The 8 Gas Laws

1.) **Ideal Gas Law**

- **PV= nRT**
- (Pressure)(Volume) = (number of Moles)(R gas constant)(Temperature)
- Sample Problem:

The helium in a 1.5 L flask at 25 °C exerts a pressure of 425 mmHg. How many moles of helium are there in the flask?

$$P = (425\ \text{mmHg})\left(\frac{1\ \text{atm}}{760\ \text{mmHg}}\right) = \mathbf{0.559\ atm}$$

V = **1.5 L**

n = ?

$$R = 0.082057\ \frac{(\text{L} \cdot \text{atm})}{(\text{mol} \cdot \text{Kelvin})}$$

T = 25 °C + 273= **298 Kelvin**

Warning: The Ideal Law requires you to use Kelvin for units of temperature!!!

$$\frac{PV}{RT} = \frac{n\cancel{RT}}{\cancel{RT}} \longrightarrow n = \frac{PV}{RT}$$

$$n = \frac{PV}{RT} = \frac{(0.559\ \text{atm})(1.5\text{L})}{(0.082057\ \text{L} \cdot \text{atm/ mol} \cdot \text{Kelvin})(298\ \text{Kelvin})} = \mathbf{0.034\ moles\ He}$$

2.) **Boyle's Law:**

- If **V** ↑ then **P** ↓ , and if **V** ↓ then **P** ↑
- $P_1V_1 = P_2V_2$
- Sample Problem:

A sample of gaseous CO_2 has a pressure of 55mmHg in a 125 mL flask. If this sample is transferred into a 650 ml flask, with the same temperature as before, what is the expected pressure of the gas?

P_1 = **55 mmHg** V_1 = **125 ml** P_2 = **?** V_2 = **650 ml**

$$\frac{P_1V_1}{V_2} = \frac{P_2\cancel{V_2}}{\cancel{V_2}} \longrightarrow P_2 = \frac{(55\ \text{mmHg})(125\ \text{ml})}{(650\ \text{ml})} = \mathbf{11\ mmHg}$$

Note: You did not have to convert any units for Boyles Law

3.) **Charles' Law**

- If **T** ↑ then **V** ↑ and if **T** ↓ then **V** ↓
- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
- Sample Problem:

If the volume of a sample of CO_2 in a gas-tight syringe is 25.0 ml at room temperature (20 °C), what is the final volume of the gas if you hold the syringe in your hand until the temperature raises to 37°C?

V_1 = **25.0 ml** V_2 = **?**

T_1 = 20°C + 273 = **293 K** T_2 = 37 °C + 273 = **310 K**

$$T_2 \cdot \frac{V_1}{T_1} = \frac{V_2}{\cancel{T_2}}\cancel{T_2} \longrightarrow V_2 = \frac{(310\text{ K})(25.0\text{ml})}{(293\text{ K})} = 26.5\text{ ml}$$

Warning: Charles Law requires you to use Kelvin for units of temperature!!!

4.) **Avagadro's Law**

- If **V** ↑ then **n** ↑ and if **V** ↓ then **n** ↓
- $\frac{V_1}{n_1} = \frac{V_2}{n_2}$
- Sample Problem:

Ammonia can be made directly from the elements according to the equation 1 $N_{2\,(g)}$ + 3 $H_{2\,(g)}$ → 2 $NH_{3\,(g)}$. If you begin with 12 L of H_2 gas what volume of N_2 gas is required for complete reaction?

V_1 = **12 L H_2 gas** V_2 = **? N_2 gas**

n_1 = **3 moles H_2** n_2 = **1 mole N_2**

$$n_2 \cdot \frac{V_1}{n_1} = \frac{V_2}{\cancel{n_2}}\cancel{n_2} \longrightarrow V_2 = \frac{(1\text{ mole }N_2)(12\text{ L }H_2)}{(3\text{ moles }H_2)} = 4\text{ L }N_2\text{ gas}$$

Note: Here you did not have to convert any units. But beware, if you are given grams – you WILL need to convert to MOLES!

5.) **Gay Lusac's Law**

- If **T** ↑ then **P** ↑ and if **T** ↓ then **P** ↓
- $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
- Sample Problem:

  If the pressure of a sample of CO_2 is 458 torr at 10°C, what is the final pressure of the gas if you heat the sample until the temperature is 50°C?

