's Law Experiment Data For Ideal Gas 1, MM = 4 g/mol

Moles (mol)	Volume (L)
0.100	1.223
0.200	2.445
0.300	3.668
0.400	4.891
0.500	6.113

Table 10. Avogadro's Law Experiment Data For Non-Ideal Gas CO_2

Moles (mol)	Volume (L)
0.100	1.226
0.200	2.452
0.300	3.678

0.400	4.904
0.500	6.131

Table 11. Dalton’s Law Experiment Data for Ideal Gases 2, 3, 4

Ideal Gas Number	**Moles (n_i)**	**Partial Pressure (P_i)**	**Total Pressure (P_T)**
2	0.093 mol	0.367 bar	1.0133
3	0.081 mol	0.319 bar	
4	0.083 mol	0.327 bar	

Table 12. Dalton’s Law Experiment Data for Ideal Gases 5, 6 7

Ideal Gas Number	**X_A**	**Partial Pressure (P_i)**	**Total Pressure (P_T)**
5	0.310	0.310 bar	1.000 bar
6	0.396	0.396 bar	
7	0.294	0.294 bar	

Table 13. Results for Ideal vs. Non-Ideal Gases Procedure

Gas	**V (L)**	**P (bar)**	**T (K)**	**n (mol)**
Ideal, low T, low p	0.1039	1.200	15.00	0.100
Ideal, low T, high p	0.009594	13.00	15.00	0.100
Ideal, high T, low p	6.236	1.200	900.00	0.100
Ideal, high T, high p	0.5756	13.00	900.00	0.100
CH_4, low T, low p	1.093	1.200	160.00	0.100
CH_4, low T, high p	0.08325	13.00	160.00	0.100
CH_4, high T, low p	2.422	1.200	350.00	0.100
CH_4, high T, high p	0.2212	13.00	350.00	0.100
CO_2, low T, low p	2.066	1.200	300.00	0.100
CO_2, low T, high p	0.179	13.00	300.00	0.100
CO_2, high T, low p	6.237	1.200	900.00	0.100

CO_2, high T, high p	0.5768	13.00	900.00	0.100

Bonus Data:

Table 4. Boyle's Law Experiment Data and Results for Non-Ideal Gas CO_2

Volume (L)	**Pressure (bar)**	**pV**
7.394	1.000	7.394
3.678	2.000	7.356
2.440	3.000	7.320
1.820	4.000	7.280
1.448	5.000	7.240
1.201	6.000	7.206
1.024	7.000	7.168
0.8907	8.000	7.126
0.7873	9.000	7.086
0.7046	10.000	7.046

Table 5. Boyle's Law Experiment Data and Results for Non-Ideal Gas CH_4

Volume (L)	**Pressure (bar)**	**pV**
7.418	1.000	7.418
3.703	2.000	7.406
2.465	3.000	7.395
1.845	4.000	7.380
1.474	5.000	7.370
1.226	6.000	7.356

1.049	7.000	7.343
0.9163	8.000	7.330
0.8130	9.000	7.317
0.7305	10.000	7.305

Table 6. Boyle's Law Experiment Data and Results for Non-Ideal Gas NH$_3$

Volume (L)	**Pressure (bar)**	**pV**
7.354	1.000	7.354
3.638	2.000	7.276
2.398	3.000	7.194
1.778	4.000	7.112
1.406	5.000	7.030
1.157	6.000	6.942
0.9787	7.000	6.851
0.8448	8.000	6.758
0.7403	9.000	6.663
0.6575	10.000	6.575

Table 7. Boyle's Law Experiment Data and Results for Non-Ideal Gas He

Volume (L)	**Pressure (bar)**	**pV**
7.435	1.000	7.435
3.720	2.000	7.440
2.481	3.000	7.443
1.862	4.000	7.448
1.490	5.000	7.450
1.242	6.000	7.452
1.065	7.000	7.455

0.9327	8.000	7.462
0.8294	9.000	7.465
0.7469	10.000	7.469

Table 8. Boyle's Law Experiment Data and Results for Van der Waals Gas N_2

Volume (L)	**Pressure (bar)**	**pV**
7.400	1.000	7.400
3.685	2.000	7.370
2.446	3.000	7.338
1.826	4.000	7.304
1.455	5.000	7.275
1.207	6.000	7.242
1.030	7.000	7.210
0.8969	8.000	7.175
0.7936	9.000	7.1424
0.7108	10.000	7.108

Graphs:

