# Chapter 10 - Physical Characteristics of Gases

## 10-1 The Kinetic-Molecular Theory of Matter

I. The Kinetic-Molecular Theory of Gases
   A. Ideal Gas
      1. An imaginary gas that perfectly fits all the assumptions of the kinetic-molecular theory
   B. Five 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
      4. There are no forces of attraction or repulsion between gas particles
      5. The average kinetic energy of gas particles depends on the temperature

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

         a. All gases at the same temperature have the same average kinetic energy
         b. Small molecules (small mass, *m*) have higher average speeds

II. The Kinetic-Molecular Theory of the Nature 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
   E. 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
   F. 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)

III. Deviations of Real Gases from Ideal Behavior

A. 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 |

**10-2 Pressure**

I. Pressure and Force

A. Pressure

1. The force per unit area on a surface

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

a. SI unit for force is the *newton* $\left(\frac{kg \cdot m}{\sec^2}\right)$

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

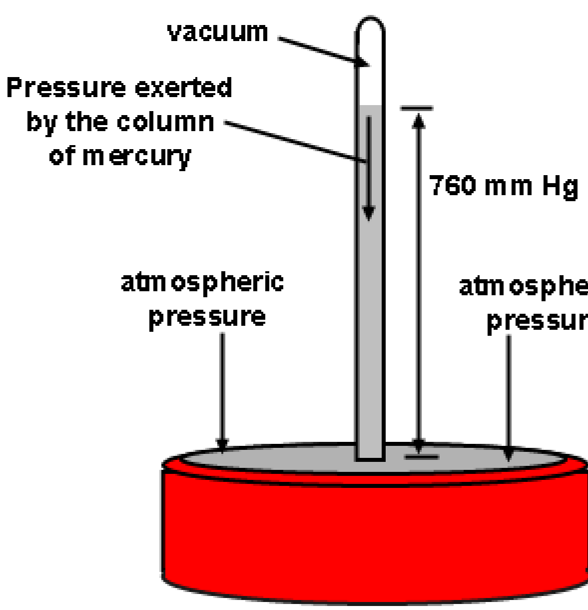
B. Composition of the Dry Atmosphere

1. 78% Nitrogen ($N_2$)
2. 21% Oxygen ($O_2$)
3. 1% other gases

C. Measuring Pressure

1. Barometer

a. The mercury barometer was invented by Evangelista Torricelli in the 1600's



D. Units of Pressure

| Table 10-1 | Units of Pressure | |
|---|---|---|
| *Unit* | *Symbol* | *Definition/Relationship* |
| Pascal | Pa | SI pressure unit<br>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<br>1 atm = 760 mm Hg = 760 torr<br>= $1.013\ 25 \times 10^5$ Pa<br>= 101.325 kPa<br>= 760 torr |
| Torr | torr | 1 torr = 1 mm Hg |

E. Standard Temperature and Pressure (STP)

1. Temperature = 0 °C
2. Pressure = 1 atm = 760 mm Hg (torr)

**10-3 The Gas Laws**

Gas Laws - Simple mathematical relationships between the volume, temperature, pressure, and quantity of a gas.

I. Boyle's Law: Pressure-Volume Relationship

A. Boyle's Law

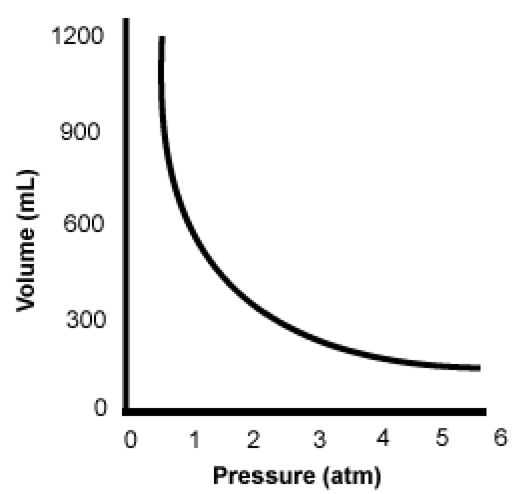
1. The volume of a fixed mass of gas varies inversely with the pressure at constant temperature
   a. Volume ⇑ as pressure ⇓
   b. Volume ⇓ as pressure ⇑

