The Ideal Gas Law encompasses all of the other Gas Laws.

\[ P \, U = n \, R \, T \]

\( n, T \) constant

\[ \frac{P}{n} = \frac{n \, R}{V} = \text{constant} \]

Boyle's

\( n, V \) constant

\[ \frac{P}{T} = \frac{RT}{n} = \text{constant} \]

Gay-Lussac's Law

\( P, T \) constant

\[ \frac{V}{n} = \frac{RT}{P} = \text{constant} \]

Avogadro's Law

\( n, P \) constant

\[ \frac{V}{T} = \frac{n \, R}{P} = \text{constant} \]

Charles' Law
STP

standard temperature and pressure
273.15K
(0.00°C)

1,000 mL gas occupies 22.4 L

\[ R = \frac{PV}{nT} = \frac{(1 \text{ atm})(22.414 \text{ L})}{(1 \text{ mol})(273.15 \text{ K})} \]

\[ R = 0.08206 \frac{\text{L atm}}{\text{mol K}} \]
Case 1

\[ P_1 V_1 = n_1 R T_1 \]

\[ P_1 V_1 = n_1 R T_1 \]

\[ P_2 V_2 = n_2 R T_2 \]

Case 2

\[ P_2 V_2 = n_2 R T_2 \]

\[ V_2 = V_1 \times \text{factor} \times \text{factor} \times \text{factor} \]

Changes "open" system

"Closed" system - no change in \( n \)
OFB, Ch. 5, p. 233, # 36

Suppose 63.6 L of iodine heptafluoride is synthesized at 300 °C and 0.459 atm. Calculate the volume occupied by the gas when it is heated to 400 °C while the pressure on the gas is simultaneously increased to 0.980 atm.

\[ V_2 = 63.6 \text{ L} \left( \frac{673 \text{ K}}{573 \text{ K}} \right) \left( \frac{0.459 \text{ atm}}{0.980 \text{ atm}} \right) \]

\[ = 34.99 \text{ L} = 35.0 \text{ L} \]
What mass in (in grams) of gaseous neon exerts a pressure of 2.00 atm in a 10.0 L tank? 

\[ PV = nRT \]

\[ (2.00 \text{ atm})(10.0 \text{ L}) = n \left( \frac{0.08206 \text{ L atm}}{\text{mol K}} \right)(273.15 \text{ K}) \]

\[ n = \frac{0.8928 \text{ mol}}{8} \]

\[ 0.8928 \text{ mol Ne} \left( \frac{20.18 \text{ g Ne}}{\text{mol Ne}} \right) = 18.015 \text{ g Ne} \]

\[ 18.09 \text{ g} \]

\[ 8 \]
Because gas particles are assumed not to interact, each gas in a mixture has the entire volume available to it. This leads to Dalton's Law of Partial Pressures.

\[ P_A V = n_A RT \]

\[ \frac{P_A}{P_{\text{tot}}} = \frac{n_A}{n_{\text{tot}}} \]

The mol fraction is written as:

\[ X_A = \frac{n_A}{n_{\text{tot}}} \]
\[ P_{\text{tot}} V = n_{\text{tot}} R T \]
\[ P_A V = n_A R T \]
\[ P_B V = n_B R T \]

\[ P_A = \text{partial pressure of A} \]

[Diagram with labels A, B, AA, AB, BA, BB]
A gas mixture contains 4.5 mol Br₂ and 33.1 mol F₂. The mixture is heated above 150 °C and the reaction shown below takes place.

\[ \text{Br}_2 (g) + 5 \text{F}_2 (g) \rightarrow 2 \text{BrF}_5 (g) \]

The reaction is stopped when 2.2 mol of bromine pentafluoride is present. What is the mol fraction of bromine present at that point? What is the partial pressure of bromine?

\[ n_{\text{Br}_2} = n_{\text{F}_2} = n_{\text{BrF}_5} \]

\[ \begin{array}{c}
4.5 \text{ mol} \\
33.1 \text{ mol} \\
0 \\
\end{array} \]

\[ \text{Net} = n_{\text{Br}_2} + n_{\text{F}_2} + n_{\text{BrF}_5} \]

\[ 3.4 \text{ mol} + 27.6 \text{ mol} + 2.2 \text{ mol} \]

\[ n_{\text{Br}_2} = 33.2 \text{ mol} \]
\[ X = \frac{7 \text{ Br}_2}{n_{\text{tot}}} = \frac{3.9 \text{ ml}}{33.2 \text{ ml}} = 0.101 \]

\[ P_{\text{Br}_2} = X \times P_{\text{tot}} \]

\[ P_{\text{Br}_2} = 0.10 \times P_{\text{tot}} \]
A sample of a gaseous binary compound of boron and chlorine weighing 2.842 g occupies 0.153 L at STP. This sample is decomposed to give solid boron and gaseous chlorine. At STP, the chlorine gas occupies 0.688 L. Determine the molecular formula of the compound.

We didn't get to this in class on Sept. 22 and it won't be done in lecture on the 24th. However, I will include a solution with Friday's notes.