Figure 1

Relationship between Pressure and Volume for Ideal Gas 1, MM= 4 g/mol

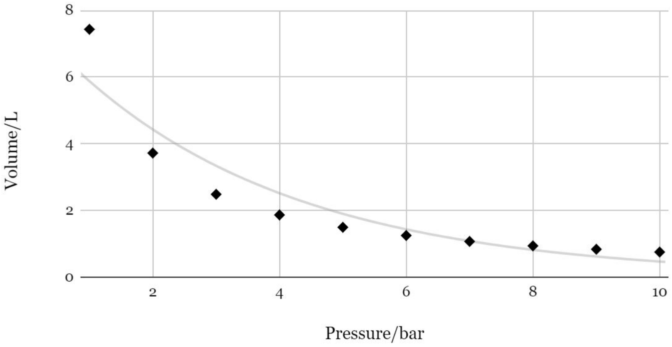


Figure 2

Relationship between Pressure and Volume for Ideal Gas 8, MM = 222g/...

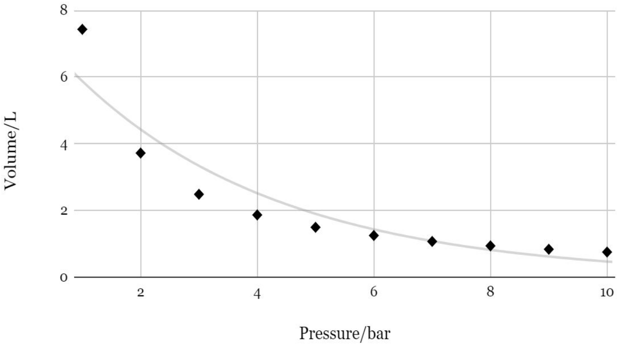


Figure 3

Relationship between Pressure and Volume for Non-Ideal Gas N2

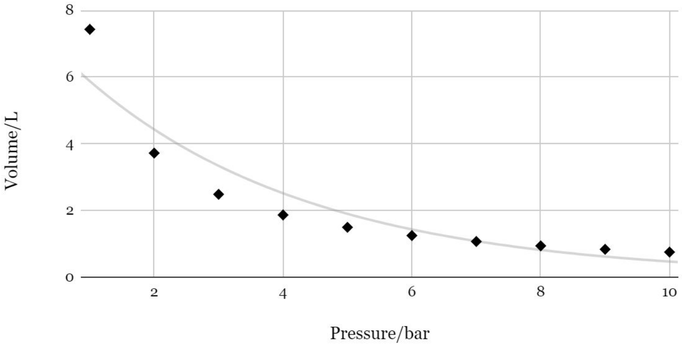


Figure 4

Relationship between Pressure and Volume for Non-Ideal Gas CO2

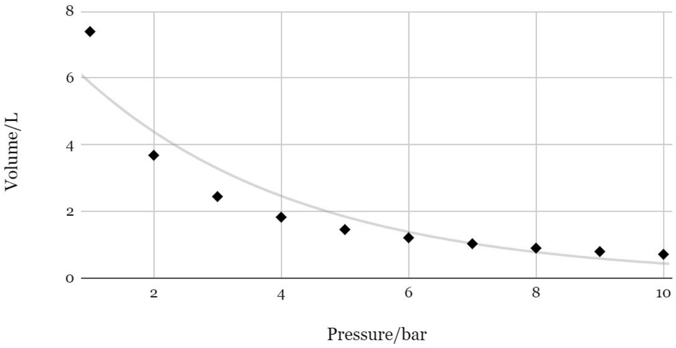


Figure 5

Relationship between Pressure and Volume for Non-Ideal Gas CH4

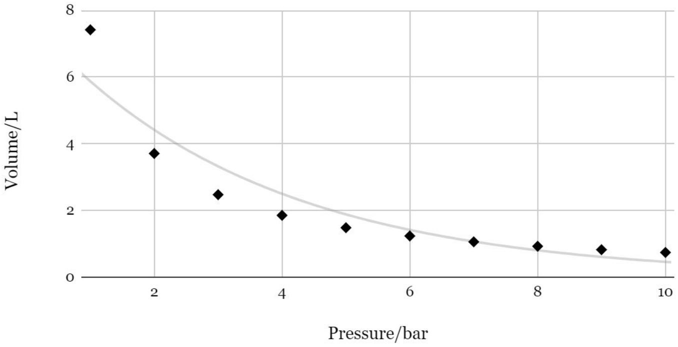


