Unit 5 – Gases and Gas Laws

13.1 The Nature of Gases

I. The Kinetic Theory and a Model for Gases

A. Assumptions of the Kinetic-Molecular Theory

1. Gases consist of large numbers of tiny particles that are far apart relative to their size
2. Gas particles undergo elastic collisions
   a. Collisions in which no energy is lost
3. Gas particles are in constant, rapid motion. They therefore possess kinetic energy, the energy of motion

II. **Gas** Pressure

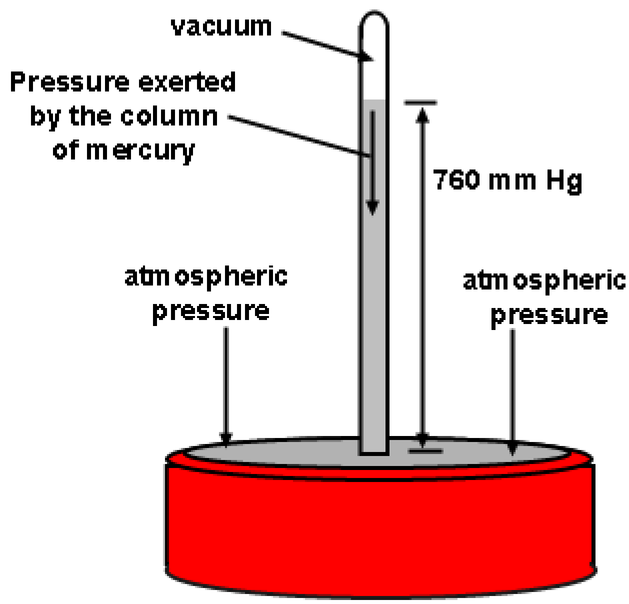
A. Pressure

1. The force per unit area on a surface

$$pressure = \frac{force}{area}$$

2. Gas molecules exert force, and therefore pressure, on any surface with which they collide

B. Units of Pressure



Units of Pressure		
Unit	*Symbol*	*Definition/Relationship*
Pascal	Pa	SI pressure unit 1 Pa = $\frac{1N}{m^2}$
Millimeter of mercury	mm Hg	Pressure that supports a 1 mm column of mercury in a barometer
Atmosphere	atm	Average atmospheric pressure at sea level and 0 °C
Torr	torr	1 torr = 1 mm Hg

C. Standard Pressure

1. 1 atm = 760 mm Hg = 760 torr = 101.3 kPa

III. Kinetic Energy and Temperature

A. Formula for Kinetic Energy

$$KE = \frac{1}{2}mv^2 \qquad m = \text{mass} \qquad v = \text{speed}$$

B. Relationship to Temperature

1. The average kinetic energy of gas particles depends on the temperature
2. All gases at the same temperature have the same average kinetic energy
   a. Small molecules (small mass, m) have higher average speeds
3. Kelvin temperature is directly proportional to the average kinetic energy of a substance
   a. 0 Kelvin = absolute zero = NO kinetic energy

14.1 Properties of Gases

I. Properties of Gases

A. Expansion
1. Gases do not have a definite shape or volume
2. Gases take the shape of their containers
3. Gases evenly distribute themselves within a container

B. Fluidity
1. Gas particles easily flow past one another

C. Low Density
1. A substance in the gaseous state has 1/1000 the density of the same substance in the liquid or solid state

D. Compressibility
1. Gases can be compressed, decreasing the distance between particles, and decreasing the volume occupied by the gas

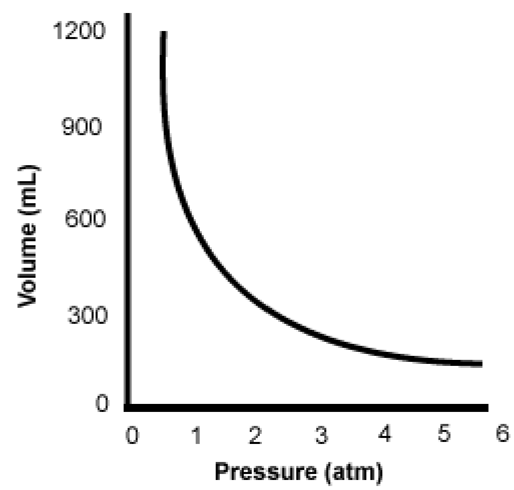
III. Factors Affecting Gas Pressure

A. Amount of Gas
1. ⇑ molecules = ⇑ collisions with walls = ⇑ pressure
2. ⇓ molecules = ⇓ collisions with walls = ⇓ pressure

B. Volume
1. ⇑ volume = ⇑ surface area = ⇓ collisions *per unit of area* = ⇓ pressure
2. ⇓ volume = ⇓ surface area = ⇑ collisions *per unit of area* = ⇑ pressure

C. Temperature
1. ⇑ temperature = ⇑ molecule speed = ⇑ frequent (and harder) collisions = ⇑ pressure
2. ⇓ temperature = ⇓ molecule speed = ⇓ frequent (and softer) collisions = ⇓ pressure

