

Chapter #5: Gases

5.2: Pressure- The Result of Molecular Collisions

- Force from the constant collision of molecules or atoms of a given area in gas creates pressure.
- Pressure depends on concentration. Higher the concentration, higher the pressure.
 - More molecules equal more collisions and more force.
- Pressure decrease with increasing altitude.

Pressure Units

- Pressure can be measured in millimeter of mercury (mmHg).
- Barometer measures pressure.
- mmHg can also be called torr after physicist Torricelli who invented the barometer.
- Atmosphere is another unit for pressure.
 - 1 atm = 760 mmHg
- 1 atm = 101325 Pa (pascal)
- 1 atm = 29.92 inHg (inches of mercury)
- 1 atm = 14.7 psi (pounds per square inch)

Manometer

- U-shaped tube with one side open to atmospheric pressure and other side with gas sample.
- Measures pressure relative to atmospheric pressure.

5.3: Simple Gas Laws- Boyle's Law, Charles's Law, and Avogadro's Law

Boyle's Law- Volume and Pressure

- There is an inverse relationship between volume and pressure.
 - Increase in one cause a decrease in other.
- Volume decrease, same number of gas particles in smaller space has more collisions and thus increase pressure.
- Temperature and amount of gas must remain constant.

Charles's Law- Volume and Temperature

- Pressure kept constant.
- Volume increase as temperature increase.
- If linear relationship graph of volume and temperature is extrapolated, gas have zero volume at 0 K. This is coldest possible temperature and there can be no negative volume.
- Increase in temperature result in increase in energy of particles and more collision, but for the pressure to remain the same, volume increases to accommodate the collisions.

Avogadro's Law- Volume and Moles

- At constant temperature and pressure.
- As concentration increase, volume increases.
- More gas particles needs more space to occupy.

5.4: Ideal Gas Law

- Combine Boyle's, Charles's. and Avogadro's Law to create ideal gas law.

$$V \propto \frac{nT}{P}$$

- Replace proportionality sign with ideal gas constant to create equation: $PV=nRT$.
- Ideal gas law for an ideal gas.
- $R= 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ (ideal gas constant)
- Gay-Lussac's law is as temperature increase, pressure also increases.
- For ideal gas law, pressure (atm), volume (L), moles (mol), and temperature (K).

5.5: Applications of Ideal Gas Law

Molar Volume at STP

- Molar volume is volume occupied by one mole.
- Standard temperature and pressure is 273K and 1 atm.
- Under STP a gas is 22.4L/1 mol.

Density of Gas

- Density (g/L) = Molar Mass / Molar Volume
- Rearrange ideal gas law to $n/V = P/RT$ to calculate density.
- Density = PM/RT

Molar Mass of Gas

- Find moles using ideal gas law and divide mass of sample by moles to find molar mass.

5.6: Mixtures of Gases and Partial Pressures

- Many gas samples are mixtures of different gases but ideal gas law assumes the different gas molecules do not interact.
- Partial pressure is pressure of individual component in mixture.
- Assume each gas component acts independently and calculate partial pressure of each component using ideal gas law and number of moles of component in mixture.
- Dalton's law of partial pressures states the sum of partial pressures equal total pressure.
 - All particles have the same average kinetic energy at a given temperature and exert the same force.
- Mole fraction = mole of component/ total moles
- Mole fraction of a component times total pressure equals partial pressure of component.

5.7: Gases in Chemical Reactions

- Use stoichiometry to calculate pressure, volume, or temperature of a product or reactant in a chemical equation.
- Given mass or moles, convert to moles of another reactant or product.
- Then use the ideal gas law to solve for pressure, volume, or temperature.

Molar Volume and Stoichiometry

- 1 mol of gas at STP equals 22.4L so use this conversion to convert between volume and grams or moles.

5.8: Kinetic Molecular Theory

- Kinetic molecular theory models gases as particles in constant motion and moves in a straight line until it a collision occurs.
- Kinetic molecular theory assumptions include:

- Size of a particle is negligibly small.
- Kinetic energy is proportional to temperature (K).
- Collision of particles is completely elastic.
- Kinetic molecular theory assumption valid for ideal gas law.
- Pressure = Force / Area ; Force = mass * acceleration
- Ideal gas law is derived from kinetic molecular theory.

Temperature and Molecular Velocities

- Kinetic Energy = 0.5(mass)(velocity)²
- Of a gas mixture at given temperature, on average lighter particles travel faster than heavier particles.

$$\text{Root mean square velocity} = \sqrt{\frac{3RT}{M}} \quad (R = 8.314 \text{ J/molK}; M \text{ (kg/mol)}; T \text{ (K)})$$

- Root mean square velocity is an average velocity and some particles move faster or slower.
- As velocity increases, velocity distribution becomes broader.
 - Increase in temperature or light particles create broader distribution.

5.9: Mean Free Path, Diffusion, and Effusion of Gases

- Gaseous particles travel at high speed but haphazard paths. Only travels short distance before it collides and changes direction.
- Mean free path is the average distance a molecule travel between collision.
 - Mean free path increase with decreasing pressure.
- Diffusion is when gas molecules spread out to concentration gradient.
 - Lighter molecules diffuse faster than heavier molecules because it has higher velocity.
- Effusion is when gas escapes container through small hole.
 - Heavier molecules effuse more slowly than lighter ones.

$$\text{rate} \propto \frac{1}{\sqrt{M}}$$

- Graham's law of effusion is the ratio of effusion rates of two different gases.

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{M_B}{M_A}}$$

5.10: Real Gases- Size and Intermolecular Forces

- Most gases act very nearly to ideal gases if:
 - Volume of gas particles is smaller than space between them.
 - Forces between has particles are not significant.

Effect of Finite Volume of Gas Particles

- At low pressure, molar volume of a gas is identical to that of an ideal gas but as pressure increases, molar volume of gas becomes greater than ideal gas.
 - At high pressure, volume of atom occupy significant portion of gas volume making actual volume greater that ideal gas law prediction.
- Ideal gas law predicts a volume that is too small.