B. Mathematical Statement of Boyle's Law

1. $PV = k$ or… $V = k\frac{1}{P}$
   where *k* is a proportionality constant
2. For identical masses of gas, at constant temperature

$P_1V_1 = k$ and
$P_2V_2 = k$

$$\therefore \boxed{P_1V_1 = P_2V_2}$$



II. Charles's Law: Volume-Temperature Relationships

A. Observations of Jacques Charles

1. Volumes of any gas at constant pressure would change by 1/273 of the original volume for every Celsius degree the temperature rose or fell from 0 °C

B. Kelvin Temperature Scale (Absolute Scale)

1. K = 273 + °C
2. °C = K - 273
3. O K = absolute zero
4. Gas volume and Kelvin temperature are directly proportional
5. Standard temperature = 0 °C = 273 K

C. Charles's Law

1. The volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature

D. Charles's Law Mathematically

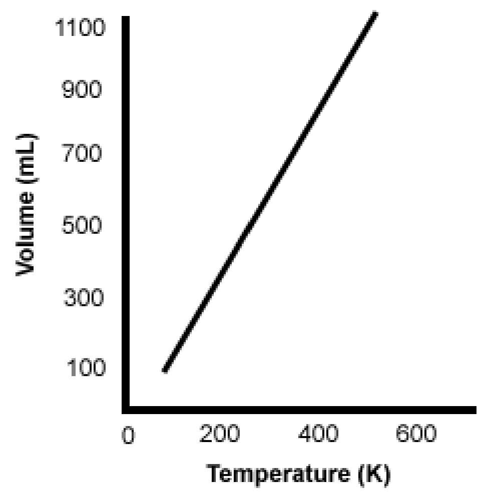
1. $V = kT$ or… $\frac{V}{T} = k$

where k is a proportionality constant and T is the temperature expressed in degrees Kelvin

2. For identical masses of gases, at constant pressure:

$$\frac{V_1}{T_1} = k \quad \text{and} \quad \frac{V_2}{T_2} = k$$

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




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

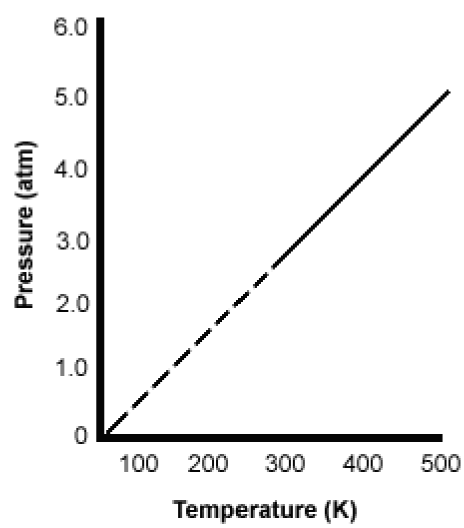
$P = kT$ or… $\frac{P}{T} = k$

where k is a proportionality constant and T is the temperature expressed in degrees Kelvin

3. For identical masses of gases, at constant volume:

$$\frac{P_1}{T_1} = k \quad \text{and} \quad \frac{P_2}{T_2} = k$$

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



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)

$$\frac{PV}{T} = k$$ where k is a proportionality constant and T is the temperature in degrees Kelvin

$$\frac{P_1V_1}{T_1} = k \quad \text{and} \quad \frac{P_2V_2}{T_2} = k \qquad \therefore \boxed{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}$$

V. 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

$$P_T = P_1 + P_2 + P_3 + ...$$

C. Gases Collected by Water Displacement

1. Gases produced in the lab are often collected by the displacement of water in a collection bottle
2. Water vapor will be present in the collected gas, and it exerts a pressure
   Water vapor pressure = $P_{H_2O}$
3. Water vapor pressure increases with temperature (Appendix A, Table-8)
4. Pressure of the dry gas

$$P_{atm} = P_{gas} + P_{H_2O} \qquad \therefore \boxed{P_{gas} = P_{atm} - P_{H_2O}}$$