14.2 The Gas Laws

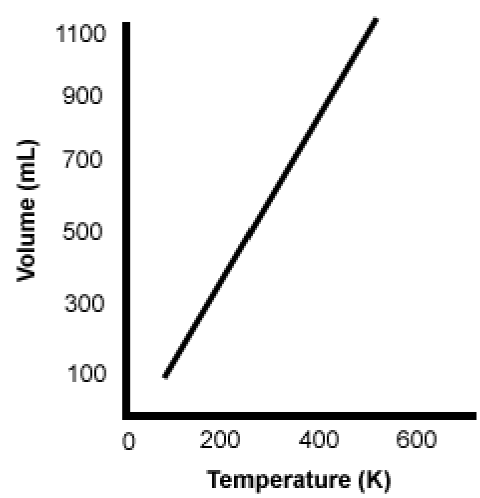


I. Boyle's Law: Pressure-Volume Relationship

A. Boyle's Law
1. The volume of a fixed mass of gas varies inversely with the pressure at constant temperature
   a. Volume ⇑ as pressure ⇓
   b. Volume ⇓ as pressure ⇑

B. Mathematical Statement of Boyle's Law
2. For identical masses of gas, at constant temperature

$$P_1V_1 = P_2V_2$$



II. Charles's Law: Volume-Temperature Relationships

A. Kelvin Temperature Scale (Absolute Scale)
1. K = 273 + °C
1. °C = K - 273
2. O K = absolute zero
3. Standard temperature = 0 °C = 273 K

B. Charles's Law
1. The volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature

C. Charles's Law Mathematically
1. For identical masses of gases, at constant pressure:

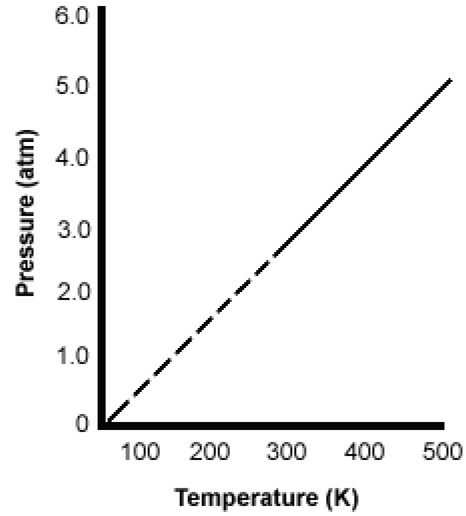
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

III. Gay Lussac's Law

A. Gay Lussac's Law

1. The pressure of a fixed mass of gas at constant volume varies directly with the Kelvin temperature
2. For identical masses of gases, at constant volume:

$$\therefore \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$



IV. The Combined Gas Law

A. The Combined Law

1. A mathematical expression of the relationship between pressure, volume and temperature of a fixed amount of gas (constant mass)
(in real life experiments, pressure, volume and temperature may all change)

$$\therefore \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

14.3 Ideal Gases

I. Ideal Gas Law

A. The mathematical relationship of pressure, volume, temperature, and the number of moles of a gas.

1. Mathematically:

$$PV = nRT$$

a. P = Pressure in atmospheres
b. V = Volume in liters
c. n = # of moles
d. T = Temperature in Kelvins

2. The ideal gas law reduces to Boyle's, Charles's, or Gay-Lussac's Law if the necessary variable is held constant

B. The Ideal Gas Constant

1. Units for R depend on units of measurement used for *P, V,* and *T*
2. For units of kilopascals, liters, and Kelvins

$$R = 8.31 \frac{L \cdot kPa}{mol \cdot K}$$

3. For units of atmospheres, liters, and Kelvins:

$$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$$

C. Finding *P, V, T* or *n*

1. Three of the four variables must be known in order to use the ideal gas law

D. Finding Molar Mass Using the Ideal Gas Law

1. $n = \dfrac{mass}{molar\ mass}$ so… $PV = \dfrac{mRT}{M}$ $\quad \therefore \quad M = \dfrac{mRT}{PV}$

E. Finding Density Using the Ideal Gas Law

1. $D = \frac{m}{V}$ and $M = \frac{mRT}{PV}$ Substituting D for $\frac{m}{V}$, you get $M = \frac{DRT}{P}$

2. Rearranging to solve for *D*:

$$D = \frac{MP}{RT}$$

II. Ideal Gases and Real Gases

A. Ideal Gas

1. An imaginary gas that perfectly fits all the assumptions of the kinetic-molecular theory

B. Real Gases

1. A gas that does not behave completely according to the assumptions of the kinetic-molecular theory.
2. Real gases occupy space and exert attractive forces on one another

Likely to behave nearly ideally
Gases at high temperature and low pressure
Small non-polar gas molecules

Likely not to behave ideally
Gases at low temperature and high pressure
Large, polar gas molecules

14.4 Gases: Mixtures and Movements

I. Dalton's Law of Partial Pressures

A. Partial Pressure

1. The pressure exerted by each gas in a mixture

B. Dalton's Law

1. The total pressure of a mixture of gases is equal to the sum of the partial pressures of the component gases

$$P_T = P_1 + P_2 + P_3 + \ldots$$

II. Graham's Law

A. Diffusion

1. Spontaneous mixing of particles of two substances caused by their random motion
2. Rate of diffusion is dependent upon:
   a. speed of particles
   b. diameter of particles
   c. attractive forces between the particles

B. Effusion

1. Process by which particles under pressure pass through a tiny opening
2. Rate of effusion is dependent upon:
   a. speed of particles (small molecules have greater speed than large molecules at the same temperature, so the effuse more rapidly)

C. Graham's Mathematical Law

$$\frac{rate\ of\ effusion\ of\ A}{rate\ of\ effusion\ of\ B} = \frac{\sqrt{Molar\ Mass\ of\ Gas\ B}}{\sqrt{Molar\ Mass\ of\ Gas\ A}}$$