Figure 6

Relationship between Pressure and Volume for Non-Ideal Gas NH3

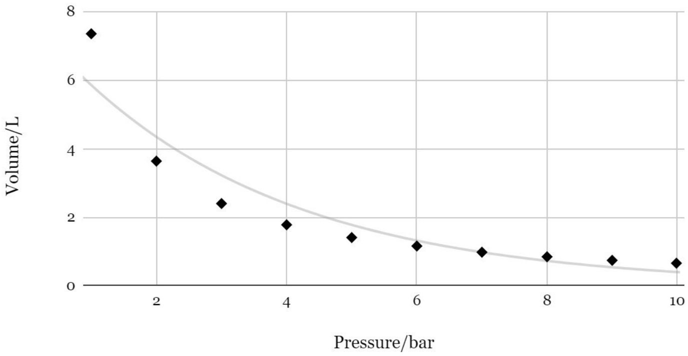


Figure 7

Relationship between Pressure and Volume for Non-Ideal Gas He

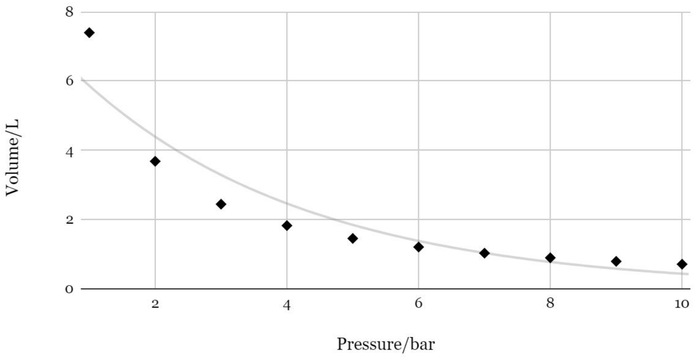


Figure 8

Relationship between Pressure and Volume for Van der Waals Gas N2

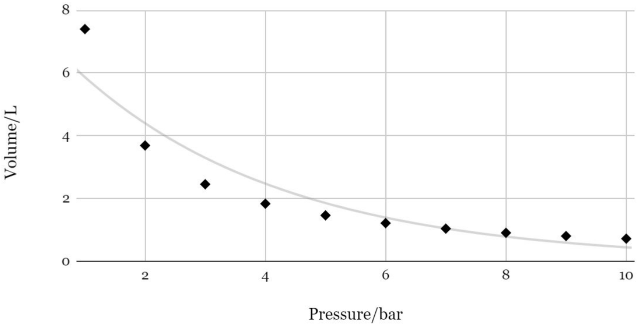


Figure 9

Relationship between Pressure and pV Constant for Non-Ideal Gas NH3

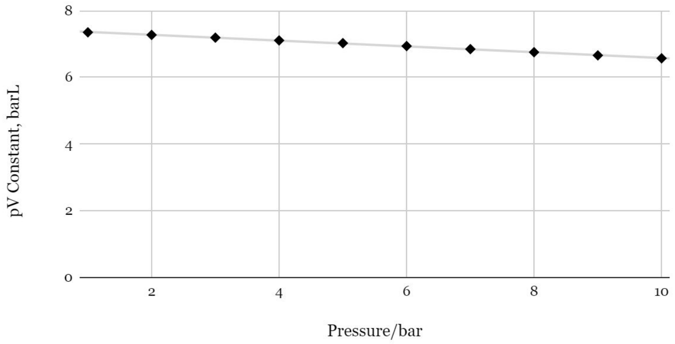


Figure 10

Relationship between Pressure and pV Constant for Ideal Gas 1 and Non-Ideal Gas He