  P_1 = 458 torr **P_2 = ?**

  T_1 = 10°C + 273 = **283 K** T_2 = 50 °C + 273 = **323 K**

  $$T_2 \cdot \frac{P_1}{T_1} = \frac{P_2}{\cancel{T_2}} \cancel{T_2} \longrightarrow P_2 = \frac{(323\ K)(458\ torr)}{(283\ K)} = 523\ torr$$

  Warning: Gay Lusac's Law requires you to use Kelvin for units of temperature!!!

6.) **Combined Gas Law**

- Since R is a constant, the following equations can be rewritten as

  $P_1V_1 = n_1RT_1$ $P_2V_2 = n_2RT_2$

  $\frac{P_1V_1}{n_1T_1} = \boxed{R}$ $\frac{P_2V_2}{n_2T_2} = \boxed{R}$

  Because both equations equal R, they also equal each other!

- $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$

- Sample Problem:

  Suppose you have a sample of a gas in a 12.5 ml container. The pressure of the gas is 685 mmHg at room temperature (22 °C). If you transfer the gas into a 25 ml container and heat it so that the final temperature of the container is 40°C, what is the final pressure?

P_1 = 685 torr

T_1 = 22°C + 273 = **295 K**

V_1 = **12.5 ml**

n_1 = **constant**

P_2 = ?

T_2 = 40 °C + 273 = **313 K**

V_2 = **25 ml**

n_2 = **constant**

$$\frac{n_2T_2}{V_2} \cdot \frac{P_1V_1}{n_1T_1} = \frac{P_2\cancel{V_2}}{\cancel{n_2T_2}} \cdot \frac{\cancel{n_2T_2}}{\cancel{V_2}}$$

where $n_1 = n_2$

$$\frac{T_2P_1V_1}{T_1V_2} = P_2 = \frac{(313\text{ K})(685\text{ torr})(12.5\text{ ml})}{(295\text{ K})(25\text{ ml})} = 363\text{ torr}$$

WARNING!!
Temperature must always be changed into units of Kelvin regardless of which gas law you are using!

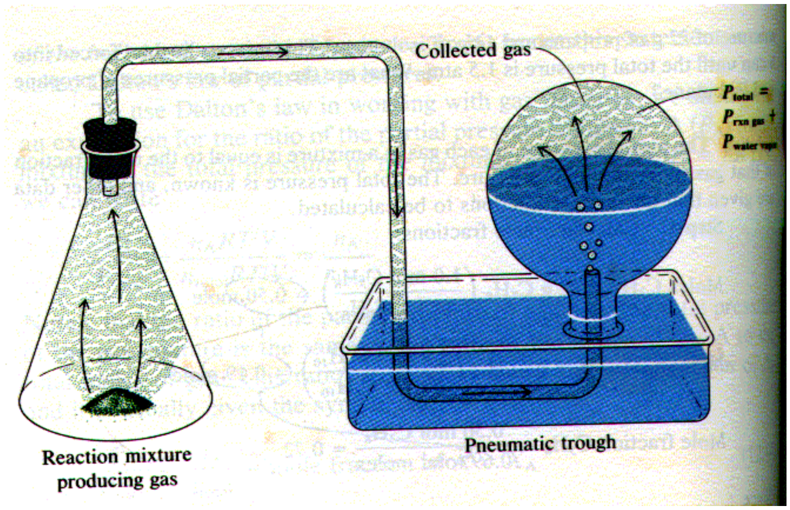
7.) **Dalton's Law**

- The pressure of a mixture of gases is the sum of the pressures of the different components.
- $P_{total} = P_1 + P_2 + P_3$ etc.
- Since the amount of pressure is directly related to the number of moles, Dalton's law can also be shown as

$$\frac{P_1}{P_{total}} = \frac{n_1}{n_{total}} = \text{mole fraction of Component \#1} = X_1$$

- Sample Problem:

  Small quantities of H_2 gas can be prepared in the lab by the following reaction: $Fe_{(s)} + 2\,HCl_{(aq)} \rightarrow FeCl_{2\,(aq)} + H_{2\,(g)}$. Assume you carried out this experiment to completion and collected 0.500 L of H_2 gas as shown below. The temperature of the gas mixture was 26°C, and the total pressure of the gas in the flask was 745 mmHg. How many total moles of gas (hydrogen + water vapor) were in the flask, what is the partial pressure of the hygrogen, and how many moles of hydrogen did you prepare?



Step 1: Use Ideal Gas Law to determine total moles in flask.

