Chemistry Chapter 9 – Properties of Gases

Warning: Be careful of units when doing proportions. Make sure you're using the correct units for the problem
Another Warning: When dealing with diatomic molecules, don't forget to double to atomic weight

1. **Boyle's Law –** at constant temperature, the **pressure** of a gas is **inversely proportional** with its **volume**
   a. **The Equation:** $Pressure_{Initial}$ x $Volume_{Initial}$ = $Pressure_{Final}$ x Volume 2_{Final}

2. **Charles Law –** At constant pressure, the **temperature in Kelvins** is **proportional** to the **volume**
   a. **The Equation:** $Volume_{Initial}$ / $Temperature_{Initial}$ = $Volume_{Final}$ / $Temperature_{Final}$

3. **Gay-Lussac's Law** – At constant volume, the **pressure** is **proportional** to the **temperature in Kelvins**
   a. **The Equation:** $Pressure_{Initial}$ / $Temperature_{Initial}$ = $Pressure_{Final}$ / $Temperature_{Final}$

4. **The Useful Gas Equation that Combines All Three Laws** – $P_{Initial} * V_{Initial} / T_{Initial} = P_{Final} * V_{Finall} / T_{Final}$

5. **Ideal Gas Law –** PV = nRT
   a. V is in Liters
   b. P is in atm
   c. T is in Kelvins
   d. N is the number of moles of gas
   e. R is the deal gas constant. Using these units, **R=0.0821Latm/mol K**

6. **STP –** Standard temperature and pressure = 273K, 1 atm (760 torr)
7. **Conversions - 1 atm** (at STP) = **760 mm/Hg = 760 Torr** = **101.3kPa** (Kilo Pascal) = **14.7 psi** ($lb/inches^2$)
   a. If a tire gauge measures a tire to be 100 mm/Hg, the actual pressure is 860 mm/Hg
   b. The gauge measures the amount of pressure ABOVE atmospheric pressure
8. The temperature in Kelvins is **proportional** to the Kinetic Energy of the gas
9. **Dalton Law of Partial Pressures** – each gas takes up the same amount of space
   a. The total Pressure of a gas = sum of the pressures of the individual gases
   b. Example: Air, a mixture of mostly O_2 and N_2
      Moles O_2/Total Moles = Pressure O_2/Total Pressure
      Moles N_2/Total Moles = Pressure N_2/Total Pressure
      Total Pressure = $Pressure_{O2}$ + $Pressure_{N2}$
10. **Diffusion –** Gas in a small area spreads over large area
11. **Effusion** – gas in large area escapes through small area
12. At constant temperature, the kinetic energy of each gas is the same.
13. **Graham's Law** – the rate of effusion is inversely proportional to the square root of the density of the gases.
   a. $M_a \times V_a^2 = M_b \times V_b^2$
   b. M - Mass
   c. V – Volume

$$\frac{\textit{Effusion Rate A}}{\textit{Effusion Rate B}} \quad \sqrt{\frac{M_B}{M_A}} =$$

14. The density of a gas is proportional to its molecular weight
15. **Ideal gases –** gasses that have no volume/attractions, Helium is the most "ideal" gas
   a. In reality, no gas is truly ideal
   b. **The Real Gas Law** – $(P + N^2\mathbf{a}/V^2)(V-n\mathbf{b})$ **–** nRT
      i. a – force of attraction
      ii. b – size of the real gas particle
16. Pressure is proportional to the force pushing on it and inversely proportional to the area of force (P=F/A)
   a. If a large force is exerted over a very large area, the force each point feels is very low
   b. If the same force is exerted over an extremely small area, the force each point feels is much larger