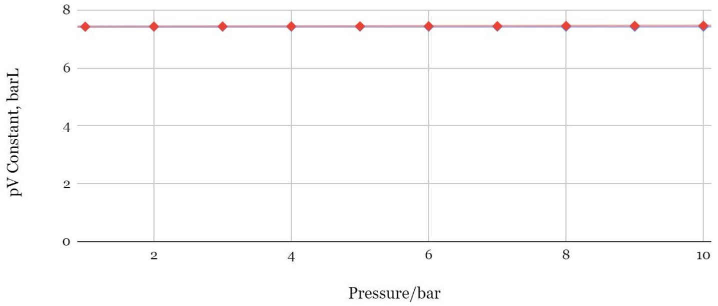


Figure 11

Relationship between Number of Moles and Volume for Ideal Gas 1

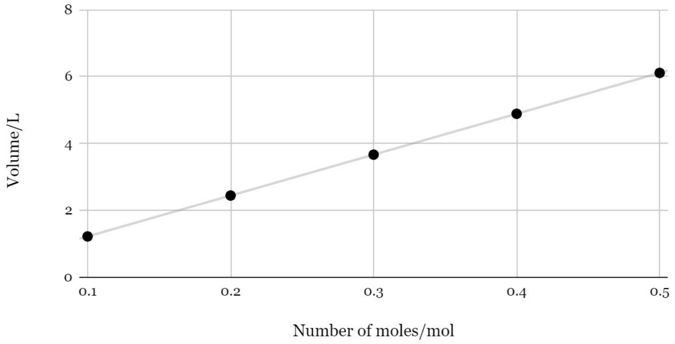
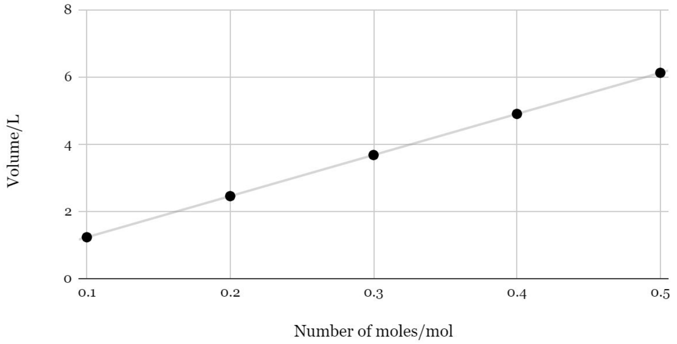


Figure 12

Relationship between Number of Moles and Volume for Non-Ideal Gas CO2



Calculations:

Sample Calculation 1. Calculating pV in the Boyle's Law Experiment for Non-Ideal Gas N_2

$$\begin{aligned} pV &= (p)(V) \\ &= (2.000 \text{ bar})(3.678 \text{ L}) \\ &= 7.356 \text{ bar} \bullet \text{L} \end{aligned}$$

Sample Calculation 2. Calculating the partial pressure, p_A for Ideal Gas 2 in Dalton's Law Experiment

$$\begin{aligned} p_A &= X_A p_{total} \\ &= \left(\frac{0.093\ mol}{0.257\ mol}\right)(1.0133 \text{ bar}) \\ &= 0.37 \text{ bar} \end{aligned}$$

Sample Calculation 3. Calculating R in the Ideal vs. Non-Ideal Gases Procedure for CH_4 at low T and p

$$\begin{aligned} R &= \frac{pV}{nT} \\ &= \frac{(1.200\ bar)(1.093\ L)}{(0.100\ mol)(160.00\ K)} \\ &= 0.08198\ \frac{bar \bullet L}{mol \bullet K} \end{aligned}$$

Sample Calculation 4. Calculating the Percentage Error for R, for an Ideal Gas at low T and p

$$\begin{aligned} \text{Percentage Error} &= \frac{|experimental\ value - accepted\ value|}{accepted\ value}\ x\ 100\% \\ &= \frac{|0.08312 - 0.0831451|}{0.0831451}\ x\ 100\% \end{aligned}$$

= 0.03019%

Results:

Table 14. Calculated Results for Ideal Gases 2, 3, and 4 in Dalton's Law Experiment

Ideal Gas Number	X_A	P_A (calculated)	P_A (measured)
2	0.362	0.367 bar	0.367 bar
3	0.315	0.319 bar	0.319 bar
4	0.323	0.327 bar	0.327 bar

Table 15. Calculated Results for Ideal Gases 5, 6, and 7 in Dalton's Law Experiment

Ideal Gas Number	X_A	P_A (calculated)	P_A (measured)
5	0.310	0.310 bar	0.310 bar
6	0.396	0.396 bar	0.396 bar
7	0.294	0.294 bar	0.294 bar

Table 16. R Constant Calculations for Ideal vs. Non Ideal Gases Procedure

Gas	Calculated R (L•bar•K/mol)	R % Error
Ideal, low T, low p	0.08312	0.03019
Ideal, low T, high p	0.08315	0.00589
Ideal, high T, low p	0.08315	0.00589
Ideal, high T, high p	0.08314	0.00613
CH_4, low T, low p	0.08198	1.4013
CH_4, low T, high p	0.06764	18.648
CH_4, high T, low p	0.08304	0.12641
CH_4, high T, high p	0.08216	1.1848
CO_2, low T, low p	0.08264	0.60749
CO_2, low T, high p	0.07757	6.7053