$\mathbf{P} = (745 \text{ mmHg})\left(\dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) = \mathbf{0.980 \text{ atm}}$

$\mathbf{V} = \mathbf{0.500 \text{ L}}$

$\mathbf{n} = ?$

$\mathbf{R} = \mathbf{0.082057} \left(\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{Kelvin}}\right)$

$\mathbf{T} = 26\ °\text{C} + 273 = \mathbf{299 \text{ Kelvin}}$

Warning: The Ideal Law requires you to use Kelvin for units of temperature!!!

$$\frac{PV}{RT} = \frac{n\cancel{RT}}{\cancel{RT}} \longrightarrow n = \frac{PV}{RT}$$

$$n = \frac{PV}{RT} = \frac{(0.980 \text{ atm})(0.500\text{L})}{(0.082057 \text{ L} \cdot \text{atm/ mol} \cdot \text{Kelvin})(299 \text{ Kelvin})} = 0.0200 \text{ total moles}$$

Step 2: Use Dalton's Law of Partial Pressures to determine moles of H_2.

$\mathbf{P_{total} = P_{hydrogen} + P_{water\ vapor}}$ (Rearrange the equation to solve for H_2.)

$\mathbf{P_{hydrogen} = P_{total} - P_{water\ vapor}}$

Note: The partial pressure of water vapor over liquid water is always 25.2 mmHg

$\mathbf{P_{hydrogen} = (745 \text{ mmHg}) - (25.2 \text{ mmHg})}$

$\mathbf{P_{hydrogen} = 719.8 \text{ mmHg}}$

Compare the pressure ratio to the mole ratio then cross-multiply & divide method to find the moles of H_2:

$$\frac{P_H}{P_{total}} = \frac{n_H}{n_{total}}$$

$$\frac{P_H}{P_{total}} = \frac{719.8 \text{ mmHg}}{745.0 \text{ mmHg}}$$

$n_H = ?$
$n_{total} = 0.0200$ moles

$n_{total} = 0.0193$ mol H_2

8.) **Graham's Law**

- At constant temperature and pressure, the rates of effusion of two gases are inversely proportional to the square roots of their molar masses.

- $$\frac{\text{Rate of Effusion of Gas \#1}}{\text{Rate of Effusion of Gas \#2}} = \sqrt{\frac{\text{Molar Mass of Gas \#2}}{\text{Molar Mass of Gas \#1}}}$$

- The rate of effusion is a count of the number of molecules moving from one location to another location in a given amount of time.
- Sample Problem:

  A sample of tetrafluoroethylene, C_2F_4, effuses through a barrier at the rate of 4.6×10^{-6} moles per hour. An unknown gas, consisting of nitrogen, oxygen, and fluorine atoms, effuses at the rate of 6.5×10^{-6} moles per hour under the same conditions. What is the molar mass of the unknown gas? Suggest a formula for the unknown gas.

Rate of Effusion of Gas #1 = **4.6×10^{-6} moles/ hour**

Rate of Effusion of Gas #2 = **6.5×10^{-6} moles/ hour**

Molar Mass of Gas #1 = formula weight of C_2F_4

Carbon = 2 x 12 g/mol = 24 g/mol
Flourine = 4 x 19 g/mol = + 76 g/mol
100 g/mol C_2F_4

$$\frac{\text{Rate of Effusion of Gas \#1}}{\text{Rate of Effusion of Gas \#2}} = \sqrt{\frac{\text{Molar Mass of Gas \#2}}{\text{Molar Mass of Gas \#1}}}$$

Rearrange equation to solve for Molar Mass of Gas #2:

$$\frac{(\text{Rate of Effusion of Gas \#1})^2}{(\text{Rate of Effusion of Gas \#2})^2} \times \text{Molar Mass of Gas \#1} = \text{Molar Mass of Gas \#2}$$

$$\frac{(4.6 \times 10^{-6} \text{ moles/ hour})^2}{(6.5 \times 10^{-6} \text{ moles/ hour})^2} \times 100 \text{ g/mol } C_2F_4 = 50 \text{ g/mol Unknown Gas \#2}$$

Suggest a possible formula for Gas #2.

Nitrogen = 1 x 14 g/mol = 14 g/mol
Oxygen = 1 x 16 g/mol = 16 g/mol
Flourine = 1 x 19 g/mol = 19 g/mol
49 g/mol → NOF; nitrosyl fluoride