Chapter 10: Gases

10.1 Characteristics of Gases

- nonmetallic, low molar mass, simple molecular formula
- vapors – liquid in gaseous state
- volume gas = volume container
- increase pressure, decrease volume
- gases for homogeneous mixtures

10.2 Pressure

- P = F/A

10.2.1 Atmosphere Pressure and the Barometer

- F = ma
- 1 pa = 1 N/M^2
- standard atmospheric pressure
    - 760mm Hg = 1.01325 x 10^5 pa
    - 1 atm = 760 mm Hg = 760 torr = 1.01325 x 10^5 pa

10.2.2 Pressures of Enclosed Gases and Manometers

- $P_{gas} = P_{h1}$ closed end manometer
- $P_{gas} + P_{h2} = P_{atm}$ open end (less than atmospheric pressure)
- $P_{gas} = P_{atm} + P_{h3}$ gas pressure exceeds atmospheric pressure

10.3 The Gas Laws

- four variable to define physical condition (state) of gas
    - T, P, V, and amount (moles, n)

10.3.1 The Pressure – Volume Relationship; Boyle's law

- Boyle's Law – the volume of a fixed quantity of gas maintained at constant temp is inversely proportional to pressure
- V = constant x 1/P or PV = constant

10.3.2 The Temperature – Volume relationship: Charles's law

- Charles's law = volume is directly proportional to absolute temperature at constant pressure
- V = constant x T or V/T = constant

10.3.3 The Quantity – Volume Relationship: Avogadro's Law

- Avogadro's hypothesis – equal volumes of gases at same temperature and pressure have equal number of molecules
- Avogadro's law = volume of gas is directly proportional to number of moles of gas at constant temperature and pressure
    - V = constant x n

10.4 The Ideal – Gas Equation

- ideal gas equation = PV = nRT
- R = gas constant
- STP = 0 degrees Celsius and 1 atm

10.4.1 Relationship Between the Ideal – Gas Equation and the Gas Laws

- PV = nRT = constant or PV = constant
- $P_1V_1 = P_2V_2$
- PV/T = nR = constant so $P_1V_1/T_1 = P_2V_2/T_2$

10.5 Further Applications of the Ideal – Gas Equation

10.5.1 Gas Densities and Molar Mass

- d = PM/RT or M = dRT/P

10.5.2 Volume of Gases in Chemical Reactions

- calculate volume of gases consumed or produced

10.6 Gas Mixtures and Partial Pressures

- Dalton's Law of Partial Pressure – total pressure of mixture of gases = sum of pressure that each would exist if alone
- $P_T = P_1 + P_2 + P_3 + \dots$
- $P_T = n_T \times RT/V$

10.6.1 Partial Pressures and Mole Fractions

- $$\frac{P_1}{P_T} = \frac{\frac{n_1 RT}{V}}{\frac{n_2 RT}{V}} = \frac{n_1}{n_2} = \text{mole fraction gas } 1 = x_1$$

10.6.2 Collecting Gases Over Water

- $P_1 = P_{gas} + P_{H2O}$

10.7 Kinetic – Molecular Theory

- kinetic molecular theory
  - 1) gases of large # molecules are in continuous random motion
  - 2) volume of Molecules negligible compared to total volume of container
  - 3) attractive and repulsive forces negligible
  - 4) energy can be transferred in collisions but average kinetic energy stays same if constant temperature
  - 5) average kinetic energy of molecules proportional to absolute temperature
- individual molecules in gases have varying speeds
- root – mean square (rms) speed, u, varies in proportion to square root of absolute temperature and inversely with square root of molar mass
- $$u = \sqrt{\frac{3RT}{M}}$$
- $\sum = \frac{1}{2} m u^2$ (average kinetic energy of gas molecules)

10.7.1 Application to the Gas Laws

- 1)effect of volume increases at constant temperature
  - pressure decreases, fewer collisions with container wall
- 2) effect of temperature increase at constant volume
  - change in momentum of collisions increase, pressure increases

10.8 Molecular Effusion and Diffusion

- effusion – escape of gas molecule through tiny hole into evacuated space
- diffusion – speed of one substance throughout space

10.8.1 Graham's law of Effusion

- $$\frac{r_1}{r_2} = \sqrt{\frac{m_2}{m_1}}$$
- rate of effusion directly proportional to rms
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- $$\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{\frac{3RT}{M_1}}{\frac{3RT}{M_2}}} = \sqrt{\frac{M_2}{M_1}}$$

10.8.2 Diffusion and Mean free path

- diffusion faster for light molecules
- diffusion slower than molecular speeds because of collisions
- mean free path – average distance traveled by a molecule between collisions

10.9 Real Gases: Deviation from Ideal Behavior

- gases deviate from ideal behavior at higher pressure
- gases deviate from ideal behavior with decrease in temperature
- molecules in ideal gas assumed to occupy no space and have no attractions for one another
- real molecules have finite volumes and attract one another
- at high pressures impact on container wall from molecules lessened
- temperature determines how effective attractive forces are; decrease in temp = more effective

10.9.1 The van der Waals equation

- van der Waals equation: $\left(P + \frac{n_2 a}{V^2}\right)(V - nb) = nRT$
- a, b different for each gas
- a, b increase with increase in mass and complexity of structure
- larger, massive molecule have larger volumes, greater intermolecular attractive forces