CO_2, high T, low p	0.08316	0.01792
CO_2, high T, high p	0.08332	0.21036

Discussion:

The goal of the Boyle's law procedure was to observe and analyze the relationship between pressure and volume in a gas at constant temperature and amount in moles. Before this experiment, it was expected that the relationship between the pressure and the volume would show an inversely proportional relationship, meaning that if the pressure were to increase by a certain factor, then the volume would decrease by the same factor, and vice versa. The results of this experiment were able to prove the initial hypothesis, as the pressure was increasing, the volume was decreasing, which in turn adds validity to Boyle’s law. The graphs in figures 1 through 8 show the relationship between pressure and volume as well as the trend line. From these graphs, it is evident that the relationship of pressure as a function of volume shows a decreasing exponential curve, showing that as the pressure increases, the volume decreases at the same rate. This is what was expected at the beginning of the experiment and it further validates Boyle’s law. While the results showed the same trend for every gas used in this experiment, what is interesting to note is that the pV constant values tend to fluctuate when observing the data from the ideal gases versus the non-ideal gases. This can be observed looking at figures 1 through 8. The data for the two ideal gases used in this experiment are shown in tables 1 and 2; when looking at the molar masses given for both ideal gases 1 and 8, it can be determined that the gases are helium and radon, respectively. By observing these two gases, it is evident that they have very similar values, which tend to stay constant as the pressure increases; this is represented graphically in figure 10. However, when analyzing the data from tables 3 to 8, it can be seen that the pV values are all different for each gas. As well, the values are not constant; they tend to decrease by a small amount as the pressure increases. This is represented graphically in figure 9. The reason why this is interesting is because the gases to which the pV values stay constant are both ideal gases, and the gases to which the pV values decrease are all non-ideal gases. A potential reason for this is that, as explained in the lab instructions, ideal gases are said to have no mass and to have no attractive or repulsive forces between particles. However because non-ideal gases do actually have volume and exert forces on each other. There is a slight deviation in the values as the pressure increases, because the gas particles get closer together and exert a greater force on each other, which moves the gas further from the truly ideal gas concept, and therefore, having different values than those of ideal gases.

The Avogadro’s law experiment was about proving the relationship between the amount in moles and the volume of a gas. It was hypothesized before the experiment that the volume and the amount in moles have a directly proportional relationship, as stated by Avogadro’s law, which means that as the volume increases, the amount in moles will increase by the same factor. The data obtained from the experiment is portrayed graphically in figures 11 and 12. It is evident that the graphs show an increasing linear relationship between volume and amount, showing that as the volume increases, the amount increases at the same rate, which is what should occur according to Avogadro’s law. It is also important to notice that the data obtained from both an ideal gas (as shown in table 9) and a non-ideal gas (as shown in table 10) are relatively similar. For both gases, the amount in moles as the volume increases are very close to the same value. This means that the values of the amount in moles will most likely be the same for any gas used in a similar Avogadro's law experiment as the volume increases or decreases. So, the type of gas does not actually matter in this experiment, all gases will show more or less the same relationship and values. Because the values are similar between non-ideal gases and ideal gases, this means that a proportionality constant (k), with units L/mol, would also be the same for both types of gases. To find this constant, the relationship between volume and amount has to be considered. Because the number of moles of a gas and its volume have a directly proportional relationship, it can be written as $n \propto V$. To determine the constant in L/mol, the volume has to be divided by the number of moles. Unlike the constant for Boyle's law, these two values cannot be multiplied together to find k, because that would imply that one value would have to increase and the other would have to decrease so that the equation would remain mathematically correct. If $k = \frac{V}{n}$, then it implies that both values would have to increase or decrease by the same amount for k to be true.

The goal of the Dalton's law of partial pressures experiment was to determine if the partial pressures obtained during the experiment is the same as the partial pressures obtained from a calculation relating the amount of moles and the pressure of a gas. This mathematical formula states that the amount in moles of one gas in a mixture (mole fraction, X) divided by the total amount in moles is equal to the partial pressure of that gas in the mixture divided by the total pressure of the mixture, $\frac{P_A}{P_T} = X$. The total pressure can be determined by the addition of the partial pressure of each gas, $P_T = P_1 + P_2 + P_3 + \ldots$ After the data had been recorded, the mole fraction of each gas (three per experiment) was used to determine the partial pressure and compare it to the original result obtained from the experiment. The partial pressure calculated for each gas in the mixture was the same exact value as the value originally given, as shown in tables 14 and 15. The fact that the values of the experimental and calculated values of the partial pressures for each gas are identical validates Dalton's theory of partial pressure. This proves that the partial pressure of each gas could have been determined by only knowing the amount in moles as well as the partial pressure of each gas in the mixture. This calculation holds true for any mixture of gases; there can be any number of different gases in a mixture with totally different amounts in moles, and the partial pressure of each gas could still be determined. As shown in tables 14 and 15, the partial pressure of a mixture of gases can be determined no matter what gas is used; however, the result of the partial pressure of a gas does depend on the mole fraction of the gas in the mixture. In other words, if all the amounts in moles of the gases are different, then each partial pressure of the gases will also be different.

The goal of the ideal versus non-ideal gases procedure was to determine the relationship between temperature (K), volume (L), and pressure (bar) of a gas at a constant amount (mol) of different gases when the values of the temperature and pressure were fluctuating at high and low amount relative to each gas, using the ideal gas formula $pV = nRT$, where R is a constant in Lbar/molK. From the data table 13, by analyzing the values, it is evident that there is a clear relationship between temperature, pressure, and volume. For example, when the temperature was set to a high value, and the pressure was set to a low value, the volume was very high compared to the other values. As well, when the temperature was low, and the pressure was high, the value of the volume was very low compared to the others. From this data table, the value of the R constant and the percentage error were calculated for each different setting for the three gases; the results are shown in table 16. In this results table, it can be observed that the ideal gas has the same R value at each setting of temperature and pressure. There are also clear deviations in certain values which are much smaller than the rest of the values. However, these deviations only occur for the two non-ideal gases used, not the ideal gas used. As well, there is only one value out of the four for each gas that shows a deviation, and it is the value in which the pressure is at a high value. This observation can be explained by the fact that an ideal gas is said to have no volume and no forces acting on neighbouring particles; this means that at a very high pressure, there would be skewed results. On the other hand, the two other gases used are non-ideal and do have a volume and exert a greater force as they get closer together, which explains why there is a deviation in the value for the R constant as the pressure increases.

Connections:

There is an article written by Gurmukh Chandan and Marco Cascella, that explains how different gas laws, including Boyle's law, Avogadro's law, Dalton's law, the ideal gas law, and many others, are related to a clinical setting and how they can be used to further understand bodily physicochemical processes.

Boyle's law can be used to describe the effects of altitude on the gases in the body, and to calculate the intra-thoracic gas volume by body plethysmography. As altitude increases, the atmospheric pressure decreases, which means that the volume would increase in enclosed spaces, according to Boyle's law. The use of Boyle's law can be useful in cases, for example, when a patient with a pneumothorax (collapsed lung) is being airlifted to a hospital. Suppose that a patient has a pneumothorax with a volume of 1500 mL at sea level, which has an atmospheric pressure of 101.3 kPa. At an altitude of 1km, which has an atmospheric pressure of 90 kPa, it can be calculated using Boyle's law that the volume of the pneumothorax will now have a volume of 1688 mL. (Chandan, G., Gascella M.)

They also related Dalton's law of partial pressures with someone climbing Mount Everest, as well as the alveolar gas equation, which is used to calculate the partial pressure of oxygen in the alveolus. For example, at sea level, the partial pressure of oxygen is 21% (21 kPa), and at the top of Mount Everest, the atmospheric pressure is 33.7 kPa. Using Dalton law, it can be calculated that the partial pressure of oxygen at this level is only 7 kPa, which is very small, and can lead to an oxygen-hemoglobin saturation of less than 80%, which is obviously dangerous for the human body (Chandan, G., Gascella M.).

Conclusion:

From the Boyle's law experiment, it was determined that the pressure and the volume of a gas have an inversely proportional relationship, meaning that as the pressure increases, the volume will decrease by the same factor, and vice versa. It was determined from the Avogadro's law experiment that there is a directly proportional relationship between the volume and the amount in moles of a gas, meaning that as the amount of moles increases, the volume also increases by the same amount. From the Dalton's law experiment, it was concluded that the partial pressure of a gas in a mixture can be determined (with a known mole fraction and total pressure) mathematically, using the formula $\frac{P_A}{P_T} = X$. From the ideal versus non-ideal gases procedure, it was determined that there is a deviation in the R constant value for a non-ideal gas at a high pressure, and that the R constant stays the same for an ideal gas.

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